STRUCTURE OF ATOM

The atomic theory of matter was first proposed on a firm scientific basis by John Dalton. His theory, called Dalton's atomic theory, regarded the atom as the ultimate particle of matter.

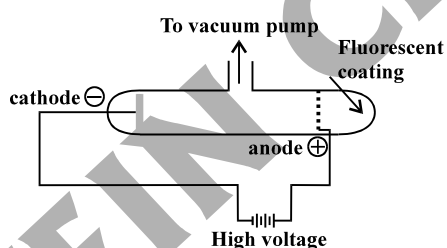
2.1 Sub-Atomic Particles :

Q. What was the results of Dalton's atomic theory ?

Solution : Dalton's atomic theory was able to explain the law of conservation of mass, law of constant composition and law of multiple proportion very successfully. However, it failed to explain the results of many experiments, for example, it was known that substances like glass or ebonite when rubbed with silk or fur generate electricity.

Q. Explain the cathode ray discharge tube experiment. What are the cathode rays and its properties ?

Solution : Faraday began to study electrical discharge in partially evacuated tubes, known as cathode ray discharge tubes. It is depicted in figure. A cathode ray tube is made of glass containing two thin pieces of metal, called electrodes, sealed in it. The electrical discharge through the gases could be observed only at very low pressures and at very high voltages. The pressure of different gases could be adjusted by evacuation. When sufficiently high voltage is applied across the electrodes, current starts flowing through a steam of particles moving in the tube from the negative electrode (cathode) to the positive electrode (anode). These were called cathode rays or cathode ray particles. The flow of current from cathode to anode was further checked by making a hole in the anode and coating the tube behind anode with phosphorescent material zinc sulphide. When these rays, after passing through anode, strike the zinc sulphide coating, a bright spot on the coating is developed (same thing happens in a television set).



The results of these experiments are summarised below.

(i) The cathode rays start from cathode and move towards the anode.

(ii) These rays themselves are not visible but their behaviour can be observed with the help of certain kind of materials (fluorescent or phosphorescent) which glow when hit by them. Television picture tubes are cathode ray tubes and television pictures result due to fluoroscence on the television screen coated with certain fluorescent or phosphorescent materials.

(iii) In the absence of electrical or magnetic field, these rays travel in straight lines.

(iv) In the presence of electrical or magnetic field, the behaviour of cathode rays are similar to that expected from negatively charged particles, suggesting that the cathode rays consists of negatively charged particles, called electrons.

(v) The characteristics of cathode rays (electrons) do not depend upon the material of electrodes and the nature of the gas present in the cathode ray tube. Thus, we can conclude that electrons are basic constituent of all the atoms.

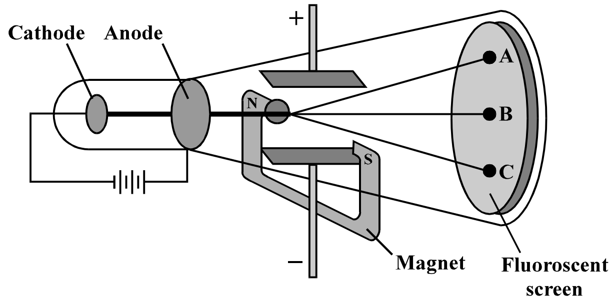
Q. Explain the experiment by J.J. Thomson which is used to find the charge to mass ratio of electron.

Solution : British physicist J.J. Thomson measured the ratio of electrical charge (e) to the mass of electron (m_e) by using cathode ray tube and applying electrical and magnetic field perpendicular to each other as well as to the path of electrons. Thomson argued that the amount of deviation of the particles from their path in the presence of electrical or magnetic field depends upon :

(i) the magnitude of the negative charge on the particle, greater the magnitude of the charge on the particle, greater is the interaction with the electric of magnetic field and thus greater is the deflection.

(ii) the mass of the particle – lighter the particle, greater the deflection.

(iii) the strength of the electrical or magnetic field – the deflection of electrons from its original path increases with the increase in the voltage across the electrodes, or the strength of the magnetic field.



When only electric field is applied, the electrons deviate from their path and hit the cathode ray tube at point A. Similarly when only magnetic field is applied, electron strikes the cathode ray tube at point C. By carefully balancing the electrical and magnetic field strength, it is possible to bring back the electron to the path followed as in the absence of electric or magnetic field and they hit the screen at point B. By carrying out accurate measurements on the amount of deflections observed by the electrons on the electric field strength or magnetic field strength, Thomson was able to determine the value of e/m_e as :

$$\frac{\mathbf{e}}{\mathbf{m_e}} = \mathbf{1.758820 \times 10^{11}\,C\,kg^{-1}}$$

Where m_c is the mass of the electron in kg and e is the magnitude of the charge on the electron in coulomb (C).

Q. What is the name of method used to determine the charge on electron ? How was the mass of electron was calculated ?

Solution : R.A. Millikan devised a method known as oil drop experiment, to determine the charge on the electrons. He found that the charge on the electron to be -1.6×10^{-19} C. The present accepted value of electrical charge is -1.6022×10^{-19} C. The mass of the electron (m_e) was determined by combining these results with Thomson's value of e/m_e ratio.

$$\mathbf{m_e = \frac{e}{e/m_e} = \frac{1.6022 \times 10^{-19}\,C}{1.758820 \times 10^{11}\,C\,kg^{-1}}} = 9.1094 \times 10^{-31}\ \text{kg.}$$

Q. What are the Canal Rays ? Also write its properties.

Solution : Electrical discharge carried out in the modified cathode ray tube led to the discovery of particles carrying positive charge, also known as canal rays. The characteristics of these positively charged particles are listed below.

(i) unlike cathode rays, the positively charged particles depend upon the nature of gas present in the cathode ray tube. These are simply the positively charged gaseous ions.

(ii) The charge to mass ratio of the particles is found to depend on the gas from which these originate.

(iii) Some of the positively charged particles carry a multiple of the fundamental unit of electrical charge.

(iv) The behaviour of these particles in the magnetic or electric field is opposite to that observed for electron or cathode rays.

The smallest and lightest positive ion was obtained from hydrogen and was called proton.

Q. Who has discovered neutrons ?

Solution : These particles were discovered by Chadwick (1932) by bombarding a thin sheet of beryllium by α-particles. When electrically neutral particles having a mass slightly greater than that of the protons was emitted. He named these particles as neutrons.

2.2 Atomic Models :

Q. Write the properties of Electron, Proton and Neutron i.e., charge, mass and charge relative to proton.

Solution :

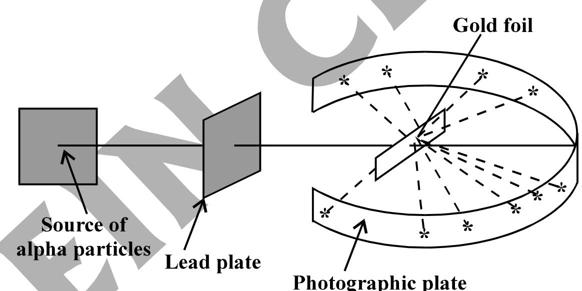
	Proton (p)	Neutron (n)	Electron (e)
Mass	1.67252×10^{-27} kg	1.67482×10^{-27} kg	9.1091×10^{-31} kg
Charge	1.60210×10^{-19} C	0	1.60210×10^{-19} C
Mass relative to the electron	1836	1839	1
Charge relative to the proton	+1	0	–1

Q. Explain the Thomson Model of Atom.

Solution : J.J. Thomson, in 1898, proposed that an atom possesses a spherical shape (radius approximately 10^{-10} m) in which the positive charge is uniformly distributed. The electrons are embedded into it in such a manner as to give the most stable electrostatic arrangement. Many different names are given to this model, for example, plum pudding, raisin pudding or watermelon. This model can be visualised as a pudding or watermelon of positive charge with plums or seeds (electrons) embedded into it. An important feature of this model is that the mass of the atom is assumed to be uniformly distributed over the atom. Although this model was able to explain the overall neutrality of the atom, but was not consistent with the result of later experiments.

