

CHAPTER 2

STRUCTURE OF ATOM

- Atom is the smallest indivisible particle of the matter. Atom is made of electron, proton and neutrons.

PARTICLE	ELECTRON	PROTON	NEUTRON
Discovery	Sir. J. J. Thomson (1869)	Goldstein (1886)	Chadwick (1932)
Nature of charge	Negative	Positive	Neutral
Amount of charge	1.6×10^{-19} Coloumb	1.6×10^{-19} Coloumb	0
Mass	9.11×10^{-31} kg	1.672614×10^{-27} kg	1.67492×10^{-27} kg

- Electrons were discovered using cathode ray discharge tube experiment.
- Nucleus was discovered by Rutherford in 1911.
- Cathode ray discharge tube experiment: A cathode ray discharge tube made of glass is taken with two electrodes. At very low pressure and high voltage, current starts flowing through a stream of particles moving in the tube from cathode to anode. These rays were called cathode rays. When a perforated anode was taken, the cathode rays struck the other end of the glass tube at the fluorescent coating and a bright spot on the coating was developed

Results:

- Cathode rays consist of negatively charged electrons.
 - Cathode rays themselves are not visible but their behavior can be observed with help of fluorescent or phosphorescent materials.
 - In absence of electrical or magnetic field cathode rays travel in straight lines
 - In presence of electrical or magnetic field, behaviour of cathode rays is similar to that shown by electrons
 - The characteristics of the cathode rays do not depend upon the material of the electrodes and the nature of the gas present in the cathode ray tube.
- Charge to mass ratio of an electron was determined by Thomson. The charge to mass ratio of an electron as $1.758820 \times 10^{11} \text{ C kg}^{-1}$
 - Charge on an electron was determined by R A Millikan by using an oil drop experiment. The value of the charge on an electron is $-1.6 \times 10^{-19} \text{ C}$.
 - The mass on an electron was determined by combining the results of Thomson's experiment and Millikan's oil drop experiment. The mass of an electron was determined to be $9.1094 \times 10^{-31} \text{ kg}$.
 - Discovery of protons and canal rays:** Modified cathode ray tube experiment was carried out which led to the discovery of protons.
 - Characteristics of positively charged particles:
 - Charge to mass ratio of particles depends on gas from which these originate

b. The positively charged particles depend upon the nature of gas present in the cathode ray discharge tube

c. Some of the positively charged particles carry a multiple of fundamental of electrical charge.

d. Behaviour of positively charged particles in electrical or magnetic field is opposite to that observed for cathode rays

- **Neutrons** were discovered by James Chadwick by bombarding a thin sheet of beryllium by α - particles. They are electrically neutral particles having a mass slightly greater than that of the protons.
- **Atomic number (Z)** : the number of protons present in the nucleus (Moseley 1913).
- **Mass Number (A)** : Sum of the number of protons and neutrons present in the nucleus.
- **Thomson model of an atom**: This model proposed that atom is considered as a uniform positively charged sphere and electrons are embedded in it. An important feature of Thomson model of an atom was that mass of atom is considered to be evenly spread over the atom. Thomson model of atom is also called as Plum pudding, raisin pudding or watermelon model. Thomson model of atom was discarded because it could not explain certain experimental results like the scattering of α - particles by thin metal foils.
- **Observations from α - particles scattering experiment by Rutherford**:
 - a. Most of the α - particles passed through gold foil undeflected
 - b. A small fraction of α - particles got deflected through small angles
 - c. Very few α - particles did not pass through foil but suffered large deflection nearly 180°
- **Conclusions Rutherford drew from α - particles scattering experiment**:
 - a. Since most of the α -particles passed through foil undeflected, it means most of the space in atom is empty
 - b. Since some of the α -particles are deflected to certain angles, it means that there is positively mass present in atom
 - c. Since only some of the α -particles suffered large deflections, the positively charged mass must be occupying very small space
 - d. Strong deflections or even bouncing back of α -particles from metal foil were due to direct collision with positively charged mass in atom
- **Rutherford's model of atom**: This model explained that atom consists of nucleus which is concentrated in a very small volume. The nucleus comprises of protons and neutrons. The electrons revolve around the nucleus in fixed orbits. Electrons and nucleus are held together by electrostatic forces of attraction.
- **Drawbacks of Rutherford's model of atom**:
 - a. According to Rutherford's model of atom, electrons which are negatively charged particles revolve around the nucleus in fixed orbits. Thus,

