3 ATOMIC STRUCTURE

Rutherford's nuclear atomic model

Q.1. Explain Rutherford's nuclear atomic model (or planetary model).

Ans : Rutherford (1911) put forward his atomic model on the basis of scattering of α- particles experiment.

i) Atom is made-up of two parts :
   a) a positively charged centre-nucleus and
   b) extra nuclear part-electrons.

ii) Nucleus contains protons & neutrons.

iii) The total mass of the atoms is almost concentrated in nucleus.

iv) Electrons are present in the extranuclear part.

v) As the atom is electrically neutral, total number of electrons is equal to the total number of protons.

vi) The electrons and nucleus are held together by electrostatics force of attraction.

vii) Electrons are revolving around the nucleus just as the planets do around the sun. Hence this model is also called planetary model.

Drawbacks of Rutherford's atomic model

Q.2. What are the drawbacks of Rutherford's atomic model ?

i) It could not explain the stability of an atom.

ii) It could not explain the position, energy and distribution of electrons around the nucleus.

Bohr's theory

Q.3. Give the postulates of Bohr's theory. OR Explain Bohr's theory of hydrogen atom.

Ans :

i) Electrons revolve around the nucleus in concentric circular paths called orbits.

ii) Each orbit corresponds to a particular energy level.

iii) As long as an electron revolves in a given orbit, it does not radiate energy. Therefore its energy remains constant or stationary (and it does not fall into the nucleus). So the orbits are called stationary states or energy levels.

iv) A definite amount of energy is absorbed when an electron jumps from lower orbit to higher orbit similarly a definite amount of energy is given out when an electron jumps from higher orbit to lower orbit.

v) An electron is not free to revolve in any orbit. Only those orbits are permitted for which the angular momentum of the electron is an integral multiple of $h/2\pi$.

$$mvr = n\left(\frac{h}{2\pi}\right)$$

Where,
- $m \rightarrow$ Mass of electron.
- $v \rightarrow$ Velocity of electron.
- $r \rightarrow$ Radius of orbit.
- $n \rightarrow$ Integer.
- $h \rightarrow$ Planck's constant.

vi) Thus, in an atom there are several orbits around the nucleus at definite distance. They are numbered as 1,2,3,4..... or K,L,M,N.....

Draw backs of BOHR's theory

Q.4. What are the drawbacks of Bohr's theory ?

Ans :

a) It fails to explain the spectra of atoms containing several electrons.

b) It fails to explain the Zeeman effect i.e. the splitting of spectral lines under the influence of magnetic field.

c) It fails to explain the Stark effect i.e. the splitting of spectral line under the influence of electric field.

d) It does not consider the dual nature of electron. i.e. particle & wave nature.

Idea of shell, sub-shells and orbitals

a) Shells :

Q.5. What are shells ?

Ans : i) Shells or orbits is a well defined circular path around the nucleus in which an electron

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revolves.

ii) These obits are also called a principle energy levels or shells.

iii) These energy levels are represented by a principle quantum number (n)

iv) These shells are numbered as 1, 2, 3, 4, ... etc. or denoted as K, L, M , N, ... respectively.

v) As the shell number increases its distance from the nucleus and energy increases.

vi) The order of energies of various shells are in the order is K > L > M > N

vii) Maximum number of electrons in any shell is given by the formula $2n^2$. where ,'n' is the orbit number.

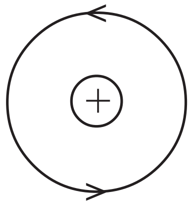


Fig. 3.1 : Shell

b) Sub-shells or sub levels :

Q.6. What are sub-shells ?

Ans :

i) A principle shell is further divided into sub-shells.

ii) Sub-shells are of four types such as s, p, d and f.

iii) Sub-shells give idea about the shape of electron cloud around the nucleus.

e.g. s - sub-shell : spherical shape
p - sub-shell : dumb - bell shape
d - sub-shell : double dumb - bell shape
f - sub-shell : complicated shape & so on.

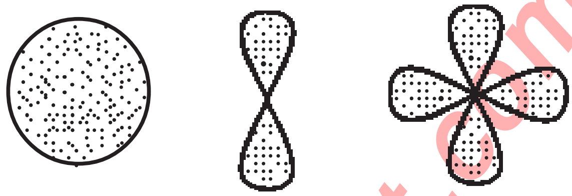
iv) Number of sub-shells present in the main energy level is equal to the orbit number.