Q. Explain the Rutherford's Nuclear Model of atom observation and conclustion.

Solution :



Rutherford and his students (Hans Geiger and Ernest Marsden) bombarded very thin gold foil with α-particles scattering experiment is represented as above. A steam of high energy α-particle from a radioactive source was directed at a thin foil (thickness ~ 100 nm) of gold metal. The thin gold foil had a circular fluorescent zinc sulphide screen around it. Whenever α-particles struck the screen, a tiny flash of light was produced at that point. It was observed that :

(i) most of the α-particle passed through the gold foil undeflected.

(ii) a small fraction of the α-particle was deflected by small angles.

(iii) a very few α-particles (~1 in 20,000) bounced back, that is, were deflected by nearly 180^0.

On the basis of the observations, Rutherford drew the following conclusions regarding the structure of atom :

(i) Most of the space in the atom is empty as most of the α-particles passed through the foil undeflected.

(ii) A few positively charged α-particles were deflected. The deflection must be due to enormous repulsive force showing that the positive charge of the atom is not spread throughout the atom as Thomson had presumed. The positive charge has to be concentrated in a very small volume that repelled and deflected the positively charged α-particles.

(iii) Calculations by Rutherford showed that the volume occupied by the nucleus is negligibly small as compared to the total volume of the atom. The radius of the atom is about 10^{-10} m, while that of nucleus is 10^{-15} m. One can appreciate this difference in size by realising that if a cricket ball represented a nucleus, then the radius of atom would be about 5 km.

Q. On the basis of Rutherford experiment explain the nuclear model of atom.

Solution : Rutherford 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.

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

Q. Define Atomic Number and Mass Number.

Solution : Atomic number (z) = number of protons in the nucleus of an atom

= number of electrons in a nuetral atom.

While the positive charge of the nucleus is due to protons, the mass of the nucleus, due to protons and neutrons. Protons and neutrons present in the nucleus are collectively known as nucleons. The total number of nucleons is termd as mass number (A) of the atom.

mass number (A) = number of protons (Z) + number of neutrons (n)

Q. What are Isobars and Isotopes ?

Solution : Isobars are the atoms with same mass number but different atomic number for example, ${}^{14}_{6}C$ and ${}^{14}_{7}N$. On the other hand, atoms with identical atomic number but different atomic mass number are known as isotopes. In other words it is evident that difference between the isotopes is due to the presence of different number of neutrons present in the nucleus.

Q. Calculate the number of protons, neutrons and electrons in ${}^{80}_{35}Br$. [NCERT Solved Example 2.1].

Solution : 35, 45

Q. The number of electrons, protons and neutrons in a species are equal to 18, 16 and 16 respectively. Assign the proper symbol to the species. [NCERT Solved Example 2.2]

Solution : 16, S, 32, ${}^{32}_{6}S^{2-}$.

Q. What are the drawbacks of Rutherford’s Model ?

Solution : (i) Inability to explain the stability of atom : According to Maxwell’s electromagnetic theory, whenever a charged particle like electron is revolving in a field of force like that of the nucleus, it loses energy continuously in the form of electromagnetic radiations. This is because when a particle is revolving, it undergoes acceleration due to change in direction even if the speed remains constant. Thus, the orbit of the revolving electron will keep on becoming smaller and smaller, following a spiral path and ultimately the electron should fall into the nucleus. In other words, the atom should collapse. However, this actually does not happen and the atom is quite stable. Thus, Rutherford model could not explain the stability of the atom.

(ii) Inability to explain the line spectra of the elements : It could not explain why elements produce line spectra.

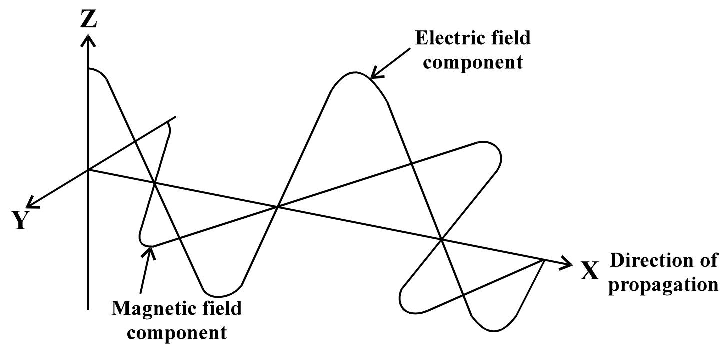
(iii) Inability to describe distribution of electrons and energies of electrons : It was unable to explain how the electrons are distributed around the nucleus and what were their energies.

2.3 Developments Leading to Bohr’s Model of Atom :

Q. Explain the properties of electromagnetic radiation.

Solution : Maxwell was again the first to reveal the light waves are associated with oscillating electric and magnetic character.

(i) The oscillating electric field 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. Simplified picture of electromagnetic wave is shown in figure.



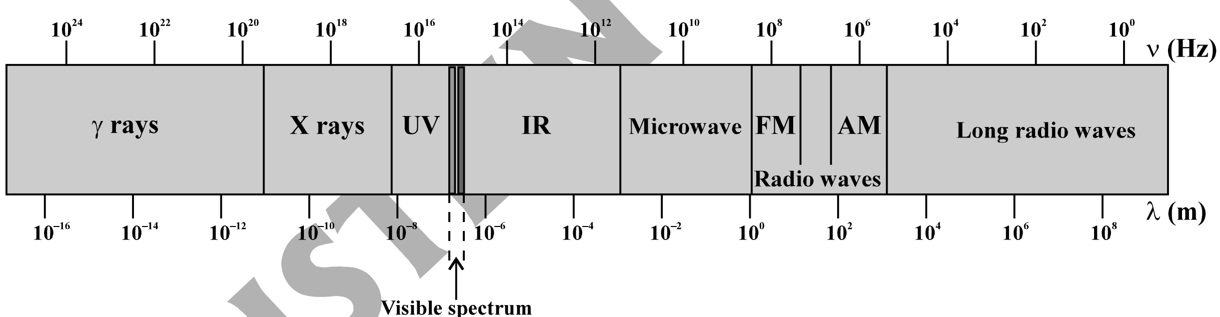
(ii) Unlike sound waves or water waves, electromagnetic waves do not require medium and can move in vacuum.

(iii) There are many types of electromagnetic radiations, which differ from one another in wavelength (or frequency). These constitute what is called electromagnetic spectrum.

These radiations are characterised by the properties, namely, frequency (v) and wavelength (λ).

Q. Explain the Electromagnetic Spectrum.

Solution : There are many types of electromagnetic radiations, which differ from one another in wavelength (or frequency). These constitute what is called electromagnetic spectrum. Different regions of the spectrum are identified by different names. Some examples are : radio frequency region around 10^6 Hz, used for broadcasting; microwave region around 10^{10} Hz used for radar; infrared region around 10^{13} Hz used for heating; unltraviolet region around 10^{16} Hz a component of sun's radiation. The small portion around 10^{15} Hz, is what is ordinarily called visible light. It is only this part which our eyes can seen (or direct). Special instruments are required to detect non-visible radiation.



Q. Define frequency.

Solution : The SI unit for frequency (v) is hertz (Hz, s^{-1}), after Heinrich Hertz. It is defined as the number of waves that pass a given point in one second.

Q. Give the relation between wavelength and frequency and velocity of light for electromagnetic ratiation.

Solution : In vaccum all types of electromagnetic radiations, regardless of wavelength, travel at the same speed, i.e., 3.0×10^8 m s^{-1} (2.997925×10^8 m s^{-1}, to be pracise). This is called speed of light and is given the symbol 'c'. The frequency (v), wavelength (λ) and velocity of light (c) are related by the equation

$c = \nu \lambda$

Q. Define wave number.

Solution : It is defined as the number of wavelengths per unit length. Its units are reciprocal of wavelength unit, i.e., m^{-1}.

Q. 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 ? [NCERT Solved Example 2.3]

Solution : 219.3 m

Q. 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) [NCERT Solved Example 2.4].

Solution : 4.0×10^{14} to 7.5×10^{14} Hz.

