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ATOMIC STRUCTURE

Atom (greek word)

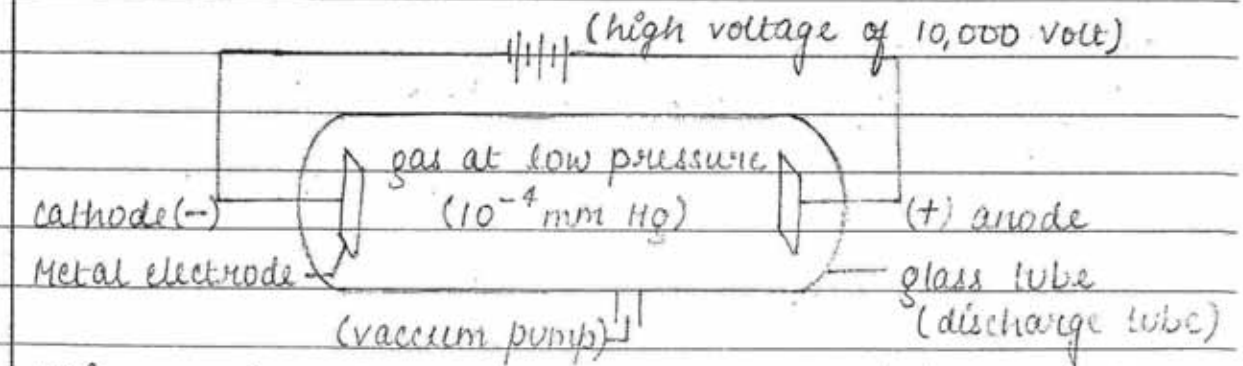
A (not) TOM (divisible) [according to John Dalton]

Atom : According to John Dalton (1803 - 1808) considered that "all matters are composed of small particles called atoms which are not further divisible".

Dalton's Atomic Theory :
It is based on two laws :
i) law of conservation of mass.
ii) law of definite / constant proportion

Features of Daltons Atomic Theory :
It includes five postulates.
MODULE 1 (Pg 157)

1. Discovery of Electron.
Cathode ray tube experiment:



This experiment was conducted by William crookes and Julius Plunker in 1879.

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Apparatus :

The cathode ray discharge tube consist of a long glass tube at the ends of which two metal electrodes are placed. those two electrodes are connected to a source of high voltage and a vaccum pump is attached to the glass tube which can control the pressure of the gas filled in the tube.

Working :

- When a high voltage of 10000 V is applied across the electrodes. In between these electrodes a gas at low pressure is present. Then, a stream of particles were observed that were flowing from cathode towards anode. These rays were termed as cathode rays.
- For the verification of cathode rays, few holes were made in the anode and the tube behind the anode was coated with a fluroscent material like zinc sulphide.
- When the cathode ray pass through the anode and strike the zinc sulphide coating, they form a bright spot over the coating. This verified that some particles were moving from cathode side towards anode side and thus these particles were known as cathode rays.

Properties of cathode rays :

MODULE 1 (Pg 137)

* Characteristics of cathode ray does not depend on the nature of the gas and the electrodes used in the discharged tube.

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* The specific charge e/m ratio (charge/mass ratio) of cathode rays (electron) does not depends on the nature of gas and electrodes present in the discharged tube. Hence, we can say that electrons are the basic constituents of all the atoms.

2 Discovery of Proton: (Anode rays or canal rays)
Goldstein observed that certain rays are moving from the side of anode towards cathode in a discharged tube. For verifying he used a perforated cathode and coated the tube behind the cathode with a fluresence material called zinc sulphide. he observed that some bright spots are formed on the zinc sulphide coating, which verified that certain rays were moving from the side of the anode towards the cathode. Hence, he named them as anode rays or canal rays.

Properties of Anode rays
MODULE 1 (Pg 157 and 158)

* The nature of anode rays depends upon the nature of the gas present in the discharged tube.
* The specific charge of anode rays depends upon the nature of the gas present in the discharged tube.
* The specific charge value for the anode rays was maximum when the hydrogen was filled in the discharge tube. It indicates that the positively charge ion from hydrogen is lightest.

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This lightest positively charged particle was named as Proton.

3. Discovery of Neutron :

James Chadwick in 1932 bombarded a thin sheet of Beryllium with fast moving alpha particles. It was observed that highly penetrating rays of neutral particles were produced. These particles were named as neutrons.

$$^{9}_{4}Be + ^{4}_{2}He \rightarrow ^{12}_{6}C + ^{1}_{0}n$$

{α - particle} {neutron}

The electron :

Properties of electron :

MODULE 1 (Pg 158)

The Proton :

Properties of proton :

MODULE 1 (Pg 158)

The Neutron :

Properties of Neutron :

MODULE 1 (Pg 159)

Other fundamental particles :

- Neutrino and Antineutrino
- Positron (positive electron $_{1}e^{0}$)
- Mesons
- Anti Proton

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Q.1 Pick out the correct sequence for the order of the mass.

i) $e > p > n > \pi$ (mu)[meson]

ii) $p > e > \pi > n$

iii) $n > p > e > \pi$

iv) $n > p > \pi > e$

Ans: iv)

Q.2 Pick out the correct sequence for e/m ratio:

i) $e > p > n > \alpha$

ii) $p > e > \alpha > n$

iii) $e > \alpha > p > n$

iv) $e > p > \alpha > n$

Ans: iv)

e/m ratio of $e = \frac{1.6 \times 10^{-19} C}{9.1 \times 10^{-28} g} = 1.76 \times 10^{8} C/g$

e/m ratio of $p = \frac{1.6 \times 10^{-19} C}{1.67 \times 10^{-24} g} = 9.58 \times 10^{4} C/g$

e/m ratio of $n = 0$

e/m ratio of $\alpha = \frac{2 \times 1.6 \times 10^{-19}}{4 \times 1.67 \times 10^{-24}} = \frac{9.58 \times 10^{4} C/g}{2}$

Q.3. Pick out the correct sequence for the e/m ratio:

i) $H^+ > D^+ > He^+ > He^{+2}$

ii) $H^+ > D^+ > He^{+2} > He^+$

iii) $H^+ > D^+ \doteq He^{+2} > He^+$

iv) $H^+ > D^+ \doteq He^+ > He^{+2}$

Ans: iii)

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e/m ratio of $H^+ = \frac{1.6 \times 10^{-19} C}{1.67 \times 10^{-24} g} = 9.58 \times 10^4 C/g$

e/m ratio of $He^+ = \frac{1.6 \times 10^{-19} C}{4 \times 1.67 \times 10^{-24} g} = \frac{9.58 \times 10^4 C/g}{4}$

e/m ratio of $He^{+2} = \frac{2 \times 1.6 \times 10^{-19} C}{4 \times 1.67 \times 10^{-24} g} = \frac{9.58 \times 10^4 C/g}{2}$

e/m ratio of $D^+ = \frac{1.6 \times 10^{-19} C}{2 \times 1.67 \times 10^{-24} g} = \frac{9.58 \times 10^4 C/g}{2}$

Representation of an atom of an element

$A \rightarrow$ Mass no

$_{Z}X \rightarrow$ Symbol of an element

$Z \rightarrow$ Atomic no.

Mass no = no. of proton + no. of neutrons = no. of nucleons

Atomic no = no. of proton = nuclear charge.

i) For a neutral atom,
no. of electrons (e) = no of proton (p) = Z.

ii) For a charged atom (ion)
no. of electrons (e) = Z - (charge on atom)

iii) No. of neutrons.
Mass no (A) − Atomic no (Z)

• For a neutral atom,
The no of electron = no. of protons.
$e = p$.

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- When an ion of an atom is formed, the electrons are transferred. Therefore, an atom and its ions differ in the no. of electron

- Mass no is always a whole number.

Q.4. Find the value of Mass no (A), Atomic no (Z), no. of proton (p), electron (e) and neutron (n).

i) $^{14}_{7}N^{-3}$

a) $A = 14$

b) $Z = 7$

c) $p^+ = Z = 7$

d) $e^- = Z - (\text{charge}) = 7 - (-3) = 7 + 3 = 10$

e) $n^0 = A - Z = 14 - 7 = 7$

ii) $^{16}_{8}O^{-2}$

a) $A = 16$

b) $Z = 8$

c) $p^+ = Z = 8$

d) $e^- = Z - (\text{charge}) = Z - (-2) = 8 + 2 = 10$

e) $n^0 = A - Z = 16 - 8 = 8$

iii) $^{23}_{11}Na$

a) $A = 23$

b) $Z = 11$

c) $p^+ = Z = 11$

d) $e^- = 11$

e) $n^0 = 23 - 11 = 12$

iv) $^{24}_{12}Mg^{+2}$

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b) $Z = 12$

c) $p^+ = Z = 12$

d) $e^- = Z - (\text{charge}) = 12 - 2 = 10.$

e) $n^0 = A - Z = 24 - 12 = 12$

v) $^{27}_{13}Al$

a) $A = 27$

b) $Z = 13.$

c) $p^+ = Z = 13.$

d) $e^- = 13$

e) $n^0 = A - Z = 27 - 13 = 14$

vi) $^{19}_{9}F^-$

a) $A = 19$

b) $Z = 9.$

c) $p^+ = Z = 9$

d) $e^- = Z - (\text{charge}) = Z - (-1) = 9 + 1 = 10$

e) $n^0 = A - Z = 19 - 9 = 10$

vii) $^{39}_{19}K^+$

a) $A = 39$

b) $Z = 19$

c) $p^+ = Z = 19$

d) $e^- = Z - (\text{charge}) = 19 - 1 = 18$

e) $n^0 = A - Z = 39 - 19 = 20$

viii) $^{40}_{20}Ca$

a) $A = 40$

b) $Z = 20$

c) $p^+ = Z = 20$

d) $e^- = 20$

e) $n^0 = A - Z = 40 - 20 = 20$

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Q.5. If in A^{-3}, the total no. of electrons are 18 and mass no. is 31, then find the value of Atomic no, proton and neutron

No of e^- = Z - (charge)

$18 = Z - (-3)$

$18 - 3 = Z$

$\therefore Z = 15$

No of $p^+ = 15$

No of $n^0 = A - Z = 31 - 15 = 16$

Some important definitions

i) Atomic number (Z) :

MODULE 1 (Pg 159)

ii) Mass number (A) :

MODULE 1 (Pg 157)

Mass number ≈ Atomic weight

- Mass no. of an atom is always a whole no. but atomic weight may be in decimal because atomic weight is the average of weights of all the isotops of that element.

element (Y_1, Y_2, Y_3); Mass (w_1, w_2, w_3); % or ratio (x_1, x_2, x_3)

Atomic weight : $\frac{w_1x_1 + w_2x_2 + w_3x_3}{x_1 + x_2 + x_3}$

iii) Isotops :

MODULE 1 (Pg 161)

iv) Isobars :

MODULE 1 (Pg 163)

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v) **Isotones / Isoneutronic / Isotonic :**

MODULE 1 (Pg 163)

vi) **Isodiaphers :**

MODULE 1 (Pg 163)

Common mistake :

we sometimes shifts the value of $(n^0 - p^+)$ to $(A-Z)$ which are supposed to be isotones

vii) **Isosters :**

MODULE 1 (Pg 163)

viii) **Isoelectronic :**

MODULE 1 (Pg 163)

Q.6. If in X^{+2}, the mass no (A) = 55 and the no. of protons (p^+) = 25, then find out the value of mass no (A), Atomic no (Z), no of protons (p^+) no of neutrons (n^0) and electrons (e^-) in X^{-3}

$X^{+2} \Rightarrow$ $A = 55$

$p^+ = 25 = Z$

$n^0 = A - Z = 55 - 25 = 30$

$e^- = Z - (\text{charge}) = 25 - 2 = 23.$

$X^{-3} \Rightarrow$ $A = 55$

$Z = 25$

$p^+ = 25$

$n^0 = A - Z = 55 - 25 = 30$

$e^- = Z - (\text{charge}) = 25 - (-3) = 28$

Q.7. If in X^{-2}, no. of $p^+ = 16$, then find the number of e^- in X^{+2}

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$X^{-2} \Rightarrow p^+ = 16 = Z$

$e^- = Z - (\text{charge}) = 16 - (-2) = 18$

$X^{+2} \Rightarrow e^- = Z - (\text{charge}) = 16 - 2 = 14$

$\because Z = p^+ = 16$

Q.8 If in ${}^{12}_{6}C$, atoms of the mass of the proton is halved, neutron is tripled and that of electron is doubled, then what will the mass of ${}^{12}_{6}C$ atom and by how much % will it change.

Initial ${}^{12}_{6}C \Rightarrow$ Mass of an atom = Mass of $(p^+ + n^\circ)$

no. of $p^+ = 6$

no. of $n^\circ = A - Z = 12 - 6 = 6$

$\because$ Mass of 1 proton = 1 amu

Mass of 6 proton = $1 \times 6 = 6$ amu

$\because$ Mass of 1 neutron = 1 amu

Mass of 6 neutron = $1 \times 6 = 6$ amu.

Total mass = $(6 + 6)$ amu = 12 amu

New ${}^{12}_{6}C \Rightarrow$ Mass of $p^+ = \frac{1}{2} \times 6$ amu = 3 amu

Mass of $n^\circ = 3 \times 6$ amu = 18 amu.

Total mass = $(3 + 18)$ amu = 21 amu

The mass of e^- is not taken into consideration because its mass is negligiable.

$$\text{Change in \%} = \frac{\text{New mass} - \text{Initial mass}}{\text{Initial mass}} \times 100$$

$$= \frac{21 - 12}{12} \times 100$$

$$= \frac{9}{12} \times 100 = 75\ \%$$

Hence, it increases by 75%.

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Q.9 If in $^{14}_{7}N$, the mass of p^+ is doubled and e^- is halfed and mass of neutron is same. Then what will be % change in Atomic weight

$^{14}_{7}N \Rightarrow$ no. of $p^+ = 7$

Mass of 1 p^+ = 1 amu

Mass of 7 $p^+ = 1 \times 7 = 7$ amu

no. of $n^\circ = 14 - 7 = 7$

Mass of 1 n° = 1 amu

Mass of 7 $n^\circ = 1 \times 7 = 7$ amu

$\therefore$ Total mass = 7 + 7 = 14 amu.

given condition.

$^{14}_{7}N \Rightarrow$ Mass of $p^+ = 2 \times 7 = 14$ amu.

Mass of n° = 7 amu.

$\therefore$ Total mass = 14 + 7 = 21 amu

$$\% \text{ change} = \frac{21 - 14}{14} \times 100$$

$$= \frac{7}{14} \times 100 = 50\%$$

Hence atomic weight increases by 50%

Q.10. If in a $^{16}_{8}O$ atom, the mass of p^+ is doubled, mass of e^- is halfed and mass of n° weigh $\frac{1}{4}$ Then, find the percentage change in the atomic weight.

