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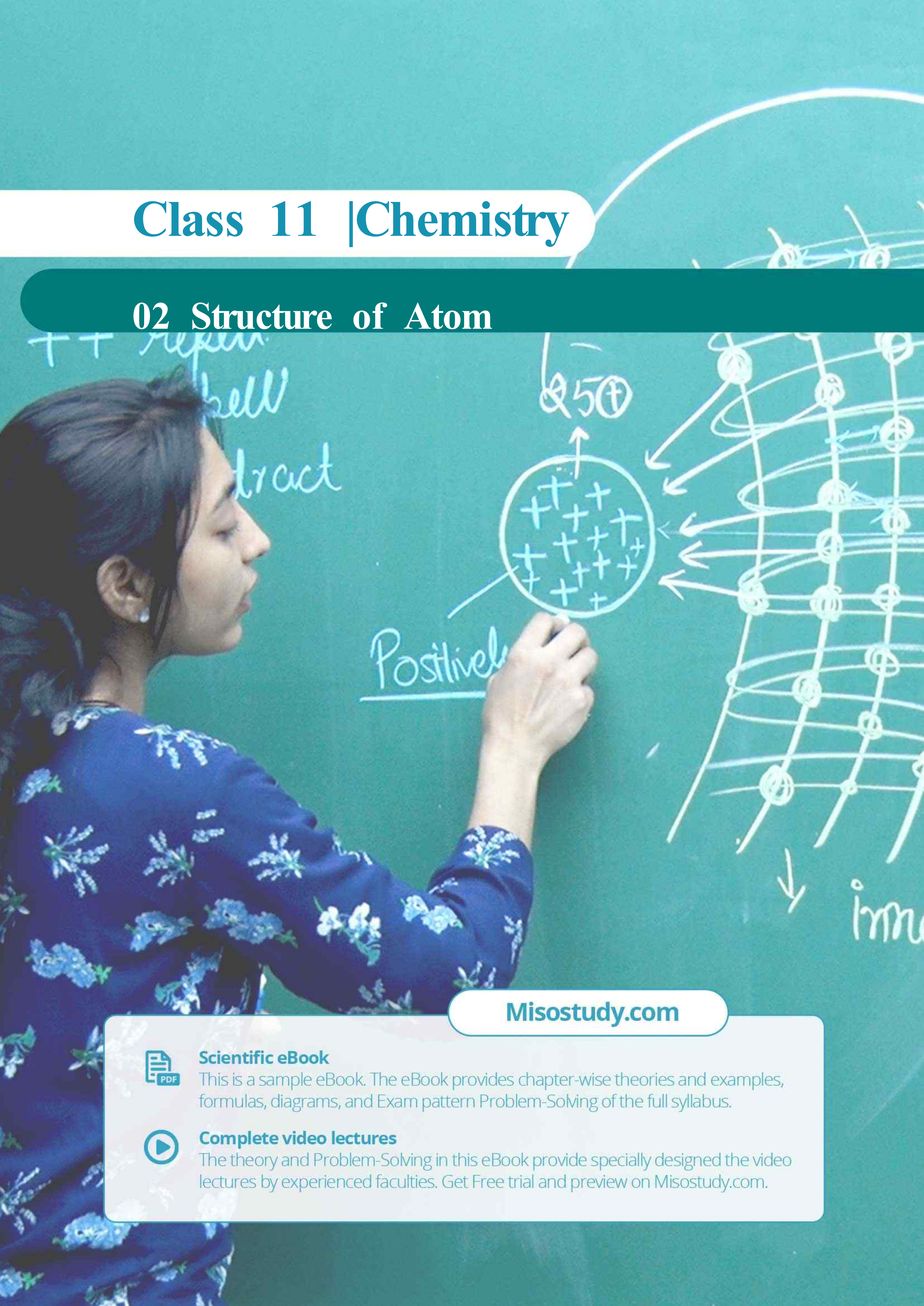

Class 11 |Chemistry

02 Structure of Atom

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01. Introduction

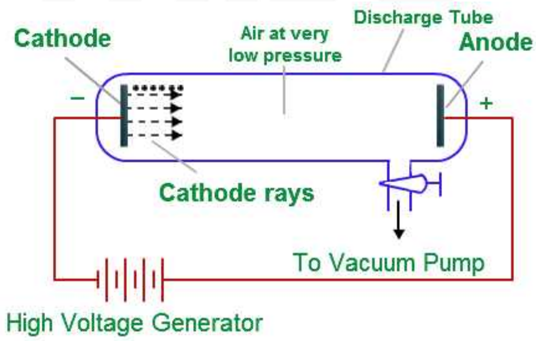
John Dalton 1808, believed that matter is made up of extremely minute indivisible particles, called atom which takes part in chemical reactions. These particle can neither be created nor be destroyed. However, modern researches have conclusively proves that atom is no longer an indivisible particle. Modern structure of atom is based on Rutherford's scattering experiment, quantization of energy and wave mechanical model.

Composition of Atom

The works of J.J. Thomson and Ernst Rutherford actually laid the foundation of the modern picture of the atom. If is now verified that the atom consists of several sub-atomic particles like electron, proton, neutron, positron, neutrino, meson etc. Out of these particles the electron, proton and the neutron are called fundamental subatomic particles.

02. ELECTRON ($_{-1}e^0$, e)

Electron was discovered by J.J. Thomson(1897) and it is a negatively charged particle.
Cathode rays were discovered by William Crooke & J.J. Thomson using a cylindrical hard glass tube fitted with two metallic electrodes. This tube was known as discharge tube. They passed electricity (10,000V) through a discharge tube at very low pressure. Blue rays emerged from the cathode. These reys were termed as Cathode rays.



Properties of Cathode rays

(i) Cathode rays travel in straight line.