Q. Calculate (a) wavenumber and (b) frequency of yellow radiation having wavelength 5800 Å. [NCERT Solved Example 2.5].

Solution : (a) 1.724×10^{4} cm^{-1} (b) 5.172×10^{14} s^{-1}.

Q. Explain the Planck's Quantum Theory.

Solution : (i) The radiant energy is emitted or absorbed not continuously but discontinuously in the form of small discrete packets of energy. Each such packet of energy is called a 'quantum'. In case of light, the quantum of energy is called a 'photon'.

(ii) The energy of each quantum is directly proportional to the frequency of the radiation, i.e.,

$$E \propto v \quad \text{or} \quad E = hv$$

where h is a proportionality constant, called Planck's constant. Its value of approx. equal to 6.626×10^{-27} erg sec or 6.626×10^{-34} joule sec.

(iii) The total amount of energy emitted or absorbed by a body will be some whole number quanta.

Hence, E = nhv where n is any integer i.e., n = 1, 2, 3, ...

Q. Explain Black Body Radiation.

Solution : Ideal black body, which emits and absorbs all frequencies is called black body and the radiation emitted by such body is called black body radiation.

Q. Explain Photoelectric effect.

Solution : (i) The electrons are ejected from the metal surface as soon as the beam of light strikes the surface, i.e., there is no time lag between the striking of light beam and the ejection of electrons from the metal surface.

(ii) The number of electrons ejected is proportional to the intensity or brightness of light.

(iii) For each metal, there is a characteristic minimum frequency, v_0 (also known as threshold frequency) below which photoelectric effect is not observed. At a frequency $v > v_0$, the ejected electrons come out with certain kinetic energy. The kinetic energies of these electrons increases with the increase of frequency of the light used.

$$\mathbf{hv = hv_0 + \frac{1}{2} m_e v^2}$$

$h\nu = \phi$ + K.E.................$\phi \rightarrow$ work function or threshold energy

K.E. $\rightarrow$ kinetic energy of electron

where m_e is the mass of the electron and v is the velocity associated with the ejected electron. Lastly, a more intense beam of light consists of larger number of photons, consequently the number of electrons ejected is also larger as compared to that in an experiment in which a beam of weaker intensity of light is employed.

Q. Explain the dual nature of electromagnetic radiation.

Solution : Some properties of light (or in general, any radiation) like interference, diffraction etc, can be explained only if we consider light to have wave nature whereas some other properties of light or any other radiation such as black body radiation and photoelectric effect can be explained if we consider light to have particle nature. Thus, light is said to have a dual nature, i.e., it behaves as a wave as well as a steam of particles.

Q. Calculate energy of one mole of photons of radiation whose frequency is 5×10^{14} Hz. [NCERT Solved Example 2.6]

Solution : 199.51 kJ mol^{-1}.

Q. A 100 watt bulb emits monochromatic light of wavelength 400 nm. Calculate the number of photons emitted per second by the bulb. [NCERT Solved Example 2.7]

Solution : 2.012×10^{20} s^{-1}.

Q. When electromagnetic radiation of wavelength 300 nm falls on the surface of sodium, electrons are emitted with a kinetic energy of 1.68×10^5 J mol^{-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 : 517 nm

Q. The threshold frequency v_0 for a metal is 7.0×10^{14} s^{-1}. Calculate the kinetic energy of an electron emitted when radiation of frequency $v = 1.0 \times 10^{15}$ s^{-1} hits the metal.

Solution : 1.988×10^{-19} J

Q. What is the difference between emission spectra and absorption spectra ?

Solution : Emission Spectrum : (i) Emission spectrum is obtained when the radiation from the source are directly analysed in the spectroscope.

(ii) It consists of bright coloured lines separated by dark spaces.

(iii) Emission spectrum can be continuous spectrum (if source emits white light) or discontinuous, i.e., line spectrum if source emits some coloured radiation.

Absorption Spectrum : (i) Absorption spectrum is obtained when the white light is first passed through the substance and the transmitted light is analysed in the spectroscope.

(ii) It consists of dark lines in the otherwise continuous spectrum.

(iii) Absorption spectrum is always discontinuous spectrum consisting of dark lines.

Q. What is quantization ?

Solution : The restriction of any property to discrete values is called quantization.

Q. (a) Define Spectrum ? (b) What is the range for the spectrum of white light ? (c) What is continuous spectrum ?

Solution : White light consists of waves with all the wavelengths in the visible range, a ray of white light is spread out into a series of coloured bands called spectrum. The spectrum of white light, that we can see, ranges from violet at 7.50×10^{14} Hz to red at 4×10^{14} Hz. Such spectrum is called continuous spectrum. Continuous because violet merges into blue, blue into green and so on.

Q. Which one, red colour or violet colour deviated least ? Give reason.

Solution : The light of red colour which has longest wavelength is deviated the least while the violet light, which has shortest wavelength is deviated the most.

Q. Define Emission spectrum.

Solution : When electromagnetic radiation interacts with matter, atoms and molecules may absorb energy and reach to a higher energy state. With higher energy, these are in an unstable state. For returning to their normal (more stable, lower energy states) energy state, the atoms and molecules emit radiations in various regions of the electromagnetic spectrum.

The spectrum of radiation emitted by a substance that has absorbed energy is called an emission spectrum.

Q. How can you produce emission spectrum ?

Solution : To produce an emission spectrum, energy is supplied to a sample by heating it or irradiating it and the wavelength (or frequency) of the radiation emitted, as the sample gives up the absorbed energy, is recorded.

Q. What is absorption spectrum ?

Solution : A continuum of radiation is passed through a sample which absorbs radiation of certain wavelenghs. The missing wavelength which corresponds to the radiation absorbed by the matter, leave dark spaces in the bright continuous spectrum.

Q. What is spectroscopy ?

Solution : The study of emission or absorption spectra is referred to as spectroscopy.

Q. Define line spectra or atomic spectra ?

Solution : The emission spectra of atoms in the gas phase consists of specific wavelengths with dark spaces between them. Such spectra are called line spectra or atomic spectra because the emitted radiation is identified by the appearance of bright lines in the spectra.

Q. What is the use of study of Line spectra ? Explain.

Solution : Line emission spectra are used in the study of electronic struture. Each element has a unique line emission spectrum. The characteristic lines in atomic spectra can be used in chemical analysis to identify unknown atoms.

Q. Explain the Line spectrum of hydrogen.

Solution : When an electric discharge is passed through gaseous hydrogen, the H_2 molecules dissociate and the energetically excited hydrogen atoms produced emit electromagnetic radiation of discrete frequencies. The hydrogen spectrum consists of several series of lines named after their discoverers.

Series	$\mathbf{n_1}$	$\mathbf{n_2}$	**Region of Spectrum**
Lyman	1	2, 3, ... ∞	Ultraviolet
Balmer	2	3, 4, ... ∞	Visible
Paschen	3	4, 5, ... ∞	Infrared
Brackett	4	5, 6, ... ∞	Infrared
Pfund	5	6, 7, ... ∞	Infrared

Maximum number of emission lines which can be obtained when excited electron in nth orbit jumps to the ground state are $\mathbf{\frac{n(n-1)}{2}}$.

Rydberg, noted that all series of lines in the hydrogen spectrum could be described by the following expression :

$$\bar{v} = 1.09 \times 10^7 \left(\frac{1}{n^2} - \frac{1}{n_2^2} \right) m^{-1} \quad (\bar{v} \text{ is wave number})$$

where $n_1 = 1, 2.....$

$n_2 = n_1 + 1, n_1 + 2 ....$

The value $1.09 \times 10^7\ m^{-1}$ is called the Rydberg constant for hydrogen.

2.4 Bohr Model for Hydrogen Atom :

Q. Write the Postulates of Bohr's Model of Atom.

Solution : The main postulates of Bohr's model of atom are as follows :

(i) An atom consists of a small, heavy positively charged nucleus in the centre and the electrons revolve around it in circular orbits.

