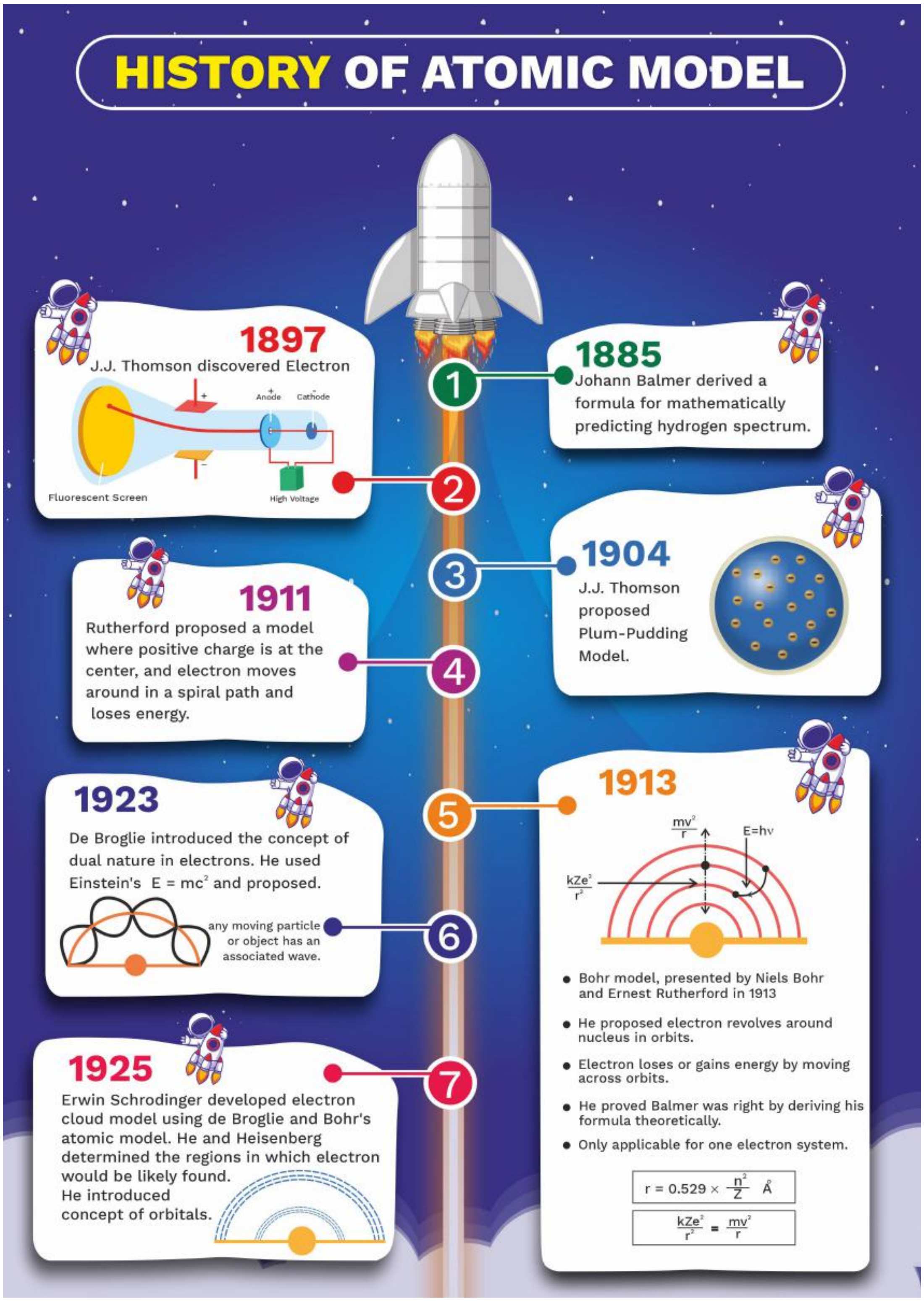
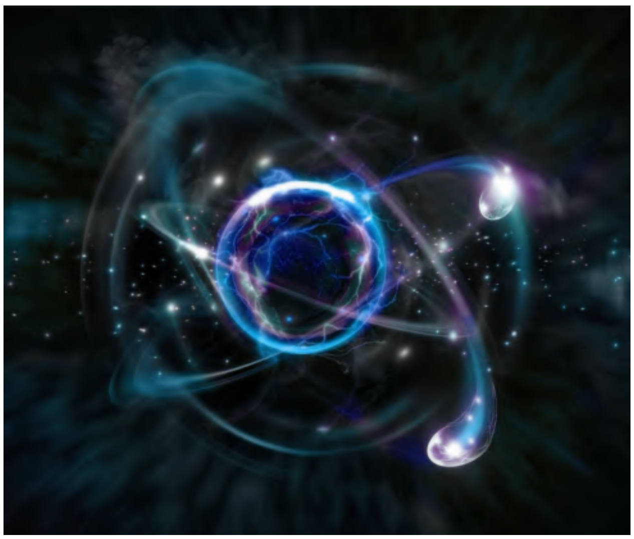
Atomic Structure



Atom

- Each element is composed of smallest particles known as 'ATOM'.
- Atom the name is derived from Greek word Atom means - 'Not to be cut'.



DALTON'S THEORY OF ATOM

J. Dalton developed his famous theory of atom in 1803. The main postulates of

- Atom was considered as a samllest, dense and hard indivisible particle.
- Each element consists of a specific type of atoms.
- The properties of elements differ because of difference in kinds of atom contained in them.
- This theory provides a satisfactory basis for law of chemical combinaiton.

Limitations of Dalton's Theory

- Dalton's theory fails to explain why atoms of different kinds should differ in valency and mass etc.
- The discovery of isobars and isotopes showed that atoms of same elements have different atomic masses (isotopes) and atoms of different kinds may have same atomic masses (isobars).

Definition

Each element is composed of smallest particles called 'ATOM'.

Concept Ladder

Democritus was the person who first suggested the existence of ATOM & coined the name ATOMOS means Not to be cut or Indivisible.

Rack your Brain

Why and how do atoms combine together to form compound atoms?

Concept Ladder

Atoms is indestructible, i.e., it cannot be created or destroyed.

- The discovery of various sub-aomic particles like protons, electrons, X-rays etc. during late 19th century lead to idea that the atom was no longer an indivisible and smallest particle of the matter.

Concept Ladder

R.S. Mullikan measured the charge on an electron by oil drop experiment. The charge on each electron is -1.602×10^{-19} C.

DISCOVERY OF FUNDAMENTAL PARTICLES

- Atom consist of several sub-atomic particles like neutron, proton, electron, neutrino, positron etc. Out of these particles, electron, proton and the neutron are called fundamental subatomic particles.

(1) Electron

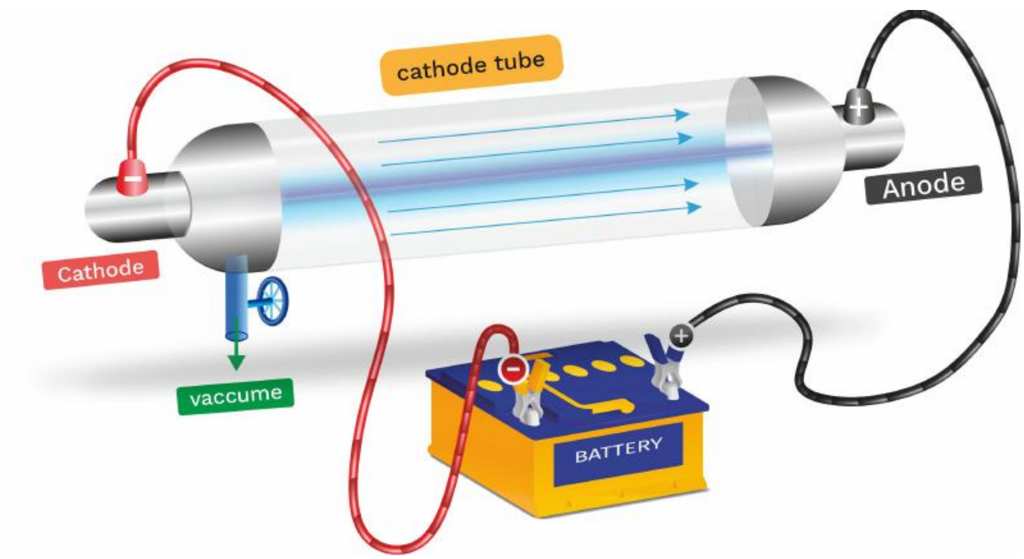
- Electron discovered by J.J. Thomson (1897) and it is negatively charged particle.

Cathode Ray Experiment

- William Crookes in 1879 studied the electrical discharge in partially evacuated tubes called as cathode ray discharge tube.

Rack your Brain

Why, cathode rays do not depend upon the nature of gas or the cathode meterial used in discharge tube?



- Discharge tube is made of glass, about 60 cm long containing two thin pieces of metals known as electrodes, sealed in it. This is called crooke's tube. Negative electrode is called as cathode and positive electrode is called anode.

- When a gas enclosed at low pressure (~10^{-4} atm) in discharge tube is subjected to a high voltage (~10,000 V), invisible rays originating from the cathode and producing a greenish glow behind the perforated anode on the glass wall coated with phosphorescent material ZnS is observed. These rays

Properties

- Cathode rays travel in straight line.
- Cathode rays produce mechanical effect, as they can rotate the wheel that placed in their path.
- Cathode rays consist negatively charged particles called as electron.
- Cathode rays travel with high speed.
- Cathode rays can cause fluorescence.
- Cathoderays heat the object on which they fall due to transfer of kinetic energy to the object.
- When cathode rays fall on heavy metals, X-rays produced.
- Cathode rays possess ionizing power that is they ionize the gas through which they pass.
- The cathode rays produce scintillation on the photographic plates.
- They can penetrate through thin metallic sheets.

Rack your Brain

Why do cathode rays produce fluorescence?

Previous Year's Question

Cathode rays have

[AIPMT]

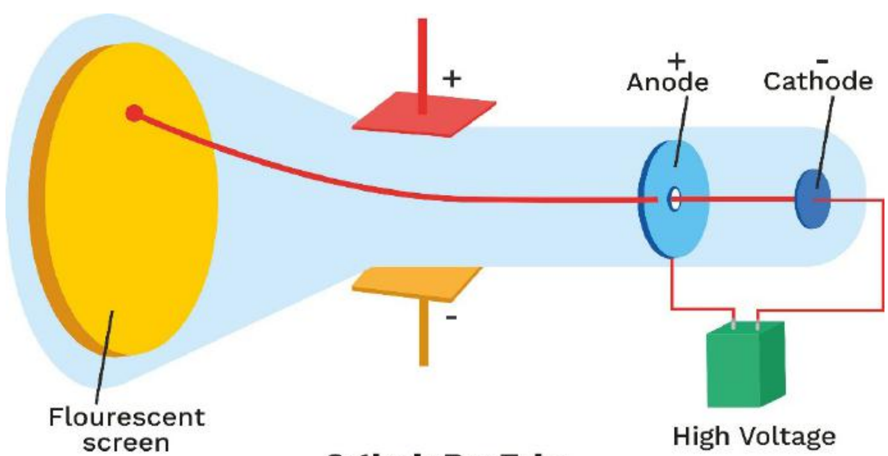
(1) Mass only
(2) Charge only
(3) No mass and charge
(4) Mass and charge both

Concept Ladder

Order of specific charge

$$\left(\frac{e}{m}\right)_n < \left(\frac{e}{m}\right)_p < \left(\frac{e}{m}\right)_e$$

$$\left(\frac{\text{mass of proton}}{\text{mass of electron}}\right) = \frac{m_p}{m_{e^-}} = 1837$$



Cathode Ray Tube

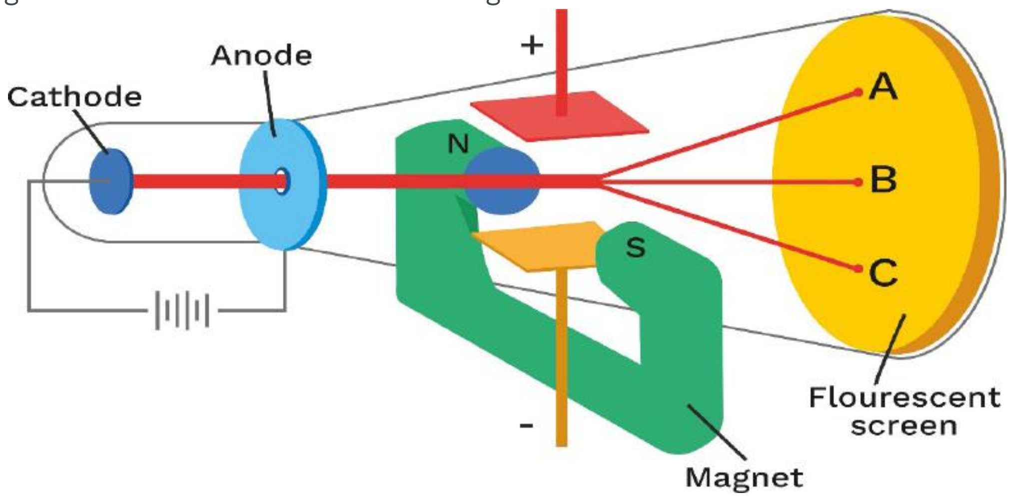
DISCOVERY OF FUNDAMENTAL PARTICLES

Charge by mass ratio of Cathode Ray

Electron is a low-mass, negatively charged particle. It can easily be swing by passing close to other electron or positive nucleus of an atom.

With the help of variation in electric and magnetic fields the ratio was determined. The apparatus is shown below.

The charge to mass ratio of the electron is given



Charge to mass ratio of electron is given by:

$e/m = 1.758820 \times 10^{11}$ C/kg

Where;

m = mass of electron in kg.

e = magnitude of the charge of electron in coulombs.

(2) Proton

- Proton was discovered by Goldstein and it is positively charged particle.

Anode Ray Experiment

Canal Ray experiment is the experiment performed by German scientist Eugen Goldstein in 1886 that led to the discovery of the proton. The discovery of proton which happened after the discovery of the electron further strengthened the structure of the atom. In the experiment, Goldstein applied high voltage across a discharge tube which had a perforated cathode. A faint luminous ray was

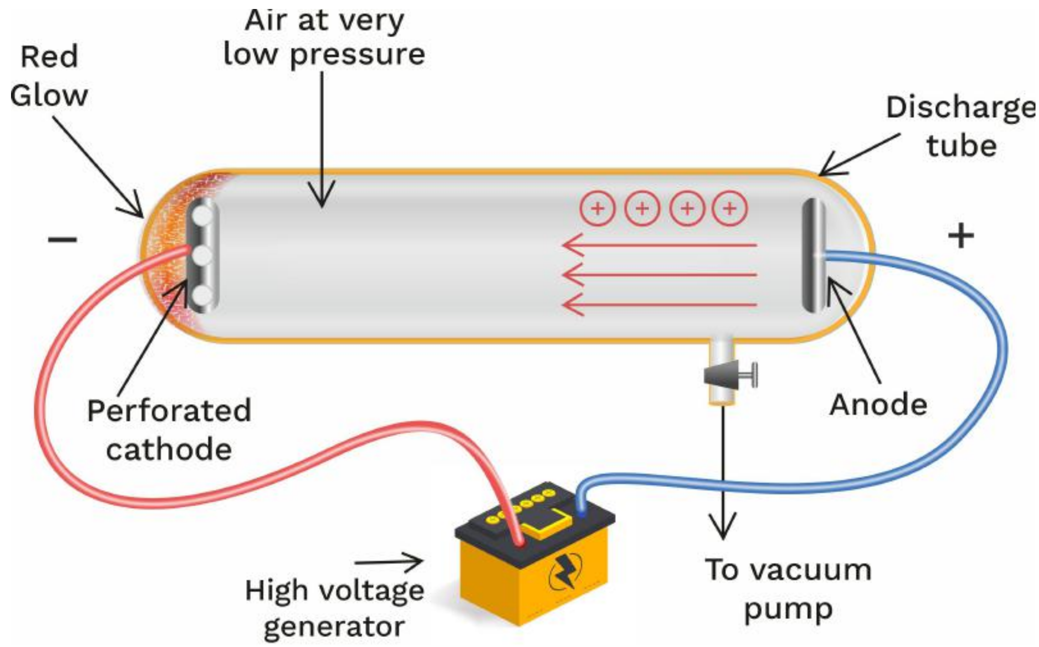
Rack your Brain

What were the important conditions maintained in the discharge tube by Goldstein?

Concept Ladder

Anode rays, e/m value is dependent upon the nature of the gas taken in the tube. It is maximum when gas present in the tube is hydrogen.

seen extending from the holes in the back of the cathode.



Properties

- Anode rays travel in wtraight line.
- Anode rays are material particle.
- Anode rays are positively charged.
- Anode rays may get swing by external magnetic field.
- Anode rays also affect the photographic plate.
- e/m ratio of anode rays is lesser than electrons.
- These anode rays produce flashes of light on ZnS screen.

(3) Neutron

- In 1932, Chadwick bombarded Be with a stream of α-particles (He^{+2}). He noticed that penetrating radiations were produced which not affected by magnetic field and electric field. These radiationss consisted of neutral particles, which were called neutrons.

Concept Ladder

Anode rays were observed by E. Goldstein but was named by E. Rutherford.

