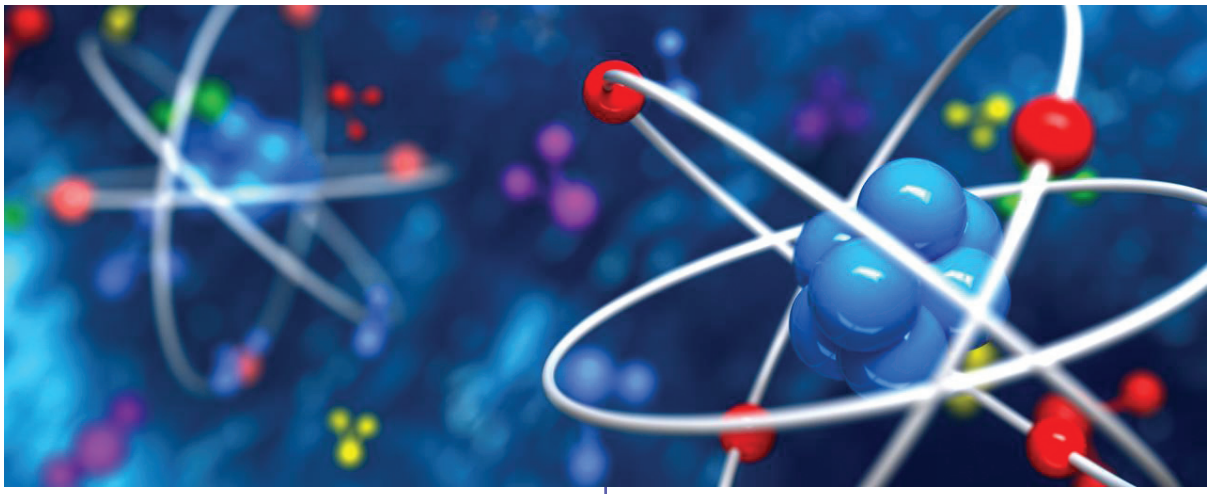
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1 STRUCTURE OF ATOM



INTRODUCTION

DALTON'S ATOMIC THEORY

* **John Dalton (1808) proposed matter to be composed of small indivisible particles called atoms, which are assumed to be structureless particles.**
* **This concept of structureless particles could not explain how atoms are different for different elements and why atoms of differnet elements combine with one another.**
* **However by the end of 19th century it was discovered that atoms consist of subatomic particles such as electrons, protons and neutrons.**
* The word atom is derived from greek word ***ATOMIO*** which means indivisible or uncutable.
* The term atom was proposed by John Dalton.
  - ➢ Matter is made up of molecules and molecules are made up of atoms.
  - ➢ Atoms are fundamental building blocks of matter.
* The matter is made of small indivisible particles called atoms, which can take part in chemical reactions.
* The atoms of the same element are identical in size, mass and in other properties in all respects.
* Atoms of different elements differ from each other in their properties and masses.
* Atoms of different elements can combine in simple ratios to form compounds.
* Atoms can be neither created nor destroyed.

Merits of Dalton's theory :

* Dalton's theory is able to explain law of conservation of mass, law of constant composition and law of multiple proportions.

Drawbacks of Dalton's theory:

* It failed to explain the internal structure of atoms
  - ➢ It could not explain how atoms of different elements differ from one another and combine with one another.
* It failed to explain the experiments like when glass or ebonite rubbed with silk or fur generate electricity.

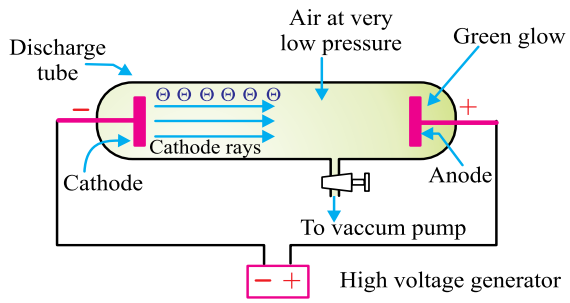
Sub atomic particles:

- ➢ Atom consists of smaller particles like electron, proton , neutron, neutrino, anti neutrino, positron antiproton, pions and measons etc...
- ➢ Electron, Proton and Neutron are considered as fundamental particles of atom.

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Electron-discovery:

* Electron was discovered by J.J.Thomson.
➢ The name eletctron was proposed by Stoney.
* The discharge tube is a sealed tube,made of glass containing two thin metal plates called electrodes.
* During the discharge tube experiment "Crookes" observed that rays were found to pass from negatively charged plate (cathode) to positively charged plate (anode).
* The electrical discharge through the gases could be observed only at very low pressures (0.01mm of Hg) and at very high voltages(10000 V).
* The rays emitted from cathode are called cathode rays.
➢ Cathode rays are stream of electrons.



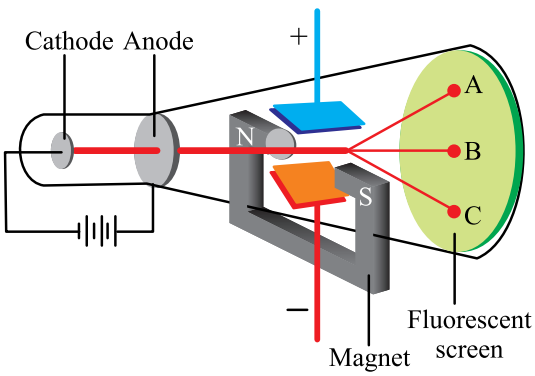
Properties of Cathode Rays :

* The cathode rays move from cathode to anode.
* These rays are invisible but their behaviour can be observed with the help of fluorescent or phosphorescent materials, which glow when hit by them.
* These rays travel in straight lines in the absence of electric and magnetic field.
* In the presence of electric field they are deflected towards anode and in magnetic field towards south pole.
* Cathode rays contain negatively charged particles called electrons.
* The nature of cathode rays is found to be independent of nature of the cathode material and nature of the gas in the tube.
* Hence electrons are considered as constituent particles of all atoms.

Charge to mass ratio of Electron $(\frac{e}{m})$:

* The charge to mass ratio is known as specific charge J.J.Thomson measured e/m ratio of the electron based on following points.
* Greater the magnitude of the charge on the particle greater is the deflection when electric and magnetic field is applied.
* Lighter the mass of the particle greater will be the deflection.
* The deflection of electrons from its original path increases when voltage increases.
* From the above points J.J.Thomson was able to determine the value of charge to mass ratio.

$$\frac{e}{m} = 1.7588\times10^{11} cKg^{-1} (or) 1.7588\times10^{8} cg^{-1}$$



➢ As the velocity of the electron increases, the specific charge of the electron decreases due to the increase of relative mass of the electron.
➢ The mass of a moving electron may be calculated by applying the formula

$$m = \frac{m_0}{\sqrt{1-\left(\frac{V}{C}\right)^2}}$$

Where m_0 = rest mass of electron
v = velocity of the electron
C = velocity of light

➢ If v = C, mass of the moving electron becomes infinity.
➢ The value of e/m of cathode rays is independent of the nature of the metal forming the cathode and the gas used in the discharge tube.
➢ This shows that electrons are universal constituents of all matter

Note : e/m value of ions w.r.t. H-atom

$$= \frac{magnitude\ of\ charge\ (relative\ charge)}{mass\ number\ (relative\ mass\ no)}$$

EXAMPLE-1: ***Find the e/m value of*** $\alpha - particle\ (He^{+2})$ ***w.r.t. H-atom?***

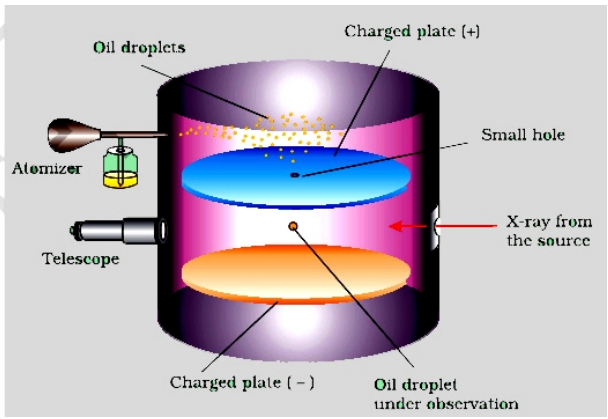
Solution: $e/m\ value\ of\ He^{+2} = \frac{2}{4} = 0.5$

Charge of Electron :

- The charge of electron is the smallest known electrical charge. It is taken as unit negative charge.
- Millikan determined the charge of the electron by an oil drop experiment.
- Charge on the oildrops was always an integral multiple of $1.60\times10^{-19}\ C$.

$$m_e = \frac{e}{e/m_e} = \frac{1.60\times10^{-19}\,C}{1.758820\times10^{11}\,ckg^{-1}} = 9.1094\times10^{-31}\,kg$$

Millikan's Oil Drop experiment



- In this , oil droplets which are in the mist form were allowed to enter through a tiny hole in the upper plate of electrical condenser.
- The downward motion of these droplets was viewed through the telescope, equipped with a micrometer eye piece.
- In chamber, the forces acting on oil drop are gravitational, electrostatic due to electrical field and a viscous drag force when the oil drop is moving.
- By measuring the rate of fall of these droplets, Millikan was able to measure the mass of oil droplets.
- The air inside the chamber was ionized by passing a beam of X-rays through it.
- The electrical charge on these oil droplets was acquired by collisions with gaseous ions.
- The fall of these charged oil droplets can be retarded, accelerated or made stationary depending upon the charge on the droplets and the polarity and strength of the voltage applied to the plate.
- By carefully measuring the effects of electrical field strength on the motion of oil droplets.
- Millikan concluded that the magnitude of electrical charge 'q' on the droplets is always an integral multiple of the electrical charge 'e' that is $q = ne$, where n = 1,2,3.....

EXAMPLE-2:

An oil drop has 8.0 × 10⁻¹⁹C charge. How many electrons does this oil drop has?

Solution :

Charge on the oil drop= $8.0\times10^{-19}\,C$

Charge on the electron= $1.6\times10^{-19}\,C$

we know that $q = ne$,

$$n = \frac{q}{e} = \frac{8\times10^{-19}}{1.6\times10^{-19}} = 5$$

EXAMPLE-3 :

In an oil drop experiment, the charges on oil drops were found as 1.5×10^{-15}, 3×10^{-15}, 4.5×10^{-15}, 6.0×10^{-15}. Calculate the magnitude of the charge on the electron.

Solution :

The magnitude of charge should be smallest & other charges should be integral multiples of that smallest charge($q = ne$). So in the problem,the smallest charge 1.5×10^{-15} and is also an integral multiple of all other charges.

case 1 : $1.5\times10^{-15} = 1\times e$

case 2 : $3\times10^{-15} = 2\times e$

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case 3 : $4.5\times10^{-15} = 3\times e$

case 4 : $6\times10^{-15} = 4\times e$

so charge on the electrons is same in all cases and it will be 1.5×10^{-15}

Discovery of Protons :

➢ Protons are discovered by Gold Stein.

* He used perforated cathode in the discharge tube and repeated J.J.Thomson experiment and observed the formation of anode rays.
* These rays also termed as positive or Canal rays.
* These are simply the positively charged gaseous ions.

➢ The name proton is suggested by Rutherford

* Properties of anode rays :

* Anode rays travel in straight line, and these are material particles.
* Anode rays are positively charged, and get deflected by external magnetic field and effect the photographic plate.
* $\frac{e}{m}$ value of these rays is smaller than that of $\frac{e}{m}$ value of electrons.
* $\frac{e}{m}$ value of anode rays depends upon nature of the gas.
* Some of the positively charged particles carry a multiple of the fundamental unti of electrical charge.
* The smallest and lightest positive ion was obtained from hydrogen is called **proton.**
* $\frac{e}{m}$ value of anode rays is maximum when the gas present in the tube is hydrogen.

$(9.58\times10^{7} C.kg^{-1}$ or $9.58\times10^{4} Cg^{-1})$

* The protons carry unit positive charge with unit mass.

➢ Rutherford obtained protons by bombarding nitrogen atomic nucleus with $\alpha - particle$.

$$^{14}_{7}N + ^{4}_{2}He \rightarrow \underset{proton}{^{1}_{1}H} + ^{17}_{8}O$$

Neutron:

* Neutrons are discovered by James Chadwick.
* When Berilium or Boron nuclides bombarded by α-particles, neutrons are formed.

➢ They have no charge but posses mass almost equal to proton.

$$\underset{\text{Beryllium}}{^{9}_{4}Be} + \underset{\alpha\text{-particle}}{^{4}_{2}He} \rightarrow \underset{\text{Carbon}}{^{12}_{6}C} + \underset{\text{Neutron}}{^{1}_{0}n}$$

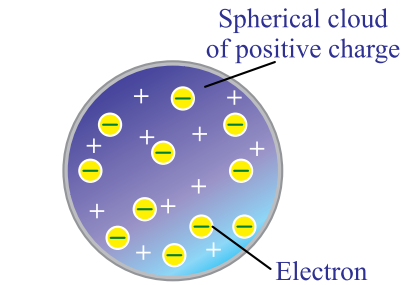
* Properties of fundamental particles

Subatomic particle	Discoveres	Mass	Relative Mass	Charge	Relative Charge	e/m
Electron	Thomson	0.000546 amu 9.18×10^{-31} Kg	$\frac{1}{1837}$	-1.602×10^{-19} C -4.8×10^{-10} esu	-1	1.76×10^{11} C/Kg
Proton	Gold Stein	1.00728 amu 1.673×10^{-27} Kg	1	1.602×10^{-19} C 4.8×10^{-10} esu	1	9.58×10^{7} C/Kg
Neutron	Chadwick	1.008665 amu 1.675×10^{-27} Kg	1	0	0	0

ATOMIC MODELS

J.J.Thomson Atomic Model:

* J.J.Thomson proposed that an atom is in spherical shape with radius approximately $10^{-10}m$ in which the positive charge is uniformly distributed.
* The electrons are embeded into it in such a manner as to give more stable electrostatic arrangment
* According to J.J.Thomson atom is like water melon.
* The positive charge is distributed like fibrous material and electrons are embeded like seeds.
* An important feature of this model is that the mass of the atom is assumed to be uniformly distributed over the atom.
* It can not explain electrical neutrality of the atom.
* This model also called plum pudding, raisin pudding model of atom.



Thomson's Plum pudding model

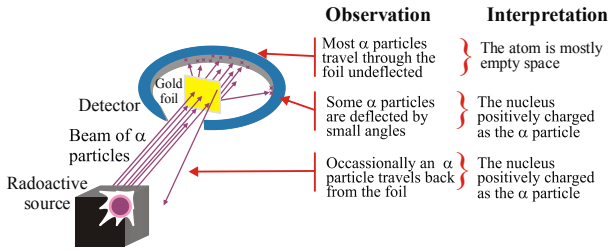
* Although this model was able to explain the overall neutrality of the atom, but was not consistent with the esults of later experiments. Thomson was awarded Nobel Prize for physics in 1906, for his theoretical and experimental investigations on the conduction of electricity by gases.

Rutherfords Model of Atom :

* Rutherford proposed atomic model based on α - ray scattering experiment.
* A beam of α - particles are passed through a thin gold foil which is enclosed with flouroscent ***ZnS*** screen capable of producing scintillations.

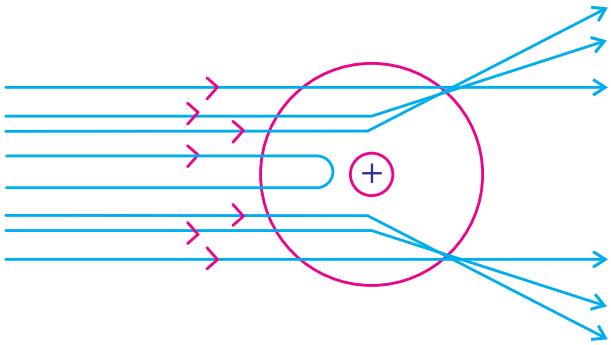
Observations - conclussions :

* Most of the α - particles passes through the foil without any deflection which indicates that most of the atom is empty.
* A small fraction of α - particles were deflected from its original path which indicates that the positive charge is at the centre of the atom and due to repulsive forces deflections takes places.



Rutherford's Gold Foil Experiment

* A very few α - particles bounced back which indicates that the whole mass is concentrated at the centre of atom. This heaviest part is named as "nucleus"



Rutherford's Alpha-ray Experiment

* On the basis of above observations and conclusions, Rutherfor d proposed the nuclear model of atom (after the discovery of protons). According to this model :

(i) The positive charge and most of the mass of the atom was densely concentrated in extremely small region. This very small portion of the atom was called nucleus by Rutherford.

(ii) The nucleus is surrounded by electrons that move around the nucleus with a very high speed in circular paths called orbits. Thus, Rutherford's model of atom resembles the solar system in which the nucleus plays the role of sun and the electrons that of revolving planets.

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(iii) Electrons and the nucleus are held together by electrostatic forces of attraction.

Drawback of Rutherford's model:

* It is against to law of electrodynamics.
* It was failed to explain stability of atom.
* According to classical electromagnetic theory, the revolving electron should loss energy continuosly and travel in a spiral path. Finally it must fall into the nucleus. But it does not happen.
* The atomic spectrum should be continuous band spectrum due to continuous loss of energybut it is a line spectrum.
* It can't explain the electronic structure of atom and energies of electrons.

*(i) According to Maxwell's electro-magnetictheory, an accelerated charge emits electromagnetic radiation and thus loses energy. Since electron is charged and is moving in circular orbit (constant acceleration), it must lose energy and spiral into the nucleus. Calculations show that the time required for collapse is 0.01 μ sec (10^{-8} sec).

*(ii) It could not explain line spectra of hydrogen. According to this model the spectra should be continuous and not discrete.

EXAMPLE-4:

1. Charge to mass ratio of cathode rays depends upon
(1) Nature of the gas in discharge tube
(2) Pressure of gas in discharge tube
(3) Nature of the material used as
(4) Velocity of electrons cathode

Solution : (4)

As velocity increases $m = \frac{m_0}{\sqrt{1-\frac{V^2}{C^2}}}$, mass increases for electron thus charge mass ratio decreases.

EXAMPLE-5 :

2. The value of charge on the oil droplets experimentally observed were -1.6×10^{-19}, -2.4×10^{-19} and -4×10^{-18}C. The value of electronic charge, indicated by these results is
(1) -1.6×10^{-19} **(2) -2.4×10^{-19}**
(3) -4×10^{-19} **(4) 0.8×10^{-19}**

Solution : (4)

The highest common factor is 0.8×10^{-19}

$\therefore$ This is the minimum charge in given conditions of experimentation and hence charge of electron

EXAMPLE-6 :

3. Which of the following statement is not correct regarding cathode rays?
(1) Cathode rays originate form the cathode
(2) Charge and mass of the particles constituting cathode rays depends upon the nature of the gas
(3) Charge and mass of the particles particles present does not depend upon the material of the cathode
(4) The ratio charge/mass of the particles is much greater than that ofanode rays.

Key : (2)

Atomic Number :

* The number of electrons or protons present in an atom of an element is called its **atomic number(Z)**.

➢ Moseley gives the relation between atomic number (Z) and frequency(ν) of the characteristic X-rays of the element by the equation

$$\sqrt{\nu} = a(Z-b)$$

Where 'a' and 'b' are constants and depends on nature of the elements.

* A neutral atom contains equal number of electrons and protons.

* **For Cation :**
Number of protons = z
Number of electrons = z – no.of electrons lost

* **For Anion :**
Number of protons = z
Number of electrons = z + no.of electrons gained

Atomic Mass number (A) :

* It is equal to sum of numbers of protons and no. of neutrons in an atom.
A = p + n

- The size of the various nuclie (r) can be calculated from radius (r) = $(1.3 \times 10^{-13})A^{1/3}$ Where A is the mass no. and r is radius of nucleus in cm.
- If nucleus is assumed to spherical, the density of nucleus (d) may be expressed as

$$d = \frac{\text{Mass of nucleus}}{\text{Volume of nucleus}} = \frac{\text{Mass no.}}{\text{Avogadro no.}} \times \frac{1}{4/3\pi r^3}$$

- The diameter of nucleus are of the order of 10^{-12} to 10^{-13} cm.
  The diameter of atom are of the order of 10^{-8} cm.
- Diameter (size) of the atom is 10^5 times the diameter of the nucleus.
- The volume of the atom is about 10^{15} times the volume of the nucleus.

* Lastly an important point to mention regarding isotopes is that *chemical properties of atoms are controlled by the number of electrons, which are determined by the number of protons in the nucleus.* Number of neutrons present in the nucleus have very little effect on the chemical properties of an element. Therefore, all the isotopes of a given element show same chemical behaviour.

* Atomic Weight :

one-twelth mass of an atom of C_6^{12} is called as atomic mass unit(a.m.u). it is also known as Dalton(Da).

EXAMPLE-7 :

What will be the difference in mass number if the number of neutrons halved and the number of electrons doubled in $^{12}_{6}C$

Solution :

Mass no is the sum of protons and neutrons

If $^{12}_{6}C \rightarrow Initial \rightarrow final$

protons :	6	- 6
Neutrons:	6	- 3
mass no :	12	- 9

Hence the decrease in mass no is 25%

EXAMPLE-8 :

Calculate the no.of protons,neutron and electrons in $^{37}_{17}Cl$

Solution :

No.of protons = Atomic number(z)=17

mass number(A)=37

No.of neutrons=A-Z=37-17=20

No.of electrons=17

EXAMPLE-9 :

Calculate the no.of protons, neutron and electron in $^{14}_{7}N^{3-}ion$

Solution:

No.of protons=atomic number(z)=7

No.of neutrons =(A-z)=14-7=7

No.of electrons in anion

$= z + magnitude\ of\ ch\arg e$ =7+3=10

EXAMPLE-10 :

The no.of electrons ,protons and neutron in a species are equal to 10,11,12 respectively. Assign proper symbol to the species.

Solution :

No.of protons=11, hence atomic no. = 11

so the element is Na.

It has one eletron less than the no.of electrons, hence it has a unit +ve charge.

No.of neutrons =12

Mass number = no.of protons + no.of neutrons

= 11 + 12 = 23

therefore the symbol of that species= $^{23}_{11}Na^{+}$

Isotopes :

* The nuclides of same element having same atomic number but different mass numbers are called isotopes.
* Isotopes show similar chemical properties but different physical and radioactive properties.
- Fractional atomic masses of elements is due to the presence of Isotopes.

Avg.atomic mass

Atomic Weight :

The atomic weight of an element is the average of mass of all the isotopes of that element.

➢ An element have three isotopes y_1, y_2 *and* y_3 and their isotopic weight are w_1, w_2, w_3 and their percentage/possibility/probability/ratio of occurance in nature are x_1, x_2, x_3 respectively, then the average atomic weight of element is

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

eg:	Cl^{35}	Cl^{37}
Probability ratio	75%	25%

$$\frac{35\times3+37\times1}{3+1} = \frac{142}{4} = 35.5 \quad \text{(or)}$$

$$= \frac{\sum \%\ abundance\ of\ isotope \times atomic\ weight}{total\ ratio}$$

eg. 1) Isotopes of hydrogen :

* Protium $\left({}_1H^1\right)$, Deuterium $\left({}_1H^2\right)$ or ${}_1D^2$ Tritium $\left({}_1H^3\right)$ or ${}_1T^3$. Protium (99.985%), Deuterium (0.015%)

* 2) Isotopes of chlorine : ${}_{17}Cl^{35}$ and ${}_{17}Cl^{37}$

* 3) isotopes of uranium : ${}^{235}_{92}U$ *and* ${}^{238}_{92}U$

EXAMPLE-11 :

The mass number of three isotopes of an element are 10,12,14 units. Their percentage abundance is 80,15 and 5units respectively. What is the atomic weight of the element?

Solution :

Percentage abundances of Isotopes=80,15,5

ratio of percentage abundances of Isotopes=16:3:1

The total ratio = 16+3+1=20

$$Avg.Awt = \frac{\sum ratio\ of\ \%\ abundance \times Atomic\ weight}{Total\ ratio}$$

$$= \frac{10\times16+12\times3+14\times1}{20} = 10.5$$

The Avg.Awt=10.5.

EXAMPLE-12 :

Naturally occuring boron consists of two isotopes whose atomic weights are 10.01 and 11.01 the atomic weight of natural boron is 10.81. Calculate the percentage of each Isotope in natural boron?

Solution :

Let the percentage of isotope with atomic weight 10.01 = x

Let the percentage of isotope with atomic weight 11.01=(100-x)

$$Avg.Awt = \frac{m_1x_1 + m_2x_2}{x_1 + x_2}$$

$$10.81 = \frac{x + 10.01 + (100 - x)11}{100}$$

x=20%

∴ The percentage of isotope with A.wt 10.01=20%

The percentage of isotope with A.wt 11.01=80.0%

Isobars :

➢ The nuclides of different elements having same mass number but different atomic numbers are called isobars.

➢ Isobars show similar physical properties but different chemical properties.

eg. 1) : ${}_6C^{14}$, ${}_7N^{14}$ 2) : ${}^{40}_{18}Ar, {}^{40}_{19}K, {}^{40}_{20}Ca$

3) : ${}^{130}_{54}Xe, {}^{130}_{56}Ba$

Isotones :

* The nuclides of the different elements with different atomic number and mass number but having same number of neutrons(A-Z) are called isotones.

➢ Isotones show different physical and chemical properties.

* **eg.**1) ${}^{30}_{14}Si, {}^{31}_{15}P, {}^{32}_{16}S$ 2) ${}^{14}_{6}C, {}^{16}_{8}O$

3) ${}^{23}_{11}Na, {}^{24}_{12}Mg$

Isodiaphers :

➢ The nuclides having same isotopic number(A-2Z) are called isodiaphers. They posses same difference of neutrons and protons (n-p).

eg. F_9^{19}, Na_{11}^{23}

Isosters :

➢ The molecules with same atomicity and same

number of electrons are called isosters.

eg. CO_2, N_2O C_6H_6 & $B_3N_3H_6$

(Atomicity = total no.of atoms in molecule)

Iso-electronic species :

➢ The molecules or ions with same number of electrons are called iso electronic species

eg1: $N^{-3}, O^{-2}, F^{-}, Ne, Na^{+}, Mg^{2+}, Al^{+3}$

eg2 : $P^{-3}, S^{-2}, Cl^{-}, Ar, K^{+}, Ca^{2+}, Sc^{+3}$

EXAMPLE-13 :

Calculate the number of electrons, protons and neutrons in (i) Phosphorus (P) atom (ii) phosphate ion. (PO_4^{3-})

Mass numbers : P = 31, O = 16

Atomic numbers : P = 15, O = 8

Solution : (1)

(i) Phosphorus atom (P)

Number of electrons = Number of protons = Atomic number = 15

Number of neutrons=Massnumber- Atomic number = 31 - 15 = 16

(ii) Phosphate ion

$PO_4^{3-} = P + 4O + 3e^{-}$

Number of electrons = $15 + 4 \times 8 + 3 = 50$

Number of protons = $15 + 4 \times 8 = 47$

Number of neutrons = $16 + 4 \times 8 = 48$

***EXAMPLE 14:* Atomic number of an element is equal to the number of**

(1) Electrons (2) Protons

(3) Neutrons (4) Electrons and protons

Key : *(2)*

***EXAMPLE 15:* The isotopes of an element differ in**

(1) The number of neutrons in the nucleus

(2) The charge on the nucleus

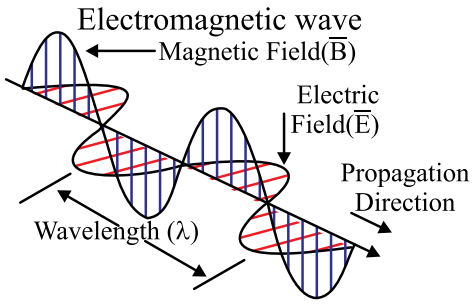
(3) The numberof extra-nuclear electrons

(4) Both the nuclear charge and the number of extranuclear electrons.

Key : (1)

Wave Nature of Electromagnetic Radiation

(i) The oscillating electric and magnetic fields produced by oscillating charged particles are perpendicular to each other and both are perpendicular to the direction of propagation of the wave.



(ii) Unlike sound waves or water waves, electromagnetic waves do not require medium and can move in vacuum.

(iii) It is now well established that there are many types of electromagnetic radiations, which differ from one nother in wavelength (or frequency). These constitute what is called electromagnetic

* **spectrum Characteristic properties of wave:**

Wave Length (λ):

* The distance between two similar points in a wave is known as wave length.

➢ The units of wave length are m, cm, A^0, nm, μ, mμ or pm.

$1A^0 = 10^{-8}\,cm = 10^{-10}\,m$

$1nm = 10^{-7}\,cm = 10^{-9}\,m = 1m\mu = 10A^0$

$1pm = 10^{-10}\,cm = 10^{-12}\,m$

Frequency (ν):

* The number of waves that pass through a given point in one second is called frequency.

* **Units :** The SI units are sec^{-1}, cycles per second (cps) or Hertz (Hz). 1cps = 1 Hz= sec^{-1}

Frequency(ν) and wavelength (λ) are related as $* \nu = \frac{C}{\lambda}$

Where C = Velocity of light

= 3 x 10^8 m/sec or 3 x 10^{10} cm/sec

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

* The number of wave lengths per unit length (or)

* The reciprocal wave length is called wave number.

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Units: cm^{-1} or m^{-1}; Wave number $\bar{v} = \frac{1}{\lambda}$

- The relation between v and $\bar{v}$ is $v = C\bar{v}$

Amplitude (a):

- The height of the crest or depth of the trough of a wave is called amplitude. **Units:** m, cm, pm
- Amplitude is a measure of the intensity or brightness of a beam of light.

Velocity (C):

* The distance travelled by a wave in one second is called its velocity. **Units:** m/sec or cm/sec.
* All types of electromagnetic radiations have the same velocity which is equal to the velocity of light $(3\times10^8 ms^{-1})$

Time Period (T) :

Time taken by a wave to pass through one point

$T = \frac{1}{\upsilon}$ second.

EXAMPLE-16 : The vividh bharati station of All india Radio, delhi,broadcasts on a frequency of 1,368 kHz (kilo hertz). Calculate the wavelength of the electromagnetic radiation emitted by transmitter. Which part of the electromagnetic spectrum does it belong to?

Solution :

$\lambda = \frac{c}{v}$ where c = velocity of E.M.R in vaccum

$= 3\times10^8 m\,\sec^{-1}$

v = *frequency* $= 1368kHz$

$= 1368\times10^3\ \sec^{-1} = \frac{3.00\times10^8 ms^{-1}}{1368\times10^3 s^{-1}} = 219.3\text{m}$

This is a characteristic radiowave wavelength.

EXAMPLE 17 :

The wavelength range of the visible spectrum extends from violet(400 nm) to red (750 nm). Express these wavelengths in frequencies (Hz). (1nm=10^{-9}m)

Solution :

Frequency of violet light

$v = \frac{c}{\lambda} = \frac{3.00\times10^8 ms^{-1}}{400\times10^{-9} m} = 7.50\times10^{14} Hz$

frequncy of red light

$v = \frac{c}{\lambda} = \frac{3.00\times10^8 ms^{-1}}{750\times10^{-9} m} = 4.00\times10^{14} Hz$

The range of visible spectrum is from 4.0×10^{14} *to* $7.5\times10^{14} Hz$ in terms of frequency

EXAMPLE-18 :

Calculate (a) Wavenumber and (b) frequency of yellow radiation having wavelength 5800A^0.

Solution :

a) Calculation of wavenumber $(\bar{v})$

$\lambda = 5800A^0 = 5800\times10^{-8} cm$

$= 5800\times10^{-10} m$

$\bar{v} = \frac{1}{\lambda} = \frac{1}{5800\times10^{-10} m}$

$= 1.724\times10^6 m^{-1} = 1.724\times10^4 cm^{-1}$

b) Calculation of the frequency (v)

$v = \frac{c}{\lambda} = \frac{3\times10^8 ms^{-1}}{5800\times10^{-10} m} = 5.172\times10^{14} s^{-1}$

***Particle Nature of Electromagnetic Radiation:**

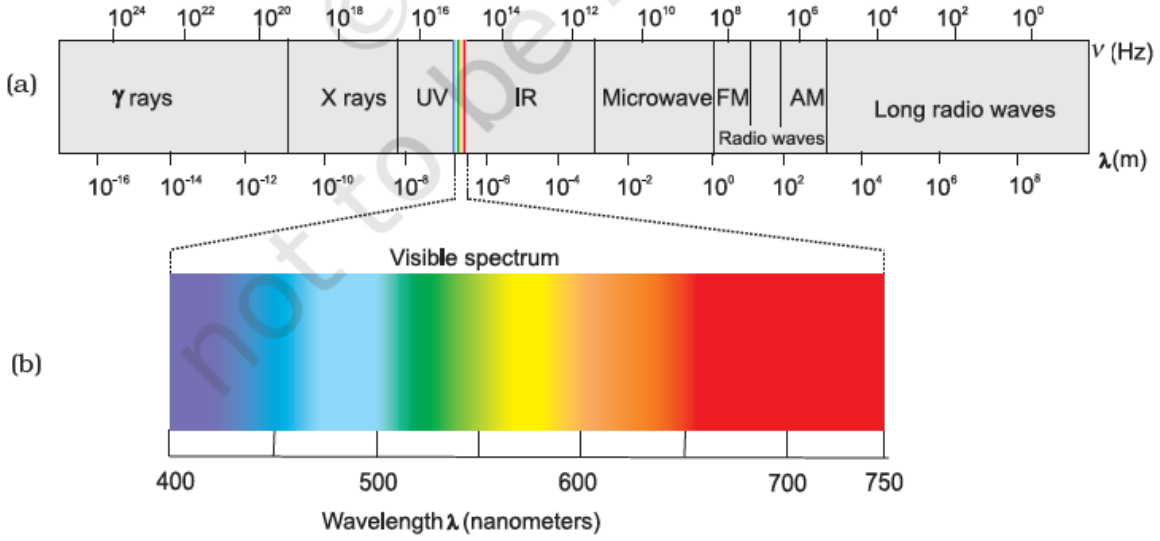
Some of the experimental phenomenon such as diffraction* and interference** can be explained by the wave nature of the electromagnetic radiation. However, following are some of the observations which could not be explained with the help of even the electromagentic theory of 19th century physics (known as classical physics):

(i) the nature of emission of radiation from hot bodies (black -body radiation)

(ii) ejection of electrons from metal surface when radiation strikes it (photoelectric effect)

(iii) variation of heat capacity of solids as a function of temperature

(iv) line spectra of atoms with special reference to hydrogen.



