$_{Z}X^{A}$

ATOMIC STRUCTURE

Electron (e^-):-

↳ Electron was discovered by JJ Thomson in 1896.
Mass of electron = 9.1×10^{-31} kg &
charge of electron = -1.6×10^{-19} C

proton (p^+):-

↳ proton was discovered by E. Goldstein in 1886.
Mass of proton = 1.67×10^{-27} kg &
charge of proton = 1.6×10^{-19} C

Neutron (n^0):-

↳ Neutron was discovered by Chadwick in 1932.
Mass of neutron = approx. 1.67×10^{-27} kg &
they don't contain any charge.

Atomic number (z):-

↳ The number of proton or electron present in a neutral atom is called atomic number. i.e., $[z = e^- = p^+]$
E.g. = $_1H$, $_{17}CL$

Mass Number (A)

↳ The Sum of proton's and neutrons present in an atom is called mass number. i.e., $\boxed{A = p^+ + n^0}$
E.g. = $^{16}_{8}O$, $^{9}_{4}Be$, $^{35}_{17}CL$

Isotopes :-

↳ Isotopes are the atoms of element having same atomic number but different mass number are called isotopes. It has different number of neutron.

Examples:- Hydrogen exist in 3 different isotopic form :-

1) protium → $^{1}_{1}H$ → (called ordinary hydrogen)
2) Deuterium → $^{2}_{1}H$ → (called heavy hydrogen)
3) Tritium → $^{3}_{1}H$ → (called radioactive hydrogen)

Isobars :-

↳ The atoms of different element having same mass number (A) but different atomic number(z) are known as Isobars.

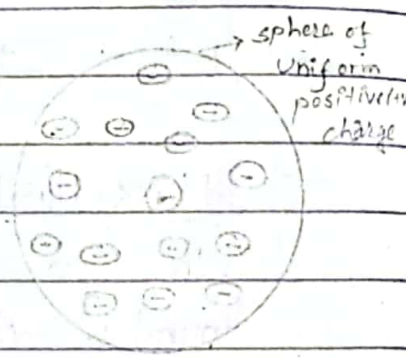
e.g. → $^{40}_{18}Ar$, $^{40}_{19}K$, $^{40}_{20}Ca$

Isotones :-

↳ The atoms of different element having same number of neutrons(n°) but different atomic number and mass number are known as isotones.

e.g. → $^{14}_{6}C$, $^{15}_{7}N$, $^{16}_{8}O$

J.J. Thomson atomic model :-



↳ Thomson proposed that an atom is a sphere of positively charge particle in which electrons or negatively charge particle are embedded like sheeds in watermelon. Which is known as pulm pudding (or Watermelon model).