Previous Year's Question

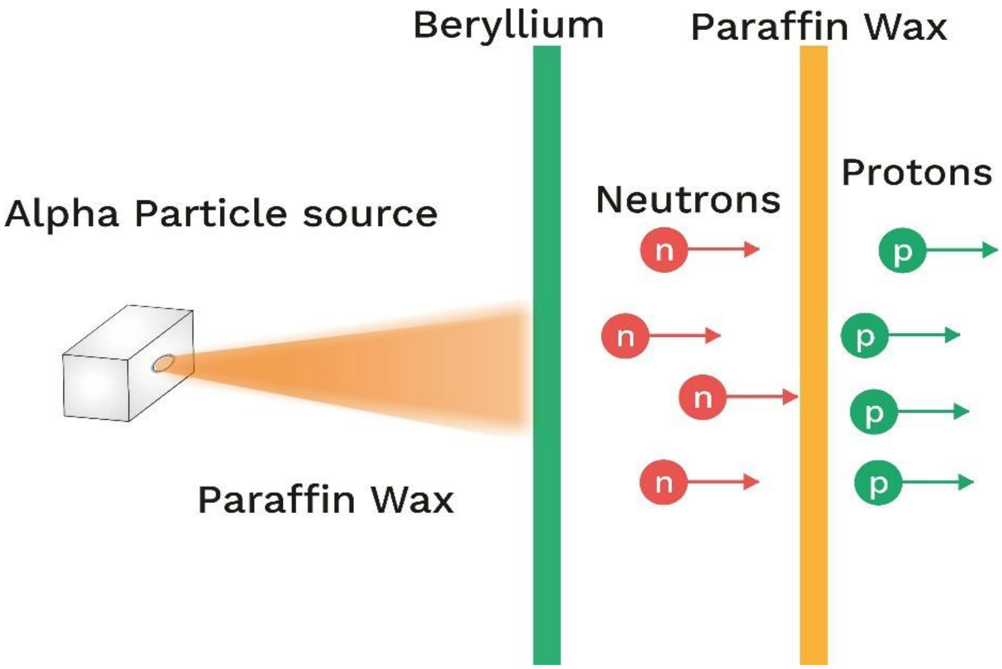
The discovery of neutron become very late because

[AIPMT]

(1) Neutrons are present in nucleus
(2) Neutrons are highly unstable particles
(3) Neutraons are chargeless
(4) Neutrons do not move

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

$${}_{5}B^{11} + {}_{2}He^{4} \rightarrow {}_{7}N^{14} + {}_{0}n^{1}$$



- These radiation was made incident on paraffin wax, a hydrocarbon having a relatively high H_2 content.
- These protons ejected from the paraffin wax (when struck by uncharged radiation) were observed with the help of ionization chamber.
- The range of the liberated protons was measured and the interaction between the uncharged radiation and the atoms of several gases was studied by Chadwick.
- The neutron is relatively massive but neutral, it is scarcely affected by the cloud of electrons surrounding the nucleus or by the positive electrical barrier of the nucleus it self, thus it can penetrate the nucleus of any element

Concept Ladder

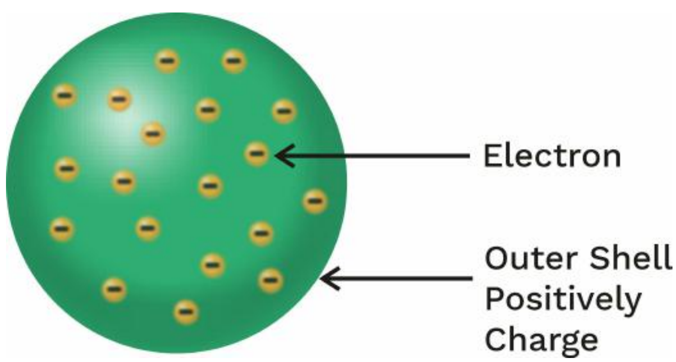
Neutron is fundamental particle of all the atomic nucleus, except hydrogen or protium.

Particles	Symbol	Mass	Charge	Discovered By
Electron	$_{-1}e^0$ or β	9.1096×10^{-31} kg 0.000548 amu	-1.602×10^{-19} Coulombs -4.803×10^{-10} esu	J.J. Thompson Stoney Lorentz 1887
Proton	$_1H^1$	1.6726×10^{-27} kg 1.00757 amu	$+1.602 \times 10^{-19}$ Coulombs $+ 4.803 \times 10^{-10}$ esu	Goldstein Rutherford 1907
Neutron	$_0n^1$	1.6749×10^{-27} kg 1.00893 amu 1 amu $\approx 1.66 \times 10^{-27}$ kg	Neutral 0	James Chadwick 1932

MODELS OF ATOM

(1) Thomson's Model

- Thomson was the first proposed a detailed model of atom.
- Thomson proposed that an atom consist of a uniform sphere of positive (+ve) charge in which electrons are present at some places.
- This model of atom is known as 'Plum-Pudding model'.



Rack your Brain

What is the name of Thomson's model?

Concept Ladder

The colloquial nickname 'plum pudding' was soon attributed to Thomson's model as distribution of electrons within its positively charged region of space reminded many scientists of raisins, then called 'plums'.

Limitations of Thomson's Theory

- An importnat drawback of this model is that the mas of the atoms is considered to be evenly spread over that atom.
- It is a static model. It does not reflect the movement of electron.

(2) Rutherford's α-Scattering Experiment

- Rutherford carried out α-particles scattering experiment by the bombardment of high speed α-particle on thin foil of gold, emitted from radius and gave the following observations, which was based on his experiment.
- The angular deflections of scattered α-particles were studied with the help of moving microscope.

Previous Year's Question

The nucleus of atom consists of **[AIPMT]**

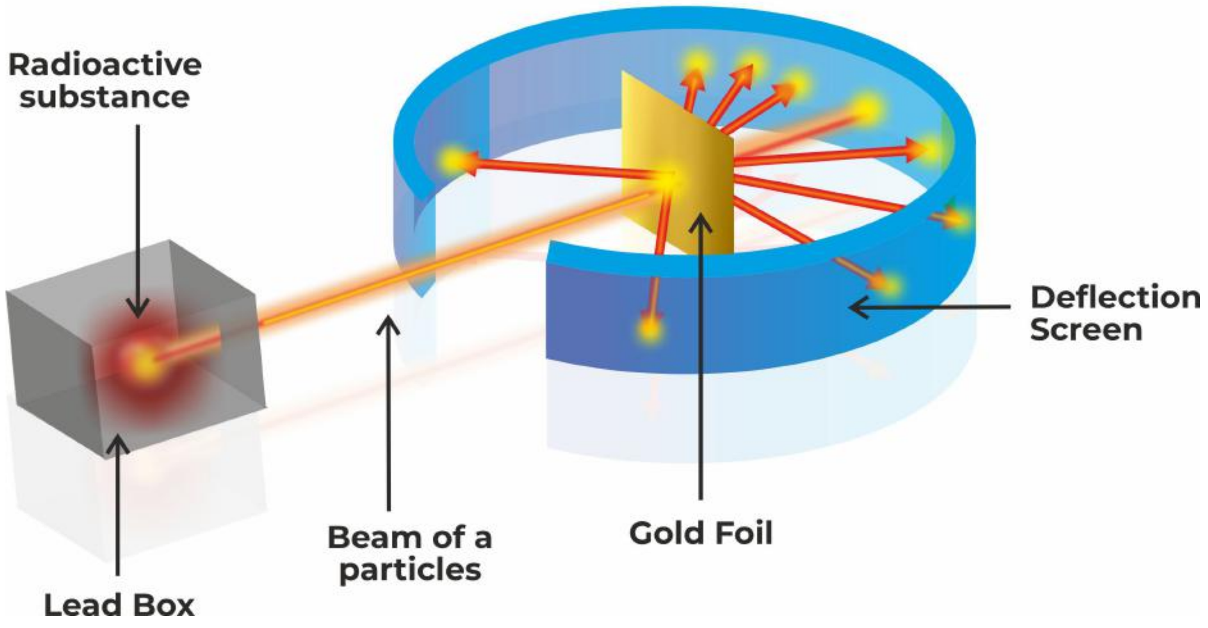
(1) Proton and neutron
(2) Proton and electron
(3) Neutron and electron
(4) Proton, neturon and electron

Concept Ladder

The relation between number of deflected particles and deflection angle θ is

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

Where, μ = deflected particles
θ = deflection angle



Observations

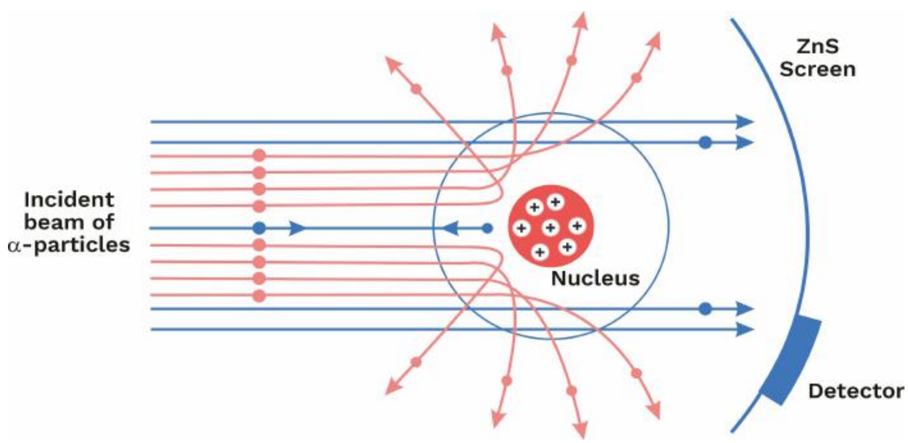
Rutherford execute a number of experiments, involving the scattering of α-particles (He^{+2}) by very thin foil of gold.

Observations were:

- Most of the α-particles (99%) passes through it, without any deviation or deflection.
- Some of the α-particles were deflected through small angles.
- Very few α-particles were deflected by large angles and occasionally an a-particle got deflected by 180°.

Concept Ladder

An atom consists of positively charge center which is known as nucleus e^- revolves around nucleus (+) like solar system



Conclusions

- Most of the α-particle (He+2) went straight through metal foil undeflected, it means that there must be very large empty space within atom.
- Since few of the α-particles were defelcted from their original path through moderate angle; it was concluded that whole of the positive (+ve) charge is concentrated and the space occupied by this positive charge is very small in the atom.
- When α-particles comes closer to this point, they bear a force of repulsion and diverge from their path.

Previous Year's Question

Rutherford experiment prove

[AIPMT]

(1) Electron (2) Proton

(3) Atom (4) Nucleus

- Positively (+ve) charged heavy mass which occupies only a small volume in an atom is known as nucleus. Nucleus is supposed to be present at the centre of atom.
- A very few of the α-particles (He^{+2}) suffered strong diversion on even returned on their path indicating that nucleus is rigid and α-particles (He^{+2}) recoil due to direct collision with heavy positively charged mass.
- As atomic number increases, the number of protons increases which increases the repulstuion and so deflection angle θ increases.
- Atom has two part

(1) Nuclear part → Mass

(2) Extra nuclear part → Size, volume

Size of nucleus = 10^{-13} cm or 10^{-15} m = 1 fermi

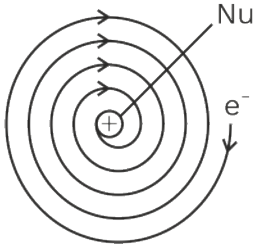
Size of atom = 10^{-8} cm or 10^{-10} m = 1 Å

$$= \frac{r_N}{r_A} = \frac{D_N}{D_A} = \frac{10^{-13}}{10^{-8}} = 10^{-5}$$

- Density of nucleus = + 10^{17} kg /m^3

Limitations of Rutherford's α-scattering experiment

- It does not obey Maxwell theory of electrodynamics. Charge particle in attractive field revolves to emits its energy. So It's loop is reduced and collide with nucleus.



Concept Ladder

$R_n \alpha (A)^{1/3}$

A = mass no

$R_n = R_0 \{A\}^{1/3}$ $\left\{ \begin{aligned} R_0 &= 1.33 \text{ fermi} \\ &= 1.33 \times 10^{-13} \text{cm} \end{aligned} \right\}$

$R_n^3 = R_0^3 (A)$

R_n = Radius of nucleus

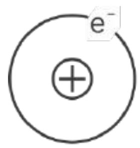
R_0 = constant

Rack your Brain

Why was Rutherford's model rejected?

Concept Ladder

- Mass Number (A) = Number of Protons (Z) + Number of Neutrons (n)
- Number of Neutrons (n) = Mass Number (A) – Number of Protons (Z)

- 

Continuous energy emits $\longrightarrow$ Spectrum $\downarrow$ Should be continuous spectrum

But in reality discontinuous spectrum occur.

ATOMIC NUMBER AND MASS NUMBER

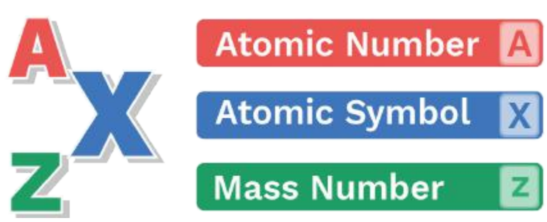
(1) Atomic Number (Z)

- It is always equal to number of protons present in the nucleus of the atom of the element. It also represents the number of electrons in the neutral atom.

  Ex: Number of protons in K is 19, so its atomic number (Z) = 19.

(2) Mass Number (A)

- The neutrons and protons present in nucleus of an atom are collectively known as nucleons.



Atomic Weight

- It is the average of average of weights of the all isotopes of that element.
- An element have two isotopes Z_1 and Z_2, that of isotopes are W_1 and W_2 and their % of occurance in nature are X_1 and X_2 respectively then the average atomic weight of element is —

$$\boxed{\text{avg. wt} = \frac{W_1X_1 + W_2X_2}{X_1 + X_2}}$$

Concept Ladder

Atomic weight may be decimal but mass number of atom always a whole number.

Rack your Brain

How do you find the atomic number?

Previous Year's Question

The number of electrons and neurons of an element is 18 and 20 respectively. Its mass number is : **[AIPMT]**

(1) 17 (2) 37

(3) 2 (4) 38

ISOTOPES, ISOBARS, ISODIPHERS, ISOSTERS AND ISOELECTRONIC

(1) Isotopes

- Isotopes given by Soddy.
- Different atoms of same elements which have similar atomic number but different mass number.
- The chemical properties are controlled by the number of electrons. Thus, isotopes of an element show same chemical behaviour but different physical properties.

Ex: ${}_1H^1$, ${}_1H^2$, ${}_1H^3$

${}_{17}Cl^{35}$, ${}_{17}Cl^{37}$

${}_6C^{12}$, ${}_6C^{13}$, ${}_6C^{14}$

${}_8O^{16}$, ${}_8O^{17}$, ${}_8O^{18}$

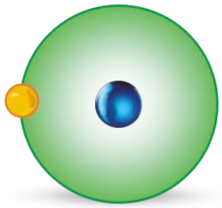
	${}_1H^1$	${}_1H^2$	${}_1H^3$
e^-	1	1	1
p^+	1	1	1
n	0	1	2
*Similar e^- , P^+ different n			
*n + p = 1,2,3			
(nucleons) different no. of nucleons			
*Nuclear charge same			

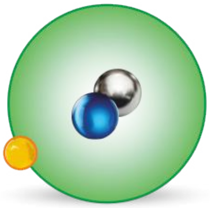
Concept Ladder

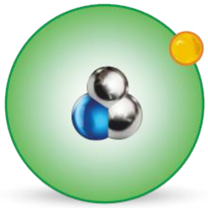
There are mainly two types of isotopes. These are radioactive and stable isotopes. Stable isotopes have a stable nuclei and do not undergo decay.

Rack your Brain

Isotopes and Isobars both do not have same value of e/m. Why?


Protium


Deuterium


Tritium

(2) Isobars

- Isotopes given by Aston.
- These are the different atoms of different elements which have similar mass number but different atomic number.
- Isobars do not show the same chemical properties.

Ex: $_{6}C^{14}$, $_{7}N^{14}$

$_{17}Cl^{37}$, $_{18}Ar^{37}$

$_{32}Ce^{76}$, $_{34}Se^{74}$

$_{26}Fe^{58}$, $_{27}Ni^{58}$

	$_{6}C^{14}$	$_{7}N^{14}$
e^-	6	7
p^+	6	7
n	8	7
*Different no of e^- , P^+ , n		
*n + p = 14, 14		
Same nucleons		
*Nuclear charge different		

(3) Neutron

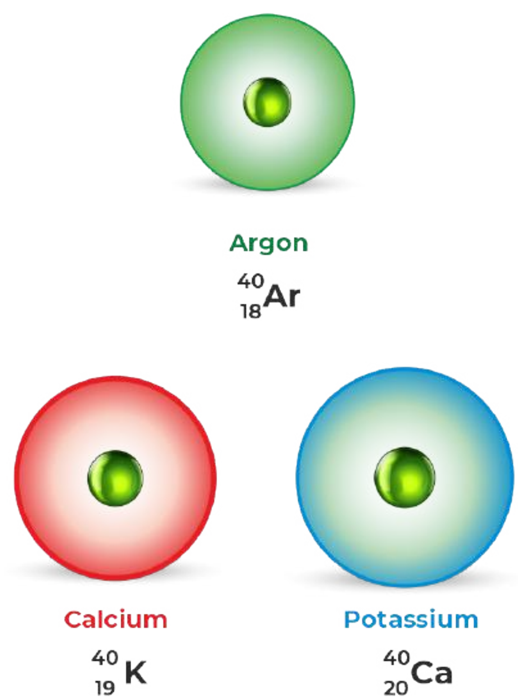
Rack your Brain

Why isobars have same physical properties?

Previous Year's Question

The nucleus of tritium contains **[AIPMT]**

(1) 1 proton + 1 neutron
(2) 1 proton + 3 neutron
(3) 1 proton + 0 neutron
(4) 1 proton + 2 neutron



(3) Isotones

The atoms of different element which have the same number of neutrons.

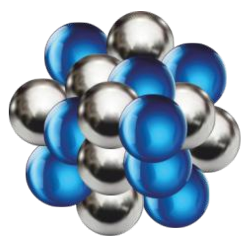
Ex:

$_{19}K^{39}$	$_{20}Ca^{40}$
n = 39 – 19	n = 40 – 20
n = 20	n = 20

$^{14}_{6}C$

No. Of Neutron

14-6 = 8



$^{16}_{8}O$

No. Of Neutron

16-8 = 8

(4) Isodiaphers

The atoms of different element which have the same difference of the number of Neutrons and protons.

	$_5B^{11}$	$_6C^{12}$	$_7N^{15}$	$_8O^{17}$
p^+	5	6	7	8
n	6	7	8	9
e^-	5	6	7	8
$(n - p^+)$	1	1	1	1

Previous Year's Question

An isotone of $_{32}Ge^{76}$ is

[AIPMT]

(1) $_{32}Ge^{77}$ (2) $_{33}As^{77}$

(3) $_{34}Se^{77}$ (4) $_{36}Sc^{77}$

Rack your Brain

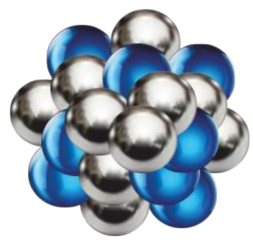
What is the difference between isotopes and isobiaphers?