- b. the electrons undergo acceleration. According to electromagnetic theory of Maxwell, a charged particle undergoing acceleration should emit electromagnetic radiation. Thus, an electron in an orbit should emit radiation. Thus, the orbit should shrink. But this does not happen.
 - c. The model does not give any information about how electrons are distributed around nucleus and what are energies of these electrons
- **Isotopes:** These are the atoms of the same element having the same atomic number but different mass number. e.g. ${}_1\text{H}^1, {}_1\text{H}^2, {}_1\text{H}^3$
 - **Isobars:** Isobars are the atoms of different elements having the same mass number but different atomic number. e.g. ${}_{18}\text{Ar}^{40}, {}_{20}\text{Ca}^{40}$
 - **Isoelectronic species:** These are those species which have the same number of electrons.
 - **Electromagnetic radiations:** The radiations which are associated with electrical and magnetic fields are called electromagnetic radiations. When an electrically charged particle moves under acceleration, alternating electrical and magnetic fields are produced and transmitted. These fields are transmitted in the form of waves. These waves are called electromagnetic waves or electromagnetic radiations.
 - **Properties of electromagnetic radiations:**
 - a. Oscillating electric and magnetic field are produced by oscillating charged particles. These fields are perpendicular to each other and both are perpendicular to the direction of propagation of the wave.
 - b. They do not need a medium to travel. That means they can even travel in vacuum.
 - **Characteristics of electromagnetic radiations:**
 - a. **Wavelength:** It may be defined as the distance between two neighbouring crests or troughs of wave as shown. It is denoted by λ .
 - b. **Frequency (ν):** It may be defined as the number of waves which pass through a particular point in one second.
 - c. **Velocity (v):** It is defined as the distance travelled by a wave in one second. In vacuum all types of electromagnetic radiations travel with the same velocity. Its value is $3 \times 10^8 \text{ m sec}^{-1}$. It is denoted by c
 - d. **Wave number:** Wave number ($\bar{\nu}$) is defined as the number of wavelengths per unit length.
 - Velocity = frequency \times wavelength $c = \nu\lambda$
 - **Planck's Quantum Theory-**
 - The radiant energy is emitted or absorbed not continuously but discontinuously in the form of small discrete packets of energy called 'quantum'. In case of light, the quantum of energy is called a 'photon'
 - The energy of each quantum is directly proportional to the frequency of the radiation, i.e. $E \propto \nu$ or $E = h\nu$ where h = Planck's constant = $6.626 \times 10^{-27} \text{ Js}$
 - Energy is always emitted or absorbed as integral multiple of this quantum. $E = nh\nu$ Where $n = 1, 2, 3, 4, \dots$

- **Black body:** An ideal body, which emits and absorbs all frequencies, is called a black body. The radiation emitted by such a body is called black body radiation.
- **Photoelectric effect:** The phenomenon of ejection of electrons from the surface of metal when light of suitable frequency strikes it is called photoelectric effect. The ejected electrons are called photoelectrons.
- Experimental results observed for the experiment of Photoelectric effect-
 - When beam of light falls on a metal surface electrons are ejected immediately.
 - Number of electrons ejected is proportional to intensity or brightness of light
 - Threshold frequency (ν_0): For each metal there is a characteristic minimum frequency below which photoelectric effect is not observed. This is called threshold frequency.
 - If frequency of light is less than the threshold frequency there is no ejection of electrons no matter how long it falls on surface or how high is its intensity.
- Photoelectric work function (W_0): The minimum energy required to eject electrons is called photoelectric work function. $W_0 = h\nu_0$
- Energy of the ejected electrons :

$$h(\nu - \nu_0) = \frac{1}{2} m_e v^2$$
- Dual behavior of electromagnetic radiation- The light possesses both particle and wave like properties, i.e., light has dual behavior. Whenever radiation interacts with matter, it displays particle like properties. (Black body radiation and photoelectric effect) Wave like properties are exhibited when it propagates (interference and diffraction)
- When a white light is passed through a prism, it splits into a series of coloured bands known as spectrum.
- Spectrum is of two types: continuous and line spectrum
 - a. The spectrum which consists of all the wavelengths is called continuous spectrum.
 - b. A spectrum in which only specific wavelengths are present is known as a line spectrum. It has bright lines with dark spaces between them.
- Electromagnetic spectrum is a continuous spectrum. It consists of a range of electromagnetic radiations arranged in the order of increasing wavelengths or decreasing frequencies. It extends from radio waves to gamma rays.
- Spectrum is also classified as emission and line spectrum.
 - Emission spectrum: The spectrum of radiation emitted by a substance that has absorbed energy is called an emission spectrum.
 - Absorption spectrum is the spectrum obtained when radiation is passed through a sample of material. The sample absorbs radiation of

certain wavelengths. The wavelengths which are absorbed are missing and come as dark lines.