Rutherford's model of atom/ α-scattering experiment/ Rutherford's atomic model:-

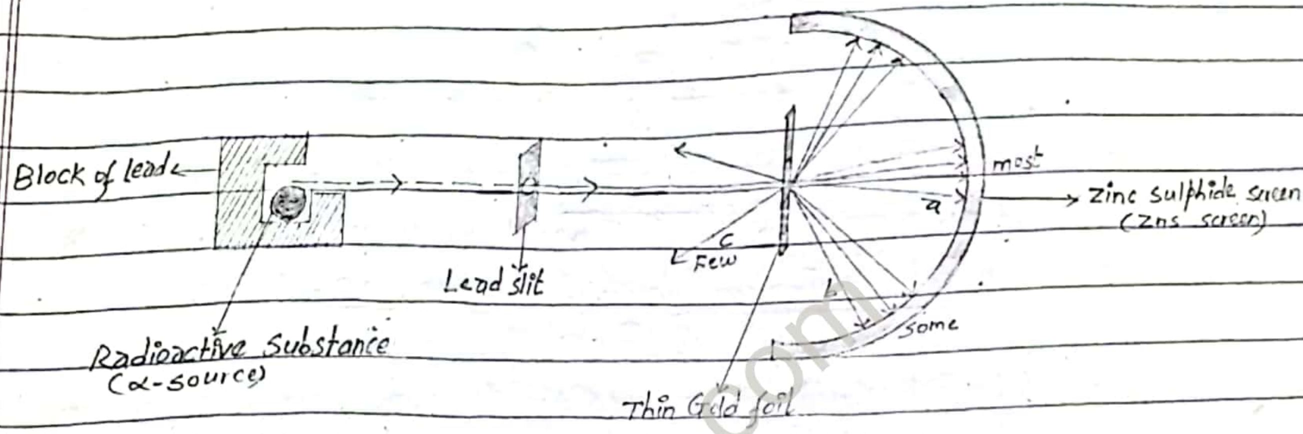


Fig: Rutherford's α-ray scattering experiment

↳ Rutherford's perform the α-particle scattering experiment for the structure of an atom and locate the position of electron & proton in an atom.

In this experiment, the source of α-particle taken inside the lead block and the beam of α-particle is allowed to strike on the thin gold foil through narrow lead slit. The gold foil is provided with circular ZnS (zinc sulphide) screen. α-particle produce flashes of light when strike on ZnS screen as shown in figure.

There are 3 types of α-particle were observed after scattering by atom contain in the gold foil:

a) Most of α-particles passed through gold foil without any deviation.
b) Some α-particles deviated through small angle.
c) Few α-particles were reflected or return back in same direction.

From this observation following conclusions are drawn:

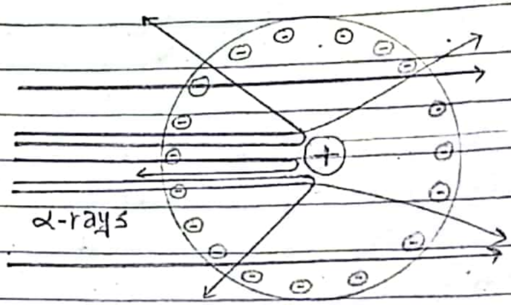


fig: Rutherford's α-ray scattering experiments focused on an atom.

☆ Observation:-/Conclusions:-

a) Most of α-particles passed through gold foil without any deviation it means there is large empty space within atom.

b) Some α-particles are deviated with small angle this shows that atom have a heavy positively charge body at it's centre.

c) Few α-particles were return back in same direction it means heavy positively charged nucleus located in very small region.

On the basis of these conclusions Rutherfords purpose nuclear model of an atom. the postulates of these nuclear model are as follows:

a) Atom consists of positively charged nucleus where mass of the atom should be concentrated.

b) Electrons are revolving around the nucleus which are held by electrostatic force of attraction.

c) An atom is electrically neutral. because the number of electrons and protons in an atom are equal.

Drawbacks/Limitat

A) According to nucleus it re spiral and fi explain about

B) It does not

c) It does not

Bohr's

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☆ Assum

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Drawbacks/Limitations of Rutherford's atomic model.

A) According to Rutherford's atomic model, electron revolve around the nucleus it rediats energy and the path of electron becomes spiral and finally falls at the nucleus. Hence it does not explain about stability of atom.

B) It does not explain hydrogen spectra.

c) It does not explain quantization of energy.

fig: falling of electron in spiral path into nucleus according to the law of electrodynamics:

Bohr's atomic model

V.V.I

↳ In 1913, Neils Bohr proposed an atomic model to overcome the limitation of Rutherford's model.

★ Assumptions/postulates of Bohr's atomic model:-

a) the electrons revolve around the nucleus in a fixed circular path called an orbit. An electron in a particular orbit or cell has a fixed amount of energy. Therefore, the orbit is also called stationary state or energy Level.