Corpuscular theory :

* According to Newton, light and other form of radiant energies are propagated in the form of corpuscules(simple invisible particle) .
* After the wave theory of light, the corpuscular theory lost significance

Black body radiation :

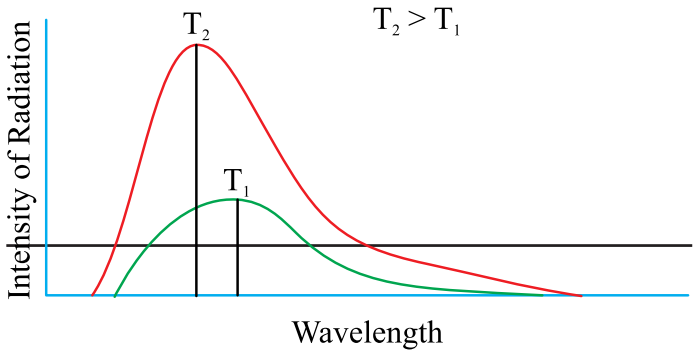
* The ideal body which emits and absorbs radiations of all frequencies, is called a black body.
* The radiation of black body is called black body radiation.

➢ A hollow sphere coated inside with platinum black, which has a small hole in its wall can act as a near black body.

eg : 1. The reddish glow from the heating element in an electric stove

eg : 2 white light emitted by the hot filament in light bulb

* A graph is obtained by plotting the intensity of radiation against wave length as shown in fig.



A Study of the curve reveals that:

* At a given temperature the intensity of radiation emitted increases, reaches a maximum value and then decreases.

* As the temperature increases the maximum intensity shifts towards the shortest wave lengths.

Planck's Quantum Theory :

* The phenomenon of the black body radiation is given by Max planck.
* The absorption or emission of radiation is mainly due to the vibrations of charged particles.
* The radiation is associated with definite amount of energy.
* The radiant energy is always in discontinuous wave form. It is absorbed or emitted discontinuosly in the form small packets of energy called quantum.
* The energy of a quantum of radiation is directly proportional to the frequency of the radiation.

$E \propto \nu$

$E = h\nu$

$$E = \frac{hc}{\lambda} = hc\bar{\nu}$$

Where

E = Energy of radiation

h = Planck's constant

$= 6.625 \times 10^{-27}$ erg – sec

$= 6.625 \times 10^{-34}$ Joule – sec (SI unit)

$= 3.99 \times 10^{-13}\ KJ$ – sec/ *mole*

* Energy is emitted or absorbed in simple integral multiples of quantum, but not fractional values

$E = nh\nu$ Where n = 1, 2, 3, 4, ...

Einstein's approach :

* According to Einstein the radiant energy is absorbed or emitted in the form of small particle is called photon. The radiation is propagated in the form of photons.
* The energy of a photon is directly proportional to frequency of the radiation.
* The energy of a photon is $E = h\nu$ *(or)* $E = \frac{hc}{\lambda}$

$$E = \frac{12375}{\lambda}$$

Where

E = energy in eV

λ = wavelength in A^0

➢ The energy of 1mole of photons is $E = Nh\nu$

Where N = Avogadro's number = 6.023×10^{23}

* Plank's quantum theory was extended by Einstein for all types of electromagnetic radiations.

Compton Effect:

➢ The increase in wave length or decrease in energy of the X - rays after scattering from an object is called the ***compton effect.***

EXAMPLE-19 :

Calculate energy of one mole of photons of radiation whose frequency is 5×10^{14} ***Hz***

Solution :

Energy (E) of one photons

$E = h\nu$

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

$\nu = 5 \times 10^{14} s^{-1}$ *(given)*

$E = (6.626 \times 10^{-34}\ Js) \times (5 \times 10^{14} s^{-1})$

$= 3.313 \times 10^{-19}\ J$

Energy of one mole of photons

$= (3.313 \times 10^{-19}\ J) \times (6.022 \times 10^{23} mol^{-1})$

$= 199.51 kJ mol^{-1}$

EXAMPLE-20 :

A 100watt bulb emits monochromatic light of wavelength 400nm. Calculate the no.of photons emitted per second by the bulb?

Solution :

Power of the bulb = 100watt = $100\ Js^{-1}$

Energy of one photon

$$E = h\nu = \frac{hc}{\lambda} = \frac{6.626 \times 10^{-34}\ Js \times 3 \times 10^{8}\ ms^{-1}}{400 \times 10^{-9}\ m}$$

$= 4.969 \times 10^{-19}\ J$

$= Number\ of\ photons\ emitted$

$$= \frac{100 Js^{-1}}{4.969 \times 10^{-19}\ J} = 2.012 \times 10^{20} s^{-1}$$

EXAMPLE-21 :

When electromagnetic radiation of wavelength 300nm falls on the surface of sodium,electrons are emitted with a kinetic energy of 1.68 × 10^5Jmol^{-1}. What is the minimum energy needed to remove an electron from sodium? What is the maximum wavelength that will cause a photoelectron to be emitted?

Solution :

The energy(E) of a 300nm photon is given by

$$h\nu = \frac{hc}{\lambda} = \frac{6.626\times10^{-34}\,Js\times3.0\times10^{8}\,ms^{-1}}{300\times10^{-9}\,m}$$

The energy of one mole of photons

$$= 6.626\times10^{-19}\,J\times6.022\times10^{23}\,mol^{-1}$$

$$= 3.99\times10^{5}\,Jmol^{-1}$$

The minimum energy needed to remove one mole of electrons from sodium.

$$= (3.99-1.68)10^{5}\,Jmol^{-1}$$

$$= 2.31\times10^{5}\,Jmol^{-1}$$

The minimum energy for one electron

$$= \frac{2.31\times10^{5}\,Jmol^{-1}}{6.022\times10^{23}\,electrons\,mol^{-1}}$$

$$= 3.84\times10^{-19}\,J$$

This corresponds to the wavelength

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

$$= \frac{6.626\times10^{-34}\,Js\times3.0\times10^{8}\,ms^{-1}}{3.84\times10^{-19}\,J}$$

=517nm this wavelength corresponds to green colour in visible spectrum.

EXAMPLE-22 :

Calculate the energy of photons of radiation whose wavelength is 5000A^0?

Solution :

$$\lambda = 5000A^{0} = 5000\times10^{-10}\,m = 5\times10^{-7}\,m$$

$$c = 3\times10^{8}\,m/\text{sec};\ h = 6.62\times10^{-34}\,Js$$

$$E = \frac{hc}{\lambda} = \frac{6.62\times10^{-34}\times3\times10^{8}}{5\times10^{-7}}$$

$$= 3.97\times10^{-19}\,J$$

EXAMPLE-23 :

What is the number of photons of light with a wave length 4000pm that provide 1J energy?

Solution:

$$wavelength\ \lambda = 4000\,pm = 4\times10^{-9}\,m$$

$$h = 6.625\times10^{-34}\,Js, c = 3.0\times10^{8}\,ms^{-1}$$

Energy of one photon, $E = \frac{hc}{\lambda}$

$$= \frac{6.625\times10^{-34}\,Js\times3\times10^{8}\,ms^{-1}}{4\times10^{-9}\,J} = 4.969\times10^{-17}\,J$$

Number of photons providing 1 joule of energy

$$= \frac{1}{4.969\times10^{-17}} = 2.01\times10^{16}$$

EXAMPLE-24 :

Calculate the energy of one mole of quanta of radiation whose frequency is 5 × 10^{10}sec^{-1}

Solution : Energy of 1 mol quantum = $Nh\nu$

$$= 6.023\times10^{23}\times6.626\times10^{-34}\times5\times10^{10}$$

$$= 19.95\,\frac{J}{mol}$$

EXAMPLE-25 :

Compare the energies of two radiations one with *l = 600nm* and other with 300nm.

Solution : $\frac{E_1}{E_2} = \frac{\lambda_2}{\lambda_1} = \frac{300}{600} = \frac{1}{2}$

The relation between two energies is $E_2 = 2E_1$

Photo electric effect :

* The phenomenon of ejection of electrons from the surface of metal when light of suitable frequency strikes is called as photo electric effect.
* The electrons ejected are called photo electrons.

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* The minimum frequency required to eject the electrons is called threshold frequency
* The photo electric effect is readily exhibited by alkali metals like K , Rb and Cs.
* A part of the energy of photon is used to escape the electron from the attractive forces and the remaining energy is used in increasing the kinetic energy of electron.

$E = W + KE$

Where W = Work function or Threshold energy

$W = h\nu_0$; $h\nu = h\nu_0 + KE$

$$K.E = h\nu - h\nu_0 = \frac{hc}{\lambda} - \frac{hc}{\lambda_0}$$

Where ν_0 = Threshold frequency

* In photo electric effect the number of photo electrons emitted is proportional to intensity of incident light.

Intensity:

➢ It is the power transfered per unit area. Its S.I. units are (W / m^2). It is used mostly for waves.

* The kinetic energy of photo electrons depends only on the frequency of incident light and not on the intensity of light.
* The minimum energy required for emission of photo electrons is called threshold energy or work function.
* For each metal, there is a characteristic minimum frequency ν_0 (also known as threshold frequency) below which photoelectric effect is not observed.
* At a frequency $\nu > \nu_0$ then photoelectric effect is observed.

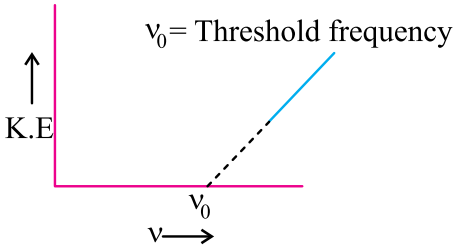
Values of work Function (W_0) for Few Metals:

Metal	Li	Na	K	Mg	Cu	Ag
W_0/eV	2.42	2.3	2.25	3.7	4.8	4.3

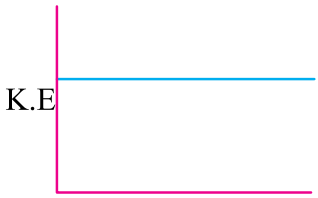
Note : Stoping potential : It is the minimum potential at which the photo electric current becomes zero. If ν_0 is the stopping potential

$eV_0 = h(\nu - \nu_0)$

* If $\nu < \nu_0$ no e^- is ejected



If $\nu = \nu_0 ; K.E = 0$



Values of work Function (W_0) for Few Metals:

Metal	Li	Na	K	Mg	Cu	Ag	Fe	Pt	W
ϕeV	2.4	2.3	2.2	3.7	4.8	4.3	4.7	6.3	4.75

EXAMPLE-26 :

The threshold frequency ν_0 for a metal is $7.0 \times 10^{14} s^{-1}$. Calculate the kinetic energy of an electron emitted when radiation of frequency $\nu = 1.0 \times 10^{15} s^{-1}$ hits the metal.

Solution :

According to Einstein's equation

$$\text{Kinetic energy} = \frac{1}{2} m_e v^2 = h(\nu - \nu_0)$$

$$= (6.626 \times 10^{-34} Js)(1.0 \times 10^{15} s^{-1} - 7.0 \times 10^{14} s^{-1})$$

$$= (6.626 \times 10^{-34} Js)(10.0 \times 10^{14} s^{-1} - 7.0 \times 10^{14} s^{-1})$$

$$= (6.626 \times 10^{-34} Js) \times (3.0 \times 10^{14} s^{-1})$$

$$= 1.988 \times 10^{-19} J$$

EXAMPLE-27 :

The minimum energy required to overcome the attractive forces between an electron and the surface of Ag metal is 5.52×10^{-19} J. What will be the maximum kinetic energy of electron ejected out from Ag which is being exposed to UV-light of $\lambda = 360 A^0$?

Solution : Energy of the photons absorbed

$$= \frac{hc}{\lambda} = \frac{6.626 \times 10^{-27} \times 3 \times 10^{10}}{360 \times 10^{-8}}$$

$$= 5.52 \times 10^{-11} erg = 5.52 \times 10^{-18} J$$

E(photon)= Work function+K.E

$K.E = 5.52\times10^{-18} - 7.52\times10^{-19}$

$= 47.68\times10^{-19}\ J$

EXAMPLE-28 :

When light of 470nm falls on the surface of potassium metal, electrons are emitted with a velocity of 6.4 × $10^4 ms^{-1}$. What is the minimum energy required to remove one mole electrons from potassium metal?

Solution: kinetic energy of photo electrons $= \frac{1}{2}mv^2$

$= \frac{1}{2}\times9.1\times10^{-31}\times(6.4\times10^4)^2 = 1.864\times10^{-21}\ J$

we know that, Absorbed energy from light = Threshold energy + kinetic energy of photoelectrons

$\frac{hc}{\lambda} = W_0 + \frac{1}{2}mv^2 = \frac{6.626\times10^{-34}\times3\times10^8}{470\times10^{-9}}$

$= W_0 + 1.864\times10^{-21}\ J = W_0 = 419.206\times10^{-21}\ J$

Energy required to eject one mole electrons

$= 419.21\times10^{-21}\times6.023\times10^{23}$

$= 252.4\times10^3\ Jmol^{-1} \quad = 252.4\ kJmol^{-1}$

EXAMPLE-29 :

When a certain metal was irradiated with light of frequency 3.2 × 10^{16}Hz,the photoelectrons emitted had twice the kinetic energy as did photoelectrons emitted when the same metal was irradiated with light of frequency 2.0 × 10^{16}Hz. Calculate v_0 for the metal?

Solution :

Applyilng photoelectric equation,

$\text{K.E} = h\nu - h\nu_0 \ or\ (\nu - \nu_0) = \frac{K.E}{h}$

$Given \quad K.E_2 = 2K.E_1$

$\nu_2 - \nu_0 = \frac{K.E_2}{h} \text{-----1} \ \&\ \nu_1 - \nu_0 = \frac{K.E_1}{h} \text{-----2}$

Dividing equation(1) with (2),

$\frac{\nu_2 - \nu_0}{\nu_1 - \nu_0} = \frac{K.E_2}{K.E_1} = \frac{2K.E_1}{K.E_1} = 2$

$or\ \nu_2 - \nu_0 = 2\nu_1 - 2\nu_0\ or\ \nu_0 = 2\nu_1 - \nu_2$

$= 2(2.0\times10^{16}) - (3.2\times10^{16}) \Longrightarrow 8.0\times10^{15}\ Hz$

EXAMPLE-30 :

Threshold wavelength of a metal is 230nm. What will be the kinetic energy of photoelectrons ejected when the metal is irradiated with wavelength 180nm?

$(h = 6.626\times10^{-34}\ J\ \text{sec})$

Solution :

Absorbed energy from light = Threshold energy +kinetic energy of photoelectrons

$\frac{hc}{\lambda} = \frac{hc}{\lambda_0} + K.E \quad \Rightarrow \quad K.E = hc[\frac{1}{\lambda} - \frac{1}{\lambda_0}]$

$= 6.626\times10^{-34}\times3\times10^8\left[\frac{1}{180\times10^{-9}} - \frac{1}{230\times10^{-9}}\right]$

$= 6.626\times10^{-34}\times3\times10^8\times10^9[\frac{1}{180} - \frac{1}{230}]$

$= 1.9878\times10^{-16}[5.55\times10^{-3} - 4.347\times10^{-3}]$

$= 2.4\times10^{-19}\ J$

SPECTRA AND ATOMIC MODELS:

* The image recorded when radiant energy is passed through a prism or grating is called a spectrum.
- The splitting of white light into seven colours is known as dispersion.
- The device used to record spectrum is called spectrograph or spectrometer.

Continuous spectrum :

* The spectrum of sun light where one colour merges into another without any gap is known as continuous spectrum.

eg: light emitted from incadescent solid

Discontinuous spectrum :

- The spectrum having distinct and well defined lines with gap (dark areas) is called discontinuous spectrum

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* Spectra is broadly divided into two types
a) Absorption spectrum b) Emission spectrum

Absorption Spectrum	Emission Spectrum
1. It is spectrum produced by transmitted lighter after absorption of radiation (excitation of electrons in a substance)	1. It is spectrum produced by emission of radiation. (de excitation of electron in a substance
2. It contains dark lines on the bright back ground.	2. It contains bright lines on the dark back ground
3. Eg: sodium vapour gives two dark lines at 5890Å and 5896Å in yellow region.	3. Eg: Sodium vapour gives two bright lines at 5890Å and 5896Å in yellow region.
4. The spectrum is discontinuous	4. The spectrummay be discontinuous.

* Emission Spectrum can be further classified into line and band spectrum (atomic and molecular spectrum).

Line Spectrum	Band Spectrum
1. The line spectrum has sharp, distinct well defined lines.	1. The band spectrum has many closed lines
2. It is the characteristic of gaseous atom and is also called atomic spectrum.	2. It is characteristic of gaseous molecules and is also called molecular spectrum.
3. It is due to transition of electrons between energy level in an atom.	3. It is due to vibration and rotational changes of atoms in a molecule.
4. It is given by inert gases, metal vapours and atoms.	4. It is given by hot metals and molecules

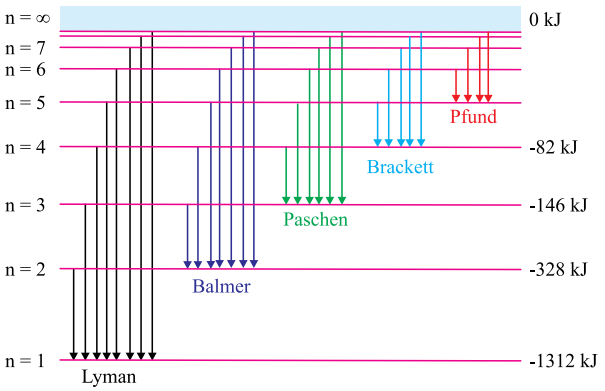
Hydrogen emission spectrum :

➢ By applying high potential difference in a discharge tube containing hydrogen gas a bright lightis obtained.

➢ This light when passed through prism ,an emerged beam of light is collected over photo graphic film. This is called H-atomic spectrum.

* It is a line spectrum and simplest of all atomic spectrum
* It contains a series of group of lines.
* They can be classified into various series.
a) Lyman series **b) Balmer series**
c) Paschen series **d) Bracket series**
e) Pfund series
* The only series visible to the naked eye is termed as the visible region of hydrogen spectrum.
* It was discovered by Balmer, so it is called Balmer series.
* **Different series of spectral lines in hydrogen emission spectrum.**

Name of the Series	n_1	n_2	Spectral region	Equation for wave number
Lyman series	1	2,3,4,5,6,7....	Ultraviolet	$\bar{v} = R \times \left(\frac{1}{1^2} - \frac{1}{n_2^2}\right)$
Balmer series	2	3,4,5,6,7....	Visible	$\bar{v} = R \times \left(\frac{1}{2^2} - \frac{1}{n_2^2}\right)$
Paschen series	3	4,5,6,7....	Near infrared	$\bar{v} = R \times \left(\frac{1}{3^2} - \frac{1}{n_2^2}\right)$
Brackett series	4	5,6,7....	Infrared	$\bar{v} = R \times \left(\frac{1}{4^2} - \frac{1}{n_2^2}\right)$
Pfund series	5	6,7....	Far Infrared	$\bar{v} = R \times \left(\frac{1}{5^2} - \frac{1}{n_2^2}\right)$



Explanation of line spectrum of Hydrogen:

* The energy gap between the two orbits is given by equation $\Delta E = E_2 - E_1$ ----------(1)
* The energy of its stationary state is given by the expression $E_n = -R_H\left(\frac{1}{n^2}\right)$ ----------(2)
* Introducing the eq (2) in equation (1)

$$\Delta E = \left(-\frac{R_H}{n_2^2}\right) - \left(-\frac{R_H}{n_1^2}\right)$$

(where n_1 and n_2 stand for initial orbit and final orbits. n_1 and n_2 may be refered as n_i and n_f)

$$\Delta E = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) 2.18\times10^{-18} J\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

* The frequency (v) associated with the absorption and emission of the photon can be evaluated by using equation

$$v = \frac{\Delta E}{h} = \frac{R_H}{h}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

On substituting R_H *and* h *values*

$$= 3.29\times10^{15}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) Hz$$

* In terms of wavenumbers $(\bar{v})$

$$\bar{v} = \frac{v}{C} = \frac{R_H}{hC}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

On substituting R_H, h *and* C *values*

$$\bar{v} = 1.09677\times10^{7}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) m^{-1}$$

➢ Rydberg constant value is not same for all the elements.
For hydrogen like species He^+, Li^{2+}, Be^{3+},
$R = 1{,}09{,}677 \times Z^2\ cm^{-1} = R_H \times Z^2$

➢ Note: $\frac{1}{R_H} = 912\ A^0$

* The first line in Balmer series is called H_α line and its wavelength is 6563 A^0.
* The second line is called H_β line and its wavelength is 4861 A^0.

➢ The spectral lines get closer when the n_2 value is increased.

* If n_2 is taken as infinity the wavelength of the limiting line(last line) in the series is obtained.

$$\bar{\nu}_\infty = \frac{1}{\lambda_\infty} = R_H\left[\frac{1}{n_1^2} - \frac{1}{\infty^2}\right] = \frac{R_H}{n_1^2}$$

* The wavelength or wave numbers of all the lines in all the series can be calculated by using Rydberg's equation.

$$\bar{v} = \frac{1}{\lambda} = R_H \times Z^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$ where R_H

= Rydberg's constant for H-atom

= 1,09,677 cm^{-1}

n_1 = Lower energy level (n_i)

n_2 = Higher energy level (n_f)

➢ Maximum number of spectral lines produced when an electron jumps from n_2 to n_1 state for a simple atom $= \frac{n(n+1)}{2}$ (or) $\sum n_2 - n_1$

➢ Number of spectral line in a series = $n_2 - n_1$

➢ Line of longest wavelength or line of shortest energy in any series of hydrogen spectrum

$$\bar{\nu}_{first} = \frac{1}{\lambda_{first}} = R_H\left[\frac{1}{n_1^2} - \frac{1}{(n_1+1)^2}\right]$$

➢ Line of shortest wavelength or line of highest energy (series limit) in any series of hydrogen spectrum $\bar{\nu}_{last} = \frac{1}{\lambda_{last}} = \frac{R}{n_1^2}$ since $n_2 = \infty$

➢ Excitation potential for

$$n_1 \rightarrow n_2 = \frac{E_{n_2} - E_{n_1}}{\text{Electronic charge}}$$

Ionisation potential for

$$n_1 \to \infty = \frac{E_{n_1}}{\text{Electronic charge}}$$

- If an electron is already present in the excited state, then the energy required to remove that electron is called separation energy

$$E_{separation} = E_{\infty} - E_{excited}$$

- The number of waves made by a Bohr electron in one complete revolution $= \frac{2\pi r}{\lambda} = \frac{2\pi r}{2\pi r} \times n = n$

EXAMPLE-31 : What are the frequency and wavelength of a photon emitted during a transition from n=5 state to the n=2state in the hydrogen atom?

Solution : Since $n_i = 5 \text{ and } n_f = 2$, this transition gives rise to a special line in the visible region of the Balmer series.

$$\Delta E = 2.18\times10^{-18} J[\frac{1}{5^2} - \frac{1}{2^2}] \quad = -4.58\times10^{-19} J$$

The frequency of the photon (taking energy in terms of magnitude) is given by

$$\nu = \frac{\Delta E}{h} = \frac{4.58\times10^{-19} J}{6.626\times10^{-34} Js} = 6.91\times10^{14} Hz$$

$$\lambda = \frac{c}{\nu} = \frac{3.0\times10^{8} ms^{-1}}{6.91\times10^{14} Hz} = 434nm$$

EXAMPLE-32 : Calculate the shortest and longest wavelength in hydrogen spectrum of Lyman series.

Solution : For Lyman series $n_1 = 1$

For shortest wavelength in lyman series(i.e,series limit), the energy difference in two states showing transition should be maximum,i.e, $n_2 = \infty$

$$\frac{1}{\lambda} = R_H[\frac{1}{1^2} - \frac{1}{(\infty)^2}] = R_H$$

$$\lambda = \frac{1}{3\times109678} = 9.117\times10^{-6} cm = 911.7A^0$$

For longest wavelength in lyman series(i.e. first line) the energy difference in two states showing transition should be minimum,i.e., $n_2 = 2$

$$\frac{1}{\lambda} = R_H[\frac{1}{1^2} - \frac{1}{(2)^2}] = \frac{3}{4}R_H$$

$$\lambda = \frac{4}{3}\frac{1}{R_H} = \frac{4}{3\times109678}$$

$$= 1215.7\times10^{-8} cm = 1215.7A^0$$

EXAMPLE-33: What transition of Li^{+2} spectrum will have same wavelength as that of second line of Balmer series in He^{+} spectrum?

Solution : $\frac{1}{\lambda_{He^+}} = \frac{1}{\lambda_{Li^+}}$

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

$$2^2\left[\frac{1}{2^2} - \frac{1}{4^2}\right] = 3^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

$$4\left[\frac{1}{4} - \frac{1}{16}\right] = 9\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right];$$

$$\frac{4}{9}\times\frac{3}{16} = \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

$$\therefore \frac{1}{n_1^2} - \frac{1}{n_2^2} = \frac{1}{12} \qquad n_1 = 3, n_2 = 6$$

$(3 \to 6)$ will have same wavelength as that of second line of Balmer series in He^{+} spectrum.

EXAMPLE-34 : Calculate the possible number of lines in the spectrum of hydrogen,when electrons return from 7th shell to 2nd shell.

Solution : $\sum(n_2 - n_1) = \sum 7 - 2 = \sum 5$

$$= 5+4+3+2+1 = 15$$

EXAMPLE - 35 : Calculate the wave number and wave length of H_α line in Bracket series of H-emission spectrum

Solution : For H_α line in Bracket series

$n_1 = 4, n_2 = 5$

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$$\bar{\nu}=R\left[\frac{1}{n_1}-\frac{1}{n_2}\right]=109677\left[\frac{1}{4^2}-\frac{1}{5^2}\right]=2467.7325\,cm^{-1}$$

$$\lambda=\frac{1}{\bar{\nu}}=4.0\times10^{-4}\,cm$$

Bohr's atomic model :

- To explain the origin of lines in hydrogen spectrum and to overcome the defects in Rutherfords model, Neils Bohr proposed his model of atom.

Postulates :

- The electron in the hydrogen atom revolves around the nucleus with definite velocity in fixed,closed, circular paths, called'orbits' or 'shells'.
- These are designed as 1,2,3,4... or K, L, M, N ..... from the side of the nucleus. [derived from Kepler's law]

* The electron revolves in the orbit whose angular momentum ($I\times\omega$) is equal to $\frac{nh}{2\pi}$.

It is given by the expression $mvr=\frac{nh}{2\pi}$

where m = mass of electron
v = velocity of electron r = radius of orbit
h = plank's constant n = 1,2,3...

- Each orbit is associated with definite amount of energy.So these are also called energy levels or energy states.

* The energy of electron does not change as long as the electron revolves in the orbit. Hence, these orbits are called 'stationary orbits' or 'stationary states'.

- The energy of electron changes only when the electron moves from one orbit to another.

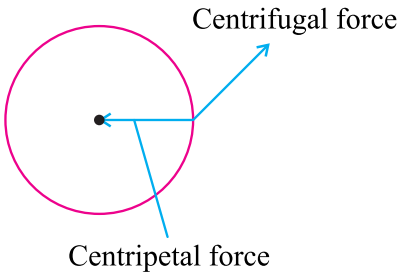
* If energy is absorbed the electron jumps from lower energy state to higher energy state, and if it lose the energy the electron jumps from higher energy state to lower energy state.

* The change in energy is given by the formula

$$\Delta E=E_2-E_1=h\nu$$

- The centripetal force on the electron is balanced by centrifugal force.

Expression for radius:



- The centripetal force of attraction between the nucleus and the electron $=\frac{-Ze^2}{r^2}$
- The centrifugal force of the electron due to revolving around the nucleus $=\frac{-mV^2}{r}$
- Expression for the radius of Bohr's orbit

$$r=\frac{n^2h^2}{4\pi^2\,m\,Ze^2}$$

- By substituting all constants, radius of orbit is

$$r=\frac{0.529\times n^2}{Z}A^0$$

For H-atom, Z = 1, hence

$$r=0.529\times n^2A^0\ ;\ r=0.529\times n^2\times10^{-8}\ cm$$

$$r=52.9\times n^2\ pm$$

Expression for energy :

- Kinetic energy of electron $=\frac{1}{2}mV^2=\frac{Ze^2}{2r}$
- Potential energy of electron $=\frac{-Ze^2}{r}$
- Total energy of electron

$$E=KE+PE=\frac{Ze^2}{2r}-\frac{Ze^2}{r}=\frac{-Ze^2}{2r}$$

- Expression for the energy of Bohr's orbit m (By substituting radius of orbit in the above expression)

$$E=\frac{-2\pi^2\,m\,Z^2e^4}{n^2\,h^2}$$

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- As we go to higher orbits, kinetic energy decreases, potential energy increases and the total energy increases.
- Energy of orbits in hydrogen atom ($Z = 1$)

$$E = \frac{-2.179 \times 10^{-11}}{n^2} \text{ ergs /atom}$$

$$= \frac{-2.179 \times 10^{-18}}{n^2} \text{ Joules /atom}$$

$$= \frac{-1312}{n^2} \text{ KJ / mole}$$

$$= \frac{-13.6}{n^2} \text{ eV / atom} \quad = \frac{-313.6}{n^2} \text{ K.cal / mole}$$

$1\text{eV} = 1.602 \times 10^{-19}$ J

- The energy of the electron in a hydrogen atom has a negative sign for all possible orbits, because the energy of the electron in the atom is lower than the energy of a free electron at rest.
- Energy of orbits for H - like species

$$E = \frac{-2.179 \times 10^{-11}}{n^2} \times Z^2 \text{ ergs} \qquad E_n = \frac{-E_1}{n^2}$$

where E_n = Energy of n^{th} orbit in hydrogen atom.

E_1 = Energy of first orbit in hydrogen atom

$n = 1, 2, 3, 4$ .....

- Similarly For Hydrogen like species

$$E_n = \frac{-E_1}{n^2} \times Z^2$$

where E_n = Energy of n^{th} orbit in other H - like species.

Z = Atomic number

E_1 = Energy of first orbit in hydrogen atom

- The ratio of PE, KE and T.E = -2 : 1 : -1

$$\frac{E_{total}}{K.E.} = \frac{(-Ze^2/2r)}{(Ze^2/2r)} = -1$$

$$\frac{E_{total}}{P.E.} = \frac{(-Ze^2/2r)}{(-Ze^2/r)} = 1/2$$

$$\frac{K.E.}{P.E.} = \frac{(Ze^2/2r)}{(-Ze^2/r)} = -1/2$$

Rydberg's constant :

$$R = \frac{2\pi^2 m Z^2 e^4}{h^3 C} = 1,09,677 \text{ x } Z^2 \text{ cm}^{-1}$$

- Difference of energy between two Bohr orbits of hydrogen atom

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

Where n_1 = lower orbit, n_2 = higher orbit

- As the value of n_2 increases, the difference of energy becomes smaller.
- The frequency of radiation absorbed (or) emitted when transitions occurs between two stationary states that differ in energy by ΔE, is given by

$$v = \frac{\Delta E}{h} = \frac{E_2 - E_1}{h}$$

This expression is commonly known as Bohr's frequency rule.

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

where Z = atomic number.

- **Calculation of the velocity of the electron in the nth orbit.**

Eliminating r from (i) and (ii) (by substituting the value of r from (ii) in (i) and rearranging), we get

For H-like particles, $v_n = \frac{2\pi Ze^2}{nh}$

On substituting the values of the constants π, e and h, we get

$$v_n = \frac{Z}{n} \times 2.188 \times 10^8 \text{ cm s}^{-1}$$

For H-atom, putting Z = 1

$$v_n = \frac{2\pi e^2}{nh} = \frac{2.188\times10^8}{n} cms^{-1}$$

Further, putting n = 1, the velocity of electron in the first orbit of H-atom will be = 2.188×10^8 cms^{-1} which is nearly 1/137th of the velocity of light.