- The study of emission or absorption spectra is referred as spectroscopy.
- Spectral Lines for atomic hydrogen:

Series	n_1	n_2	Spectral Region
Lyman	1	2, 3, 4, 5 ...	Ultraviolet
Balmer	2	3, 4, 5 ...	Visible
Paschen	3	4, 5 ...	Infrared
Brackett	4	5, 6 ...	Infrared
Pfund	5	6, 7 ...	Infrared

- Rydberg equation

$$\bar{\nu} = 109,677 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right) \text{ cm}^{-1}$$

$$R = \text{Rydberg's constant} = 109677 \text{ cm}^{-1}$$

- **Bohr's model for hydrogen atom:**

- An electron in the hydrogen atom can move around the nucleus in a circular path of fixed radius and energy. These paths are called orbits or energy levels. These orbits are arranged concentrically around the nucleus.
- As long as an electron remains in a particular orbit, it does not lose or gain energy and its energy remains constant.
- When transition occurs between two stationary states that differ in energy, the frequency of the radiation absorbed or emitted can be calculated

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

- An electron can move only in those orbits for which its angular momentum is an integral multiple of $h/2\pi$

$$m_e v r = n \cdot \frac{h}{2\pi} \quad n = 1, 2, 3, \dots$$

- The radius of the n th orbit is given by $r_n = 52.9 \text{ pm} \times \frac{n^2}{Z}$
- energy of electron in n th orbit is :

$$E_n = -2.18 \times 10^{-18} \left(\frac{Z^2}{n^2} \right) \text{ J}$$

- **Limitations of Bohr's model of atom:**

- Bohr's model failed to account for the finer details of the hydrogen spectrum.
- Bohr's model was also unable to explain spectrum of atoms containing more than one electron.

- **Dual behavior of matter:** de Broglie proposed that matter exhibits dual behavior i.e. matter shows both particle and wave nature. de Broglie's relation is

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

- **Heisenberg's uncertainty principle:** It states that it is impossible to determine simultaneously, the exact position and exact momentum (or velocity) of an electron. The product of their uncertainties is always equal to or greater than $h/4\pi$.

Mathematically $\Delta x \times \Delta p \geq \frac{h}{4\pi}$
 where Δx = uncertainty in position,
 Δp = uncertainty in momentum

- Heisenberg's uncertainty principle rules out the existence of definite paths or trajectories of electrons and other similar particles
- Failure of Bohr's model:
 - a. It ignores the dual behavior of matter.
 - b. It contradicts Heisenberg's uncertainty principle.
- Classical mechanics is based on Newton's laws of motion. It successfully describes the motion of macroscopic particles but fails in the case of microscopic particles.

Reason: Classical mechanics ignores the concept of dual behaviour of matter especially for sub-atomic particles and the Heisenberg's uncertainty principle.

- **Quantum mechanics** is a theoretical science that deals with the study of the motions of the microscopic objects that have both observable wave like and particle like properties.
- Quantum mechanics is based on a fundamental equation which is called Schrodinger equation.
- Schrodinger's equation: For a system (such as an atom or a molecule whose energy does not change with time) the Schrödinger equation is written as:

$$\hat{H}\Psi = E\Psi$$

\hat{H} is the Hamiltonian operator

E is the total energy of the system

Ψ represents the wave function which is the amplitude of the electron

Wave

- When Schrödinger equation is solved for hydrogen atom, the solution gives the possible energy levels the electron can occupy and the corresponding wave function(s) of the electron associated with each energy level. Out of the possible values, only certain solutions are permitted. Each permitted solution is highly significant as it corresponds to a definite energy state. Thus, we can say that energy is quantized.
- ψ gives us the amplitude of wave. The value of ψ has no physical significance.
- ψ^2 gives us the region in which the probability of finding an electron is maximum. It is called probability density.

- Orbital: The region of space around the nucleus where the probability of finding an electron is maximum is called an orbital.
- Quantum numbers: There are a set of four quantum numbers which specify the energy, size, shape and orientation of an orbital. To specify an orbital only three quantum numbers are required while to specify an electron all four quantum numbers are required.
- Principal quantum number (n):** It identifies shell, determines sizes and

N	1	2	3	4
Shell no.:	K	L	M	N
Total number of orbitals in a shell = n^2	1	4	9	16
Maximum number of electrons = $2n^2$	2	8	18	32

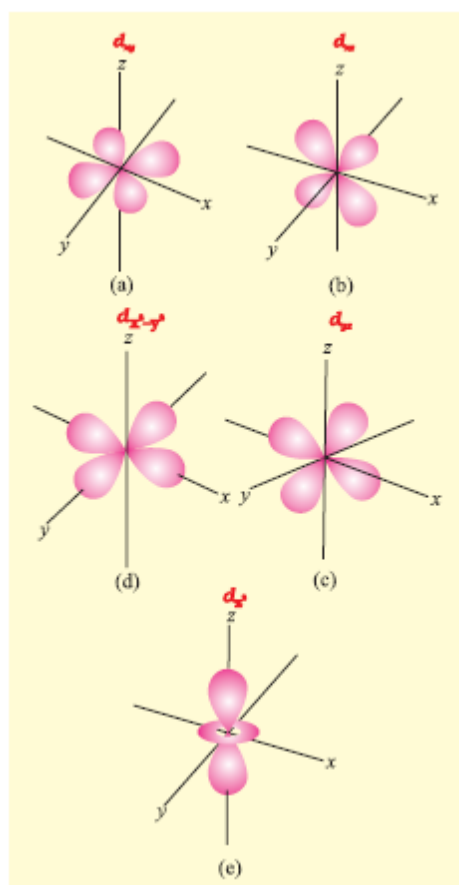
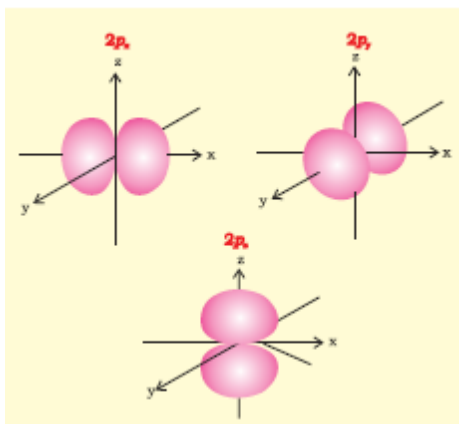
energy of orbitals