Previous Year's Question

The number of protons, neutrons and electrons in $^{175}_{71}Lu$, respectively are :

[NEET-2020]

(1) 71, 104 and 71
(2) 104, 71 and 71
(3) 71, 71 and 107
(4) 175, 104 and 71



$^{14}_{6}C$

No. Of Neutron

14-6 = 8

$^{18}_{8}O$

No. Of Neutron

18-8 = 10

difference in protons and neutrons

8-6=6 **10-8=2**

(5) Isosters

The molecules which have the same number of atoms and electrons.

	CO_2	N_2O
Total Atoms	3	3
No. of e^-	$6 + 8 \times 2 = 22\ e^-$	$7 \times 2 + 8 = 22\ e^-$

(6) Isoelectronic

Atoms / ions/ molecules having similar no. of e^-

Ex:

H_2O and NH_3

$10e^-$ $10e^-$

Cl^- and Ar

$18e^-$ $18e^-$

$B_3N_3H_6$ and C_6H_6

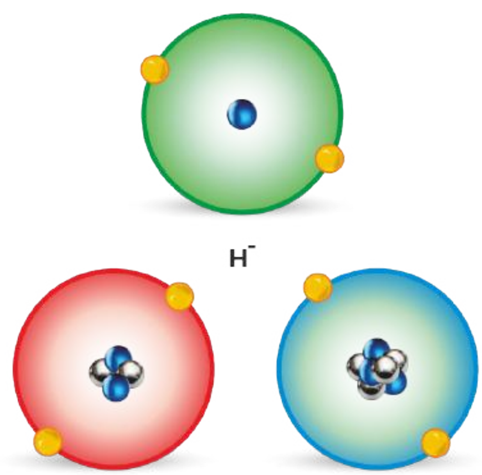
$42e^-$ $42e^-$

Previous Year's Question

Be^{2+} is isoelectronic with **[AIPMT-2014]**

(1) Mg^{2+} (2) Na^+

(3) Li^+ (4) H^+

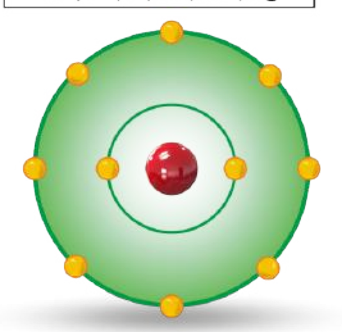


H^-

He **Li^+**

ISOELECTRONIC

$N^{3-}, O^{2-}, F^-, Ne, Na^+, Mg^{+2}$



(7) Relative Abundance

Isotopes of an element occur in different percentages in nature, which is termed as relative abundance.

Using this relative abundance the average atomic mass of the element can be calcualted.

Ex: ^{35}Cl and ^{37}Cl

(75% and 25% abundancve respectively)

WAVE

- A wave motion is a mean of transfer of energy from one point to another point without any conveying of matter between the points.
- When we throw the piece of stone particles on water surface in a pond, we observe circles of ever increasing radius, till these strike on the wall of the pond.
- When we put piece of cork on the surface of this water, we observe that the cork moves up and down as the wave passes, but the piece does not travel along with the waves.

A wave is characterized by six characterstics

(1) Wavelength (λ)

- It is defined as distance between two nearest crest or nearest through. It is measured in term of a Å (Angstrom), pm (Picometer), nm (nanometer), cm (centimeter), m (meter)

$1Å = 10^{-10}$ m, $1Pm = 10^{-12}$ m

$1nm = 10^{-9}$ m, $1cm = 10^{-2}$ m

(2) Frequency (ν)

- Frequency of a wave is a number of waves which pass through a point in 1 sec. it is measured in term of Hertz (Hz), sec^{-1}, or cycle per second (cps)

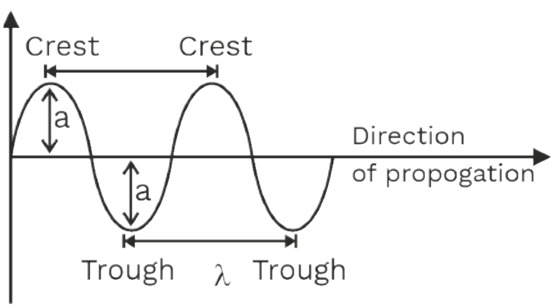
1 Heartz = 1 sec^{-1} = 1 cps.

Concept Ladder

The upper most point of the wave is called crest and the lower most point is called trough.

Rack your Brain

What frequency is harmful to humans?



Previous Year's Question

The frequency of an electromagnetic radiation is 2×10^6 Hz. What is the wavelength in metres

[AIPMT]

(1) 6.0×10^{14} (2) 1.5×10^4
(3) 1.5×10^2 (4) 0.66×10^{-2}

(3) Time period (T)

- Time taken for one complete oscillation of wave is known as period (T). Time taken by wave to travel a distance equal to one wavelength. If C is the speed of wave, then $C = \frac{\lambda}{T}$

(4) Wave Number ($\bar{\nu}$)

- Number of wavelength per unit length. $\bar{\nu} = \frac{1}{\lambda}$

(5) Amplitude (A)

- It is the height of depth or crest of a through of a wave.

(6) Velocity (C)

- It is defined as distance covered by a wave in 1 sec. $\nu = \frac{C}{\lambda}$

Electromagnetic Waves (EMW)

- It contains electric and magnetic field.
- Energy is always transferred in the form of waves with the speed of light (3×10^8 m/s).
- It's a pure energy waves.
- It does not contain mass no medium is required for transmission.
- Direction of propagation is perpendicular from both electric field and magnetic field.
- There are various types of electromagnetic waves (radiation) which differs from one another in wavelengths.
  Ex: Cosmic Rays, g-rays , X-rays, U.V, visible, I.R, Micro, Radio.

Concept Ladder

The amplitude of a wave is related to the amount of energy it carries. The sound is perceived as louder if the amplitude increases, and softer if the amplitude decreases.

Rack your Brain

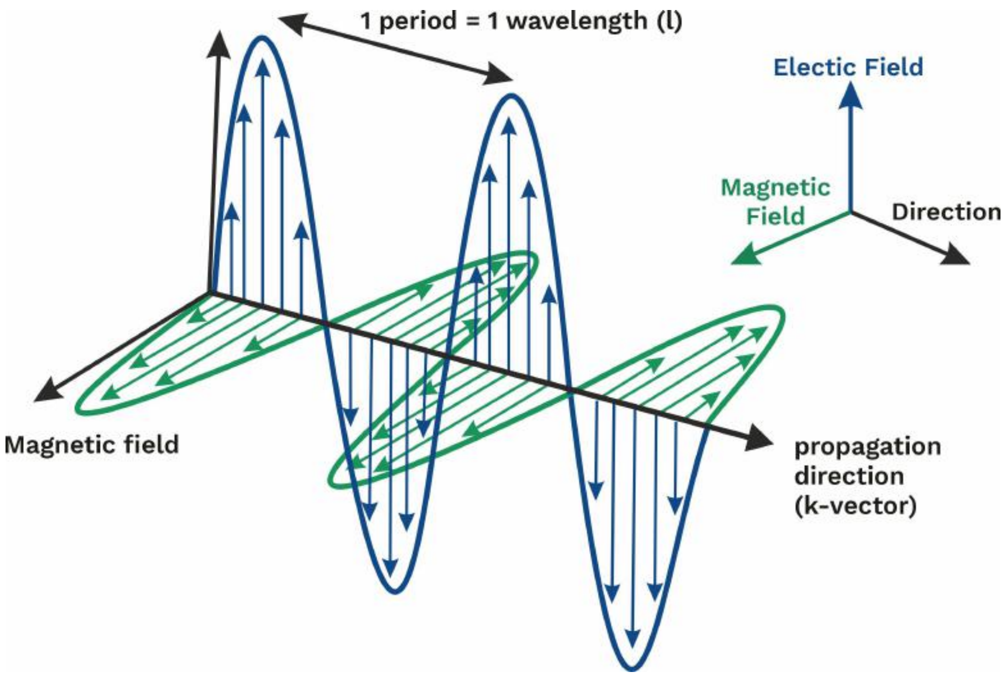
What is the speed of EMW through the vacuum?

Concept Ladder

Near infrared waves are used in remote controls. Far infrared waves are radiant heat. Gamma rays have the greatest energy.

Rack your Brain

What would happen if there was no electromagnetic spectrum?



Electromagnetic Spectrum

- Arrangement of varous types of electrognetic radiations in order of their increasing (or decreasing) wavelengths or frequencies is known as electromagnetic spectrum.

Maxwell Theory of Electromagnetic Wave

- All the radiations have wave nature which explains interference (linear superposition) and diffraction.
- They consist of oscillating electric and magnetic field perpendicular to each other and to the direction of propagation.
- All the radiations (radio waves, micro waves, infra red waves, visible, UV, X-rays, g-rays) travel at the speed of light in vaccum.
- Energy of electromagnetic wave is proportional to amplitude and not linked with frequence of waves.

Rack your Brain

How important are EM waves in our lifes?

Previous Year's Question

Electromagnetic radiation with maximum wavelength is

[AIPMT]

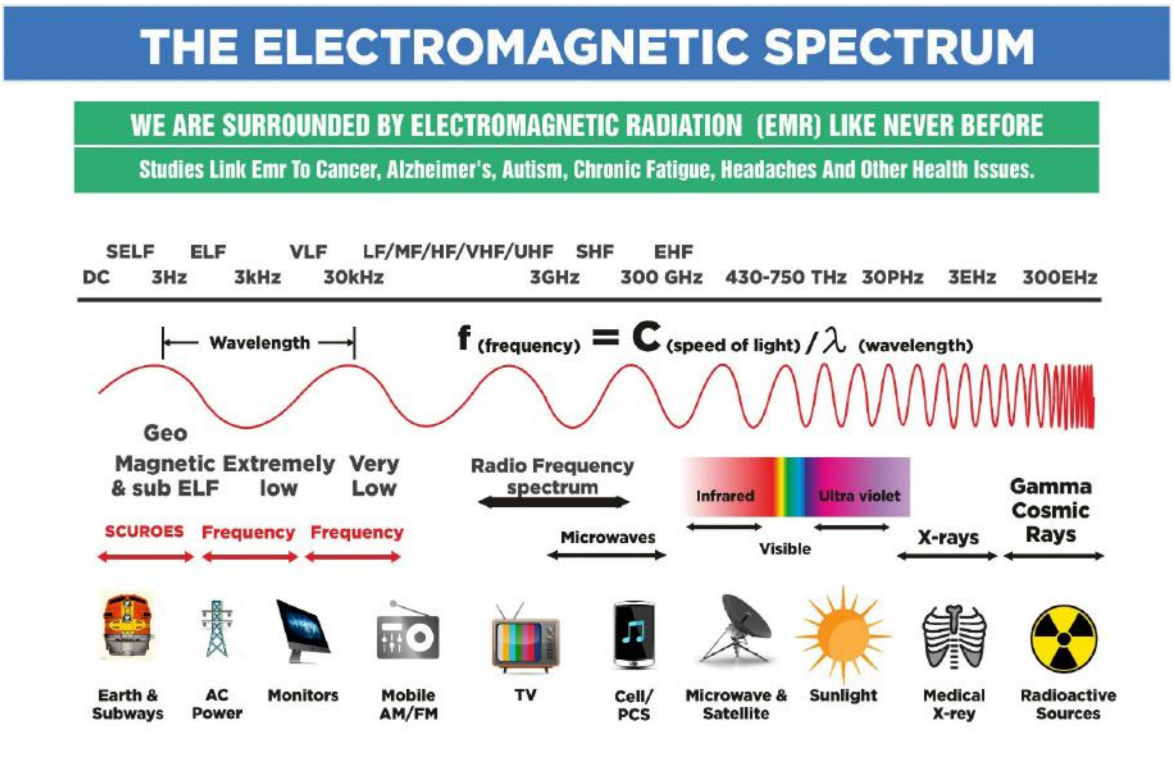
(1) Ultraviolet
(2) Radiowave
(3) X-ray
(4) Infrared

Limitations of Maxwell Theory of Electromagnetic Wave

- Phenomenon of black body radiations.
- Photoelectric effect.
- Line spectra of atoms

Rack your Brain

Why was Maxwell theory not accepted?



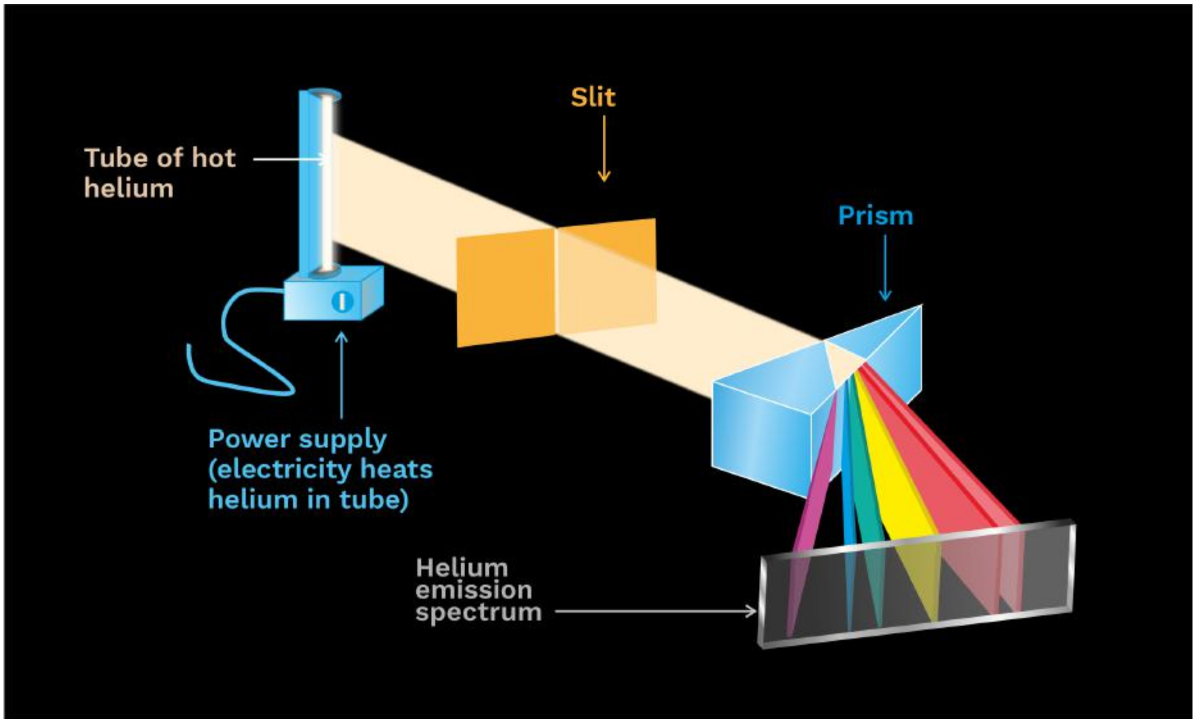
SPECTRUM

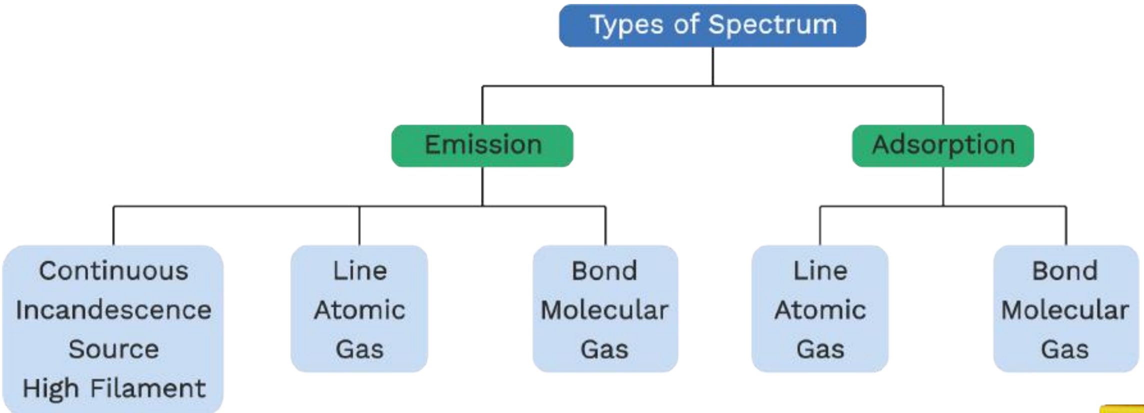
Low energy waves ⟵⟶ High energy waves

- When light coming from a source is scatterd by a prism, light of different wavelength are deviated through different angles and get separated. This phenomenon is known as dispersion and such a dispersed light may be received on photographic plate or it's may be viewed directly by eye. A collection of dispersed light giving its wavelength composition is called a spectrum.

Concept Ladder

Most of the light in the universe is invisible to our eyes. The light we can see, made up of the individual colors of the rainbow.





Types of Spectrum

(1) Emission spectrum

Any substance on heating gets excited on absorbing energy or at a very high temperature or afterwhich radiations are emitted from the substance. The radiations when analysed with the help of spectroscope, spectral lines are obtained.

Emission spectrum may be classified as :

Concept Ladder

Emission spectroscopy is referred to as optical emission spectroscopy because of the light nature of what is being emitted.

(i) Continuous spectrum

When sunlight is passed through a prism, it gets dispersed into continuous bands of different colours.

(ii) Line spectrum

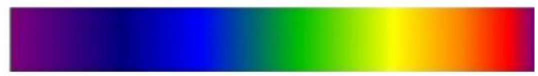
If the radiations obtained by the excitation of a substance are analysed with the help of a spectroscope, a series of thin bright lines of specific colours are obtained. There is dark space in between two consecutive lines. This type of spectrum is called line spectrum or atomic spectrum.

(2) Absorption spectrum

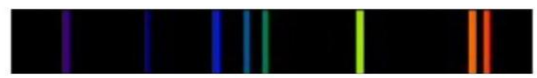
When white light of an incandescent substance is passed through any other substance, this substance absorbs the radiations of particular wavelength from the white light. On analysing the transmiited light we obtain a spectrum in which dark lines of specific wave lengths are observed. These lines constitute the absorption spectrum.