Date: / /

b) Only those orbits are permitted for which angular momentum of electron is integral multiple of $\frac{h}{2\pi}$.

i.e., $mvr = \frac{nh}{2\pi}$

Where,
m = mass of electron, v = Velocity of electron,
r = radius of atom and h = planck's Constant = 6.62×10^{-34} J.sec
n = principle quantum number, where, n = 1, 2, 3, 4, -----

The angular momentum of an electron is constant in a particular orbit. This is known as quantization of angular momentum.

c) An electron neither gain nor loss any amount of energy as long as it remains in particular orbit. Energy is emitted or absorbed when an electron jumps from one energy level to another. When an electron jumps from lower energy state to higher energy state, it absorbs a certain amount of energy. When it jumps from higher energy state to lower energy state it emits a certain amount of energy. The energy is emitted or absorbed in the form of radiation which can be observed in the form of spectra.

The amount of energy absorbs or emitted is given by the difference in energy of two energy level.

i.e., $\Delta E = E_2 - E_1 = h\nu = \frac{hc}{\lambda}$ $[\because \nu = \frac{c}{\lambda}]$

Where,
h = planck's Constant, ν = frequency of radiation
c = Velocity of light, λ = wave Length of radiation.

d) The centrifugal force of electron is balanced by electrostatic force of attraction between electron and nucleus.

Success of Bohr's Theory

* this theory explains the stability of an atom.
* this theory explains successfully the atomic spectrum of hydrogen atom.
* Energy and radius of the n^{th} orbit of a monoelectronic system (H, Li^{++}) can be calculated.

Explanation of emission of Hydrogen spectrum on the basis of Bohr's atomic model :-

↳ Hydrogen atom gives five series of line spectra on the basis of Bohr's atomic theory. According to Bohr's atomic theory, an atom possess/contains different orbit with different energy. The electrons in the orbits absorbs energy when it jumps higher energy level and emits energy when it returns back to lower energy level. Hydrogen gives absorbtion as well as emission spectra. The wave length of emitted spectral line is given by:

$$\frac{1}{\lambda} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

Where, λ = wave length of radiation, n_1 = Lower energy level, n_2 = Higher energy level, R = Rydberg Constant = 109677 cm^{-1}

Based on the relation, the emission of spectral line involved in the electronic transition. The spectral lines are diagrammatically shown in figure ~~below~~; below;