For e.g.
First orbit contains only one sub-shell i.e. s
Second orbit contains two sub-shells i.e. s, p
Third orbit contains three sub-shells i.e. s, p, d
Fourth orbit contains four sub-shells i.e. s, p, d and f.

v) Maximum number of electrons present in different sub-shells are : s^2, p^6, d^{10}, f^{14}.

vi) The order of energies of sub-shells is :
s < p < d < f.

vii) The letters s, p, d and f originate from the terms sharp, principal, diffuse and fundamental which were old spectroscopic terms.



s - subshell p - subshell d - subshell

Fig.3.2

c) Orbitals

Q.7. What is orbital ?

Ans:

i) It is the three dimensional region in space around nucleus in which the probability of finding the electron is maximum.

ii) The number of orbitals in a given sub-shell are given below :

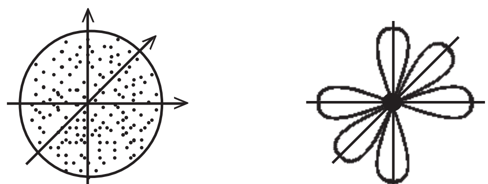
Sub -shell	**s**	**p**	**d**	**f**
No. of orbitals	**1**	**3**	**5**	**7**

iii) Maximum number of orbital in each shell is given by n^2 (where n = orbit number).

Orbit number	**1**	**2**	**3**	**4**
No. of orbitals	**1**	**4**	**9**	**16**

iv) Maximum no. of electrons in an orbital is 2.

v) Both the electrons present in an orbital have opposite spins clockwise (↑) and anti - clock (↓) .



s - orbital p - orbital

Fig. : 3.3

Shapes of s & p - orbitals

Q. 8. Explain the followings

i) s-orbitals ii) p-orbitals.

i) s-orbital : s-subenergy level contains only one orbital i.e. s-orbital.

ii) s-orbital is spherically symmetrical around the nucleus

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thus it is non - directional.

iii) It can accomodate two electrons with opposite spins.

iv) The size of s-orbital depends upon the principle quantum number i.e. is in order 4s > 3s > 2 s > 1s.

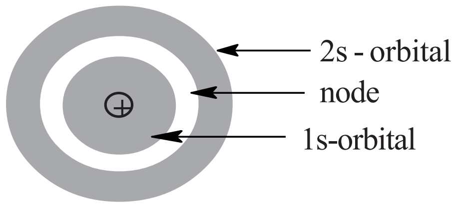


Fig. 3.4 : Shape of s - orbital.

ii) p-orbitals :

i) p-subenergy level contains three orbitals. i.e. px, py and pz.

ii) Each p-orbital is dumbell shaped. It consists of two lobes, one lobe on either side of the nucleus.

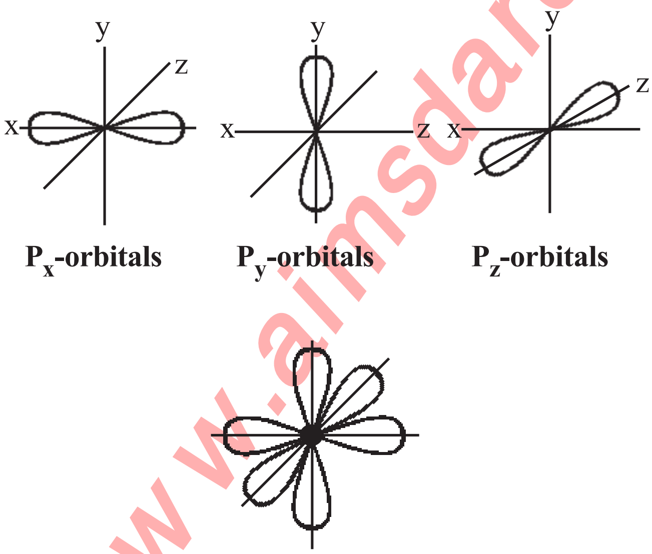
iii) The p_x, p_y and p_z orbitals are directed along x,y and z-axis respectively.

iv) Each p-orbital can accomodate two electrons with opposite spins.

v) The p_x, p_y and p_z orbitals in the same subshell have same energy.

vi) The size an energy of p -orbital increases with increase in principle energy level i.e. in order 4p > 3p > 2p

vii) The size ,shape & energy of the three obritals are same.



The three p-orbitals

Fig. 3.5 : p-orbitals (shape & orientation)

QUANTUM NUMBERS

Q.9. What are quantum numbers ? Name them.