- Azimuthal quantum number (l):** Azimuthal quantum number. 'l' is also known as orbital angular momentum or subsidiary quantum number. l. It identifies sub-shell, determines the shape of orbitals, energy of orbitals in multi-electron atoms along with principal quantum number and orbital angular momentum, i.e., $\sqrt{l(l+1)} \frac{h}{2\pi}$. The number of orbitals in a subshell = $2l + 1$. For a given value of n, it can have n values ranging from 0 to n-1. Total number of subshells in a particular shell is equal to the value of n.

Subshell notation	s	p	d	f	g
Value of 'l'	0	1	2	3	4
Number of orbitals	1	3	5	7	9

- Magnetic quantum number or Magnetic orbital quantum number (ml):** It gives information about the spatial orientation of the orbital with respect to standard set of co-ordinate axis. For any sub-shell (defined by 'l' value) $2l+1$ values of ml are possible. For each value of l, $m_l = -l, -(l-1), -(l-2), \dots, 0, 1, \dots, (l-2), (l-1), l$
- Electron spin quantum number (ms):** It refers to orientation of the spin of the electron. It can have two values $+1/2$ and $-1/2$. $+1/2$ identifies the clockwise spin and $-1/2$ identifies the anti-clockwise spin.
- The region where this probability density function reduces to zero is called nodal surfaces or simply nodes.
- Radial nodes:** Radial nodes occur when the probability density of wave function for the electron is zero on a spherical surface of a particular radius. Number of radial nodes = $n - l - 1$
- Angular nodes:** Angular nodes occur when the probability density wave function for the electron is zero along the directions specified by a particular angle. Number of angular nodes = l
- Total number of nodes = $n - 1$

- Degenerate orbitals: Orbitals having the same energy are called degenerate orbitals.
- Shape of p and d-orbitals



- **Shielding effect or screening effect:** Due to the presence of electrons in the inner shells, the electron in the outer shell will not experience the full positive charge on the nucleus. So, due to the screening effect, the net positive charge experienced by the electron from the nucleus is lowered and is known as effective nuclear charge. Effective nuclear charge experienced by the orbital decreases with increase of azimuthal quantum number (l).
- **Aufbau Principle:** In the ground state of the atoms, the orbitals are filled in order of their increasing energies

- **n+l rule**-Orbitals with lower value of (n+l) have lower energy. If two orbitals have the same value of (n+l) then orbital with lower value of n will have lower energy.
- The order in which the orbitals are filled is as follows:
1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 4f, 5d, 6p, 7s...
- **Pauli Exclusion Principle**: No two electrons in an atom can have the same set of four quantum numbers. Only two electrons may exist in the same orbital and these electrons must have opposite spin.
- **Hund's rule of maximum multiplicity**: Pairing of electrons in the orbitals belonging to the same subshell (p, d or f) does not take place until each orbital belonging to that subshell has got one electron each i.e., it is singly occupied.
- **Electronic configuration of atoms**: Arrangement of electrons in different orbitals of an atom. The electronic configuration of different atoms can be represented in two ways.
 - a. $s^a p^b d^c$ notation.
 - b. Orbital diagram: each orbital of the subshell is represented by a box and the electron is represented by an arrow (\uparrow) a positive spin or an arrow (\downarrow) a negative spin.
- **Stability of completely filled and half filled subshells**:
 - a. Symmetrical distribution of electrons- the completely filled or half filled sub-shells have symmetrical distribution of electrons in them and are more stable.
 - b. Exchange energy- The two or more electrons with the same spin present in the degenerate orbitals of a sub-shell can exchange their position and the energy released due to this exchange is called exchange energy. The number of exchanges 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.

ONE MARK QUESTIONS

1. Neutrons can be found in all atomic nuclei except in one case. Which is this atomic nucleus and what does it consist of?

Ans. Hydrogen atom. It consists of only one proton.

2. Calculate wave number of yellow radiations having wavelength of 5800 \AA .

Ans. Wave number = $1/\text{wavelength}$

$$\text{Wavelength} = 5800 \text{ \AA} = 5800 \times 10^{-10} \text{ m}$$

$$\text{Wave number} = 1/5800 \times 10^{-10} \text{ m} = 1.72 \times 10^6 \text{ m}^{-1}$$

3. What are the values of n and l for 2p orbital?

Ans. $n=2$ and $l=1$

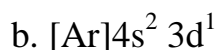
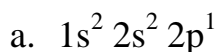
4. Which of the following orbitals are not possible? 1p, 2s, 3f and 4d

Ans. 1p and 3f are not possible.