$^{16}_{8}O \Rightarrow$ no. of p^+ = no of n° = 8

Mass of 1 p^+ = Mass of 1 n° = 1 amu

no of $(p^+ + n^\circ) = 8 + 8 = 16$.

mass of $(p^+ + n^\circ) = 8 + 8$ amu = 16 amu

given condition,

$^{16}_{8}O \Rightarrow$ mass of $p^+ = 2 \times 8 = 16$ amu

mass of $n^\circ = \frac{1}{4} \times 8 = 2$ amu

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mass of $(p^+ + n^0) = 16 + 2 = 18$ amu

Change in percentage $= \frac{18 - 16}{16} \times 100.$

$= \frac{2}{16} \times 100 = 12.5\%$

Hence, atomic weight increases by 12.5%

Q.11 Assuming that atomic weight of $^{12}_{6}C$ is 150 units from the atomic table, then what will be the weight of $^{16}_{8}O$ atom and $^{14}_{7}N$.

$^{12}_{6}C \Rightarrow$ no. of p^+ + no of $n^0 = 12$.

mass of $(p^+ + n^0) = 12$ amu.

given condition,

$^{16}_{8}O \Rightarrow$ no of p^+ + no. of $n^0 = 16$.

mass of $(p^+ + n^0) = 16$ amu.

$\because$ 12 amu = 150 units

1 amu $= \frac{150}{12}$ units and.

16 amu $= \frac{150}{12} \times 16 = 200$ units

$^{14}_{7}N \Rightarrow$ no. of p^+ + no of $n^0 = 14$

mass of $(p^+ + n^0) = 14$ amu

$\therefore$ 14 amu $= \frac{150}{12} \times 14 = 175$ units.

Q.12. Chlorine exist in the nature in the form of two isotops $^{35}_{17}Cl$ and $^{37}_{17}Cl$. If the ratio of their occurence in the nature is 3:1 respectively. Determine the average atomic weight of Cl.

Atomic weight $= \frac{w_1x_1 + w_2x_2}{x_1 + x_2}$

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$$\frac{35\times3 + 37\times1}{3+1} = \frac{105+37}{4}$$

$$= \frac{142}{4} = 35.5$$

Q.13. Boron has two isotops $^{10}_{5}B$ and $^{11}_{5}B$. If their relative percentage of occurence in nature is 80% and 20% respectively. then, determine its average atomic weight.

$$\text{Atomic weight} = \frac{w_1x_1 + w_2x_2}{x_1 + x_2}$$

$$= \frac{10\times80 + 11\times20}{80+20}$$

$$= \frac{800+220}{100} = \frac{1020}{100} = 10.2$$

Q.13. The average atomic weight of an element Y is 41.8. If it exist in the form of two isotops ^{40}Y and ^{42}Y. then, calculate their percentage of occurence in the nature.

$$\text{Atomic weight} = \frac{w_1x_1 + w_2x_2}{x_1 + x_2}$$

percentage of occurence = $x\%$ and $(100-x)\%$

$$41.8 = \frac{40\times x + 42\times(100-x)}{x+(100-x)}$$

$$41.8 = \frac{40x + 4200 - 42x}{100}$$

$$4180 - 4200 = -2x$$

$$-20 = -2x$$

$$\therefore x = \frac{-20}{-2} = 10\%$$

$$(100-x) = 100-10 = 90\%$$

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Q.14. An element has three isotops and their weights are 11, 12 and 13 units respectively. If their relative percentage in nature is 85, 10 and 5 respectively then find out its average atomic weight.

$$\text{Average Atomic weight} = \frac{w_1x_1 + w_2x_2 + w_3x_3}{x_1 + x_2 + x_3}$$

$$= \frac{11 \times 85 + 12 \times 10 + 13 \times 5}{85 + 10 + 5}$$

$$= \frac{935 + 120 + 65}{100}$$

$$= \frac{1120}{100} = 11.2$$

Models of atoms :

i) Thomson's model of atom (1904)

MODULE 1 (Pg 165)

- He was the first scientist to propose a detailed model of an atom.
- According to his model, electrons are embedded in positively charged sphere of an atom
- An atom is electrically neutral because it contains equal number of protons and electrons.
- This model of an atom is also known as
  - Plum - Pudding Model
  - Raisin Pudding Model
  - Water Melon Model

Drawbacks :

- The mass and the positive charge of the atom is considered to be evenly spread.

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- It represented a static model that is electrons were stable.

ii) Rutherford's α-scattering experiment:

Observation:

MODULE 1 (Pg 165)

- Most of the α-particles passed undeflected.
- Few of the α-particles got deflected through small angles.
- A few of the α-particle (1 out of 20,000) retraced back its path.

Conclusion:

- Most of the particles passed undeflected, indicating that lots of spaces in an atom was empty.
- The positively charged were not uniformly distributed, as thomson said because only few particles got deflected. The positive charge was concentrated in a small region in an atom which ~~was dot~~ later came to be known as ~~neutrons~~ nucleus.
- The size of the nucleus is much smaller as compared to the size of an atom

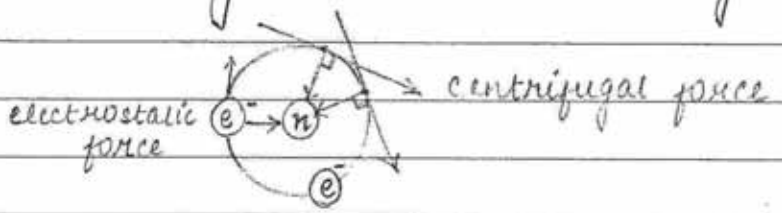
MODULE 1 (Pg 165 and 166)

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The proposed structure of an atom :

- The positive charge and mass of the entire atom is concentrated in a small region known as nucleus in the centre of the atom.
- The negatively charged electrons revolves around the nucleus in the fixed orbit in the same way like planets revolves around the sun. Hence, it is also known as 'Planetary model of an atom'.
- The electrostatic force of attraction acts between the electrons (e^-) and nucleus of an atom and this forced is balanced by centrifugal force acting on electron in tangential direction



APPLICATION of Rutherford Model :

MODULE 1 (Pg 166)

i) $\frac{\text{Radius of an atom}}{\text{Radius of nucleus}} = \frac{r_A}{r_N} = \frac{10^{-10}\ m}{10^{-15}\ m} = 10^5\ m$

$\therefore r_A = 10^5\ r_N$

Radius of an atom is 10^5 times the radius of a nucleus.

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ii) $\frac{\text{Volume of an atom}}{\text{Volume of the nucleus}} = \frac{4/3\pi(r_A)^3}{4/3\pi(r_N)^3} = \frac{(10^{-10})^3}{(10^{-15})^3}$

$= \frac{10^{-30}}{10^{-45}} = 10^{15}$

$\therefore$ Vol^m of an atom $= 10^{15}$ vol^m of the nucleus.

iii) The radius of the nucleus is proportional to the cube root of the number of nucleons (mass no.)

$\therefore R \propto A^{1/3}$

$\Rightarrow R = R_0 A^{1/3}$ cm

where $R_0 = 1.33 \times 10^{-13}$ cm (constant)

A = mass no ($p^+ + n^\circ$)

R = Radius of the nucleus.

iv) The relation between number of deflected particles and deflected angle θ.

$\therefore \mu \propto \frac{1}{\sin^4 \frac{\theta}{2}}$ (θ increases, μ decreases)

where μ = deflected particles

θ = deflected angle.

Q.15. Calculate the radius of a nucleus that contains 64 nucleons in it.

$\because R = R_0 A^{1/3}$, $R_0 = 1.33 \times 10^{-13}$ cm

$= 1.33 \times 10^{-13}$ cm $\times (64)^{1/3}$

$= 1.33 \times 10^{-13}$ cm $\times (4)^{3 \times \frac{1}{3}}$

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$= 1.33 \times 10^{-13}$ cm $\times$ 4

$= 5.32 \times 10^{-13}$ cm

Q.16. If 120 particles are deflected at an angle of 60° calculate the number of deflected particle at an angle of 90°

$$\because \mu \propto \frac{1}{\sin^4 \frac{\theta}{2}}$$

$$120 = \frac{K}{\sin^4 \frac{60^\circ}{2}} = \frac{K}{\sin^4(30^\circ)} \quad \text{— (I)}$$

$$\mu = \frac{K}{\sin^4 \frac{90^\circ}{2}} = \frac{K}{\sin^4(45^\circ)} \quad \text{— (II)}$$

Dividing (I) from (II)

$$\frac{120}{\mu} = \frac{\frac{K}{\sin^4(30^\circ)}}{\frac{K}{\sin^4(45^\circ)}}$$

$$\Rightarrow \frac{120}{\mu} = \frac{K}{\sin^4(30^\circ)} \times \frac{\sin^4(45^\circ)}{K} = \frac{(1/\sqrt{2})^4}{(1/2)^4}$$

$$\therefore \frac{120}{\mu} = \frac{1}{4} \times \frac{16}{1} = 4$$

$$\therefore \mu = \frac{120}{4} = 30$$

∴ Number of deflected particle = 30

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Drawbacks :

MODULE 1 (Pg 166 and 167)

- He was not able to explain the stability of an atom

- The loss of energy by electron was not continuous in an atom. because if electrons loses its energy continuously, the spectrum observed should be continuous but the actual observed spectrum consist of defined lines of definite frequency

- He was not able to explain the arrangement of electrons and their energy.

ELECTROMAGNETIC WAVES (EM WAVES) OR radiant energy / electromagnetic radiation :

MODULE 1 (Pg 167)

- It is a form of energy transmitted from one body to another in a continuous manner in the form of waves are called electromagnetic waves.

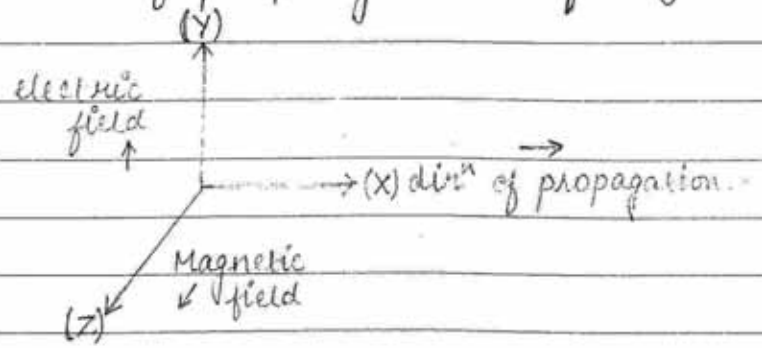
- It is of two types :
  i) Wave theory : proposed by Maxwell
  ii) Particle theory : proposed by Planck

- It does not require a material medium for proporation and it travels with the speed of the light (3×10^8 m/s).

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- If a charged particle undergoes acceleration, it ~~gets both electric~~ generates both electric as well as magnetic field which are perpendicular to each other and are also perpendicular to the direction of propagation of light



Characteristics of wave :
MODULE 1 (Pg 167)

a) Wavelength : (λ) (lambda)
It is the distance between two consecutive crest or two consecutive trough.

$1\ m = 10^2\ cm = 10^9\ nm = 10^{12}\ pm = 10^{10}\ \mathring{A}$
$1\ cm^{-1} = 10^2\ m^{-1}$ and $1\ cm = 10^{-2}\ m$

b) Frequency : (ν) (nu)
It is the number of waves which pass through a point in 1 sec.

$1\ Hz = 1\ sec^{-1} = 1\ cps.$

c) Time period : (T)
Time taken by a wave to pass through 1 point.

$T = \frac{1}{\nu}\ sec$

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d) Velocity : (c) (speed of light = 3×10^8 m/s)
It is the distance covered by a wave in 1 sec
Frequency is inversely proportional to λ.

$$\lambda \propto \frac{1}{\nu} \quad = \quad \frac{c}{\nu} \text{ (constant)}$$

c = speed of light in vaccum = 3×10^8 m/sec.

e) Wave number : ($\bar{\nu}$) (nu bar)
It is number of waves present in 1 cm and it is reciprocal of wave length.

$$\bar{\nu} = \frac{1}{\lambda}$$

$1\ cm^{-1} = 100\ m^{-1}$

f) Amplitude : (a)
It is height of crest or depth of though

Q.17 Calculate the wave number of a radiation whose frequency is 5×10^{16} Hz

$$\bar{\nu} = \frac{1}{\lambda}$$

$$\lambda = \frac{c}{\nu} = \frac{3 \times 10^8 \text{ m/s}}{5 \times 10^{16} \text{ Hz}}$$

$$\therefore \bar{\nu} = \frac{5 \times 10^{16-8}}{3 \text{ m}} = 1.6 \times 10^8\ m^{-1}$$

Q-18. A radio station is broadcasting at 100 ~~mgH~~ MHz. If the distance between the station and the receiver is 300 km. Then, how long would it take the signal to reach the receiver. Find wavelength and wave number

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$\nu = 100\ MgH = 100 \times 10^6 = 10^8\ Hz.$

$\lambda = 300\ km = 300 \times 1000 = 3 \times 10^5\ m$

Time $(T) = \dfrac{\text{displacement } (\lambda)}{\text{velocity } (c)}$

$= \dfrac{3 \times 10^5}{3 \times 10^8} = 10^{5-8} = 10^{-3}\ sec.$

$\lambda = \dfrac{c}{\nu} = \dfrac{3 \times 10^8}{3 \times 10^8} = 10^{8-5} = 10^3\ 3m$

$\bar{\nu} = \dfrac{1}{\lambda} = \dfrac{1}{3m} = 0.33\ m^{-1}$

Q-19. The wave number of a beam of light is $400\ cm^{-1}$ what is the wavelength of light in nm. also, find out the frequency of light.

$\bar{\nu} = 400\ cm^{-1}$

$\lambda = \dfrac{1}{\bar{\nu}} = \dfrac{1}{400}\ cm = \dfrac{1}{400} \times 10^{-2}\ m$

$= \dfrac{1}{400} \times 10^{-2} \times 10^9\ nm.$ $\{1\ cm = 10^{-2}\ m,\ 1\ m = 10^9\ nm\}$

$\lambda = \dfrac{10^7}{4 \times 10^2}\ nm = \dfrac{10^{7-2}}{4} = \dfrac{10^5}{4} = 25000\ nm.$

$\nu = \dfrac{c}{\lambda} = \dfrac{3 \times 10^8\ m/s}{25000\ nm} = \dfrac{3 \times 10^8\ m/s}{25 \times 10^3 \times 10^{-9}\ m}$

$= \dfrac{3 \times 10^8\ m/s}{25 \times 10\ m} = \dfrac{3 \times 10^8}{25 \times 10^{-6}}\ s.$

$= 0.12 \times 10^{14} = 1.2 \times 10^{13}\ s.$

Q.20 Calculate the wave number and the frequency of a radiation whose wavelength is 6000 Å.

