ATOMIC STRUCTURE

Chapter 2

Teaching Periods	Assessment	Weightage
10	1	8

Students will be able to:

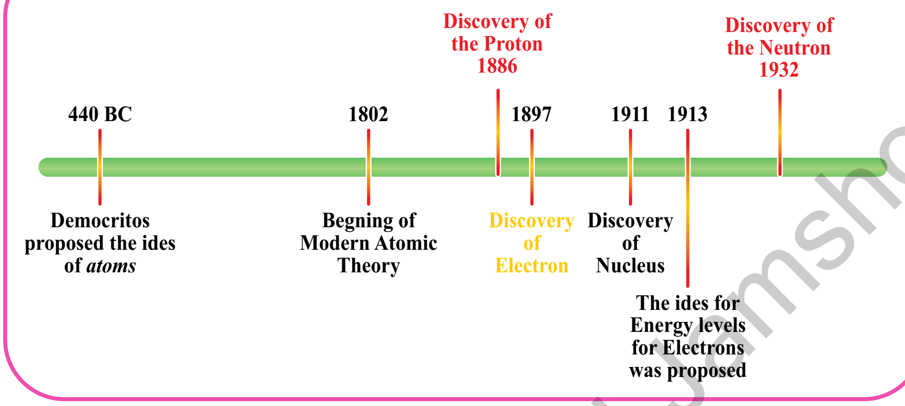
- **Describe** properties of sub atomic particles.
- **Summarize** Bohr's atomic theory.
- **Understand** to use Bohr's atomic model for calculating radii of orbit energy, frequency and wave Number of radiations emitted or absorbed by an electron.
- **Describe** spectrum and relate discrete line spectrum of hydrogen to energy levels of electrons in the hydrogen atom.
- **Explain** production, properties, types and uses of X-rays.
- **Explain** the uses of nuclear radiation in health, agricultural etc.
- **Define** photon as a unit of radiation energy.
- **Describe** the concept of orbitals.
- **Explain** the significance of quantized energies of electrons.
- **Distinguish** among principal energy levels, energy sub levels, and atomic orbitals.
- **Describe** the general shapes of s, p, and d orbitals.
- **Describe** the hydrogen atom using the Quantum Theory.
- **Understand** to use the Aufbau Principle, the Pauli Exclusion Principle, and Hund's Rule to write the electronic configuration of the elements.
- **Describe** the orbits of hydrogen atom in order of increasing energy.
- **Explain** the sequence of filling of electrons in atoms.
- **Describe** radioactivity and uses of nuclear radiation in daily life.

INTRODUCTION

In history, first time two Greek philosophers Leucippus and Democritus (440 B.C) gave the idea that matter is composed of invisible, indivisible and infinite number of building blocks called atomos (Greek a ; non, tom ; break). This ancient theory of atom was based on philosophical reasoning rather than scientific basis. They said that different atoms and their different combinations produce different types of matters (solids, liquids and gases). Most philosophers of that time did not give weightage to their concepts about matter and composition but favoured the Aristotelian concept that "Matter is composed of four elements- earth, water, air and fire".

The idea of atom was revisited and studied upon many scientist and philosophers, however it was John Dalton (1803) who recognized as the introducer of modern atomic theory.

Timeline for the discoveries and development regarding subatomic particles and atomic structure



After 12 years of discovery of electron as a fundamental entity that dwell in an atom, Robert Andrews Millikan, in 1909 performed an ingenious oil drop experiment with charged oil drop getting suspended in electric field having its weight balanced by the electrostatic force, he found out that the charge on every oil drop was an integral multiple of 1.6 x 10^{-19} Coulomb which encouraged him to deduce that this must be the charge of Thomson's electrons.

In 1911, Ernest Rutherford laid the foundation that the atom consists of an extremely charged region which he called nucleus wherein the most of the mass of an atom is concentrated while electrons revolve in the extra-nuclear part.