Also, from equation (i), it may be seen that

$$v^2 = \frac{Ze^2}{mr} \text{ which shows that } v \propto \frac{1}{\sqrt{r}}$$

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

By Bohr postulate, $mvr = n\frac{h}{2\pi}$ or $v = \frac{nh}{2\pi mr}$

Number of revolutions per sec

$$= \frac{\text{Velocity of the electron}}{\text{Circumference of the orbit}} = \frac{v}{2\pi r}$$

$$= \frac{nh}{2\pi mr}\times\frac{1}{2\pi r} \text{ (On substituting the value of } v) = \frac{nh}{4\pi^2 mr^2}$$

∴ **Time taken for one revolution**

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

No. of revolutions per sec

$$= \frac{nh}{4\pi^2 m}\times\left(\frac{4\pi^2 mZe^2}{n^2h^2}\right)^2$$

$$= \frac{4\pi^2 mZ^2e^4}{n^3h^3} = -\frac{E_1}{h}\left(\frac{2}{n^3}\right)$$

where $E_1 = \frac{2\pi^2 mZ^2e^4}{h^2}$ is the energy of the first shell. Calculation of number of waves in any orbit

No. of waves in any orbit

$$= \frac{\text{Circumference of that orbit}}{\text{Wavelength}} = \frac{2\pi r}{\lambda} = \frac{2\pi r}{h/mv}$$

$$\left(\because \lambda = \frac{h}{mv} \text{ by de Broglie relation}\right)$$

$$= \frac{2\pi rmv}{h} = \frac{2\pi}{h}(mvr) = \frac{2\pi}{h}\frac{nh}{2\pi} = n$$

Ionisation potential:

- For hydrogen atom, ionization potential $= \frac{E_1}{n^2}$
- For H - like species, Ionisation potential $= \frac{E_1\times Z^2}{n^2}$.
- Ionisation potential of an atom or ion $= 13.6\left[\frac{Z^2}{n^2}\right]eV$

EXAMPLE-36 : Calculate the radius of Bohr's 3rd orbit in Li^{+2} ion.

Solution : *we know that,* $r_n = \frac{n^2}{Z}\times0.529A^0$

when n=3 and Z = 3,

$$r_3 = \frac{3^2}{3}\times0.529A^0 = 3\times0.529A^0 = 1.587A^0$$

EXAMPLE-37 : Calculate the velocity (in cm/sec) of an electron placed in the third orbit of the hydrogen atom. Also calculate the number of revolutions per second that this electron makes around the nucleus.

Solution : Radius of 3rd orbit = $3^2\times0.529\times10^{-8}$

$$= 4.761\times10^{-8} cm$$

$$mvr = n\frac{h}{2\pi} \text{ or } v = \frac{nh}{2\pi mr}$$

$$= \frac{3\times6.624\times10^{-27}}{2\times3.14\times(9.108\times10^{-28})(4.761\times10^{-8})}$$

$$= 0.729\times10^8 cm/\text{sec}$$

$$\text{Time taken for one revolution} = \frac{2\pi r}{v}$$

Number of revolutions per second

$$= \frac{1}{\frac{2\pi r}{v}} = \frac{v}{2\pi r} = \frac{0.729\times10^8}{2\times3.14\times4.761\times10^{-8}}$$

$$= 2.4\times10^{14} revolution/\text{sec}$$

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EXAMPLE 38 : Calculate the energy associated with the first orbit of He^+. What is the radius of this orbit?

Solution : $E_n = -\frac{(2.18\times10^{-18}J)Z^2}{n^2} atom^{-1}$

$for\ He^+, n=1, Z=2$

$$\frac{(2.18\times10^{-18}J)2^2}{1^2} = -8.72\times10^{-18}J$$

The radius of the orbit is given by equation

$$r_n = \frac{(0.0529nm)n^2}{Z} \quad \text{since n=1,Z=2}$$

$$r_n = \frac{(0.0529nm)1^2}{2} = 0.02645nm$$

EXAMPLE-39 :

The ionisation energy of hydrogen atom is 13.6eV. What will be the ionisation energy of He^+ and Li^{+2} ions?

Solution:

Ionisation energy = -(energy of the 1st orbit)

Energy of the 1st orbit of hydrogen = -13.6eV

Energy of the 1st orbit of $He^+ = -13.6\times Z^2$

(Z for $He^+ = 2$)

$= -13.6\times eV = -54.4eV$

so, Ionisation energy of

$He^+ = -(-54.4) = 54.4eV$

$Energy\ of\ 1st\ orbit\ of\ Li^{2+} = -13.6\times9$

$= -122.4eV$

$Ionisation\ energy\ of\ Li^{2+} = -(-122.4)$

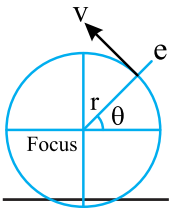
$= 122.4eV$

DRAWBACKS OF THE BOHR'S MODEL:

- It failed to explain the line spectra of atoms or ions having more than one electron.
- * It fails to account the fine spectra details (doublet, that is two closely spaced lines) of the hydrogen atom.
- * It failed to explain Zeeman effect and Stark effect.
- * The splitting of spectral lines of an atom into a group of fine lines under the influence of magnetic field is called ***Zeeman effect.***
- * The splitting of spectral lines of an atom into group of fine lines under the influence of an electric field is called ***Stark effect.***
- Bohr model of the hydrogen atom, not only ignores dual behaviour of matter but also contradicts Heisenberg uncertainity principle.

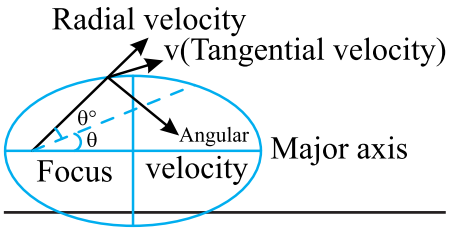
SOMMERFELD EXTENSION OF THE BOHR'S MODEL

- According to sommerfeld electron revolve around the nucleus in the Elliptical Orbits.
- Circular orbit is a special case of elliptical orbit when the length of major axis becomes equal to the length of minor axis then the shape of orbit will be circular



r = Constant

θ = Varible



r = Variable

θ = variable

- If electrons revolve in elliptical orbit then its angular momentum shows two components

1. Radial component

$J_r = \frac{n_r h}{2\pi}$ (where n_r=radial quantum number.)

$[n_r = (n-1).....0]$ **n = Shell number**

2. Azimuthal Components

$J_\phi = \frac{n_\phi h}{2\pi}$ (n_ϕ=Azimuthal quantum number.)

[n_ϕ=1,2,3,4,....n] n=shell number

So total Angular momentum $(J_n) = J_r + J_\phi$

$$\frac{nh}{2\pi} = \frac{n_r h}{2\pi} + \frac{n_\phi h}{2\pi} \qquad n = n_r + n_\phi$$

where n = principal quantum number

Ex. Let n =4 then $n = n_r + n_\phi$

$4 = 3 + 1$ $\quad$ $4 = 2 + 2$

$4 = 1 + 3$ $\quad$ $4 = 0 + 4$

The length of major axis represent by $n_r + n_\phi$ i.e, n and length of minor axis represent by n_ϕ

The path of electron

$$K = \frac{n_r + n_\phi}{n_\phi} = \frac{n}{n_\phi} = \frac{\text{Length of major axis}}{\text{Length of minor axis}}$$

If n = 4 then n_ϕ = 1 ,2, 3, 4

K= $\underbrace{\frac{4}{1}, \frac{4}{2}, \frac{4}{3}}_{\text{3 Elliptical path}}$, $\underbrace{\frac{4}{4}}_{\text{1 circular path}}$

If n = 5 then n_ϕ =1, 2, 3, 4, 5

K=

$\underbrace{\frac{5}{1}, \frac{5}{2}, \frac{5}{3}, \frac{5}{4}}_{\text{4 Elliptical path}}$, $\underbrace{\frac{5}{5}}_{\text{1 circular path}}$

If n = 1,then Elliptical path
= (n – 1) = (1 -1) =0

Circulars path = 1

In nth orbit : Number of elliptical path = (n – 1)

Number of circular path = 1

Note : Electron in the first orbit always revolve in circular path.

DUAL NATURE OF MATTER AND LIGHT

de-Broglie's wave theory :

* The wave nature of electron was first proposed by de-Broglie.
* According to de-Broglie theory all moving particles have wave properties.

➢ Wave properties are important only for particles of small mass and high velocity.

From Planck's quantum theory $E = h\nu = \frac{hc}{\lambda}$ .....(1)

Einstein's mass energy relationship is $E = mC^2$ .....(2)

from equation (1) & (2),

$$\frac{hc}{\lambda} = mc^2 \; or \; \frac{h}{\lambda} = mc \; or \; \lambda = \frac{h}{mc}$$

where 'c' is the velocity of light.

* If the velocity of micro particle is 'v' the de-Broglie's equation is $\lambda = \frac{h}{mv} = \frac{h}{p}$

where; λ = wave length

h = Planck's constant = $6.625 \times 10^{-34} J.sec$

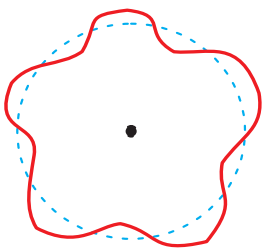
ν = Velocity of the particle

$m\nu$ = p = Momentum of the particle

de-Broglie's concept and Bohr's theory :

➢ Two types of waves are possible for an electron moving around the nucleus in the circular path

A) A standing or stationary or non-energy radiating wave:



i.e., $2\pi r = n\lambda$

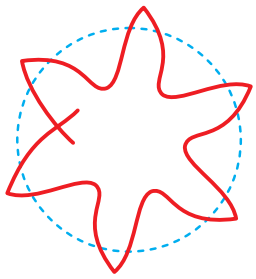
$\lambda = \frac{2\pi r}{n}$ ($\therefore$ n= integer or whole number)

we know that, $\lambda = \frac{h}{mv} \Rightarrow \frac{2\pi r}{n} = \frac{h}{mv}$

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

Hence de-Broglie's theory and Bohr's theory are in agreement with each other.

b) Non stationary or energy radiating wave.



In this $2\pi r \neq n\lambda$ hence, such an orbit cannot exist.

de-Broglie's applications :

- Number of waves in an orbit = n
- Number of revolutions of an electron per second in an orbit $= \frac{\text{Velocity of electron}}{\text{circumference}}$
- Relation between kinetic energy and wave length of a moving particle

$$\lambda = \frac{h}{\sqrt{2KE \times m}}$$

- Let a charged particle like electron be accelerated with a potential of V, then

$$KE = eV \quad \therefore \lambda = \frac{h}{\sqrt{2eVm}}$$

- For an electron $\lambda = \frac{12.27}{\sqrt{V}} A^0$
- The de Broglie wavelength for an electron in a given orbit = $3.33 \times n$ A^0

EXAMPLE-40 :

What will be the wavelength of a ball of mass 0.1kg moving with a velocity of 10ms^{-1}

Solution :

According to de-Broglie equation

$$\lambda = \frac{h}{mv} = \frac{6.626 \times 10^{-34} Js}{(0.1kg)(10ms^{-1})}$$

$$= 6.626 \times 10^{-34} m (J = kgm^2 s^{-2})$$

EXAMPLE-41 :

The mass of an electron is 9.1 × 10^{-25}J, If its K.E. is 3.0 × 10^{-25}J, calculate its wavelength.

Solution : since K.E $= \frac{1}{2}mv^2$

$$v = \left(\frac{2K.E}{m}\right) = \left(\frac{2 \times 3.0 \times 10^{-25} kgm^2 s^{-2}}{9.1 \times 10^{-31}}\right)^{\frac{1}{2}} = 812 ms^{-1}$$

$$\lambda = \frac{h}{mv} = \frac{6.626 \times 10^{-34} Js}{(9.1 \times 10^{-31} kg)(812 ms^{-1})}$$

$$= 8967 \times 10^{-10} m = 896.7 nm$$

EXAMPLE-42 :

Calculate the mass of a photon with wavelength 3.6 A^0

Solution : $\lambda = 3.6A^0 = 3.6 \times 10^{-10} m$

Velocity of photon= velocity of light

$$m = \frac{h}{\lambda v} = \frac{6.626 \times 10^{-34} Js}{(3.6 \times 10^{-10} m)(3 \times 10^8 ms^{-1})}$$

$$= 6.135 \times 10^{-33} kg$$

EXAMPLE-43 :

Calculate the de- Broglie wavelength of an electron travelling at 1% of the speed of light.

Solution :

de-Broglie wavelength can be calculated as,

$\lambda = \frac{h}{mv}$ .....(1) where, $h = 6.626 \times 10^{-34} Js$

$m = 9.1 \times 10^{-31} kg$

$v = \frac{1}{100} \times 3 \times 10^8 = 3 \times 10^6 m/\text{sec}$

substituting these values in eq(1), we get

$$\lambda = \frac{6.626 \times 10^{-34}}{9.1 \times 10^{-31} \times 3 \times 10^6}$$

$$= 242.7 \times 10^{-12} m = 242.7 pm$$

EXAMPLE-44 :

Two particles A and B are in motion. If the wavelength associated with particle A is 5×10^{-8}m, calculate the wavelength associated with particle B if its momentum is half of A.

Solution : we know , $\lambda = \frac{h}{p}$

$$\frac{\lambda_A}{\lambda_B} = \frac{p_B}{p_A} ------1 \quad \lambda_A = 5\times10^{-8} m,$$

$$p_B = \frac{1}{2} p_A, i.e, p_A = 2p_B$$

$$\therefore from\, eqn(i) \quad \frac{5\times10^{-8}}{\lambda_B} = \frac{p_B}{2p_B}$$

$$\lambda_B = 10^{-7} m$$

EXAMPLE 45 :

An electron beam emerges from an accelerator with kinetic energy 100eV. what is its de- Broglie wavelength?

$[m = 9.1\times10^{-31} kg, h = 6.6\times10^{-34} Js,$
$1eV = 1.6\times10^{-19} J]$

Solution : Kinetic energy of electron =100eV

$= 100\times1.6\times10^{-19} J = 1.6\times10^{-17} J$ we know,

$$\lambda = \frac{h}{\sqrt{2KE.m}} = \frac{6.626\times10^{-34}}{\sqrt{2\times1.6\times10^{-17}\times9.1\times10^{-31}}}$$

$= 1.228\times10^{-10} m$

EXAMPLE-46 :

The kinetic energy of an electron is 4.55×10^{-25}J. Calculate the wavelength.

$[h = 6.6\times10^{-34} Js;$
$mass\ of\ electron = 9.1\times10^{-31} kg]$

Solution : $K.E = \frac{1}{2}mv^2 = 4.55\times10^{-25}$

$$or \frac{1}{2}\times9.1\times10^{-31}\times v^2 = 4.55\times10^{-25}$$

$$or\, v^2 = \frac{2\times4.55\times10^{-25}}{9.1\times10^{-31}}$$

$v = 10^3 ms^{-1}$

Applying de Broglie equation

$$\lambda = \frac{h}{mv} = \frac{6.6\times10^{-34}}{9.1\times10^{-31}\times10^3} = 0.72\times10^{-6} m$$

EXAMPLE-47 :

The dual nature of radiations was proposed by

(1) Max Planck **(2) de-Broglie**
(3) Einstein **(4) Schrodinger**

Solution : 3

EXAMPLE-48 :

Which of the following statement is not correct?

(1) All electromagnetic radiations travel with the same velocity

(2) Matter waves generally have velocity les than electromagnetic waves

(3) Matter waves are emitted bymaterial particles

(4) Electromagnetic waves are associated with electric and magnetic fields but matter waves are not.

Solution : 3

EXAMPLE-49 :

Which of the following has largest de Broglie wavelength given that all have equal velocity?

(1) CO_2 molecule **(2) NH_3 molecule**
(3) Electron **(4) Proton**

Solution : 3

Heisenberg's Uncertainity Principle

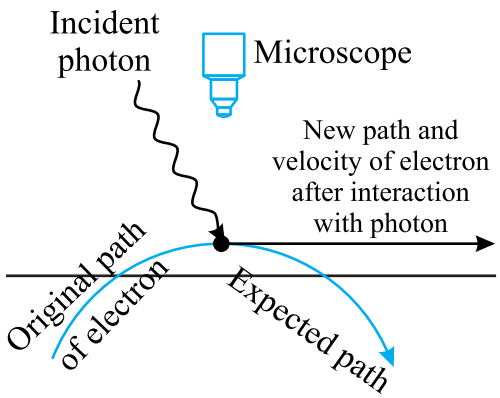
- Explanation Behind the Heisenberg's Uncertainty Principle
- If we suppose that we try to measure the position of the electron. To see its position we need small wavelength light. We know that light consists of photons. So, photons strike the electron, the reflected photon in the form of light reaches our eye and we are able to see the electron. No doubt, in this way we are able to know the position of the electron.
- We know that the electron is a point charge and almost dimensionless, therefore to know its position we must use wavelength smaller than the dimensions of the electron. Therefore shorter the wavelength, will give greater the accuracy.

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➢ But shorter wavelength is associated with high frequency and high energy.

$$Energy\ \alpha \frac{1}{Wavelength}$$

➢ Therefore when this high energy photon strikes the electron, it changes the speed as well as direction of the electron i.e., momentum changes. Thus, it is clear that momentum cannot be found with certainty in this case, but position is known.



Principle:

* It is impossible to determine simultaneously, the exact position and exact momentum (or velocity) of an electron" It is called Heisenberg's uncertainity principle. (or)
* It is impossible to determine both the position and momentum of the electron simultaneously and accurately.
* It is given by the expression

$$\Delta x . \Delta p \geq \frac{h}{4\pi} \Rightarrow \Delta x . m \Delta v \geq \frac{h}{4\pi}$$

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

where Δx = uncertainity in position

Δp = uncertainity in momentum

Δv = uncertainity in velocity

m = mass of the particle

h = Planck's constant

* It states that if one is determined with high accuracy, then the other becomes uncertain i.e. If position of the electron is determined $(\Delta x = 0)$, then $\Delta v = \infty$

similarly If velocity of the electron is determined $(\Delta v = 0)$, then $\Delta x = \infty$

* The product of the uncertainities is inversely proportional to mass of the particle.

$$\Delta x . \Delta v \propto \frac{1}{m}$$

* The uncertainty principle is mainly applicable for microscopic particles. If A and B are two particles then

$$\frac{(\Delta x . \Delta v)_A}{(\Delta x . \Delta v)_B} = \frac{m_B}{m_A}$$

Significance of uncertainity principle :

* It rules out the existence of definite paths and leads to probability which can be given by quantum mechanical model.

eg. If uncertainty in position is of only $10^{-8} m$, then the uncertainty velocity would be

$$\frac{10^{-4} m^2 s^{-1}}{10^{-8} m} \approx 10^{+4} ms^{-1} \text{ (a large value)}$$

EXAMPLE-50 :

A microscope using suitable photons is employed to locate an electron in an atom within a distance of $0.1A^0$. What is the uncertainty involved in the measurement of its velocity?

Solution: $\Delta x \Delta p = \frac{h}{4\pi}$ (or) $\Delta x m \Delta v = \frac{h}{4\pi}$

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

$$\Delta v = \frac{6.626 \times 10^{-34} Js}{4 \times 3.14 \times 0.1 \times 10^{-10} m \times 9.11 \times 10^{-31} kg}$$

$$= 0.579 \times 10^{7} ms^{-1} (1J = 1kgm^2 s^{-2})$$

$$= 5.79 \times 10^{6} ms^{-1}$$

EXAMPLE-51 :

A golf ball has a mass of 40g, and a speed of 45 m/s. If the speed can be measured

within accuracy of 2%, calculate the uncertainty in the position.

Solution :

The uncertainty in the speed is 2%, i,e

$$45\times\frac{2}{100}=0.9ms^{-1}$$

using the equation $\Delta x=\frac{h}{4\pi m\Delta v}$

$$=\frac{6.626\times10^{-34}\,Js}{4\times3.14\times40\times10^{-3}\,g\times0.9ms^{-1}}$$

$$=1.46\times10^{-33}\,m$$

This is nearly 10^{18} times smaller than the diameter of a typical atomic nucleus. As mentioned earlier for large particles, the uncertainty principle sets no meaningful limit to the precision of measurments.

EXAMPLE 52 :

An electron has a speed of 40ms^{-1} accurate upto 99.99%. What is the uncertainty in its location?

Solution : Given, $\Delta v=\frac{0.01}{100}\times40=0.004ms^{-1}$

we know, $\Delta x.\Delta v\geq\frac{h}{4\pi m}$ $\Delta x\geq\frac{h}{4\pi m\Delta v}$

$$\Delta x=\frac{6.625\times10^{-34}}{4\times3.14\times9.1\times10^{-31}\times0.004}$$

$$\Delta x=1.45\times10^{-2}\,m$$

EXAMPLE 53 : Heisenberg's uncertainty principle rules out the exact simultaneous measurement of

(1) probability and intensity
(2) energy and velocity
(3) charge density and radius
(4) positionand momentum

EXAMPLE 54 :

If uncertainty Δx in position is along X-axis, then uncertainty in momentum (Δp) is along

(1) X-axis **(2) Y-axis**
(3) Z-axis **(4) any axis**

EXAMPLE 55 : The uncertainty found from the uncertainty principle $(\Delta x.\Delta p=h/4\pi)$ is

(1) the minimum value
(2) the maximum value
(3) the exact value
(4) only an approximate value

Quantum Mechanical Model of Atom & Its Important Features :

* The fundamental equation of quantum mechanics was developed by Schrodinger.
- It explains three dimensional concept of moving electron.
- This equation is based on de-Broglie's wave equation and Heisenberg's uncertainty principle.
- Schrodinger's wave equation is written as

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

where ψ = Wave function(amplitude of the wave)
m = Mass of electron h = Planck's constant
E = Total energy of the electron
V = Potential energy of electron
x, y, z are co - ordinates.

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

∇^2 is the Laplician operator

- $\hat{H}\psi=E\psi$ where $\hat{H}$ is a mathematical operator called Hamiltonian operator.

where $\hat{H}$ is $V-\frac{h^2}{8\pi^2m}\left(\frac{\partial^2}{\partial x^2}+\frac{\partial^2}{\partial y^2}+\frac{\partial^2}{\partial z^2}\right)$

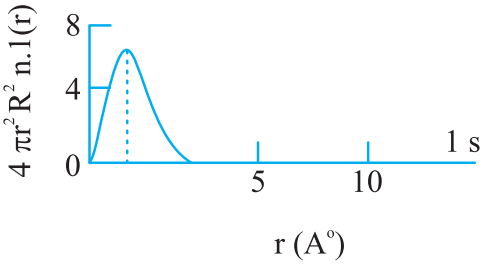
* The Schrodinger wave equation gives principal, azimuthal and magnetic quantum number but not the spin quantum number

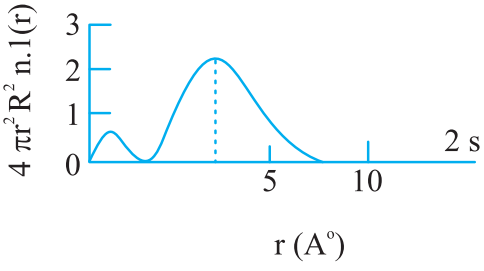
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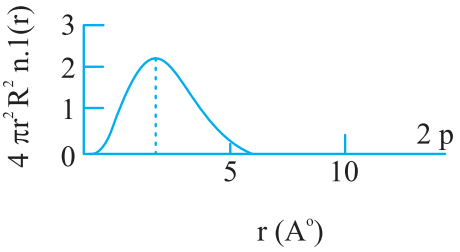
RADIAL PROBABILITY DISTRIBUTION CURVES

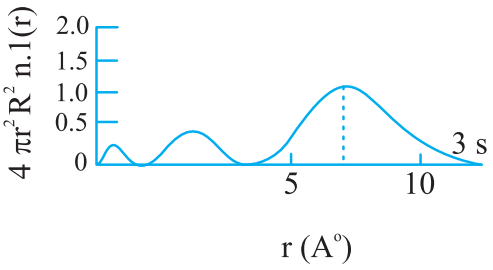
- $R(r)$ the radial function, which depends on the quantum numbers n and l.
- The radial function R has no physical meaning, but R^2 gives the probability of finding the electron at a given point r from the nucleus.
- R^2 or ψ^2 is radial probability density
- The probability of finding electron at a distance r around the nucleus is $4\pi r^2 R^2$. This is called the radial distribution function.
- The probability of the electron in a small radial volume dv being at a distance r from the nucleus is $4\pi r^2 R^2 \ dr$. This is called the radial probability distribution function (D).
- Graphs of the radial distribution function for hydrogen plotted against r are shown in Figures These disgrams show that the probability is zero at the nucleus (as $r = 0$), and by examining the plots for $1s, 2s$ and $3s$ that the most probable distance increases markedly as the principal quantum number increases.

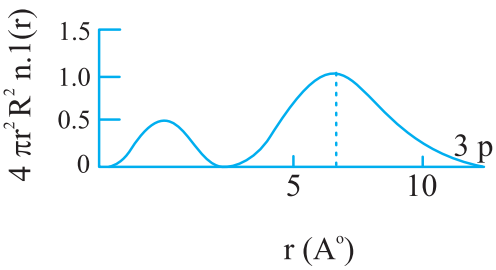
  Furthermore, by comparing the plots for $2s$ and $2p$, or $3s, 3p$ and $3d$ it can be seen that the most probable radius deccreases slightly as the subsidiary quantum number increases.

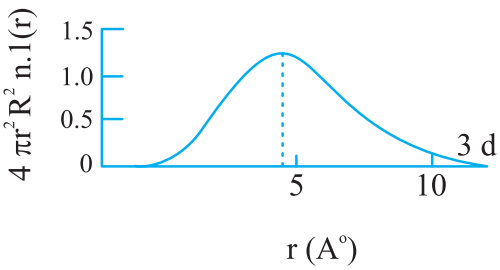




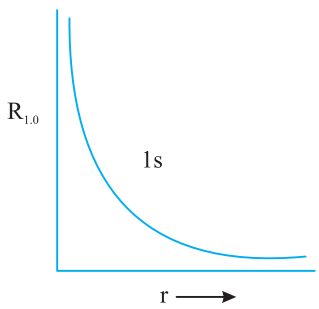


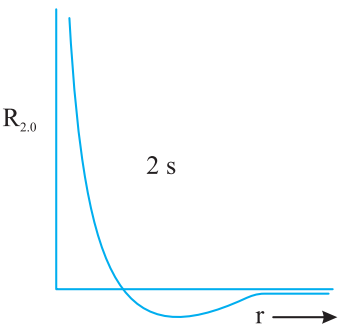


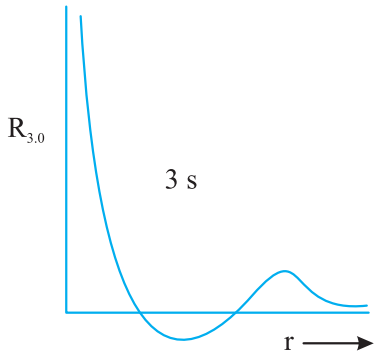




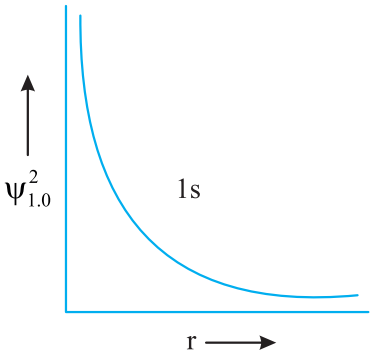
ψ or 'R' Vs 'r' graphs :

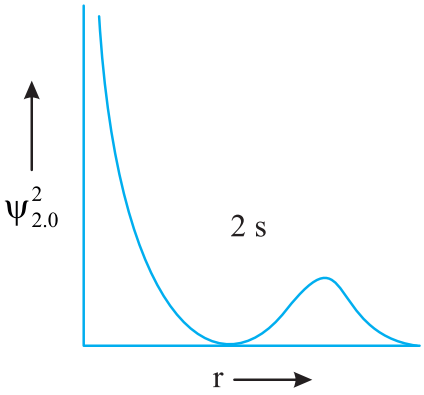






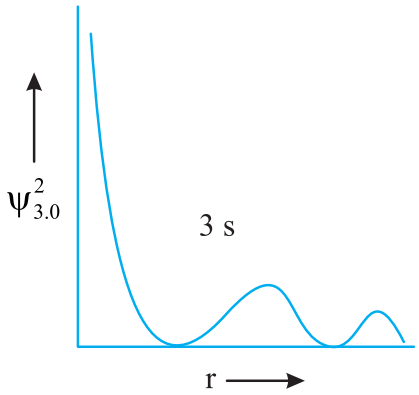
ψ^2 or R^2 Vs 'r' graphs :





Nodes or Radial nodes or nodal surfaces

- It is the region where probability of finding electron density (ψ^2) is zero .
- Number of radial nodes = n - l - 1
  where n = principal quantum number
  l = Azimuthal quantum number
  Eg. For 1s, 2p,3d,4f no of nodes =0
  For 2s, 3p,4d,5f no of nodes =1
  For 3s, 4p,5d,6f no of nodes =2



Note: The Schrodinger wave equation gives principal, azimuthal and magnetic quantum number but not the spin quantum number

Important features of the Quantum Mechanical Model of Atom:

* It states that the energy of electron in an atom is quantized.
* It explains the probability of finding the electron around the nucleus in three dimensionally.
* An atomic orbital is the wave function ψ for an electron in an atom.
* **Significance of Ψ :** It is a wave function.It corresponds to energy state which contains all information about electron.
* **SIGNIFICANCE OF ψ^2 :** It is a probability function. It indicates maximum probability of finding an electron at a certain point in an atom.

- The probability of finding an electron at a certain distance from the nucleus is called ***radial probability.***
- The curves obtained by plotting probability function $D = 4\pi r^2 \, dr \, \Psi^2$ and radial distance (r) are called radial probability distribution curves.
- Number of peaks obtained in a curve = n - l
  where n = principal quantum number
  l = Azimuthal quantum number
- The nodal surface of 2s orbital exists at a distance of $2a_0$ from the nucleus. Where a_0 is the Bohr radius $(0.529\,A^0)$

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➢ The curve for 2s orbital has two peaks the curve passes through lower maximum at 0.53 A^0 and higher maximum at 2.6A^0 radial distance.

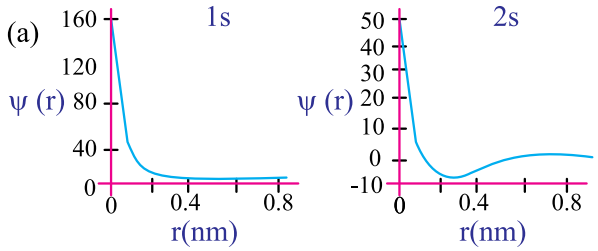
Shapes of Atomic Orbitals :

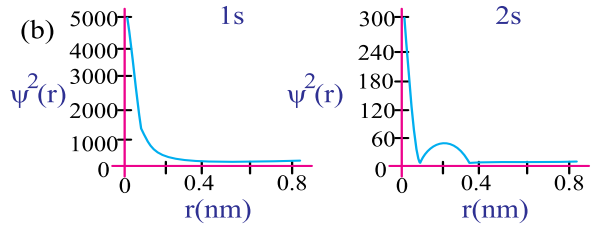
* The space around the nucleus, where the probability of finding electron (ψ^2) is maximum is called an ***orbital.***

➢ The maximum probability of finding an electron in an orbital is 95 %

S-orbital :

* The variation of ψ^2 as function of r for 1s and 2s orbitals is given as follows





* It may be noted that (from b-graph) for 1s orbital the probablity density is maximum at the nucleus and it decreases sharply as move away from it.

* For 2s orbital the probability density first decreases sharply to zero (node) and again starts increasing.

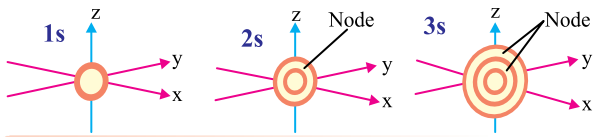
* The region where probability density (ψ^2) is zero called nodes (or) nodal surface.

* For ns-orbital, $(n-l)$ nodes are present.
eg: for 2s: 2-1 = 1, 3s : 3-1 = 2
for 4s : 4-1 = 3, ns : n-1

* The shape of s orbital (l = 0,m = 0) is spherical.It is a non directional orbital.

* Size of s- orbital increases with increase in 'n', that is $4s > 3s > 2s > 1s$.

➢ The shapes of s-orbitals are



P-orbital :

➢ In a p - sub shell, the three orbitals are represented as p_x, p_y, and p_z. These are degenerate orbitals.

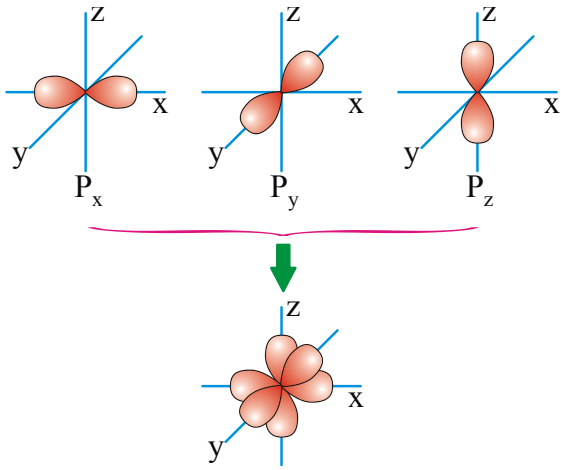
➢ The shape of a p - orbital ($l = 1$) is dumbell.

➢ p - orbitals are oriented along the axes. So they are directional orbitals.

Orbital	:	p_x	p_y	p_z
m	:	±1	±1	0

➢ With the increase of principle quantum number size, and energy of 'p' orbitals increases $4p > 3p > 2p$.

* The shapes of p-orbitals are



➢ p-sub shell is three fold degenerate.

d-ORBITAL :

* In a d - sub shell, the five orbitals are represented as $d_{xy}, d_{yz}, d_{zx}, d_{x^2-y^2}$ and d_z^2.