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$\lambda = 6000\ \text{Å}$

$1\ \text{Å} = 10^{-10}\ m$

$\therefore 6000\ \text{Å} = 6000 \times 10^{-10} = 6 \times 10^{-7}\ m.$

$\bar{\nu} = \frac{1}{\lambda} = \frac{1}{6 \times 10^{-7.3}\ m} = 0.16 \times 10^{7}\ m^{-1}$

$= 1.66 \times 10^{6}\ m^{-1}$

$\upsilon = \frac{c}{\lambda} = \frac{3 \times 10^{8}\ m/s}{6 \times 10^{-7}\ m}$

$= 0.5 \times 10^{15}\ Hz$

$= 5 \times 10^{14}\ Hz$

The electromagnetic spectrum:

MODULE 1 (Pg 168)

- on the basis of increasing wavelength.

cosmic ray's < γ rays < x-rays < Ultra-violet < visible spectrum < Infrared < Micro wave < Radio wave.

Planck's Quantum theory:

MODULE 1 (Pg 168)

- The radiant energy is emitted or absorbed discontinuously by a body in the form of small discrete parti packects of energy called quantum.

- In case of light, the smallest packet is 'photon'.

energy of 1 mole quantum:

$E = N h\nu = \frac{N h c}{\lambda}$ $1\ mol^{-1}$ or $erg\ mol^{-1}$

$\therefore N = 6.023 \times 10^{23}$.

{'mol^{-1}' signifies that, we need to calculate in 1 mole so instead of n, it will be $N = 6.023 \times 10^{23}$}

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$$E = \frac{2.859}{\lambda \text{ cm}} \text{ (calorie mol}^{-1}\text{)}$$

Q.21 Calculate the wavelength of photon whose energy is 1 electron volt.

1 electron volt (eV) $= 1.6 \times 10^{-19}$ J

$\because E = \frac{hc}{\lambda} \quad \therefore \lambda = \frac{hc}{E}$

$$= \frac{6.626 \times 10^{-34} \times 3 \times 10^{8}}{1.6 \times 10^{-19}}$$

$$= \frac{19.878 \times 10^{-7}}{1.6} = 12.4 \times 10^{-7} \text{ m.}$$

Q.22 If the wavelength of a photon whose energy is 'E' is 6000 Å, then find out the wavelength of a photon whose energy is 3E.

$$E_1 = \frac{hc}{\lambda_1} \quad \therefore \frac{E_1}{E_2} = \frac{\lambda_2}{\lambda_1}$$

$$E_2 = \frac{hc}{\lambda_2}$$

$$\frac{E}{3E} = \frac{\lambda_2}{6000} \Rightarrow \frac{1}{3} = \frac{\lambda_2}{6000}$$

$$\therefore \lambda_2 = \frac{6000}{3} \text{ Å} = 2000 \text{ Å}$$

Q.23. 3×10^{8} photons of a certain light radiation are found to produce 1.5 joule of energy. Calculate the wavelength of the radiation. (m)

$\because E = \frac{nhc}{\lambda} \quad \therefore \lambda = \frac{nhc}{E}$

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$$\lambda = \frac{6.626 \times 10^{-34} \times 3 \times 10^{8} \times 3 \times 10^{8}}{1.5}$$

$$= \frac{6.626 \times 9 \times 10^{-18}}{1.5} \text{ m}$$

$$= 39.7 \times 10^{-18} \text{ m}$$

24. A 100 watt bulb emits monochromatic light of wavelength 1986 nm. Calculate the no of photon emitted per sec by the bulb.

1 watt = 1 J/sec

∴ 100 watt = 100 J/sec.

1 nm = 10^{-9} m

∴ 1986 nm = 1986×10^{-9} m

$$E = \frac{nhc}{\lambda}$$

$$100 = n \times \frac{6.626 \times 10^{-34} \times 3 \times 10^{8}}{1986 \times 10^{-9}}$$

$$1986 \times 10^{-9} \times 10^{2} = 6.626 \times 3 \times 10^{-26} \times n$$

$$1986 \times 10^{-7} \lambda = 19.86 \times 10^{-26} \times n$$

$$19.86 \times 10^{-5} \times = 10^{-26} \times 19.86 \times n$$

$$\therefore n = \frac{10^{-5}}{10^{-26}} = \frac{10^{26}}{10^{5}}$$

$$n = 10^{26-5} = 10^{21}$$

Unit convention.

- 1 calorie = 4.2 J
- 1 Joule = 10^{7} erg
- 1 electron volt = 1.6×10^{-19} J

$E = \frac{nhc}{\lambda}$, unit of λ and c should be in m and m/s respectively. and unit of h should be according to energy.

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Bohr's atomic model :

MODULE 1 (Pg 169)

Some important formulae:

Coulombic force: $\frac{Kq_1q_2}{r^2}$

Centrifugal force: $\frac{mv^2}{r}$

Angular momentum : mvr

Bohr's Postulates :

MODULE 1 (Pg 169).

It includes six postulates.

- nth excited state will represent (n+1)th energy level

  ΔE = $E_{higher} - E_{lower}$ = energy of a quantum.

- During transition, definit amount of energy is released and absorbed.

Application of Bohr's model

i) Radius of various orbits (shell):

MODULE 1 (Pg 170)

$$r_n = 0.529 \times \frac{n^2}{Z} \text{ Å}$$

where, n = no of shells and

Z = atomic no.

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Case I : If n is constant, $r \propto \frac{1}{n}$

$$\therefore \frac{r_1}{r_2} = \frac{Z_2}{Z_1}$$

Case II : If Z is constant, $r \propto Z$

$$\therefore \frac{r_1}{r_2} = \frac{(n_1)^2}{(n_2)^2}$$

Case III : neither (n) nor (Z) is constant

$$\therefore \frac{r_1}{r_2} = \frac{(n_1)^2}{(n_2)^2} \times \frac{Z_2}{Z_1} \text{ or } \frac{(n_1)^2}{Z_1} \times \frac{Z_2}{(n_2)^2}$$

25. If the radius of the 1st orbit of hydrogen atom is x Å, then the radius of its 3rd orbit will be.

$$\frac{r_1}{r_2} = \frac{(n_1)^2}{(n_2)^2}$$

$$\Rightarrow \frac{x}{r_2} = \frac{(1)^2}{(3)^2} = \frac{1}{9}$$

$$\therefore r_2 = 9x.$$

$\therefore$ radius of 3rd orbit $= 9x$

26 Calculate the radius of 3rd orbit of Li^{+2}

$$r_n = \left(0.529 \times \frac{n^2}{Z}\right) \text{Å}$$

$$= \left(0.529 \times \frac{3^2}{3}\right) \text{Å}$$

$$= 0.529 \times \frac{9}{3} = 1.578 \text{ Å}$$

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Q.27 Calculate the ratio of radius of 2nd to 3rd orbit for hydrogen atom.

$$\frac{r_1}{r_2} = \frac{(n_1)^2}{(n_2)^2}$$

$$\frac{r_1}{r_2} = \frac{(2)^2}{(3)^2} = \frac{4}{9} = 4:9$$

$$r_1 : r_2 = 4:9$$

Q.28 Calculate the radius ratio of 3rd orbit of hydrogen atom to 2nd orbit of He^+

$$\frac{r_1}{r_2} = \frac{(n_1)^2}{z_1} \times \frac{z_2}{(n_2)^2}$$

$$= \frac{(3)^2}{1} \times \frac{2}{(2)^2} = \frac{9 \times 2}{4} = \frac{9}{2}$$

$$\therefore r_1 : r_2 = 9:2$$

Q.29 Calculate the diameter and circumference of 3rd orbit of Li^{+2}.

$$r_n = \left(0.529 \times \frac{n^2}{z}\right) \text{Å}$$

$$= 0.529 \times \frac{(3)^2}{3} = \frac{0.529 \times 9}{3}$$

$$r_n = 1.587 \text{ Å}$$

diameter $= r_n \times 2 = (1.587 \times 2)$ Å $= 3.174$ Å

circumference $= 2\pi r$

$= (2 \times 3.14 \times 1.587)$ Å

$= 9.965$ Å

Q.30. The ratio of radius of two bore orbit for hydrogen atom is 4:1, then the orbit will be.

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i) K and L
ii) N and L
iii) L and K
iv) Both (ii) and (iii)
Ans: iv)

$$\frac{r_1}{r_2} = \frac{(n_1)^2}{(n_2)^2} = \frac{(2)^2}{(1)^2} = 4:1$$

31 Find out the ratio of radius of 2nd orbit of Li^{+2} to 3rd orbit of He^+

$$\frac{r_1}{r_2} = \frac{(n_1)^2}{Z_1} \times \frac{Z_2}{(n_2)^2}$$

$$\frac{r_1}{r_2} = \frac{(2)^2}{3} \times \frac{2}{(3)^2} = \frac{4}{3} \times \frac{2}{9} = \frac{8}{27}$$

$\therefore r_1 : r_2 = 8 : 27$

32 Find out the ratio of circumference of 1st excited state of Li^{+2}, 2nd excited state of H and 3rd excited state of He^+.

radius of $Li^{+2} : H : He^+$

$$0.529 \frac{n^2}{Z} = \frac{(2)^2}{3} : \frac{(3)^2}{1} : \frac{(4)^2}{2}$$

$$= \frac{4}{3} : \frac{9}{1} : \frac{16}{2}$$

circumference of $Li^{+2} : H : He^+$

$$2\pi r = 2\pi \times 0.529 \times \frac{4}{3} : 2\pi \times 0.529 \times 9 : 2\pi 0.529 \times 8$$

$$= \frac{4}{3} : 9 : 8$$

$$= 4 : 27 : 24$$

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Q.33 Find out the ratio of radius of 2nd orbit of hydrogen to 1st orbit of Li^{+2} to 3rd orbit of He^+.

Radius of H : Li^{+2} : He^+

$$0.529 \times \frac{n^2}{Z} = \frac{(2)^2}{1} : \frac{(1)^2}{3} : \frac{(3)^2}{2}$$

$$= 4 : \frac{1}{3} : \frac{9}{2}$$

$$= 24 : 2 : 27$$

Q.34 Calculate the angular momentum of an electron in 4th orbit of He^+.

Angular velocity = $mvr = \frac{nh}{2\pi}$

$$= \frac{4h}{2\pi} = \frac{2h}{\pi}$$

ii) Velocity of electron in Bohr orbit:

MODULE 1 (Pg 171)

$$v = \left(2.18 \times 10^6 \frac{Z}{n}\right) m/s.$$

OR

$$v = \left(2.18 \times 10^8 \frac{Z}{n}\right) cm/s$$

where, n = no. of shells and

Z = atomic no.

Case I: If Z is constant, $v \propto \frac{1}{n}$

$$\therefore \frac{v_1}{v_2} = \frac{n_2}{n_1}$$

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Case II: If n is constant, $v \propto Z$

$$\frac{v_1}{v_2} = \frac{Z_1}{Z_2}$$

Case III: neither (n) nor (Z) is constant.

$$\frac{v_1}{v_2} = \frac{Z_1}{n_1} \times \frac{n_2}{Z_2}$$

Q.35 Calculate the ratio of velocity of electron in 2nd orbit of Li^{+2} to 3rd orbit of He^{+}

$$\frac{v_1}{v_2} = \frac{Z_1}{n_1} \times \frac{n_2}{Z_n}$$

$$= \frac{3}{2} \times \frac{3}{B2} = \frac{9}{4} = 9:4$$

$\therefore v_1 : v_2 = 9:4$

Q.36 Calculate the velocity of an electron in 3rd orbit of Li^{+2}

$$v = \left(2.18 \times 10^6 \frac{Z}{n}\right) m/s$$

$$= \left(2.18 \times 10^6 \times \frac{3}{3}\right) m/s.$$

$$= 2.18 \times 10^6 \ m/s.$$

Q.37 If v_1, v_2, v_3 and v_4 are the velocities of an electron in 1st orbit of H, He^{+}, Li^{+2}, Be^{+3} respectively, then pick out the correct relation.

i) $v_1 < v_2 < v_3 < v_4$

ii) $v_1 > v_2 > v_3 > v_4$

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iii) $v_1 = v_2 = v_3 = v_4$

iv) None of these.

Ans) i)

$$v_1 : v_2 : v_3 : v_4$$

$$\frac{z_1}{n_1} : \frac{z_2}{n_2} : \frac{z_3}{n_3} : \frac{z_4}{n_4}$$

$\because n$ = constant

$\therefore v \propto z$

$1 : 2 : 3 : 4$

$\therefore v_1 < v_2 < v_3 < v_4$

Q. 38 If v_1, v_2, v_3 and v_4 are the velocities of an electron in 1st, 2nd, 3rd and 4th orbit of H atom respectively. then pick out the correct relation.

i) $v_1 < v_2 < v_3 < v_4$

ii) $v_1 > v_2 > v_3 > v_4$

iii) $v_1 = v_2 = v_3 = v_4$

iv) None of these

Ans) ii

$$v_1 : v_2 : v_3 : v_4$$

$\because z$ = constant

$$\therefore v \propto \frac{1}{n}$$

$\therefore v_1 > v_2 > v_3 > v_4$

Q-39. If v_1, v_2, v_3 and v_4 are the velocities of an electrons in 1st orbit of H, 2nd orbit of He, 3rd Orbit of Li^{+2} and 4th Orbit of beryllium, then pick out the correct

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i) $v_1 < v_2 < v_3 < v_4$

ii) $v_1 > v_2 > v_3 > v_4$

iii) $v_1 = v_2 = v_3 = v_4$

iv) None of these

Ans) iii)

$$v_1 : v_2 : v_3 : v_4$$

Here, $\frac{Z}{n}$ = constant.

$$\therefore \frac{v_1}{n_1} : \frac{v_2}{n_2} : \frac{v_3}{n_3} : \frac{v_4}{n_4} = K$$

$$\therefore v_1 = v_2 = v_3 = v_4$$

iii) Total energy of electron in Bohr model:

MODULE 1 (Pg 171)

Potential energy: $\frac{-KZe^2}{r}$

Kinetic energy: $\frac{KZe^2}{2r}$

Total energy: Potential energy + Kinetic energy

$$= \frac{-KZe^2}{r} + \frac{KZe^2}{2r}$$

$$= \frac{-KZe^2}{2r}$$

$$\therefore T.E = -KE = \frac{PE}{2}$$

$$T.E = (-2.18 \times 10^{-18}) \times \frac{Z^2}{n^2} \text{ J/Atom}$$

$$= -13.6 \times \frac{Z^2}{n^2} \text{ eV/Atom}$$

$$\therefore E(\text{energy}) \propto -\frac{Z^2}{n^2}$$

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$= \frac{-13.6 \times 1^2}{2^2} = \frac{-13.6}{4}$ eV/atom.