(ii) Cathode rays produce mechanical effect, as they can rotate the wheel placed in their path.

(iii) Cathode rays consist of negatively charged particles known as electron.

(iv) Cathode rays travel with high speed.
(v) Cathode rays can cause fluorescence.
(vi) Cathode rays heat the object on which they fall due to transfer of kinetic energy to the object.
(vii) When cathode rays fall on heavy metals, X-rays are produced.
(viii) Cathode rays possess ionizing power i.e., they ionize the gas through which they pass.
(ix) The cathode rays produce scintillation on the photographic plates.
(x) They can penetrate through thin metallic sheets.
(xi) The nature of these rays does not

03. Thomson's Model of Atom [1904]

- Thomson was the first to propose a detailed model of the atom.
- Thomson proposed that an atom consists of a uniform sphere of positive charge in which the electrons are distributed more or less uniformly.
- This model of atom is known as "Plum-Pudding model" or "Raisin Pudding Model" or "Water Melon Model".

Drawbacks:

- An important drawback of this model is that the mass of the atoms is considered to be evenly spread over that atom.
- It is a static model. It does not reflect the movement of electron.
- It could not explain the stability of an atom.

04. Rutherford's Scattering Experiment

Rutherford observed that

(i) Most of the α-particles (nearly 99.9%) went straight without suffering any deflection.
(ii) A few of them got deflected through small angles.
(iii) A very few (about one in 20,000) did not pass through the foil at all but suffered large deflections (more than 90°) or even came back in the direction from which they have come i.e. a deflection of 180°.

Following conclusions were drawn from the above observations-

(i) Since most of the α-particle went straight through the metal foil undeflected, it means that there must be very large empty space within the atom.
(ii) Since few of the α-particles were deflected from their original paths through moderate angles; it was concluded that whole of the +ve charge is concentrated and the space occupied by this positive charge is very small in the atom.
  - When α-particles come closer to this point, they suffer a force of repulsion and deviate from their paths.
  - The positively charged heavy mass which occupies only a small volume in an atom is called nucleus. It is supposed to be present at the centre of the atom.

(iii) A very few of the α-particles suffered strong deflections on even returned on their path indicating that the nucleus is rigid and α-particles recoil due to direct collision with the heavy positively charged mass.

Drawbacks of rutherford model-

(i) This theory could not explain stability of atom. According to Maxwell electron loose its energy continuously in the form of electromagnetic radiations. As a result of this, the e- should loose energy at every turn and move close and closer to the nucleus following a spiral path. The unlimited result will be that it will fall into the nucleus, thereby making the atom unstable.

(ii) If the electrons loose energy continuously, the observed spectrum should be continuous but the actual observed spectrum consists of well defined lines of definite frequencies. Hence the loss of energy by electron is not continuous in an atom.

05. Distance of closest approach :

When the α-particles approaches the nucleus to made a head-on collision with a the nucleus, the α-particle approaches the nucleus until coulombic potential energy of repulsion, $k\dfrac{Z_1Z_2e^2}{r}$, becomes equal to its initial K.E., $\dfrac{1}{2}$ m.v^2.

Thus $\dfrac{1}{2}mv^2 = k\dfrac{Z_1Z_2e^2}{r}$

Hence, the distance of closest approach, $r = \dfrac{k\,z_1\,z_2\,e^2}{\left(\dfrac{1}{2}mv^2\right)}$.

The nucleus must be further smaller than the distance of closest approach.

06. Moseley Experiment (Discovery of Atomic Number)

Moseley (1912-1913), investigated the X-rays spectra of 38 different elements, starting from aluminium and ending in gold. He measured the frequency of principal lines of a particular series (the α-lines in the K series) of the spectra. It was observed that the frequency of a particular spectral line gradually increased with the increase of atomic mass of the element. But, it was soon realised that the frequency of the particular spectral line was more precisely related with the serial number of the element in the periodic table which he termed as atomic number (Z). He presented the following relationship:

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

where, v=frequency of X-rays, Z=atomic number, 'a' and 'b' are constants. When the values of square root of the frequency were plotted against atomic number of the elements producing X-rays, a straight line was obtained.

07. Some Atomic Terms

Atomic number = Number of unit positive charge on nucleus = Number of protons in nucleus = Number of electrons in neutral atom.
Two different elements can never have identical atomic number.

Mass number(A) = Number of protons + Number of neutrons or Atomic number (Z)
Hence Number of neutrons = A – Z.
The atom of an element X having mass number (A) and atomic number (Z) may be represented by a symbol,

Isotopes : Atoms of an element with the same atomic number but different mass number.

eg.	$_1H^1$	$_1H^2$	$_1H^3$
	Protonium	deuterium	Tritium

Isodiapheres : The elements which have same value of (n – p) is called Isodiapheres.

eg.	$_7N^{14}$	$_8O^{16}$
Values of (n - p)	0	0

Isotone : Elements which contain same no. of neutron is called **Isotone**.

eg.	$_{14}Si^{30}$	$_{15}P^{31}$	$_{16}S^{32}$
number of neutrons	16	16	16

Isobar : The two different atoms which have same mass number but different atomic number is called Isobar.

eg.	$_{18}Ar^{40}$	$_{19}K^{40}$	$_{20}Ca^{40}$

Isoelectronic : Ion or atom or molecule which have the same mass number of electron is called Isoelectronic species.

eg.	$_{17}Cl^{-}$	$_{18}Ar$	$_{19}K^{+}$	$_{19}Ca^{+2}$
No. of electrons	18	18	18	18

08. Bohr's Atomic Model

This model was based on the quantum theory of radiation and the classical law of physics. It gave new idea of atomic structure in order to explain the stability of the atom and emission of sharp spectral lines.

Postulates

(i) The atom has a central massive core nucleus where all the protons and neutrons are present. The size of the nucleus is very small.