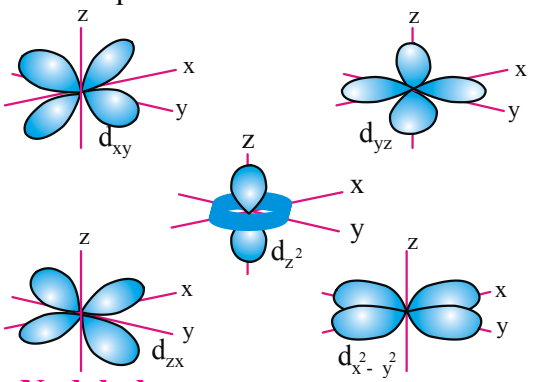
➢ These are degenerate orbitals.

➢ The shape of a d - orbital ($l = 2$) is double dumb bell.

➢ d_{xy}, d_{yz} and d_{zx} orbitals are oriented in between the axes. $d_{x^2-y^2}$ and d_z^2 orbitals are oriented along the axes.

Orbital :	d_{xy}	d_{yz}	d_{zx}	$d_{x^2-y^2}$	dz^2
m :	±2	±1	±1	±2	0

- d-sub shell is five fold degenerate.
- The shapes of d-orbitals are



Nodal planes :

- The plane where the probability of finding the electrons is zero $(\psi^2 = 0)$ is called a ***nodal plane.***

* Number of nodal planes (or) angular nodes for an orbital = l.

Orbital	No. of Nodal Planes	Nodal Plane
S	0	Nil
P_x	1	YZ
P_y	1	ZX
P_z	1	XY
d_{xy}	2	YZ, ZX
d_{yz}	2	ZX, XY
d_{xz}	2	XY, YZ
$d_{x^2-y^2}$	2	in between YZ & ZX Planes

- d_{z^2} orbital has no nodal plane ,because it has torous ring.It has two nodal cones above and below the plane.
- When the number of nodal planes increases, the energy of the orbital increases. So the energy order of the orbitals is s < p < d < f

* Number of radial nodes = n - l - 1
where n = principal quantum number
l = Azimuthal quantum number

Quantum Numbers :

- Four quantum numbers are required for the complete explanation of electrons in an atom.
  1. Principal quantum number
  2. Azimuthal quantum number
  3. Magnetic quantum number
  4. Spin quantum number

Principal Quantum Numbers (n) :

- It was proposed by NeilsBohr

* The values of n =1, 2, 3, 4 ..... or K, L,M, N ..... respectively

* It indicates the size and energy of the orbit.

* With the increase of 'n', size and energy of orbital increases

- The maximum number of electrons in an orbit = $2n^2$

* Total number of orbitals = n^2
(where n = no.of the orbit)

- Angular momentum of an electron in an orbit
$= n\frac{h}{2\pi}$

Azimuthal Quantum Numbers (l) :

- It was proposed by Sommerfeld
- The values of l depends on 'n' .

* The values are = 0, 1, 2, ...(n -1).

* The values of l represents various sub shells. When l = 0, 1, 2, 3 ... the orbitals are s, p, d, f ..... sub shells respectively.

- The energies are in the order of s < p < d < f .

* It indicates the shape of an orbital and angular momentum of electron.

* Total number of sub shells in an energy level = n

- Angular momentum of the electron in an orbital
$= \frac{h}{2\pi}\sqrt{l(l+1)} = \bar{h}\sqrt{l(l+1)}$
where h = Planck's constant
l = Azimuthal quantum number

n value	l value	Sub-shell notation
1	0	1s
2	0	2s
	1	2p
3	0	3s
	1	3p
	2	3d
4	0	4s
	1	4p
	2	4d
	3	4f

Magnetic Quantum number (m_l):

➢ It was proposed by Lande.

* The values of m depends on 'l'.

* The values ranges from $= +l \ldots 0 \ldots -l$

* The total 'm' values $= 2l + 1$

➢ The total number of 'm' values indicates the total number of orbitals in the subshell.

➢ The number of orbitals in s, p,d,f, g and h sub shells are 1, 3, 5, 7, 9 and 11 respectively.

➢ It indicates the orientation of orbitals in space.

➢ The number of orbitals in an energy level $= n^2$

* The number of orbitals in a sub shell $= 2l + 1$

➢ Maximum number of electrons in a subshell $= 2(2l + 1)$ where l = Azimuthal quantum number.

Value of l	0	1	2	3	4	5
Sub-shell	s	p	d	f	g	h
No.of ortbitals $(2l+1)$	1	3	5	7	9	11
No.of electrons $2(2l+1)$	2	6	10	14	18	22

Spin Quantum Number (m_s):

* It was proposed by Goudsmit and Uhlenbeck.

* The values of spin quantum number are independent. The values of $s = +\frac{1}{2}$ and $-\frac{1}{2}$

* For each value of 'm', there can be two 's' values.

➢ It indicates the direction of the spin of the electron.

* The clock wise $(\uparrow)$ direction spin is represented by $+\frac{1}{2}$ and anticlock wise $(\downarrow)$ direction spin is represented by $-\frac{1}{2}$

➢ Spin anuglar momentum of the electron $= \frac{h}{2\pi}\sqrt{s(s+1)}$; where 's' is total spin.

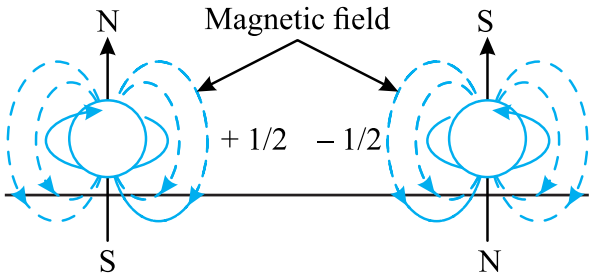
* Maximum number of electrons in an orbital = 2.

➢ The maximum number of electrons present in s, p, d and f shells are 2, 6, 10 and 14 respectively.

Total spin of an atom $= n \times \frac{1}{2}$

(n = number of unpaired electrons)

➢ Each orbital can accommodate two electrons with opposite spin or spin paired; paired electrons cancel the magnetic moment and develop mutual magnetic attraction as shown in figure.



Electrons having same spin are called spin parallel and those having opposite spin are called spin paired.

Spin paired ⟶ 

Spin parallel ⟶ 

EXAMPLE - 56 :

The magnetic quantum number for the valence electron of Caesium is

(1) 3 (2) 0 (3) – 3

(4) Any number between +3 to –3

Key : (2)

EXAMPLE - 57 :

The magnetic quantum number represents
(1) Size of the orbital
(2) Spin angular momentum
(3) Orbitalangular momentum
(4) Spatial orientation of orbital.

Key : (4)

EXAMPLE - 58 :

Which of the following statements concerning the four quantum numbers is false?
(1) n gives idea of the size of an orbital
(2) l gives the shape of an orbtial
(3) m gives the energy of the electron in the orbital
(4) s gives the direction of spin of the electron in the s-orbital.

Key: (3)

EXAMPLE 59 :

What is the total number of orbitals associated with the principle quantum number n=3?

Solution :

For n=3, the possible values of l are 0,1 and 2. Thus there is one 3s orbital (n=3, $l=1$ and m_l $-1,0,+1$); there are five 3d orbitals (n=3, $l=2$ *and* $m_l=-2,-1,0,+1,+2$).

The same value can also be obtained by using the relation; number of orbitals $=n^2, i,e\, 3^2=9$

EXAMPLE-60 :

Using s, p, d, f notations, describe the orbital with the following quantum numbers

a) n = 2, l = 1 ***b) n= 4, l = 0,***
c) n = 5, l = 3 ***d) n = 3, l = 2***

Solution :

	n	*l*	orbital
a)	2	1	*2p*
b)	4	0	*4s*
c)	5	3	*5f*
d)	3	2	*3d*

Energy of Orbitals :

* The energy of an electron in a hydrogen atom is determined by the principal quantum number. Thus the energy of the orbitals increases as follows:
1s < 2s = 2p < 3s = 3p = 3d < 4s = 4p = 4d = 4f

➢ The energy of an electron in a multielectron atom depends not only on its principal quantum number (shell), but also on its azimuthal quantum number (subshell).

* The energy of an electron is given by $(n+l)$ value.

* The lower the value of $(n+l)$ for an orbital, the lower is its energy. If two orbitals have the same value of $(n+l)$, the orbital with lower value of *'n'* will have the lower energy.

* Energies of the orbitals in the same subshell decrease with increase in the atomic number (Z_{eff}).

eg. energies of 2s orbital in IA group

$E_{2s}(H) > E_{2s}(Li) > E_{2s}(Na) > E_{2s}(K)$

Aufbau Principle :

* ***In the ground state of the atoms,the orbitals are filled in order of their increasing energies.***

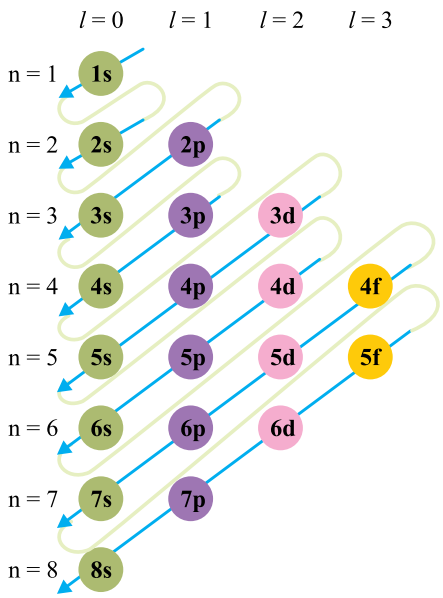
* It means , among the available orbitals, the orbitals of lowest energy are filled first.

➢ The energy value of an orbital increases as its (n + *l*) value increases.

➢ If two orbitals have the same value for (n + *l*), the orbital having lower 'n' value is first filled.

* The increasing order of energy's different orbitals is as follows.
1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s< .....

Moellar diagram of order of filling of orbitals :



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Pauli's Exclusion Principle:

* ***It gives the restrictions in filling of various orbitals with electrons.***
* It states that - No two electrons in an atom can have the same set of all four quantum numbers.
* The two electrons in a given orbital may have same values of n, *l* and m, but different 's' value.
* It says that , an orbital cannot accommodate more than two electrons. **eg.** For 2s orbital,



For 1st electron, n = 2, *l* = 0, m = 0, s = $+\frac{1}{2}$

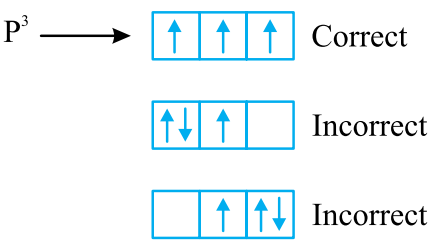
For 2nd electron, n = 2, *l* = 0, m = 0, s = $-\frac{1}{2}$

* The maximum number of electrons present in s, p, d and f shells are 2, 6, 10 and 14 respectively.

Hund's Rule :

* It deals with the filling up of degenerate orbitals with electrons (degenerate = orbitals of equal energy)
* The orbitals having the same values for n and *l* are called degenerate orbitals.

➢ **Hund's rule of maximum multiplicity** It states that electrons are distributed among the orbitals of a subshell in such a way as to give the maximum number of unpaired electrons with parallel spins.



* The pairing of electrons in the p,d and f ***orbitals*** start with the entry of 4th, 6th and 8th electrons respectively.

Electronic Configuration of Atoms :

* The distribution of electrons into orbitals of an atom is called Electronic Configuration.

➢ ***It is represented as nl^x notation or moellar diagram***

* The electrons in the inner shells are called core electrons.
* The electrons that are added to the electronic shell with highest principal quantum number are called valence electrons.

ELECTRONIC CONFIGURATIONS FROM Z = 21 to 30

Elements	Symbol	Atomic number	Electronic configuration
Scandium	Sc	21	$[Ar]3d^1 4s^2$
Titanium	Ti	22	$[Ar]3d^2 4s^2$
Vinadium	V	23	$[Ar]3d^3 4s^2$
Chromium	Cr	24	$[Ar]3d^5 4s^1$
Manganese	Mn	25	$[Ar]3d^5 4s^2$
Iron	Fe	26	$[Ar]3d^6 4s^2$
Cobalt	Co	27	$[Ar]3d^7 4s^2$
Nickel	Ni	28	$[Ar]3d^7 4s^2$
Copper	Cu	29	$[Ar]3d^7 4s^2$
Zinc	Zn	30	$[Ar]3d^7 4s^2$

In the above table the electronic configuraiton of chromimum and copper is exceptional

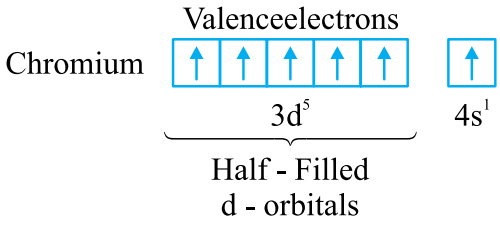
Exceptional configurations of chromium and copper:-

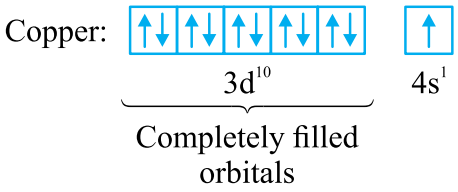
* The electronic configurations of chromium (Cr) and copper (Cu) are different from expected electronic configuration.

	Expected	Actual
Chromium	$[Ar]3d^4 4s^2$	$[Ar]3d^5 4s^1$
Copper		$[Ar]3d^9 4s^2$

* These excepted configurations are due to the fact that the half-filled and full-filled orbitals are very stable and nature loves stability and so do the elements. So, the unexcepted electronic configurations are acquired to gain stability. As in case of chromium, in its actual electronic configuration, d-orbitals are half-filled.

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Causes of stability of completely filled and Half-filled subshells. The reasons for the stability of completely filled and half-filled subshells are as follows:

***1. Symmetrical distribution of electrons:**

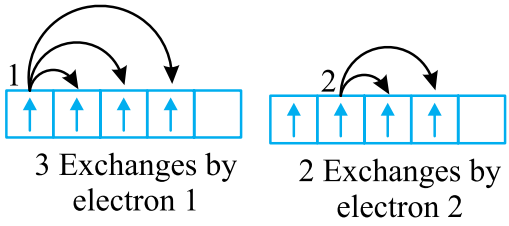
it is a known fact that the symmetry leads to stability. So, the half-filled and fully-filled orbitals are symmetrical, hence stable. Moreover this type of electrnic configuration is even favoured by energy factor. The electrons in the 3d subshell have equal energy, but they just differ in their special distribution. This results in less shielding of one another.Due to small shielding the electrons are puller closer to the nucleus. And as the electrons move closer to the nucleus, their energy decreases (more negative). Further the decrease in energy leads to stability.

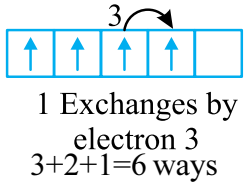
***2. Exchange energy :**

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 of energy. More are the number of electrons with identical spin more are the number of ways of exchanging with other electrons, more will be the energy released called exchange energy. Moreover, we know that decrease in energy leads to stability.

CASE 1:

if the electronic configuration of chromium was $[Ar]3d^4 4s^2$, then electrons could be exchanged in only six ways as shown below.

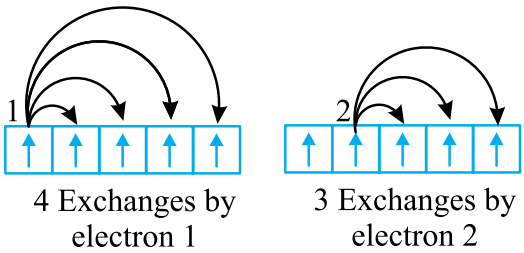


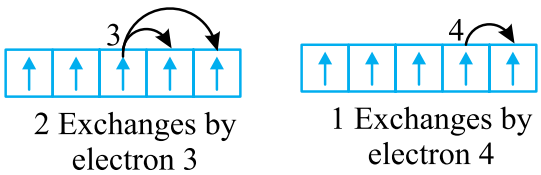


3+2+1=6 ways

Possible exchanges for a d^4 configuration

Case II: Actual configuration of chromium is





Possible exchanges for a d^5 configuration so, in total 4+3+2+1=10 ways of exchange

Here, we conclude that in actual half-filled configuration of chromium more are the number of exchanges i.e 10 in contrast to 6 exchanges in the expected configuration.

Similarly the exchange energy of d^{10} electronic configuration is more than for d^9 configuration. Therefore the d^{10} electronic configuration of copper is stable.

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CLASS ROOM EXERCISE
TEACHING QUESTIONS

SUB - ATOMIC PARTICALS AND ATOMIC MODELS

1. Which of the folowing is never true for cathode rays?
(1) They possess kinetic energy
(2) They are electromagnetic waves
(3) The produce heat
(4) They produce mechanical pressure.

2. The mass of a neutron is
(1) same as that of proton
(2) slightly less than that of a proton
(3) slightly morethan that of a proton
(4) much different from that of proton

3. The mass of the neutron is of the order of
(1) 10^{-23} kg (2) 10^{-24} kg
(3) 10^{-26} kg (4) 10^{-27} kg

4. Isobars have same number of
(1) protons (2) electrons
(3) nucleons (4) neutrons.

5. The total number of neutrons in dipositive zinc ion with mass number 70 is
(1) 34 (2) 40 (3)36 (4) 38

6. The introduction of a neutron into the nucleus of an atom would lead to a change in
(1) The number of electrons also
(2) The chemical nature of the atom
(3) Its atomic number (4) Its atomic weight

7. One of the fundamental particles is missing in one of the isotopes of hydrogen atom. The particle and isotope are respectively
1) Neutron, protium 2) Neutron, tritium
3) Proton, protium 4) Electron, tritium

8. The charge of an electron is $1.6\times10^{-19}C$ what will be the value of charge on Na^{+} ion.
1) $1.6\times10^{-19}C$ 2) $3.2\times10^{-19}C$
3) $2.4\times10^{-19}C$ 4) $10\times1.6\times10^{-19}C$

9. Which of the following is correct for cathode rays in discharge tube
1) Independent of the nature of the cathode
2) Independent of the nature of the gas
3) Deflection is observed in presence of electric and magnetic field
4) All the above

10. The specific charge for a cathode ray.
1) Has the smallest value when the discharge tube is filled with H_2
2) Is constant
3) Varies with the atomic number of gas in the discharge tube
4) Varies with the atomic number of an element forming the cathode ray

11. The specific charge for positive rays is much less than the specific charge for cathode rays. This is because:
1) Positive rays are positively charged
2) Charge on positive rays is less
3) Positive rays comprise ionised atoms whose mass is much higher
4) Experimental method for determination is wrong.

12. If S_1 be the specific charge (e/m) of cathode rays and S_2 be that of positive ray then which is true?
1) $S_1 = S_2$ 2) $S_1 > S_2$
3) $S_1 < S_2$ 4) Any one of these

13. The massive particle among the following is
1) α - particle 2) Deuteron
3) Proton 4) β - particle

14. Which of the following statements about the electron is incorrect?
1) It is a negatively charged particle.
2) The mass of electron is equal to the mass of neutron.
3) It is a basic constituent of all atoms.
4) It is a constituent of cathode rays.

15. Ernest Rutherford's model of the atom didn't specifically include the _____.
1) Proton 2) Electron
3) Nucleus 4) Neutron

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16. **The conclusions of Rutherford scattering experiment does not include:**
    1) α – particle can come within a distance of the order of $10^{-14} m$ of the nucleus.
    2) The radius of the nucleus is less than 10^{-14} m
    3) Scattering follows Coulomb's law
    4) The (+) vely charged particles of an atom move with extremely high velocities.

17. **A neutral atom, with atomic number greater than one consists of**
    1) Protons only
    2) Protons and neutrons
    3) Neutrons & electrons
    4) Neutrons, electrons & protons

18. **The ratio between the neutrons present in carbon atom and silicon atom with mass numbers 12 and 28 is**
    1) 7 : 3 2) 3 : 7 3) 1 : 2 4) 2 : 1

19. **Many elements have non-integral masses. This is because**
    1) Their isotopes have different atomic number
    2) Their isotopes have different masses
    3) Their isotopes have non-integral masses
    4) Their constituents, protons, electrons and neutrons combine to give fractional masses

20. **Among the following which is not isoelectronic with others**
    1) HF 2) H_2O 3) NH_3 4) CO

21. **Set of iso electronic ions among the following is**
    1) Na^+, Cl^-, O^{-2} 2) K^+, Ca^{+2}, F^-
    3) Cl^-, K^+, S^{-2} 4) H^+, Be^{+2}, Na^+

22. **The value of charge on the oil droplets experimentally observed were -1.6×10^{-19} and -4×10^{-19} coulomb. The value of the electronic charge, indicated by these results is**
    1) 1.6×10^{-19} 2) -2.4×10^{-19}
    3) -4×10^{-19} 4) -0.8×10^{-19}

23. **The charge to mass ratio of α – particle is approximately to the charge to mass ratio of proton is**
    1) Half 2) Twice 3) 4 times 4) 6 times

24. **The increasing order of specific charge of electron (e), proton (p), alpha particle (α) and neutron (n) is**
    1) e, p, n, α 2) n, p, e, α
    3) n, α, p, e 4) n, p, α, e

25. **The nitride ion in lithium nitride is composed of**
    1) 7 protons + 7 electrons
    2) 10 protons + 7 electrons
    3) 7 protons + 10 electrons
    4) 10 protons + 10 electrons

26. **The wrong statement among the following is**
    1) Nitrogen atom, nitride ion have same atomic number
    2) Aluminium atom and its ion have same mass number
    3) Iron atom, ferrous ion have same electron configuration
    4) Nuclear charge is same in both chlorine atom, chloride ion

27. **In which of the following species both cation and Anion have same number of electrons**
    1) CaO 2) KBr 3) NaF 4) MgS

28. **An atom contains electrons, protons and neutrons. If the mass of each neutrons is halved, and each electron is doubled, then the atomic mass of $_{12}Mg^{24}$**
    1) Gets doubled
    2) Approximately remain same
    3) Approximately get reduced by 5%
    4) Approximately get reduced by 25%

29. **Mass of one mole of electrons is**
    1) 55mg 2) 55g 3) 0.55mg 4) 0.55g

30. **The e/m ratio of cathode rays is x unit, when hydrogen is filled in the discharge tube. What will be its value when deuterium (D_2) is filled in it?**
    1) x unit 2) x/2 unit 3) 2x unit 4) x/4 unit

31. **α –particles are projected towards the following metals, with the same kinetic energy. Towards which metal, the distance of closest approach is minimum?**
    1) $Cu(Z=29)$ 2) $Ag(Z=47)$
    3) $Au(Z=79)$ 4) $Ca(Z=20)$

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32. Which of the following nuclear reactions will generate an isotope

1)Neutron particle emission

2) Positron emission

3) α – particle emission

4) β – particle emission

33. Chlorine exists in two forms, *Cl – 37* and *Cl – 35* but its atomic mass is 35.5. This indicates the ratio of *Cl – 37* and *Cl – 35* is approximately.

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

34. The mass numbers of three isotopes of an element are 10,12,14 units.Thier percentage abundance is 80,15, and5 respectively.What is the atomic weight of the element?

1) 10.5 2) 11.5 3) 12.5 4) 13.5

35. An ion with mass number 56 contains 3 units of positive charge and 30.4% more neutrons than electrons. Assign the symbol to this ion

1) $^{55}_{26}Fe^{3+}$ 2) $^{57}_{26}Fe^{3+}$ 3) $^{59}_{26}Fe^{3+}$ 4) $^{56}_{26}Fe^{3+}$

36. Which of the following conclusions could not be derived from Rutherford's α-particle scattering experiement?

(1) Most of the space in the atom is empty.

(2) The radius of the atom is about 10^{-10} m while that of nucleus is 10^{-15} m.

(3) Electrons move in a circular path of fixed energy called orbits.

(4) Electrons and the nucleus are held together by electrostatic forces of attraction.

37. Which of the following statement is not correct about the characteristics of cathode rays?

(1) They start from the cathode and move towards the anode.

(2) They travel in straight line in the absence of an external electrical or magnetic field.

(3) Characteristics of cathode rays do not depend upon the material of electrodes in cathode ray tube.

(4) Characteristics of cathode rays depend upon the nature of gas present in the cathode ray tube.

38. Which of the following statements about the electron is incorrect?

(1) It is a negatively charged particle.

(2) The mass of electron is equal to the mass of neutron.

(3) It is a basic constituent of all atoms.

(4) It is a constituent of cathode rays.

39. Which of the following properties of atom could be explained correctly by Thomson Model of atom?

(1) Overall neutrality of atom.

(2) Spectra of hydrogen atom.

(3) Position of electrons, protons and neutrons in atom.

(4) Stability of atom.

40. Two atoms are said to be isobars if.

(1) they have same atomic number but different mass number.

(2) they have same number of electrons but different number of neutrons.

(3) they have same number of neutrons but different number of electrons.

(4) sum of the number of protons and neutrons is same but the number of protons is different.

41. Isodiaphers are

1) $_9F^{19}$,$_{11}N^{23}$ 2) $_{15}P^{31}$,$_{16}S^{32}$

3) $_{17}Cl^{35}$,$_{17}Cl^{37}$ 4) $_{18}Ar^{40}$,$_{19}K^{40}$

42. Calculate the wavelength of photon having energy 5 e.V.

1) $2.47\times10^{-6}cm$ 2) $2.47\times10^{-5}cm$

3) $24.7\times10^{-5}cm$ 4) $24.7\times10^{-6}cm$

43. An element with mass number 81 contains 31.7% more neutrons as compared to protons. It symbol is

1) $_{40}X^{81}$ 2) $_{81}X^{31.7}$

3) $_{40}X^{81}$ 4) $_{35}X^{81}$

44. The fraction of volume occupied by the nucleus with respect to the total volume of an atom is

1) 10^{-15} 2) 10^{-5}

3) 10^{-30} 4) 10^{-10}

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45. Which of the following is iso-electronic with neon

1) O^{2-} 2) F^{+} 3) Mg 4) Na

46. Rutherfords experiment, which established the nuclear model of atom, used a beam of:

1) â-particle which impinged on a metal foil and get absorbed
2) ã-rays, which impinged on a metla foil and ejected electron
3) Helium atoms, which impinged on a metla foil and got scattered
4) Helium nuclie, which impinged on a metal foil and got scattered

KEY

1. (2)	2. (3)	3. (4)	4. (3)
5. (2)	6. (4)	7. (1)	8. (1)
9. (4)	10. (2)	11. (3)	12. (2)
13. (1)	14. (2)	15. (4)	16. (4)
17. (4)	18. (2)	19. (2)	20. (4)
21. (3)	22. (4)	23. (2)	24. (3)
25. (3)	26. (3)	27. (3)	28. (4)
29. (3)	30. (1)	31. (4)	32. (1)
33. (3)	34. (1)	35. (4)	36. (3)
37. (4)	38. (2)	39. (1)	40. (4)
41. (3)	42. (2)	43. (4)	44. (1)
45. ()	46. (4)		

ELECTROMAGNETIC THEORY PLANK'S QUANTUM THEORY AND PHOTO ELECTRIC EFFECT

47. A photon of 300 nm is absorbed by a gas and then reemitted two photons. One of the remitted photon has wavelength 496 nm, then the other photon has

1) 2.625×10^{-19} erg 2) $2.625\times10^{-19}\ J$
3) $2.625\ J$ 4) 2.625 Cal

48. The time period of a light is $2.0\times10^{-10}\ s$. The wavelength for this wave is

1) 0.06m 2) 6m 3) 0.03m 4) 0.3m

49. A 100 watt bulb emits monochromatic light of wavelength 400nm. Calculate the number of photons emitted per second by the bulb.

1) $20.12\times10^{20}\ s^{-1}$ 2) $2.012\times10^{20}\ s^{-1}$
3) $4.969\times10^{-19}\ s^{-1}$ 4) $49.69\times10^{-19}\ s^{-1}$

50. All types of electromagnetic radiations possess same

1) Wave length 2) Frequency
3) Energy
4) Velocity when they passed through vacuum

51. The radiation having maximum wave length is

1) Ultraviolet rays 2) Radio waves
3) X-rays 4) Infra-red rays

52. Electromagnetic radiation, which of the following has greater wavelength than visible light

1) U.V rays 2) I.R rays
3) Gamma rays 4) X-rays

53. The product of which of the following is equal to the velocity of light

1) Wave length and wave number
2) Wave length and frequency
3) Frequency and wave number
4) Wave length and amplitude

54. Which of the following relates to photons both as wave motion and as a stream of particles?

1) Interference 2) $E = mc^2$
3) Diffraction 4) $E = hv$

55. The photoelectric emission from a surface starts only when the light incident upon the surface has certain minimum

1) Intensity 2) Wavelength
3) Frequency 4) Velocity

56. A surface ejects electrons when hit by green light but not when hit by yellow light. Will electrons be ejected if the surface is hit by red light

1) Yes 2) No
3) Yes, if the red bream is quite intense
4) Yes, if the red beam continues to fall upon

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57. **Kinetic energy of photo electrons is independent on --------- of incident radiation.**
1) Wavelength 2) Wave number
3) Frequency 4) Intensity

58. **Energy levels A, B, C of a certain atoms corresponding to increasing values of energy level i.e., $E_A < E_B < E_C$. If λ_1, λ_2 and λ_3 are the wavelengths of radiations corresponding to the transitions C to B, B to A and C to A respectively which of the following statement is correct?**
1) $\lambda_3 = \lambda_1 + \lambda_2$ 2) $\lambda_3 = \frac{\lambda_1\lambda_2}{\lambda_1 + \lambda_2}$
3) $\lambda_1 + \lambda_2 + \lambda_3 = 0$ 4) $\lambda_3^2 = \lambda_1^2 + \lambda_2^2$

59. **Planck's constant has the same dimensions as that of**
1) Power 2) Work
3) Radiant energy 4) Angular momentum.