(ii) Out of a large number of circular orbits theoretically possible around the nucleus, the electrons revolve only in those orbits which have a fixed value of energy. Hence, these orbits are called energy levels or stationary states. The word stationary does not mean that the electrons are stationary but it means that the energy of the electron revolving in a particular orbit is fixed and does not change with time. The different energy levels are numbered as 1, 2, 3, 4...etc, or designated as K, L, M, N, O, P.....etc, starting from the shell closest to the nucleus.

$$E_n = -\frac{21.8 \times 10^{-19}}{n^2} J/atom = -\frac{13.6}{n^2} eV/atom$$

The radii of the stationary states of the hydrogen atom are given by the expression

$r_n = a_0 n^2$

For H-like particles, the radii of the stationary states are given by the expression

$$r_n = \frac{a_0 n^2}{Z}, \text{ where } a_0 = 52.9 \text{ pm}$$

(iii) Since the electrons revolve only in those orbits which have fixed values of energy, hence electrons in an atom can have only certain definite or discrete values of energy and not any value of their own. This is expressed by saying that the energy of an electron is quantized.

(iv) Like energy, the angular momentum of an electron in an atom can have certain definite or discrete values and not any value of its own. The only permissible values of angular momentum are given by the expression :

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

i.e., angular momentum of the electron is an integral multiple of h/2 π. Here, m is the mass of the electron, v is the tangential velocity of the revolving electron, r is the radius of the orbit, h is the Planck's constant and n is any integer. In other words, the angular momentum of the electron can be h/2π, 2h/2π, 3h/2π...etc. This means that like energy, the angular momentum of an electron in an atom is also quantized.

(v) When the electrons in an atom are in their lowest (normal) energy state, they keep on revolving in their respective orbits without losing energy because energy can neither be lost not gained continuously. This state of atom is called normal or ground state.

(vi) Energy is emitted or absorbed only when the electrons jump from one orbit to the other. For example, when energy is supplied to an atom by subjecting it to electric discharge or high temperature, an electron in the atom may jump from its normal energy level (ground state) to some higher energy level by absorbing a definite amount of energy. This state of atom is called excited state. Since the life time of the electron in the excited state is short, it immediately jumps back to the lower energy level by emitting energy in the form of light of suitable frequency or wavelength. The amount of energy emitted or absorbed is given by the difference of energies of the two energy levels concerned, i.e., $\Delta E = E_2 - E_1$.

Q. Write the expression of radius of stationary state and energy of stationary state in terms of Rydberg constant.

Solution : Radius of nth orbit

$$r_n = 0.53\frac{n^2}{Z}\text{ Å} \text{ where Z = atomic number}$$

Energy of the electron in the nth orbit

$$E_n = -13.6\frac{Z^2}{n^2}\text{ ev},\quad E_n = -(2.18\times10^{-18})\frac{Z^2}{n^2}(J)$$

$E_n = -R_H\frac{z^2}{n^2}J$ where R_H is called Rydberg constant and its value is 2.18 × 10^{-18} J.

Q. What does the negative electronic energy (E_n) for hydrogen atom mean ?

Solution : This negative sign means that the energy of the electron in the atom is lower than the energy of a free electron at rest. A free electron at rest is an electron that is infinitely far away from the nucleus and is assigned the energy value of zero. Mathematically, this corresponds to setting n equal to infinity so that $E_\infty = 0$. As the electron gets closer to the nucleus (as n decreases), E_n becomes larger in absolute value and more and more negative. The most negative energy value is given by n = 1 which corresponds to the most stable orbit. We call this the ground state.

Q. What are the frequency and wavelength of a photon emitted during a transition from n = 5 state to the n = 2 state in the hydrogen atom ? [NCERT Solved Examples 2.10]

Solution : 6.91 × 10^{14} Hz, 434 nm

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

Solution : –8.72 × 10^{-18} J, 0.02645 nm

Q. What are limitations of Bohr's Model ?

Solution : Bohr's model of the hydrogen atom was no doubt an improvement over Rutherford's nuclear model, as it could account for the stability and line spectra of hydrogen atom and hydrogen like ions (for example, He^+, Li^{2+}, Be^{3+}, and so on. However, Bohr's model was too simple to account for the following points.

(i) It fails to account for the finer details (doublet, that is two closely spaced lines) of the hydrogen atom spectrum observed by using sophisticated spectroscopic techniques. This model is also unable to explain the spectrum of atoms other than hydrogen, for example, helium atom which possesses only two electrons.

Further, Bohr's theory was also unable to explain the splitting of spectral lines in the presence of magnetic field (Zeeman effect) or an electric field (Stark effect).

(ii) It could not explain the ability of atoms to form molecules by chemical bonds.

2.5 Towards Quantum Mechenical Model of the Atom :

Q. Explain the Dual Behaviour of Matter.

Solution : de Broglie proposed that matter, like radiation, should also exhibit dual behaviour i.e., both particle and wavelike properties. This means that just as the photon has momentum as well as wavelength, electrons should also have momentum as well as wavelength, de Broglie, from this analogy, gave the following relation between wavelength (λ) and momentum (p) of a material particle $\lambda = \frac{h}{mv} = \frac{h}{p}$, where m is the mass of the particle, v its velocity and p its momentum.

Q. What will be the wavelength of a ball of mass 0.1 kg moving with a velocity of 10 m s^{-1} ? [NCERT Solved Example 2.12]

Solution : 6.626×10^{-34} m

Q. The mass of an electron is 9.1 × 10^{-31} kg. If its K.E. is 3.0 × 10^{-25} J, calculate its wavelength. [NCERT Solved Example 2.13]

Solution : 896.7 nm

Q. Calculate the mass of a photon with wavelengths 3.6 Å. [NCERT Solved Example 2.14]

Solution : 6.135×10^{-29} kg

Q. Explain the Heisenberg's Uncertainty Principle.

Solution : It states that it is impossible to determine simultaneously, the exact position and exact momentum (or velocity) of an electron.

Mathematically, it can be given as $\Delta x \times \Delta p_x \geq \frac{h}{4\pi}$ or $\Delta x \times \Delta(mv_x) \geq \frac{h}{4\pi}$ or $\Delta x \times \Delta v_x \geq \frac{h}{4\pi m}$

where Δx is the uncertainty in position and Δp_x (or Δv_x) is the uncertainty in momentum (or velocity) of the particle. If the position of the electron is known with high degree of accuracy (Δx is small), then the velocity of the electron will be uncertain [$\Delta(v_x)$ is large.

Q. A microscope using suitable photons is employed to locate an electron in an atom within a distance of 0.1 Å. What is the uncertainty involved in the measurement of its velocity ? [NCERT Solved Example 2.15]

Solution : 5.79×10^{6} m s^{-1}

Q. A golfball 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. [NCERT Solved Example 2.16].

Solution : 1.46×10^{-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 measurements].

2.6 Quantum Mechanical Model of Atom :

Q. What are the important features of the Quantum Mechanical Model of Atom ?

Solution : Quantum mechanical model of atom is the picture of the structure of the atom, which emerges from the application of the Schrodinger equation to atoms. The following are the important features of the quantum mechanical model of atom :

(i) The energy of electrons in atoms is quantized (i.e., can only have certain specific values), for example when electrons are bound to the nucleus in atoms.

(ii) The existence of quantized electronic energy levels is a direct result of the wave like properties of electrons and are allowed solutions of Schrodinger wave equation.

(iii) Both the exact position and exact velocity of an electron in an atom cannot be determined simultaneously (Heisenberg uncertainty principle). The path of an electron in an atom therefore, can never be determined or known accurately. That is why, as you shall see later on, one talks of only probability of finding the electron at different points in an atom.