5. Write the electronic configuration of the element having atomic number 24.

Ans. $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$

6. What atoms are indicated by the following electronic configurations?



Ans. a. Boron b. Scandium

7. What is the relationship between frequency and wavelength of light?

Ans. velocity of light = frequency x wavelength. Frequency and wavelength are inversely proportional to each other.

8. State Pauli Exclusion Principle.

Ans. No two electrons in an atom can have the same set of four quantum numbers or an orbital can have maximum two electrons and these must have opposite spin.

9. When α - rays hit a thin foil of gold, very few α - particles is deflected back.

What does it prove?

Ans. There is a very small heavy body present within the atom.

10. What is the difference between a quantum and a photon?

Ans. The smallest packet of energy of any radiation is called a quantum whereas that of light is called photon.

TWO MARKS QUESTIONS

1. Write the complete symbol for the atom with the given atomic number (Z) and mass number(A). (a) $Z = 17, A = 35$ (b) $Z = 92, A = 233$

Ans. (a) ${}^{35}_{17}\text{Cl}$ (b) ${}^{233}_{92}\text{U}$

2. Using s,p,d and f notation, describe the orbital with the following quantum numbers-

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

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

3. How many electrons in an atom have the following quantum numbers?

a. $n=4, m_s = -1/2$ b. $n=3, l=0$

Ans. (a) 16 electrons (b) 2 electrons.

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

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

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

$x + 1.317x = 81$ or $2.317x = 81$

$x = 35$

Thus proton = 35, i.e., atomic no. = 35

Hence symbol is ${}^{81}_{35}\text{Br}$

5. (i) The energy associated with the first orbit in the hydrogen atom is $-2.18 \times 10^{-18} \text{ J/atom}$. What is the energy associated with the fifth orbit

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

Ans. (i) $E_n = -2.18 \times 10^{-18} / n^2$ $E_5 = -2.18 \times 10^{-18} / 5^2 = -8.72 \times 10^{-20} \text{ J}$

(ii) For H atom, $r_n = 0.529 \times n^2$ $r_5 = 0.529 \times 5^2 = 13.225 \text{ \AA} = 1.3225 \text{ nm}$

6. Explain, giving reasons, which of the following sets of quantum numbers are not possible.

(a) $n=0, l=0; m_l = 0, m_s = +1/2$ (c) $n=1, l=0; m_l = 0, m_s = -1/2$

(b) $n=1, l=1; m_l = -0, m_s = +1/2$ (d) $n=2, l=1; m_l = 0, m_s = +1/2$

Ans. (a) Not possible because $n \neq 0$ (c) Not possible because when $n=1, l \neq 1$

(b) Possible (d) Possible

7. (a) What is the lowest value of n that allows g orbitals to exist?

(b) An electron is in one of the $3d$ orbitals, Give the possible values of n, l and m_l for this electron.

Ans. (a) minimum value of $n = 5$

(b) $n=3, l=2, m_l = -2, -1, 0, +1, +2$

8. Calculate the total number of angular nodes and radial nodes present in 30 orbitals.

Ans. For $3p$ orbitals, $n=3, l=1$

Number of angular nodes $= l = 1$

Number of radial nodes $= n - l - 1 = 3 - 1 - 1 = 1$

9. Mention the drawbacks of Rutherford's atomic model.

Ans. 1. It could not explain the stability of an atom.

2. It could not explain the line spectrum of H -atom.

10. State de-Broglie concept of dual nature of matter. How is the dual nature of electron verified?

Ans. Just as light has dual nature, every material particle in motion has dual nature (particle nature and wave nature). The wave nature has been verified by Davisson and Germer's experiment whereas particle nature by scintillation experiment.

THREE MARKS QUESTIONS

1. State (a) Hund's Rule of maximum Multiplicity (b) Aufbau Principle (c) $n+l$ rule

Ans. (a) Pairing of electrons in the orbitals belonging to the same subshell (p, d or f) does not take place until each orbital belonging to that subshell has got one electron each i.e., it is singly occupied.

(b) In the ground state of the atoms, the orbitals are filled in order of their increasing energies

(c) Orbitals with lower value of $(n+l)$ have lower energy. If two orbitals have the same value of $(n+l)$ then orbital with lower value of n will have lower energy.

2. Write down the quantum numbers n and l for the following orbitals

a. $2p$ b. $3d$ c. $5f$

Ans. a. $n=2, l=1$ b. $n=3, l=2$ c. $n=5, l=3$

3. Write the 3 points of difference between orbit and orbital.

Ans.