60. **The number of photons of light having wave number x in 1 J of energy source is (Planck's constant = h, velocity of light = C)**
1) hcx 2) hc/x 3) $\frac{x}{hc}$ 4) $\frac{1}{hcx}$

61. **The work function of a photoelectric material is 3.3 eV. It threshold frequency will be**
1) $4\times10^{11}\,Hz$ 2) $7.96\times10^{14}\,Hz$
3) $5\times10^{33}\,Hz$ 4) $4\times10^{11}\,Hz$

62. **In photo electric effect, the energy photon striking a metallic surface is $5.6\times10^{-19}\,J$. The kinetic energy of the ejected electrons is $12.0\times10^{-20}\,J$. The work function is**
1) $6.4\times10^{-19}\,J$ 2) $6.8\times10^{-19}\,J$
3) $4.4\times10^{-19}\,J$ 4) $6.4\times10^{-20}\,J$

63. **The kinetic energy of electrons ejected by using light having frequency equal to threshold frequency (v_0) is :**
1) hv_0 2) Almost zero
3) Very large 4) h/v_0

64. **The frequency of a wave light is $1.0\times10^6\,sec^{-1}$. The wave length for this wave is**
1) 3×10^4 cm 2) 3×10^{-4} cm
3) 6×10^4 cm 4) 6×10^6 cm

65. **The energy per quantum associated with light of wave length 250×10^{-9} m is**
1) 7.95×10^{-19} J 2) 7.95×10^{-26} J
3) 3.93×10^{-26} J 4) 3.93×10^{-19} J

66. **What is the energy of photons that corresponds to a wave number of $2.5\times10^{-5}\,cm^{-1}$**
1) 2.5×10^{-20} erg 2) 5.1×10^{-23} erg
3) 4.97×10^{-21} erg 4) 8.5×10^{-2} erg

67. **Nitrogen laser produces a radiation at a wavelength of 337.1 nm. If the number of photons emitted is 5.6×10^{24}. Calculate the power of this laser**
1) $3.33\times10^6\,J$ 2) $3.33\times10^5\,J$
3) $1.56\times10^6\,J$ 4) $15.6\times10^8\,J$

68. **The ratio of energies of photons with wavelengths $2000A^0$ and $4000A^0$ is**
1) 1/4 2) 4 3) 1/2 4) 2

69. **An Electro magnetic radiation of wavelength 242nm is just sufficient to ionise a sodium atom. Calculate the ionisation energy of sodium in kJ/mol?**
1) 494.5 2) 594.5 3) 694.5 4) 794.5

70. **When a metal is irradiated with light of frequency $4.0\times10^{16}\,s^{-1}$ the photo electrons emitted had six times the K.E as the K.E of photo electron emitted when the metal was irradiated with light of frequency $2.0\times10^{16}\,s^{-1}$. The calculate the critical frequency of the metal.**
1) $2.0\times10^{16}\,s^{-1}$ 2) $1.6\times10^{16}\,s^{-1}$
3) $3.0\times10^{16}\,s^{-1}$ 4) $4.2\times10^{16}\,s^{-1}$

71. In photo electric effect,if the energy required to over come the attractive forces on the electron(work function) of *Li, Na* and *Rb* are 2.41eV, 2.3eV and 2.09eV respectively, the work function of "*K*" could approximately be in eV
1) 2.52 2) 2.2 3) 2.35 4) 2.01

72. Neon gas emits at 616 nm. The distance travelled by this radiaton in 30 sec is.
1) $9\times10^{7}\,m$ 2) $9\times10^{9}\,m$
3) $4.5\times10^{9}\,m$ 4) $7\times10^{9}\,m$

73. A quantum of light having energy E has wavelength equal to $7200A^{0}$. The frequency of light which corresponds to energy equal to 3E, is
1) $1.25\times10^{14}\,s^{-1}$ 2) $1.25\times10^{15}\,s^{-1}$
3) $1.25\times10^{13}\,s^{-1}$ 4) $1.25\times10^{14}\,s^{-1}$

74. Bond dissociation energy of AB molecules is 300 kJ/mole. The number of moles of photons of wavelength $6625A^{0}$ requires to dissociate 3 moles of AB molecules is
1)1 2) 2 3) 4 4) 5

75. A certain dye absorbs light of $l = 4530A^{0}$ and wavelength of fluoroscence light is $5080A^{0}$. Assuming that under given condition 47% of the absorbed energy is reemited out as fluoroscence the ratio of quanta emitted out to the number of quanta absorbed.
1) 0.527 2) 1.5 3) 52.7 4) 3

76. A 25 watt bulb emits monochromatic yellow light of wavelength of $0.57\mu m$. Calculate the rate of emission of quanta per second.
1) 7.16×10^{-19} 2) 7.16×10^{19}
3) 3.5×10^{15} 4) 6.5×10^{19}

77. An electromagnetic radiation of wavelength 484nm is just sufficient to ionise a sodium atom. Calculate the ionisation energy of sodium in kJ/mol?
1) 246.9 2) 594.5 3) 694.5 4) 794.5

78. If the radiation source has the duration of 2 nano seconds and the number of photons emitted is 2.5×10^{15}. The energy of the source is
1) $8.282\times10^{10}\,J$ 2) $4.141\times10^{-10}\,J$
3) $6.262\times10^{-9}\,J$ 4) $8.282\times10^{-10}\,J$

79. A gas absorbs photon of 355nm and emits at two wavelength if one of the emission is at 680nm, the order is at
1)1035nm 2)325nm 3)743nm 4)518nm

80. The threshold frequency for a metal is $7.0\times10^{14}\,s^{-1}$. What is the kinetic energy of an electron emitted when radiation of frequency $1.0\times10^{15}\,S^{-1}$ hits the metal
1) $1.988\times10^{-17}\,J$ 2) $1.988\times10^{19}\,J$
3) $3.988\times10^{-19}\,J$ 4) $1.988\times10^{-19}\,J$

81. Iodine molecule dissociates into atoms after absorbing light of 4500 ^{0}A. If one quantum of radiation is absorbed by each molecule, calculate the kinetic energy of Iodine atoms.(Bond energy of I_2 = 240 kJ mol^{-1})
1) 2.16 ´ 10^{-20} J 2) 4.1 ´ 10^{-20}J
3) 3.12 ´ 10^{-14} J 4) 2.16 ´ 10^{-22} J

82. A photon of wavelength $4\times10^{-7}\,m$ strikes on metal surface, The work function of the metal is $2.13eV$. The velocity of the photo electron is
1) $5.67\times10^{6}\,ms^{-1}$ 2) $5.67\times10^{5}\,ms^{-1}$
3) $5.67\times10^{-5}\,ms^{-1}$ 4) $5.67\times10^{-6}\,ms^{-1}$

83. When electromagnetic radiation of wavelength 300nm falls on the surface of sodium, electrons are emitted with a Kinetic energy of $1.68\times10^{5}\,Jmol^{-1}$. The maximum wave length that will cause a photo electron to be emitted is
1) 51.7nm 2) 517nm 3) 427nm 4) 62nm

84. Photo electric emission is observed from a surface for frequencies v_1 and v_2 the KE in two cases are in ratio 1 : K, then the threshold frequency v_0 is given by
1) $\frac{v_2 - v_1}{K-1}$ 2) $\frac{Kv_1 - v_2}{K-1}$
3) $\frac{Kv_2 - v_2}{K-1}$ 4) $\frac{v_2 - v_2}{K}$

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85. **Light of certain wavelength strikes on a metal surface with intensity x and the metal emits y electrons per second of average energy, z. What will happen to y and z if x is doubled?**

1) y will be doubled and z will become half
2) y will remian same and z will be doubled
3) Both y and z will be doubled
4) y will be doubled but z will remain same

86. **Threshold frequency of metal is f_0. When light of frequency $\upsilon = 2f_0$ is incident on the metal plate, velocity of electron emitted in V_1. When a plate frequency of incident radiation is $5f_0$, V_2 is velocity of emitted electron , then $V_1:V_2$ is**

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

87. **If λ_0 is the threshold wavelength for photoelectric emission, λ the wavelength of light falling on the surface of a metal and m is the mass of the electron, then the velocity of ejected electron is given by**

1) $\left[\frac{2h}{m}(\lambda_0-\lambda)\right]^{1/2}$ 2) $\left[\frac{2hc}{m}(\lambda_0-\lambda)\right]^{1/2}$

3) $\left[\frac{2hc}{m}\left\{\frac{\lambda_0-\lambda}{\lambda\lambda_0}\right\}\right]^{1/2}$ 4) $\left[\frac{2hc}{m}\left\{\frac{1}{\lambda_0}-\frac{1}{\lambda}\right\}\right]^{1/2}$

88. **The ejection of the photo electron from the silver metal in the photo electric effect experiment can be stopped by applying the voltage of 0.35 vol. When the radiation 256.7 nm is used. The work function for silver metal is**

1) 4.48 ev 2) 3.35 ev
3) 44.8 ev 4) 22.4 ev

89. **The MRI (magnetic resonace imaging) body scanners used in hospitals. Operate with 400 MHz radio frequency energy. The wavelength corresponding to this radio frequency is**

1) 0.75 m 2) 0.75 cm
3) 1.5 m 4) 2 cm

90. **Photon of which lights has maximum energy**

1) Red 2) Blue
3) Violet 4) Green

91. **The value of Planck's constant is 6.63 x 10^{-34}Js. The velocity of light is 3 x 10^8 m/sec. Which value is closest to the wavelength in nanometer of a quantum of light with frequency of 8 x10^{15} se^{-1}**

1) 5 x 10^{-18} 2) 4 x 10^1
3) 3 x 10^7 4)) 2 x 10^{-25}

KEY

47. (2)	48. (1)	49. (2)	50. (4)
51. (2)	52. (2)	53. (2)	54. (4)
55. (3)	56. (2)	57. (4)	58. (2)
59. (4)	60. (4)	61. (2)	62. (3)
63. (2)	64. (1)	65. (1)	66. (3)
67. (1)	68. (4)	69. (1)	70. (2)
71. (2)	72. (2)	73. (2)	74. (4)
75. (1)	76. (2)	77. (1)	78. (4)
79. (3)	80. (4)	81. (2)	82. (2)
83. (2)	84. (2)	85. (4)	86. (2)
87. (3)	88. (1)	89. (1)	90. (3)
91. (2)			

ATOMIC SPECTRA
HYDROGEN SPECTRUM
BOHR'S ATOMIC MODEL

92. **The best evidence that electrons are arranged in definite orbits or energy levels is based on the observation that**

1) Atomic spectra consist of discrete lines and not continuous bands
2) Electrons in the beta ray have high kinetic energy
3) The penetrating power of cathode ray electrons depends upon the voltage used to produce them
4) Electrons revolve around the nucleu

93. **The band spectrum is caused by**

1) Molecules
2) Atoms
3) Any substance in solid state
4) Any substance in liquid state

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94. The hydrogen spectrum from an incandescent source of hydrogen is:
1) An emission band spectrum
2) An emission line spectrum
3) An absorption band spectrum
4) An absorption line spectrum

95. Transition of electron from M-shell to K-shell results in the emission of
1) Cosmic rays 2) Infrared rays
3) Ultraviolet rays 4) X-rays

96. Which of the following transition is associated with coloured spectral line
1) n = 5 to n = 3 2) n = 4 to n = 2
3) n = 2 to n = 1 4) n = 3 to n = 1

97. Values of n_1 and n_2 for H_β spectral line in the hydrogen emission spectrum
1) 1 and 2 2) 2 and 3
3) 3 and 2 4) 2 and 4

98. Rydberg constant is
1) Same for all elements
2) Different for different elements
3) A universal constant
4) Is different for lighter elements but same for heavier elements

99. The value of the total energy of an electron in the hydrogen atom is given by
1) mv^2 2) $1/2\ mv^2$ 3) $-e^2/2r$ 4) $-mv/r^2$

100. The ratio between potential energy and total energy of an electron in H-atom according to Bohr atom
1) 1 : -1 2) 1 : 1 3) 1 : 2 4) 2: 1

101. The equation corresponding to the wave number of spectral line in the Bracket series
1) $R[(1/2^2)-(1/4^2)]$
2) $R[(1/4^2)-(1/5^2)]$
3) $R[(1/3^2)-(1/5^2)]$
4) $R[(1/6^2)-(1/4^2)]$

102. The wave length of a spectral line for an electronic transition is inversely related to
1) Velocity of electron undergoing transition
2) Number of elctrons undergoing transaction
3) The difference in energy levels involved in the transition
4) None of these

103. ΔE value is maximum in
1) $E_2 - E_1 = \Delta E$ 2) $E_3 - E_2 = \Delta E$
3) $E_4 - E_3 = \Delta E$ 4) $E_5 - E_4 = \Delta E$

104. Which of the following transitions are not allowed in the normal electronic emission spectrum of an atom?
1) $2s \rightarrow 1s$ 2) $2p \rightarrow 1s$
3) $3d \rightarrow 4p$ 4) $5p \rightarrow 3s$

105. In hydrogen spectrum, the spectral line of Balmer series having lowest wavelength is
1) H_α – line 2) H_β – line
3) H_γ – line 4) H_δ – line

106. In Hydrogen atom electron is present in the N shell. If it loses energy, a spectral line may be observed in the region
1) Infra-red 2) Visible
3) Ultra-violet 4) All the above

107. The electron present in 5th orbit in excited hydrogen atoms returned back to ground state. The no. of lines which appear in Lyman series of hydrogen spectrum
1) 5 2) 10 3) 4 4) 6

108. Which of the following gives neither emission spectrum nor absorption spectrum?
1) He^+ 2) H_2 3) H^+ 4) He

109. The following electronic transition corresponds to the shortest wave length (n = no. of orbit)
1) $n_5 \rightarrow n_1$ 2) $n_5 \rightarrow n_3$
3) $n_5 \rightarrow n_2$ 4) $n_5 \rightarrow n_4$

110. Which of the following electron transition in hydrogen atom will require the energy equivalent to its ionization energy ?
1) from n = 1 to n = 2 2) from n = 2 to n =3
3) from n = 1 to n = 3
4) from n = 1 to $n = \infty$

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111. If the mass of the electron is reduced to half the Rydberg constant

1) Remains unchanged 2) Becomes half
3) Becomes double
4) Becomes one fourth

112. According to Bohr's theory, the angular momentum for an electron of 5th orbit is;

1) $10h/\pi$ 2) $5h/2\pi$
3) $25h/\pi$ 4) $5\pi/2h$

113. The change in velocity when hydrogen electron jumps from K shell to L shell is

1) One-half of its original velocity
2) Twice to its original velocity
3) One-quarter of its original velocity
4) Equal to its original velocity

114. Each hydrogen atom is excited by giving 10.2 eV. The maximum number of spectral lines in the emission is equal to

1) 1 2) 2 3) 3 4) 4

115. Consider the following statements

I) Bohr's theory can also be used to explain the spectra of He^+ ion

II) Energy of an electron in the first Bohr orbit of hydrogen atom is $-13.6\ eV$

III) Bohr's theory is only applicable to hydrogen atom and not to any other species

IV) The energy of an electron in a hydrogen atom is quantised

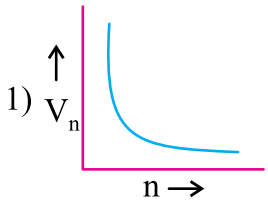
The correct statements are

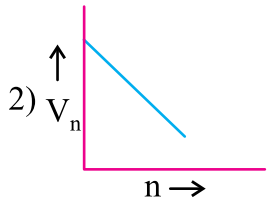
1) I, II, IV 2) II, III 3) II, IV 4) All

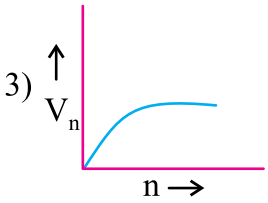
116. The ionisation potential of H-atom is $13.6\ eV$. It is exposed to electromagnetic radiation of wavelength $1028A^0$ and gives out induced radiations, then

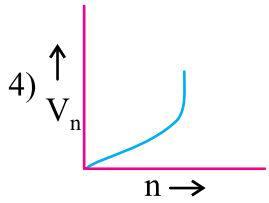
1) Longest wavelength of induced is $6568A^0$
2) Lowest wavelength of induced radiation is $102A^0$
3) Longest wavelength of induced radiation is $3252A^0$
4.) Longest wavelength of induced is $1216A^0$

117. Which of the following curves may represent the speed of electron in a H-atom as a function of principal quantum number (n)

1) 

2) 

3) 

4) 

118. The difference in angular momentum associated with the electron in two successive orbits of hydrogen atom is

1) $h/2\pi$ 2) h/π
3) $h/2$ 4) $(n-1)h/2\pi$

119. Properties of electrons that are quantized in Bohr's atomic model are

1) Mass and energy
2) Energy and angular momentum
3) Angular momentum and mass
4) Mass and charge

120. When greater number of excited hydrogen atoms reach the ground state, then

1) More number of lines are found in Lyman series
2) The intensity of lines in Balmer series increase
3) The intensity of lines in Lyman series increase
4) Both the intensity and number of lines in Lyman series increases

121. To which of the following is Bohr's theory applicable

I) He^+ **II) Li^{+2}**
III) Tritium **IV) Be^{+2}**
The correct combination is

1) III, IV 2) I, II, III, IV
3) I, II 4) I, II, III

122. The ratio of highest possible wavelength to lowest possible wavelength of Lyman series is

1) 4/3 2) 9/8 3) 27/5 4) 16/5

123. If the wave number of the first line in the Balmer series of hydrogen atom is 15000 cm^{-1}, the wave number of the first line of the Balmer series of Li^{2+} is

1) $1.35\times10^{5}\,cm^{-1}$ 2) $1.66\times10^{9}\,cm^{-1}$

3) $13.5\times10^{5}\,cm^{-1}$ 4) $1.43\times10^{4}\,cm^{-1}$

124. What is the lowest energy of the spectral line emitted by the hydrogen atom in the Lyman series? (h=Planck's constant; C=Velocity of light; R=Rydberg constant).

1) $\frac{5hcR}{36}$ 2) $\frac{4hcR}{3}$

3) $\frac{3hcR}{4}$ 4) $\frac{7hcR}{144}$

125. The ionization energy of H atom is x kJ. The energy required for the electron to jump from n = 2 to n =3 will be :

1) *5x* 2) *36x/5* 3) *5x/36* 4) *9x/4*

126. When the electron of 5th orbit jumps into the first orbit, the number of spectral lines produced in hydrogen spectrum is

1) 5 2) 10 3) 20 4) 1

127. The Ratio of m^{th} to n^{th} wavelength of Lyman series in H-spectrum is equal to

1) $\frac{\lambda_m}{\lambda_n}=\frac{(m^2-1)\times n^2}{(n^2-1)\times m^2}$

2) $\frac{\lambda_m}{\lambda_n}=\frac{(n^2-1)\times m^2}{(m^2-1)\times n^2}$

3) $\frac{\lambda_m}{\lambda_n}=\frac{(m+1)^2}{(n+1)^2}\times\frac{(n+1)^2-1}{(m+1)^2-1}$

4) $\frac{\lambda_m}{\lambda_n}=\frac{(n+1)^2}{(m+1)^2}\times\frac{(m+1)^2-1}{(n+1)^2-1}$

128. Which of the following relationship is correct?

1) E_1 of $H=\frac{1}{2}E_2$ of $He^{+}=\frac{1}{3}E_3$ of $Li^{+2}=\frac{1}{4}$ of E_4 of Be^{+3}

2) E_1 of $H=E_2$ of $He^{+}=E_3$ of $Li^{+2}=E_4$ of Be^{+3}

3) E_1 of $H=2E_2$ of $He^{+}=3E_3$ of $Li^{+2}=4E_4$ of Be^{+3}

4) E_1 of $H=\frac{2}{3}E_2$ of $He^{+}=\frac{4}{3}E_3$ of $Li^{+2}=\frac{5}{4}E_4$ of Be^{+3}

129. What is the wavelength of a photon emitted during a transition from n = 5 state to the *n* = 2 state in the hydrogen atom

1) 434nm 2) 234nm

3) 476nm 4) 244nm

130. Which one of the following transitions of an electrons in hydrogen atom emits radiation of the lowest wavelength

1) $n_2=\alpha\,to\,n_1=2$ 2) $n_2=4\,to\,n_1=3$

3) $n_2=2\,to\,n_1=1$ 4) $n_2=5\,to\,n_1=3$

131. The velocity of electron in first orbit of H-atom as compared to the velocity of light is:

1) $\frac{1}{10}th$ 2) $\frac{1}{100}th$ 3) $\frac{1}{1000}th$ 4) *Same*

132. In a collection of H-atoms, all the electrons jump from n = 5 to ground level finally (directly or indirectly), without emitting any line in Balmer series. The number of possible different radiations is

1) 10 2) 8 3) 7 4) 6

133. What is likely to be principal quantum number for a circular orbit of diameter 20.6 nm of the hydrogen atom. If we assume Bohr orbit to be the same as that represented by the principal quantum number?

1) 10 2) 14 3) 12 4)16

134. If the radius of the first Bohr orbit of Hydrogen atom is 'x', then the de-Broglie wavelength of electron in third orbit is nearly.

1) $2\pi x$ 2) $6\pi x$ 3) $9x$ 4) $x/3$

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135. A single electron in an ion has ionization energy equal to 217.6eV. What is the total number of neutrons present in one ion of it?
1) 2 2) 4 3) 5 4) 9

136. The ionisation energy for the Hydrogen atom in the ground state is $2.18\times10^{-18}\ Jatom^{-1}$. The energy required for the following process $He^{+}(g)\rightarrow He^{2+}(g)+e^{-}$ is
1) $8.72\times10^{-18}\ Jatom^{-1}$
2) $8.72\times10^{-19}\ Jatom^{-1}$
3) $4.35\times10^{-18}\ Jatom^{-1}$
4) $2.62\times10^{-19}\ Jatom^{-1}$

137. If the diameter of carbon atom is $0.15nm$, the number of carbon atoms which can be placed side by side is a straight line across length of 20 cm is
1) 13.3×10^{9} 2) 1.33×10^{9}
3) 6.2×10^{9} 4) 1.33×10^{7}

138. An electronic transition in hydrogen atom results in the formation of H_a line of hydrogen in lyman series the energies associated with the electrons in each of the orbits involved in the transition(in K.Cal/mol) are
1) $-313.6, -34.84$ 2) $-313.6, -78.4$
3) $-78.4, -34.84$ 4) $-78.4, -19.6$

139. The wavelength of a spectral line emitted by hydrogen atom in the lyman series is 16/15R cm. What is the value of n_2
1) 2 2) 3 3) 4 4) 1

140. 1.8 gm of H-atom sample are excited to radiations. The study of spectra indicates that 27% of the atoms are in 3rd energy level, 15% of atoms in 2nd energy level and rest in ground state
1) Number of atoms in 2nd energy level is 1.626×10^{23} atoms
2) Number of atoms in 3rd energy level is 2.9268×10^{23} atoms
3) Number of atoms in Ground state is 6.2872×10^{23} atoms
4) All are correct

141. Calculate the energy emitted when electrons of 1.0 gram atom of hydrogen undergo transition giving the spectral line of lowest wave energy in the visible region of its atomic spectrum.
1) $n_2 = 3$ to $n_1 = 2$; E = 182.8 KJ
2) $n_2 = 2$ to $n_1 = 1$; E = 155.8 KJ
3) $n_2 = 3$ to $n_1 = 1$; E = 180.8 KJ
4) $n_2 = 4$ to $n_1 = 2$; E = 182.5 KJ

142. In Bohr series of lines of hydrogen spectrum, the third line from the red end corresponds to which one of the following inter-orbit jumps of the electron for Bohr orbits is an atom of hydrogen.
1) $5\rightarrow2$ 2) $4\rightarrow1$ 3) $2\rightarrow5$ 4) $3\rightarrow2$

143. The energy difference between the states of n = 2 and n = 3 is 'E' eV in Hydrogen atom. The ionization potential of H atom is
1) 3.2 E 2) 7.2 E 3) 5.6 E 4) 13.2 E

144. In a certain electronic transition in the hydrogen atom from an initial state (1) to a final state (2), the difference in the orbital radius ($r_1 - r_2$) is 24 times the first Bohr radius. Identify the transition.
1) $5\rightarrow1$ 2) $25\rightarrow1$ 3) $8\rightarrow3$ 4) $1\rightarrow5$

145. I.E of He^{+} is $19.6 \times 10^{-18}\ J\ atom^{-1}$. The energy of the Ist stationary state (n=1) of Li^{+2} is
1) 4.41×10^{-16} J atom^{-1}
2) -4.41×10^{-17} J atom^{-1}
3) -2.2×10^{-15} J atom^{-1} –
4) -8.83×10^{-17} J atom^{-1}

146. When a hydrogen atom emits a photon of energy 12.1 eV, the orbital angular momentum changes by
1) $\frac{h}{2\pi}$ 2) $\frac{3h}{2\pi}$ 3) $\frac{h}{\pi}$ 4) $\frac{2h}{\pi}$

147. Energy of an electron is given by $E = -2.178 \times 10^{-18} J\left(\frac{z^2}{n^2}\right)$ wavelength of light required to excite an electron in an hydrogen atom from level n=1 to n=2 will be $[h = 6.62 \times 10^{-34} J\,s \& C = 3 \times 10^{8} \frac{m}{s}]$

1) $1.214 \times 10^{-7} m$ 2) $2.816 \times 10^{-7} m$
3) $6.5 \times 10^{-7} m$ 4) $8.5 \times 10^{-7} m$

148. According to Bohr's theory ,the angular momentum of electron in the fifth Bohr orbit is?

1) $\frac{25h}{\pi}$ 2) $\frac{1.0h}{\pi}$ 3) $\frac{10h}{\pi}$ 4) $\frac{2.5h}{\pi}$

149. In an atom, two electrons move around the nucleus in circular orbits of radii R and 4R. The ratio of the time taken by them to complete one revolution is

1) 1:4 2) 4:1 3) 1:8 4) 8:1

150. The ionisation energy of hydrogen atom is $1.312 \times 10^{6} J\,mol^{-1}$. The energy required to excite the electron in the atom from $n_1 = 1$, $n_2 = 2$ is

1) $8.51 \times 10^{5} J\,mol^{-1}$ 2) $6.56 \times 10^{5} J\,mol^{-1}$
3) $7.56 \times 10^{5} J\,mol^{-1}$ 4) $9.84 \times 10^{5} J\,mol^{-1}$

151. The approximate size of the nucleus of $^{64}_{28}Ni$ is

1) 3fm 2) 4fm 3) 5fm 4) 2fm

152. If the value of E=-78.4 Kcal/mole, the order of the orbit in hydrogen atom is

1) 4 2) 3 3) 2 4) 1

153. If velocity of an electron is 1st orbit of H atom is V, what will be the velocity of 3rd orbit of Li^{+2}

1) V 2) V/3 3) 3V 4) 9V

154. Spectrum given by the street lights of sodium vpour

1) Emission line spectrum
2) Absorption band spectrum
3) Emission band spectrum
4) Absorption band specturm

155. The energy of hydrogen atom in its ground state is -13.6 eV. The energy of the level corresponding to n = 5

1) -.054eV 2) -5.40 eV
3) -0.85eV 4) -2.72eV

156. Total no. Of lines in Lyman series of H spectrum will be (where n=no. Of orbits)

1) n 2) n-1 3) n-2 4) n(n+1)

157. The wavelength of the radiation emitted, when in a hydrogen atom electron falls from infinity of stationary state 1, would be

1) 9.1 x 10^{-8}nm 2) 192nm
3) 406nm 4) 91nm

158. The ionization enthalpy of hydrogen atom is 1.312 x 10^{6} J mol^{-1}. The energy required to excite the electron in the atom from n=1 to n=2 is

1) 8.51 x 10^{5} J mol^{-1} 2) 6.56 x 10^{5} J mol^{-1}
3) 7.56 x 10^{5} J mol^{-1} 4) 9.84 x 10^{5} J mol^{-1}

159. The energy required to break one mole of Cl-Cl bonds in Cl_2 is 242 kJ mol^{-1}. The longest wavelength of light capable of breaking a single Cl-Cl bonds is (C=3 x 10^{8} ms^{-1} and N_A=6.02 x 10^{23} mol^{-1})

1) 494 nm 2) 594nm 3) 640 nm 4) 700 nm

160. Ionisation energy of He^{+} us 19.6 x 10^{-18} J $atom^{-1}$. The energy of the first stationary state (n=1) of H atom is:

1) 8.82 x 10^{-17} J $atom^{-1}$
2) 4.41 x 10^{-16} J $atom^{-1}$
3) -4.9 x 10^{-18} J $atom^{-1}$
4) -2.2 x 10^{-15} J $atom^{-1}$

161. The radius of which of the following orbit is same as that of the first Bohr's orbit of hydrogen atom?

1) He^{+}(n=2) 2) Li^{2+}(n=2)
3) Li^{2+}(n=3) 4) Be^{3+}(n=2)

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162. The kinetic energy of an electron in the second Bohr orbit of a hydrogen atom is [a_0 is Bohr radius]

1) $\frac{h^2}{4\pi ma_0^2}$ 2) $\frac{h^2}{16\pi ma_0^2}$

3) $\frac{h^2}{32\pi ma_0^2}$ 4) $\frac{h^2}{32\pi ma_0}$

KEY

92. (1)	93. (1)	94. (2)	95. (3)
96. (2)	97. (4)	98. (2)	99. (3)
100. (4)	101. (2)	102. (3)	103. (1)
104. (3)	105. (4)	106. (4)	107. (3)
108. (3)	109. (1)	110. (4)	111. (2)
112. (2)	113. (1)	114. (1)	115. (1)
116. (1)	117. (1)	118. (1)	119. (2)
120. (3)	121. (4)	122. (1)	123. (1)
124. (3)	125. (3)	126. (2)	127. (2)
128. (2)	129. (1)	130. (3)	131. (2)
132. (4)	133. (2)	134. (2)	135. (3)
136. (1)	137. (2)	138. (2)	139. (3)
140. (4)	141. (1)	142. (1)	143. (2)
144. (1)	145. (2)	146. (3)	147. (1)
148. (4)	149. (3)	150. (4)	151. (3)
152. (3)	153. (1)	154. (1)	155. (1)
156.(2)	157. (4)	158. (2)	159. (1)
160. (3)	161. (4)	162. (3)	

DUAL NATURE OF MATTER AND LIGHT DE-BROGLIE'S AND HEISENBERG UNCERTAINITY PRINCIPLE

163. An electron travels with a velocity of v ms^{-1}. For a proton to have the same de Broglie wavelength, the velocity will be approximately

(1) $\frac{v}{1840}$ (2) $\frac{1840}{v}$ (3) 1840 v (4) v

164. Which of the folowing relates to photons both as wave motion and as a stream of particles?

(1) Interference (2) $E = mc^2$

(3) Diffraction (4) $E = hv$

165. If electron, hydrogen, helium and neon nuclei are all moving with the velocity of light, then the wavelengths associated with these particles are in the order

(1) Electron > hydrogen > helium > neon

(2) Electron > helium > hydrogen > neon

(3) Electron < hydrogen < helium < neon

(4) Neon < hydrogn < helium < electron

166. If the uncertainty in the position of an electron is zro, the uncertainty in its momentum would be

(1) zero (2) greater than h/4 π

(3) less than h/4 π (4) infinite

167. If a light of wavelength λ hits the moving electron, the uncertainty in measurement of its position will be

(1) greater than λ (2) less than λ

(3) equal to λ

(4) It can have any value

168. Diffraction of the electron beam is an evidence of the fact that

1) Electrons repel each other

2) Light has wave properties

3) Electron has wave property

4) Electron has momentum

169. Wave properties are only important for particles having

1) High mass and low velocities

2) Low mass and no velocity

3) High mass and high velocities

4) Low mass and high velocities

170. Which of the following is responsible to rule out the existence of definite paths or trajectories of electrons?

1) Pauli's exclusion principle.

2) Heisenberg's uncertainty principle.

3) Hund's rule of maximum multiplicity.

4) Aufbau principle.

171. Which of the following has the largest de Broglie wavelength provided all have equal velocity?

1) Carbon dioxide molecule

2) Ammonia molecule

3) Oxygen molecule

4) Nitrogen molecule

172. Among the following particles, which will have the shortest wavelength when accelerated by one million eV?

1) Neutron 2) Tritium atom
3) α-particle 4) Electron

173. If the wavelength of the electron is numerically equal to the distance travelled by it in 1 sec, then

1) $\lambda = \frac{h}{p}$ 2) $\lambda = \sqrt{\frac{h}{m}}$
3) $\lambda = \frac{h}{m}$ 4) $\lambda = \sqrt{\frac{h}{p}}$

174. When uncertainty in position and momentum are equal, then uncertainty in velocity is :

1) $\sqrt{\frac{h}{\pi}}$ 2) $\frac{1}{2}\sqrt{\frac{h}{\pi}}$
3) $\frac{1}{2m}\sqrt{\frac{h}{\pi}}$ 4) $2m\sqrt{\frac{h}{\pi}}$

175. If E_e, E_α, and E_p represents the kinetic energies of an electron alpha particle and a proton respectively, each moving with same deBroglie wavelength then :

1) $E_e = E_\alpha = E_p$ 2) $E_e > E_\alpha > E_p$
3) $E_\alpha < E_P < E_e$ 4) $E_e = E_P < E_n$

176. Calculate the wavelength (in nm) associated with a proton moving at $1.0\times10^3 m/s$. The mass of proton is $1.67\times10^{-27} kg$ and h is $6.63\times10^{-34} Js$

1) 0.032 nm 2) 2.5 nm
3) 14.0 nm 4) 0.4 nm

177. The de-Broglie wavelength for a proton with a velocity 15% of the speed of light is

1) $8.8\times10^{-12} m$ 2) $8.8\times10^{-15} cm$
3) $8.8\times10^{-15} m$ 4) $4.4\times10^{-15} cm$

178. The velocities of two particles A and B are 0.05 and 0.02m/s respectively. The mass of B is five times the mass of A. The ratio of their de-Broglies wavelength is

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

179. The mass of an electron is m, its charge e and it is accelerated from rest through a potential difference V. The velocity of electron will be calculated by formula :

1) $\sqrt{V/m}$ 2) $\sqrt{eV/m}$
3) $\sqrt{(2eV/m)}$ 4) None of these

180. The uncertainity in the positions of an electron and proton is equal, the ratio of the uncertainities in the velocity of an electron and proton is

1) 10^3:1 2) 1:1836
3) 3672:1 4) 1836:1

181. A ball of mass 200 gm is moving with velocity of 10 $m.s^{-1}$. If the error in measurement of velocity is 0.1%, the uncertainity in its position is

1) 3.3.x 10^{-31} m 2) 3.3 x 10^{-27} m
3) 5.3 x 10^{-25} m 4) 2.64 x 10^{-32} m

182. The kinetic energy of an electron accelerated from rest through a potential difference of 5V will be

1) 5J 2) 5erg 3)5eV 4) 8×10^{-10} eV

183. In an atom, an electron is moving with a speed of 600m/s with an accuracy of 0.005% certainity with which the position of the electron can be located is

$[h = 6.6\times10^{-34} J\,s, m = 9.1\times10^{-31} kg]$

1) $1.52\times10^{-4} m$ 2) $5.1\times10^{-3} m$
3) $1.92\times10^{-3} m$ 4) $3.84\times10^{-3} m$

184. The kinetic energy of electron is $3.0\times10^{-25} J$. The wave length of the electron is

1) $7965A^0$ 2) $4625A^0$ 3) $91A^0$ 4) $8967A^0$

185. Uncertainity in the position of an electron (mass = 9.1×10^{-31} kg) moving with velocity 300m^{-1}, accurate upto 0.001%, will be ($h = 6.63\times10^{-34}$ Js)

1) 19.2 $\times$ 10^{-2} m 2) 5.76 $\times$ 10^{-2} m
3) 1.92 $\times$ 10^{-2} m 4) 3. 83 $\times$ 10^{-2} m

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186. An electron of a velocity 'x' is found to have a certain wavelength. The velocity to be possessed by the neutron to have half the de Broglie wavelength possessed by electron is :
1) $x/1840$ 2) $x/920$ 3) $3680x$ 4) $x/3680$

187. What possibly can be ratio be of the de Broglie wavelengths for two electron having the same initial energy and accelerated through 50 volts and 200 volts?
1) 3:10 2)10:3 3) 1:2 4) 2:1

188. An α – particle is accelerated through a potential difference of V volts from rest. The de-Broglie's wavelength associated with it is
1) $\sqrt{\frac{150}{V}}A^0$ 2) $\frac{0.286}{\sqrt{V}}A^0$
3) $\frac{0.101}{\sqrt{V}}A^0$ 4) $\frac{0.983}{\sqrt{V}}A^0$

QUANTUM MECHANICAL MODEL OF ATOM

189. The designation of a subshell with n = 4 and l =3 is
(1) 4s (2) 4p (3) 4d (4) 4f

190. The shape of the orbital with l = 1 and m = 0 is
(1) spherical (2) dumb-bell
(3) clover-leaf (4) dough-nut

191. Which of the following orbitals is not symmetrical about the Z-axis?
(1) p_z (2)d_z2 (3) s (4) d_{xz}

192. Ψ^2 (psi) the wave function represents the probability of finding electron. Its value depends
1) Inside the nucleus
2) Far from the nucleus
3) Near the nucleus
4) Upon the type of orbital

193. In the Schrodinger wave equation ψ represents
1) Orbitals 2) Wave function
3) Amplitude function 4) Both 2 & 3

194. The electron density of $3d_{xy}$ orbital in YZ plane is
1) 50 % 2) 95 % 3) 33.33 % 4) Zero

195. In an orbital, the signs of lobes indicate the
1) Sign of the wave function
2) Sign of the probability distribution
3) Presence or absence or electron
4) Sign of charge

196. The number of radial nodes, nodal planes for an orbital with n = 4 ; l = 1 is
1) 3, 1 2) 2, 1 3) 2, 0 4) 4, 0

197. The number of angular and radial nodes of 4d orbital respectively are
1)3,1 2)1,2 3)3,0 4)2,1

198. The number of radial nodes of 3s and 2p orbitals respectively are
1) 0,2 2)2,0 3)1,2 4)2,1

199. The basis of quantum mechanical model of an atom is
1) Angular momentum of electron
2) Qantum numbers
3) Dual nature of electron
4) Black body radiation

200. In the plots of radial distribution function for the hydrogen 3s orbital versus 'r', the no. of peaks are
1) 3 2) 2 3) 1 4) 0

201. Which of the following conditions is incorrect for
1) ψ must be single valued at any particular point
2) ψ must be positive
3) ψ must be a continuous function of its coordinates
4) None of the above

202. Consider the following statements :
1) Electron density in XY plane in $3d_{x^2-y^2}$ orbital is zero
2) Electron density in XY plane in $3d_z^2$ orbital is zero
3) 2s orbital has only one spherical node
4) For $2p_z$ orbital YZ is the nodal plane
The correct statements are :
1) 2 and 3 2) 1,2,3,4 3) Only 2 4) 1 & 3

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203. Which of the following statements is correct?