Ans : *The numbers which indicate state of the electron are called quantum numbers.* There are four quantum numbers,

i) Principal quantum number (n)

ii) Azimuthal quantum number (l)

iii) Magnetic quantum number (m)

iv) Spin quantum number (s)

Principal Quantum Number (n)

Q.10. Explain the principal quantum number.

i) Principal quantum number (n) indicates the main energy level of an electron.

ii) It describes the size of electron cloud.

iii) It is denoted by 'n'

iv) Principle quantum number (n) can have positive integer values like 1,2,3,4,....

v) Higher the value of 'n' greater is the energy & distance of the orbit from the nucleus.

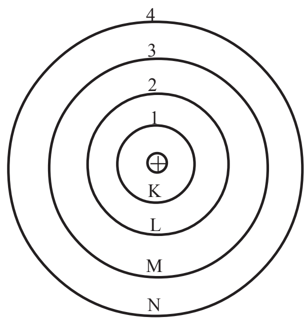


Fig. 3.6 : Electron shells.

Azimuthal Quantum Number

Q.11. What do you know about azimuthal quantum number.

Ans :

i) Azimuthal quantum number (l) indicates the sub - energy level of electron (i.e. s,p,d,f, etc.)

ii) It describes the shape of the electron cloud i.e. spherical, dumbell shaped etc.

iii) It is doneted by '*l*'

iv) Values of 'l' depend upon the principal quantum number 'n'.

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v) Values of 'l' range from 0 to (n–1) for given value of n.

v) The 0, 1, 2 & 3 values of 'l' represent s, p, d & f subshells respectively.

Principal quantum no. (n) (shell no.)	Total values of Azimuthal quantum no. (l) (sub-shell)	Values of Azimuthal quantum no. (sub-shell)	Remark
i) n = 1 (i.e. K shell)	One value	0 (i.e. 1 s)	1st shell contains only one sub shell i.e. (1s)
ii) n = 2 (i.e. L shell)	Two values	0 , 1 (2s) (2p)	2nd shell contains 2 sub-energy levels 2s, 2p
iii) n = 3 (i.e. M shell)	Three values	0, 1, 2 (3s) (3p) (3d)	3rd shell contains 3 sub shells 3s, 3p, 3d.
iv) n = 4 (i.e. N shell)	Four values	0, 1, 2, 3 4s 4p 4d 4f	4th energy level contains 4 sub energy levels 4s,4p,4d,4f

Magnetic Quantum number (m)

Q.12. Explain the magnetic quantum number.

i) The magnetic quantum number indicates orbital of an electron in the subenergy level.

ii) Values of m depends upon azimuthal quantum number

iii) It is denoted by 'm'

iv) Total values of m for each value of $l = 2l + 1$

v) The values of 'm' range from $-l$ ... 0 ... $+l$

Subenergy level (l)	Total values of m = $2l+1$	Values of 'm' (orbitals)	Remark
$l = 0$ (s-subshell)	One	0 (s- orbital)	Thus s - subshell contains only one orbital i.e. s- orbtial.
$l = 1$ (p-subshell)	Three	-1 (p_x) 0 (p_y) +1 (p_z)	Thus p - subshell contains three orbitals p_x, p_y , p_z.
$l = 2$ (d-subshell)	Five	–2 (d_{xy}) –1 (d_{yz}) 0 (d_{zx}) +1 ($d_{x^2-y^2}$) +2 (d_{z^2})	Thus d - subshell contains five orbitals i.e. d_{xy} , d_{yz}, d_{zx} , $d_{x^2-y^2}$, d_{z^2}.
$l = 3$ (f-subshell)	Seven	-3, -2, -1, 0, 1, 2 ,3	Thus f - sub energy level contains seven orbitals.

Spin quantum no (s)

Q.13. Explain the spin quantum number.

Ans :

i) An electron rotates about its own axis (i.e. spinning) while revolving around the nucleus.

ii) It is denoted by 's'.

iii) An electron can spin in clockwise or anticlockwise direction.

iv) Spin quantum number indicates directions of electron spins.

v) The spin quantum number can have two values for each value of 'm' (+ ½ & -½)

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Quantum numbers and distrubution of electrons : Following table gives the distribution of electrons in various shells,subshells and orbitals on the basis of quantum numbers.