$= -3.4$ eV/atom

Energy of an electron in 3rd orbit

$= \frac{-13.6 \times 1^2}{3^2} = \frac{-13.6}{9}$ eV/atom

$= -1.51$ eV/atom

Energy of an electron in 4th orbit

$= \frac{-13.6 \times 1^2}{4^2} = \frac{-13.6}{16}$ eV/atom

$= -0.85$ eV/atom

Energy of an electron in 5th orbit

$= \frac{-13.6 \times 1^2}{5^2} = \frac{-13.6}{25}$ eV/atom

$= -0.54$ eV/atom

Energy of an electron at infinity.

$= \frac{-13.6 \times 1^2}{\infty^2} = 0$

Infinity ($n=\infty$) — 0

4th excited state (O) ($n=5$) — −0.54 eV

3rd excited state (N) ($n=4$) — −0.85 eV

2nd excited state (M) ($n=3$) — −1.51 eV

1st excited state (L) ($n=2$) — −3.4 eV

ground state (K) ($n=1$) Most stable — −13.6 eV

Minimum energy

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$$T.E. \propto -\frac{Z^2}{n^2}$$

Case I: If n is constant, $E \propto -Z^2$

$\therefore$ # $Z \uparrow \Rightarrow E \downarrow$

$$\therefore \frac{E_1}{E_2} = \frac{(Z_1)^2}{(Z_2)^2}$$

Case II: If Z is constant, $E \propto -\frac{1}{n^2}$

$\therefore$ $n \uparrow \Rightarrow E \uparrow$

$$\therefore \frac{E_1}{E_2} = \frac{(n_2)^2}{(n_1)^2}$$

Case III: neither (n) nor (Z) is constant

$$\frac{E_1}{E_2} = \frac{(Z_1)^2}{(n_1)^2} \times \frac{(n_2)^2}{(Z_1)^2}$$

Q. 40. calculate the energy of an electron in 1st 2nd, 3rd, 4th and 5th orbit of H atom. And also calculate its energy at infinity.

Energy of an electron in 1st Orbit:

$$-13.6 \times \frac{Z^2}{n^2} \text{ eV/atom}$$

$$= -13.6 \times \frac{1^2}{1^2} = -13.6 \text{ eV/atom}$$

Energy of an electron in 2nd Orbit:

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Energy difference :

- (K-L) shell : $-3.4-(-13.6) = 13.6-3.4 = 10.2$ eV
- (L-M) shell : $-1.51-(-3.4) = 3.4-1.51 = 1.89$ eV
- (M-N) shell : $-0.85-(-1.51) = 1.51-0.85 = 0.66$ eV
- (N-O) shell : $-0.54-(-0.85) = 0.85-0.54 = 0.31$ eV

a) Ionization energy :
MODULE 1 (Pg 169)

Minimum amount of energy required to remove an electron from its ground state to infinity.

$IE = E_\infty - E_1$
$= 0 - E_1 = -E_1$ (Kinetic energy in 1st she

For eg: In case of hydrogen atom.
in 1st shell, $-(-E_1) = -(-13.6)$ eV
$= +13.6$ eV

- Ionization energy is equal to the energy of an electron in its ground state with a reverse sign.

- Ionization energy is always positive (+ve).

b) Seperation energy :
MODULE 1 (Pg 170)

Minimum amount of energy required to remove an electron from its excited state to infinity is called seperation energy

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$SE = E_\infty - E_n \quad [n = 2, 3, 4, 5 \ldots\ldots]$

$0 - E_n = -E_n$

For eg: In case of hydrogen atom

1st S.E $= -E_2 = +3.4\ eV$

2nd S.E $= -E_3 = +1.51\ eV$

- The energy of an electron in that state from which it is removed with a reverse sign.

- Seperation energy is always positive (+ve).

c) Excitation Energy :

MODULE 1 (Pg 170)

The amount of energy required to shift an electron from ground state to any other excited state.

$EE = E_n - E_1 \quad [n = 2, 3, 4, 5 \ldots\ldots]$

#

For eg: In case of hydrogen atom

1st $EE = E_2 - E_1 = -3.4 - (-13.6)$

$= 13.6 - 3.4 = 10.2\ eV$

2nd $EE = E_3 - E_1 = -1.51 - (-13.6)$

$= 13.6 - 1.51 = 12.09\ eV$

3rd $EE = E_4 - E_1 = -0.85 - (-13.6)$

$= 13.6 - 0.85 = 12.75\ eV$

* The total energy of an electron is always (-ve) ~~positive~~ and its maximum value is at infinity. ($E_\infty = 0$)

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For any particular species : (applicable of all)

- $TE_1 < TE_2 < TE_3 < TE_4$
- $KE_1 > KE_2 > KE_3 > KE_4$
- $PE_1 < PE_2 < PE_3 < PE_4$
- $E_2 - E_1 > E_3 - E_2 > E_4 - E_3 > E_5 - E_4$

* TRICKS : (for any species)

i) E_n (any species) $= E_n$ (H atom) $\times z^2$

ii) $\Delta E = (E_2 - E_1)$ (any species) $= \Delta E\,(E_2 - E_1)$ (H atom) $\times z^2$

iii) $E_n = \frac{E_1}{n^2}$ * [Do not use either of the tricks in ratio]

Q. 41. Calculate the energy of an electron in 2nd orbit of Li^{+2}.

$$E = -13.6 \times \frac{z^2}{n^2} = -13.6 \times \frac{9}{4}\ eV$$

$$= -30.6\ eV \quad \underline{OR}$$

Using trick (1),

$$E_2\ (Li^{+2}) = E_2\ (H) \times z^2$$

$$= -3.4 \times 9 = -30.6\ eV$$

Q.42 The energy of an e^- in 1st Orbit of H-atom is $-E$ eV. Calculate the energy in 3rd orbit

$$\frac{E_1}{E_2} = \frac{(n_2)^2}{(n_1)^2}$$

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$$\frac{-E}{E_2} = \frac{9}{1}$$

$$\therefore E_2 = \frac{-E}{9} \quad \text{OR}$$

Using trick (iii)

$$E_n = \frac{E_1}{n^2} \Rightarrow E_2 = \frac{-E}{9}$$

43. Calculate the difference of energy between 3rd and 2nd Orbit of He^+.

$$\Delta E = E_3 - E_2$$

$$= -\frac{13.6 \times 4}{9} - \left(-\frac{13.6 \times 4}{4}\right)$$

$$= -13.6\left[\frac{4}{9} - \frac{4}{4}\right]$$

$$= -13.6 \times \frac{16-36}{36} = -13.6 \times \frac{(-20)}{36}$$

$$= 13.6 \times \frac{5}{9} = 7.56 \text{ eV} \quad \text{OR}$$

Using trick (ii).

$$\Delta E_{(3-2)}^{He^+} = \Delta E_{(3-2)}^{H} \times Z^2$$

$$= -1.51 - (-3.4) \times 4$$

$$= (3.4 - 1.51) \times 4 = 7.56 \text{ eV}$$

44 Calculate the ratio of ΔE between 3rd and 2nd orbit of He^+ to 2nd ~~orbit of Li^{+2}~~ and 1st Orbit of Li^{+2}

$$\frac{E_3 - E_2}{E_2 - E_1} = \frac{-13.6 \times Z^2/n^2 + 13.6\, Z^2/n^2}{-13.6 \times Z^2/n^2 + 13.6\, Z^2/n^2}$$

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$\Delta E_1\ (E_3 - E_2) = -13.6 \times \frac{4}{9} + 13.6 \times \frac{4}{4}$
(He^+)

$\Delta E_2\ (E_2 - E_1) = -13.6 \times \frac{9}{4} + 13.6 \times \frac{9}{1}$
(Li^{+2})

$\therefore \frac{E_3 - E_2}{E_2 - E_1} = \frac{-\frac{4}{9} + 1}{-\frac{9}{4} + 9.}$

$= \frac{5}{9} \times \frac{4}{27} = \frac{20}{243.}$

$\frac{\Delta E\ (He^+)}{\Delta E\ (Li^{+2})} = \frac{20}{243} \Rightarrow 20 : 243.$

Q.45. Calculate the energy of an electron in ground state of H, He^+ and Li^{+2}

- $E_1\ (H) = -13.6 \times \frac{1}{1} = -13.6$ eV
- $E_1\ (He^+) = -13.6 \times \frac{4}{1} = -54.4$ eV
- $E_1\ (Li^{+2}) = -13.6 \times \frac{9}{1} = -122.4$ eV

Q.46 Calculate the ionization energy of H, He^+, Li^{+2}.

- $-(-E_1)_H = -(-13.6\ eV) = +13.6$ eV
- $-(-E_1)_{He^+} = -(-13.6 \times 4) = +13.6 \times 4$
  $= +54.4$ eV
- $-(-E_1)_{Li^{+2}} = -(-13.6 \times 9) = +13.6 \times 9$
  $= +122.4$ eV

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Q.47 Calculate the energy of an e^- in IIIrd excited state of Li^{+2}

$$E = -13.6 \times \frac{Z^2}{n^2}$$

$$= -13.6 \times \frac{9}{16} = -7.65 \text{ eV}$$

Q.48 Calculate the 2nd excitation energy of He^+.

2nd excitation energy $= E_3 - E_1$

$$\Delta E \ (E_3 - E_1) = -\frac{13.6 \times 4}{9} - \left(\frac{-13.6 \times 4}{1}\right)$$

$$= +13.6 \times 4 - \frac{13.6 \times 4}{9}$$

$$= 13.6 \times 4 \left[1 - \frac{1}{9}\right] = 13.6 \times 4 \times \frac{8}{9}$$

$$= 1.51 \times 32 = 48.36 \text{ eV} \quad \text{OR}$$

Using trick (ii)

$$\Delta E_{(3-2)}^{He^+} = \Delta E_{(3-2)}^{H} \times Z^2$$

$$= 12.09 \times 4 = 48.36 \text{ eV}$$

Q.49 How much minimum energy should be absorbed by an e^- of H atom in ground state to reach the excited state.

a) 13.6 eV
b) 3.4 eV
c) 12.09 eV
d) 10.2 eV

Ans) d)

1st excited state $= E_2$

$\therefore \Delta E_{(2-1)} = -3.4 - (-13.6)$

$= 13.6 - 3.4 = 10.2$ eV

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Q.50 In which of the following condition transition, minimum energy is released?

i) $n = 1$ to $n = 2$

ii) $n = 1$ to $n = \infty$

iii) $n = 3$ to $n = 1$

iv) $n = 3$ to $n = 2$

Ans) iv)

Released: transition from higher energy level to lower energy level.

Q.51 In which of the following transition, radiation of maximum energy is absorbed?

i) $n = 1$ to $n = 2$

ii) $n = 1$ to $n = 3$.

iii) $n = \infty$ to $n = 1$

iv) $n = 1$ to $n = \infty$

Ans) iv)

Absorbed: transition from lower energy level to higher energy level.

Q.52 The ionization energy of an atom is 100 eV then calculate,

i) Energy of 2nd orbit.

ii) 1st excitation energy

iii) Amount of energy required to excite an e^- from $n = 1$ to $n = 2$

iv) calculate the frequency of photon absorbed when e^- excites from $n = 1$ to $n = 2$

$IE = 100$ eV

$-E_1 = 100$ eV ($\because IE = -E_1$)

$\therefore E_1 = -100$ eV

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i) $E_n = \frac{E_1}{n^2} = \frac{-100}{4} = -25\ eV\ (E_2)$.

ii) 1st excited state = $E_2 - E_1$
$= -25\ eV - (-100\ eV)$.
$= 100\ eV - 25\ eV = 75\ eV$

iii) $E_2 - E_1 = 75\ eV$

iv) $\because \Delta E = h\nu$
$\therefore \nu = \frac{\Delta E}{h} = \frac{(E_2 - E_1)(1.6 \times 10^{-19})}{6.626 \times 10^{-34}}$
$\nu = \frac{75 \times 1.6 \times 10^{15}}{6.626}$
$= 18 \times 10^{15}\ Hz$

Q.53. The ionization energy of H is $1.3 \times 10^6\ J\ mol^{-1}$. Calculate the ~~frequency~~ energy required to excite an e^- from $n=1$ to $n=2$

$IE = 1.3 \times 10^6\ J\ mol^{-1}$
$\therefore E_1 = -1.3 \times 10^6\ J\ mol^{-1}$ [$\because IE = -E_1$]

Using trick (iii)

$E_n = \frac{E_1}{n^2} = \frac{-1.3 \times 10^6}{4}$

$E_2 = -0.328 \times 10^6\ J\ mol^{-1}$

$\therefore$ Required energy = $E_2 - E_1$

$= (-0.328 \times 10^6) - (-1.3 \times 10^6)\ J\ mol^{-1}$
$= 1.3 \times 10^6 - 0.328 \times 10^6$
$= 10^6 (1.3 - 0.328)$
$= 0.984 \times 10^6\ J\ mol^{-1}$

Q.54 If the 2nd exciation energy of an atom is 1210 eV, then find out its atomic no.

$E_3 - E_1 = 1210\ eV$

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Using trick (ii),

$$\Delta E_{(3-1)}^{'Z'} = \Delta E_{(3-1)}^{H} \times Z^2$$

$$1210 = -1.51 - (-13.6) \times Z^2$$

$$1210 = 13.6 - 1.51 \times Z^2.$$

$$\therefore Z^2 = \frac{12.09}{1210} \simeq 100.$$

$$\therefore Z = \sqrt{100} = 10.$$

Hence, atomic no = 10

Q.55 Calculate the potential energy of an e^- in IInd excited state of He^+

$$TE_n^{He^+} = TE_n^{H} \times Z^2 \quad \text{(Using trick i.)}$$

$$= -1.51 \times 4 = -6.04 \text{ eV}$$

$$\because \frac{PE}{2} = TE$$

$$\therefore PE = 2TE$$

$$= 2 \times (-6.04) = -12.08 \text{ eV and.}$$

$$KE = -TE = -6.04 \text{ eV}$$

Q.56 ☆ If the potential energy of an e^- in IInd orbit is for Li^{+2} is x, then find out the total energy for He^+ in the 2nd orbit.

$$\because \frac{PE}{2} = TE$$

$$(PE_2)_{Li^{+2}} = x$$

$$\therefore (TE_2)_{Li^{+2}} = \frac{x}{2}$$

Using, $\frac{E_1}{E_2} = \frac{(Z_1)^2}{(Z_2)^2}$

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$\Rightarrow \frac{E_1 (Li^{+2})}{E_1 (He^+)} = \frac{9}{4}$

$\Rightarrow \frac{x/2}{E_1 (He^+)} = \frac{9}{4}$

$\therefore E_1 (He^+) = \frac{x}{2} \times 4 \times \frac{1}{9} = \frac{2x}{9}$

$(TE_2)_{He^+} = \frac{2x}{9}$

Q.57 Calculate the 1st seperation energy for Li^{+2}

$\Delta E = -(-E_n)$, where $n = 2$

$= -(-E_2)$ (1st seperation energy).