(iv) An atomic orbital is the wave function ψ for an electron in an atom. Whenever an electron is described by a wave function, we say that the electron occupies that orbital. Since many such wave functions are possible for an electron, there are many atomic orbitals in an atom. These "one electron orbital wave functions" or orbitals form the basis of the electronic structure of atoms. In each orbital, the electron has a definite energy. An orbital cannot contain more than two electrons. In a multi-electron atom, the electrons are filled in various orbitals in the order of increasing energy. For each electron of a multi-electron atom, there shall, therefore, be an orbital wave function characteristic of the orbital it occupies. All the information about the electron in an atom is stored in its orbital wave function ψ and quantum mechanics makes it possible to extract this information out of ψ.

(v) The probability of finding an electron at a point within an atom is proportional to the square of the orbital wave function i.e., $|\psi|^2$ at the point. $|\psi|^2$ is known as probability density and is always positive. From the value of $|\psi|^2$ at different points within an atom, it is possible to predict the region around the nucleus where electron will most probably be found.

Q. What is the difference between orbit and orbital ?

Solution :

	Orbit		**Orbital**
1.	It is circular or eliptical path traced by an electron while revolving round the nucleus of atom giving fixed value of the distance of e^- from the nucleus.	1.	It is the region in space where there is high probability of finding the electron.
2.	It violates the Heisenberg principle	2.	It does not violate the Heisenberg principle
3.	It is not in accordance to the dual character of matter	3.	It is in accordance of dual character of matter.

Q. Explain briefly the four quantum numbers.

Solution : Quantum numbers : Each orbitals is designated by three quantum number n, *l* and m.

The principle quantum number (n) :

It determines the size and to a large extent the energy of the orbital. The larger the value of n, the larger the energy of the orbital. Principle quantum number also identifies the shell i.e., $\mathbf{n} = \underset{K}{\underset{\uparrow}{1}}, \underset{L}{\underset{\uparrow}{2}}, \underset{M}{\underset{\uparrow}{3}}, \underset{N}{\underset{\uparrow}{4}}$ shell.

There are n^2 orbitals in a shell. All the orbitals of a given volume of n constitute a single shell of atom. Each shell consists of one or more subshells or sublevels. The number of subshells in a principal shell is equal to the value of n.

Azinuthal or Subsidiary quantum number : (*l*)

Each subshell in a shell is designated by *l*. *l* can have n values ranging from o to (n – 1) e.g. when n = 1, *l* = 0, n = 2, *l* = 0, 1 etc. i.e. for n = 1, *l* = 0 it means there is one subshell for n = 2, *l* = 0, 1. It means there are two subshell.

Subshells corresponding to different values of *l* are represented by following symbols

l =	0, 1, 2, 3, 4, 5
notations	s, p, d, f, g, h

Each subshell consists of one or more orbitals. The number of orbitals in a subshell is given by $(2l + 1)$ e.g. In any $l = 0$ subshell, there are 2(0) + 1 = 1 orbitals. For $l = 1$ subshell there are 2(1) + 1 = 3 orbitals. In any $l = 2$ subshell, there are 2(2) + 1 = 5 orbitals.

In other words,

subshell notation	=	s, p, d, f, g
value of *l*	=	0, 1, 2, 3, 4
Number of orbitals	=	1, 3, 5, 7, 9

The orbital angular momentum = $\sqrt{l(l+1)}\,\dfrac{h}{2\pi}$

The quantum number *l* also gives the shape of the orbital in the subshell.

Magnetic quantum number (m) :

It gives information about the orientation of the orbital.

For any subshell (defined by l values), $(2l + 1)$ values of m are possible and these values are given as

$m = -l, -(l-1), ...\ O, ...\ +(l-1), +l$

e.g. for $l = 2$(d), m can have total $= 2(2) + 1 = 5$ values. These values are –2, –1, 0, +1, +2 [Five orbitals]

For $l = 1$(p) $m = 2(1) + 1 = 3$ values. These values are –1, 0, +1 [Three orbitals] i.e., there are three preffered orientation of p in space. Thus each orbital is defined by set of values of n, l and m

e.g. If 4s is given then $n = 4$, $l = 0$

If 5p is given then $n = 5$, $l = 1$

Spin quantum number : Electrons spin about the axes. Some spinning in one direction and some in other direction. Two orientation of electron which is possible is one is in clockwise direction and other one is in anticlockwise direction. They are represented by two arrows ↑ (spin up) and ↓ (spin down). The two spins have either +½ value or –½ value.

Q. What is the total number of orbitals associated with the principal quantum number n = 3 ? [NCERT Solved Example 2.17]

Solution : 9

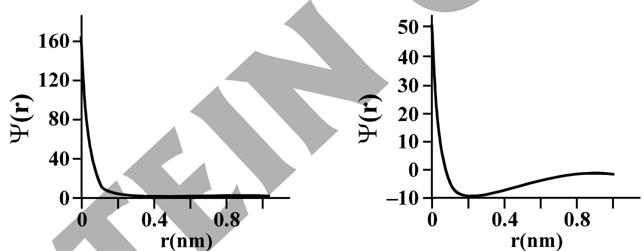
Q. Using s, p, d, f notations, describe the orbitals with the following quantum numbers

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

Solution : 2p, 4s, 5f, 3d

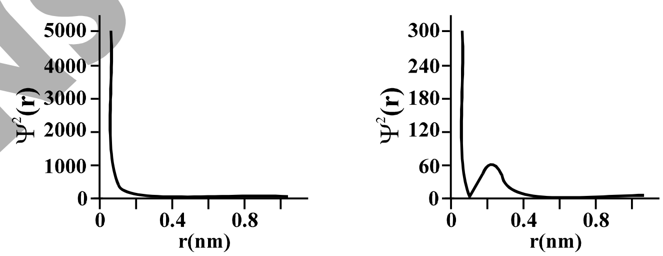
Q. Draw the plot of orbital wave function Ψ(r) and Ψ²(r) as a function of distance 'r' of the electron from the nucleus for 1s and 2s orbitals.

Solution : (a)



1s 2s

(b)



1s 2s

For 1s orbital the probability density is maximum at the nucleus and it decreases sharply as we move away from it. On the other hand, for 2s orbital the probability density first decreases sharply to zero and again starts increasing. After reaching a small maxima it decreases again and approaches zero as the value of r increases further. The region where this probability density function reduces to zero is called nodal surfaces or simply nodes.

Q. Define nodes.

Solution : The region where probability density of electron reduces to zero is called nodal surfaces or simply nodes. The total number of nodes are given by (n – 1), i.e., sum of l angular nodes and $(n - l - 1)$ radial nodes.

Q. Define the rules for filling of electron into orbitals of an atom briefly.

Solution : The filling of electron into different orbitals takes place according to following three rules :

1. Aufbau Principle : In the ground state of the atoms, the orbitals are filled in order of their increasing energies.

In other words,electrons first occupy the lowest energy orbital available to then and then enter to higher energy orbital.

The order of increase of energy of orbitals can be calculated by $(n + l)$ rule.

Lower the value of $(n + l)$ for an orbital, the lower is its energy.

If two orbitals have the same $(n + l)$ value, the orbital with lower value of n has the lower energy.

2. Pauli Exclusion Principle : The number of electrons to be filled in various orbitals is restricted by this principle.

No two electrons in an atom can have the same set of four quantum numbers.

e.g. If an electron in an atom has particular set of quantum number say $n = 1$, $l = 0$, $m = 0$, $s = +½$ then no other electron in an atom can have this set of quantum number.

3. Hund's Rule of maximum multiplicity : This rule deals with the filling of electrons into orbitals belonging to same subshell.

Orbital of same subshells like p_x, p_y, p_z are of equal energy and these are called as degenrate orbitals. Similarly d has five orbitals of same energy i.e., d_{xy}, d_{yz}, d_{zx}, $\mathbf{d_{x^2-y^2}}$, $\mathbf{d_{z^2}}$, f has 7 degenrate orbitals etc.

Hund's Rule States that : pairing of electrons in the orbitals belonging to the same subshell (p, d or f) does not takes place until each orbital belonging to that subshell has got one electron each i.e. is singly occupied.

e.g. $\overset{\uparrow}{\mathbf{p_x}}$ $\overset{\uparrow}{\mathbf{p_y}}$ $\overset{\uparrow}{\mathbf{p_z}}$ as there are 3-orbitals the pairing of electron starts in p with the entry of 4th electron.