Orbit	orbital
1. An orbit is a well defined circular path around the nucleus in which the electrons revolve	1. An orbital is the three dimensional space around the nucleus within which the probability of finding an electron is maximum (upto 90 %)
2. It represents the planar motion of an electron around the nucleus	2. It represents the three dimensional motion of an electron around the nucleus
3. All orbits are circular and disc like	3. Different orbitals have different shapes, i.e., s -orbitals are spherically symmetrical, p -orbitals are dumb-bell shaped and so on.

4. State Heisenberg's uncertainty principle. calculate the uncertainty in the position of an electron if the uncertainty in its velocity is 5.7×10^5 m/s.

Ans. It states that it is impossible to determine simultaneously, the exact position and exact momentum (or velocity) of an electron. The product of their uncertainties is always equal to or greater than $h/4\pi$.

$$\Delta x \times \Delta v = h/4\pi$$

$$\Delta x = \frac{h}{4\pi \times \Delta v} = \frac{6.6 \times 10^{-34}}{4 \times 3.14 \times 9.1 \times 10^{-31} \times 5.7 \times 10^5} = 1.0 \times 10^{-10} \text{ m}$$

5. Write 3 points of differences between electromagnetic waves and matter waves.

Electromagnetic waves	Matter waves
1. These are associated with electric and magnetic fields	1. These are not associated with electric and magnetic field.
2. They do not require any medium for propagation.	2. They require medium for propagation
3. They travel with the same speed as that of light	3. They travel with lower speeds not constant for all matter waves

6. (i) Calculate the number of electrons which will together weigh one gram.
(ii) Calculate the mass and charge of one mole of electrons

Ans. (i) Mass of one electron = 9.10939×10^{-31} kg

$$\therefore \text{Number of electrons that weigh } 9.10939 \times 10^{-31} \text{ kg} = 1$$

$$\text{Number of electrons that will weigh } 1 \text{ g} = (1 \times 10^{-3} \text{ kg})$$

$$= \frac{1}{9.10939 \times 10^{-31} \text{ kg}} \times (1 \times 10^{-3} \text{ kg})$$

$$= 0.1098 \times 10^{-3+31}$$

$$= 0.1098 \times 10^{28}$$

$$= 1.098 \times 10^{27}$$

$$\text{(ii) Mass of one electron} = 9.10939 \times 10^{-31} \text{ kg}$$

$$\text{Mass of one mole of electron} = (6.022 \times 10^{23}) \times (9.10939 \times 10^{-31} \text{ kg})$$

$$= 5.48 \times 10^{-7} \text{ kg}$$

Charge on one electron = 1.6022×10^{-19} coulomb

Charge on one mole of electron = $(1.6022 \times 10^{-19} \text{C}) (6.022 \times 10^{23})$
 $= 9.65 \times 10^4 \text{ C}$

7. Find energy of each of the photons which
 (i) correspond to light of frequency $3 \times 10^{15} \text{ Hz}$.
 (ii) have wavelength of 0.50 \AA .

Ans.(i) Energy (E) of a photon is given by the expression,

$$E = h\nu$$

Where,

h = Planck's constant = $6.626 \times 10^{-34} \text{ Js}$ ν = frequency of light = $3 \times 10^{15} \text{ Hz}$

Substituting the values in the given expression of E :

$$E = (6.626 \times 10^{-34}) (3 \times 10^{15}) \quad E = 1.988 \times 10^{-18} \text{ J}$$

(ii) Energy (E) of a photon having wavelength (λ) is given by the expression,

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

h = Planck's constant = $6.626 \times 10^{-34} \text{ Js}$

c = velocity of light in vacuum = $3 \times 10^8 \text{ m/s}$

Substituting the values in the given expression of E :

$$E = \frac{(6.626 \times 10^{-34})(3 \times 10^8)}{0.50 \times 10^{-10}} = 3.976 \times 10^{-15} \text{ J}$$

$$\therefore E = 3.98 \times 10^{-15} \text{ J}$$

8. 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 $n = 2$?

Ans. Then $n_i = 4$ to $n_f = 2$ transition will give rise to a spectral line of the Balmer series. The energy involved in the transition is given by the relation,

$$E = 2.18 \times 10^{-18} \left[\frac{1}{n_i^2} - \frac{1}{n_f^2} \right]$$

Substituting the values in the given expression of E :

$$\begin{aligned} E &= 2.18 \times 10^{-18} \left[\frac{1}{4^2} - \frac{1}{2^2} \right] \\ &= 2.18 \times 10^{-18} \left[\frac{1-4}{16} \right] \\ &= 2.18 \times 10^{-18} \times \left(-\frac{3}{16} \right) \end{aligned}$$

$$E = -(4.0875 \times 10^{-19} \text{ J})$$

The negative sign indicates the energy of emission.

$$\text{Wavelength of light emitted } (\lambda) = \frac{hc}{E}$$

$$\left(\text{since } E = \frac{hc}{\lambda} \right)$$

Substituting the values in the given expression of λ :

$$\begin{aligned} \lambda &= \frac{(6.626 \times 10^{-34})(3 \times 10^8)}{4.0875 \times 10^{-19}} \\ \lambda &= 4.8631 \times 10^{-7} \text{ m} \\ &= 486.3 \times 10^{-9} \text{ m} \\ &= 486 \text{ nm} \end{aligned}$$

9. 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 (iii) Identify the element .