$= +E_2$

$(TE_2)_{Li^{+2}} = (TE_2)_H \times Z^2$ (Using trick i)

$= -3.4 \times 9 = -30.6$ eV

$\therefore$ 1st seperation energy $= -(-TE_2)_{Li^{+2}}$

$= -(-30.6)$

$= 30.6$ eV

Q.58 The potential energy of an e^- in H atom is -6.8 eV then determine that the e^- will be present in which excited state.

$PE = -6.8$ eV

$\therefore TE = \frac{-6.8}{2} = -3.4$ eV $\left(\because TE = \frac{PE}{2}\right)$

$E_n = \frac{-13.6 \times Z^2}{n^2}$ $\therefore n^2 = \frac{-13.6 \times 1}{-3.4} = 4$

$n^2 = 4$ $\therefore n = \sqrt{4} = 2$, $n = 2$, 1st E.S.

Using trick i,

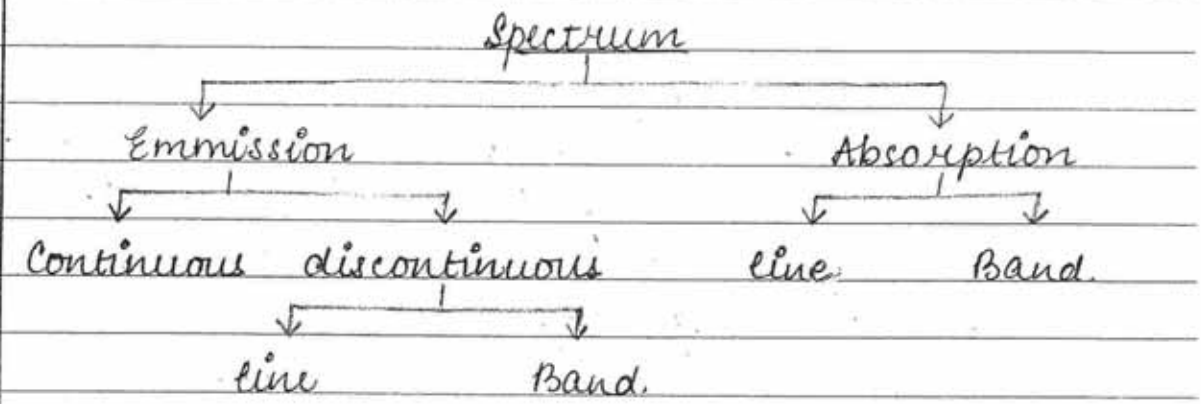
$TE_1 = \frac{-6.8}{2} = -3.4$ eV, (1st excited state)

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Spectrum:

When a radiation is passed through a prism, it splits into radiations (electromagnetic waves) of various wavelength. When these radiations are made to fall on a photographic film, the impression produced is known as spectrum.



i) **Emmission spectrum:**

When a radiation from light-emiting source like sun, bulb, burner, a gas in the discharged tube, a heated substance etc is passed through a prism and then made to fall on a photo-graphic film, the impression so obtained is known as the emmission spectrum.

For eg:

- The spectrum of sunlight (white light) is a continuous emmission spectrum.
- The spectrum of atoms is a line emmission spectrum. (discontinuous)
- The spectrum of molecule is a band emmission spectrum. (discontinuous)

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* The spectrum of an element is the fingerprint of that element i.e. the spectrum of two element can't be same.

* The spectrum of same electron species can be ~~same~~ similar but can't be same.

ii) Absorption Spectrum:
When a radiation is passed through an absorbing sample and then it is made to ~~fall~~ pass through a prism and then made to fall ~~upon~~ a photographic film. The impression obtained is known as absorption spectrum.

- when a radiation is passed through an atomic gas, then the obtained spectrum is absorbed line spectrum.
- when a radiation is passed through a molecule, then the obtained spectrum is absorbed band spectrum.

* In absorption spectrum, dark lines are obtained in otherwise bright spectrum
* These dark line corresponds to the wavelength absorbed by the sample.
* The absorbed spectrum can be called the photo-negative of the initial spectrum.



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Hydrogen spectrum (emmission line)

MODULE 1 (Pg 172)

- Total no. of spectral line (TSL) where e^- tends from n_2 to n_1 (emmission line)

$$\therefore \text{TSL} = \frac{(n_2 - n_1)(n_2 - n_1 + 1)}{2}$$

- For any transition:

i) 1st line is called 'α' line
$\therefore n_2 = n_1 + 1$

ii) 2nd line is called 'β' line
$\therefore n_2 = n_1 + 2$

iii) 3rd line is called 'γ' line
$\therefore n_2 = n_1 + 3$.

iv) For any 'x' line:
$\therefore n_2 = n_1 + x$.

- Last / Marginal / Limiting line:
$\therefore n_2 = \infty$

- Criteria / Basis : 1st line ↓ and last line ↓

i)	Energy : (E) :	Minimum	Maximum
ii)	Frequency : (υ) :	Minimum	Maximum
iii)	Wavelength : (λ) :	Maximum	Minimum

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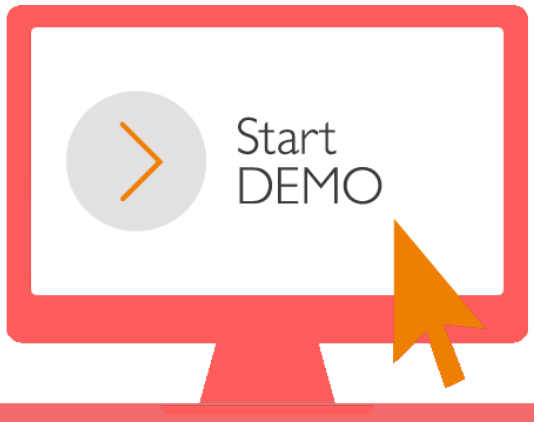
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59. Calculate total no of spectral lines when an e^- transits from 6th energy level to ground state. Also calculate the no of lines in each series and also determine the no. of lines formed in ultra, visible and Infrared region

- $TSL = \frac{(n_2 - n_1)(n_2 - n_1 + 1)}{2}$

$= \frac{(6-1)(6-1+1)}{2}$

$= \frac{5 \times 6}{2} = 15$ lines.

- No. of lines in each series:

i) Lyman's series (UV regions): $\because n_2 = n_1 + 1$

$\therefore n_1 = n_2 - 1 \Rightarrow 6 - 1 = 5$

ii) Balmer's series (visible regions): $\because n_2 = n_1 + 2$

$\therefore n_1 = n_2 - 2 \Rightarrow 6 - 2 = 4$

iii) Paschen's series (IR region): $n_2 = n_1 + 3$.

$\therefore n_1 = n_2 - 3 \Rightarrow 6 - 3 = 3$.

iv) Brackett series: $n_1 = n_2 - 4 \Rightarrow 6 - 4 = 2$

v) Pfund series: $n_1 = n_2 - 5 \Rightarrow 6 - 5 = 1$

vi) Humphery series: $n_1 = n_2 - 6 \Rightarrow 6 - 6 = 0$

- Total number of lines in ultra region: 5
- Total number of lines in visible regions: 4
- Total number of lines in Infrared regions:

$3 + 2 + 1 = 6$

60. Calculate total no of spectral line when an e^- transits from 7th energy level to 1st excited state. Also, calculate the no of lines in each series and also determine the number of lines

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formed in ultra, visible and Infrared regions

- $T.S.L = \frac{(n_2 - n_1)(n_2 - n_1 + 1)}{2}$

$= \frac{(7-2)(7-2+1)}{2}$

$= \frac{5 \times 6}{2} = 15$ lines

- No of lines in each series:

i) Lyman's series: $n_1 = n_2 - 1 = 0$ (Lyman's series)

ii) Balmer's series: $n_1 = n_2 - 2 = 7 - 2 = 5$

iii) Paschen's series: $n_1 = n_2 - 3 = 7 - 3 = 4$

iv) Brackett series: $n_1 = n_2 - 4 = 7 - 4 = 3$

v) Pfund series: $n_1 = n_2 - 5 = 7 - 5 = 2$

vi) Humphery series: $n_1 = n_2 - 6 = 7 - 6 = 1$

- Total no. of lines in ultra regions = 0
- Total no. of lines in visible regions = 5
- Total no. of lines in Infrared regions = $4 + 3 + 2 + 1 = 10$

Q.61 Calculate total no. of spectral line when an e^- transmits from 6th energy level to 3rd energy levels. Also calculate the no. of lines in each series and also determine the no of lines formed in ultra, visible and Infrared regions.

- $T.S.L = \frac{(n_2 - n_1)(n_2 - n_1 + 1)}{2}$

$= \frac{(6-3)(6-3+1)}{2}$

$= \frac{3 \times 4}{2} = 6$ lines

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- No of lines in each series:

Lyman's series: $n_1 = n_2 - 1 = 0$ (Lyman's series)

Balmer's series: $n_1 = n_2 - 2 = 0$ (Balman's series)

Paschen's series: $n_1 = n_2 - 3 = 6 - 3 = 3$.

Brackett's series: $n_1 = n_2 - 4 = 6 - 4 = 2$

Pfund series: $n_1 = n_2 - 5 = 6 - 5 = 1$

Humphery series: $n_1 = n_2 - 6 = 6 - 6 = 0$.

- No. of lines in Ultra regions: 0
- No of lines in visible regions: 0
- No of lines in Infrared regions: $3+2+1 = 6$.

62 An e^- in hydrogen jumps from higher orbit (n_2) to ground state taking multiple transitions If the total no. of lines formed in the spectrum are 21, then find out the value of n_2 ?

$$T.S.L. = \frac{(n_2 - n_1)(n_2 - n_1 + 1)}{2}$$

$$21 = \frac{(n_2 - 1)(n_2 - 1 + 1)}{2}$$

$$42 = n_2 (n_2 - 1)$$

$$(n_2)^2 - n_2 - 42 = 0.$$

$$(n_2)^2 - 7n_2 + 6n_2 - 42 = 0.$$

$$n_2(n_2 - 7) + 6(n_2 - 7) = 0.$$

$n_2 = -6$ and $n_2 = 7$

$n_2 = -6$ (not possible)

$\because$ No. of lines can't be negative.

$\therefore n_2 = 7.$

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Q.63. Determine the transitions made by the e^- from n_2 energy level to n_1 energy level for the following lines.

i) 5th line in the Lyman's series.

ii) γ line in the Brackett's series

iii) α line in the Humphery's series.

iv) β line in the Paschen's series

v) 4 th line in the Balmer's series.

vi) last line in the Pfund series.

i) $n_2 = n_1 + 5$ ($\because n_1 = 1$ for Lyman series)
$= 1 + 5 = 6$. (6th to 1st)

ii) $n_2 = n_1 + 3$. ($\because n_1 = 3$ for Brackett series)
$= 4 + 3 = 7$ (7th to 4th)

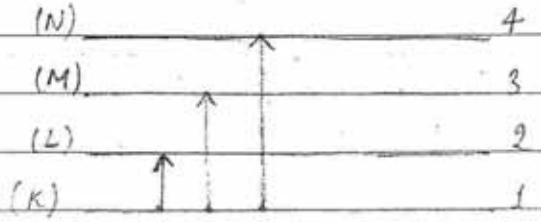
iii) $n_2 = n_1 + 1$ ($\because n_1 = 1$ for Humphery series)
$= 6 + 1 = 7$ (7th to 6th).

iv) $n_2 = n_1 + 2$ ($\because n_1 = 2$ for Paschen series)
$= 3 + 2 = 5$ (5th to 3rd)

v) $n_2 = n_1 + 4$ ($\because n_1 = 4$ for Balmer series)
$= 2 + 4 = 6$ (6th to 2nd).

vi) $n_2 = n_1 + 5$ ($\because$ last line in Pfund series)
$= 5 + \infty = \infty$ (∞ to 5th)

* In the absorption spectrum of hydrogen only dark lines of lyman's series are obtained. This verifies / shows that an e^- does not make multiple transitions during excitations.



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Q-64 If an atom contains 5 permitted Orbit, then find out the total no. of lines formed in emmission and absorption spectrum

$n_2 = 5$ and $n_1 = 1$.

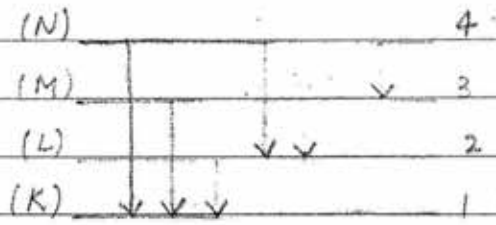
$$T.S.L = \frac{(n_2 - n_1)(n_2 - n_1 + 1)}{2}$$

$$= \frac{(5-1)(5-1+1)}{2}$$

$$= \frac{4 \times 5}{2} = 10 \text{ lines.}$$

- Total no. of lines in emmission spectrum = 10 lines.
- Total no. of lines in absorption spectrum = = Lyman's series = $(n_2 - 1)$ = $5 - 1 = 4$ lines.

Q.65 In the emmission spectrum of hydrogen, following lines are obtained.



How many of these lines will be obtained in absorption spectrum?

No of lines obtained in absorption spectrum = No of lines of Lyman's series.

Hence,

Total no of lines obtained = 3

where, $n_1 = 1$.

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Rydberg formula:

MODULE 1 (Pg-173)

i) $\bar{\upsilon} = \frac{1}{\lambda} = RZ^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$

ii) $\upsilon = RCZ^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$

iii) $E = -\frac{RChZ^2}{n^2}$

iv) $\Delta E = RChZ^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$

where $R = 1.1 \times 10^5\ cm^{-1}$ or $1.1 \times 10^7\ m^{-1}$

$C = 3 \times 10^8\ m/s.$

$h = 6.626 \times 10^{-34}\ J$

Z = atomic no., and.

n_1 = no. of lower energy level

n_2 = no. of higher energy level.

v) $\frac{1}{R} = 912\ \text{Å}.$

Q.66 Calculate the wavelength for the 1st line of the lyman's series in Hydrogen atom

$n_1 = 1$ and $n_2 = n_1 + 1 = 1 + 1 = 2$

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

$= R(1)^2\left[\frac{1}{1^2} - \frac{1}{2^2}\right]$

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$$= R\left[1 - \frac{1}{4}\right]$$

$$\frac{1}{\lambda} = R\left[\frac{3}{4}\right] = \frac{3R}{4}$$

$$\therefore \lambda = \frac{4}{3R} \text{ units.}$$

67 Calculate the frequency of radiation of the Balmer line of Hydrogen atom

$n_1 = 2$ and $n_2 = n_1 + 1 = 2 + 1 = 3.$

$$\therefore \upsilon = RCZ^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

$$= RC(1)^2\left[\frac{1}{(2)^2} - \frac{1}{(3^2)}\right]$$

$$= RC\left[\frac{1}{4} - \frac{1}{9}\right]$$

$$\upsilon = RC\left[\frac{5}{36}\right] = \frac{5RC}{36} \text{ units}$$

68 An e^- shows transition according to first line of Paschen in Hydrogen atom. then, calculate the energy of the electronic transitions.