Q. Write the increasing order of energies of orbitals according to $(n + l)$ rule.

Solution : 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p....

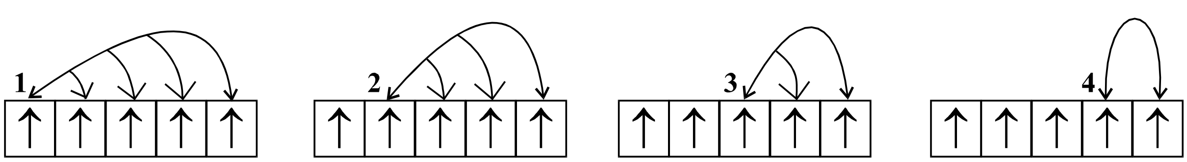
Q. What is the cause of stability of completely filled and half filled sub-shells ?

Solution : The completely filled and completely half filled sub-shells are stable due to the following reasons :

1. Symmetrical distribution of electrons : It is well known that symmetry leads to stability. The completely filled or half filled subshells have symmetrical distribution of electrons in them and are therefore more stable. electrons in the same subshell (here 3d) have equal energy but different spatial distribution. Consequently, their shielding of one-another is relatively small and the electrons are more strongly attracted by the nucleus.

2. Exchange Energy : The stabilizing effect arises whenever two or more electrons with the same spin are present in the degenerate orbitals of a subshell. These electrons tend to exchange their positions and the energy released due to this exchange is called exchange energy. The number of exchanges that can take place is maximum when the subshell is either half filled or completely filled. As a result the exchange energy is maximum and so is the stability.

You may note that the exchange energy is at the basis of Hund's rule that electrons which enter orbitals of equal energy have parallel spins as far as possible. In other words, the extra stability of half-filled and completely filled subshell is due to : (i) relatively small shielding, (ii) smaller coulombic repulsion energy, and (iii) larger exchange energy.



4 exchange by electron 1 **3 exchange by electron 2** **2 exchange by electron 3** **1 exchange by electron 4**

Possible exchange for a d^5 configuration

NCERT EXERCISE

2.1 (i) Calculate the number of electrons which will together weighted one gram.

(ii) Calculate the mass and charge of one mole of electrons.

2.2 (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 mass of a 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 STP. Will the answer change if the pressure and temperature are changed ?

2.3 How many neutrons and protons are there in the following nuclei ?

$$^{13}_{6}C,\ ^{16}_{8}C,\ ^{24}_{12}Mg,\ ^{56}_{26}Fe,\ ^{88}_{38}Sr$$

2.4 Write the complete symbol for the atom with the given atomic number (Z) and atomic mass (A)

(i) Z = 17, A = 35, (ii) Z = 92, 233, (iii) Z = 4, A = 9.

2.5 Yellow light emitted from a sodium lamp has a wavelength of (λ) of 580 nm. Calculate the frequency (ν) and wave number $(\bar{\nu})$ of yellow light.

2.6 Find the energy of each of the photons which

(i) correspond to light of frequency 3×10^{15} Hz.

(ii) have wavelength of 0.50 Å.

2.7 Calculate the wavelength, frequency and wavenumber of a light wave whose period is 2.0×10^{-10} s.

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

2.9 A photon of wavelength 4×10^{-7} m strikes on metal surface, the work function of the metal being 2.13 eV. Calculate (i) the energy of the photon (eV), (ii) the kinetic energy of the emission and (iii) the velocity of the photoelectron (1eV = 1.6020×10^{-19} J).

2.10 Electromagnet radiation of wavelength 242 nm is just sufficient to ionise the sodium atom. Calculate the ionisation energy of sodium in kJ mol^{-1}.

2.11 A 25 watt bulb emits monochromatic yellow light of wavelength of 0.57 µm. Calculate the rate of emission of quanta per second.

2.12 Electrons are emitted with zero velocity from a metal surface when it is exposed to radiation of wavelength 6800 Å. Calculate threshold frequency (v_0) and work function (W_0) of the metal.

2.13 What is the wavelength of light emitted when the electron in a hydrogen atom undergoes transition from an energy level with n = 4 to an energy level with n = 2 ?

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

2.15 What is the maximum number of emission lines when the excited electron of a H atom in n = 6, drops to the ground state ?

2.16 (i) The energy associated with the first orbit in the hydrogen atom is -2.18×10^{-18} J $atom^{-1}$. What is the energy associated with the fifth orbit ?

(ii) Calculate the radius of Bohr's fifth orbit for hydrogen atom.

2.17 Calculate the wavelength for the longest wavelength transition in the Balmer series of atomic hydrogen.

2.18 What is the energy in joules, required to shift the electron of the hydrogen atom from the first Bohr orbit to the fifth Bohr orbit and what is the wavelength of the light emitted when the electron returns to the ground state ? The ground state electron energy is -2.18×10^{-11} ergs.

2.19 The electron energy in hydrogen atom is given by $E_n = (-2.18 \times 10^{-18}) / n^2$ joules. Calculate the energy required to remove an electron completely from the n = 2 orbit. What is the longest wavelength of light in cm that can be used to cause this transition ?

2.20 Calculate the wavelength of an electron moving with a velocity of 2.05×10^{7} m s^{-1}.

2.21 The mass of an electron is 9.1×10^{-31} kg. If its K.E. is 3.0×10^{-25} J, calculate its wavelength.

2.22 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.

2.23 (i) Write the electronic configuration of the following ions :
(a) H^{-} (b) Na^{+} (c) O^{2-} (d) F^{-}
(ii) What are the atomic numbers of elements whose outermost electrons are represented by
(a) $3s^{1}$ (b) $2p^{3}$ (C) $3p^{5}$
(iii) Which atoms are indicated by the following configurations ?
(a) [He] $2s^{1}$ (b) [Ne] $3s^{2}$ $3p^{3}$ (c) [Ar] $4s^{2}$, $3d^{1}$

2.24 What is the lowest value of n that allows 'g' orbitals to exist ?

2.25 An electron is in one of the 3d-orbitals. Give the possible values of n, l and m_l for this electron.

2.26 An atom of an element contains 29 electrons and 35 neutrons. Deduce (i) the number of protons and (ii) the electronic configuration of the element.

2.27 Give the number of electrons in the species H_2^{+}, H_2 and O_2^{+}.

2.28 (i) An atomic orbital has n = 3. What are the possible values of l and m_l ?
(ii) List the quantum numbers (m_l and l) of electrons for 3d orbital.
(iii) Which of the following orbitals are possible ?
1p, 2s, 2p and 3f

2.29 Using s, p, d notations, describe the orbital with the following quantum numbers.
(a) n = 1, l = 0; (b) n = 3, l = 1; (c) n = 4, l = 2 (d) n = 4, l = 3

2.30 Explain giving reasons, which of the following sets of quantum numbers are not possible.
(a) n = 0, l = 0, m_l = 0, m_s = + ½ (b) n = 1, l = 0, m_l = 0, m_s = – ½
(c) n = 1, l = 1, m_l = 0, m_s = + ½ (d) n = 2, l = 1, m_l = 0, m_s = – ½
(e) n = 3, l = 3, m_l = –3, m_s = + ½ (f) n = 3, l = 1, m_l = 0, m_s = + ½

2.31 How many electrons in an atom may have the following quantum numbers ?
(a) $n = 4, m_s = -\frac{1}{2}$ (b) n = 3, l = 0

2.32 Show that the circumference of the Bohr orbit for the hydrogen atom is an integral multiple of the de Broglie wavelength associated with the electron revolving around the orbit.

2.33 What transition in the hydrogen spectrum would have the same wavelength as the Balmer transition n = 4 to n = 2 of He^{+} spectrum ?

2.34 Calculate the energy required for the process :

$$He^{+}(g) \rightarrow He^{2+}(g) + e^{-}$$

The ionization energy for the H atom in the ground state is 2.18×10^{-18} J $atom^{-1}$.