Ans.(i) For an atom to be neutral, the number of protons is equal to the number of electrons.

\therefore Number of protons in the atom of the given element = 29

(ii) The electronic configuration of the atom is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$

(iii) Copper

10. Give the number of electrons in the species H_2^+ , H_2 and O_2^+

Ans. Number of electrons present in hydrogen molecule (H_2) = $1 + 1 = 2$

\therefore Number of electrons in $\text{H}_2^+ = 2 - 1 = 1$

Number of electrons in $\text{H}_2 = 1 + 1 = 2$

Number of electrons present in oxygen molecule (O_2) = $8 + 8 = 16$

\therefore Number of electrons in $\text{O}_2^+ = 16 - 1 = 15$

FIVE MARKS QUESTIONS WITH ANSWERS

1. What are the draw backs of Bohr's atomic model? 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.

Ans. 1. Bohr's model failed to account for the finer details of the hydrogen spectrum.

2. Bohr's model was also unable to explain spectrum of atoms containing more than one electron.
3. Bohr's model was unable to explain Zeeman effect and Stark effect i
4. Bohr's model could not explain the ability of atoms to form molecules by chemical bonds

Since a hydrogen atom has only one electron, according to Bohr's postulate, the angular momentum of that electron is given by:

$$mvr = n \frac{h}{2\pi} \dots\dots\dots (1) \quad \text{Where, } n = 1, 2, 3, \dots$$

According to de Broglie's equation:

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

$$\text{or } mv = \frac{h}{\lambda} \dots\dots\dots (2)$$

Substituting the value of 'mv' from expression (2) in expression (1):

$$\frac{hr}{\lambda} = n \frac{h}{2\pi}$$

$$\text{or } 2\pi r = n\lambda \dots\dots\dots (3)$$

Since '2 πr ' represents the circumference of the Bohr orbit (r), it is proved by equation (3) that the circumference of the Bohr orbit of the hydrogen atom is an integral multiple of de Broglie's wavelength associated with the electron revolving around the orbit.

2. State photo electric effect. 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.

Ans. Photoelectric effect: The phenomenon of ejection of electrons from the surface of metal when light of suitable frequency strikes it is called photoelectric effect. The ejected electrons are called photoelectrons. It is given that the work function (W_0) for caesium atom is 1.9 eV.

$$W_0 = \frac{hc}{\lambda_0}$$

(a) From the expression, $W_0 = \frac{hc}{\lambda_0}$, we get:

$$\lambda_0 = \frac{hc}{W_0}$$

Where,

λ_0 = threshold wavelength

h = Planck's constant

c = velocity of radiation

Substituting the values in the given expression of (λ_0):

$$\lambda_0 = \frac{(6.626 \times 10^{-34} \text{ Js})(3.0 \times 10^8 \text{ ms}^{-1})}{1.9 \times 1.602 \times 10^{-19} \text{ J}}$$

$$\lambda_0 = 6.53 \times 10^{-7} \text{ m}$$

Hence, the threshold wavelength λ_0 is 653 nm.

(b) From the expression, $W_0 = h\nu_0$, we get:

$$\nu_0 = \frac{W_0}{h}$$

Where,

ν_0 = threshold frequency

h = Planck's constant

Substituting the values in the given expression of ν_0 :

$$\nu_0 = \frac{1.9 \times 1.602 \times 10^{-19} \text{ J}}{6.626 \times 10^{-34} \text{ Js}}$$

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

$$\nu_0 = 4.593 \times 10^{14} \text{ s}^{-1}$$

Hence, the threshold frequency of radiation (ν_0) is $4.593 \times 10^{14} \text{ s}^{-1}$.

(c) According to the question:

Wavelength used in irradiation (λ) = 500 nm

Kinetic energy = $h(\nu - \nu_0)$

$$= hc \left(\frac{1}{\lambda} - \frac{1}{\lambda_0} \right)$$

$$= (6.626 \times 10^{-34} \text{ Js})(3.0 \times 10^8 \text{ ms}^{-1}) \left(\frac{\lambda_0 - \lambda}{\lambda \lambda_0} \right)$$

$$= (1.9878 \times 10^{-26} \text{ Jm}) \left[\frac{(653 - 500)10^{-9} \text{ m}}{(653)(500)10^{-18} \text{ m}^2} \right]$$

$$= \frac{(1.9878 \times 10^{-26})(153 \times 10^9)}{(653)(500)} \text{ J}$$

$$= 9.3149 \times 10^{-20} \text{ J}$$

Kinetic energy of the ejected photoelectron = $9.3149 \times 10^{-20} \text{ J}$

Since K.E

$$= \frac{1}{2}mv^2 = 9.3149 \times 10^{-20} \text{ J}$$

$$v = \sqrt{\frac{2(9.3149 \times 10^{-20} \text{ J})}{9.10939 \times 10^{-31} \text{ kg}}}$$

$$= \sqrt{2.0451 \times 10^{11} \text{ m}^2 \text{ s}^{-2}}$$

$$v = 4.52 \times 10^5 \text{ ms}^{-1}$$

Hence, the velocity of the ejected photoelectron (v) is $4.52 \times 10^5 \text{ ms}^{-1}$.