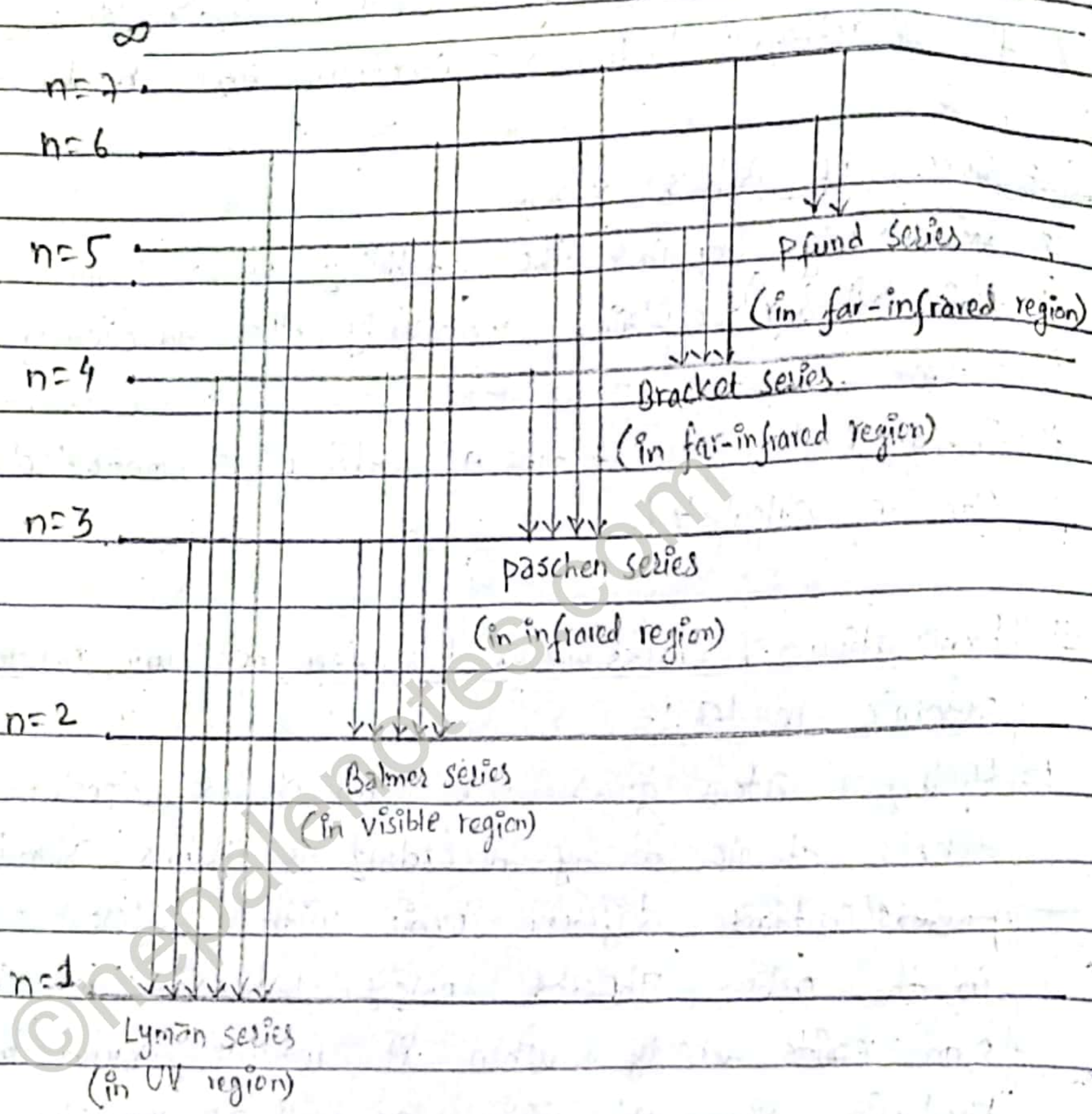


fig: Hydrogen spectra

IMP

Lyman series:-

↳ The series of spectral lines produced when electron jumps from higher energy level ($n_2 = 2,3,4,---,\infty$) to the lower energy level ($n_1=1$) is called lyman series. This spectral lines falls on UV region.

Balmer Series:-

↳ The series of spectral lines produced when electron jumps from higher energy level ($n_2 = 3,4,5,---,\infty$) to the lower energy level ($n_1 = 2$) is called Balmer series. This spectral lines falls on visible region.

paschen series:-

↳ The series of spectral lines produced when electron jumps from higher energy level ($n_2 = 4,5,6,---,\infty$) to the lower energy level ($n_1=3$) is called paschen series. This spectral lines falls on infrared(IR) region.

Bracket series:-

↳ The series of spectral lines produced when electron jumps from higher energy level ($n_2 = 5,6,7,---,\infty$) to the lower energy level($n_1=4$) is called Bracket series. This spectral lines falls on far-infrared region.

Pfund series:-

↳ The series of spectral lines produced when electron jumps from higher energy level ($n_2 = 6,7,8,---,\infty$) to the lower energy level ($n_1=5$) is called Pfund series. This spectral lines falls on far-infrared region.

Limitation of Bohr's atomic model:-

* Bohr's model fails to explain fine structure of spectral line.
* Bohr's model fails to explain spectral of multiple electron system.
* Bohr's model fails to explain further splitting line spectra into fine structure in the presense of magnetic field (Zeeman effect) and in the presense of electric field (stark effect).
* It does not explain dual nature of an electron.
* No mathematical background to explain the relation $mvr = \frac{nh}{2\pi}$.

De-Broglie Concept:-

↳ In 1924, Louis De-Broglie suggested that matter in motion like light have a duel character i.e., It behaves as particle as well as wave. According to this concept electron sometimes behaves as as particle and some times as wave.

By using this idea of wave nature of matter De-Broglie dessive an equation for wave length of electron.

According to Einstein's relation;

$$E = mc^2 \quad \text{----- (i)}$$

Where, E = energy, m = mass of particle, c = velocity of light

According to planck's statement

$$E = h\nu \quad \text{--- (ii)}$$

Where, E = energy, h = planck's constant, ν = frequency of radiation.