$n_1 = 3$ and $n_2 = n_1 + 1 = 3 + 1 = 4$

$$\because \Delta E = RCh\,Z^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

$$= RCh(1)^2\left[\frac{1}{3^2} - \frac{1}{4^2}\right]$$

$$= RCh\left[\frac{1}{9} - \frac{1}{16}\right]$$

$$= RCh\left[\frac{7}{144}\right] = \frac{7RCh}{144} \text{ units}$$

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Q.69. An e^- shows transition according to first brackett line in Hydrogen atom. then calculates its wave no.

$n_1 = 4$ and $n_2 = n_1 + 1 = 4 + 1 = 5$

$$\because \bar{\upsilon} = RZ^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

$$= R(1)^2\left[\frac{1}{4^2} - \frac{1}{5^2}\right]$$

$$= R\left[\frac{1}{16} - \frac{1}{25}\right]$$

$$= R\left[\frac{9}{400}\right] = \frac{9R}{400} \text{ units}$$

Q.70 An e^- shows transition according to the limiting line of the Pfund series in Hydrogen atom. then, calculate the wavelength of this in Å.

$n_1 = 5$ and $n_2 = \infty$

$$\because \bar{\upsilon} = \frac{1}{\lambda} = RZ^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

$$= R(1)^2\left[\frac{1}{5^2} - \frac{1}{\infty^2}\right]$$

$$= R\left[\frac{1}{25} - 0\right] \qquad \left[\because R = \frac{1}{912\,\text{Å}}\right]$$

$$\frac{1}{\lambda} = \frac{R}{25}$$

$$\therefore \lambda = \frac{25}{R} = \left[\frac{25 \times 912}{1}\right] \text{Å}$$

$$= 22800\ \text{Å}$$

Q.71 If the shortest wavelength of Lyman's series of Li^{+2} is x. then find out the

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maximum wavelength of Balman's series of ~~hydrogen atom~~. He^+

For Lyman's series.

$n_1 = 1$ and $n_2 = \infty$ (maximum wavelength)

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

$$= R(3)^2\left[\frac{1}{1^2} - \frac{1}{\infty^2}\right]$$

$$= 9R(1-0) = 9R$$

$$\therefore \lambda = \frac{1}{9R} = x.$$

$$\therefore R = \frac{1}{9x} \quad - (1)$$

For Balman's series

$n_1 = 2$ and $n_2 = 2+1 = 3$. (min wavelength)

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

$$= R(2)^2\left[\frac{1}{2^2} - \frac{1}{3^2}\right]$$

$$= 4R\left[\frac{1}{4} - \frac{1}{9}\right]$$

$$\frac{1}{\lambda} = 4R\left[\frac{5}{36}\right] = \frac{5R}{9}$$

$$\therefore \lambda = \frac{9}{5R}$$

Putting R from eq (1) we get.

$$\lambda = \frac{9}{5 \times \frac{1}{9x}} = \frac{9 \times 9x}{5}$$

$$= \frac{81x}{5} \text{ units.}$$

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Q.72 If the maximum frequency of Balmer's series of H-atom is x, then find out the maximum frequency of Paschen series of He^+.

For Balmer's series.

$$\nu = RCZ^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] \quad (n_1 = 2, n_2 = \infty)$$

$$= RC(1)^2\left[\frac{1}{2^2} - \frac{1}{\infty^2}\right]$$

$$= \frac{RC}{4}\left[\frac{1}{4} - 0\right] = \frac{RC}{4} = x.$$

$$\therefore RC = 4x \quad -(I)$$

For Paschen's series.

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

$n_1 = 3$ and $n_2 = \infty$

$$= RC(2)^2\left[\frac{1}{3^2} - \frac{1}{\infty^2}\right]$$

$$= 4RC\left[\frac{1}{9} - 0\right] = \frac{4RC}{9} \quad -(II)$$

By putting eq (I) in (II), we get

$$= \frac{4 \times 4x}{9} = \frac{16x}{9}$$

Limitations of Bohr Model:

MODULE 1 (Pg 173)

- Fine spectral lines indicates the presence of sub-energy shells in an orbit and their energy will be nearly same.

It includes five drawbacks.

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Sommerfield Extension of the Bohr model:

MODULE 1 (Pg - 174)

Sommerfield theory was related to the explanation of fine spectral line i.e the existence of the sub orbits. Sommerfield introduced the concept of sub-energy level in an orbit according to sommerfield. in a main energy shell the energies of the sub shells are slightly different from another.

Hence, on jumping of an electron from one energy level to another energy level will involve slightly different amount of energy as it will depend on the subshell

∴ Number of fine lines in a spectral line resulting from a transition between n_2 and n_1 energy level = $n_2 \times n_1$

de Broglie concept (Dual nature of matter)

MODULE 1 (Pg - 174)

$$\lambda = \frac{h}{mv} = \frac{h}{p}$$

where, λ = wavelength

h = plank's constant

m = mass

v = velocity

p = momentum

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~~Important~~

$$\lambda = \frac{h}{mc}$$

when particle ~~pt~~ travels with the velocity of light

Important point concerned with de-broglie concept:

MODULE 1 (Pg - 175)

$$\lambda = \frac{h}{\sqrt{2m\,K.E}}$$

when a charged particle carrying Q coulomb is accelerated by applying potential difference V, then,

MODULE 1 (Pg - 175)

$$\lambda = \frac{h}{\sqrt{2mQV}}$$

- The wave nature of electron was varified experimentally by Davisson and Germer in a 'Crystal diffraction of cathode rays' in 1927. (electron microscope are used on this basis)

- An object undergoes diffraction has wave nature and it is associated with sub micro-scopic particles like electron, proton etc., when it is in motion.

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• Number of waves made by an electron when it revolves in a particular Bohr orbit = Orbit number (n).

$\therefore \quad n = \frac{2\pi r}{\lambda}$

Q.73. Calculate the wave length number of an electron moving with kinetic energy of 4.55×10^{-25} J

$\because \quad \lambda = \frac{h}{\sqrt{2m\,KE.}}$

$= \frac{6.626 \times 10^{-34}}{\sqrt{2 \times 9.1 \times 10^{-31} \times 4.55 \times 10^{-25}}}$

$= \frac{6.626 \times 10^{-34}}{\sqrt{9.1 \times 9.1 \times 10^{-56}}}$

$= \frac{6.626 \times 10^{-34}}{9.1 \times 10^{-28}}$

$= 0.72 \times 10^{-6}$ m.

Q.74 Two particles A and B are in motion. the wavelength of particle A is 5×10^{-5} m. If the mass of the particle B is 50% of A and its velocity is 25% of A, then calculate the wavelength of particle B. ?

$\lambda_A = \frac{h}{m_A v_A} = 5 \times 10^{-5}$ m

$\lambda_B = \frac{h}{\frac{m_A}{2} \times \frac{v_A}{4}}$

$\therefore \quad \frac{\lambda_A}{\lambda_B} = \frac{h / m_A v_A}{8h / m_A v_A} = \frac{h}{8h}$

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$$\therefore \frac{\lambda_A}{\lambda_B} = \frac{h}{8h} = \frac{1}{8}$$

$$\therefore \lambda_B = \lambda_A \times 8$$

$$= 5 \times 10^{-5} \times 8$$

$$= 40 \times 10^{-5} \text{ m}$$

Q.75. Calculate the de broglie wavelength of a particle of mass 10 grams with a velocity of 200 m/s.

$$\lambda = \frac{h}{mv}, \quad m = 10 \text{ gms} = 10 \times 10^{-3} \text{ kg}$$

$$= \frac{6.626 \times 10^{-34}}{10 \times 10^{-3} \times 200}$$

$$= \frac{6.626 \times 10^{-34}}{2 \times 10^{-3} \times 10^{3}}$$

$$= 3.313 \times 10^{-34} \text{ m}$$

Q.76 Calculate the de broglie wavelength of an e^- moving in 2nd orbit of H atom.

$$\lambda = \frac{2\pi r_n}{n}$$

$$= \frac{2 \times \pi \times 0.529 \times 4 \text{ Å}}{2}$$

$$= 2.11\pi \text{ Å}$$

Heisenberg uncertainity Principle:

MODULE 1 (Pg - 176)

According to Heisenberg uncertainity principle "It is impossible to measure simultaneously the exact position and the exact momentum of a body as small as an electron"

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$\Delta x \cdot \Delta p \geq \frac{h}{4\pi}$ OR $\Delta x \cdot \Delta v \geq \frac{h}{4\pi m}$

where $\frac{h}{4\pi} = 0.527 \times 10^{-34}$ J sec.

λ (Incident light) α Size of an object

* Heisenberg ruled out the concept of fixed no. of orbit given by Bohr. According to heisenberg, we can only specify a region in space where the probability of finding electron is maximum

77 If the uncertainity in position of a particle is 0. then find out uncertainity in its momentum

$\because \Delta x \cdot \Delta p = \frac{h}{4\pi}$

$\therefore \Delta p = \frac{h}{4\pi \Delta x} = \frac{h}{4\pi \times 0}$

$= \frac{h}{0} = \infty$

$\therefore \Delta p = \infty$

Q.78 The uncertainity in position of an e^- is 1 Å then calculate the uncertainity in its velocity (m/s)

$\Delta x \cdot \Delta p = \frac{h}{4\pi}$

$\therefore \Delta v = \frac{h}{4\pi m \Delta x}$

$= \frac{0.527 \times 10^{-34}}{9.1 \times 10^{-31} \times 10^{-10}}$ $\left[\because 1\text{Å} = 10^{-10}\text{ m}\right]$

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$$\frac{0.527 \times 10^{-34}}{9.1 \times 10^{-41}} = 5.7 \times 10^{5} \text{ m/s.}$$

Q 79. An e^- is moving with a velocity of 300 m/s. if the uncertainity in velocity is 0.01% calculate the uncertainity in position.

$v = 300$ m/s.

$$\therefore \Delta v = \frac{300}{100} \times 0.01 = 0.03 \text{ m/s}$$

$$\therefore \Delta x = \frac{h}{4\pi m \Delta v}$$

$$= \frac{0.527 \times 10^{-34}}{9.1 \times 10^{-31} \times 0.03.}$$

$$= \frac{0.527 \times 10^{-3}}{9.1 \times 3} = \frac{5.7}{3} \times 10^{-3.}$$

$$= 1.9 \times 10^{-3} \text{ m}$$

Q.80. The uncertainity in position of an electron is double the uncertainity in position of He atom. If the uncertainity in momentum of electron is x then, the uncertainity in momentum of He atom.

$$\Delta p(H) = \frac{h}{4\pi \Delta x(He)}$$

$$\Delta p(e^-) = \frac{h}{4\pi \Delta x(e^-)}$$

$$\therefore \frac{\Delta p(H)}{\Delta p(e^-)} = \frac{\Delta x(e^-)}{\Delta x(H)} = \frac{2\,\Delta x(H)}{\Delta x(H)}$$

$$\therefore \frac{\Delta p(H)}{x} = 2.$$

$$\therefore \Delta p(H) = 2x.$$

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Wave mechanical ~~Orde~~ Model:

- It was based on the dual nature of electron i.e. the particle as well as the wave nature.

- This model was given by Schrodinger and was represented by the following equation

$$\frac{\partial^2\psi}{\partial x^2} + \frac{\partial^2\psi}{\partial y^2} + \frac{\partial^2\psi}{\partial z^2} + \frac{8\pi^2 m}{h^2}[E-V]\psi = 0$$

Or

$$\nabla^2\psi + \frac{8\pi^2 m}{h^2}(E-V) = 0.$$

Where ∇^2 = Laplacian operator
ψ (Psi) = wave function (Amplitude of wave)
x, y, z = space-coordinate
E = Total energy
V = Potential energy
∂ = dell

- The solution of the Schrodinger equation gives a set of three quantum number n, l and m.

Orbitals:

- It is a region in space where the probability of finding electron is maximum (greater than 95%) or it is an electron cloud which has a definite size, shape and orientation.

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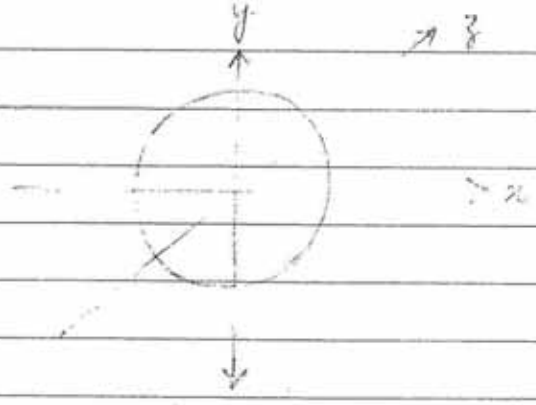
- An orbital is represented as follows: $[\psi_{nlm}]$

Shapes of Atomic orbitals:

i) 's' - orbital.

MODULE 1 (Pg - 175)

- The shape of the 's' orbitals is represented as a symmetrical sphere
- The 's' orbitals are symmetrical along all the axis.
- The 's' orbitals are non-directional in nature.



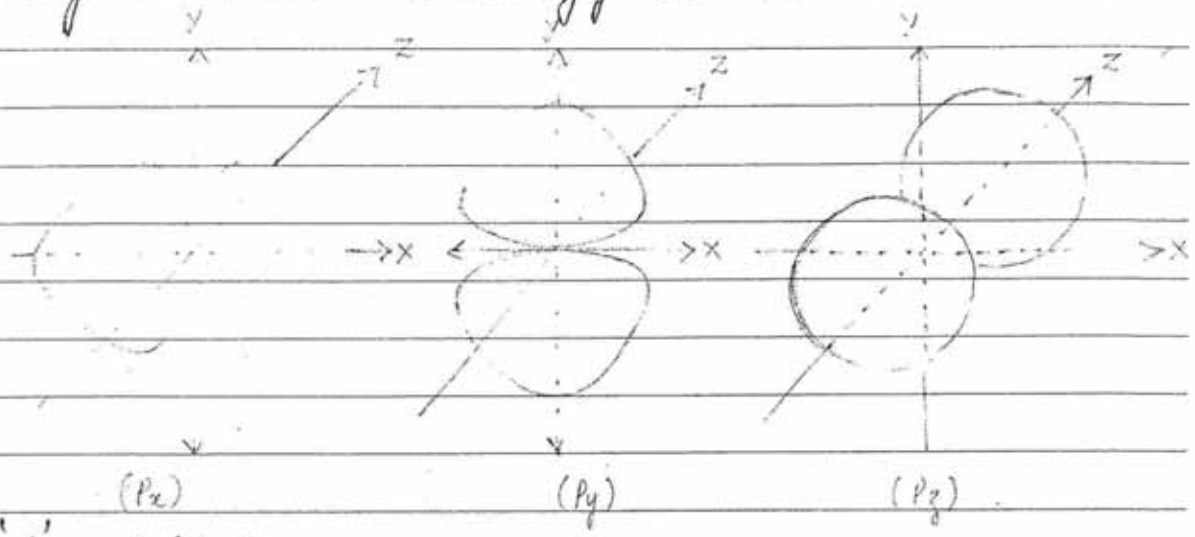
ii) 'p' orbital

MODULE 1 (Pg - 175)

- The shape of a p-orbital is ~~a dumb~~ represented as a dumb-bell (2 lobes)
- There are three types of p-orbital, p_x, p_y and p_z which are directed along each axis.
- The angle between p-orbitals is 90°
- p-orbitals are directional in nature.