By 1920, scientists knew and believed that most of the mass of atom was located in the central nuclear core and in 1932, English chemist James Chadwick informed the world about the third subatomic particle that was neutral and a part of nucleus which he named neutron.

2.1 SUBATOMIC PARTICLES AND THEIR CHARACTERISTICS

More than hundred subatomic particles have been discovered in an atom such as electron, proton, neutron, positron, mesons, hyperons, neutrino antineutrino etc, however, only electron, proton and neutron are considered to be the fundamental particles of an atom which play an important role for the determination of physical and chemical properties of elements. Almost all of the mass of an atom exists in nucleus and nucleus was discovered by Rutherford in (1911 A.D). Except protium (isotope of hydrogen), nuclei of all other atoms contain neutrons.

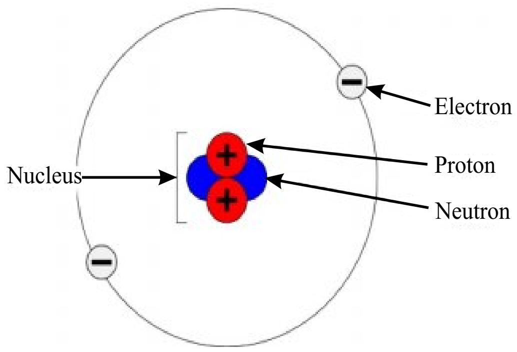


Fig 2.1: Electrons, protons and neutrons of Helium atom

Electrons:

The subatomic particle electron was discovered by J.J. Thomson in 1897 A.D, while studying on cathode rays experiment. He called the cathode rays beam as electrons Thereafter, Millikan with his "Oil Drop Experiment" calculated the charge on an electron.

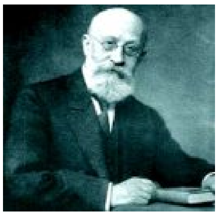
Electrons revolve in orbits and occupy about 100,000 times greater volume than nucleus but form less than one percent total mass of an atom. They carry negative charge whose magnitude is equal to positive charge of protons. Electrons are attracted by protons and this attractive force (electrostatic force) keeps electrons constantly moving around nucleus. The mass of electron is nearly 1836 times less than proton and 1839 times less than neutron.

J.J Thomson (1897 A.D)

Electronic configuration refers to the presence of electrons in various shells and sub-shells. This arrangement of electrons affects the atomic stability, melting points, boiling points, density etc. They are also involved in chemical bond formation.

Protons:

Protons are positively charged subatomic particles which are located in the nuclei of all elements. They were discovered by Goldstein (1886 A.D), a German Physicist, by using perforated cathode in Crook's discharge tube. He observed that positive ions were formed when cathode rays hit the gaseous atoms in discharge tube. Rutherford called these positively charged particles as 'PROTONS'. With respect to mass and charge, the protons of all elements are identical to each other. Atomic number of elements is related only to protons while mass number is the sum of protons and neutrons (About 99.94% mass of an atom exists in nucleus). Protons and electrons have same magnitude of charge but have different masses. All elements have different number of protons.



Goldstein (1886 A.D)

Neutrons:

The discovery of neutrons was made by James Chadwick (1932 A.D), in an artificial radioactivity experiment. He bombarded alpha particles from Polonium on Beryllium, a stable element and noted that some highly penetrating neutral particles were produced. These particles were named as neutrons.

$${}_4Be^9 + {}_2He^4 \rightarrow {}_6C^{12} + {}_0n^1 \text{ [neutron]}$$

Like protons, neutrons are also the part of atomic mass but neutron has no charge. Neutrons of all elements have same nature and mass. They are bounded with each other and protons by the nuclear forces. Under certain circumstances, when electrostatic force (between protons) overcomes the nuclear force, nucleus becomes unstable and fission reaction is occurred.

James Chadwick (1932 A.D)

Neutrons are slightly heavier than protons. They are not deflected by electric or magnetic fields. However, the isotopes of same element always have the same number of protons but different number of neutrons. Further, the stability of nucleus depends upon the neutrons. The number of neutrons is determined by subtracting the number of protons from the mass number.