Rack your Brain

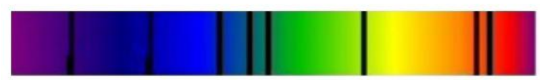
Why is the sun a continuous spectrum?



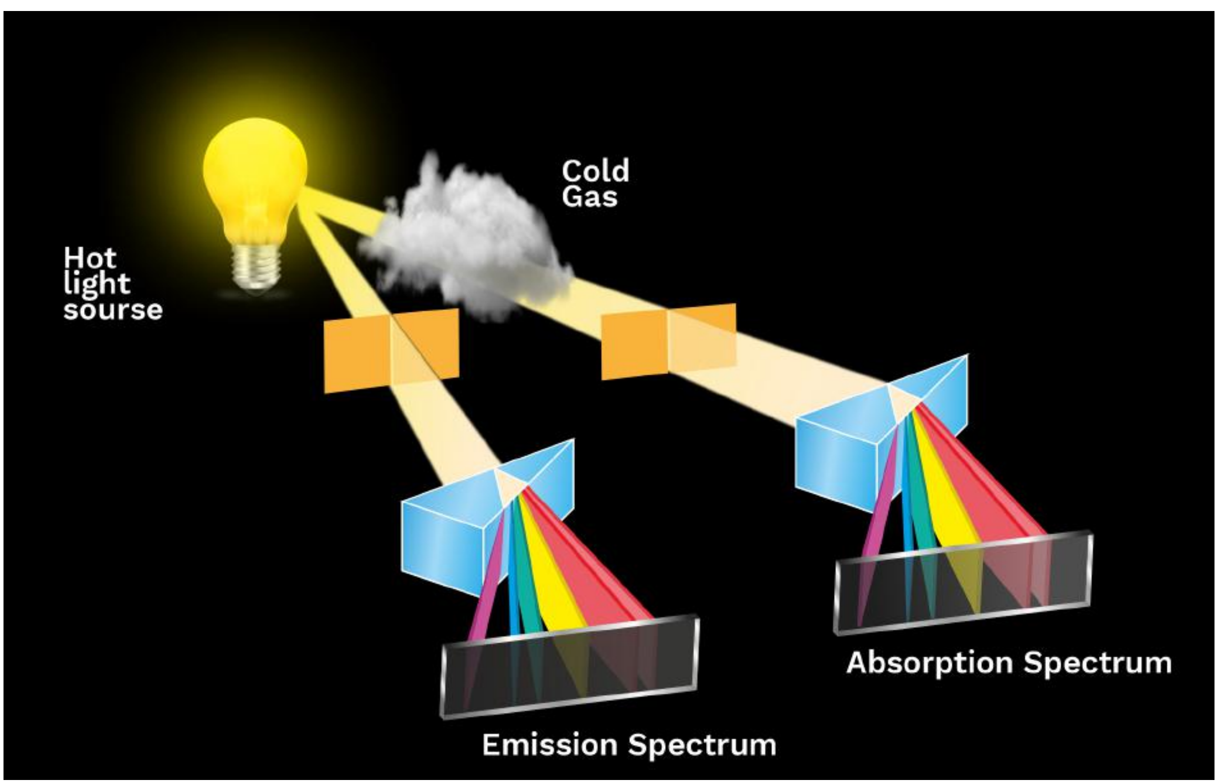
Continuous Spectrum



Emission Spectrum



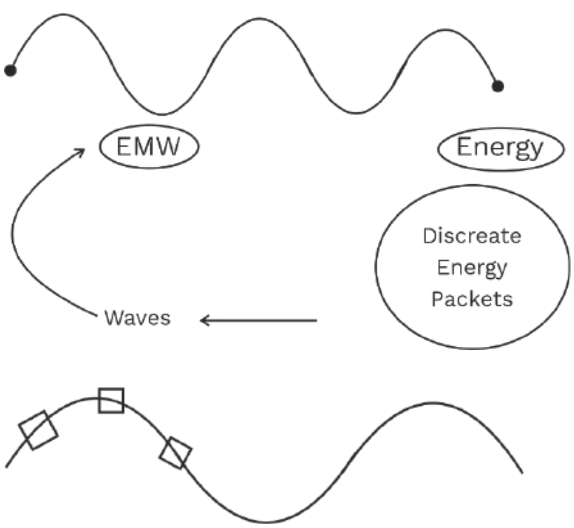
Absorption Spectrum



PLANCK'S QUANTUM THEORY

Diffraction and interference are explained by wave nature of electromagnetic radiation. However, some of the observations are given which could not be explained with the help of even the electromagnetic theory 19th century physics (konwn as classical physics) :

- Nature of emission of radiation from hot bodies (black -body radiation).
- Ejection of electrons from metal surface when radiation strikes it (photoelectric effect)
- Variation of heat capacities of solids is a function of temperature.
- Line spectra of atoms with special reference to hydrogen.
- According to this theory, atoms or molecules can emit or absorb energy only in discrete quantities (small packets) and not in any arbitrary amount. Planck gave name quantum to the smallest quantity of energy that can be emitted in the form of E.N. radiation.



Quanta
Quantum

Definition

The radiant energy emitted or absorbed by a body discontinuously in the form of small discrete packets. These packets are called quantum.

Concept Ladder

Body can emit or absorb energy as hν, 2hν, ... but it can not emit or absorb energy in fractional values of hν such as 1.5 hν, 2.5 hν.

Previous Year's Question

Calculate the energy in joule corresponding to light of wavelength 45 nm.

[NEET-2014]

(1) 6.67×10^{15}
(2) 6.67×10^{11}
(3) 4.42×10^{-15}
(4) 4.42×10^{-18}

- Energy of photon is proportional to frequency and is given by

  $E \propto \nu$

  $E = h\nu$; h = Planck's constant

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

  $E = \frac{hc}{\lambda}$; $\left\{\upsilon = \frac{c}{\lambda}\right\}$

- A body can emit or absorb energy only in terms of the integral multiples of quantum, i.e.

  $E = n.\ h\nu$, where n = 1, 2, 3, ...

Black Body Radiation

- Black body radiation phenomenon first given by Max Planck in 1900.
- The ideal body, which emits and absorbs radiations of all frequencies, is called black body and the radiation emitted by that body is known as black body radiation.

Previous Year's Question

The value of Planck's constant is 6.63×10^{-34} Js. The speed of light is 3×10^{17} nm s^{-1}. Which of the given values is closest to the wavelength in nanometer of a quantum of light with frequency of 6×10^{15} s^{-1}?

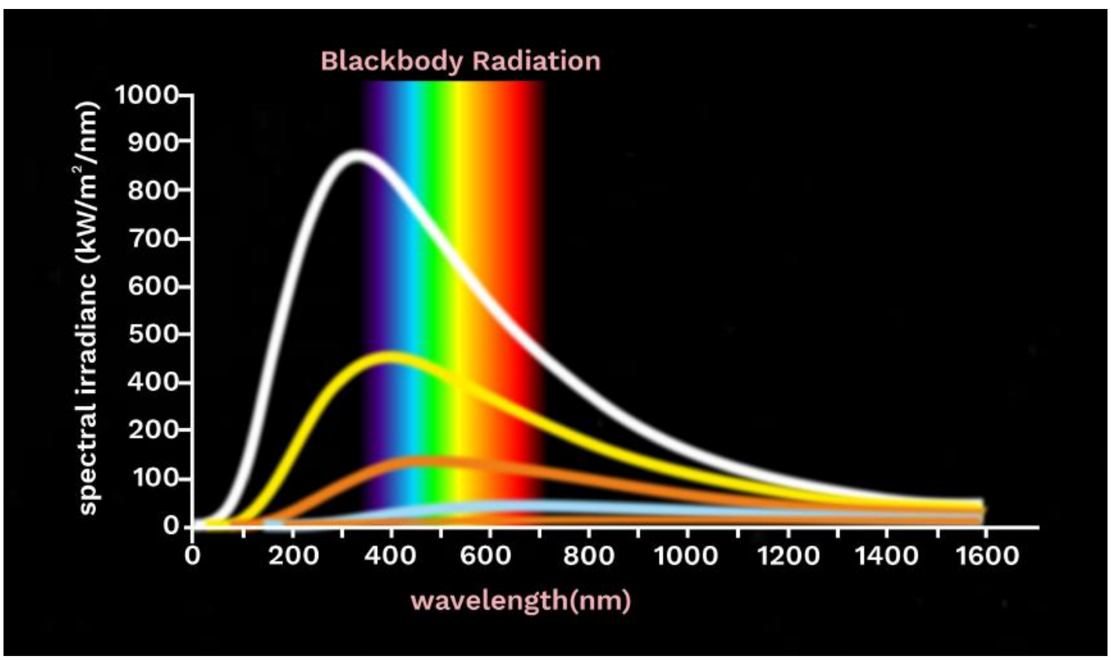
[NEET-2013]

(1) 50 (2) 75

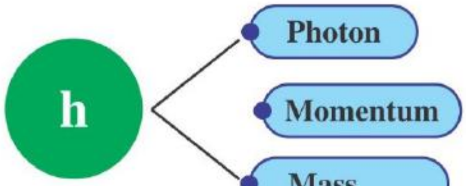
(3) 10 (4) 25

Concept Ladder

Blackbody is a surface that absorbs all radiant energy falling on it. The term arises because incident visible light will be absorbed rather than reflected, and therefore the surface will appear black.



Quanta Packet



Energy Transfer

Q1 3×10^8 photons of a certain light radiation are found to produce 1.5 J of energy. Calculate the wave length of light radiation ?

A1 $E_{Total} = n \times \{h\upsilon\}$

$$E_{Total} = n \times \left\{\frac{hc}{\lambda}\right\}$$

$$1.5\text{ J} = 3 \times 10^8 \left\{\frac{6.6 \times 10^{-34} \times 3 \times 10^8}{\lambda}\right\}$$

$\lambda = 3.96 \times 10^{-17}$ m

Q2 100 watt bulb emits monochromatic light of wave length = 400 nm. Calculate the no. of photons emitted per second?
(1) 5×10^{20} (2) 3×10^{20} (3) 4×10^{20} (4) 2×10^{20}

A2 (4)

t = 1 s

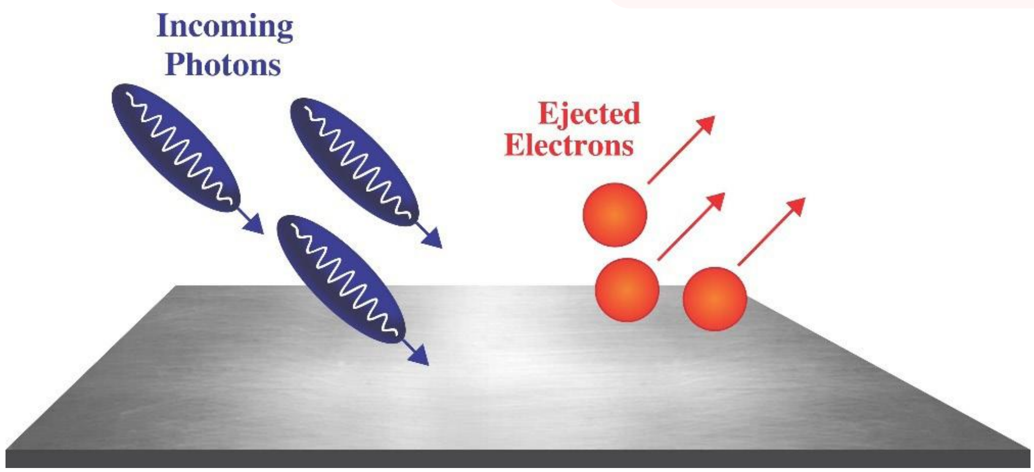
$$E_{total} = \frac{nhc}{\lambda}$$

$$100 = \frac{n \times 6.6 \times 10^{-34} \times 3 \times 10^8}{400 \times 10^{-9}}$$

$n = 2 \times 10^{20}$ no. of photons per second

PHOTOELECTRIC EFFECT

The ejection of electrons when light of certain minimum frequency called as threshold frequency is incident on a metal surface is called as photoelectric effect.



Rack your Brain

Which metal is best for photoelectric effect?

Threshold Frequency

The minimum frequency of incident light which can cause photo electric emission i.e. this frequency is just able to eject electrons with out giving them additional energy.

Work Function

The minimum quantity of energy which is required to remove an electron to infinity from the surface of a given solid, usually a metal.

Incident energy = Work Function (ϕ) + K.E.$_{max}$

$$E_i = \phi + (K.E.)_{max}$$

$$h\upsilon = h\upsilon_0 + \frac{1}{2}m_e\upsilon^2$$

Where me is mass of the electron and v is the velocity associated with the ejected electron.

Concept Ladder

The maximum kinetic energy of photoelectrons depends on the frequency of incident radiation; but, it is independent of the intensity of light used.

Some facts of Photoelectric Effect

- There is zero time lag between incidence of light and emission of photoelectrons.

Previous Year's Question

In photoelectric effect, the kinetic energy of photoelectrons increases linearly with the

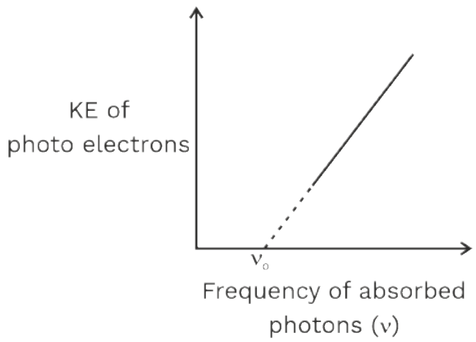
[AIPMT]

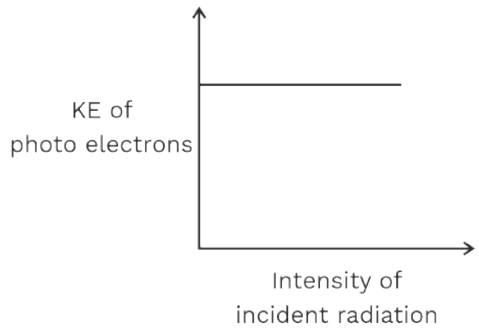
(1) Wavelength of incident light
(2) Frequency of incident light
(3) Velocity of incident light
(4) Atomic mass of an element

- For the emission of photoelectrons, frequency of incident light must be equal to or greater than the threshold frequency.
- Rate of emission of photoelectrons from a surface of metal is directly proportional to the intensity of incident light.

Concept Ladder

The minimum potential at which the photoelectric current becomes zero is called stopping potential.





Q3 The threshold frequency ν_0 for a metal is $6 \times 10^{14}\ s^{-1}$. Calculate the kinetic energy of an electron emitted when radiation of frequency $\nu = 1.1 \times 10^{15}\ s^{-1}$ hits the metal.

A3

$$K.E. = \frac{1}{2} m_e V^2 = h(\nu - \nu_0)$$

$\therefore$ K.E. $= (6.626 \times 10^{-34})(1.1 \times 10^{15} - 6 \times 10^{14})$

$\therefore$ K.E. $= (6.626 \times 10^{-34})(5 \times 10^{14}) = 3.313 \times 10^{-19}$J

BOHR'S ATOMIC MODEL

It is a quantum mechanicla model. This model based onquantum theory of radiation and classical law of physics. This model explain the stability of the atom and emission of sharp spectral lines.

Rack your Brain

Do emitted photoelectrons have same kinetic energy?

Bohr's Atomic Model

- Worked with JJ Thomson and found flaws in his theory.
- He proposed electron revolves around nucleus in orbits.
- Electron is stabilized by centripetal and electrostatic force.
- Electron don't lose energy in an orbits.
- Electron loses or gains energy by moving across orbits.
- He proved Blamer was right by derving his formula theoretically.
- Only applicable for one electron systems.
- Failed to predict dual nature of electron.

1913

Bohr Model's Postulates

- Atom has a central core nucleus where the protons and neutrons are present. Size of the nucleus is very small.
- Negatively charged electron are revolving around the nucleus in the same wasy as the planets are revolving around the sun. The path of electron is circular.
- Electrons can revolve only in those orbits whose angular momentum (mvr) is integral multiple of $\frac{h}{2\pi}$.

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

- Absorption or emission of radiation by an atom takes place when an electron jumps from on stationary orbit to another.
- The radiation is emitted or absorbed as a single quantum (photon) whose energy is equal to the difference in energy of the electron in the two orbitals involved. Thus, $\Delta e = h\nu$, where h = Planck's constant and n = frequency of the radiant energy. Hence the spectrum of the atom will have certain fixed frequence.
- The lowest energy state (n = 1) is called the group state. After absorption of energy,electron gets excited and jumps to an outer orbit. It has to fall back to a lower orbit with the release of energy.

Concept Ladder

Bohr's theory satisfactorily explains the spectra of species having opne electron, viz. H, He^{+}, Li^{2+} etc.

Rack your Brain

Angular momentam is integral multiple of $h/2\pi$. Why not fractional multiple is possible?

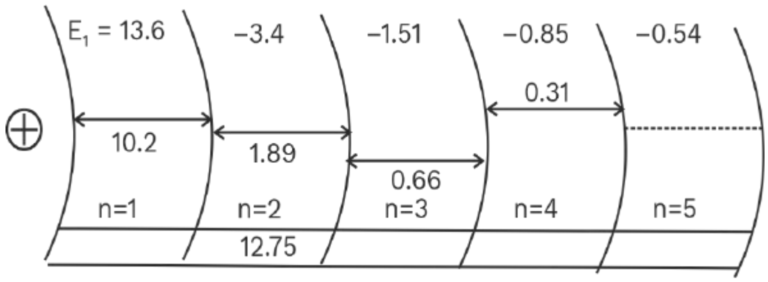
Concept Ladder

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

$$r = 0.529 \times \frac{n^2}{Z} A$$

$$v = 2.188 \times 10^6 \times \frac{Z}{n} m/sec$$

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

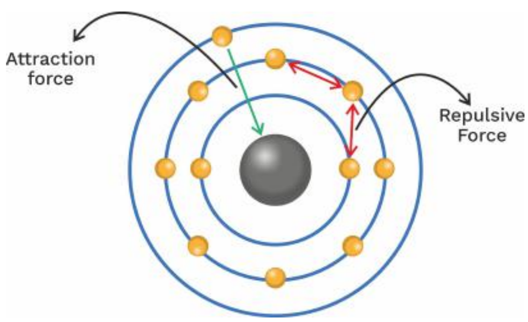


$$mvr = \frac{1h}{2\pi} \quad \frac{2h}{2\pi} \quad \frac{3h}{2\pi} \quad \frac{4h}{2\pi} \quad \frac{5h}{2\pi}$$

Radius of the Bohr's Orbit

Let us consider an electron of mass 'm' and charge 'e' is revolving around nucleus having charge 'Ze' (where Z is atomic number & e is charge) with a linear or tangential velocity of 'v'. Further, let us consider that 'r' is radius of orbit in which electron is revolving.