1) An orbital describes the path of an electron in an atom
2) An orbital is a region where the electron is not located
3) An orbital is a function which gives the probabilities of finding the electron in a given region
4) All the above

204. For an electron in a hydrogen atom, the wave function ψ is proportional to exp, where a_0 is the Bohr's radius. What is the ratio of the probability of finding the electron at the nucleus to the probability of finding it at a_0?

1) e 2) e^2 3) $1/e^2$ 4) zero

205. The electron density between 1s and 2s is

1) High 2) Low
3) Zero 4) Abnormal

206. Which of the following statement(s) is/are correct about angular nodes

1) They are independent from the radial wave function
2) They are directional in nature
3) The number of angular nodes of orbital is equal to azimuthal quantum number.
4) All are correct

207. Which one of the following conditions is incorrect for a well behaved wave function (ψ)

1) ψ must be finite
2) ψ must be single valued
3) ψ must be infinite
4) ψ must be continuous

208. The number of radial nodes for 3*p* orbital is

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

209. Number of angular nodes for 4*d* orbital is

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

210. In ψ_{321} the sum of angular momentum, spherical nodes and angular mode is:

1) $\frac{\sqrt{6}h+4\pi}{2\pi}$ 2) $\frac{\sqrt{6}h}{2\pi}+3$

3) $\frac{\sqrt{6}h+2\pi}{2\pi}$ 4) $\frac{\sqrt{6}h+8\pi}{2\pi}$

211. Which of the following statement in relation to the hydrogen atom is correct?

1) 3s, 3p and 3d orbitals all have the same energy
2) 3s and 3p orbitals are of lower energy than 3d orbital
3) 3p orbital is lower in energy than 3d orbital
4) 3s orbital is lower in energy than 3p orbitals

KEY

163. (1)	164. (4)	165. (1)	166. (4)
167. (3)	168. (2)	169. (4)	170. (2)
171. (2)	172. (3)	173. (2)	174. (3)
175. (3)	176. (4)	177. (3)	178. (1)
179. (3)	180. (4)	181. (4)	182. (3)
183. (3)	184. (4)	185. (3)	186. (2)
187. (4)	188. (3)	189. (4)	190. (2)
191. (4)	192. (4)	193. (4)	194. (4)
195. (1)	196. (2)	197. (4)	198. (2)
199. (3)	200. (1)	201. (2)	202. (1)
203. (3)	204. (4)	205. (3)	206. (4)
207. (3)	208. (4)	209. (3)	210. (1)
211. (3)			

QUANTUM NUMBERS

212. How may electrons can have n = 4 and l = 2 ?

(1) 6 (2) 10 (3) 18 (4) 32

213. Which shell would be the first to have a g-sub-shell?

(1) 4th (2) 5th (3) 6th (4) 7th

214. If m has seven values, the value of azimuthal quantum number should be

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

215. The quantum number which determines the number of sub-energy levels in any main energy level is

1) n 2) l 3) m 4) s

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216. Among the various quantum numbers (n, *l*, m, s)describing an electron which can have the largest value
1) n 2) *l* 3) m 4) s

217. The angular momentum of an electron in an atom depends on
1) m 2) *l* 3) n 4) All

218. A 3d electron having s = + 1/2 can have a magnetic quantum number
1) + 2 2) + 3 3) - 3 4) - 4

219. The $2p_x$, $2p_y$ and $2p_z$ orbitals of an atom have identical shapes but differ in their
1) Size 2) Shape
3) Orientation 4) Spin

220. The orbital with maximum number of possible orientations
1) s 2) p 3) d 4) f

221. The quantum number which cannot say any thing about an orbital is
1) n 2) *l* 3) m 4) s

222. Which is not an atomic orbital?
1) *2d* 2) *5p* 3) *3p* 4) *4d*

223. The quantum number in which the valence electrons of magnesium differs in
1) m 2) n 3) *l* 4) s

224. The set of quantum numbers not possible to an electron is
1) 1,1,1, +1/2 2) 1,0,0, +1/2
3) 1, 0, 0, - 1/2 4) 2, 0,0, +1/2

225. The quantum number *l* and the number of electrons (n) in the sub level are related by
1) $n = 2l + 1$ 2) $l = 2n + 1$
3) $n = 4l + 2$ 4) $n = 2l^2$

226. The set of quantum numbers
n=3, l=2, m_l= 0
1) Describes an electron in a 2*s* orbital
2) Is not allowed
3) Describes an electron in a 3*p* orbital
4) Describes one of the five orbitals of a similar type

227. The orbital having minimum 'm' value
1) Spherical in shape 2) Dumbell in shape
3) Double dumbell in shape 4) Tetrahedral

228. An orbital made of four lobes can have the following quantum numbers
1) n = 2, l = 2, m = 0
2) n = 3, l = 1, m = –2
3) n = 3, l = 2, m = 0
4) n = 3, l = 3, m = –3

229. Number of electrons of manganese with magnetic quantum number value '0' is
1) 1 2) 8 3) 12 4) 13

230. The total number of m values for n=4 is
1) 8 2) 16 3) 12 4) 20

231. Choose the incorrect statement :
1) The shape of an atomic orbital depends upon the azimuthal quantum number
2) The orientation of an atomic orbitals depend upon the magnetic quantum number
3) The energy of an electron in an atomic orbital of multi-electron atom depends on principal quantum number.
4) The number of degenerate atomic orbitals of one type depends on the value of azimuthal and magnetic quantum numbers.

232. Which of the following statements on quantum numbers is not correct?
1) Quantum numbers n, l, m and m_s are needed to describe an electron in an atom completely.
2) Quantum numbers n, l, m and s are obtained by solving the Schrodinger wave equation.
3) A subshell in an atom can be designated with two quantum numbers n and l.
4) The maximum value of l is equal to n - 1 and that of m is $\pm l$.

233. Which of the following sets of quantum numbers represents the highest energy of an atom?
1) n = 3, l = 1, m = 1, s = + 1/2
2) n = 3, l = 2, m = 1, s = + 1/2
3) n = 4, l = 0, m = 0, s = + 1/2
4) n = 3, l = 0, m = 0, s = + 1/2

234. In an atom the order of increasing energy of electrons with quantum numbers
i)n=4, l = 1 ii)n=4,l =0
iii)n=3,l =2 iv)n=3,l =1 is
1)iii< i < iv < ii 2)ii < iv <i < iii
3) i < iii < ii < iv 4) iv < ii < iii <i

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235. The angular wave function depends upon quantum numbers.
1) n and l 2) l and m 3) l and s 4) m and s

236. The minimum angular momentum of an electron with the magnetic quantum numbers –1, 0, +1
1) $\sqrt{\frac{3}{2}\frac{h}{\pi}}$ 2) $\frac{h}{\pi}$ 3) $\frac{2h}{\pi}$ 4) $\frac{3}{2}\frac{h}{\pi}$

237. The electrons occupying the same orbital have the same values for all the quantum number except for
1) n 2) l 3) m 4) s

238. In order to designate an orbital in an atom the no. of quantum no. required
1) One 2) Two 3) Three 4) Four

239. The max. number of electron that can be accomodated in all the orbitals for which $l = 3$ is
1) 2 2) 6 3) 10 4) 14

240. During ionisation of copper atom, the quantum numbers of electron removed maybe
1) $n = 4\ l = 1 \quad s = +\frac{1}{2}$
2) $n = 3\ l = 0 \quad s = -\frac{1}{2}$
3) $n = 4\ l = 0 \quad s = +\frac{1}{2}$
4) $n = 4\ l = 2 \quad s = -\frac{1}{2}$

241. The spin magnetic momentum of electrons in an ion is 4.84 BM. Its total spin will be
1) ± 1 2) ± 2 3) $\geq \sqrt{\frac{h}{4\pi}}$ 4) ± 2.5

242. The maximum number of sub levels,orbitals and electrons in N shell of an atom are respectively
1) 4, 12, 32 2) 4, 16, 30
3) 4, 16, 32 4) 4, 32, 64

243. In multi electron atom, which of the following orbitals described by the three quantum numbers will have the same energy in the absence of magnetic and electric fields.
a) n = 1, l = 0, m = 0
b) n = 2, l = 0, m = 0
c) n = 2, l = 1, m = 1
d) n = 3, l = 2, m = 1
e) n = 3, l = 2, m = 0
1) a & c 2) b & c 3) c & d 4) d & e

244. The values of four quantum numbers of valence electron of an element X is n = 4, l = 0, m = 0, s = 1/2 The element is
1) *K* 2) *Ti* 3) *Na* 4) *Sc*

245. Given

K	L	M	N
2	8	11	2

The number of electrons present in $l = 2$ is
1) 3 2) 6 3) 5 4) 4

246. An impossible set of four quantum number of an electron is
1) n = 4, l = 2, m = -2, s = + 1/2
2) n = 4, l = 0, m = 0, s = +1/2
3) n = 3, l = 2, m = -3, s = + 1/2
4) n = 5, l = 3, m = 0, s = - 1/2

247. Total number of orbitals associated with third shell will be _____.
(1) 2 (2) 4 (3) 9 (4) 3

248. Orbital angular momentum depends on
(1) l (2) n and l
(3) n and m (4) m and s

249. Which of the following sets of quantum numbers represents the highest energy of an atom
1) n = 3, l = 1, m = 1, s = +1/2
2) n = 3, l = 2, m = 1, s = +1/2
3) n = 4, l = 0, m = 0, s = +1/2
4) n = 3, l = 0, m = 0, s = +1/2

250. What is likely to be principal quantum number for a circular orbit of diameter 20nm of the hydrogen atom if we assume Bohr orbit be the same as that represented by the principal quantum number ?
1) 10 2) 14 3) 12 4) 16

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251. The orbital with zero angular momentum is

1) s 2) p 3)d 4) f

KEY

212. (2)	213. (2)	214. (3)	215. (1)
216. (1)	217. (2)	218. (1)	219. (3)
220. (4)	221. (4)	222. (1)	223. (4)
224. (1)	225. (3)	226. (4)	227. (3)
228. (3)	229. (4)	230. (2)	231. (3)
232. (2)	233. (2)	234. (4)	235. (2)
236. (2)	237. (4)	238. (3)	239. (4)
240. (3)	241. (2)	242. (3)	243. (4)
244. (1)	245. (1)	246. (3)	247. (3)
248. (1)	249. (2)	250. (2)	251. (1)

ELECTRONIC CONFIGURATION

252. According to $(n+l)$ rule after completing 'np' level the electron enters into

1) (n – 1) d 2) (n + 1) s
3) nd 4) (n + 1) p

253. If Pauli's exclusion principle is not known, the electronic arrangement of lithium atom is

1) $1s^2\ 2s^1$ 2) $1s^1\ 2s^2$
3) $1s^3$ 4) $1s^2\ 2s^1\ 2p^1$

254. Any p - orbital can accommodate upto

1) Four electrons
2) Two electrons with parallel spins
3) Six electrons
4) Two electrons with opposite spins

255. Due to which of the following reasons the nitrogen shows three unpaired electrons

1) Hund's rule 2) Aufbau principle
3) Pauli's principle
4) Heisenberg's principle

256. Mg^{+2}, Al^{+3} have identical _____

1) Configuration 2) Atoms
3) Ions 4) Molecules

257. The maximum number of electrons in an atom which can have $(n+l)=4$

1) 2 2) 6 3) 8 4) 18

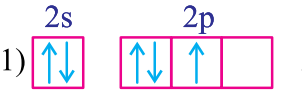
258. The valency shell electron configuration of an atom is $4s^2\ 4p^5$. The maximum no. of electrons having parallel spin in this configuration are

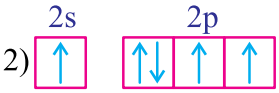
1) 7 2) 4 3) 3 4) 5

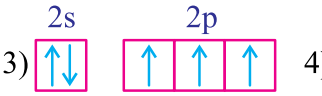
259. The rule that explains the reason for chromium to have $[Ar]3d^5\ 4s^1$ configuration instead of $[Ar]\ 3d^4s^2$?

1) Pauli's exclusion principle 2) Aufbau Principle
3) Hund's rule
4) Heisenberg principle

260. The orbital diagram in which the aufbau principle is violated is

1) 

2) 

3) 

4) 

261. The electronic configuration of an element with atomic number 64 is

1) $6s^2\ 5d^1\ 4f^7$ 2) $6s^2\ 5d^0\ 4f^8$
3) $6s^1\ 5d^0\ 4f^7$ 4) $6s^1\ 5d^2\ 4f^7$

262. How many 'd' electrons are present in Cr^{2+} ion?

1) 4 2) 5 3) 6 4) 3

263. Which of the following statements is incorrect?

1) Extra stability of half filled and completely filled orbitals among s and p block elements is reflected in trends of IE across a period.
2) Extra stability of half filled and completely filled orbitals among s and p block elements is reflected in E.A. trends across a period.
3) Aufbau principle is incorrect for cases where energy difference between ns and (n – 1)d subshell is larger
4) Extra stability to half filled subshell is due to higher exchange energies.

264. The ion that is most stable

1) Fe^+ 2) Fe^{2+} 3) Fe^{3+} 4) Fe^{4+}

265. Which has the same number of s-electrons as the d-electrons in Fe^{2+} ?

1) Li 2) Na 3) N 4) P

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266. Which of the following statements are incorrect

I) There are five unpaired electrons in Fe^{+3}(z=26)

II) Fe^{+3}, Mn^{+} and *Cr* all having 24 electrons have same value of magnetic moment

III) Copper (*I*) chloride is coloured salt

IV) Every coloured ion is paramagnetic

1) I & II 2) II & III

3) III & IV 4) I & IV

267. Sum of electronic spins of all electrons with the configuration $3d^7$ is

1) + 3/2 2) + 5/2 3) + 7/2 4) 9/2

268. Which one of the following pairs of ions have the same electronic configuration?

1) Cr^{+3}, Fe^{+3} 2) Fe^{+3}, Mn^{+2}

3) Fe^{+3}, CO^{+3} 4) Sc^{+3}, Cr^{+3}

269. How many electrons are present in the M-shell of an atom of the element with atomic number Z=24?

1) 5 2) 6 3) 12 4) 13

270. The atomic numbers of elements X,Y and Z are 19, 21 and 25 respectively. The number of electrons present in the M shells of these elements follow the order.

1) Z>X>Y 2) X>Y>Z

3) Z>Y>X 4)Y>Z>X

271. Which one of the following sets correctly represents the in case in the paramagnetic property of the ions.

1) $Cu^{+2} > V^{+2} > Cr^{+2} > Mn^{+2}$

2) $Cu^{+2} < Cr^{+2} < V^{+2} < Mn^{+2}$

3) $Cu^{+2} < V^{+2} < Cr^{+2} < Mn^{+2}$

4) $V^{+2} < Cu^{+2} < Cr^{+2} < Mn^{+2}$

272. Which of the following options does not represent ground state electronic configuration of an atom?

(1) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^8 4s^2$

(2) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^9 4s^2$

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

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

273. Which of the following is responsible to rule out the existence of definite paths or trajectories of electrons?

(1) Pauli's exclusion principle.

(2) Heisenberg's uncertainty principle.

(3) Hund's rule of maximum multiplicity.

(4) Aufbau principle.

274. The pair of ions having same electronic configuration is _____.

(1) Cr^{3+}, Fe^{3+} (2) Fe^{2+}, Mn^{2+}

(3) Fe^{3+}, Co^{3+} (4) Sc^{3+}, Cr^{3+}

275. For the electrons of oxygen atom, which of the following statements is correct?

(1) Zeff for an electron in a 2*s* orbital is the same as Zeff for an electron in a 2*p* orbital.

(2) An electron in the 2*s* orbital has the same energy as an electron in the 2*p* orbital.

(3) Zeff for an electron in 1*s* orbital is the same as Zeff for an electron in a 2*s* orbital.

(4) The two electrons present in the 2*s* orbital have spin quantum numbers m_s but of opposite sign.

276. The spin only magnetic moment [in BM] of Ni^{+2} in aqueous solution would be (atomic no of *Ni*=28)

1) 2.84 2) 4.9 3) 0 4) 1.73

277. The magnetic moment of cobalt of the compound $Hg[Co(SCN)_4]$ is [Given :Co^{+2}]

1) $\sqrt{3}$ 2) $\sqrt{8}$ 3) $\sqrt{15}$ 4) $\sqrt{24}$

KEY

252. (2) 253. (3) 254. (4) 255. (1)
256. (1) 257. (3) 258. (2) 259. (3)
260. (2) 261. (1) 262. (1) 263. (3)
264. (3) 265. (4) 266. (2) 267. (1)
268. (2) 269. (4) 270. (3) 271. (3)
272. (2) 273. (2) 274. (2) 275. (4)
276. (1) 277. (3)

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HINTS & SOLUTIONS

7. $H_1^1(\text{Pr}otium)$ number of neutrons = 1-1=0
12. mass of electron is less
17. Electron,proton and neutron are the fundametal particles
18. n = A-Z
20. $HF, H_2O and\ NH_3$ have 10 electrons
22. q = ne
23. e/m value of ion $= \dfrac{magnitude\ of\ ch\arg e}{mass\ number}$
25. N^{3-}
26. Neutral atom and ions will have different no.of electrons.
27. No.of electrons = $atomic\ no \pm ch\arg e$
28.

	p	n	mass
Initial	12	12	24
Final	12	6	18

$= \dfrac{6\times100}{24}$ =25% reduced.

30. e/m value of cathode rays independent on the nature of gas
31. Isotopes are species are having same number of protons but different number of neutrons
32. $\text{Avg.at.wt} = \dfrac{\sum \%\text{abundance}\times\text{Atomicweight}}{\text{Totalratio}}$
34. $\text{Avg.at.wt} = \dfrac{\sum \%\text{abundance}\times\text{Atomicweight}}{\text{Totalratio}}$
35. Let the number of electrons in $A^{+3} = x$
56 = x+3+1.304x
42. $\lambda = \dfrac{12375}{E}A^0$
47. $E_2 = hc[\dfrac{1}{\lambda} - \dfrac{1}{\lambda_1}]$
48. $\lambda = \dfrac{c}{\nu} \quad \because \nu \propto \dfrac{1}{t} \Rightarrow \lambda = ct$
49. $E = \dfrac{Nhc}{\lambda} \Rightarrow N = \dfrac{E\lambda}{hc}$
51. $|X-rays|U.V\ rays|I.R|Radio$
$\rightarrow \lambda \uparrow$
52. $\gamma-rays|X-rays|U.V\ rays|I.R|Radio$
$\rightarrow \lambda \uparrow$
53. $\nu = \dfrac{c}{\lambda} \Rightarrow c = \nu\lambda$
56. Red light has less frequency than green. So, it will not cause ejection of electrons
58. $E = \dfrac{hc}{\lambda}$
60. $E = Nhc\bar{\nu}\ ; N = \dfrac{E}{hc\bar{\nu}}$
61. $W = h\nu_0\ ; \nu_0 = \dfrac{W}{h}$
62. $h\nu = W + K.E\ ;\ W = h\nu - K.E$
64. $\nu = \dfrac{c}{\lambda}$
65. $E = \dfrac{hC}{\lambda}$
66. $E = hc\bar{\nu}$
67. $E = \dfrac{Nhc}{\lambda}$
68. $\dfrac{E_1}{E_2} = \dfrac{\lambda_2}{\lambda_1}$
69. $E = \dfrac{hc}{\lambda} = E\times N_A$
70. $\dfrac{K.E_2}{K.E_1} = \dfrac{\nu_2 - \nu_0}{\nu_1 - \nu_0}$
71. As the size of atom increases, energy required to over come the attractive forces on the outer most electron decreases.
72. $v = \dfrac{d}{t} \Rightarrow d = v\times t\ \ [\because v = 3\times10^8 m/\sec]$
73. $\nu_1 = \dfrac{c}{\lambda_1}; \nu_2 = \dfrac{c}{\lambda_2} \qquad \dfrac{E_1}{E_2} = \dfrac{\nu_1}{\nu_2}$
74. $E = \dfrac{Nhc}{\lambda}$

75. $E=\frac{Nhc}{\lambda}$; $\frac{E_1}{E_2}=\frac{\lambda_2}{\lambda_1}$

76. $E=\frac{Nhc}{\lambda}$

77. $E=\frac{hc}{\lambda}$ Energy of one mole= $E\times N_A$

78. $E=Nh\nu$ $(\because \nu=\frac{1}{t}); E=\frac{Nh}{t}$

79. $E_{total}=E_1+E_2$

$\frac{hc}{\lambda}=\frac{hc}{\lambda_1}+\frac{hc}{\lambda_2}\Rightarrow\frac{1}{\lambda}=\frac{1}{\lambda_1}+\frac{1}{\lambda_2}$

$\frac{1}{353}=\frac{1}{680}+\frac{1}{\lambda_2}$

80. $h\nu=h\nu_0+kE\,;K.E=h(\nu-\nu_0)$

81. B.E for 1 molecule $=\frac{240\times10^3}{6\times10^{23}}=40\times10^{-20}J$

$=4\times10^{-19}J$

Energy of incident light

$E=\frac{hc}{\lambda}=4.41\times10^{-19}J$

$K.E=E-W$ $K.E=0.41\times10^{-19}$

$=4.1\times10^{-20}J$

82. $h\nu=W+\frac{1}{2}mv^2\,;\frac{hC}{\lambda}=W+\frac{1}{2}mv^2$

$(\because 1eV=1.6\times10^{-19}J)$

$W=2.13\times1.6\times10^{-19}J$

83. $h\nu=W+K.E\,;W=\frac{hc}{\lambda}-K.E$

$W=\frac{hc}{\lambda_0}\Rightarrow\lambda_0=\frac{hc}{W}$

84. $hv_1=hv_0+1\Rightarrow h\left[v_1-v_0\right]=1$ ----------(1)

$hv_2=hv_0+k\Rightarrow h\left[v_2-v_0\right]=k$ ---------- (2)

$\frac{(1)}{(2)}=\frac{1}{K}=\frac{v_1-v_0}{v_2-v_0}$

$v_2-v_0=Kv_1-Kv_0$ $Kv_0-v_0=Kv_1-v_2$

$v_0\left[K-1\right]=Kv_1-v_2$ $v_0=\frac{Kv_1-v_2}{K-1}$

85. no. of photo electrons ejected α intensity of radiation but KE is independent of intensity of radiation.

86. $K.E=h(v-v_0)\text{Þ}\frac{1}{2}mv^2=h(v-v_0)$

$v^2=\frac{2h(v-v_0)}{m}\,;v_1^{\,2}=\frac{2h(2f_0-f_0)}{m}$ -----(1)

$v_2^{\,2}=\frac{2h(5f_0-f_0)}{m}$ --------2

87. $\frac{hc}{\lambda}=h\frac{c}{\lambda_0}+\frac{1}{2}mv^2$

$hc\left[\frac{1}{\lambda}-\frac{1}{\lambda_0}\right]=\frac{1}{2}mv^2$

$V=\left(\frac{2hc}{m}\left[\frac{\lambda_0-\lambda}{\lambda_0\lambda}\right]\right)^{1/2}$

88. $h\nu=w+K.E\,;w=h\nu-K.E$

$=\frac{hc}{\lambda}-K.E=\frac{6.625\times10^{-34}\times3\times10^8}{256.7\times10^{-9}}-0.35$

89. $V=\frac{C}{\lambda}=\frac{3\times10^8}{400\times10^6}$

$=0.75$ m

91. $V=\frac{C}{\lambda}$

$\lambda=\frac{3\times10^8}{8\times10^{15}}$

0.375×10^{-7}

$37.5\times10^{-9}m$

$37.5\text{nm}\approx4\times10^1$

95. Lyman series of line are observed in uv-region.

96. n=4 to n=2 falls in visible region

97. For H_β, $\Delta n=n_2-n_1=2$

100. $\frac{P.E}{T.E}=\frac{\frac{-e^2}{r}}{\frac{-e^2}{2r}}=2$

101. For bracket series $n_1=4,n_2=5,6,7...$

103. On moving away from nucleus energy difference between successive orbit decrease

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104. Lower orbit to higher orbit

105. $\lambda \propto \frac{1}{\Delta n}$

108. It has no electron

109. $\lambda \propto \frac{1}{\Delta n}$

110. $R = \frac{2\pi^2 m Z^2 e^4}{Ch^3}, \quad R \propto m$

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

113. $V_n = \frac{v_1}{n}$

114. $E_2 - E_1 = 10.2ev$,so electron is excited to second orbit ,therefore no.of spectral lines $= \frac{n(n-1)}{2}$

116. Energy of given radiation

$= (E) = \frac{hc}{\lambda} = 12.02ev$

$E = \frac{-13.6}{n^2}(e.V) \quad n = 3 \quad \Delta E = \frac{hc}{\lambda}$

$a)\lambda_{1\ induced} = \frac{hc}{E_3 - E_1}$

$b)\lambda_{2\ induced} = \frac{hc}{E_3 - E_2}$

$c)\lambda_{2\ induced} = \frac{hc}{E_2 - E_1}$

117. $V \propto \frac{1}{n}$

118. $mvr = \frac{nh}{2\pi}$, b/n two successive orbits

$n_2 - n_1 = 1$

122. $E = \frac{hc}{\lambda}; \frac{1}{\lambda} = R[\frac{1}{n_1^2} - \frac{1}{n_2^2}]$

123. $\bar{v} = \bar{v}_H \times z^2$

124. $\Delta E = Rhc\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$

125. $E_{ionisation} = -E_1$

$E_1 = -x \qquad E_2 = -\frac{x}{4} \qquad E_3 = -\frac{x}{9}$

$\Delta E = \frac{-x}{9} - \frac{x}{4} = \frac{x}{4} - \frac{x}{9} = \frac{5x}{36}$

126. Number of spectral lines $= \frac{n(n-1)}{2}$

127. $\bar{v} = \frac{1}{\lambda} = R[\frac{1}{n_1^2} - \frac{1}{n_2^2}]$

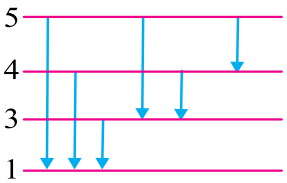
128. $E\alpha \frac{Z^2}{n^2}$

129. $\bar{v} = \frac{1}{\lambda} = R[\frac{1}{n_1^2} - \frac{1}{n_2^2}]$

130. For any series the first line have lowest wave length and highest energy

131. $C = 3\times10^{10} cm/\sec;$

$V = 2.18\times10^{8} cm/\sec;$

132. 

133. $r_n = 0.529\times n^2 A^0$; diameter = 2r ;

$r = \frac{diameter}{2}$

134. $r_n = n^2 r_1;$

$mv_n r_n = \frac{nh}{2\pi}$

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

135. Ionization energy

$-217.6 = -13.6\times\frac{Z^2}{1^2}; Z = 4m$

So, it is ${}^{9}_{4}Be^{3+}$; no. of neutrons 9 - 4 = 5

136. $[IE]_z = [IE]_H \frac{z^2}{n^2}$

137. No. of carbon items that can be placed in a straight line $= \frac{length}{diameter}$

138. $E_n = \frac{-313.52}{n^2} \times Z^2$

139. $\frac{1}{\lambda} = R[\frac{1}{n_1^2} - \frac{1}{n_2^2}] \Rightarrow \frac{15R}{16} = R[\frac{1}{1^2} - \frac{1}{n_2^2}]$

140. No.of atoms $= n \times N_0 \times \frac{\%}{100}$

141. Balmer series - visible region,for the line of balmer series energy is maximum

142. Balmer series - visible region for 3^{rd} line

$n_1 = 2,\ n_2 = 5$

143. $\frac{E_1}{3^2} - \frac{E_1}{2^2} = E \Rightarrow \frac{-5E_1}{36} = E$

$E_1 = -7.2E\ ;\ I.E = +E_1$

$= +7.2\ E$

144. $r_1 - r_2 = 0.529\left(n_1^2 - n_2^2\right)$

145. $I.E = -E_1;\ E_n \propto \frac{z^2}{n^2} \Rightarrow \frac{(E_1)_{He^+}}{(E_1)_{Li^{+2}}} = \frac{(z_{He^+})^2}{(z_{Li^{+2}})^2}$

146. $\Delta E = E_3 - E_1 = 12.1eV$

Electron jumps from 3rd orbit to 1st orbit

$mvr = \frac{3h}{2\pi} - \frac{h}{2\pi}$

147. $\Delta E = \frac{hc}{\lambda} = 2.178\times10^{-18}(\frac{1}{n_1^2} - \frac{1}{n_2^2})$

$\frac{6.6256\times10^{-34}\times3\times10^8}{\lambda} = 2.178\times10^{-18}(\frac{1}{1^2} - \frac{1}{2^2})$

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

149. According to Kepler's law $t^2 \propto R^3$

150. $\Delta E = E_2 - E_1 \quad [\because I.E = -E_1]$

$= \frac{-1.312\times10^6}{2^2} - (\frac{-1.312\times10^6}{1})$

151. $\gamma = R_0(A)^{\gamma_3}$ γ = radius of necleus

R_0 = constant value (1.25)

A = atomic mass

$\gamma = 1.25(64)^{\gamma_3} \quad \gamma = 1.25\times4 = 5\,fm$

152. $E_n = \frac{-313.6}{n^2}$

$n^2 = \frac{-313.6}{78.4}$

$n^2 = 4$ n = 2

153. $V \times \frac{z}{n}$

155. $E_n = \frac{-13.6}{n^2} z^2 eV\ /\ atm = \frac{-13.6}{2.5} \times 1$

156. n - 1

157. $\overline{V} = \frac{1}{\lambda} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$

$= R\left(\frac{1}{1}\right) \eta = \infty;\ \lambda = \frac{1}{R} = 91nm$

158. $E_H\left(\frac{1}{n_1 2} - \frac{1}{n_2 2}\right)$

$= 1.312\times10^6\left(\frac{1}{1} - \frac{1}{4}\right)$

159. $E = \frac{N_A \times C}{X}\lambda = \frac{6.23\times10^{23}\times6.62510^{-34}\times3\times10^8}{242\times10^3}$

$\lambda = 494nm$

160. $(IE)_{H^{+2}} = \frac{1}{4}\times19.6\times10^{-18}$

171. $\lambda = \frac{h}{mv};\ \lambda \propto \frac{1}{m}$

172. $\lambda = \frac{h}{\sqrt{2meV}};\ \lambda \propto \frac{1}{\sqrt{eV}}$

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173. $\lambda = V$; $\therefore \lambda = \frac{h}{mV}$; $\lambda = \frac{h}{m \times \lambda}$

$\lambda^2 = \frac{h}{m}$ $\lambda = \sqrt{\frac{h}{m}}$

174. $\Delta x.\Delta p \geq \frac{h}{4\pi}$; $\Delta x^2 \geq \frac{h}{4\pi}$; $\Delta x \geq \sqrt{\frac{h}{4\pi}}$

$\Delta v \geq \frac{h}{4\pi.m.\Delta x}$; $\Delta v \geq \frac{1}{2m}\sqrt{\frac{h}{\pi}}$

175. $\lambda = \frac{h}{\sqrt{2mkE}}$

176. $\lambda = \frac{h}{mv}$

177. $\lambda = \frac{h}{mv}$

178. $\lambda = \frac{h}{mv}$; $\frac{\lambda_A}{\lambda_B} = \frac{m_B.v_B}{m_A.v_A}$

179. $\frac{1}{2}mv^2 = ev$

180. $\frac{\Delta x_e.m_e.\Delta v_{e^-}}{\Delta x_p.m_p.\Delta vp} = 1$; $\frac{\Delta v_e}{\Delta vp} = \frac{m_p}{m_e}$

181. $\Delta x.m.\Delta v = \frac{h}{4\pi}$; $\Delta x = \frac{h}{4\pi.m.\Delta v}$

182. $\frac{1}{2}mv^2 = e.v$ $V^2 = \frac{2eV}{m}$ $V = \sqrt{\frac{2eV}{m}}$

183. $\Delta x.m.\Delta v \geq \frac{h}{4\pi}$; $\Delta x \geq \frac{h}{4\pi.m.\Delta v}$

$\Delta x \geq \frac{6.6256 \times 10^{-34}}{4 \times 3.14 \times 9.1 \times 10^{-31} \times \frac{600 \times 0.005}{100}}$

184. $\lambda = \frac{h}{\sqrt{2mkE}}$

185. $\Delta x = \frac{h}{4\pi m.\Delta V}$

186. $\frac{1}{\frac{1}{2}} = \frac{m_n v_n}{m_e.v_e}$, $V_n = \frac{2 \times m_e \times v_e}{mn}$

187. $\frac{\lambda_1}{\lambda_2} = \sqrt{\frac{eV_2}{eV_1}} = \sqrt{\frac{200}{50}} = \frac{2}{1}$

188. $\lambda = \frac{n}{\sqrt{4mer}} = 6.625 \times 10^{-34}$

$\sqrt{4 \times 4 \times 1.6 \times 10^{-27} \times 1.6 \times 10^{-19} \times V}$

$= \frac{1.035 \times 10^{-11} m}{\sqrt{V}} = \frac{0.1035}{\sqrt{V}} A^o$

196. No of radial nodes $= n - l - 1$
No of nodal planes $= l$

197. No.of angular nodes $= l = 2$
No.of radial nodes $= n - l - 1 = 4 - 2 - 1 = 1$

198. No.of spherical/radial nodes = n-l -1
for s = orbital, $l = 0$.

200. No.of peaks= n –l

208. $(n - l - 1)$

209. l

210. $\psi_{321} : n = 3, l = 2, m = 1$

Angular momentum $= \frac{h}{2\pi}\sqrt{l(l+1)} = \frac{\sqrt{6}h}{2\pi}$

Spherical nodes $= 3 - 2 - 1 = 0$;
Angular node $= 2$

Sum of all the above $= \frac{\sqrt{6}h}{2\pi} + 2 = \frac{\sqrt{6}h + 4\pi}{2\pi}$

215. No of sub energy levels=n

218. For 3d orbital $l = 2$;m= -2 to+2.

220. f - orbitals has 7 orientations.