Shell or energy level (n)	**Subshell or subenergy level (l)**	**Orbitals (m)**	**No.of e^- in subshell (s) = 2 (2*l* + 1)**	**Total e^- in a shell ($2n^2$)**
K shell (n = 1)	0 (1s-suben- ergy level)	0 (1s-orbital)	+ ½ ,-½ = 2	2
L shell (n = 2) 8	0 (2s) 1 (2p)	0 (2s) -1 ($2p_x$) 0 ($2p_y$)	+ ½ ,-½ = 2 + ½ , -½ }	+ ½ , -½ 6
M shell (n = 3)	0 (3s) 1 (3p) 2 (3d)	+1 ($2p_z$) 0 (3s) -1 ($3p_x$) 0 ($3p_y$) +1 ($3p_z$) -2 (3dxy) -1 (3dyz) 0 (3dzx) 1 (3dx^2-y^2) 2 (3dz^2)	+ ½ , -½ + ½ ,-½ = 2 + ½ ,-½ } + ½ ,-½ 6 + ½ ,-½ + ½ ,-½ + ½ ,-½ + ½ ,-½ } 10 + ½ ,-½ + ½ ,-½	18
N shell (n = 4)	0 (4s) 1 (4p) 2 (4d) 3 (4f)	0 (4s) -1 ($4p_x$) 0 ($4p_y$) +1 ($4p_z$) -2 (4dxy) -1 (4dyz) 0 (4dzx) 1 (4dx^2-y^2) 2 (4dz^2) -3 -2 -1 0 1 2 3	+ ½ ,-½ = 2 + ½ ,-½ } + ½ ,-½ 6 + ½ ,-½ + ½ ,-½ + ½ ,-½ } + ½ ,-½ 10 + ½ ,-½ + ½ ,-½ + ½ ,-½ + ½ ,-½ + ½ ,-½ } + ½ ,-½ 14 + ½ ,-½ + ½ ,-½ + ½ ,-½	32

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Electronic Configuration

Chemical properties of an atom depend upon number of electrons and their distribution in shells, sub-shells

Q.14. Define electronic configuration.

Ans : It is the distribution of electrons in shell, sub-shell & orbitals of an atom.

The electronic configuration is mainly explained by following rules

1) Aufbau principle

2) Pauli's exclusion principle

3) Hund's rule of maximum multiplicity

Aufbau Principle

(German word, Aufbau : builiding up)

Q. 15. Explain Aufbau Principle.

Ans : i) According to aufbau principle, *"In the gound state the atoms of the orbitals are filled with electrons in order of the increasing energies."*

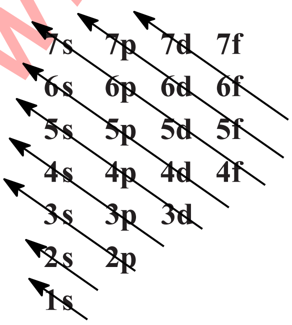
ii) Electron enters the orbitals according to following ($(n + l)$) rules.

a) The orbitals with lowest ($n + l$) value is filled first.
e.g. out of 3d and 4s-obtitals, 4s orbital is occupied first as its (n + 1) value ($4 + 0 = 4$) is lower than 3d-orbital ($3 + 2 = 5$).

b) When two or more orbitals have same ($n + l$) value the orbital with lowest value of n is filled first.
e.g. 3d and 4p orbitals have same $n + l$ value i.e. 5 In this case 3d will be occupied before the 4p orbital as it has lower n value.

iii) The order of energy of different orbitals in an atom is given below.
1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d and so on.

iv) The sequence given above can be represented by diagram as shown in below.



Pauli's Exclusion principle

Q. 16. State Pauli's Exclusion principle.

Ans : *"No two electrons in an atom can have the same set of all the four quantum numbers."*

Two electrons in an atom can have at the most three quantum numbers the same, but one quantum number will be different.

From this it is clear that an orbital can accomodate two electrons with opposite spins.

Hund's rule of maximum multiplicity

Q. 17. State Hund's rule of maximum multiplicity ?

Ans : "*Electron pairing does not take place in any orbital of same type until all the orbitals contain a single electron each*".

e.g. Electronic configuration of neon (Z = 10)

10) Ne → $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}$
(Z = 10) $1s^2$ $2s^2$ $2p_x^{\,2}\ 2p_y^{\,2}\ 2p_z^{\,2}$

Q.18.Explain why chromium and copper show exceptional electronic configurations ?