(ii) The electron in an atom revolve around the nucleus in certain discrete orbits. Such orbits are known as stable orbits or non – radiating or stationary orbits.

(iii) An electron can move only in those permissive orbits in which the angular momentum (mvr) of the electron is an integral multiple of $h/2\pi$ Thus,

$$mvr = n\frac{h}{2\pi}$$

Where, m = mass of the electron, r = radius of the electronic orbit, v = velocity of the electron in its orbit.

(iv) The angular momentum can be $\frac{h}{2\pi}, \frac{2h}{2\pi}, \frac{3h}{2\pi}, ...\frac{nh}{2\pi}$. This principal is known as quantization of angular momentum. In the above equation 'n' is positive integer which has been called as principal quantum number. It can have the values n=1,2,3,...... (form the nucleus). Various energy levels are designed as K(n=1), L(n=2), M(n=3),......etc. Since the electron present in these orbits is associated with some energy, these orbits are called energy levels.

(v) The emission or absorption of radiation by the atom takes place when an electron jumps from one stationary orbit to another.

09. Calculation of the radius of the Bohr's orbit :

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

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

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

Where : $K = \text{constant} = \frac{1}{4\pi\epsilon_0} = 9\times10^9\ Nm^2/C^2$

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

Hence $\frac{mv^2}{r} = \frac{KZe^2}{r^2}$ or, $v^2 = \frac{KZe^2}{mr}$ ...(1)

From the postulate of Bohr, $mvr = \frac{nh}{2\pi}$ or, $v^2 = \frac{n^2h^2}{4\pi^2m^2r^2}$ ...(2)

From equation (1) and (2) :

$\therefore\ r = \frac{n^2h^2}{4\pi^2mKZe^2}$ On putting the value or e, h, m,

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

10. Calculation of velocity of an electron in Bohr's orbit

Velocity of the revolving electron in n^{th} orbit is given by

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

Putting the value of r in above equation then $\quad v = \dfrac{nh \times 4\pi^2 mZe^2}{2\pi mn^2h^2}$, $\quad v = \dfrac{2\pi Ze^2}{nh}$

on putting the values of e and h,

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

11. Calculation of energy of an electron :

The total energy of an electron revolving in a particular orbit is -
T. E. = K. E. + P. E.

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

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

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

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

Substituting the value of mv^2 in the equation (3)

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

So, $\quad$ T.E. $= -\dfrac{KZe^2}{2r}$

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

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

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

$E_n = \dfrac{2\pi^2 Z^2 e^4 mk^2}{n^2h^2}$

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

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

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

12. Relation between P. E., K. E. & T. E. :

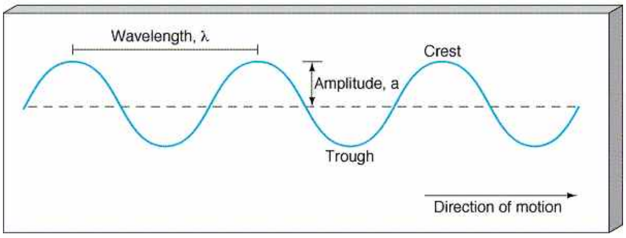
$P.E. = -\dfrac{kZe^2}{r}$, $\quad K.E. = \dfrac{1}{2}\dfrac{kZe^2}{r}$, $\quad T.E. = -\dfrac{1}{2}\dfrac{kZe^2}{r}$,

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

13. Details About Waves

A wave motion is a means of transferring energy from one point to another point without any actual transportation of matter between these points. When we throw a piece of stone on the surface of water in a pond, we observe circles of ever increasing radius, till they strike the wall of the pond.

14. Some important terms related with wave motion.



(i) **Wave length (λ):** The distance between two adjacent crest or trough of the wave (or the distance between two similar neighbouring points)

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

(iii) **Frequency (v):** Number of oscillations per unit time is called frequency. $v\frac{C}{\lambda}$

(iv) **Wave number $(\bar{v})$:** Number of wavelength per unit length. $(\bar{v}) = \frac{1}{\lambda}$

(v) **Amplitude (A):** It is the height of crest or depth of a trough of a wave.

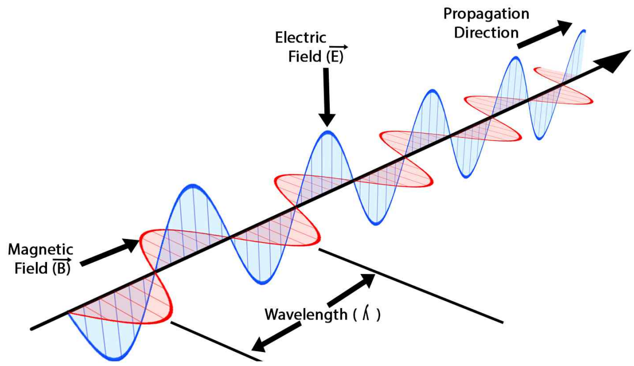
15. Characteristics of electromagnetic radiations:

(i) All electromagnetic waves move or travel with the same speed equal to that of light.

(ii) They do not require any medium to propagate.