222. 2d orbital is not possible

229. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$ in each sub energy level one orbital is with m = 0

233. Higher $(n + l)$ indicates higher energy

234. The order of increase of energy can be calculated from (n+l) rule. If two orbitals have same value of (n+l) , the orbital with lower value of n will be filled first.

239. no.of electrons $= 4l + 2$

241. $\mu = \sqrt{4S(S+1)}$

242. no.of sub shell = n,no.of orbitals = n^2,
no.of electrons = $2n^2$

243. Same (n+l) value indicates same energy level.

247. n^2

248. $\sqrt{l(l+1)}.\frac{h}{2\pi}$

249. Conceptual

250. $\gamma = \frac{diameter}{2} = \frac{20}{2} = 10nm$

$\gamma = \frac{diameter}{2} = 0.529n^2A^o$

$10\times10^{-9} = 0.529\times10^{-10}n^2$

264. Half filled d-orbital is stable

267. sum of spin quantum no = $5\times\frac{1}{2}+2\left(-\frac{1}{2}\right)$

271. paramagnetic property depends upon the number of unpaired electrons ,higher the no.of unpaired electrons , higher the paramagnetic property.

276. $\mu_B = \sqrt{n(n+2)}$ $Ni^{+2} = 3d^8$
Number of unpaired electrons (n) = 2

277. $\mu = \sqrt{n(n+2)}$
n = n of unpuried electory

STUDENT EXERCISE
PRACTICE QUESTIONS

INTRODUCTION AND ATOMIC MODELS

1. An oil drop has 6.39×10^{-19} charge. What will be the number of electrons in this drop?
1) 2 2) 4 3) 8 4) 16

2. Charge of one mole of alpha particle is
1) + 2 units 2) +1 units
3) + 2 faraday 4) + 2 coulombs

3. The ratio of e/m of proton and α - particle is
1) 2 : 1 2) 1 : 2 3) 1 : 1 4) 1 : 3

4. Charge of the species with 17 protons, 18 neutrons and 18 electrons is
1) + 1 2) - 1 3) - 2 4) None

5. An isotone of $^{76}_{32}Ge$ is:
a) $^{77}_{32}Ge$ **b)** $^{77}_{33}As$ **c)** $^{77}_{34}Se$ **d)** $^{78}_{34}Se$
1) Only a and b 2) Only b and d
3) Only b and c 4) b, c and d.

6. The number of neutrons present in the deuterium isotope of hydrogen is
1) 2 2) 3 3) 5 4) 1

7. Which has highest specific charge?
1) Na^+ (A =23) 2) Mg^{2+} (A=24)
3) Al^{3+} (A =27) 4) Si^{4+} (A =28)

8. a-particles are projected towards the following metals, with the same kinetic energy. Towards which metal, the distance of closest approach is minimum?
1) $Zn(z=30)$ 2) $Cd(Z=48)$
3) $Hg(Z=80)$ 4) $Al(Z=13)$

9. The mass numbers of three isotopes of an element are 11,12,13 units.Thier percentage abundance is 80, 15, and 5 respectively.What is the atomic wieght of the element?
1)10.25 2)11.25 3)12.25 4)13.25

10. Boron has two istopes B^{10} & B^{11} whose relative abudances are 20% & 80% respectively avg.atomic wieght of Boron is?
1)10 2)11 3)10.5 4)10.8

11. If the wavelength of green light is about $5000A^0$, then the frequency of its wave is
1) 16×10^{14} sec^{-1} 2) 16×10^{-14} sec^{-1}
3) 6×10^{14} sec^{-1} 4) 6×10^{-14} sec^{-1}

12. The energy of photon of light having frequency of $3\times10^{15}s^{-1}$ is
1) $1.99\times10^{-18}J$ 2) $1.99\times10^{-17}J$
3) $1.99\times10^{-17}ergs$ 4) $1.99\times10^{-18}ergs$

13. What is the energy of photons that corresponds to a wave number of $5\times10^{-5}m^{-1}$
1) $99.384\times10^{-30}J$ 2) $993.84\times10^{-30}J$
3) $9.9384\times10^{-30}J$ 4) $0.99384\times10^{-30}J$

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14. **Suppose 10^{-17} J of energy is needed by the interior of human eye to see an object. How many photons of green light (λ =550nm) are needed to generate this minimum amount of energy?**
1) 14 2) 28 3) 39 4) 42

15. **The ratio of the energies of two different radiations whose frequencies are 3 x 10^{14} Hz and 5 x 10^{14} Hz is**
1) 3 : 5 2) 5 : 3 3) 3 : 1 4) 5 : 1

16. **Which one of the following frequency of radiation(in Hz) has a wavelength of 600nm**
1) 2×10^{13} 2) 5×10^{16}
3) 2×10^{14} 4) 5×10^{14}

ELECTROMAGNETIC THEORY
PLANK'S QUANTUM THEORY AND PHOTO ELECTRIC EFFECT

17. **Identify the incorrectly matched set from the following**

List-I	List-II
1) Wavelength	Nanometers
2) Frequency	Hertz
3) Wavenumber	m^{-1}
4) Velocity	ergs

18. **If λ_1 *and* λ_2 are the wavelength of characteristic X-rays and gamma rays respectively, then the relation between them is**
1) $\lambda_1 = \frac{1}{\lambda_2}$ 2) $\lambda_1 = \lambda_2$
3) $\lambda_1 > \lambda_2$ 4) $\lambda_1 < \lambda_2$

19. **The characteristic not associated with Planck's theory is :**
1) Radiations are associated with energy
2) The magnitude of energy associated with a quantum is proportional to frequency
3) Radiation energy is neither emitted nor absorbed continuously.
4) Radiation energy is neither emitted nor absorbed discontinuously

20. **Ultraviolet light of 6.2 eV falls on Caesium surface (work function = 1.2 eV). The kinetic energy (in electron volts) of the fastest electron emitted is approximately**
1) 5 eV 2) 4 eV 3) 3 eV 4) 2 eV

21. **Visible light photons do not show Compton effect because they**
1) Move very slowly 2) Have no momentum
3) Have very less mass
4) Have larger wavelength

22. **As the frequency of the light increases, the momentum of its Photon**
1) Increases 2) Decreases
3) Remains same 4) Cannot be predicted

23. **In photoelectric effect, the energy of the photon striking a metallic surface is 5.6 × 10^{-19} J. The kinetic energy of the ejected electrons is 12.0 × 10^{-20}J. The work function is :**
1) $6.4\times10^{-19}\,J$ 2) $6.8\times10^{-19}\,J$
3) $4.4\times10^{-19}\,J$ 4) $6.4\times10^{-20}\,J$

24. **An Electro magnetic radiation of wavelength 484nm is just sufficient to ionise a sodium atom.Calculate the ionisation energy of sodium in kJ/mol approximately?**
1) 494.5 2) 246.9 3) 989.0 4) 794.5

ATOMIC SPECTRA
HYDROGEN SPECTRUM
BOHR'S ATOMIC MODEL

25. **The line spectra of two elements are not identical because :**
1) The elements do not have the same number of neutrons
2) They have different mass numbers
3) Their outermost electrons are at different energy levels
4) All of the above.

26. **Among the first lines of Lyman, Balmer, Paschen and Brackett series in hydrogen atomic spectra which has higher energy?**
1) Lyman 2) Balmer
3) Paschen 4) Brackett

27. **When the atomic electron is at infinite distance from the nucleus, its energy is**
1) infinity 2) zero 3) negative 4) positive

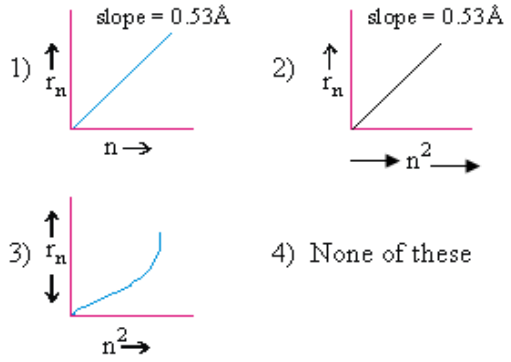
28. **The wave number of the H_α line in Balmer series of hydrogen spectrum is**
1) 5R / 36 2) 3R / 16
3) 21R / 100 4) 3R / 4

29. **The electronic transition that emits maximum energy is [n = represents orbit]**
1) $n_5 \rightarrow n_4$ 2) $n_4 \rightarrow n_3$
3) $n_2 \rightarrow n_1$ 4) $n_3 \rightarrow n_2$

30. **With increasing atomic number of a single electron species, the energy difference between two orbits**
1) Increases 2) Decreases
3) Remains constant
4) First increases followed by a decrease

31. **The ratio of the radius of the Bohr orbit for the electron orbiting the hydrogen nucleus that of the electron orbiting the deuterium nucleus is approximately**
1) 1:1 2) 1 : 2 3) 2 : 1 4) 1 : 4

32. **Which of the following curves may represent the radius of orbit (r_n) in a H-atoms as a function of principal quantum number(n)**



4) None of these

33. **How much energy is required to ionise a H atom if the electron occupies n = 5 orbit?**
1) 5.44 ev 2) 10.8 ev
3) 0.544 ev 4) 1.08 ev

34. **If the speed of electron in the Bohr's first orbit of hydrogen in x, the speed in the third orbit is**
1) x/9 2) x/3 3) 3x 4) 9x

35. **Which of the following lines will have a wave number equal in magnitude to the value of R in the H - Spectral series**
1) Limiting line of Balmer series
2) Limiting line of Lyman series
3) First line of Lyman series
4) First line of Balmer series

36. **The wave number of first line in Balmer series of Hydrogen is 15,200 cm^{-1} the wave number of first line in Balmer series of Be^{3+}**
1) $2.43 \times 10^5 cm^{-1}$ 2) $3.43 \times 10^5 cm^{-1}$
3) $4.43 \times 10^5 cm^{-1}$ 4) $5.43 \times 10^5 cm^{-1}$

37. **What transition in the hydrogen spectrum would have the same wavelength as the Balmer transition *n=4* to *n=2* of He^+ spectrum ?**
1) $n_1 = 1$; $n_2 = 2$ 2) $n_1 = 2$, $n_2 = 3$
3) $n_1 = 3$, $n_2 = 2$ 4) $n_1 = 2$, $n_2 = 4$

38. **The wave number for the longest wavelength transition in the Balmer series of atomic hydrogen is**
1) $15.2\times10^6 m^{-1}$ 2) $13.6\times10^6 m^{-1}$
3) $1.5\times10^6 m^{-1}$ 4) $1.3\times10^6 m^{-1}$.

39. **The ionization potential of hydrogen atom is 13.6 eV. The wavelength of the energy radiation required for the ionization of H-atom**
1) 1911 nm 2) 912 nm
3) 68 nm 4)91.2 nm

40. **A gas of mono atomic hydrogen is excited by an energy of 12.75 eV/atom. Which spectral lines of the following are formed in Lyman, Balmer and Paschen series respectively.**
1) 3, 2, 1 2) 2, 3, 1 3) 1, 3, 2 4)1, 2, 3

41. **The wave length of the radiation emitted by Hydrogen when compared to He^+ ion is**
1) 2 times that of He^+ ion
2) 3 times that of He^+ ion
3) 4 times that of He^+ ion 4) Same as He^+

42. **The energy of the second Bohr orbit of hydrogen atom is - 3.41 eV. The energy of the second orbit of He^+ would be**
1) – 0.85eV 2) –13.6eV
3) –1.70eV 4) -6.82eV

43. **If the diameter of carbon atom is $0.15nm$, the number of carbon atoms which can be placed side by side is a straight line across length of 10.0 cm is**
1) 66.66×10^7 2) 66.66×10^8
3) 6.2×10^9 4) 1.33×10^7

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44. The ionization energy of the ground state of hydrogen atom is 2.18×10^{-18} J. The energy of an electron in its second orbit would be

1) -1.09×10^{-18} J 2) -2.18×10^{-18} J

3) -4.36×10^{-18} J 4) -5.45×10^{-19} J

45. The velocity of an electron in the first Bohr orbit of hydrogen atom is $2.19\times10^{6}\,ms^{-1}$. Its velocity in the second orbit would be

1) $1.10\times10^{6}\,ms^{-1}$ 2) $4.38\times10^{6}\,ms^{-1}$

3) $5.5\times10^{5}\,ms^{-1}$ 4) $8.76\times10^{6}\,ms^{-1}$

46. Energy of electron moving in the second orbit of He^{+}ion is

1) -13.6ev 2) -3.4 ev

3) -1.51 ev 4) -0.84 ev

47. According to Bohr's theory of hydrogen atom

1) There is only fixed set of allowed orbitals for the electron
2) The allowed orbitals of the electrons are elliptical in shape
3) The moment of an electron from one allowed toanother allowed orbital is forbidden
4) No light is emitted as long as the electron remains in an allowed orbital

48. The ratio of radius of 2nd and 3rd Bohr orbit is

1) 3 : 2 2) 9 : 4 3) 2 : 3 4) 4 : 9

49. According to Bohr's theory, which one of the following values of angular momentum of hydrogen atom is not permitted.

1) $\frac{1.25h}{\pi}$ 2) $\frac{h}{\pi}$ 3) $\frac{1.5h}{\pi}$ 4) $\frac{0.5h}{\pi}$

KEY

1. (2)	2. (3)	3. (1)	4. (2)
5. (2)	6. (4)	7. (4)	8. (4)
9. (2)	10. (4)	11. (3)	12. (1)
13. (3)	14. (2)	15. (1)	16. (4)
17. (4)	18. (3)	19. (4)	20. (1)
21. (4)	22. (1)	23. (3)	24. (2)
25. (3)	26. (1)	27. (2)	28. (1)
29. (3)	30. (1)	31. (1)	32. (2)
33. (3)	34. (2)	35. (2)	36. (1)
37. (1)	38. (3)	39. (4)	40. (1)
41. (3)	42. (2)	43. (1)	44. (4)
45. (1)	46. (1)	47. (4)	48. (4)
49. (1)			

DUAL NATURE OF MATTER AND LIGHT DE-BROGLIE'S AND HEISENBERG UNCERTAINITY PRINCIPLE

50. If moving with equal speeds, the longest wavelength of the following matter waves is that for a (an)

1) Electron 2) α-particle

3) Proton 4) Neutron

51. The momentum of radiation of wavelength 0.33 nm is ..... kgm sec^{-1}

1) 2×10^{-24} 2) 2×10^{-12}

3) 2×10^{-6} 4) 2×10^{-48}

52. The mass of the electrons 9.8 x 10^{-28}gram and uncertainty in the velocity equal to 2x10^{-3}cm/sec. The uncertainty in the position of an electron is (h=6.62 x10^{-27}ergsec)

1) $2.9\times10^{+2}$ cm 2) 2.9×10^{-2} cm

3) $2.9\times10^{-12}\,cm^{-1}$ 4) $2.9\times10^{+12}\,cm^{-1}$

53. The velocity of an electron with de Broglie wavelength of 1.0 x 10^{2} nm is :

1) $7.2\times10^{5}\,cm/sec$ 2) $72\times10^{5}\,cm/sec$

3) $7.2\times10^{4}\,cm/sec$ 4) $3.6\times10^{5}\,cm/sec$

54. The wave length of a electron with mass $9.1\times10^{-31}kg$ and kinetic energy $3.0\times10^{-25}J$ is

1) 89.67nm 2) 8.96nm

3) 456.7nm 4) 896.7nm

55. A cricket ball of 0.5 Kg is moving with a velocity of 100m per sec. the wavelength associated with its motion is

1) 1/100 m 2) 6.6×10^{-34} m

3) 1.32×10^{-35} m 4) 6.6×10^{-28} m

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56. A microscope using suitable photons is employed to locate an electron in an atom within a distance of $0.1A^0$. What is the uncertainty involved in the measurement of its velocity?

1) $2.69\times10^{6}\,ms^{-1}$ 2) $5.79\times10^{5}\,ms^{-1}$
3) $5.79\times10^{6}\,ms^{-1}$ 4) $4.62\times10^{6}\,ms^{-1}$

57. The mass of photon with wave length $3.6A^0$ is

1) $6.135\times10^{-33}\,kg$ 2) $6.135\times10^{-27}\,kg$
3) $4.126\times10^{-29}\,kg$ 4) $4.126\times10^{-25}\,kg$

58. If the velocity of electron in Bohr's first orbit is $2.19\times10^{6}\,ms^{-1}$. The de-Broglie's wavelength is

1) 332 pm 2) 313 pm
3) 3.32 pm 4) 3.13 pm

59. Uncertainity in position of a particle of 25 gram in space is 10^{-5}m. Hence uncertainity in velocity (m/sec) is

($h = 6.6\times10^{-34}\,J$ - sec)

1) 2.1×10^{-28} 2) 2.1×10^{-34}
3) 0.5×10^{-34} 4) 5×10^{-24}

60. An electron, a proton and an alpha particle have K.E. of 16E, 4E and E respectively. What is the qualitative order of their de-Broglie wavelengths:

1) $\lambda_e > \lambda_p > \lambda_\alpha$ 2) $\lambda_p = \lambda_\alpha > \lambda_e$
3) $\lambda_p < \lambda_c < \lambda_\alpha$ 4) $\lambda_\alpha < \lambda_e = \lambda_p$

61. The wavelengths of electron waves in two orbits is 3 : 5. The ratio of kinetic energy of electrons will be

1) 25:9 2) 5:3 3) 9:25 4) 3:5

QUANTUM MECHANICAL MODEL OF ATOM

62. Which of the following statements is not correct?

1) The wave function depicting the dependence on r involves two quantum numbers n and l.
2) The wave function depicting the angular dependence involves two quantum numbers l and m.
3) The spin quantum number is not the outcome of the Schrodinger equation.
4) The lowest energy state of an atom corresponds to n = 0

63. In a main energy level, the orbital with more number of nodal planes will be .....

1) Higher energy 2) Lower energy
3) Either 1 or 2 4) Neither 1 nor 2

64. Choose the correct statement among the following :

1) ψ^2 represents the atomic orbital
2) The number of peaks in radial distribution is n-1
3) A node is a point in space around nucleus where the wave function ψ has zero value
4) All of the these

65. The maximum probability of finding electron in the d_{xy} orbital is

1) Along with x - axis
2) Along the y - axis
3) At an angle of 45^0 from the X and Y axis
4) At an angle of 90^0 from the x and y axis

66. Which of the following statements regarding an orbital are correct

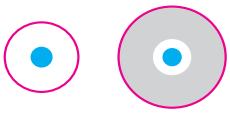
1) An orbital is a definite trajectory around the nucleus in which electron can move
2) An orbital always has spherical trajectory
3) An orbital is the region around the nucleus where there is a 90 – 95% probability of finding all the electrons of an atom
4) An orbital is characterized by 3 quantum numbers n, l and m

67. Which of the following statements on the atomic wave function ψ is not correct?

1) ψ may be a real valued wave function
2) ψ may be in some cases be a complex function
3) ψ has a mathematical significance only
4) ψ is proportional to the probability of finding an electron

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68. The probability density plots of 1s and 2s orbitals are given in figure



1s 2s

The density of dots in region represents the probability density of finding electrons in the region.On the basis of above diagram which of the following statements is incorrect?

1) 1s and 2s orbitals are spherical in shape
2) The probability of finding the electron is maximum near the nucleus.
3) The probability of finding the electron at a given distance is equal in all directions.
4) The probability density of electrons for 2s orbitals decreases uniformly as distance from the nucleus increases.

QUANTUM NUMBERS

69. What is the full degeneracy of the $n = 3$ state of a H-atom in the absence of a magnetic field?

1) 4 2) 10 3) 8 4) 18

70. For the azimuthal quantum number 'l' the total number of magnetic quantum numbers is given by

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

3) $l = \frac{2m+1}{2}$ 4) $l = \frac{2m-1}{2}$

71. How many sets of four quantum numbers are possible for electrons present in He^{2-} anion

1) 2 2) 4 3) 5 4) 7

72. The set of quantum numbers n = 2, $l = 2$, $m_l = 0$

1) Describes an electron in a 2*s* orbital
2) Describes one of the five orbitals of a similar type
3) Describes an electron in a 2*p* orbital
4) Is not allowed.

73. The sub-energy level which can accommodate maximum number of electrons with parallel spin values is

1) 4p 2) 6s 3) 3d 4) 6p

74. The azimuthal quantum number and the principal quantum number of the 17th electron are

1) $l = 1$, n = 3 2) $l = 3$, n = 2
3) $l = 1$, n = 17 4) $l = 2$, n = 1

75. The lowest orbital in which an electron with Azimuthal quantum no. value 3 is

1) 4 2) 5 3) 1 4) 6

76. The quantum numbers n = 3, $l = 1$, m = +1 and s = +1/2 represent the unpaired electron present in

1) Sodium atom 2) Aluminium atom
3) Fluorine atom 4) Potassium atom

77. The magnetic quantum number m for the outermost electron in the Na atom,is

1)0 2)2 3)3 4)1

78. Which of the following quantum numbers is correct for an electron in 4f-orbital

1) n = 4, $l = 3$, m = +1, s = +1/2
2) n = 4, $l = 3$, m = +4, s = +1/2
3) n = 4, $l = 4$, m = +1, s = +1/2
4) n = 4, $l = 2$, m = –2, s = +1/2

79. The maximum number of electrons with spin value +1/2 in the orbital with azimuthal quantum number value $l = 2$ and magnetic quantum number m = +2 is .....

1) 5 2) 6 3) 3 4) 1

80. Which one of the following set of quantum numbers is not possible for a 4p electron?

1) n =4, $l = 1$, m = +1, $m_s = +\frac{1}{2}$

2) n =4, $l = 1$, m = 0, $m_s = +\frac{1}{2}$

3) n = 4, $l = 1$, m = 2, $m_s = +\frac{1}{2}$

4) n = 4, $l = 1$, m = -1, $m_s = +\frac{1}{2}$

81. The total number of electrons present in all the s orbitals, all the p orbitals and all the d orbitals of cesium ion are respectively.

1) 6, 26, 10 2) 10,24,20
3) 8, 22, 24 4) 12, 20, 23

82. The quantum numbers + 1/2 and - 1/2 for the electron spin represent

1) rotation of electron in clockwise and anti clockwise direction respectively
2) rotation of electron in anti-clockwise and clockwise direction respectively
3) magnetic moment of the electron pointing up and down respectively
4) two quantum mechanical spin states which have no classical analogue

83. The correct set of quantum numbers for the unpaired electron of Chlorine atom

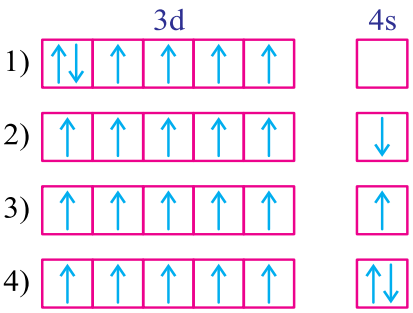
1) 2,0,0,+1/2 2) 2,1,-1,+1/2
3) 3,0,0,+1/2 4) 3,1,-1, $\pm$ 1/2

84. The quantum number which explain the line spectra observed as doublets in case of hydrogen and alkali metals and doublets & triplets in case of alkaline earth metals is

1) Spin 2) Azimuthal
3) Magnetic 4) Principal

ELECTRONIC CONFIGURATION

85. Which of the following arrangements of electrons is most likely to be stable?



86. Aufbau principle fails to explain the configuration of element with atomic number

1) 18 2) 21 3) 24 4) 27

87. Total number of electrons in any energy level is

1) $\sum_{l=0}^{l=n-1} 2(2l+1)$ 2) $\sum_{l=0}^{l=n} 2(2l+1)$

3) $\sum_{l=0}^{l=n+1} 2(2l+1)$ 4) $\sum_{l=1}^{l=n-1} 2(2l+1)$

88. The atomic number at which filling of a g-orbital is likely to begin is :

1) 121 2) 116 3) 106 4) 124

89. n and *l* values of an orbital A are 3 and 2 and another orbital B are 5 and 0. The energy of

1) B is more than A 2) A is more than B
3) A and B are same
4) A is four times than B

90. Number of unpaired electrons of neutral manganese atom and its divalent ion are in the ratio (the atomic number of manganese is 25 and it loses two electrons to form the divalent ion)

1) 1 : 1 2) 25 : 23 3) 5 : 3 4) 3 : 53

91. Which of the following electrons is most tightly bound by the nucleus

1) 4p 2) 5s 3) 4d 4) 5d

92. $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5$ is not the electron configuration of

1) Mn^{3+} 2) Fe^{3+} 3) Cr^{+} 4) Co^{4+}

93. Which one of the following statements is correct?

1) 2's' orbital is spherical with two nodal planes
2) The de Broglie wavelength (λ) of a particle of mass 'm' and velocity 'V' is equal to mV/h
3) The principal quantum number (n) indicates the shape of the orbital
4) The electronic configuration of phosphorous is given by $[Ne]\ 3s^2\ 3p^1_x\ 3p^1_y\ 3p^1_z$

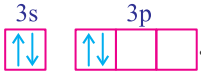
94. Which of the following has maximum number of unpaired electrons ?

1) Zn 2) Fe^{2+} 3) Ni^{3+} 4) Cu^{+}

95. The successive elements belonging to the 3d-series have the same number of electrons in the d-sub-shell. The elements are

1) Ti & V 2) V & C
3) Cr & Mn 4) Mn & F

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96. The electronic configuration in the valence shell of silicon is 

. The rule violated is

1) Auf-bau principle 2) Pauli's rule
3) Hund's rule 4) All

97. In potassium, the order of energy levels is

1) 3s > 3d 2) 4s < 3d
3) 4s > 4p 4) 4s = 3d

98. An element has 2 electrons in K shell, 8 electrons in L shell, 13 electrons in M shell and one electron in N shell. The element is

1) *Cr* 2) *Fe* 3) *V* 4) *Ti*

99. A compound Vanadium has a magnetic moment of 1.73BM. The electronic configuration ofVanadium ion in the compound is_____.

1) $[Ar]3d^2$ 2) $[Ar]3d^1 4s^0$
3) $[Ar]3d^3$ 4) $[Ar]3d^0 4s^1$

100. A transition metal 'X' has a configuration *[Ar] 3d*^{4} in its +3 oxidation state, its atomic number is

1) 25 2) 26 3) 22 4) 19

101. Which one of the following ions has same number of unpaired electrons as those present in V^{+3} ion.

1) Fe^{+3} 2) Ni^{+2} 3) Mn^{+2} 4) Cr^{+3}

KEY

50. (1)	51. (1)	52. (1)	53. (1)
54. (4)	55. (3)	56. (3)	57. (1)
58. (1)	59. (1)	60. (1)	61. (1)
62. (4)	63. (1)	64. (4)	65. (3)
66. (3)	67. (4)	68. (4)	69. (4)
70. (2)	71. (2)	72. (4)	73. (3)
74. (1)	75. (1)	76. (2)	77. (1)
78. (1)	79. (4)	80. (3)	81. (2)
82. (4)	83. (4)	84. (1)	85. (3)
86. (3)	87. (1)	88. (1)	89. (1)
90. (1)	91. (1)	92. (1)	93. (4)
94. (2)	95. (3)	96. (3)	97. (2)
98. (1)	99. (2)	100. (1)	101. (2)

HINTS & SOLUTIONS

1. q = ne
2. 1 mole of positive charge = 1F
3. specific charge = e/m
4. 1 electron is more

7. $e/m = \frac{magnitude\ of\ charge}{mass\ number}$

9. $Avg.at.wt = \frac{\sum \%\,abundance \times Atomic\,weight}{Total\,ratio}$

10. $Avg.at.wt = \frac{\sum \%\,abundance \times Atomic\,weight}{Total\,ratio}$

11. $\nu = \frac{c}{\lambda}$

12. $E = h\nu$

13. $E = hc\bar{\nu}$

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

15. $\frac{E_1}{E_2} = \frac{\nu_1}{\nu_2}$

16. $\nu = \frac{c}{\lambda}$

20. K.E = E-W

23. $E = w_0 + K.E$

24. $E = \frac{hc}{\lambda} = E \times N_A$

27. $E_n = \frac{-2\pi^2 mZ^2 e^4}{n^2 h^2}$; *if* $n = \infty$ *then* $E = 0$

28. $\bar{\nu} = \frac{1}{\lambda} = R[\frac{1}{n_1^2} - \frac{1}{n_2^2}]$

29. The energy difference b/n first and second orbits is more.

30. $\Delta E = RhC\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] \times Z^2$

$\Delta E \propto Z^2$

31. $r \propto \frac{n^2}{z}$

32. $r \propto \frac{n^2}{z}$

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33. $-E_n = I.E = \frac{13.6}{n^2} \times z^2$

34. $V_n = \frac{V_1}{n}$

35. $\bar{v} = R[\frac{1}{n_1^2} - \frac{1}{n_2^2}]$

36. $\bar{v} = \bar{v}_H \times Z^2$

37. $R \times Z^2 \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$

38. $\bar{v} = \frac{1}{\lambda} = R_H \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$

39. $\Delta E = \frac{hc}{\lambda}$

40. $\Delta E = E_4 - E_1 = 12.75 ev/atom$,then the electron is excited to fourth orbit.

41. $\frac{1}{\lambda} \propto Z^2$

42. $\frac{E_1}{E_2} = \frac{Z_1^2}{Z_2^2}$

43. No. of carbon atoms items that can be placed in a straight line $= \frac{length}{diameter}$

44. $E_n = \frac{-E_1}{n^2}$

45. $v_n = \frac{v_1}{n}$

46. $E_n = \frac{-13.6}{n^2} \times Z^2$

48. $\frac{r_1}{r_2} = \frac{n_1^2}{n_2^2}$

49. $mvr = \frac{nh}{2\pi}$ *where* $n = 1, 2, 3...$

50. $\lambda \propto \frac{1}{m}$

51. $p = \frac{h}{\lambda}$

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

53. $\lambda = \frac{h}{mv}$

54 $\lambda = \frac{h}{\sqrt{2mkE}}$

55. $\lambda = \frac{h}{mv}$

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

57 $\lambda = \frac{h}{mv}$

58. $\lambda = \frac{h}{mv}$

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

60. $\lambda = \frac{h}{\sqrt{2mkE}}$

61 $\frac{\lambda_1}{\lambda_2} = \sqrt{\frac{KE_2}{KE_1}}$

62. n=0 is not possible.

63. The orbital with more no.of nodal planes will have more energy.

70. $m = 2l + 1$

71. no.of sets = no.of electrons

73. 3d has 5 parallel spin values.