Number of neutrons = Mass number – Number of protons

Table 2.1 Properties of subatomic particles

Particles	Mass (a.m.u)	Mass (g)	Relative charge	Electric charge (coulomb)
Electron (e)	0.00055	9.11×10^{-28}	-1	-1.602×10^{-19}
Proton (p)	1.0073	1.673×10^{-24}	+1	$+1.602 \times 10^{-19}$
Neutron (n)	1.0086	1.674×10^{-24}	0	0

Self Assessment

(i) How many electrons protons and neutrons are present in $^{65}_{30}Zn$?

(ii) How many time the mass of electron is lesser than the mass of proton and neutron?

2.2 INTRODUCTION AND APPLICATION OF BOHR'S ATOMIC THEORY


Neil Bohr (1913 A.D)

Rutherford became successful in presenting that atoms have tiny positive nucleus and negatively charged electrons revolve around it. He compared the movement of electrons around nucleus like planets around sun. His comparison of atomic model with solar system was defective because electrons are charged bodies while planets are charge less. If electrons revolve like planets they should lose energy continuously. As a result, electrons become closer and closer to nucleus and finally fall into it. Further, if electron continuously emits energy it should form continuous spectrum. Actually, both phenomenons did not happen. Neither electrons fall into nucleus nor continuous spectrum is formed. Contrary, line spectrum is formed. These complications were explained by Danish Physicist Neil Bohr (1913 A.D). His atomic Model was based on Planck's quantum theory.

Postulates of Bohr's theory are given below.

(i) Electrons revolve around the nucleus in circular orbits situated at fixed distance from nucleus with definite energy.

(ii) As long as electron remains in an appropriate orbit, it neither loses, nor gains energy. Hence each orbit has a fixed energy level, however the energy of orbits increases with the increased distance from nuclei.

(iii) During excitation, electron absorbs some quantized energy and jump from a lower energy orbit to an appropriate higher energy orbit but when it returns back to lower energy orbit it emits quantized energy.

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

Here 'υ' is frequency of radiation and **"h"** is plank constant.

(iv) Electron can move only in those orbits in which the angular momentum of electron (mvr) is integral multiple of nh/2π.

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

Here '**m**' and '**v**' are the mass and velocity of electron and '**r**' is the radius of orbit.

2.2.1 Calculation of Radius, Energy, Frequency and Wave Number

The truth of the basic assumption of Bohr's theory was established when applied to the single electron system like Hydrogen and ions like He^{+1}, Li^{+2} etc.

Hydrogen atom consists of a single electron revolving around the nucleus.

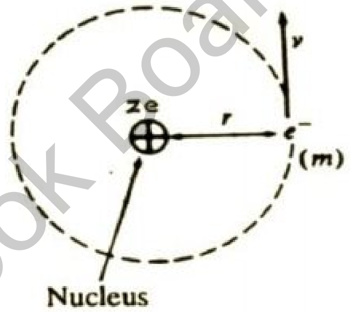


Fig 2.2: Hydrogen atom

Let '**e**' and '**m**' are the charge and the mass of electron, '**Ze**' is the charge on nucleus and '**r**' is the radius of an orbit in which electron is moving with a velocity '**v**'. According to Coulomb's law, the electrostatic force between the nucleus and electron is written as:

$$F = K\frac{Ze^{+}e^{-}}{r^2}$$

Where K is proportionality constant. It is equal to $\frac{1}{4\pi\varepsilon_o}$

Hence attractive force between nucleus and electron can be written as

$$F = \frac{Ze^2}{4\pi\varepsilon_o r^2}$$

But at the same time the centrifugal force on electron will be $\frac{mv^2}{r}$.