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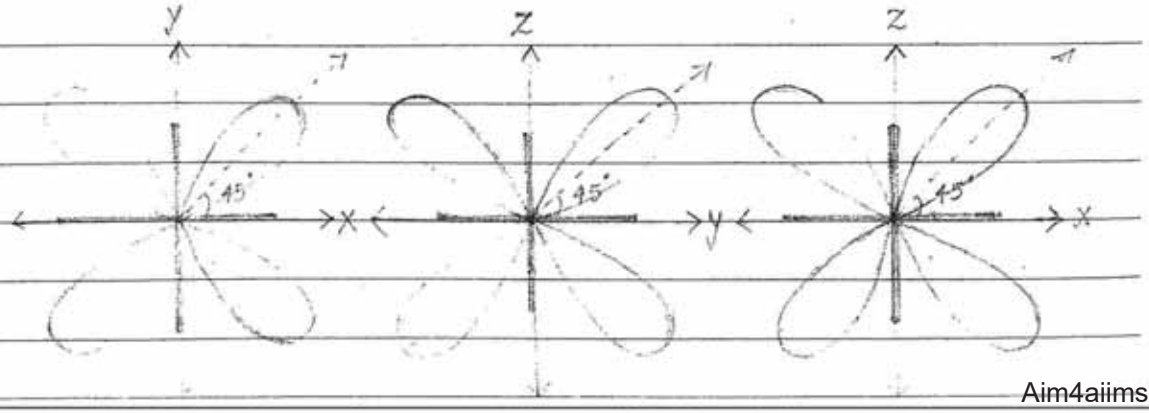
- In the absence of an external field, the energy of the p-orbitals in the same orbit is equal and hence they are called as '3 fold degenerate orbitals degenerate: Same energy levels.



iii) 'd'- orbital

- The shape of a 'd'- orbital is represented as a double dumb-bell (4 lobes)
- There are five types of 'd-orbitals' Between the axis: (d_{xy}, d_{yz}, d_{xz}) and along the axis: ($d_{x^2-y^2}$, d_{z^2})
- These d-orbitals are categorised into two parts:

a. Between the axis.

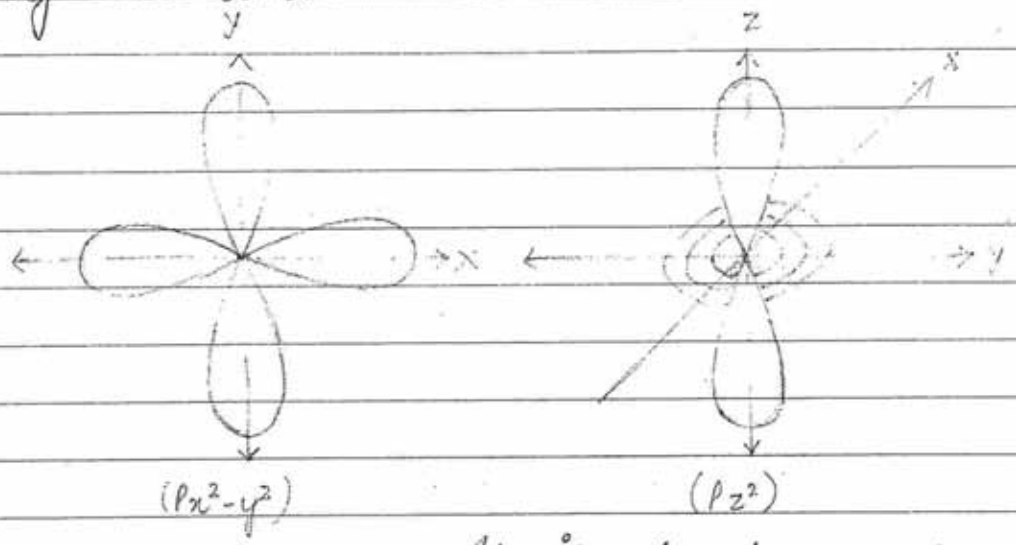


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these d-orbitals are symmetrical and they make an angle of 45° with each axis.

b. Along the axis.



(Px^2-y^2) (Pz^2)

It is also known as baby soother or doughnut, because it consists of two lobes along with an e- ring in x-y plane

- d-orbitals are directional in nature,
- In absence of an external field, the energy of the d-orbital in the same orbit are equal. hence, they are called as 5-fold degenerate orbitals.

iv) f-orbital

- The shape of f-orbital is highly complex in nature.
- There are seven f-orbital

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Quantum Number :

MODULE 1 (Pg 176)

- Quantum number are the address of an electron.
- Quantum number determines the position and the energy of an electron in an atom.
- An orbital is designated by three quantum no. n, l and m
- An electron is designated by four quantum number n, l, m and s

- n = Principal Quantum numbers.
  Shell (K, L, M, N .....)
- l = Azimuthal Quantum numbers.
  Sub-Shell (s, p, d, f .....)
- m = Magnetic Quantum number.
  Orbitals (s, p_x, p_y, p_z, d_{xy}, d_{yz}, d_{xz}, $d_{x^2-y^2}$, d_{z^2} .....)
- s or m_s = Spin Quantum number.

Principal Quantum number (n)

MODULE 1 (Pg - 176)

- It was given by Bohr.
- It represents the name, size and energy of a shell.
- The possible value of n are positive integers. (1, 2, 3, 4 .....) and nomenclatured as (K, L, M ....)
- It also represents the distance of e^- from the nucleus.

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- As the value of n increases, the energy and the distance of an electron from the nucleus also increases.
- Total number of electron in a shell are given by $2n^2$.

2. Azimuthal / Angular / Subsidary / Secondary Quantum number (l)

MODULE 1 (Pg - 177)

- It was given by Summerfield.
- It signifies a subshell and determines the shape of an orbital.
- The possible value of l are from 0 to $(n-1)$

When $l = 0$ (s - subshell),
when $l = 1$ (p - subshell),
when $l = 2$ (d - subshell),
when $l = 3$ (f - subshell),
when $l = 4$ (g - subshell) and so on

- Relation between n and l :

When $n = 1$, $l = 0$ (s - subshell)
when $n = 2$, $l = 0, 1$ (s, p - subshell)
when $n = 3$, $l = 0, 1, 2$ (s, p, d - subshell)
when $n = 4$, $l = 0, 1, 2, 3$. (s, p, d, f - subshell)
It shows that forth shell contains four subshell i.e s, p, d and f.

- Hence, nth shell = n subshell.

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- It represents the orbital angular momentum of an electron.
  Orbital angular momentum = $\frac{\sqrt{l(l+1)}\,h}{2\pi}$

  $s = 0 \quad (\because l = 0)$

  $p = \frac{h}{\sqrt{2}\pi} \quad (\because l = 1)$

  $d = \frac{\sqrt{6}\,h}{2\pi} \quad (\because l = 2)$

  $f = \frac{\sqrt{12}\,h}{2\pi} = \frac{\sqrt{3}\,h}{\pi} \quad (\because l = 3).$

- Number of electron in a subshell = $2(2l+1)$

  $s = 2 \quad (\because l = 0)$

  $p = 6 \quad (\because l = 1)$

  $d = 10 \quad (\because l = 2)$

  $f = 14 \quad (\because l = 3)$

- Number of orbitals in a subshell = $(2l+1)$

  $s = 1 \quad (\because l = 0)$

  $p = 3 \quad (\because l = 1)$

  $d = 5 \quad (\because l = 2)$

  $f = 7 \quad (\because l = 3)$

- Maximum number of electron in an orbital are two.

- It represents/signifies the energy of a subshell in a particular orbit.

  When l = 0 1 2 3

  $s < p < d < f$

  Its applicable only for multi electron species

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- In single electron species like hydrogen, $s = p = d = f$. because effective nuclear charge will be same

Q. 81. Calculate the orbital angular momentum of an electron in p-subshell.

Orbital angular momentum $= \frac{\sqrt{l(l+1)}\,h}{2\pi}$

for p-subshell, $l = 1$

$$= \frac{\sqrt{1(1+1)}\,h}{2\pi}$$

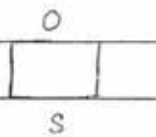
$$= \frac{\sqrt{2}\,h}{2\pi} = \frac{\sqrt{2}\,h}{2\pi} \times \frac{h}{\sqrt{2}\pi}$$

3. Magnetic Quantum number (m)

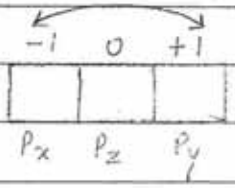
MODULE 1 (Pg - 178)

- It was given by Linde.
- It represents the orientation of an orbital
- The possible values of m is from $(-l)$ to $(+l)$ including zero.

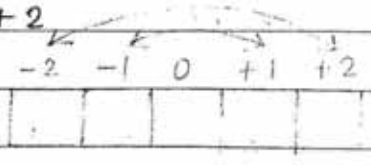
when $l = 0$, $m = 0$.
(s-subshell → 1 orbital)



when $l = 1$, $m = -1, 0, +1$
(p-subshell → 3 orbitals)



when $l = 2$, $m = -2, -1, 0, +1, +2$
(d subshell → 5 orbitals)



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- Number of values of 'm' in a subshell : $(2l+1)$

4 Spin Quantum number (s)

MODULE 1 (Pg - 179)

- It was given by Uhlenbeck and Goudschmidt
- It represent the direction of rotation of an electron about its own axis.
- The possible value of s are $\left(+\frac{1}{2}\right)$ and $\left(-\frac{1}{2}\right)$ that represent clockwise and anti-clockwise rotation. ↑↓
- It represent the spin angular momentum of an electron.

when $s = +\frac{1}{2}$

spin angular momentum $= \sqrt{s(s+1)}\,\frac{h}{2\pi}$

when $s = +\frac{1}{2}$ $= \sqrt{\frac{1}{2}\left(\frac{1}{2}+1\right)}\,\frac{h}{2\pi}$

$= \sqrt{\frac{1}{2}\left(\frac{3}{2}\right)}\,\frac{h}{2\pi}$

$= \sqrt{\frac{3}{4}}\,\frac{h}{2\pi} = \frac{\sqrt{3}h}{2\times 2\pi} = \frac{\sqrt{3}h}{4\pi}$

when $s = -\frac{1}{2}$

spin angular momentum is not possible.

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- Total number of spin $= \pm\frac{1}{2} \times$ (no of unpaired e^-)

Formula's Association :

i) Total no of electron in a shell : $2n^2$
ii) No. of Orbitals in a shell : n^2
iii) Total no. of electron in a subshell : $2(2l+1)$
iv) No. of orbitals in a subshell : $(2l+1)$
(Total no of values of 'm' in a subshell)
v) Maximum no. of electron in a orbital : 2
vi) Orbital angular momentum (μ_l) : $\sqrt{l(l+1)}\,\frac{h}{2\pi}$
vii) spin angular momentum (μ_s) : $\sqrt{s(s+1)}\,\frac{h}{2\pi}$
viii) Total angular momentum (μ) : $(\mu_l) + (\mu_s)$
ix) nth shell : n subshell

Q.82 Find out the values of n, l and m for the following orbitals.

i) $4d_{z^2}$:
$n = 4 \quad l = 2 \quad m = 0$

ii) 5s :
$n = 5 \quad l = 0 \quad m = 0$

iii) $3p_z$:
$n = 3 \quad l = 1 \quad m = 0$

iv) $2p_x$:
$n = 2 \quad l = 1 \quad m = \pm 1$

Q.83 Identify the orbitals for the following value of n, l and m.

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i) $n = 4 \quad l = 1 \quad m = -1$

Orbital : $4p_x$ or $4p_y$

ii) $n = 3 \quad l = 2 \quad m = 0$

Orbital : $3d_{z^2}$

iii) $n = 2 \quad l = 0 \quad m = 0$

Orbital : 2s

iv) $n = 3 \quad l = 2 \quad m = 2$

Orbital : $3d_{x^2-y^2}$ or $3d_{xy}$

v) ψ_{420}

here $n = 4$, $l = 2$, $m = 0$

$\therefore$ orbital : $4d_{z^2}$

84 which of the following set of quantum number is possible :

i) $n = 3 \quad l = 2 \quad m = -3 \quad s = +1/2$

ii) $n = 3 \quad l = 3 \quad m = 2 \quad s = -1/2$

iii) $n = 3 \quad l = 1 \quad m = 0 \quad s = +1/2$

iv) $n = 3 \quad l = 1 \quad m = 0 \quad s = +1$

Ans : (iii) $n = 3$, $l = 1$, $m = 0$ and $s = +1/2$

It is possible because $m \neq -3$ and $l \neq 3$ and $s \neq +1$.

$\because$ $l < n$ and $m = (-l)$ to l and $s = +1/2$

Q.85 which of the following orbitals are not possible.

i) 1p [×]

Not possible because if $n = 1$, then $l = 0$

$\because l = 0$, $\therefore$ only 's-orbital' is possible

ii) 2s [✓]

Possible because if $n = 2$, then $l = 0, 1$

when $l = 0$ i.e ~~'s-orbital'~~ 's-subshell'

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iii) 3f [×]

Not possible because if $n = 3$ then $l = 0, 1, 2$ (s, p and d subshell).

Hence, 'f-subshell' is not possible

iv) 4f [✓]

Possible because if $n = 4$ then $l = 0, 1, 2, 3$

hence $l = 3$, 'f-subshell'.

v) 2d [×]

Not possible because if $n = 2$, then $l = 0, 1$ (s, p subshell)

Q.86 Calculate the total no of electrons which can be present in a 4(th) d subshell

Total no. of e^- = $2(2l+1)$

$= 2(2\times2+1)$

$= 2(4+1) = 2\times5 = 10$

Q.87 Find out the ~~value of~~ total no. of 'm' for an e^- present in the 5th shell.

$\therefore n = 5$

$l = 0, 1, 2, 3, 4$

when $l = 0$, $m = 0$

when $l = 1$, $m = -1, 0, +1$

when $l = 2$, $m = -2, -1, 0, +1, +2$

when $l = 3$, $m = -3, -2, -1, 0, +1, +2, +3$

when $l = 4$, $m = -4, -3, -2, -1, 0, +1, +2, +3, +4$

hence total value of $m = 1+3+5+7+9 =$ 25 electrons.

OR

no. of e^- in a shell = $2n^2 = 2\times5^2$

$= 2\times25$ electrons.