2.35 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 length of scale of length 20 cm long.

2.36 2×10^{8} atoms of carbon are arranged side by side. Calculate the radius of carbon atom if the length of this arrangement is 2.4 cm.

2.37 The diameter of zinc atom is 2.6 Å. Calculate (a) radius of zinc atom in pm and (b) number of atoms present in a length of 1.6 cm if the zinc atoms are arranged side by side lengthwise.

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

2.39 In Millikan's experiment, static electric charge on the oil drop has been obtained by shining X-rays. If the static electric charge on the oil drop is -1.282×10^{-18} C, calculate the number of electrons present on it.

2.40 **In Rutherford's experiment, generally the thin foil of heavy atoms, like gold, platinum etc. have been used to be bombarded by the α-particles. If the thin foil of light atoms like aluminium etc. is used, what difference would be observed from the above results ?**

2.41 **Symbols $^{79}_{35}Br$ and ^{79}Br can be written, whereas symbols $^{35}_{79}Br$ and ^{35}Br are not acceptable. Answer briefly.**

2.42 **An element with mass number 81 contains 31.7% more neutrons as compared to protons. Assign the atomic asymbol.**

2.43 **An ion with mass number 37 possesses one unit of negative charge. If the ion contains 11.1% more neutrons than the electrons, find the symbol of the ion.**

2.44 **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.**

2.45 **Arrange the following type of radiations in increasing order of frequency :**

(a) Radiation from microwave open (b) amber light from traffic signal (c) radiation from FM radio (d) cosmic rays from outer space and (e) X-rays.

2.46 **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.**

2.47 **Neon gas is generally used in the sign boards. If it emits strongly at 616 nm, calculate (a) the frequency of the emission, (b) distance travelled by this radiation in 30 s, (c) energy of quantum and (d) number of quanta present if it produces 2 J of energy.**

2.48 **In astronomical observations, signals observed from the distant stars are generally weak. If the photon detector receives a total of 3.15×10^{-18} J from the radiations of 600 nm, calculate the number of photons received by the detector.**

2.49 **Lifetimes of the molecules in the excited states are often measured by using pulsed radiation source of duration nearly in the nano second range. If the radiation source has the duration the of 2 ns and the number of photons emitted during the pulse source is 2.5×10^{15}, calculate the energy of the source.**

2.50 **The longest wavelength doublet absorption transition is observed at 589 and 589.6 nm. Calculate the frequency of each transition and energy difference between the two excited states.**

2.51 **The work function for caesium atom is 1.9 eV. Calculate (a) the threshold wavelength and (b) the threshold frequency of radiation. (c) If the caesium element is irradiated with a wavelength of 500 nm, calculate the kinetic energy and the velocity of the ejected photoelectron.**

2.52 **Following results are observed when sodium metal is irradiated with different wavelengths. Calculate (a) threshold wavelength and, (b) Planck's constant**

λ(nm)	500	450	400
$v \times 10^{-5}$ (cm s^{-1})	2.55	4.35	5.35

2.53 **The ejection of the photoelectron from the silver metal in the photoelectric effect experiment can be stopped by applying a voltage of 0.35 V when the radiation of 256.7 nm is used. Calculate the work function for silver metal.**

2.54 **If the photon of the wavelength 150 pm strikes an atom and one of its inner bound electrons is ejected out with a velocity of 1.5×10^{7} m s^{-1}, calculate the energy with which it is bound to the nucleus.**

2.55 **Emission transitions in the Paschen series end at orbit n = 3 and start from orbit n and can be represented as $v = 3.29 \times 10^{15}\ \text{Hz}\left[\frac{1}{3^2} - \frac{1}{n^2}\right]$. Calculate the value of n if the transition is observed at 1285 nm. Find the region of the spectrum.**

2.56 **Calculate the wavelength for the emission transition if it starts from the orbit having radius 1.3225 nm and ends at 211.6 pm. Name the series to which this transition belongs and the region of the spectrum.**

2.57 **Dual behaviour of matter proposed by de Broglie led to the discovery of electron microscope often used for the highly magnified images of biological molecules and other type of material. If the velocity of the electron in this microscope is 1.6×10^6 m s^{-1}, calculate de Broglie wavelength associated with the electron.**

2.58 **Similar to electron diffraction, neutron diffraction microscope is also used for the determination of the structure of molecules. If the wavelength used here is 800 pm, calculate the characteristic velocity associated with the neutron.**

2.59 **If the velocity of the electron in Bohr's first orbit is 2.19×10^6 m s^{-1}, calculate the de Broglie wavelength associated with it.**

2.60 **The velocity associated with a proton moving in a potential difference of 1000 V is 4.37×10^5 m s^{-1}. If the hockey ball of mass 0.1 kg is moving with this velocity, calculate the wavelength associated with this velocity.**

2.61 **If the position of the electron is measured within an accuracy of ± 0.002 nm, calculate the uncertainly in the momentum of the electron. Suppose the momentum of the electron is $\frac{h}{4\pi_m \times 0.05\,nm}$, is there any problem in defining this value.**

2.62 **The quantum numbers of six electrons are given below. Arrange them in order of increasing energies. If any of these combination(s) has/have the same energy :**

(i) $n = 4, l = 2, m_l = -2, m_s = -\frac{1}{2}$ **(ii)** $n = 3, l = 2, m_l = 1, m_s = +\frac{1}{2}$

(iii) $n = 4, l = 1, m_l = 0, m_s = +\frac{1}{2}$ **(iv)** $n = 3, l = 2, m_l = -2, m_s = -\frac{1}{2}$

(v) $n = 3, l = 1, m_l = -1, m_s = +\frac{1}{2}$ **(vi)** $n = 4, l = 1, m_l = 0, m_s = +\frac{1}{2}$

2.63 **The bromine atom possesses 35 electrons. It contains 6 electrons in 2p orbital, 6 electrons in 3p orbital and 5 electron in 4p orbital. Which of these electron experiences the lowest effective nuclear charge ?**

2.64 **Among the following pairs of orbitals which orbital will experience the larger effective nuclear charge ? (a) 2s and 3s, (b) 4d and 4f (c) 3d and 3p.**

2.65 **The unpaired electrons in Al and Si are present in 3p orbital. Which electrons will experience more effective nuclear charge from the nucleus ?**

2.66 **Indicate the number of unpaired electrons in : (a) P, (b) Si, (c) Cr, (d) Fe, and (e) Kr.**

2.67 **(a)** **How many sub-shells are associated with n = 4 ?**

(b) **How many electrons will be present in the sub-shells having m_s value of $-\frac{1}{2}$ for n = 4 ?**

ANSWERS

2.1 (i) 1.099×10^{27} electrons (ii) 5.48×10^{-7} kg, 9.65×10^{4}C

2.2 (i) 6.022×10^{24} electrons (ii) (a) 2.4088×10^{21} neutrons (b) 4.0347×10^{-6} kg (iii) (a) 1.2044×10^{22} protons (b) 2.0145×10^{-5} kg. There is not effect of temperature and pressure

2.3 7,6 : 8. 8 : 12, 12 : 30, 26 : 50, 38

2.4 (i) $^{35}_{17}Cl$ (ii) $^{233}_{92}U$ (iii) $^{9}_{4}Be$

2.5 $5.17 \times 10^{14} s^{-1}$, $1.72 \times 10^{6} m^{-1}$

2.6 (i) 1.988×10^{-18} J (ii) 3.9756×10^{-15} J

2.7 $5 \times 10^{9}\ s^{-1}$, 6.0×10^{-2} m, 16.66 m^{-1}

2.8 2.012×10^{16} photon

2.9 (i) 3.10 eV (ii) 0.97 eV (iii) $5.84 \times 10^{5}\ m\ s^{-1}$

2.10 494 kJ mol^{-1}

2.11 $7.18 \times 10^{19}\ s^{-1}$

2.12 $4.41 \times 10^{14}\ s^{-1}$, 2.92×10^{-19} J

2.13 486 nm

2.14 8.72×10^{-20} J, thus energy required to remove from the first orbit of hydrogen atom is more than the energy required to remove the electron from the firth orbit