89. If $n + l$ is same, small n value indicate lower energy level

99. Magnetic momentum = $\sqrt{n(n+2)}$ BM

WORK AT HOME

PREVIOUS EXAMS QUESTIONS

1. Given, the mass of electron is 9.11×10^{-31} kg and planck's constant is 6.626×10^{-34} Js. The uncertainty involved in the measurement of velocity within a distance of $0.1\,\overset{0}{A}$ is : [AIPMT 2006]

1) $6.65 \times 10^{-35}\, m$ 2) $5.79 \times 10^{5}\, ms^{-1}$
3) $5.79 \times 10^{6}\, ms^{-1}$ 4) $5.79 \times 10^{7}\, ms^{-1}$

2. The number of possible orientations of an electron in an atom is given by its: [AIPMT 2006,07]

1) spin quantum number
2) spin angular momentum
3) magnetic quantum number
4) orbital angular momentum

3. The de Broglie wavelength of a particle of mass 1 g and velocity $100 ms^{-1}$ is [AFMC 2006]

1) $6.63 \times 10^{-33}\, m$ 2) $6.63 \times 10^{-34}\, m$
3) $6.63 \times 10^{-35}\, m$ 4) $6.65 \times 10^{-35}\, m$

4. The de Broglie wavelength asociated with a ball of mass 1 kg having kinetic energy 0.5J is: [AIIMS 2006]

1) $6.626 \times 10^{-34}\, m$ 2) $13.20 \times 10^{-34}\, m$
3) $10.38 \times 10^{-21}\, m$ 4) $6.626 \times 10^{-34}\, \overset{0}{A}$

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5. Consider the following sets of quantum numbers:

	n	l	m	s
i)	3	0	0	+1/2
ii)	2	2	1	+1/2
iii)	4	3	-2	-1/2
iv)	1	0	-1	-1/2
v)	3	2	3	+1/2

Which of the sets of quantum numbers is not possible? [AIPMT 2007]

1) i, ii, iii and iv 2) ii, iv and v
3) i and iii 4) ii, iii and iv

6. The measurement of electron position is associated with an uncertainty in momentum which is equal to $1\times10^{-18}\ g\ cm\ s^{-1}$. The uncertainty of electron velocity is: (Mass of electron is $9\times10^{-28}\ g$) [AIPMT 2007]

1) $1\times10^{8}\ cm\ s^{-1}$ 2) $1\times10^{11}\ cm\ s^{-1}$
3) $1\times10^{9}\ cm\ s^{-1}$ 4) $1\times10^{6}\ cm\ s^{-1}$

7. The ground state energy of H-atom is -13.6eV. When its electron is in the 1^{st} excited state, its excitation energy is: [AIPMT 2008]

1) 0 eV 2) 3.4 eV
3) 6.8 eV 4) 10.2 eV

8. Two nuclei have their mass number in the ration 1:3. The ratio of their nucleur densities would be : [AIPMT 2008]

1) 1 : 1 2) 1 : 3 3) 3 : 1 4) $(3)^{1/3}$:1

9. If uncertainty in position and momentum are equal, then uncertainty in velocity is : [AIPMT 2008]

1) $\frac{1}{m}\sqrt{\frac{h}{\pi}}$ 2) $\sqrt{\frac{h}{\pi}}$ 3) $\frac{1}{2m}\sqrt{\frac{h}{\pi}}$ 4) $\sqrt{\frac{h}{2\pi}}$

10. The uncertainties in the velocities of two particles A and B are 0.05 and 0.02 ms^{-1} respectively. The mass of B is five times to that of the mass of A. What is the ratio of uncertainties $\frac{\Delta x_A}{\Delta x_B}$ in their positions? [AIPMT 2008]

1) 2 2) 0.25 3) 0.4 4) 1

11. The de Brogllie wavelength of a particle with mann 1 kg and velocity 100 ms^{-1} is : [JIPMER 2008]

1) $6.6\times10^{-33}m$ 2) $6.6\times10^{-36}m$
3) $3.3\times10^{+33}m$ 4) $3.3\times10^{-36}m$

12. The maximum number of electrons in a subshell of an atom is determined by the following : [AIPMT 2009]

1) $4l+2$ 2) $2l+1$ 3) $4l-2$ 4) $2n^2$

13. Which of the following sets of quantum number represent an impossible arrangement? [AIPMT 2009]

	n	l	m	s
1)	3	2	-2	$+\frac{1}{2}$
2)	4	0	0	$-\frac{1}{2}$
3)	3	2	-3	$+\frac{1}{2}$
4)	5	3	0	$-\frac{1}{2}$

14. The energy absorbed by each molecule A_2 of a substance is 4.4×10^{-19} J and bond energy per molecule is 4.0×10^{-19} J. The kinetic energy of the molecule per atom will be : [AIPMT 2009]

1) $2.0\times10^{-20}\ J$ 2) $2.2\times10^{-19}\ J$
3) $2.0\times10^{-19}\ J$ 4) $4.0\times10^{-20}\ J$

15. The energy of electron in first Bohr's orbit of H-atom is -13.6 eV. The possible energy value of the excited state for electron in Bohr's orbits of hydrogen is: [JIPMER 2007]

1) -6.8 eV 2) - 4.2 eV
3) -3.4 eV 4) +6.8 eV

16. which of the following ions has electronic configuration $[Ar]3d^6$? [AIPMT 2010]

1) Ni^{3+} 2) Mn^{3+} 3) Fe^{3+} 4) Co^{3+}

17. The energies E_1 *and* E_2 of two radiations are 25 eVand 50 eV respectively. The relation between their wavelengths, i.e, λ_1 *and* λ_2 will be: [AIPMT 2011]

1) $\lambda_1 = \frac{1}{2}\lambda_2$ 2) $\lambda_1 = \lambda_2$

3) $\lambda_1 = 2\lambda_2$ 4) $\lambda_1 = 4\lambda_2$

18. According to Bhor theory, which of the following transitions in the hydrogen atom will give rise to the least energetic photon? [AIPMT 2011]

1) n = 6 to n =1 2) n = 5 to n = 4
3) n = 6 to n = 5 4) n = 5 to n = 3

19. Total number of atomic orbitals in 4th energy level of an atom is : [AIPMT 2011]

1) 4 2) 8 3) 16 4) 32

20. If n = 6, the correct sequence for filling of electrons will be [AIPMT 2011]

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

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

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

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

21. The orbital angular mimentum of ap-electron is given as: [AIPMT 2012]

1) $\sqrt{6}\times\frac{h}{2\pi}$ 2) $\frac{h}{\sqrt{2}\pi}$ 3) $\sqrt{3}\times\frac{h}{2\pi}$ 4) $\sqrt{\frac{3}{2}}\times\frac{h}{\pi}$

22. The correct set of four quantum numbers for the valence electron of rubedium atom (Z=37) is: [AIPMT 2012]

1) 5, 1, 1, $+\frac{1}{2}$ 2) 6, 0, 0, $+\frac{1}{2}$

3) 5, 0, 0, $+\frac{1}{2}$ 4) 5, 1, 0, $+\frac{1}{2}$

23. Maximum number of electrons in a subshell with $l = 3$ and n =4 is : [AIPMT 2012]

1) 14 2) 16 3) 10 4) 12

24. Which is maximum number 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) 12

25. Based on the equation,

$E = -2.178\times10^{-18} J\left(\frac{n^2}{Z^2}\right)$, certain conclusions are written. Which of these is not correct? [NEET 2013]

1) Equation can be used to calculate the change in energy when the electron changes orbital.
2) For n = 1, the electron has a more negativ eenrgy than it does for n = 6, which means that the electron is more loosely bound in the smallest allowed orbital.
3) The negative sign simply means that the energy of electron bound to the nucleus is lower than it would be if the electron were at the infinite distance from the nucleus.
4) Larger tha value of n, larger is the orbit radius.

26. The value of planck's constant is 6.63×10^{-34} Js. The speed of light is $3\times10^{17}\ nms^{-1}$. Which value is closest to the wavelength in nanometer of a quantum of light with frequency $6\times10^{15}\ s^{-1}$? [NEET 2013]

1) 50 2) 75 3) 10 4) 25

27. What is the maximum number of orbitals that can be identified with the following quantum numbers? $(n = 3, l = 1, m_e = 0)$. [AIPMT 2014]

1) 3 2) 4 3) 1 4) 2

28. Calculate the energy in joule corresponding to light of wavelength 45 nm: (Planck's constant $h = 6.63\times10^{-34}$ js; speed of light $c = 3\times10^{8}\ ms^{-1}$) [AIPMT 2014]

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

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29. **The number of d-electrons in Fe^{2+} (Z=26) is not equal to the number of electrons in which one of the following?[AIPMT 2015]**

1) p-electrons in Cl(Z=17)
2) d-electrons in Fe(Z=26)
3) p-electrons in Ne(Z=10)
4) s - electrons in Mg (Z=12)

30. **The angular momentum of electron in d subshell is equal to : [AIPMT 2015]**

1) $\sqrt{2}\hbar$ 2) $2\sqrt{3}\hbar$ 3) $0\hbar$ 4) $\sqrt{6}\hbar$

31. **Two electrons occupying the same orbital are distinguished by [NEET 2016]**

1) Principal quantumnumber
2) Magnetic quantumnumber
3) Azimuthal quantumnumber
4) Spin quantumnumber

32. **Which of the following pairs of d-orbitals will have electron density along the axes? [NEET - 2016]**

1) d_{z^2}, d_{xz} 2) d_{xz}, d_{yz}

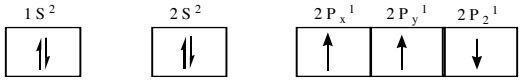
3) $d_{z^2}, d_{x^2-y^2}$ 4) $d_{xy}, d_{x^2-y^2}$

33. **Which one the wrong staement? [NEET-2017]**

1) The uncertainty principle is $\Delta E \times \Delta t \, z \frac{h}{4\pi}$
2) Half filleld & fully filled orbitals have greater stability due to greater exchange energy, greater symmetry and more balanced arrangement.
3) The energy of 2s - orbital is less then 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 = velocity of the particle

34. **Which one is a wrong statement? [NEET-2018]**

1) Total orbital angular momentum of electron in s-orbital in equal to zero.
2) An orbital is designated by 3 quantum numbers while an e- in an atom is designated by 4 quantum numbers.
3) The electonic configuration of N atom is



4) The value of m for d_{z^2} is zero.

35. **4d, 5p, 5f and 6p orbitals are arranged in order of decreasing energy. the correct option is.. [NEET-2019]**

1) 6p > 5f > 4d > 5p
2) 5f > 6p > 4d > 5p
3) 5f > 6p > 5p > 4d
4) 6p > 5f > 5p > 4d

36. **Which of the following series of transitions in the spectrum of hydrogen atom falls in visible regeon? [NEET-2019]**

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

KEY

1. (3)	2. (3)	3. (1)	4. (1)
5. (2)	6. (3)	7. (4)	8. (1)
9. (3)	10. (1)	11. (2)	12. (1)
13. (3)	14. (1)	15. (3)	16. (4)
17. (3)	18. (3)	19. (3)	20. (2)
21. (2)	22. (3)	23. (1)	24. (2)
25. (2)	26. (1)	27. (3)	28. (2)
29. (1)	30. (4)	31. (4)	32. (3)
33) 3	34) 3	35) 3	36) 4

HINTS & SOLUTIONS

1. $\Delta x.\Delta p \geq \frac{h}{2\pi}$

3. $\lambda = \frac{h}{mv}$

4. $\lambda = \frac{h}{\sqrt{2KE}}$

6. $\Delta x.\Delta p \geq \frac{h}{2\pi}$

7. $E_n = \frac{-13.6 \times z^2}{n^2}$

9. $\Delta x.\Delta p \geq \frac{h}{2\pi}$

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10. $\frac{\Delta x_A}{\Delta x_B} = \frac{m_B \Delta V_B}{m_A \Delta V_A}$ 11. $\lambda = \frac{h}{mv}$

15. $E_n = \frac{-13.6 \times z^2}{n^2}$ 17. $\frac{E_1}{E_2} = \frac{\lambda_2}{\lambda_1}$

18. low(n+l)has lower energy

19. n^2

21. $\sqrt{l(l+1)}\frac{h}{2\pi}$

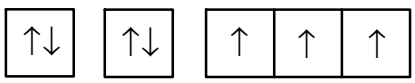
23. $l = 3$ is f sub energy level

24. any orbital has only two electrons

26. $E = h.\frac{c}{\lambda}$ 28. $E = h.\frac{c}{\lambda}$

33. In case of hydrogen like atoms, energy depends on the principle quantum number only. Hence, 2s-orbital will have energy equal to 2p-orbital.

34. According to Hund's rule of max. Multiplicity, the correct configuration of N is,



$1s^2$ $2s^2$ $2p_x^1$ $2p_y^1$ $2p_z^1$

35. $5f > 6p > 5p > 4d$

36. Balmer series of transitions in the spectron of hydrogen atom falls in visible region.

MULTI MODEL QUESTIONS

In the questions that follows two statements are given. Reason is supported to be the explanation for Assertion. Study both the statements and then mark your answers. according to the codes given below. Mark your answer as

1) Both (A) and (R) are true and (R) is the correct explanation of (A)
2) Both (A) and (R) are true and (R) is not the correct explanation of (A)
3) (A) is true but (R) is false
4) (A) is false but (R) is true

1. **Assertion(A):** The energy of ultraviolet radiation is greater than the energy of infrared radiation
Reason(R): The velocity of ultraviolet radiation is greater than the velocity of infrared radiation.

2. **A :** Fe^{+3} ion is more stable than Fe^{+2}
R : Fe^{+3} ion has more number of unpaired electrons than Fe^{+2}

3. **A :** The kinetic energy of the photo electron ejected increases with increase in intensity of incident light.
R: Increase in intensity of incident light increases the rate of emission.

4 **A :** Threshold frequency is a characteristic for a metal
R : Threshold frequency is a maximum frequency required for the ejection of electron from the metal surface.

5. **A :** Line spectrum of Li^{+2} and He^{+} are identical
R : Isoelectronic species produce identical spectrum

6. **A :** Hydrogen has only one electron in its orbit but produces several spectral lines.
R : There are many excited energy levels available in a sample of Hydrogen gas.

7. **A :** It is not essential that all the lines available in the emission spectrum will also be available in the absorption spectrum
R : The spectrum of hydrogen atom is only absorption spectrum

8. **A :** In an atom, the velocity of electrons in the higher orbits keeps on decreasing
R : Velocity of electron is inversely proportional to the radius of the orbit

9. **A :** The radial probability of 1s electron first increases, till it is maximum at $0.53A^0$, and then decreases.
R : Bohr's radius for the first orbit is $0.53A^0$

10. **A :** Wavelength of limiting line of lyman series is less than wavelength of limiting line of Balmer series.
R : Rydberg constant value is same for all elements

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11. **A :** The faster a particle moves, the greater its momentum and the shorter is the wave length that is associated with it.
R : Because $\lambda = hp^{-1}$ and $p = mv$

12. **A :** An electron cannot exist in the nucleus
R : The deBroglie wavelength of an electron is much smaller than the diameter of the nucleus

13. **A :** The position of electron can be determined with the help of an electronic microscope
R : The product of uncertainty in momentum and the uncertainty in the position of an electron cannot be less than a finite limit

14. **A :** It is not possible to predict position and the velocity of an electron exactly and simultaneously
R : Electron moving with high speed possesses both the particle nature and the wave nature

15. **A :** deBroglie equation has significance for any microscopic or submicroscopic particles
R : deBroglie wavelength is inversely proportional to the mass of the object.

16. **A :** There are two nodal regions in 3s-orbital.
R : There is no nodal plane in 3s-orbital.

17. **A :** A spectral line will be observed for a $2p_x$–$2p_y$ transition
R : The energy of $2p_x$ and $2p_y$ orbitals is the same

18. **A :** The P_x orbital has maximum electron density along the x axis and its nodal plane is yz plane
R : For a given atom, for all values of n, the p-orbitals have the same shape, but the overall size increase as n increases

19. **A :** Electrons may be considered as particles and waves
R : An electron in an atom may be described as occupying at atomic orbital or by a wave function ψ, which is a solution to the schrodinger wave equation

20. **A :** ψ indicates the amplitude of electron - wave
R : ψ^2 denotes probability of finding an electron in the space around the nucleus

21. **A :** The electronic configurations in which all of the orbitals of the same sub shell are either completely filled or are exactly half filled are more stable
R : The completely filled or exactly half filled shells possess a symmetrical distribution of electrons and allow their maximum number of exchanges

22. **A :** An orbital cannot have more than two electrons
R : The two electrons in an orbital create opposite magnetic field

23. **A :** P – orbital can accomadate 6 electrons
R : No two electrons in atom can have same set of four quantum numbers

24. **A :** Ground state electronic configuration of chromium atom is (Ar) $4s^1 3d^5$
R : Exchange energy is more wtih(*Ar*) $4s^1 3d^5$than (Ar) $4s^2 3d^4$

25. **A :** nodal plane of p_x atomic orbital is yz plane
R : In p_x atomic orbital election density is zero in the yz plane

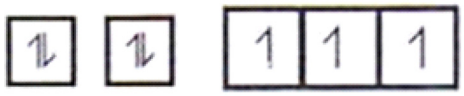
26. **A :** No two electrons in an atom can have the same values of four quantum numbers,
R : No two electrons in an atom can be simultaneously in the same shell, same subshell same orbital's and have same spin

27. **A :** p-orbital has dumb-bell shape
R : Electrons present in p-orbtial can have one of three values for 'm' i.e. +1,-1, 0

28. **A :** The ground state configuration of Cr is $3d^5 4s^1$
R : A set of exactly half filled obitals containing parallel spin arrangement provide extra stability

29. **A :** Limiting line the balmer series has a wavelength of 36.4 m.
R : Limiting lines is obtained for a jump of electron form $n=\infty$ to $n=2$ for Balamer series

30. **A :** The electronic configuration of nitrogen atom is represented as



Not as



R : The configuration of ground state of an atom is the one which has the greatest multiplicity

31. **A :** The configuration of B atom cannot be $1s^2\ 2s^3$

R : Hund's rule demands that the configuration should display maximum multiplicity

32. **A :** 2p orbital do not have spherical nodes

R : The number of spherical nodes in p-orbital's is given by (n-2)

33. **A:** in Rutheford's gold foil experiment, very few α - particles are delfected back

R: Nucleus present inside the atom is heavy

34. **A:** The are two spherical nodes in 3s-orbital

R: There is no angular node in 3s-orbtial

Matrix Match Type

Following questions contains statements given in two columns. Which have to be matched. Statements in column I have to be matched with statements in column II

35.

List - I	List - II
A) Velocity of light	1) Energy particle
B) Plank's constant	2) Energy packet
C) Wave number	3) 3×10^8m / sec
D) Photon	4) 6.625×10^{-34}J–sec
	5) cm^{-1}

The correct match is

	A	B	C	D
1)	2	3	4	5
2)	1	2	3	4
3)	2	4	5	1
4)	3	4	5	1

36.

List - I	List - II
I) $h\nu = W + K.E$	a) Quantization of angular momentum
II) $E = h\nu$	b) Wave numbers of Balmer series
III) $\bar{\nu} = R\left[\frac{1}{2^2} - \frac{1}{n^2}\right]$	c) Quantum theory
IV) $m.v.r = \frac{h}{2\pi}$	d) Photo electric effect

The correct match is

	I	II	III	IV
1)	a	b	c	d
2)	c	d	b	a
3)	d	c	b	a
4)	b	d	a	c

37.

List - I	List - II
I) Wave number	**a)** ms^{-1}
II) Frequency	**b) nm**
III) Wavelength	**c)** s^{-1}
IV) Velocity	**d)** m^{-1}

The correct match is

	I	II	III	IV		I	II	III	IV
1)	a	b	c	d	2)	d	c	b	a
3)	b	c	d	a	4)	c	d	b	a

38.

List - I	List - II
A) Energy	1) $\frac{2\pi ze^2}{nh}$
B) Velocity	2) $\frac{-2\pi^2 mz^2 e^4}{n^2h^2}$
C) Rydberg constant	3) $\frac{2\pi^2 mz^2 e^4}{h^3c}$
D) Radius	4) $\frac{n^2 h^2}{4\pi^2 mz\ e^2}$
	5) $\frac{-4\pi^2 mz^2\ e^4}{n^2\ h^2}$

The correct match is

1) A = 2 B = 4 C = 5 D = 1

2) A = 2 B = 1 C = 3 D = 4

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3) A = 3 B = 2 C = 1 D = 4
4) A = 4 B = 3 C = 1 D = 5

39.

List - I	List - II
I) Radial probability distribution curve of 3s orbital	a) $1.1A^0$
II) Distance of maximum probability of 1s electron	b) 1s orbital
III) Radial node for a 2s electron	c) 3 peaks, 2 radial nodes
IV) No spherical nodes	d) $0.53A^0$

The correct match is

	I	II	III	IV		I	II	III	IV
1)	a	b	c	d	2)	c	d	a	b
3)	b	a	d	c	4)	d	a	b	c

40.

List - I	List - II
A) nodal plane	1) $\frac{Ze^2}{2r}$
B) p-orbital	2) $\lambda = \frac{h}{mv}$
C) deBroglie	3) Spherical
D) Kinetic energy	4) Probability of e^- is zero
	5) Dumb bell

The correct match is

	A	B	C	D		A	B	C	D
1)	4	5	2	1	2)	2	4	3	5
3)	1	5	3	2	4)	3	1	4	2

41.

List - I	List - II
I) ψ^2 depends only upon distance	a) p–orbitals
II) ψ^2 depends upon distance and on one direction	b) d–orbital
III) ψ^2 depends upon distance and on two directions	c) f–orbital
IV) ψ^2 depends upon distance and on three directions	d) s –orbitals

The correct match is

	I	II	III	IV		I	II	III	IV
1)	d	c	b	a	2)	c	b	a	d
3)	d	a	b	c	4)	d	a	c	b

42.

List - I	List-II
A) No of electrons present in an orbit	1) 2
B) Number of orbitals in an orbit	2) n
C) Number of electrons in an orbital	3) n^2
D) Number of sub shells in an orbit	4) $2n^2$

The correct match is

	A	B	C	D		A	B	C	D
1)	4	2	1	3	2)	1	2	3	4
3)	4	3	1	2	4)	2	1	3	4

43.

List - I	List - II
I) The electrons of same orbital differ in *'s'* value	a) Hund's rule
II) Order of orbitals is *2s, 2p, 3s, 3p, 4s*	b) Stability of completely filled sublevel
III) E.C of N is $1s^2 2s^2 2p_x^1 2p_y^1 2p_z^1$	c) Pauli's principle
IV) E.C of Cu is $[Ar]\ 4s^1 3d^{10}$	d) Aufbau principle

The correct match is

	I	II	III	IV		I	II	III	IV
1)	a	b	c	d	2)	d	c	b	a
3)	c	d	a	b	4)	c	b	d	a

44.

Column-I	Column-II
A) 7s	P) Maximum energy
B) 4d	Q) Maximum number of electrons
C) 5d	R) 3 sub shells
D) 4p	S) Minimum number of orbitals

45.

Column-I	Column-II
A) Radial function (R)	P) n
B) Angular function (θ)	Q) l
C) Angular function (ϕ)	R) m
D) Quantised angular	

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momentum S) s

46. Match species given in Column I with the electronic configuration given in Column II.

Column I	Column II
(i) Cr	(a) $[Ar]3d^8 4s^0$
(ii) Fe^{2+}	(b) $[Ar]3d^{10} 4s^1$
(iii) Ni^{2+}	(c) $[Ar]3d^6 4s^0$
(iv) Cu	(d) $[Ar]3d^5 4s^1$
	(e) $[Ar]3d^6 4s^2$

	i	ii	iii	iv
1)	a	b	c	d
2)	d	c	a	b
3)	c	a	d	b
4)	b	c	a	d

47. Which of the following sets of quantum numbers are correct?

	n	l	m
(1)	1	1	+2
(2)	2	1	+1
(3)	3	2	-2
(4)	3	4	-2

1) 1&3 2) 2&3 3) 3&4 4) 1&4

48. In which of the following pairs, the ions are iso-electronic?

1) Na^+, Mg^{2+} (2) Al^{3+}, O^-

(3) Na^+, O^{2-} (4) N^{3-}, Cl^-

1) 1&2 2) 1&3 3) 2&3 4) 2&4

49. Which of the following statements concerning the quantum numbers are correct?

(1) Angular quantu m number determines the three dimensional shape of the orbital.

(2) The principal quantum number determines the orientation and energy of the orbital.

(3) Magnetic quantum number determines the size of the orbital.

(4) Spin quantum number of an electron determines the orientation of the spin of electron relative to the chosen axis.

1) 1&2 2) 1&4 3) 2&3 4)2&4

50. Match the following species with their corresponding ground state electronic configuration.

Atom/ion	Electron configuration
i) Cu	a) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$
ii) Cu^{2+}	b) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2$
iii) Zn^{2+}	c) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$
iv) Cr^{3+}	d) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$
	e) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^3$

The correct answer is :

	i	ii	iii	iv
1)	c	d	a	e
2)	d	a	e	c
3)	a	b	c	d
4)	c	b	d	e

51. Match the quantum numbers with the information provided by these.

Quantum number	Information provided
(i) Principal quantum number	(a) orientation of the orbital
(ii) Azimuthal quantum number	(b) energy and size of orbit
(iii) Magnetic quantum number	(c) spin of electron
(iv) Spin quantum number	(d) shape of the orbital

	i	ii	iii	iv
1)	b	a	d	c
2)	b	d	a	c
3)	a	c	d	b
4)	a	b	c	d

52. Match the following rules with their statements:

Rules

(i) Hund's Rule

(ii) Aufbau Principle

(iii) Pauli Exclusion Principle

(iv) Heisenberg's Uncertainty Principle

Statements

(a) No two electrons in an atom can have the same set of four quantum numbers.

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(b) Half-filled and completely filled orbitals have extra stablity.

(c) Pairing of electrons in the orbitals belonging to the samesubshell does not take place until each orbital is singly occupied.

d) It is impossible to determine the exact position and exact momentum of a subatomic particle simultaneously.

(e) In the ground state of atoms, orbitals are filled in the order of their increasing energies.

The correct answer is

	i	ii	iii	iv
1)	c	e	d	a
2)	e	c	d	a
3)	c	e	a	d
4)	e	d	a	b

53. Match the following

(i) X-rays (a) $v = 10^0 - 10^4\ Hz$

(ii) UV (b) $v = 10^{10}\ Hz$

(iii) Long radio waves (c) $v = 10^{16}\ Hz$

(iv) Microwave (d) $v = 10^{18}\ Hz$

	i	ii	iii	iv
1)	d	c	a	b
2)	c	a	b	d
3)	d	b	a	c
4)	a	d	b	c

54. Match the following

List - I	List-II
a. energy of ground state of He^+	i) $-6.04\ Ev$
b. Potential energy of I orbit of H-atom	ii) $-27.2\ eV$
c. kinetic energy of II excited state of He^+	iii) $8.7\times10^{-18}\ J$
d. ionisation potential of He^+	iv) $-54.4\ eV$

1) a-i, b-ii, c-iii, d-iv
2) a-iv, b-iii, c-ii, d-i
3) a-iv, b-ii, c-i, d-iii
4) a-ii, b-iii, c-i, d-iv

KEY

01) 3	02) 2	03) 4	04) 3	05) 1
06) 1	07) 3	08) 1	09) 2	10) 3
11) 1	12) 3	13) 4	14) 1	15) 1
16) 2	17) 4	18) 2	19) 2	20) 2
21) 1	22) 3	23) 4	24) 1	25) 1
26) 1	27) 1	28) 1	29) 1	30) 1
31) 2	32) 1	33) 2	34) 2	35) 4
36) 3	37) 2	38) 2	39) 2	40) 1
41) 3	42) 3	43) 3		

44) A-P,S; B-Q; C-Q; D-R

45) A-P,Q; B-Q,R; C-R; D-Q,S

46. 2	47. 2	48. 2	49. 2	50. 1
51. 2	52. 3	53. 1	54. 3	

HINTS & SOLUTIONS

7. The minimum frequency required to eject an electron from the surface of metal is called threshold frequency.

8. Due to same electronic configuration.

10. H-spectrum can be both emission and absorption

15. $\lambda = \frac{h}{mv}$

21. 2P orbitals are degenerate.

22. The two lobes are oriented along x-axis.

24. ψ^2 is the probability function

ψ is the wave function.

34. no. of electrons in an orbit = $2n^2$

no. of orbitals in an orbit = n^2

no. of electrons in an orbital = 2

no. of subshells = n

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CONCEPT BOOSTER

APPLICATIONS OF RUTHERFORD MODEL

On the basis of scattering experiments, Rutherford proposed model of the atom, which is known as nuclear atomic model. According to this model-

(i) An atom consists of a heavy positively charged nucleus where all the protons and neutrons are present. Protons & neutros are collectively reffered to as nucleons. Almost whole of the mass of the atom is contributed by these nucleon. The magnitude of the +ve charge on the nucleus is different for different atoms.

(ii) The volume of the nucleus is very smalland is only a minute fraction of the total volume of the atom. Nucleus has a diamter of the order of $10^{-12} to 10^{-13}$ cm and the atom has a diameter of the order of 10^{-8} cm.

$$\frac{D_A}{D_N} = \frac{\text{Diameter of the atom}}{\text{Diameter of the nucleus}}$$

$$= \frac{10^{-8}}{10^{-13}} = 10^5, \quad D_A = 10^5 D_N$$

➢ Thus diameter (size) of the atom is 10^5 times the diameter of the nucleus.

The radius of a nucleus is proportional to the cube root of the number of nucleons with it.

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

➢ Where $R_0 = 1.33 \times 10^{-13}$ (a constant) and
A = mass number (p+n) and
R=radius of the nucleus

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

(iii) There is an empty space around the nucleus called extra nuclear part. In this part electrons are present. The number of electrons in an atom is always equal to number of protons present in the nucleus. As the nuclear part of atom is responsible for the mass of the atom, the extra nuclear part is responsible for its volume.

The volume of the atom is about 10^{15} times the volume of the nucleus.

$$\frac{\text{volume of the atom}}{\text{volume of the nucleus}} = \frac{(10^{-8})^3}{(10^{-13})^3} = 10^{15}$$

*(iv) Electrons revolve around the nucleus in closed orbits with high speeds. The centrifugal force acting on the revolving e^- is being counter balanced by the force of attraction between the electrons and the nucleus. This model was similar to the solar system, the nucleus representing the sun and revolving electrons as planets.

CONCEPT BOOSTER

Nature of light :

* It can be explained by two theories
a) wave theory of light
b) corpuscular theory of light

Wave theory :

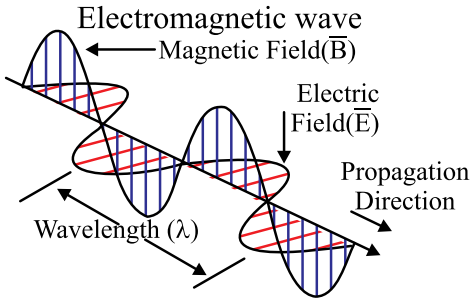
➢ According to Huygens Light and other forms of radiant energy propagate through space in the form of waves.

* Maxwell proposed when electrically charged particle moves under acceleration,alternating electrical and magnetic fields are produced and transmitted.

* These are transmitted in the form of wave and are associated with electric and magnetic fields are called electromagnetic radiation(EMR) or electromagnetic waves.

* Both electric and magnetic fields are perpendicular to each other and to the direction of the propagation.

* These electromagnetic waves do not require any medium and can travel in vaccum.



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POINTS

- The atomic theory of matter was first proposed by john Datlon,known as Daltons atomic thoery .
- According to Modern atomic theory atoms can be further devided in to subatomic particles i.e electrons,protons and neutrons.
- The value of charge to mass ratio of electron is $1.758820\times10^{11}ckg^{-1}$
- Mulliken determined the charge of the electron by an oil drop experiment
- Charge on the oildrops was always an integral multiple of $1.60\times10^{-19}c$

  $m_e = 9.1094\times10^{-31}kg$
- $\frac{e}{m}$ value of anode rays is maximum when the gas present in the tube is hyderogen
- **Isotopes:** Atoms with identical atomic number but different mass numbers are known as isotopes.

  Eg: ${}_1H^1, {}_1H^2, {}_1H^3$
- **Isobars:** Atoms with same mass number with different atomic number are known as isobars. Eg: ${}_6C^{14}\ {}_7N^{14}$
- The number of electrons or protons present in an atom of an element is called its atomic number.
- Atomic number is equal to the nuclear positive charge of an element.
- Number of neutrons=A-Z.
- The wave length of visible light is 3800-7600 A^0
- Black body is perfect absorber and perfect emitter of radiant of energy.
- Substance absorb or emit light discontinuously in the form of small packets or bundles.The smallest packet of energy is called quantum.
- When light is exposed to clean metallic surface,electron are ejected from the surface.This effect is called photo electric effect. $h\nu = W + KE$
- The wave length or wave number of varoius lines in the visible region can be expressed by an equation.

  $$\bar{\nu} = \frac{1}{\lambda} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$
- Maximum number of lines produced when an electron jumps from nth level to ground level

  $$= \frac{n(n-1)}{2}.$$
- Expression for the radius of Bohrs orbit

  $$r = \frac{n^2h^2}{4\pi^2mZe^2}$$
- Radius of orbits in H atom like ions

  $$r = \frac{0.529\times n^2}{Z}A^0$$
- Total energy of electron =KE+PE=

  $$= \frac{Ze^2}{2r} - \frac{Ze^2}{r} = \frac{-Ze^2}{2r}$$
- de Broglie's equation is $\lambda = \frac{h}{mv} = \frac{h}{p}$
- Number of waves in an orbit =n
- The uncertainty principle equation is

  $$\Delta x.\Delta p \geq \frac{h}{4\pi}$$
- ψ indicates the amplitude of the wave.
- ψ^2 indicates maximum probability of finding an electron at a certain point .
- The plane where the probability of finding the electron is zero $(\psi^2 = 0)$ is called a nodal plane.
- Number of radial nodes $= n-l-1$ where n = principal quantum number.
- Chromium (Z=24) and copper (Z=29) have anomalous electronic configuration.