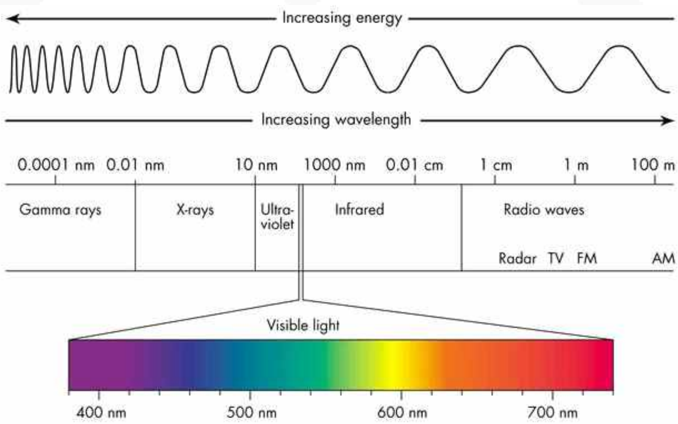
(iii) These consist of electric and magnetic field that oscillate in the direction perpendicular to each other and to the direction in which the wave propagates (as shown in above diagram)

Electromagnetic Wave



16. Electromagnetic Spectrum

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



17. Planck's Quantum Theory

According to this theory atoms or molecules could emit or absorb energy only in discrete quantities (small packets) and not in any arbitary amount. Planck gave the name quantum to the smallest quantity of energy that can be emitted in the form of E.M. radiation.
The energy of a photon is proportional to its frequency and is given by
$E = h\upsilon$ where $h = 6.626\times10^{-34}$ J sec
A body can emit or absorb energy only in terms of the integral multiples of quantum, i.e.
$E = n.\ h\upsilon$, where n = 1,2,3,.....

i.e. a body can emit or absorb energy as $h\upsilon$, $2h\upsilon$ ..... but it can not emit or absorb energy in fractional values of $h\upsilon$ such as $1.5\ h\upsilon$, $2.5\ h\upsilon$.
Einstein supported the planck's theory and explained the photoelectric effect considering that electromagnetic radiations also propagate in the form of photon. Energy of each photon depends on frequency of light $\left(E = h\upsilon = \frac{hc}{\lambda}\right)$.

Since wave character of light explains the interference and diffraction phenomenon while the particle character explains, black body radiations and photoelectric effect, the light was considered to have wave as well as particle character [Dual character of light.]

(i) Wave nature : diffraction, interference, polarisation.
(ii) Particle nature : photoelectric effect.

18. Hydrogen Spectrum

Hydrogen spectrum is an example of line or atomic emission spectrum. When an electric discharge is passed through hydrogen gas at low pressure, a bluish light is emitted. All these lines of H-spectrum have Lyman, Balmer, Paschen, Barckett, Pfund and Humphrey series. These spectral series were named by the name of scientist discovered them. To evalute wavelength of various H-lines Rydberg introduced the following expression,

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

Where R is constant known as Rydberg's constant its value is 109, 67800 m^{-1}
Although H-atom consists only one electron yet it's spectra consist of many spectral lines because electrons in different hydrogen atoms absorb different amount of energies and are excited to different energy levels. Since life time of electrons in these excited states is very small, they return to some lower energy level or even to the ground state in one or more jumps.

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

19. Lyman Series

(a) It is a first series of spectral series of H.
(b) It's value of $n_1 = 1$ and $n_2 = 2,3,4$, where 'n_1' is ground state and 'n_2' is called excited state of electron present in *H*-atom.
(c) If the electron goes to $n_1 = 1$ from $n_2 = 2$ — first Lyman series
If the electron goes to $n_1 = 1$ from $n_2 = 3$ — Second Lyman series
If the electron goes to $n_1 = 1$ from $n_2 = 4$ — third Lyman series ----- so on.

(d) $\frac{1}{\lambda} = R_H\left[\frac{1}{1^2} - \frac{1}{n_2^2}\right]$ where $n_2 > 1$ always.

(e) For marginal line of limiting line $n_2 = \infty$. Hence the wavelength of marginal line $= \frac{n_1^2}{R_H}$ for all series. So, for lyman series it is $\frac{1}{R_H}$.

20. Balmer Series :

(i) It is second series of H-spectral series.
(ii) It was found out in 1892 in visible region by **Balmer.**
(iii) Blamer series was found out before all series because it was found in visible region.
(iv) It's value of $n_1 = 2$ and $n_2 = 3,4,5$ ....
(v) If the electron goes to $n_1 = 2$ from $n_2 = 3$ — first Balmer series
If the electron goes to $n_1 = 2$ from $n_2 = 4$ — Second Balmer series
If the electron goes to $n_1 = 2$ from $n_2 = 5$ — third Balmer series ----- so on.

(vi) The wavelength of marginal line of Balmer series $= \frac{n_1^2}{R_H} = \frac{2^2}{R_H} = \frac{4}{R_H}$.

(vii) $\frac{1}{\lambda} = R_H\left[\frac{1}{2^2} - \frac{1}{n_2^2}\right]$ where $n_2 > 2$ always.

21. Paschen Series :

(i) It is the third series of H – spectrum.
(ii) It was found out in infra red region by ***Paschen.***
(iii) It's value of $n_1 = 3$ and $n_2 = 4,5,6$ ....
(iv) If the electron goes to $n_1 = 3$ from $n_2 = 4$ — first Paschen series
If the electron goes to $n_1 = 3$ from $n_2 = 5$ — Second Paschen series
If the electron goes to $n_1 = 3$ from $n_2 = 6$ — third Paschen series ----- so on.
(v) The wavelength of marginal line of paschen series $= \frac{n_1^2}{R_H} = \frac{3^2}{R_H} = \frac{9}{R_H}$.

(vi) $\frac{1}{\lambda} = R_H\left[\frac{1}{3^2} - \frac{1}{n_2^2}\right]$ where $n_2 > 3$ always.

22. Brackett Series :

(i) It is fourth series of H – spectrum.
(ii) It was found out in infra red region by **Brackett.**
(iii) It's value of $n_1 = 4$ and $n_2 = 5,6,7$ ....

(iv) If the electron goes to $n_1 = 4$ from $n_2 = 5$ — first Brackett series
If the electron goes to $n_1 = 4$ from $n_2 = 6$ — Second Brackett series
If the electron goes to $n_1 = 4$ from $n_2 = 7$ — third Brackett series ----- so on.