3. (a) 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 lists:

1. $n = 4, l = 2, m_l = -2, m_s = -1/2$
2. $n = 3, l = 2, m_l = 1, m_s = +1/2$
3. $n = 4, l = 1, m_l = 0, m_s = +1/2$
4. $n = 3, l = 2, m_l = -2, m_s = -1/2$
5. $n = 3, l = 1, m_l = -1, m_s = +1/2$
6. $n = 4, l = 1, m_l = 0, m_s = +1/2$

(b) Among the following pairs of orbitals which orbital will experience the larger effective nuclear charge? (i) $2s$ and $3s$, (ii) $4d$ and $4f$, (iii) $3d$ and $3p$

Ans. (a) For $n = 4$ and $l = 2$, the orbital occupied is $4d$.

For $n = 3$ and $l = 2$, the orbital occupied is $3d$.

For $n = 4$ and $l = 1$, the orbital occupied is $4p$.

Hence, the six electrons i.e., 1, 2, 3, 4, 5, and 6 are present in the $4d, 3d, 4p, 3d, 3p$, and $4p$ orbitals respectively.

Therefore, the increasing order of energies is $5(3p) < 2(3d) = 4(3d) < 3(4p) = 6(4p) < 1(4d)$.

(b) Nuclear charge is defined as the net positive charge experienced by an electron in the orbital of a multi-electron atom. The closer the orbital, the greater is the nuclear charge experienced by the electron (s) in it.

(i) The electron(s) present in the $2s$ orbital will experience greater nuclear charge (being closer to the nucleus) than the electron(s) in the $3s$ orbital.

(ii) $4d$ will experience greater nuclear charge than $4f$ since $4d$ is closer to the nucleus.

(iii) $3p$ will experience greater nuclear charge since it is closer to the nucleus than $3f$.

4. (i) The unpaired electrons in Al and Si are present in $3p$ orbital. Which electrons will experience more effective nuclear charge from the nucleus?
 (ii) Indicate the number of unpaired electrons in: (a) P, (b) Si, (c) Cr, (d) Fe

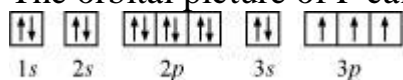
Ans. (i) the electrons in the $3p$ orbital of silicon will experience a more effective nuclear charge than aluminium.

(ii) (a) Phosphorus (P):

Atomic number = 15

The electronic configuration of P is: $1s^2 2s^2 2p^6 3s^2 3p^3$

The orbital picture of P can be represented as:



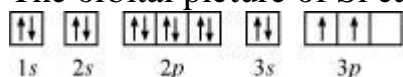
From the orbital picture, phosphorus has **three** unpaired electrons.

(b) Silicon (Si):

Atomic number = 14

The electronic configuration of Si is: $1s^2 2s^2 2p^6 3s^2 3p^2$

The orbital picture of Si can be represented as:



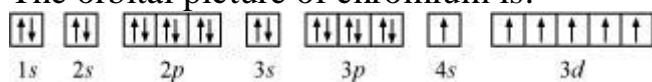
From the orbital picture, silicon has **two** unpaired electrons.

(c) Chromium (Cr):

Atomic number = 24

The electronic configuration of Cr is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$

The orbital picture of chromium is:



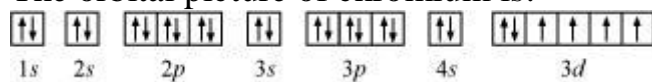
From the orbital picture, chromium has **six** unpaired electrons.

(d) Iron (Fe):

Atomic number = 26

The electronic configuration is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$

The orbital picture of chromium is:



From the orbital picture, iron has **four** unpaired electrons.

HOTS QUESTIONS WITH ANSWERS

1. Give the name and atomic number of the inert gas atom in which the total number of d-electrons is equal to the difference between the numbers of total p and total s electrons.

Ans. electronic configuration of Kr (atomic no.=36) = $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$

Total no. of s-electrons = 8, total no. of p-electrons = 18. Difference = 10

No. of d- electrons = 10

2. What is the minimum product of uncertainty in position and momentum of an electron?

Ans. $h/4\pi$

3. Which orbital is non-directional ?

Ans. s- orbital

4. What is the difference between the notations l and L ?

Ans. l represents the sub-shell and L represent shell.

5. How many electrons in an atom can have $n + l = 6$?

Ans. 18

6. An anion A^{3+} has 18 electrons. Write the atomic number of A.

Ans. 15

7. Arrange the electron (e), protons (p) and alpha particle (α) in the increasing order for the values of e/m (charge/mass).

Ans. $\alpha < p < e$