From eq^n(i) and eq^n(ii); we get;

$$h\nu = mc^2$$

or, $h \cdot \frac{c}{\lambda} = mc^2$

$$\Rightarrow \lambda = \frac{h}{mc} = \frac{h}{\text{moment of photon}} = \frac{h}{P}$$

For electron; $\lambda = \frac{h}{mv}$.

Where, m = mass of electron & v = Velocity of electron
If the mass of the object becomes smaller then λ is significant and that mass behaves as the wave Similarly, if the mass of the object is about a reasonable mass then that mass can be treated (act) (behave) as a particle.

Heisenberg's uncertainity principle :-

↳ It suggested that, "It is not possible to measure exact position and exact momentum (or velocity) simultaneously". This principle states that, " It is impossible to determined the position and momentum of microscopic particle accurately at the same time."

Mathematically;

$$\Delta x \cdot \Delta P \geq \frac{h}{4\pi} \quad \text{------- (i)}$$

Where,

Δx = Uncertainity in position, ΔP = Uncertainity in momentum &
h = planck's constant

from above equation, if momentum of particle is measure accurately i.e., $\Delta P = 0$; then there will be infinite Uncertainity in the position.

i.e., $\Delta x = \frac{h}{4\pi \cdot \Delta P} = \frac{h}{0} = \infty$

Since, $\Delta P = m \cdot \Delta v$

putting value of ΔP in eqn(i); we get,

or, $\Delta x \cdot m \cdot \Delta v = \frac{h}{4\pi}$

$$\Rightarrow \boxed{\Delta x \cdot \Delta v = \frac{h}{4\pi m}}$$

Hence, this principle has no significance in our daily life as it is applicable for moving microscopic particle.

Orbit and orbitals :-

Orbitals	Orbit
• It is the region around the nucleus where probability of finding electron is maximum.	• It is the circular path around nucleus in which electron revolve around the nucleus.
• It is denoted by letter's s, p, d & f.	• It is denoted by principle quantam number, i.e. n=1,2,3,4 or K, L, M, N.
• Orbitals have different shape, e.g:- 's'-orbital have spherical shape. 'p'-orbital have dumb-bell shape.	• All orbits are circular.
• Orbital have directional character except 's-orbital	• Orbit don't have directional character.
• One orbital can accommodate maximum two electrons	• One orbit can accommodate maximum $2n^2$ number of electron. where, n is the principle quantan number.

Shapes of orbitals :-

↳ 's'-orbital have spherical shape and it does not have directional character. The size of 's'-orbital increases with increase in principle quantum number.

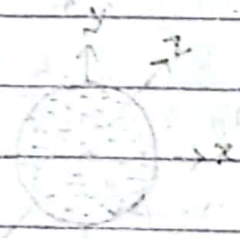
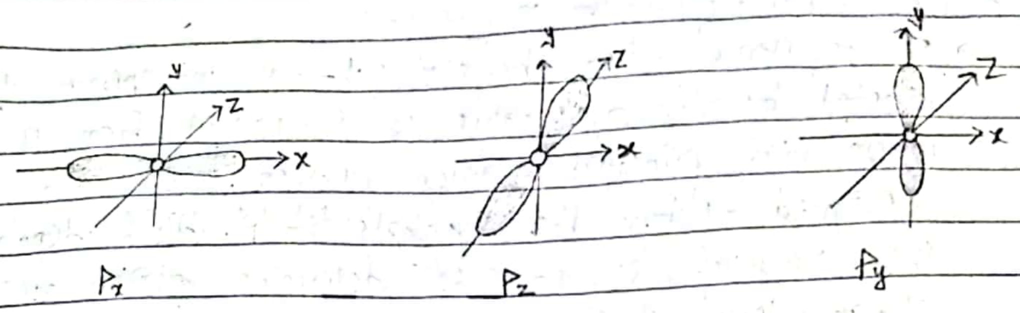


fig: s-orbital

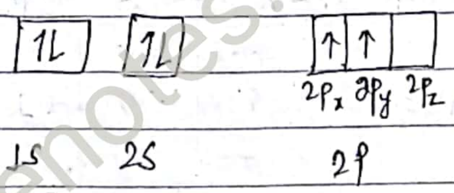
All the 'p'-orbital have dumb-bell shape with two lobs on opposite side of nucleus. p-orbital are designed as P_x, P_y, P_z. p-orbital has directional character.