(v) The wavelength of marginal line of Brackett series $= \frac{n_1^2}{R_H} = \frac{4^2}{R_H} = \frac{16}{R_H}$.

(vi) $\frac{1}{\lambda} = R_H\left[\frac{1}{4^2} - \frac{1}{n_2^2}\right]$ where $n_2 > 4$ always.

23. Pfund Series :

(i) It is fifth series of H – spectrum.
(ii) It was found out in infra red region by Pfund.
(iii) It's value of $n_1 = 5$ and $n_2 = 6,7,8$ ....
(iv) If the electron goes to $n_1 = 5$ from $n_2 = 6$ — first Pfund series
If the electron goes to $n_1 = 5$ from $n_2 = 7$ — Second Pfund series
If the electron goes to $n_1 = 5$ from $n_2 = 8$ — third Pfund series ----- so on.

(v) The wavelength of marginal line of Pfund series $= \frac{n_1^2}{R_H} = \frac{5^2}{R_H} = \frac{25}{R_H}$.

(vi) $\frac{1}{\lambda} = R_H\left[\frac{1}{5^2} - \frac{1}{n_2^2}\right]$ where $n_2 > 5$ always.

24. Calculation of Rydberg Constant from bohr's atomic model

Suppose that an electron transit from first energy. level to second energy level. Then, the change of energy. is given by

$\triangle E = E_{n_2} - E_{n_1}$

or, $= \left[\frac{-2\pi^2 mZ^2e^4k^2}{n_2^2 h^2}\right] - \left[\frac{-2\pi^2 mZ^2e^4k^2}{n_1^2 h^2}\right]$ or, $\frac{hc}{\lambda} = \frac{-2\pi^2 mZ^2e^4k^2}{h^2} \times \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$

or, $\frac{1}{\lambda} = \frac{2\pi^2 mZ^2e^4k^2}{ch^3}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] = R_H \cdot Z^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$

or, $R_H = \frac{2\pi^2 me^4k^2}{ch^3}$ Rydberg constant

25. Particle and Wave Nature of Electron

In 1924,de Broglie proposed that an electron, like light, behaves both as material particle and as a wave. This proposal gave a new theory, known as wave mechanical theory of matter. According to this theory, the electrons, protons and even atoms, when in motion, posses wave properties.

de Broglie derived an expression for calculating the wavelength of the wave associated with the electron. According to Planck's equation

$$E = h\upsilon = h.\frac{c}{\lambda} \quad .....(1)$$

The energy of a photon on the basis of Einstein's mass-energy relationship is

$$E = mc^2 \quad .....(2)$$

Where, c is the velocity of the electron.
From (1) and (2) equation

$$h \times \frac{c}{\lambda} = mc^2$$

$$\lambda = \frac{h}{mc} = \frac{h}{p} \quad .....(3)$$

Momentum of the moving electron is inversely proportional to its wavelength.
Let kinetic energy of the particle of mass 'm' is E.

$$E = \frac{1}{2}mv^2$$

$$2Em = m^2v^2$$

$$\sqrt{2Em} = mv = p\,(\text{momentum})$$

$$\lambda = \frac{h}{p} = \frac{h}{\sqrt{2Em}} \quad .....(4)$$

Let a charged particle, with charge q be accelerated with a potential of V; then the kinetic energy may be given as : E=qV

$$\frac{1}{2}mv^2 = qV$$

$$m^2v^2 = 2qVm$$

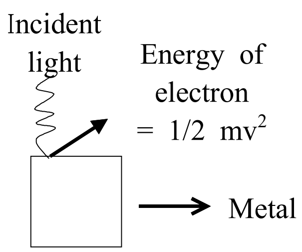
$$mv = \sqrt{2qVm}$$

and,

$$\lambda = \frac{h}{\sqrt{2qVm}} \quad .....(5)$$

26. Photoelectric Effect

Emission of electrons from a metal surface when exposed to light radiations of appropriate wavelength is called photoelectric effect. The emitted electrons are called photoelectrons.



Work function or threshold energy may be defined as the minimum amount of energy required to eject electrons from a metal surface. According to Einstein,
Maximum kinetic energy of the ejected electron = absorbed energy – work function

$$\frac{1}{2}mv_{max}^2 = hv - hv_0 = hc\left[\frac{1}{\lambda} - \frac{1}{\lambda_0}\right]$$

where, v_0 and λ_0 are threshold frequency and threshold wavelength respectively.

Stopping Potential

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

If V_0 is the stopping potential, then

$$eV_0 = h(v - v_0) \text{ or } V_0 = \frac{h(v - v_0)}{e} = \frac{K.E._{max}}{e}$$

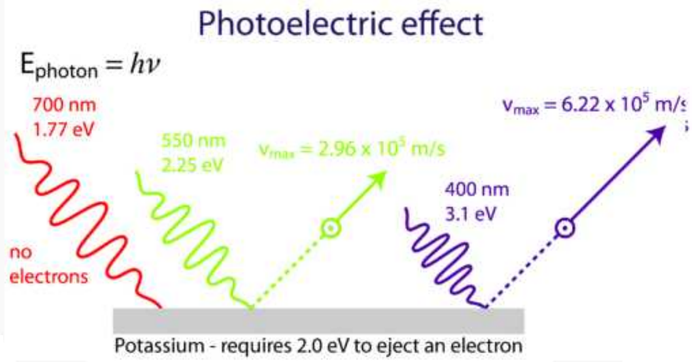
Some facts of Photoelectric Effect

(i) There is no time lag between incidence of light and emission of photoelectrons.

(ii) For emission of photoelectrons, the frequency of incident light must be equal to or greater than the threshold frequency.

(iii) Rate of emission of photoelectrons from a metal surface is directly proportional to the intensity of incident light.