According to Coulomb's law,electrostatic force of attraction (F) between moving electron and nucleus is –



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

Where : K = constant $= \frac{1}{4\pi\varepsilon_o} = 9 \times 10^9 \, Nm^2 / C^2$

and the cnetripetal force $F = \frac{mv^2}{r}$

Hence $\frac{mv^2}{r} = \frac{KZe^2}{mr}$

or $\quad v^2 = \frac{KZe^2}{mr} \quad$ ......(1)

From the postulate of Bohr,

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

or $\quad v^2 = \frac{n^2h^2}{4\pi^2m^2r^2} \quad$ ......(2)

Previous Year's Question

If r is the radius of the first orbit, the radius of n^{th} orbit of H-atom is given by

[AIPMT]

(1) rn^2 (2) rn
(3) r/n (4) r^2n^2

From equation (1) and (2)

$$\therefore \qquad r = \frac{n^2h^2}{4\pi^2 mKZe^2}$$

On putting value or e, h, m,

$$r = 0.529 \times \frac{n^2}{Z} A$$

Velocity of an electron in Bohr's Orbit

The total energy of an electron is revolving in a particular orbit is —

$$mvr = \frac{nh}{2\pi} \qquad v = \frac{nh}{2\pi mr}$$

Purring the value of r in above equation then

$$v = \frac{nh \times 4\pi^2 mZe^2}{2\pi mn^2h^2}$$

$$v = \frac{2\pi Ze^2}{nh}$$

on purring the values of e and h

$$v = 2.188 \times 10^6 \times \frac{Z}{n} m/sec$$

Calculation of energy of an electron

The total energy of an electron revolving in a particular orbit is —

T.E. = K.E. + P.E.

The K.E. of an electron $= \frac{1}{2}mv^2$

and the P.E. of an electron $= -\frac{KZe^2}{r}$

$$\text{Hence, T.E.} = \frac{1}{2}mv^2 - \frac{KZe^2}{r} \qquad \text{......(3)}$$

$$\text{But } \frac{mv^2}{r} = \frac{KZe^2}{r^2} \quad \text{or} \quad mv^2 = \frac{KZe^2}{r}$$

Substituting value of mv2 in the equation (3)

$$\text{T.E.} = \frac{KZe^2}{2r} - \frac{KZe^2}{r} = \frac{KZe^2}{r}$$

$$\text{So, T.E.} = -\frac{KZe^2}{2r}$$

Rack your Brain

What will be the value of Kinetic energy and Potential Energy at n = ∞?

Concept Ladder

Bohr's atomic model explained the stability of an atom. According to Bohr, an electron revolving in a particular orbit cannot lose energy. Therefore, emission of radiation is not possible as long as the electron remains in one of its energy levels and hence there is no cause of insatbility in his model.

Previous Year's Question

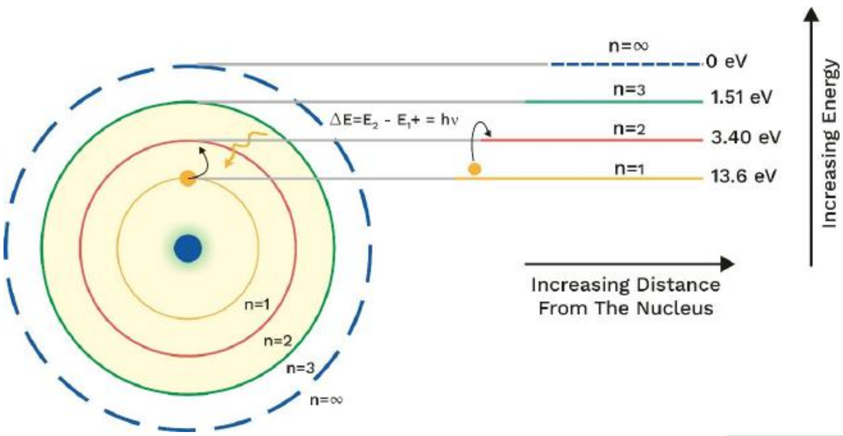
The energy of second Bohr orit of the hydrogen atom is –328 kJ mol^{-1}; hence the energy of fourth Bohr orbit would be

[AIPMT]

(1) – 41 kJ mol^{-1}
(2) – 82 kJ mol^{-1}
(3) – 164 kJ mol^{-1}
(4) – 1312 kJ mol^{-1}

Substituting value of 'r' in the equation of T.E.

$$E = -\frac{kZe^2}{2} \times \frac{4\pi^2 Ze^2 mk}{n^2h^2} = -\frac{2\pi^2 Z^2 e^4 mk^2}{n^2h^2}$$



Thus, the total energy of an electorn in n^{th} orbit is given by

$$E = \frac{2\pi^2 Z^2 e^4 mk^2}{n^2H^2}$$

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

$$= -21.8 \times 10^{-19} \times \frac{Z^2}{n^2} \text{ J / atom}$$

$$= -313.6 \times \frac{Z^2}{n^2} \text{ Kcal / mole}$$

Relationship between P.E., K.E. & T.E.

$$\text{P.E.} = -\frac{KZe^2}{r}, \ \text{K.E.} = \frac{1}{2}\frac{KZe^2}{r}, \ \text{T.E.} = -\frac{1}{2}\frac{KZe^2}{r}$$

$$\text{T.E.} = \frac{\text{P.E.}}{2} = -\text{K.E.}$$

Previous Year’s Question

Based on equation E = – 2.178 × 10^{-18} J(Z^2/n^2), certain conclusions are written. Which of them is not correct?

[NEET-2013]

(1)) Equation can be used to calculate change in energy when electron changes orbit.

(2) For n = 1, electron has a more negative energy than it does for n = 6 which means that electron is more loosely bound

(3) Negative sign in equation simply means that energy of electron bound tot he nucleus is lower than it would be if electrons were at the infinite distance from nucleus.

(4) Larger the value of n, larger would be the orbit radius..

Calculation of the number of revolutions of the electron in an orbit per second

$$\text{Number of revolutions per sec.} = \frac{\text{velocity of the electron}}{\text{Circumference of the orbit}}$$

$$= \frac{v}{2\pi r} = \frac{nh}{2\pi mr} \times \frac{1}{2\pi r} \qquad \left[\text{On substituting the value of v from } mvr = \frac{nh}{2\pi}\right]$$

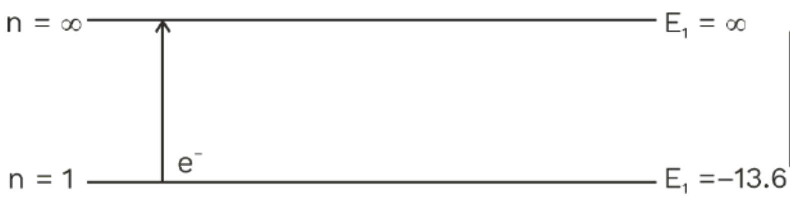
$$= \frac{nh}{4\pi^2 mr^2}$$

$$\text{No. of revolutions per second} = \frac{nh}{4\pi^2 mr^2} = \frac{nh}{4\pi^2 m} \times \left(\frac{4\pi^2 mze^2 k}{n^2h^2}\right)^2$$

$$= \frac{4\pi^2 mz^2 e^4 k^3}{n^3 h^3}$$

Ionization Energy / Ionization potential

Minimum energy required by an electron to leave the ground state:



$IE/IP = E_\infty - E_1$
$= 0 - E_1$
$= -E_1$

For Ex: H

$(IP)_H = -(13.6) = 13.6$

$\boxed{IP = -E_1}$ — H, He^+, Li^{+2}

Separation Energy

Minimum amount of energy required to escape an e^- from an excited state

$2 \longrightarrow \infty$

$3 \longrightarrow \infty$

Previous Year's Question

In hydrogen atom, energy of first excited state is −3.4eV. Then find out K.E. of same orbit of hydrogne atom.

[AIPMT]

(1) +3.4 eV (2) +6.8 eV
(3) −13.6 eV (4) +13.6 eV

Failure of Bohr Model

- The theory was very successful in predicting and accounting the energies of line spectra of hydrogen i.e. one electron system. It could not explain theline spectra of atoms containing more than one electron.
- Theory does not explain the presence of multiple spectral lines.
- Theory does not explain splitting of spectral lines in magnetic field (Zeeman effect) and in electric field (Stark effect). Intensity of spectral lines was also not explained by the Bohr atomic model.
- This theory could not explain uncertainty principle.

The Bohr-Sommerfeld Theory

For explaining fine structures of spectral lines Sommerfeld introduced two modification in Boht's theory.

- According to Sommerfeld, path of an electron around nucleus, is an ellipse with nucleus at one of the foci. Circular orbit is special case of the ellipse.
- Velocity of electron moving in an elliptical orbit varies at different parts of the orbit. This causes relavitistic variation in mass of moving electron. Therefore, he took into account relativistic variation of mass of electron with velocity. Therefore, this is known as relativistic atom model.

Limitations of the Bohr-Sommerfeld Theory

- Sommerfeld's theory was able to give an explanaiton of the fine structure of the spectral line of hydrogen atom. But he could not predict the correct of spectral lines.

Concept Ladder

Bohr-Sommerfeld theory described the atom in terms of two quantum numbers, while Bohr ahd originally used only one quantum number.

Rack your Brain

What is the success of the Sommerfeld model?

Previous Year's Question

Who modified Bohr's theory by introducing elliptical orbits for electron path?

[AIPMT]

(1) Rutherford
(2) Thomson
(3) Hund
(4) Sommerfeld

Q4 Which one is in correct about the angular momentum?

(1) $mvr = \frac{h}{2\pi}$ (2) $mvr = \frac{h}{\pi}$ (3) $mvr = \frac{3h}{2\pi}$ (4) $mvr = \frac{3h}{4\pi}$

A4 (4)

(1) $n = 1$ (2) $n = 2 \quad \frac{2h}{2\pi}$

(3) $mvr = \frac{3h}{2\pi} \quad n = 3$ (4) $mvr = \frac{3h}{4\pi} \quad n = 3/2$

Q5 For H-atom, calculate radius

A5 $Z = 1$

$$(r_n)_H = 0.529\left[\frac{n^2}{Z}\right]Å$$

$$(r_1)_H = 0.529\left[1^2\right]Å$$

$$(r_2)_H = 0.529\left[2^2\right]Å$$

$$(r_3)_H = 0.529\left[3^2\right]Å$$

$$(r_n)_H = 0.529\left[n^2\right]Å$$

$$r_n = n^2\{r_1\}$$

$$(r_3)_H = (9r_1)_H \quad r_n = 16r_1$$

$$r_5 = 25r_1$$

Q6 The P.E. of an e^- in the H^- atom is –3.02 eV. Then find out

(1) K.E. (2) T.E. (3) Orbit (4) Radius

A6 ()

$PE = -x \quad = 3.02$

$KE = \frac{+x}{2} \quad = \frac{+3.02}{2} = +.51$

$$T.E = -\frac{x}{2} = \frac{-3.02}{2} = -1.51$$

$$= \frac{-13.6Z^2}{n^2} = -1.51$$

$$\frac{-13.6}{n^2} = -1.51$$

$n^2 = 9$; $n = 3$ $\quad [r_3 = 0.529 \times \frac{9}{1} Å]$

Q7 If I.P of H atom is 'x' eV. Then calculate the required energy to exited e^- from 2^{nd} to 3^{rd} orbit :

(1) $\frac{2}{36}x$ (2) $\frac{1}{8}x$ (3) $\frac{5}{36}x$ (4) $\frac{7x}{3.6}$

A7 (3)

$I.P = -E_1 = x$

$E_1 = -x = -13.6$

$x = 13.6$

$$E_3 - E_2 = \left(\frac{-13.6}{3^2}\right) - \left(\frac{-13.6}{2^2}\right) = 13.6\left(\frac{1}{4} - \frac{1}{9}\right) = x\left(\frac{5}{36}\right)$$

HYDROGEN SPECTRUM

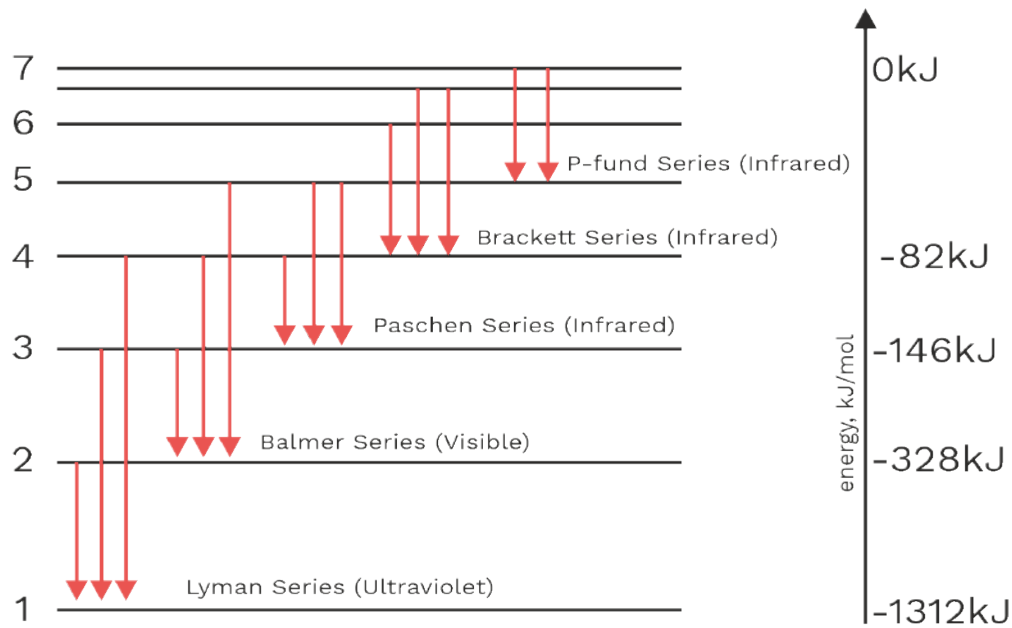
- Hydrogen spectrum is an example of atomic or line emission spectrum.
- Whenever an electric discharge is passed to hydrogen gas at low pressure, a bluem light is emitted. The light shows discontinuous line spectrum of several isolated sharp lines through prism.
- All these lines of H-spectrum have Lyman, Balmer, Paschen, Brackett, Pfund and Humphrey series.
- Wavelength of various H-lines Rydberg introduced the following expression,

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

R is a Rydberg's constant its value is 109,67800 m^{-1}.

Concept Ladder

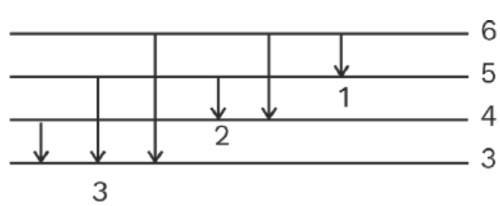
The spectral series are important in astronomy for detecting the presence of hydrogen and calculating red shifts.



Total Number of spectral Lines

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

Ex In the H- atom if an e^- moves from 6^{th} to 3^{rd} orbit by transition is multi steps. Then find
(i) Total no of spectra lines
(ii) No of lines in visible region
(iii) No of lines in UV region
(iv) No of lines is Brackett series
(v) No of lines in Infrared series
(vi) No of lines in Paschen series



Sol Lyman $\rightarrow$ O, Balmer $\rightarrow$ O (UV)
(IR) Paschen $\rightarrow$ 3 Brackett $\rightarrow$ 2 (IR)
P– Fund $\rightarrow$ 1 (IR)

Concept Ladder

Maximum number of lines produced when an electron jumps from nth level to group level $= \frac{n(n-1)}{2}$

Rack your Brain

Why the line spectra of two elements are not identical?

SERIES	7	6	5	4	3	2
Lyman	6	5	4	3	2	I (UV)
Balmer	5	4	3	2	1	0 visible
Paschen	4	3	2	1	0	Infared Ray
Brackett	3	2	1	0		
P-fund	2	1	0			
Humphry	1	0				
	(21)	**(15)**	**(10)**	**(6)**	**(3)**	**(1)**

Q8 A certain electron transition from an excited state of H-atom in one or more steps give rise of 5 lines in U.V region. Then how many lines does this transition produce in IR region?