Degenerate Orbital:-

↳ the orbitals of an atom having equal energy are known as degenerate orbitals.

e.g. In Carbon:-



Quantum Number:-

↳ These are the identification numbers which give Complete information about an electron in an atom. Four quantum numbers are required for the Complete indentification of an electron. They are:- [1] principal quantum number (n), [2] Azimuthal quantum number (l), [3] Magnetic quantum number (m), [4] Spin quantum number (s)

[1] principal quantum number (n):-

↳ It represents the main energy level (shell/orbit). It is denoted by 'n'. It can have only positive integer value, i.e., n=1,2,3,4--- and it can also be represented by letters K,L,M,N. This quantum number helps to determine size and energy of an electron. The maximum number of electron that each can hold is given by; $2n^2$ rule.

[2] Azimuthal quantum number(l):-

↳ It represents the shape of sub-shell and energy also. It is denoted by 'l'. The value of 'l' ranges from '0' to (n-1). Where 'n' is principal quantum number. The sub-shell are designed with the symbols 's', 'p', 'd', 'f' depending upon the value of 'l'. It also determine orbital angular momentum of an electron is given by;

$$mvr = \sqrt{l(l+1)} \times \frac{h}{2\pi}$$

#

Designation of sub-shell	s	p	d	f
l value	0	1	2	3

#

shell	principal q.n. (n)	Azimuthal q.n.(l)	no. of sub-shell
K	n=1	0 (s-subshell)	1
L	n=2	0 (s-subshell) 1 (p-subshell)	2
M	n=3	0 (s-subshell) 1 (p-subshell) 2 (d-subshell)	3

[3] Magnetic quantum number (m):-

↳ It represents the orientation of orbital in the space and total number of orbital in the sub-shell. It is denoted by letter 'm'. The. The value of 'm' ranges from $(-l)$ to $(+l)$ including 0. The total value of 'm' = $(2l+1)$

e.g:-

s-orbital;
$l = 0$, s-subshell
'm' ranges from $(-l)$ to $(+l)$
-0 to $+0$
$\Rightarrow \boxed{m = 0}$
i.e., 'm' has only one value 0.

p-orbital;
$l = 1$,
'm' ranges from (-1) to $(+1)$
$\Rightarrow \boxed{m = -1, 0, 1}$
i.e., 'm' has 3 value.

d-orbital;
$l = 2$,
'm' ranges from (-2) to $(+2)$
$\Rightarrow \boxed{m = -2, -1, 0, 1, 2}$
i.e, 'm' has '5' value.

\#

Shell	principal q.no. (n)	Azimuthal q.no. (l)	Magnetic q.no (m).
K-Shell	n = 1	0 (s-subshell)	0 (s-sub orbital)
L shell	n = 2	0 (s-sub shell)	0 (s-orbital)
		1 (p-subshell)	−1 (P_x − orbital)
			0 (P_y − orbital)
			+1 (P_z − orbital)
M shell	n = 3	0 (s-sub shell)	0 (s-orbital)
		1 (p-subshell)	−1 (P_x − orbital)
			0 (P_y − orbital)
			+1 (P_z − orbital)
		2 (d-subshell)	−2, −1, 0, +1, +2 → d-orbital

[4] Spin quantum numbers (s):-

↳ It represents the spin of an ~~atom~~ electron about its on axis. It is denoted by letter 's'. The value of Spin quantum numbers (s) are $+\frac{1}{2}$ & $-\frac{1}{2}$ for clock wise and anticlock wise direction of an electron respectively.