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



27. Heisenberg's Uncertainty Principle

According to this principle it is impossible to measure simultaneously the exact position and exact momentum of a body as small as an electron. If uncertainty of measurement of position is $\triangle x$ and uncertainty of measurement of momentum is $\triangle p$ or $m\triangle v$, then according to Heisenberg.

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

where h is planck's constant

Like de Broglie equation, this principle has significance only for microscopic particles.

28. Wave Mechanical Model of Atom

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

Significance of Ψ :
The wave function may be regarded as the amplitude function expressed in terms of coordinates x, y and z. The wave function may have positive of negative values depending upon the values of coordinates

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

29. Quantum Numbers

Quantum numbers are to specify and display to complete information about size, shape ans orientation of the orbital. These are principal, azimuthal and magnetic quantum number, which follows directly from solution of schrodinger wave equation.

30. Principal Quantum Number (n) :

(i) It was proposed by Bohr and denoted by 'n'.
(ii) It determines the average distance between electron and nucleus, means it is denoted the size of atom.
(iii) It determine the energy of the electron in an orbit where electron is present.
(iv) The maximum number of an electron in an orbit represented by this quantum number as $2n^2$.

31. Azimuthal Quantum Number of Angular Quantum Number (*l*) -

(i) It was proposed by Sommerfield and denoted by '*l*'.
(ii) It determines the number of subshells or sublevels to which the electron belongs.
(iii) It tells about the shape of subshells.
(iv) It also expresses the energies of subshells s<p<d<f (Increasing energy).
(v) The value of *l* is integral values upto (n-1), starting from zero where 'n' is the number of principle shell.

(vi)

Value of l	0	1	2	3
Name of subshell	s	p	d	f
Shape of orbital	spherical	Dumbbell	Clover leaf (Except d_{z^2} doughtnut)	Complex

(vii) It represent the orbital angular momentum, which is equal to $\frac{h}{2\pi}\sqrt{l(l+1)}$.

(viii) The number of electrons in subshell = 2(2l+1).

(ix) For a given value of 'n' the total value of 'l' is always equal to the value of 'n'.

32. Magnetic Quantum Number (m) :

(i) It gives the number of permitted orientation of subshells.
(ii) The value of m varies from $-l$ to $+l$ through zero.
(iii) Degenerate orbitals – Orbitals having the same energy are known as degenerate orbitals. e.g. for p subshell P_x, P_y and P_z are degenerate orbital.
(iv) The number of degenerate orbitals of s subshell = 0.

33. Spin quantum number (s) :

(i) The value of 's' is $+\frac{1}{2}$ or $-\frac{1}{2}$, which is signified as the spin or rotation or direction of electron on it's axis during the movement.
(ii) The spin may be clockwise or anticlockwise.
(iii) It represents the value of spin angular momentum is equal to $\frac{h}{2\pi}\sqrt{s(s+1)}$.
(iv) Maximum spin of an atom = $\frac{1}{2}\times$number of unpaired electron.

34. Shape and size of orbitals

An orbital is the region of space around the nucleus within which the probability of finding an electron of given energy is maximum (90-95%). The shape of this region (electron cloud) gives the shape of the orbital. It is basically determined by the azimuthal quantum number l, while the orientation of orbital depends on the magnetic quantum number (m).

35. s-orbital (l = 0):

These orbitals are spherical and symmetrical about the nucleus. The probability of finding the electron is maximum near the nucleus and keeps on decreasing as the distance from the nucleus increases. There is vacant space between two successive s-orbitals known as radial node. But there is no radial node for 1s orbital since it is starting from the nucleus.

36. p-orbital (l = 1):

The probability of finding the p-electron is maximum in two lobes on the opposite sides of the nucleus. This gives rise to dumb-bell shape for the p-orbital l = 1.
Hence, m = –1, 0, +1. Thus, p-orbital have three different orientations. These are designated as p_x, p_y & p_z depending upon whether the density of electron is maximum along the x y and z axis respectively.

37. d-orbital (l = 2):

For d-orbitals, l = 2. Hence m = –2, –1, 0, +1, +2. Thus there are 5d orbitals. They have relatively complex geometry. Out of the five orbitals, the three (d_{xy}, d_{yz}, d_{zx}) project in between the axis and the other two d_{z^2} and $d_{x^2-y^2}$ lie along the axis.

38. Spherical nodes :

The spherical surface where probability of finding the electron is zero, is called spherical nodes.
No. of spherical nodes in any orbital = n–l–1

39. Nodal Plane :

This is a plane passing through the nucleus where the probability of finding the electron is zero.
Number of nodal plane in a orbital = l

Orbital	Nodal plane
p_x	yz
p_y	xz
p_z	xy
d_{xy}	yz,zx
d_{yz}	xy,xz
d_{zx}	xy,yz

40. Electronic Configuration

The distribution of electrons in different orbitals of atom is known as electronic configuration of the atoms. Filling up of orbitals in the ground state of atom is governed by the following rules:

41. Aufbau Principle

According to this principle, "In the ground state, the atomic orbitals are filled in order of increasing energies i.e. in the ground state the electrons first occupy the lowest energy orbitals available".
Lower the value of $n+l$, lower is the energy of the orbital and such an orbital will be filled up first.

When two orbitals have same value of $(n+l)$ the orbital having lower value of "n" has lower energy and such an orbital will be filled up first.
Thus, order of filling up of orbitals is as follows:
$1s<2s<2p<3s<3p<4s<3d<4p<5s<4d<5p<6s<4f<5d$

42. Pauli's Exclusion Principle

According to this principle, "No two electrons in an atom can have same set of all the four quantum numbers n, l, m and s.
According to this principle an orbital can accommodate at the most two electrons with spins opposite to each other. It means that an orbital can have 0, 1, or 2 electron. If an orbital has two electrons they must be of opposite spin.