(1) 6 (2) 7 (3) 8 (4) 9

A8 (1)

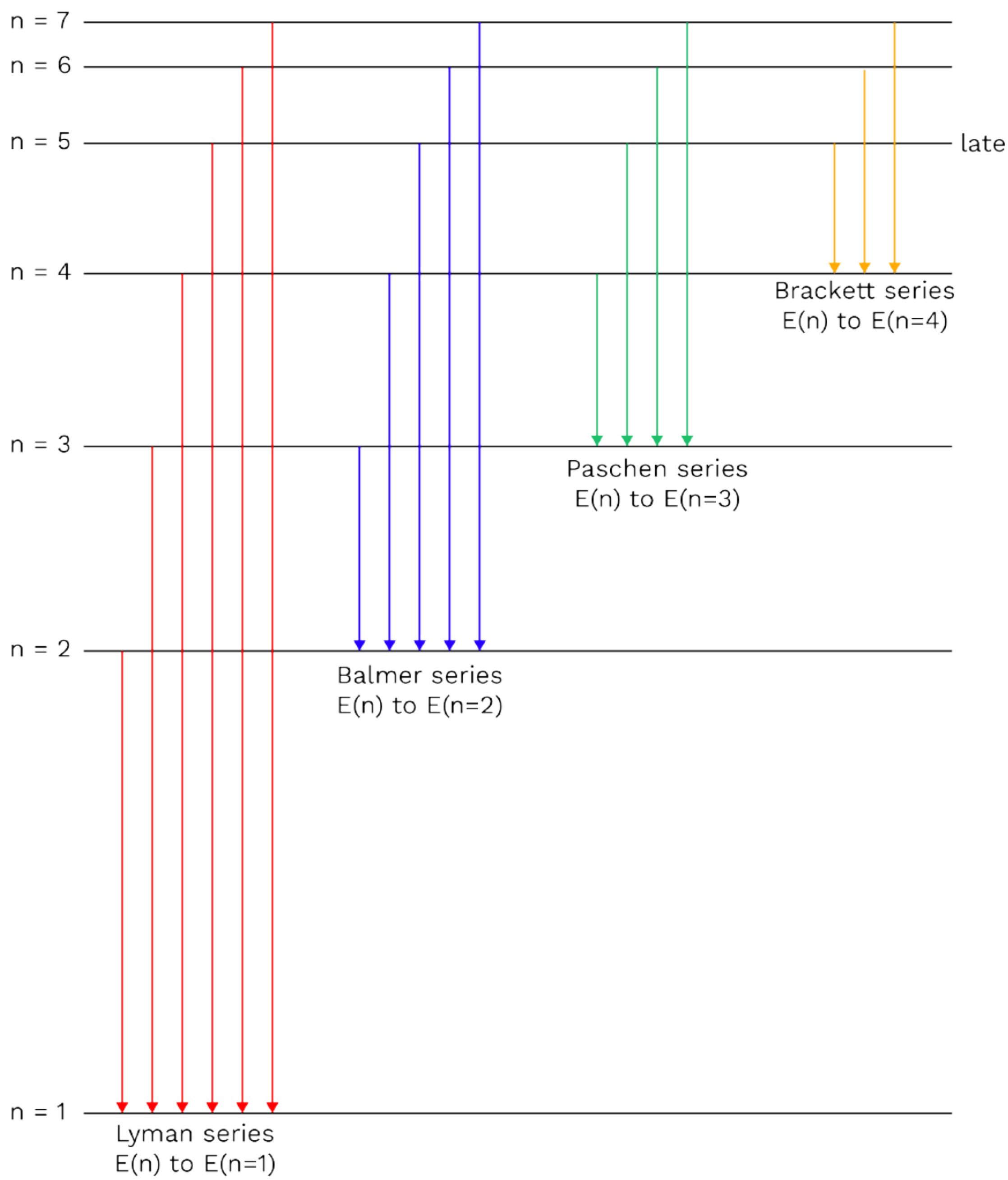
5	$\to$ U.V	4	$\to$ visible	3	$\to$ IR
2	$\to$ IR	1	$\to$ IR		
IR = $\to$ 4	UV				
3	visible	2	IR	1	IR

Q9 The e^- in a hydrogen atom transition from the Bohr orbit 5 to the orbit 2. Calculate the wavelength of photon emitted during transition. Given h = 6.6×10^{-34} J/sec.
C= 3×10^{8} m/s. $R_H = 2.18 \times 10^{-19}$ J

A9

$$R = \frac{R_H}{hc} \qquad \frac{1}{\lambda} = RZ^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

Electron transitions for the Hydrogen atom



Ex Emission Line spectrum   Any particular series Lyman series

$$\Delta E = h\upsilon = \frac{hc}{\lambda}$$

Lyman last line ($\infty \rightarrow 1$)
$\Delta E \rightarrow$ max $\upsilon \rightarrow$ max

$$* \frac{1}{\lambda} = Rz^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$n_2 \rightarrow n_1$ (λ)
R = 109678 cm^{-1}
= 1097.0 cm^{-1}

$$\frac{1}{R} = 912Å$$

Previous Year's Question

Which of the following series of transitions in the spectrum of hydrogen atoms falls in visible region: **[NEET-2019]**

(1) Brackett series
(2) Lyman series
(3) Balmer series
(4) Paschen series

Q10 Calculate the λ of 1st lined and last line of Balmer series in H-spectrum ?

A10

$$\frac{1}{\lambda} = RZ^2\left(\frac{1}{2^2} - \frac{1}{3^2}\right)$$

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

$$\frac{1}{\lambda} = \frac{5R}{36}$$

$$\lambda = \frac{36}{5} \times \left(\frac{1}{R}\right) = \frac{36}{5} \times 912Å$$

Last line

$\infty \rightarrow L$

$$\frac{1}{\lambda} = R\left\{\frac{1}{2^2} - \frac{1}{\infty^2}\right\}$$

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

$$\lambda = \frac{4}{R}$$

= 4 × 912 Å

WAVE MECHANICAL MODEL OF AN ATOM

The model consists of following :

(A) de-Broglie concept (Dual nature of Matter)

(B) Heisenberg's Uncertainty principle.

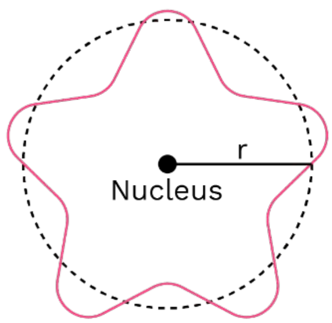
DUAL NATURE OF MATTER (WAVE NATURE OF ELECTRON)

Concept Ladder

Circumference of orbit is equal to integral multiple of wavelength (λ), i.e., $2\pi r = n\lambda$.

De Broglie

Intorduced the concept of dual nature in electron. He used Einstein's $E = mc^2$ and proposed any moving particle or object has an associated wave.



- An electron in its path is associated with a wavelength.
- The wavelength depends on the mass
- $E = \frac{hc}{\lambda}$
- $E = mc^2$
- $\lambda = \frac{h}{mc}$

1923

$\lambda \propto \frac{1}{p}$ or $\lambda = \frac{h}{p}$ (Here h = Planck's constant,

p = momentum of electron)

$\because$ Momentum (p) = Mass (m)×Velocity (v)

$\therefore$ $\lambda = \frac{h}{mv}$ We know that according to Bohr theory, $mvr = \frac{nh}{2\pi}$

or $2pr = \frac{nh}{mv}$

$\therefore$ $2pr = n\lambda$

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

Previous Year's Question

In hydrogen atom, the de Broglie wavelength of an electron in the second Bohr orbit is

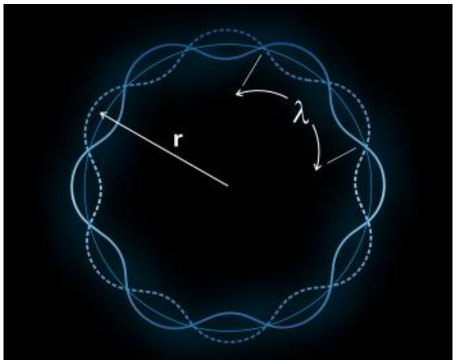
[Given that Bohr radius, a_0 = 52.9 pm]

[NEET-2019]

(1) 211.6 pm (2) 211.6 π pm

(3) 52.9 π pm (4) 105.8 π pm

From above description it is clear that according to de-Broglie there is similarity between wave theory and Bohr theory.



Heisenberg uncertainty principle

When an electron is considered to be a wave as suggested by de-Broglie, it is not possible to identify the exact position and velocity of the electron more precisely at a given instant since the wave extends throughout a region of space.

$\Delta x.Dp \geq \frac{h}{4\pi}$ or $\Delta x.\ m\Delta v \geq \frac{h}{4\pi}$ or $\Delta x.\Delta v \geq \frac{h}{4\pi m}$

or $\Delta t \times \Delta x \times \frac{\Delta p}{\Delta t} \geq \frac{h}{4\pi}$ Where h is Planck's constant.

$F \times \Delta t \times \Delta x \geq \frac{h}{4\pi}$ $\quad \Delta E \times \Delta t \geq \frac{h}{4\pi}$

Previous Year's Question

Which one is the wrong statement?

[NEET-2017]

(1) The unceratnity principle is

$$\Delta E \times \Delta t \geq \frac{h}{4\pi}$$

(2) Half filled and fully filled orbitals have greater stability due to greater exchange energy, greater symmetry and more balanced arrangement.

(3) Then energy of 2s-orvbital is less than the energy of 2p-orbital in case of hydrogen like atoms.

(4) de-Broglie's wavelength is given by $\lambda = \frac{h}{mv}$ where m = mass of particle, v = group velocity of the a particle

Rack your Brain

Why possiblity of the electron in the nucleus is zero?

Q11 A ball weight 25 g moves with a velocity of 6.6×10^4 cm/s then find out the de-Broglie λ associated with it.

A11

$$\lambda = \frac{h}{mv} = \frac{6.6 \times 10^{-34} \times 10^7 erg}{25 \times 6.6 \times 10^4 cm/s} s = 0.04 \times 10^{-31} \text{ cm} = 4 \times 10^{-33} \text{ cm}$$

Q12 If the uncertainity in position of a moving particle is 0 then find out Δp.

A12

$$\Delta x \, \Delta p \geq \frac{h}{4\pi}$$

$$\text{or } \Delta p \geq \frac{h}{4\pi \Delta x}$$

$$\text{or } \Delta p \geq \frac{h}{4\pi \times 0}$$

$$\text{or } \Delta p \geq \infty$$

WAVE MECHANICAL MODEL OF ATOM

- This model was developed by Erwin Schrodinger in 1926.
- This atomic model is based on particle and wave nature of electron is known as wave mechanical model of the atom.
- This model describes the electron as a three-dimensional wave in the electronic field of positively charged nucleus.
- Schrodinger derived an equation which described wave motion of an electron. The differential equation is :

$$\frac{d^2\Psi}{dx^2} + \frac{d^2\Psi}{dy^2} + \frac{d^2\Psi}{dz^2} + \frac{8\pi^2 m}{h^2}(E - V)\Psi = 0$$

where x, y and z are cartesian co-ordinates of the electron

m = mass of the elecron

E = total energy of electron

V = potential energy of electro

n h = planck's constna

Ψ = wave function of the electron

$$\nabla^2\Psi + \frac{8\pi^2 m(K.E.)}{h^2}\Psi = 0$$

Significance of Ψ

Wave function is regarded as the amplitude function expressed in terms of coordination x, y and z. Wave function can have negative or

Previous Year's Question

The uncertainty in momentum of an electron is 1 × 10^{-5} kg m/s. The uncertainty in its position will be

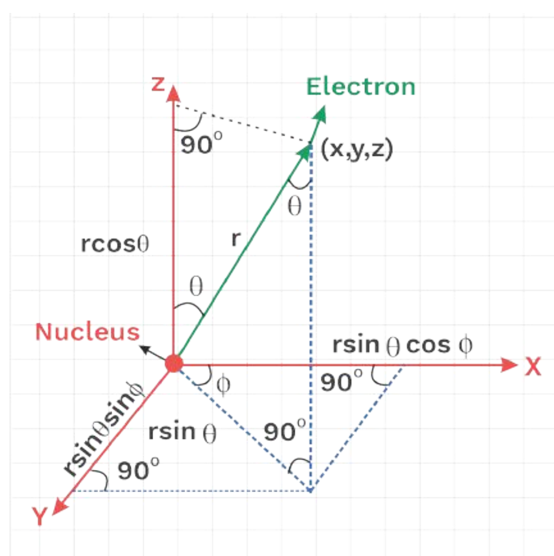
[AIPMT]

(1) 5.27 × 10^{-30} m

(2) 1.05 × 10^{-26} m

(3) 1.05 × 10^{-28} m

(4) 5.25 × 10^{-28} m



Concept Ladder

Berhaviour of electrons and other microscopic particles, a new branch of science called quantum mechanics was developed.

positive values depending upon the values of coordinates.

Significance of Ψ^2

Ψ^2 is a probability factor. It describes the probability of fi8nding an electron within a small space. The spacein which there is maximum probability of finding an electron is termed as orbital.

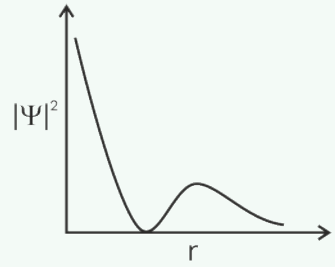
Variations of Radial Wave Function (R)

(i) Plots of radial wave function R against the distance r

The variation of the radial part of the orbital wave funcitons for 1s, 2s and 2p orbitals. The radial funcitn value changes form positive to zero then to negative. The region where this function reduces to zero is called nodal surfaces or simply nodes.

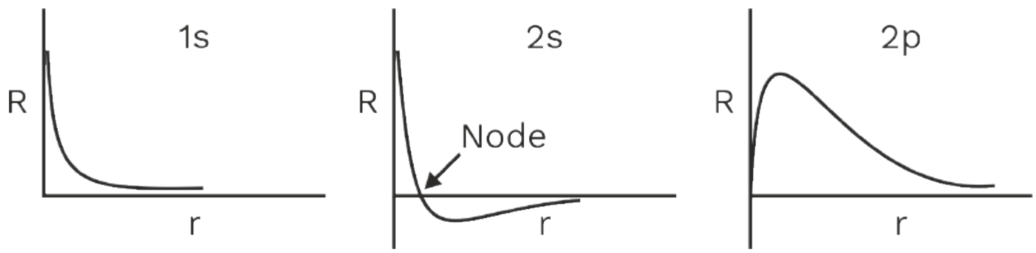
Previous Year's Question

The graph between $|\Psi|^2$ and r (radial distance) is shown below. This represents



[AIPMT]

(1) 3s orbital (2) 1s orbital
(3) 2p orbital (4) 2s orbital

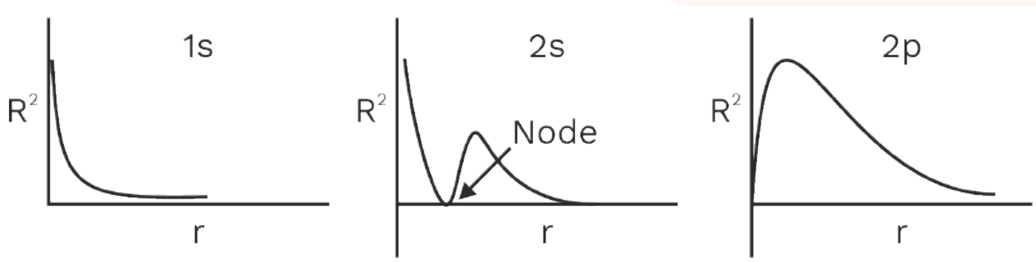


(ii) Radial probability density (R^2)

Square of radial wave function, R^2, is the measure of the probability of finding the electron in a unit volume around a particular point and is called probability density.

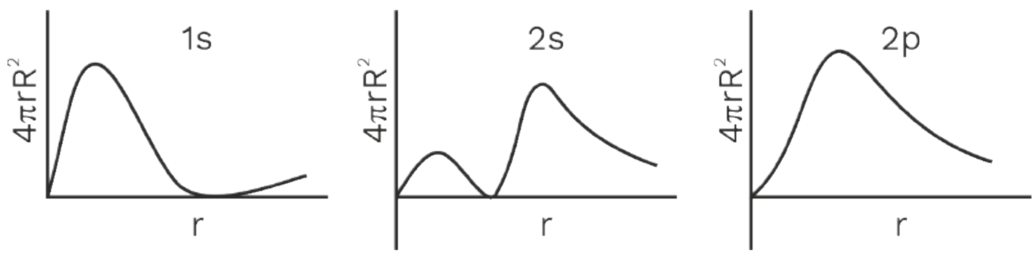
Rack your Brain

What is the significance of the wave function Ψ?



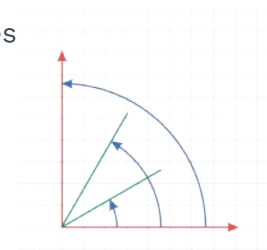
(iii) Radial distribution function, ($4\pi r^2R^2$)

Probability density of finding electron at a point at a distance r form the nucleus. Since the atoms have spherical symmetry, it is more useful to discuss the probability of finding the electron in a sphericla shell between the radii r + dr and r.



Node

Nodal point (ϕ=0)

Nodal plane	Nodal plane / Angular Node = l
Nodal surface	Nodal surface / Radial Nodes = n – l – 1
Radial Nodes	Radial Nodes →→→→→→→→
Angular Nodes	

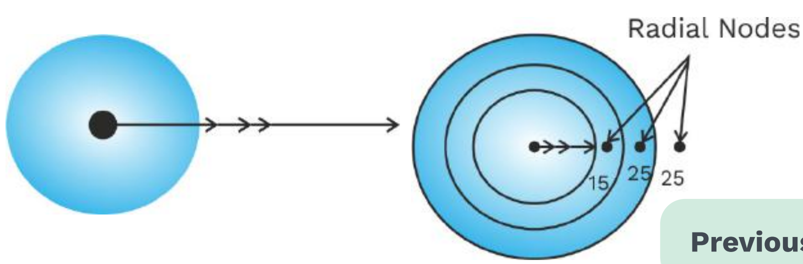
Rack your Brain

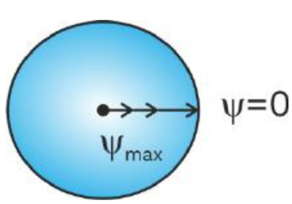
How do you find number of nodal planes in a molecular orbital?