Ans : i) Cr : Expected configuration of chromium is

$Cr_{(Z=24)} \rightarrow 1s^2, 2s^2 2p^6, 3s^2 3p^6, 4s^2, 3d^4$

The 3d-orbital is not half filled. Hence it has less stability. One 4s electron to enter into 3d orbital to make 4s and 3d orbitals half filled so chromium acquire extra stability and its electronic configuration becomes.

$Cr_{(Z=24)} \rightarrow 1s^2, 2s^2 2p^6, 3s^2 3p^6, 4s^1, 3d^5$

ii) Cu : Expected configuration of copper is

$Cu_{(Z=29)} \rightarrow 1s^2, 2s^2 2p^6, 3s^2 3p^6, 4s^2, 3d^9$

The 3d-orbital is neither half filled nor completely filled. Hence it has less stability. One 4s electron to enter into 3d orbital to make it completely filled 4s orbital half filled so copper acquire extra stability and its electronic configuration becomes.

$Cu_{(Z=29)} \rightarrow 1s^2, 2s^2 2p^6, 3s^2 3p^6, 4s^1, 3d^{10}$

Q.19. Give electronic configurations of first ten elements.

Ans :

1) H → $1s^1$
(Z = 1)

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2) He → $1s^2$
(Z = 2)

3) Li → $1s^2, 2s^1$
(Z = 3)

4) Be → $1s^2, 2s^2$
(Z = 4)

5) B → $1s^2, 2s^2 2p_x^1 2p_y^0 2p_z^0$
(Z = 5)

6) C → $1s^2, 2s^2 2p_x^1 2p_y^1 2p_z^0$
(Z = 6)

7) N → $1s^2, 2s^2 2p_x^1 2p_y^1 2p_z^1$
(Z = 7)

8) O → $1s^2, 2s^2 2p_x^2 2p_y^1 2p_z^1$
(Z = 8)

9) F → $1s^2, 2s^2 2p_x^2 2p_y^2 2p_z^1$
(Z = 9)

10) Ne→ $1s^2, 2s^2 2p_x^2 2p_y^2 2p_z^2$
(Z = 10)

Q.20. Write electronic configuration of the elements from atomic number 11 to 18 .

Ans :

Na (Z = 11) → $1s^2, 2s^2\ 2p^6, 3s^1$.

Mg (Z = 12) → $1s^2, 2s^2\ 2p^6, 3s^2$.

Al (Z = 13) → $1s^2, 2s^2\ 2p^6, 3s^2\ 3p_x^1\ 3p_y^0\ 3p_z^0$

Si (Z = 14) → $1s^2, 2s^2\ 2p^6,\ 3s^2\ 3p_x^1\ 3p_y^1\ 3p_z^0$

P (Z = 15) → $1s^2, 2s^2\ 2p^6, 3s^2\ 3p_x^1\ 3p_y^1\ 3p_z^1$.

S (Z = 16) → $1s^2, 2s^2\ 2p^6, 3s^2\ 3p_x^2\ 3p_y^1\ 3p_z^1$.

Cl (Z=17) → $1s^2, 2s^2\ 2p^6, 3s^2\ 3p_x^2\ 3p_y^2\ 3p_z^1$.

Ar (Z=18) → $1s^2, 2s^2\ 2p^6, 3s^2\ 3p_x^2\ 3p_y^2\ 3p_z^2$.

Q.21. Give the electronic configuration of potassium and calcium.

Ans :

1) K (Z=19) → $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6, 4s^1$
2) Ca (Z=20) → $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6, 4s^2$

Electronic configuration of first series of transition metals.

Q.22. Write electronic configurations of the elements from Scandium (Z = 21) to Zinc (Z = 30).

At.no.	Element	Electronic Configuration
21	Scandium (Sc) →	$1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ 3d^1, 4s^2$
22	Titanium (Ti) →	$1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ 3d^2, 4s^2$
23	Vanadium (V) →	$1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ 3d^3, 4s^2$
24	Chromium (Cr) →	$1s^2, 2s^2\ 2p^6, 3s^2, 3p^6\ \mathbf{3d^5}, \mathbf{4s^1}$
25	Manganese (Mn) →	$1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ 3d^5, 4s^2$
26	Iron (Fe) →	$1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ 3d^6, 4s^2$
27	Cobalt (Co) →	$1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ 3d^7, 4s^2$
28	Nickel (Ni) →	$1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ 3d^8, 4s^2$
29	Copper (Cu) →	$1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ \mathbf{3d^{10}}, \mathbf{4s^1}$
30	Zinc (Zn) →	$1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ 3d^{10}, 4s^2$

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