43. Hund's Rule of Maximum Multiplicity

According to this rule "Electron filling will not take place in orbitals of same energy until all the available orbitals of a given subshell contain one electron each with parallel spin".

44. Extra stability If half filled and completely filled orbitals

Symmetry of orbitals
If the shift of an electron from one orbital to another orbital differing slightly in energy results in the symmetrical electronic configuration. It becomes more stable. For example p^3, d^5, f^7 configurations are more stable then their near once.
Exchange energy
In case of half filled and completely filled orbitals, the exchange energy is maximum and is greater than the loss of energy due to the transfer of electron from a higher to a lower sublevel e.g. from 4s to 3d orbitals in case of Cu and Cr.

CBSE Exam Pattern Exercise
Subjective Questions (1)

(Q. No. 1 to 2) One Mark

1. Which of the following are isoelectronic species, i.e., those having the same number of electrons?

   Na^+, K^+, Mg^{2+}, Ca^{2+}, S^{2-}, Ar.

2. A certain particle carries 2.5×10^{-16} C of static electric charge. Calculate the number of electrons present in it.

(Q. No. 3 to 4) Two Marks

3. An element with mass number 81 contains 31.7% more neutrons as compared to protons. Assign the atomic symbol.

4. If the diameter of a carbon atom is 0.15 nm, calculate the number of carbon atoms which can be placed side by side in a straight line across the length of scale of length 20 cm long.

(Q. No. 5 to 6) Three Marks

5. How much energy is required to ionise a hydrogen atom if the electron occupies n = 5 orbit? Compare your answer with the ionisation energy of hydrogen atom (energy required to remove the electron from n = 1 orbit).

6. The quantum numbers of four electrons are given below. Arrange them in order of increasing energies. List if any of these combination (s) has/have the same energy:
   (i) $n = 4,\ l = 2,\ m_l = -2,\ m_s = -1/2$
   (ii) $n = 3,\ l = 2,\ m_l = -1,\ m_s = +1/2$
   (iii) $n = 4,\ l = 1,\ m_l = 0,\ m_s = +1/2$
   (iv) $n = 3,\ l = 2,\ m_l = -2,\ m_s = -1/2$

(Q. No. 7) Four Marks

7. If the position of the electron is measured within an accuracy of ±0.002nm, calculate the uncertainty in the momentum of the electron. Suppose the momentum of the electron is $h/(4\pi \times 0.05)$ nm, is there any problem in defining this value?

(Q. No. 8 to 10) Five Marks

8. What transition in the hydrogen spectrum would have the same wavelength as the Balmer transition, $n = 4$ to
   $n = 2$ of He^+ spectrum?

9. The work function for caesium atom is 1.9 eV. Calculate (a) the threshold wavelength and (b) the threshold frequency of the radiation. If the caesium element is irradiated with a wavelength

500 nm, calculate the kinetic energy and the velocity of the ejected photoelectron.

10.
(i) Calculate the total number of electrons present in one mole of methane.
(ii) Find (*a*) the total number and (*b*) the total mass of neutrons in 7 mg of ^{14}C (Assume that the mass of neutron = 1.675×10^{-27} kg)
(iii) Find (a) the total number and (b) the total mass of protons in 34 mg of NH_3 at S.T.P. (Assume the mass of proton = 1.6726×10^{-27} kg)
Will the answer change if temperature and pressure are changed?

Answer & Solution

Q1

No. of electron are: Na^{+} = 11–1 = 10, K^{+}=19–1 = 18, Mg^{2+} = 12 – 2 = 10, Ca^{2+} = 20–2 = 18, S^{2-}=16+2 = 18, Ar = 18.

Hence, isoelectronis species are N^{a+} and Mg^{2+}; K^{+}, Ca^{2+}, S^{2-} and Ar.

Q2

Charges carried by one electron $=1.6022\times10^{-19}$C

$\therefore$ Electrons present in particle carrying 2.5×10^{-16}C charges $=\dfrac{2.5\times10^{-16}}{1.6022\times10^{-19}}=1560$

Q3

Mass number = 81, *i.e.,* $p + n = 81$

If protons = x, then neutrons $=x+\dfrac{31.7}{100}\times x=1.317x$

$\therefore$ x + 1.317 x=81 or 2.317 x= 81 or $x=\dfrac{81}{2.317}=35$

Thus , protons= 35, *i.e.,* atomic no. = 35

Hence, the symbol is $^{81}_{35}$Br

Q4

Diameter of carbon atom= 0.15 nm = 0.15×10^{-9} m 1.5×10^{-10} m

Length along which atoms are to be placed = 20 cm 20×10^{-2} m = 2×10^{-1} m

$\therefore$ No. of C-atoms which can be placed along the line $=\dfrac{2\times10^{-1}}{1.5\times10^{-10}}=1.33\times10^{9}$

Q5

$$E_n=\frac{21.8\times10^{-9}}{n^2}\text{ J atom}^{-1}$$

For ionization from 5[th] orbit, $n_1 = 5$, $n_2 = \infty$

$\therefore$ $\Delta E = E_2 - E_1 = -21.8\times10^{-19}\left(\dfrac{1}{n_2^2}-\dfrac{1}{n_1^2}\right)= 21.8\times10^{-19}\left(\dfrac{1}{n_1^2}-\dfrac{1}{n_2^2}\right)=21.8\times10^{-19}$

$\left(\frac{1}{5^2}-\frac{1}{\infty}\right)$=**8.72×10^{-20}J**

For ionization from 1st orbit, n_1 = 1, n_2 = ∞

ΔE' =21.8×10^{-19} $\left(\frac{1}{1^2}-\frac{1}{\infty}\right)$=**21.8×10^{-19} J**

$$\frac{\Delta E'}{\Delta E}=\frac{21.8\times10^{-19}}{8.72\times10^{-20}}=25$$

Thus, the energy required to remove electron from 1st orbit is 25 times than that required to removed electron from 5th orbit.