Definition

The plane / surface in which probability of finding electron is zero is called as Node.

s- Subshell





Previous Year's Question

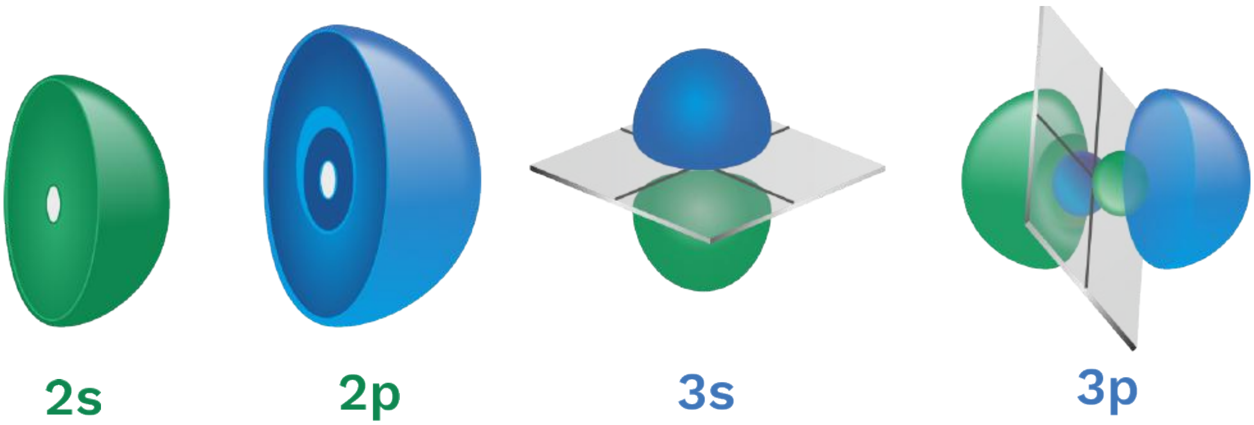
Orbital having 3 angualr nodes and 3 total nodes is

[NEET-2019]

(1) 5p (2) 3d

(3) 4f (4) 6d

Radial Nodes (n –l –1)



Angular Node l = 0

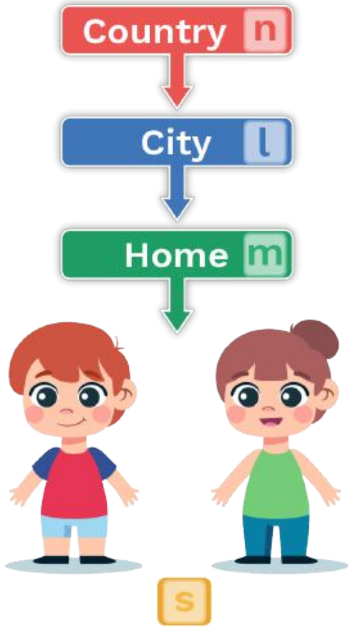
RN = n – l – 1 = n = 2

n – 2 [RN]

1s	0	$2p_x$	0
2s	1	$3p_x$	1
3s	2	$4p_x$	2

QUANTUM NUMBERS

- Quantum numbers are to specify and display to complete information about size, shape and orientation of the orbital.



Concept Ladder

The number of nodes is always one less than the principal quantum number : Nodes = n - 1. In the first electron shell, n = 1. the 1s orbital ha no nodes. The 3s, 3p and 3d orbitals have two nodes, etc.

Definition

To obtain the information about an electron Identification numbers are required. These numbers are called Quantum numbers.

Types of Quantum Numbers

(1) Principal Quantum Number (n)
(2) Azimuthal Quantum Number (l)
(3) Magnetic Quantum Number (m)
(4) Spin Quantum Number (s)

(1) Principal Quantum Number (n)

- It is the most important quantum number as it determines the size and to large extent the energy of the orbital.
- The average energy of the electron is directly proportional to the principal quantum number.
- The size of an orbital will increase an increase in the principal quantum number.
- Maximum number of electrons in a shell is given by $2n^2$.

Principal Quantum Number (n)

Represents the orbit number in an action.

It is denoted by letter ‘n’.

Exclusive identity of an e–

*n → name, size energy of shell

n = 1 K Shell
n = 2 L
n = 3 M
n = 4 N
n = 5 O
n = 6 P

Angular momentum (J/L) mvr= $\frac{nh}{2\pi}$

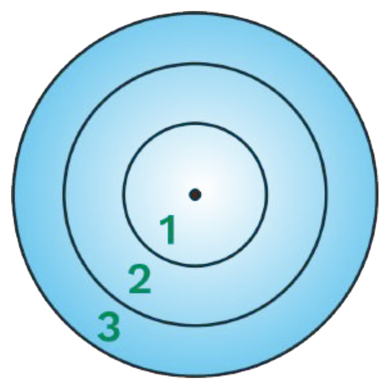
$J_3 > J_2 > J_1$

Max. No of e^- in a shell = $2n^2$

K n = 1 (2)
L n = 2 (8)
M n = 3 (18)
N n = 4 (32)

Concept Ladder

Quantum number are important because they can be determine the elctron configuration, probale location of electrons, ionization energy, atomic radius.



Previous Year’s Question

The orientation of an atomic orbital is governed by

[NEET]

(1) principal quantum number
(2) azimuthal quantum number
(3) spin quantum number
(4) magnetic quantum number

(2) Azimuthal Quantum Number (l)

- This is also known as orbital angular momentum or subsidiary quantum number.
- Azimuthal quantum number (l) gives the information about subshell in which the electron is located.

- Orbital Angular Momentum $\left\{\sqrt{\ell(\ell+1)}\,\frac{h}{2\pi}\right\}$

$\sqrt{\ell(\ell+1)}\ \hbar$

s	l = 0	→	0	
p	l = 1	→	$\sqrt{2}$	$\{\hbar\}$
d	l = 2	→	$\sqrt{6}$	$\{\hbar\}$
f	l = 3	→	$\sqrt{12}$	$\{\hbar\}$

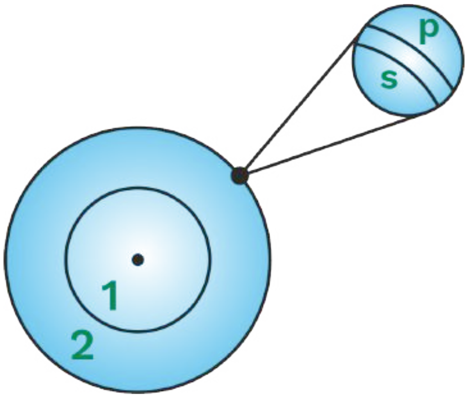
Concept Ladder

The name 'azimuthal quantum number' for l was originally introduced by Sommerfeld, who refiend Bohr's semi-classical model by replacing circular orbits with elliptic ones. The spherical orbitals were similar (in the lowest-energy state) to a rope oscillating in a large 'horizontal' circle.

Azimuthal Quantum Number (ℓ)

Represents the shape of an orbital in atom.

It is denoted by letter 'ℓ' and its value vary from 0 to 'n'.



Value of l	0	1	2	3
Sub-shell notation	s	p	d	f
Number (2l + 1) of orbitals	1	3	5	7

- Max. No. of e^- in a subshell = (4l + 2)

s	l = 0	2	for any given n, l will be from [0 to n =–1]
p	l = 1	6	
d	l = 2	10	
f	l = 3	14	

Shape of subshell

- s, p, d and f are taken from spectroscopic terms sharp, principal, diffuse, and fundamental, respectively.
- In a multi electron atom, the energy associated with an electron depends both on n and l.

n	l	Subshell notation
1	0	1s
2	0	2s
2	1	2p
3	0	3s
3	1	3p
3	2	3d
4	0	4s
4	1	4p
4	2	4d
4	3	4f

Previous Year's Question

What is the maximum numbers of electrons that can be associated with the following set of quantum numbers?

n = 3, l = 1 and m = –1

[NEET-2013]

(1) 4 (2) 2

(3) 10 (4) 6

Concept Ladder

The magnetic quantum number describes the energy levels available within a subshell and yields the projection of the orbitgal angular momentum along a specified axis.

Previous Year's Question

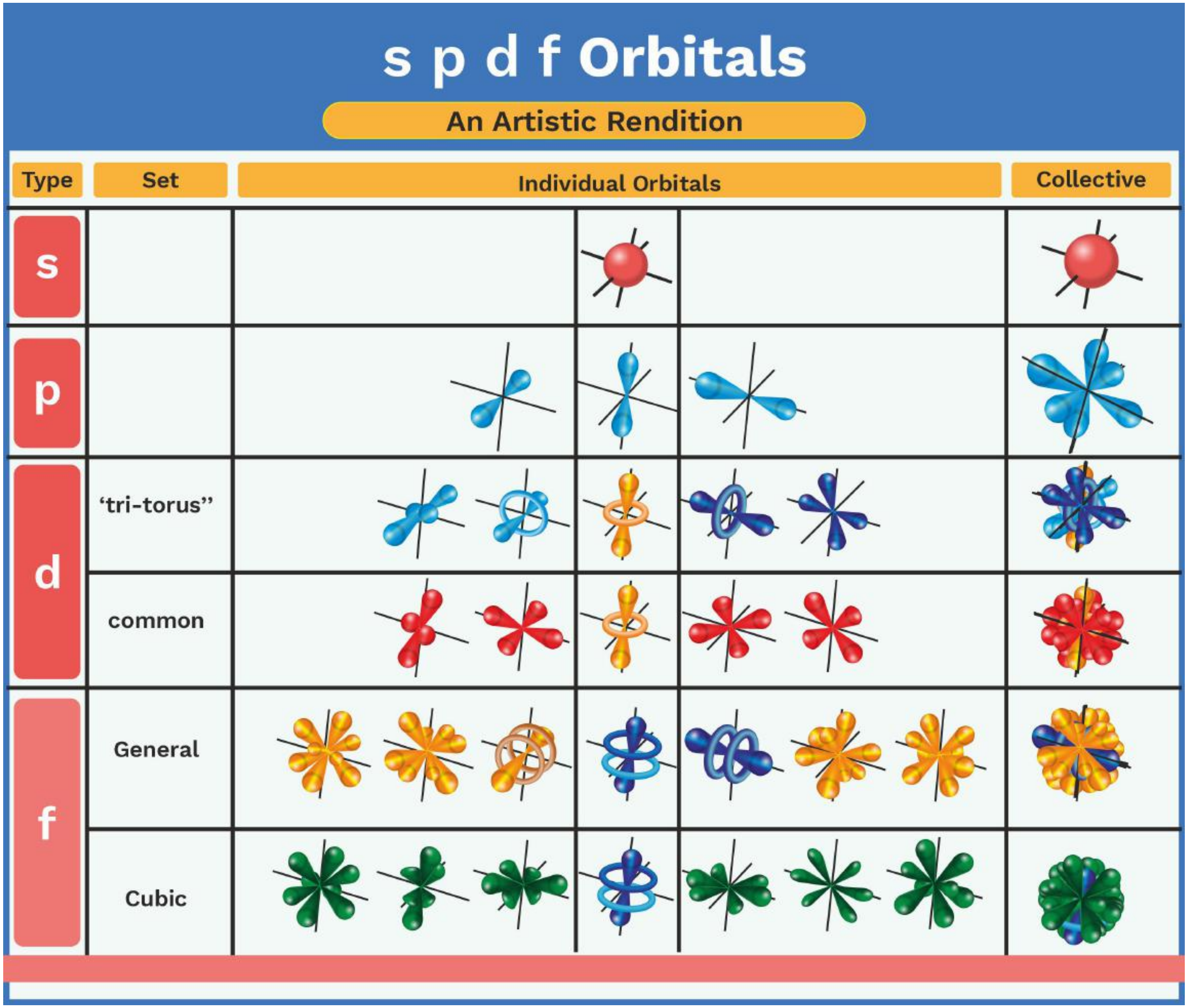
The following quantum number are possible for how many orbitals?

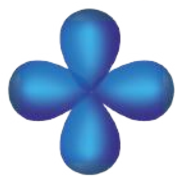
n = 3, l = 2 and m = +2

[NEET]

(1) 1 (2) 2

(3) 3 (4) 4


s p d f Orbitals
An Artistic Rendition
Type
Set
Individual Orbitals
Collective
s
p
d
'tri-torus"
common
f
General
Cubic

$l = 0$	s		Spherical
$l = 1$	p		Dumbbell
$l = 2$	d		Double dumbbell
$l = 3$	f		Complicated

(3) Magnetic Quantum Number (m)

- Magnetic quantum number was proposed by Land.
- The magnetic quantum number gives information about the spatial orientation of orbitals. These different orientations are called orbitals.
- It is denoted by m and its value depend on l values.
- The possible value of m range from $-l$ through 0 to $+l$.
- Total number of orbitals = $(2l + 1)$
- Number of obitals in a shell is n^2.

(4) Spin Quantum Number (s)

- In 1925, George Uhlenbeck and Samuel Goudsmit proposed the presence of the fourth quantum number and depicted it as the elctron spin quantum number.
- Two electrons that have different s, values $\pm\frac{1}{2}$ and both electrons have opposite spins.
- An orbital can hold maximum two electrons.
- Spin angular momentum is depicted by the symbol μ_s. The value of μ_s:

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

Concept Ladder

Zeeman effect : Splitting of lines of atomic spectrum in magnetic field.
Stark effect : Splitting of lines of atomic spectrum in electric field.

Previous Year's Question

Which of the following pairs of d-orbitals will have electron denstiy along the axis?

[NEET-2012]

(1) d_{z^2}, d_{xz} (2) d_{xz}, d_{yz}
(3) $d_{z^2}, d_{x^2-y^2}$ (4) $d_{xy}, d_{x^2-y^2}$

Previous Year's Question

The correct set of four quantum number for the valence electron of rubidium atom (Z = 37) is

[NEET-2012]

(1) 5, 1, 1, +1/2 (2) 6, 0, 0, +1/2
(3) 5, 0, 0, +1/2 (4) 5, 1, 0, +1/2

- If all the electrons in an atom or molecule are paired. They behave as dimagnetic substance. It is weakly repelled by the magnetic field.
- If atoms or molecules of a substance have one or more unpaired electrons, it behaves as a paramagnetic substance. it is weakly attracted by the magnetic field.
- Magnetic moment $= \sqrt{n(n+2)}$ BM

  Where, n = number of unpaired electrons

Concept Ladder

The spin of electrons is responsible for most of the magnetic properties of atoms, molecules, or ions. Due to their spin, electrons behave as tiny magnets.

Difference between orbit and orbital

Orbit	Orbital
1. An orbit refers to the circular path in which an electron revolves around the nucleus.	**1.** An orbital refers to the region of space having the maximum probability of finding an electron around the nucleus.
2. An orbit represents the motion of an electron arougn the nucleus in a plane.	**2.** An orbital represents the motion of an electron around nucleus in three-dimensional space.
3. An orbit (n) can accommodate a maximumn of $2n^2$ electrons.	**3.** An orbital can accommodate a maximum of two electrons.
4. Orbits are desinated as K, L, M etc. or 1, 2, 3etc., from the nucleus outwards.	**4.** Orbitals are designated as d_{xy}, d_{yz}, d_{zx}, d_{x2-y2}, d_{z2}, p_x, p_y, p_z etc.
5. Orbits are circular in shape.	**5.** Orbitals have different shapes, e.g., s orbitals are spherically symmetrical whereas p orbitals are dumbbell shaped.

QUANTUM NUMBER

THE ELECTRONS ADDRESS

n, l, m Similarly Quantum Numbers are used to identify the position of an electron in an atom.

GPS IS USED TO TRACK ANYONE AT ANY PLACE ON THE EARTH

QUANTUM NUMBER

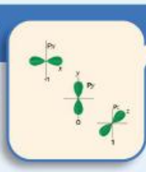
1. Principal Quantum Number (n)

Represents the orbit number in an atom. It is denoted by letter 'n'.

2. Azimuthal Quantum Number (l)

Represents the shape of an orbital in atom. it is denoted by letter 'l' and its value vary from 0 to 'n'.

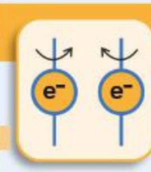
3. Magnetic Quantum Number(m_l)



Represents the orientation of an orbital in the space. it is denoted by letter 'm_l' and its value vary from 'l' to '$-l$'

x

4. Spin Quantum Number(m_s)



Represents the spin of an electron. It is denoted by m_s and each electron has an orbital either $\frac{1}{2}$ or $-\frac{1}{2}$

DIFFERENCE BETWEEN BOTH ATOMIC MODELS

Bohr's model

Is a 2-Dimensional model. Therefore he used only Principal quantum number (n) to identify the position of an electron In an atom.

Schrodinger's model

is a 3-Dimensional model. Therefore he used n. l. m. to identify the position of an electron in an atom.

Hide something at n = 1, l = 1, m_l = 0, m_s = $\frac{1}{2}$

Quantum numbers have some restrictions. It's not possible to find an electron at every possible combination of n. l. m. s. So you will never find an electron at above point.

No two electron in an atom have same Quantam Number.

n = 2 l = 1 m = 0 s = $\frac{-1}{2}$ ← | | → n = 2 l = 1 n = 0 s = $\frac{1}{2}$

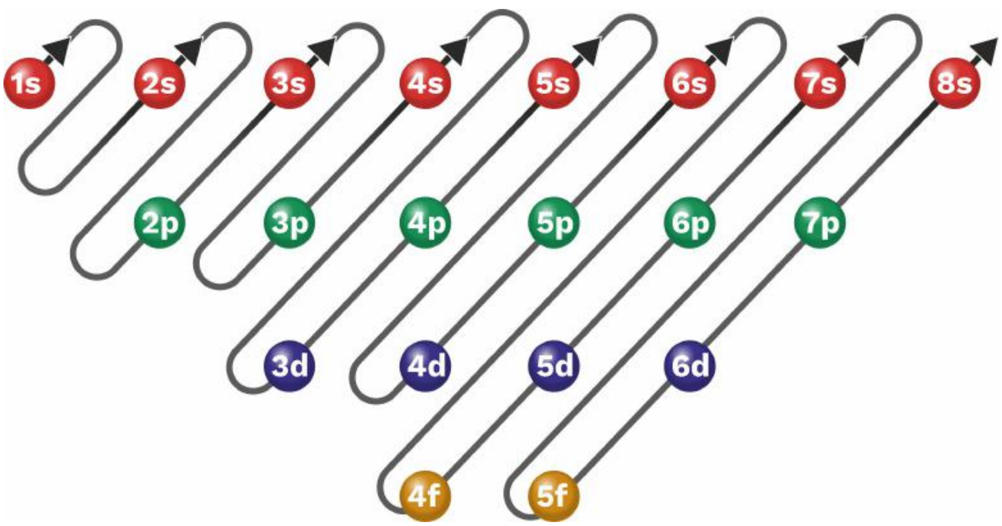
RULES FOR FILLING OF ORBITALS IN AN ATOM

(1) Aufbau Principle

The principle states that electrons are added progresively to the various orbitals in the order of increasing energies. The electrons first occupy the lowest energy orbital avaiable to them and enter into higher energy orbitals only after lower energy orbitals are filled.

Concept Ladder

Afbau principle proposed by Niels Bohr in the early 1920s, the principle was a tool for obtaining a picture of the atomic constitution, i.e., the arrangement of electrons on orbits around the nucleus.



(n +l) Rule (For multi electron species)

The subshell with lowest (n + l) value is filled up first, when two or more subshell have same (n + l) value then the subshell with lowest value of n is filled up first.

(2) Hund's Maximum Multiplicity Rule

According to the Hund's rule orbital available in the subshell are first filled singly with parallel spin electron before they begin to pair this means that pairing of electrons occurs with the introduction of second electron.