$+\frac{1}{2}$ indicate :- ↑ (upward spin)

$-\frac{1}{2}$ indicate :- ↓ (downward spin)

SQ: 1 [A] Find the value of n, l, m of 3s, 2s, 2p & 3p?

☆ Soln:-

For '3s' :-

n = 3,

l = 0 [s-subshell]

m = 0 [s-orbital]

For '2s'

n = 2

l = 0 [s-subshell]

m = 0 [s-orbital]

For '2p'

n = 2,

l = 1, [p-subshell]

m = 0, -1, +1 [p-orbital]

For 3p

n = 3,

l = 1, [p-subshell]

m = -1, 0, +1 [p-orbital]

Aufbau principle:-

↳ It states that, "The electrons are filled in the orbital in the order of their increasing energy." According to this principle, orbitals with lower energy is filled first followed by filling of orbital of higher energy. The energy of an orbital depends upon 'n' & 'l' value. Thus, filling of orbital can be explain on the basis of (n+l) value.

1) The orbital having lower value of (n+l) has lower energy and filled first. e.g., '4s' orbital is filled before '3d' orbital.

For '4s', $4+0=4$

For '3d', $3+2=5$

2) When two orbital have same (n+l) value the orbital with lower 'n' value has lower energy and filled first.

e.g. '4p' orbital filled before '5s' orbital.

For 4p; $4+1=5$

For 5s; $5+0=5$

pauli's Exclusion principle:-

↳ It states that, "No two electrons in an atom can have same value for all four quantum numbers." According to this principle, two electrons in a atom can have same value of three quantum number (i.e. n, l & m) at most but the fourth quantum number (i.e. s) should be different.

e.g.:

According to the concept of quantum number & pauli's exclusion principle, two electrons in '1s' orbital of 'he' can have following electronic configuration:-

$1s^2$

Electron	n	l	m	s
1st	1	0	0	$+1/2$
2nd	1	0	0	$-1/2$

Thus, two electrons Can have Same value for (n, l & m) but different for (s). This means that s-orbital Can accommodat (contain) two electrons with opposite spin and Cannot hold more than two electrons.

Hund's rule of Multiplicity:-

↳ It State that, "pairing in degenerate atomic orbitals does not take place whenever all the degenerate atomic orbitals are not Singly filled with parallel Spin."

e.g:-

Let us take in example in which three electrons are to be filled in p-orbital.

* Nitrogen(N):- $1s^2\ 2s^2\ 2p^3$

[↑↓] [↑↓] [↑↓ | ↑ |] ✗

1s 2s $2p_x\ 2p_y\ 2p_z$

[↑↓] [↑↓] [↑ | ↑ | ↑] ✓

* Neon (Ne):- $1s^2\ 2s^2\ 2p^6$

[↑↓] [↑↓] [↑↓ | ↑↓ | ↑↓]

1s 2s $2p_x\ 2p_y\ 2p_z$

Note, [May be it states, "In the orbital of same subshell, electrons are filled Sigly first, before pairing starts."]

Electronic Configuration:-

↳ The systematic distribution of electron in different shell, orbital and sub shell in order of their energy is known as electronic configuration.

Half filled and completly filled orbital are more stable then, partially filled orbital because they have lower energy which can be explain as:-

1) Half filled and completely filled orbital have symmetrical distribution of electrons. This symmetry lower the energy of the system and stability increases.

2) When the orbital are half filled and completely filled the exchange of position by the electrons lowers the energy of the system and stability increases.

Hence, the orbital having $p^3, p^6, d^5, d^{10}, F^7$ & F^{14} electronic configuration are relatively more stable.

* Cromium (Cr) :- (24) :-
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$

* Cuppor (Cu) :- (29) :-
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^{10}$

* Scandium (Sc) :- (21) :-
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^1$

← 1s
← 2s
← 3s 2p
← 4s 3p
← 5s 4p 3d
← 6s 5p 4d
← 7s 6p 5d 4f
← 7p 6d 5f