Q6

The orbitals occupied by the electron are

(i) 4 *d*

(ii) 3 *d*

(iii) 4 *p*

(iv) 3 *d*

(v) 3 *p*

(vi) 4 *p*

Their energies will be in the order:

(v) < (ii) = (iv) = < (vi) = (iii) < (i)

Q7

Δ *x* = 0.002 nm = 2×10^{-3} nm = 2×10^{-12} m

$$\Delta x\times\Delta p=\frac{h}{4\pi}\ \therefore\ \Delta p=\frac{h}{4\pi\Delta x}=\frac{6.626\times10^{-34}\ \text{kg m}^2\text{s}^{-1}}{4\times3.14\times(2\times10^{-12}\text{m})}=\mathbf{2.638\times10^{-23}\ kg\ m\ s^{-1}}$$

$$\text{Actual momentum}=\frac{h}{4\pi\times0.05\,\text{nm}}=\frac{h}{4\pi\times5\times10^{-11}\text{m}}=\frac{6.626\times10^{-34}\text{kg m}^2\text{s}^{-1}}{4\times3.14\times5\times10^{-11}\text{m}}$$

= 1.055×10^{-24} kg m s^{-1}

It cannot be defined as the actual magnitude of the momentum is smaller than the uncertainty.

Q8

For H-like particles in general $\overline{V}=\frac{2\pi^2mZ^2e^4}{ch^3}\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right)=RZ^2\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right)$

∴For He$^+$ spectrum, for Balmer transition, *n* = 4 to *n* = 2.

$$\overline{V}=\frac{1}{\lambda}=RZ^2\left(\frac{1}{2^2}-\frac{1}{4^2}\right)=R\times4\times\frac{3}{16}=\frac{3R}{4}$$

For hydrogen spectrum $\overline{V}=\frac{1}{\lambda}=R\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right)=\frac{3}{4}R$ or $\frac{1}{n_1^2}-\frac{1}{n_2^2}=\frac{3}{4}$

which can be so far n^{-1} = 1 and n_2 = 2. *i.e.,* the transition is from ***n* = 2 to n = 1**

Q9

(i) Work function (W_0) = h ν_0

$$\therefore\nu_0=\frac{W_0}{h}=\frac{1.9\times1.602\times10^{-19}\text{J}}{6.626\times10^{-34}\text{J s}}=4.59\times10^{14}s^{-1}$$ (1 eV = 1.602 × 10^{-19}J)

(ii) $\therefore \lambda_0 = \frac{c}{v_0} = \frac{3.0\times10^8\ ms^{-1}}{6.59\times10^{-14}\ s^{-1}} = 6.54\times10^{-7}m$ = 654×10^{-9} m = **654 nm**

(iii) K.E. of ejected electron = $h\ (v\text{-}v_0)$ = $hc\left(\frac{1}{\lambda}-\frac{1}{\lambda_0}\right)$

$= (6.626\times10^{-34}\ \text{J s})\ (3.0\times10^{8}\ \text{m s}^{-1})\left(\frac{1}{500\times10^{-9}\text{m}}-\frac{1}{654\times10^{-9}\text{m}}\right)$

$= \frac{6.626\times3.0\times10^{-26}}{10^{-9}}\left(\frac{154}{500\times654}\right)\text{J} = \mathbf{9.36\times10^{-20}}\ \text{J}$

$\text{K.E.} = \frac{1}{2}m\ v^2 = 9.36\times10^{-20}\text{J or kg m}^2\ \text{s}^{-2}$

$\therefore \quad \frac{1}{2}\times(9.11\times10^{-31}\ \text{kg})\ v^2 = 9.36\times10^{-20}\ \text{kg m}^2\ \text{s}^{-2}$

or $v^2 = 2.055\times10^{11}\ \text{m}^2\ \text{s}^{-2} = 20.55\times10^{10}\ \text{m}^2\ \text{s}^{-2}$ or v=**4.53×10^5ms^{-1}.**

Q10

(i) 1 molecule of CH_4 contains electrons = 6+4=10

$\therefore$ 1 mole, *i.e.,* 6.022×10^{23} molecules will contains electrons = **6.022×10^{23}**

(ii)

(a) 1 g atom of ^{14}C = 14 g = 6.022×10^{23} atoms = $(6.022\times10^{23})\times8$ neutrons.)

(as each ^{14}C atom has 14−6=8 neutrons)

Thus, 14 g or 14000 mg have $8\times6.022\times10^{23}$ neutrons

$\therefore$ 7 mg will have neutron$= \frac{8\times6.022\times10^{23}}{14000}\times7 = 2.4088\times10^{21}$

(b) Mass of 1 neutrons = 1.675×10^{-27} kg

$\therefore$Mass of 2.4088×10^{-21} neutrons = (2.4088×10^{21}) $(1.675\times10^{-27}$ kg) = **4.0347×10^{-6} kg**

(iii)

(a) 1 mol of NH_3 = 17 g NH_3 = 6.022×10^{23} molecules of NH_3

= $(6.022\times10^{23})\times(7+3)$ protons = 6.022×10^{24} protons

$\therefore$ 34 mg, *i.e.,* 0.034 g NH_3 $= \frac{6.022\times10^{24}}{17}\times0.034 = 1.2044\times10^{22}$protons.

(b) Mass of one proton = 1.6726×10^{-27} kg

$\therefore$Mass of 1.2044×10^{22} protons = $(1.6726\times10^{27})\times(1.2044\times10^{22})$ kg = **2.0145×10^{-5} kg**

There is no effect of temperature and pressure.