Previous Year's Question

In a given atom no two electrons can have the same valus for all the four quantum numbers. This is called **[AIPMT]**

(1) Hund's Rule
(2) Aufbau Principle
(3) Uncertainty Principle
(4) Pauli's Exclusion Principle.

(3) Pauli Exclusion Principle

According to this rule no two electron in an atom can have same values of all four quantum numbers.

Ex:

$_6C^{12} \rightarrow$ $1s^2$ $2s^2$ $2p^2$

P_x P_z P_y

n	1	2	2
l	0	0	1
m	0	0	+1, 0, –1
s	+½, –½	+½, –½	+½, +½

Concept Ladder

The Pauli exclusion principle helps explain a wide variety of physical phenomena. Electrons have to stack within an atom, i.e. have different spins while at teh same electron orbital.

ELECTRONIC CONFIGURATION OF ATOMS

In an atom, electrons are distributed among various orbitals very much in accordance with rules governing the filling of different orbitals.

Ex Notation form of Na — $1s^2 2s^2 2p^6 3s^1$

Orbital diagram from of Na

Condensed form of Na — [Ne] $3s^1$

Definition

Distribution of electrons into orbitals of an atom is called its electronic configuration.

STABILITY ORDER OF COMPLETELY FILLED AND HALF FILLED SUB-SHELLS

The ground state electronic configuration of atom corresponds to the lowest energy state and gives higher stability.

The electronic configuration of most of the atoms follows the basic rules.

Certain elements such as Cr or Cu do not follow the rules because the two sub-shells 4s and 3d slightly differ in energy, i.e., 4s is slightly lower in energy than 3d orbital. So the valence electronic configuration are $3d^5 4s^1$ and $3d^{10} 4s^1$ respectivley, and not $3d^4 4s^2$ and $3d^9 4s^2$.

The extra stability of half-filled and fully filled electronic configuration can be explained in

Previous Year's Question

If n = 6, the correct sequence for filling of electrons will be

[AIPMT-2011]

(1) $ns \rightarrow (n-2)f \rightarrow (n-1)d \rightarrow np$

(2) $ns \rightarrow (n-1)d \rightarrow (n-2)f \rightarrow np$

(3) $ns \rightarrow (n-2)f \rightarrow np \rightarrow (n-1)d$

(4) $ns \rightarrow np \rightarrow (n-1)d \rightarrow (n-2)f$

terms of symmetry and exchange energy.

Symmetrical distribution of electrons

The electronic configurations in which all the orbitals of the same sub-shell are either completely filled or half filled have relatively more symmetrical distribution of electrons. So their shielding of one another is relatively small and the electrons are more strongly attracted by the nucleus.

Ex: Cr

(less symmetrical)

(more symmetrical)

Exchage energy

Exchange means shifing of electrons from one orbital to another within same sub-shell. Energy gets released when electrons exchange their positions and the energy is called exchange energy.

For maximum number of exchanges, the maximum the energy released and the maximum the stabilisation.

Half-filled and fully-filled degenerate orbitals have more number of elctron exchanges, and consequently, they have larger exchange energy of stabilisation.

Ex Cr (For $3d^4 4s^2$)

Previous Year's Question

The electronic configuratino of Cu (Z = 29) is

[AIPMT]

(1) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9$
(2) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$
(3) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 4p^6 5s^2 5p^1$
(4) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 4p^6 3d^3$

Concept Ladder

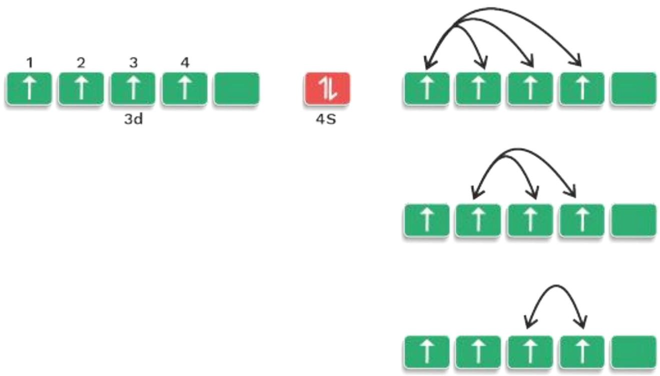
Stability of electronic configuration depends on :
(1) Half-filled and Full-filled
(2) Symmetrical distribution
(3) Exchange Energy

Previous Year's Question

The outer electronic configuratino of Gd (Z = 64) is

[NEET-2013]

(1) $4f^5 5d^4 6s^1$ (2) $4f^7 5d^1 6s^2$
(3) $4f^3 5d^5 6s^2$ (4) $4f^4 5d^5 6s^1$



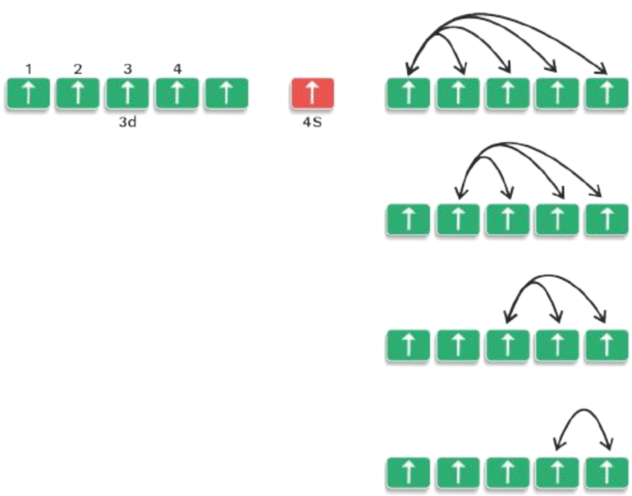
3 exchanges for electron 1

2 exchanges for electron 2

1 exchanges for electron 3

Total exchanges = 3 + 2 + 1 = 6

Cr (For $3d^5 4s^1$)



4 exchanges for electron 1

3 exchanges for electron 2

2 exchanges for electron 3

1 exchanges for electron 4

Total exchanges = 4 + 3 + 2 + 1 = 10

Q13 If the mass of e^- is assumed to be doubled the mass of proton is doubled and mass of neutron is halved, Then calculate the atomic wt. of $_8O^{16}$ and % by which it is increased?

(1) 20 amu, 25% (2) 30 amu, 25 %
(3) 40 amu, 25% (4) 45 amu, 25 %

A13 (1)

$1e^-$ = negligible, $1P^+$ = 1 amu, 1n ≈ 1 amu

$_8O^{16}$

$e^- = 8 \Rightarrow 0$ amu, $p^+ = 8 \Rightarrow 8$ amu, $n = 8 \Rightarrow 8$ amu

Total 16 amu

Double

e^- = X, p^+ = 16 amu, n = 4 amu

Total 20 amu

% increment: $\frac{4}{16} \times 100 = 25\%$

Q14 If the mass of proton is halved, mass of neutron is tripled and mass of e^- remains unchanged. Then calculate the atomic weight of $_6C^{12}$ and the % increment:

(1) 22 amu, 75% (2) 21 amu, 75%
(3) 20 amu, 50% (4) 15 amu, 60%

A14 (2)

$_6C^{12}$

$e^- = \longrightarrow$ X

$p^+ = \longrightarrow$ 6 $\longrightarrow$ 3 amu

$n = \longrightarrow$ 6 $\longrightarrow$ 18 amu

Total 12 21 amu (% increment = $\frac{9}{12} \times 100$ = 75%)

Q15 If an element has two isotopes forms 18 and 20 and the % abundance in nature are 20% and 80% respectively. Then calculate the avg. atomic wt. of element ?

(1) 21.6 (2) 19.6 (3) 20.6 (4) 22.6

A15 (2)

$X^{A^{18}}$ $X^{A^{20}}$

$$\text{Average atomic weight} = 18 \times \frac{20}{100} + 20 \times \frac{80}{100} = 19.6$$

Q16 The wave number of a beam of light is 400 cm^{-1} calculate the wavelength in terms of nanometer. Also find its frequency.

A16 $\bar{\upsilon} = 400 \text{ cm}^{1}$

$$\bar{\upsilon} = \frac{1}{\lambda} = 400 \text{ cm}^{-1}$$

$$\lambda = \frac{1}{400}\text{cm} = \frac{1}{400 \times 10^{-7}\text{nm}}$$

$$\upsilon = \frac{c}{\lambda} = \frac{3 \times 10^{8}}{\frac{1}{400} \times 10^{-2}\text{m}}$$

Q17 A radio station is broadcasting program at 10^8 frequency if the distance b/w radio station and the receiver is 3 lakh meter then, how long would it take the signal to reach the receiver. Find out the value of λ and $\bar{\upsilon}$?

(1) 3m, 0.22 m^{-1}
(2) 3m, 0.33 m^{-1}
(3) 4m, 0.33 m^{-1}
(4) 1m, 0.30 m^{-1}

A17 (2)

$$t = \frac{\{300000 \text{ m}\}}{3 \times 10^{8}\text{m/s}} \frac{1}{10^{3}} \text{S}$$

$$\upsilon = \frac{c}{\lambda}$$

$$10^{+8} = \frac{3 \times 10^{8}}{\lambda} \Rightarrow \lambda = 3$$

$$\bar{\upsilon} = \frac{1}{\lambda} = \frac{1}{3}\text{m}^{-1}$$

Q18 Radius of 1st Bohr orbit of H–atom is 0.52×10^{-8} . Calculate the radius of 1st orbit of He ?

(1) 0.16×10^{-8} cm
(2) 0.36×10^{-8} cm
(3) 0.46×10^{-8} cm
(4) 0.26×10^{-8} cm

A18 (4)

$$r_n \propto \frac{n^2}{Z}$$

$$\frac{2}{1} = \frac{(r_1)_H}{(r_1)_{He^+}} = \frac{n_1^2 / z_1}{n_2^2 / z_2}$$

$$r_{He^+} = \frac{(r_1)_H}{2} = \frac{1^2 / 1}{1^2 / 2} = 2$$

$$= \frac{0.52 \times 10^{-8}}{2} = 0.21 \times 10^{-8} \text{ cm}$$

	Orbit	Z
H	1	1
He^+	1	2

$$r \propto \frac{1}{Z}$$

Q19 I.P of H^- atom is 13.6 eV. Calculate the I.P of He^+ & Li^{+2}

A19

H I.P = $-E_1$ $E_1 = -13.6$

He^+ I.P = $-E_1$ $E_1 = -13.6 \times 4$

Li^{+2} I.P = $-E_1 = +(9 \times 13.6) = -13.6 \times 9$

I.P = $\{13.6\} \times Z^2$

Q20 A photon of energy 12.09 eV is completely absorbed by a H–atom initially in ground state find out the orbit in which electron is revolving

(1) 2 (2) 3 (3) 4 (4) 5

A20 (2)

$$-13.6 + 12.09 = -1.51 = -13.6 \frac{(1)}{n^2}$$

$n^2 = 9$

$n = 3$

Q21 An e^- of an atom having z = 5 revolves around a nucleus the energy required to excite the e^- from 3^{rd} to 4^{th} orbit will be :

(1) 12.5 (2) 13.5 (3) 14.5 (4) 16.5

A21 (4)

$$E_4 - E_3 = \left[\frac{-13.6\times 25}{4^2}\right] - \left[\frac{-13.6\times 25}{3^2}\right]\text{eV}$$

$$= 13.6 \times 25\left[\frac{1}{9}-\frac{1}{16}\right] \quad = 13.6 \times 25 \times \left(\frac{7}{16\times 9}\right)\text{eV}$$

$= 16.5$

Trick

$\rightarrow$ ( ) × Z^2 ( )
Z = 5
$3 \rightarrow 4$
10.2 1.89 0.660.31
0.66 × 25 = 16.5

Q22 Compare the velocities of e^- in the first excited state of He^+ and 2nd excited state of Li^{+2}

(1) 2 : 1 (2) 1 : 2 (3) 1 : 1 (4) 2 : 3

A22 (3)

$\rightarrow$	He^+	Li^{+2}	
Z	2	3	
n	2	3	$[V \propto \frac{Z}{n}]$
	1 = 1		

Q23 The relative abundance of two rubidium isotopes

$_{37}Rb^{85}$ $_{37}Rb^{87}$
75% 25%

(1) 75.5 (2) 86 (3) 86.5 (4) 87.5

A23 (2)

Find out average atomic wt.
85 87
75% 25%
86

Chapter Summary

- Atom is the fundamental unit of matter which is further indivisible i.e. atom can neither be created nor be destroyed.
- Atomic theory of matter was first proposed by John Dalton.
- Discovery of subatomic particles namely electron and proton. Electron is discovered through cathode ray discharge tube experiment. They consist of negatively charged particles which are known as electrons. The characteristics of cathode rays is independent of electrodes and nature of gas present in cathode tube.
- Measurement of e/m for electron. The value of e/m has been found to be 1.7588 $\times 10^{11}$ C/Kg.
- Charge on an electron is 1.602×10^{-19} C.
- Mass of the e^- can be calculated from the value of e/m and the value of e.
- The characteristics of anode rays or canal rays is dependent on nature of gas present in cathode tube. Charge to mass ratio of particle depends on the gas from which they generated.
- Atoms are made of three particles electron, protons and neutrons.
- **Rutherford's Model:** Atom of an element consist of a small positive charged nucleus situated at the centre of the atom. Electrons are distributed in different concentric circular paths around the nucleus called orbits. Atomic radius is of the order 10–10 m while nucleus is 10^{-15} m.
- Atomic no. of an element = Total no. of protons present in the nucleus.
- Protons and neutrons present in nucleus collectively called nucleons.
- Mass no. of an element = No. of protons + No. of neutrons
- **Isotopes:** They have same atomic number but different atomic weight and have same chemical properties.
- **Isobar:** The different atoms which have same atomic masses but different atomic number are called as Isobar.
- **Isotone:** When elements have same number of electrons of neutron are called as Isotones.
- **Isoelectronic:** Ion or atom or molecule or species which have the same number of electrons are called as isoelectronic species.
- **Electromagnetic Wave Radiation:** The oscillating electrical/magnetic field are electromagnetic radiations. Both electrical and magnetic field are perpendicular to each other.
- Order of wavelength in electromagnetic spectrum
- Cosmic rays < λ-rays < X-rays < Ultraviolet rays < Visible < Infrared < Micro waves < Radio waves.
- **Plank's Quantum Theory**: Some important phenomena such as interference and diffraction are generally explained by wave nature of electromagnetic radiation..

(i) Nature of emission of radiation from the surface of hot bodies (black - body radiation)
(ii) Ejection of electrons from the surface of metal happens when radiation strikes it (photoelectric effect)

- **Photoelectric Effect (P.E.E.):**

The ejection of electrons when light of certain minimum frequency called as threshold frequency is incident on a matel surface is called as photoelectric effect.
Incident energy = Work function (ϕ) + K.K.max¬
$E_i = \phi + (K.E.)_{max}$
$h\nu = h\nu_0 + (1/2)\, m_e v^2$
where m_e is mass of electron and v is the velocity associated with the ejected electron.

- **Bohr's Atomic Model:** Energy of an electron remains constant as long as it stays in same orbit called stationary Orbit. A fixed amount of energy is associated with each stationary orbit and hence it is called Energy level.

Energy of an electron:

$$T.E. = \frac{P.E.}{2} = -K.E.$$

$$T.E. = -13.6 \times \frac{Z^2}{n^2} eV / atom$$

- **Ground state (G.S.):** In any single electron species n = 1 is called ground state.
- **Excited state (E.S.):** In single electron species n > 1 is called excited state.
- For the nth shell = (n - 1) the excited state
- **Excitation energy:** Energy required to excite an electron from its ground state to any excited state is called excitation energy.
- Wave Mechanical Model of An Atom: WAVE MECHANICAL MODEL OFAN ATOM:

The Dual Nature of Matter (The Wave Nature of Electron)
De-Broglie Equation (Dual nature of matter and radiation):

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

$$\lambda = \frac{h}{mv} = \frac{h}{\sqrt{2m(K.E.)}}$$

If a charged particle Q is accelerated through potential difference V from rest then De-broglie wavelength is

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

- The circumference of the nth orbit is equal to n times the wavelength of the electron. $2\pi rn = n\lambda$.
- **Heisenberg's Uncertainity Principle:** It is impossible to obtain simultaneously both position and velocity (or momentum) of a microscopic particle with absolute accuracy.

$$\Delta x.\Delta p \geq \frac{h}{4\pi} \quad \text{or} \quad m\ \Delta x.\Delta v \geq \frac{h}{4\pi} \quad \text{or} \quad \Delta x.\Delta v \geq \frac{h}{4\pi m}$$

- **Quantum Mechanical Model:**

  The Schrodinger Equation:

$$\frac{\partial^2\Psi}{\partial x^2} + \frac{\partial^2\Psi}{\partial y^2} + \frac{\partial^2\Psi}{\partial z^2} + \frac{8\pi^2 m}{h^2} \quad \text{or} \quad \nabla^2\Psi + \frac{8\pi^2 m}{h^2}(E - V)\ \Psi = 0$$

- **Quantum Numbers:**

(i) Principal quantum number (n): Number of orbitals present in nth shell = n^2.
The maximum number of electrons which can be present in a principal energy shell is equal to $2n^2$.
(ii) Azimuthal quantum number (l):
Number of orbitals in a given subshell = 2l + l
Maximum number of electrons in particular subshell = 2 × (2l + l)
Orbital angular momentum $L = \frac{h}{2\pi}\sqrt{\ell(\ell+1)} = \hbar\sqrt{\ell(\ell+1)}$

(iii) Magnetic quantum number (m): It describes the orientations of the subshells. It can have values from –l to +l including zero, i.e., total (2l + l) values.
(iv) Spin quantum number (s): It describes the spin of the electron. It has values +1/2 and –1/2.
(+) signifies clockwise spinning and (–) signifies anticlockwise spinning.

- **Shape of The Orbitals:**

Nodal plane and Nodal surface :- The space where probability of finding an e^- is zero.
Nodal plane = l ; Nodal surface = n - l – 1

- **Stability of Completely Filled and Half-filled Subshells:**

Symmetrical distribution of electrons: Due to small shielding, the electrons are pulled closer to the nucleus and this decrease in energy leads to stability.
Exchange energy: Electrons with the same spin have a tendency to exchange their positions when they are present in the degenerate orbitals of a subshell. The energy released during this exchange is called exchange energy