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88. Find out the no. of e^- in M shell having the value of magnetic quantum no as -1

when $m = -1$

~~then $l = 1$~~

M shell, $\therefore n = 3$.

$\therefore l = 0, 1, 2$

when $l = 0$, $m = 0$

when $l = 1$, $m = (-1)\ 0, +1$

when $l = 2$, $m = -2, (-1), 0, +1, +2$

~~hence, total no of e^- i.e value of 'm' =~~

$1 + 3 + 5 = 9$

Total no. of e^- having magnetic quantum as $-1 = 2 \times 2 = 4$ electrons.

89. In orbitals of $n = 3$ and $l = 2$, find out the total no of electrons having the value of spin quantum no. as $-1/2$.

When $n = 3$, $l = 2$ ~~$\neq$ $\neq$~~

Total no. of orbital : 3d

Total no. of e^- in 3d = 10

$\therefore$ no of e^- having $s = -1/2 = 5$

90. Find out the maximum no. of orbitals in the 4th shell having an electron with a value of spin quantum no $+1/2$

Total no. of ~~orbital~~ electrons in a shell $= 2n^2$

$\because n = 4$

$\therefore$ no of ~~orbitals~~ electrons $= 2 \times 4^2 = 16 \times 2 = 32\ e^-$

$\because$ spin quantum no $= +1/2$

$\therefore$ no of electrons $= \frac{32\ e^-}{2} = 16\ e^-$

Hence, 16 electron can have $+1/2$ spin quantum no.

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Q.91 For $n+l = 5$, find out the no. of subshell, total no of orbital and total no of electro

i) Total no. of subshell;

Possible outcomes = $5+0 = 5s$

$= 4+1 = 4p$

$= 3+2 = 3d.$

3

Hence, no. of subshell = 3.

ii) Total no. of orbitals :

In 5s = 1

In 4p = 3

In 3d = 5

9

Hence, no of orbitals = 9.

iii) Total no of e^- :

In 5s = 2

In 4p = 6

In 3d = 10

18.

Hence, no of electrons = 18

Q.92 If The value of $n+l$ is not more than 3, then find out the total no. of subshell, total no. of orbitals and total no of electrons.

Possible outcomes :

when $n+l = 3$.

$3+0 = 3$, (3s ~~orb. subshell~~ orbital)

$2+1 = 3$. , (2p ~~subshell~~ orbital)

when $n+l = 2$

$2+0 = 2$, (2s ~~subshell~~ orbital)

when $n+l = 1$

$1+0 = 1$, (1s ~~subshell~~ orbital)

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Hence,
total no. of subshell = 4
total no of orbital = 1 + 3 + 1 + 1
= 6
total no of electrons = 12

Nodes:

- It is a region in space where the probability of finding e^- is zero.
- It is of two types:

i) Radial node / spherical node / Nodal surface:
value of radial node = $(n-l-1)$

ii) Angular node / Non-spherical node / nodal plane:
value of angular node = l (azimuthal q.n.)

∴ Value of total no. of nodes = (value of radial) + (angular node)
= $(n-l-1) + l$
= ~~$(n-l)$~~ $(n-1)$

Q. 93 Calculate the number of radial, angular and total no. of nodes for the following:

i) 4d : $n = 4$ and $l = 2$
Radial node = $(n-l-1) = 4-2-1 = 1$
Angular node = $l = 2$
Total node = $(n-1) = 4-1 = 3$

ii) 5s : $n = 5$ and $l = 0$
Radial node = $(n-l-1) = 5-0-1 = 5-1 = 4$

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Angular node = (l) = 0
Total node = $(n-1)$ = 5 - 1 = 4

iii) ~~3d~~ 3p : $n = 3$ and $l = 1$
Radial node = $(n-l-1) = 3-1-1 = 1$
Angular node = $(l) = 1$
Total node = $(n-1) = 3-1 = 2$

• Number of nodal plane (l) :

i) s - orbital : $s = 0$
ii) p - orbital : p_x = yz plane
p_y = xz plane
p_z = xy plane
iii) d - orbital : d_{xy} = xz plane and yz plane
d_{yz} = xy plane and xz plane
d_{xz} = xy plane and yz plane
$d_{x^2-y^2}$ = 0
d_{z^2} = 0

Rules of filling orbitals :

i) Aufbau Principle :

ii) $(n+l)$ rule :

iii) Pauli's exclusion principle :
he stated that no 2 electrons can have the same value of all four quantum numbers
eg : In 2s orbital.
[1↓] [1↑] (not possible)

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In 2s orbital: $n_1 = 2$ and $n_2 = 2$

(eg:) $l_1 = 0$ and $l_2 = 0$

$m_1 = 0$ and $m_2 = 0$

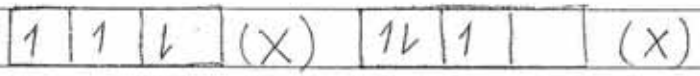
$s_1 = +\frac{1}{2}$ and $s_2 = -\frac{1}{2}$

iv) Hund's rule of Maximum multiplicity:
Electrons are distributed in a way to give maximum no. of unpaired electrons with parallel spin.

Reasons:

- due to symmetrical distribution of electrons the inner electronic repulsion becomes minimum.

- As the number of electrons with the same spin increases, the number of exchanges will increase and thus the exchange energy will increase which ultimately leads to increase in the stability of an electron.



(opposite spin) (Pairing)

Result:
Half filled and full filled orbitals are more stable.

eg: less stable $\{ p^3 < p^6 \}$ more stable
$\{ d^5 < d^{10} \}$
$\{ f^7 < f^{14} \}$

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Electronic configuration of ions :

- Cations are positively charged and anions are negatively charged.
- Electronic configuration of anions are same as that of Aufbau Principal but for the electronic configuration of cations the electrons from the outermost shell (n) with higher energy loses earlier.

eg: $_{30}Zn$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ \boxed{4s^2}$

$\therefore {}_{30}Zn^{+2}$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}$

$_{24}Cr$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ \boxed{4s^1\ 3d^5}$

$\therefore {}_{24}Cr^{+2}$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^4$

Some general points :

i) The elements having unpaired electrons are called paramagnetic and the elements with no unpaired electrons are called diamagnetic

ii) Magnetic movement / Paramagnetic movement $= \sqrt{n(n+2)}\ BM$

where, n = no. of unpaired electrons.

BM = bohr magneton (constant)

$BM = 9.27 \times 10^{-24}$ Joule/sec.

$\therefore \sqrt{n(n+2)}\ BM = 9.27 \times 10^{-24}\ J/s.$

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iii) Total spin = $\pm \frac{1}{2} \times$ no. of unpaired e^-.

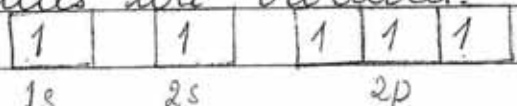
iv) Kernel : It is the remaining part of an atom after the removal of outermost shell.

v) Core ~~energy~~ charge : Remaining portion after Kernel.
i.e. = Atomic no − no. of e^- in kernel.

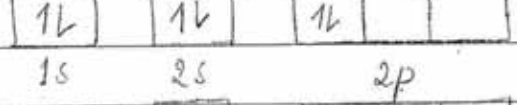
94. For which element of the periodic table., Pauli's exclusion principal is not applicable ?
It is not applicable for hydrogen (H_2) because it doesn't contains (minimum number) ~~of~~ two electrons in its shell.

95 Determine the first element of periodic table for which hund's rule is applicable.
carbon, because from this element, the pairing of the electron starts.

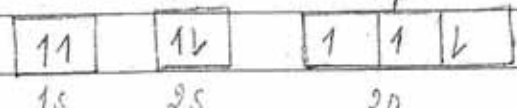
96. In the following configuration, which rules are violated.

i) 

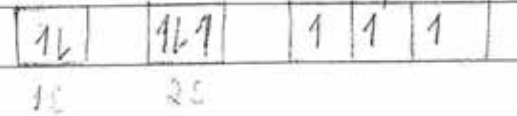
Aufbau principal

ii) 

Hunds rule

iii) 

Hunds and Pauli's rule

iv) 

Pauli's Principal

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Q.97 If principal quantum no. more than 4 is not possible, then find out the total no. of possible element in nature?

$n \leq 4$

Total no. of electron in a shell = $2n^2$

when $n = 1$; no of $e^- = 2(1)^2 = 2$

when $n = 2$; no of $e^- = 2(2)^2 = 8$

when $n = 3$; no of $e^- = 2(3)^2 = 18$

when $n = 4$; no of $e^- = 2(4)^2 = 32$

$\therefore 2 + 8 + 18 + 32 = 60$

Hence, total no. of possible elements = 60.

Q.98 If principle quantum no. more than 4 is not possible in the electronic configuration of an element then find out the maximum no. of electrons.

$n \leq 4$

configuration : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p$

Hence, maximum no of electrons present in the element = ~~36~~ $2+2+6+2+6+2+10+6$

= 36 electrons.

Q.99 In an atom, how many electrons can have these value of quantum no.

$n = 4$; $l = 2$; $m = 1$ and $s = +1/2$

Only one electron can have the above values because according to Pauli's exclusion rule, 'No two electrons can have the same value of all the four quantum number'.

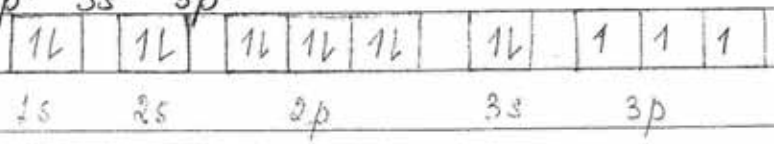
Q.100 Write down the electronic configuration of Phosphorus?

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i) Find out the total no. of e^- in phosphorus for which

electronic configuration of phosphorus :

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$



- $n = 3$

  no. of e^-s = 3 + 2 = 5 electrons
- $n = 3$ and $l = 1$

  no of e^-s = 3 electrons
- $n = 3$ and $m = 0$

  no of e^-s = 3 electrons
- $n = 2$ and $m = -1$

  no of e^-s = 2 electrons
- $l = 1$

  no. of e^-s = 6 + 3 = 9 electrons
- $l = 0$ and $s = +1/2$

  no. of e^-s = 3 electrons

ii) Total spin : $\pm \frac{1}{2} \times$ no. of unpaired electron

$= \frac{\pm 1}{2} \times 3 = \frac{\pm 3}{2}$

iii) Magnetic moment : $\sqrt{n(n+2)}$ BM.

$= \sqrt{3(3+2)}$ BM

$= \sqrt{3 \times 5}$ BM $= \sqrt{15}$ BM

(iv) Write down the set of four quantum numbers for :

4 th electron.

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$n = 2$; $l = 0$ $m = 0$ $s = \pm 1/2$

- 11th electron

$n = 3$; $l = 0$ $m = 0$ $s = \pm 1/2$

- 15th electron

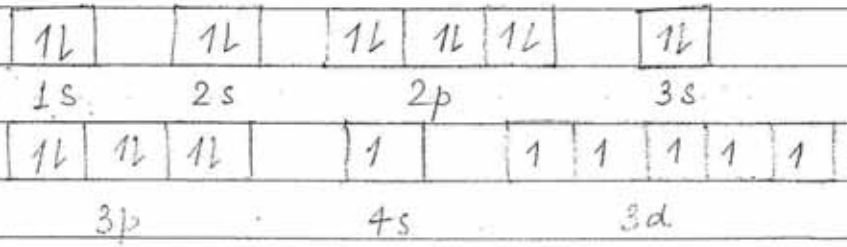
$n = 3$; $l = 1$ $m = +1, 0, -1$ $s = \pm 1/2$

Q.101 For chromium atom determine the followings :

∵ Chromium (Cr) is an exception,

∴ its electronic configuration =

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$



- No of unpaired electrons : 1+5 = 6 electrons
- No of paired electrons : 2+2+6+2+6 = 18 electron
- No of pair of electron : 9 electron pair
- No of s, p and d electron : s-subshell = 7 p subshell = 12 e^-s and d subshell = 5e^-

~~Total no of paramagnetic movement :~~

~~= total no. of unpaired electrons = 6 electrons~~

- No. of p-orbital : 3+3 = 6 orbital
- No of e^- with (n+l) = 4 : (4s and 3p orbitals)

= 6 + 1 = 7 electrons

- No of e^- with $m = 0$: (1s, 2s, 3s, 4s, $3p_z$ and $3d_{z^2}$ Orbitals)

= 2 + 2 + 2 + 2 + 2 + 1 + 1 = 12 electrons

- No of e^- with $l = 1$ and $s = +1/2$: (2p and 3p orbitals) = 3 + 3 = 6 electrons.

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- no. of electrons with $n = 2$ and $l = 0$ (2s orbital) = 2 electrons.
- no. of electrons with $n = 3$ and $m = -2$ (3d orbital) = 1 electron
- Total spin : $\pm \frac{1}{2} \times$ no. of unpaired e^-

$$= \pm \frac{1}{2} \times 6 = \pm 3$$

- Paramagnetic movement : $\sqrt{n(n+2)}$ BM

$$= \sqrt{6(6+2)} \text{ BM}$$

$$\sqrt{6 \times 8} \text{ BM} = \sqrt{48} \text{ BM}$$

- Write down the configuration for 19th and last electron for chromium.

For 19th electron of chromium,

$\because 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ \boxed{4s^1}$

$\therefore n = 4,\ l = 0,\ m = 0,\ s = \pm 1/2$

For last electron of chromium,

$\because 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ \boxed{3d^5}$

$\therefore n = 3,\ l = 2,\ m = -2 \text{ to } +2,\ s = \pm 1/2$

Q.102 Calculate the total spin and magnetic movement of Fe^{+2}, Zn^{+2} and Cr^{+3}

i) Fe^{+2} : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6$

$\therefore$ Total spin = $\pm \frac{1}{2} \times$ no. of unpaired e^-

$$= \pm \frac{1}{2} \times 4 = \pm 2$$

Magnetic movement = $\sqrt{n(n+2)}$ BM

$$= \sqrt{4(4+2)} \text{ BM}$$

$$= \sqrt{4 \times 6} \text{ BM} = \sqrt{24} \text{ BM}$$

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ii) Zn^{+3} : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}$.

Total spin = $\pm \frac{1}{2} \times$ no. of unpaired e^-

$= \pm \frac{1}{2} \times 0 = 0$

Magnetic movement = $\sqrt{n(n+2)}$ BM

$= \sqrt{0(0+2)}$ BM $= 0$

iii) Cr^{+3} : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^3$

Total spin = $\pm \frac{1}{2} \times$ no. of unpaired electrons.

$= \pm \frac{1}{2} \times 3 = \pm \frac{3}{2}$

Magnetic movement = $\sqrt{n(n+2)}$ BM

$= \sqrt{3(3+2)}$ BM

$= \sqrt{3 \times 5}$ BM $= \sqrt{15}$ BM