2.15 15

2.16 (i) 8.72×10^{-20} J (ii) 1.3325 nm

2.17 $1.523 \times 10^{6}\ m^{-1}$

2.18 (i) 2.08×10^{-11} ergs (ii) 956 Å

2.19 3647 Å

2.20 3.55×10^{-11} m

2.21 8967 Å

2.22 Na^{+} and Mg^{2+} are isoelectronic species ; K^{+}, Ca^{2+}, S^{2-} and Ar are isoelectronic species

2.23 (i) (a) H^{-} : $1s^2$ (b) Na+ : $1s^2, 2s^2, 2p^6$ (c) O^{2-} : $1s^2, 2s^2, 2p^6$ (d) F^{-} : $1s^2, 2s^2, 2p^6$ (ii) (a) Sodium (atomic number 11) (b) Nitrogen (atomic number 7) (c) Chromium (atomic number 26) (iii) (a) Li (b) P (c) Sc

2.24 n = 5, g-orbitals

2.25 $n = 3, l = 2, m_l = -2, -1, 0, + 1 + 2$ (any one value)

2.26 (i) Number of electrons = 29 i.e., atomic number = 29, Number of protons = 29 (ii) Electronic configuration : $1s^2, 2s^2, 2p^6, 3s^2, 3p^6\ 3d^{10}, 4s^1$ (half-filled and completely filled are more stable)

2.27 $H_2^{+} = 1$, $H_2 = 2$, $O_2^{+} = 15$

2.28 (i) l : 0, 1, 2 (ii) m_l : –2, –1, 0, +1, +2 (iii) 2s, 2p

2.29 (a) 1s (b) 3p (c) 4d (d) 4f

2.30 (a) Not possible, as n cannot have a value of 0 (b) Possible (c) Not possible, because when n = 1, $l \neq 1$ (d) Possible (e) Not possible, because when n = 3, $l \neq 3$ (f) Possible

2.31 (a) only 16 electrons (b) only 2 electrons

2.33 $n_1 = 1$ and $n_2 = 2$

2.34 8.72×10^{-18} J

2.35 1.33×10^{9}

2.36 6 nm

2.37 (a) 1.3×10^{4} pm (b) 1.23×10^{6}

2.38 1563

2.39 8

2.40 More number of K-particles will pass as the nucleus of the lighter atoms is small, smaller number of K-particles will be deflected as a number of positive charges is less than on the lighter nuclei.

2.41 For a given element the number of protons is the same for the isotopes, whereas the mass number can be different for the given atomic number.

2.42 $^{81}_{35}Br$

2.43 $^{37}_{17}Cl^{-1}$

2.44 $^{56}_{26}Fe^{3+}$

2.45 Cosmic rays > X-rays > amber colour > mocrowave > FM

2.46 3.3×10^{6} J

2.47 (a) $4.87 \times 10^{14}\ s^{-1}$ (b) 9.0×10^{9} m (c) 32.27×10^{-20} J (d) 6.2×10^{18}

2.48 10

2.49 8.828×10^{-10} J

2.50 3.46×10^{-22} J

2.51 (a) 652 nm (b) $4.598 \times 10^{14}\ s^{-1}$ (c) 3.97×10^{-13} J, $9.33 \times 10^{8}\ ms^{-1}$

2.52 $\lambda_0 = 530.8$ nm, $\cong 531$ nm

2.53 4.3 eV

2.54 7.62×10^{3} eV

2.55 n = 5

2.56 434 nm, visible region

2.57 455 pm

2.58 $4.94 \times 10^{4}\ m\ s^{-1}$

2.59 332 pm

2.60 1.51×10^{-27} m

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

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

2.63 4p

2.64 (i) 2s (ii) 4d (iii) 3p

2.65 Si

2.66 (a) 3 (b) 2 (c) 6 (d) 4 (e) zero

2.67 (a) n = 4, *l* = 0, 1, 2, 3 (b) 16

ADDITIONAL QUESTIONS AND PROBLEMS

Q. Write down some most important postulates of Bohr's model of an atom.

Q. Compare the shapes of 1s, 2s and 3s orbitals.

Q. What is a photon ? How is the energy of photon related to its

(a) frequency (b) wavelength

Q. Calculate the

(i) mass of 1 mole of electrons and

(ii) charge of 1 mole of electrons

Q. Mention the total number of electrons protons and neutrons in each of the following

(i) N^{-3} (ii) PO_4^{-3} (iii) CO_2 (iv) PO_3^{-1}

(v) NO

Q. Why d^5 and d^{10} configurations are more stable than d^4 and d^6 configuration ?

Q. Write down electronic configuration of the following

(a) H^{-1} (b) C (c) Na^+

(d) Fe^{+3} (e) Ne(Z = 10) (f) Ca(Z = 20)

(g) Cr (Z = 24) (h) Cu(Z = 29) (i) Ni^{+2} (Z = 28)

(j) N^{-3} (Z = 7)

Q. How many unpaired electrons are present in

(a) Cr^{+3} (b) Ni^{+2} (c) Cr (d) Cu

Q. What is the maximum electrons that can be filled in the M-shell ? Give reasons ?

Q. Differentiate the term orbit and orbital.

Q. Describe Rutherford's model of an atom and mention its defects. Discuss the solution proposed by Bohr.

Q. Write notes on

(a) Pauli exclusion principle (b) Hund's rule

(c) Aufbau principle (d) Heisenberg's undertaintly principle

(e) Dual nature of electron

Q. What is meant by nucleus of an atom ? Give an account of the scattering experiment which led to the discovery of the nucleus.

Q. What are the quantum numbers ? Explain the siginficance of each quantum number.

Q. What are electromagnet waves ? Explain the following properties of waves

(a) Wavelength (b) Frequency

(c) Wave number

How are these interrelated ?

Q. Derive an expression for the energy of an electron. Discuss the drawbacks of the Bohr's model.

Q. Give the essential postulates of Bohr's model of an atom. How did it explain

(a) the stability of an atom

(b) origin of the spectral lines in H-atom ?

Q. What are orbitals ? Discuss the shape and features of s, p and d-orbitals.

Q. State and explain Pauli exclusion principle.

Q. State Heisenberg's uncertainty principle and prove that it is vaid only for the microscopic particles.

Q. Develop de-Broglie equation. How it is experimentally verified ?

Q. State and explain Hund's rule of maximum multiplicity.

Q. **Calculate the wavelength of a photon associated with an energy of 1 eV.**

A. 12.41×10^{-5}

Q. **How many photons of light having wavelength of 500 nm are necessary to provide 1 Joule of energy.**

A. 2.51×10^{18} photons

Q. **A photon of 3000 Å is absorbed by a gas and then re-emitted as two photons. One photon is red (λ = 7600 Å). What would be the wavelength of the other photon ?**

A. 4956 Å

Q. **Calculate the wavelength associated with an electron ($m = 9.1 \times 10^{-31}$ kg) moving with a velocity of 10^8 m/sec.**

A. 7.25×10^{-10} m

Q. **Calculate the momentum of a particle which has a de-Broglie wavelength of 3 Å.**

A. 2.2×10^{-24} kg m/sec

Q. **Calculate the uncertainty in position of an electron if the uncertainty in its velocity is 5.7×10^{-5} m/sec.**

A. 1×10^{-10} m

Q. **A photon of wavelength 4×10^{-7} m strikes on metal surface, the work function of metal being 2.13 eV. Calculate**

(i) **the K.E. of the photon and**

(ii) **the velocity of photon electron**

A. (i) 0.97 eV (ii) 5.85×10^5 m/s

Q. **The I.P. of He^{+1} is 19.6×10^{-18} J/atom. Calculate the energy of the 1st stationary state of Li^{+2}.**

A. 4.41×10^{-17} J

Q. **If the energy difference between the ground state of an atom and its excited state is 4.4×10^{-19} J, what is the wavelength of photon required to produce this transition ?**

A. 4.5×10^{-7} m