As far as electron remains in the same orbit, the two opposite forces are equal.

$$\frac{mv^2}{r} = \frac{Ze^2}{4\pi\varepsilon_o r^2}$$

$$mv^2 = \frac{Ze^2}{4\pi\varepsilon_o r} \text{ ----------- (i)}$$

$$r = \frac{Ze^2}{4\pi\varepsilon_o mv^2} \text{ -------- (ii)}$$

According to Bohr's thoery

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

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

Taking square on both sides,

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

Now put the value of v^2 in equation (ii)

$$r = \frac{Ze^2 4\pi^2 m^2 r^2}{4\pi\varepsilon_o mn^2h^2}$$

By rearranging, we get

$$r = \frac{\varepsilon_o n^2h^2}{Z\pi me^2} \text{ -------------- (iii)}$$

Where ε_o = Vacuum permittivity constant = 8.84×10^{-12} $C^2/J.m$

h = plank constant = 6.625×10^{-34} J.S

m = mass of electron = 9.11×10^{-31}kg

e = charge of electron = 1.602×10^{-19} C

Assembling all constants in equation (iii) we get

$$r = \frac{a^o n^2}{z} \text{ --------- (iv)}$$

and $$a^o = \frac{\varepsilon_o h^2}{Z\pi me^2}$$

a^o is Bohr's constant and its value is 0.529×10^{-8} cm or 0.529

Equation (iv) is used for the determination of radius of n^{th} orbit of hydrogen atom and hydrogen like ions viz He^+, Li^{+2} etc.

Example 2.1

Find the radius of 4th orbit of electron in Hydrogen atom.

Solution:

The radius of n = 4 for hydrogen atom can be determined by using Bohr's equation.

$$r = \frac{a^o n^2}{Z}$$

But for hydrogen Z = 1 hence

$$r = \frac{0.529 \times (4)^2}{1}$$

$$r = 8.46 Å$$

Calculation of Energy of electron in n^{th} orbit:

In any orbit of an atom, the total energy of an electron is the sum of potential energy (P.E) and kinetic energy (K.E).

$$E_{total} = K.E + P.E$$

P.E is the work done in bringing the electron from infinity to a point at a distance **r** from the nucleus and can be calculated as.

$$P \cdot E = \text{work done} = -\text{ force } \times \text{ displacement}$$

$$P \cdot E = \frac{-Ze^2}{4\pi\varepsilon_o r^2} \times r$$

$$P \cdot E = \frac{-Ze^2}{4\pi\varepsilon_o r}$$

Here negative sign indicates that P.E decreases when electron is brought from infinity to a point at a distance **r**.

Now total energy of electron is written as

$$E_{total} = \left(\frac{1}{2}mv^2\right) + \left(-\frac{Ze^2}{4\pi\varepsilon_o r}\right)$$

$$E = \left[\frac{1}{2}mv^2\right] - \left[\frac{Ze^2}{4\pi\varepsilon_o r}\right] \text{ - - - - - - - - - - - } (v)$$

The value of mv^2 is taken from eq(i) and insert in eq(v)

$$E = \frac{1}{2}\left[\frac{Ze^2}{4\pi\varepsilon_o r}\right] - \left[\frac{Ze^2}{4\pi\varepsilon_o r}\right]$$

$$E = \frac{-Ze^2}{8\pi\varepsilon_o r} \text{ - - - - - - - - - - - } (vi)$$

Taking the value of **r** from equation (iii) and put in equation (vi)

$$E = \frac{-Ze^2}{8\pi\varepsilon_o} \times \frac{\pi m Ze^2}{\varepsilon_o n^2 h^2}$$

$$E = \frac{-mZ^2e^4}{8\pi\varepsilon_o^2 n^2 h^2} \text{ ----------(vii)}$$

But for hydrogen atom, Z = 1, thus

$$E = \frac{-me^4}{8\varepsilon_o^2 h^2}\left(\frac{1}{n^2}\right) \text{ --------------- (viii)}$$

$$E = -K\left(\frac{1}{n^2}\right) \text{ ---------- (ix)}$$

Here K is a factor assembled by various constants present in equation (viii). Its value is 2.18×10^{-18} J/atom or 1312.8 KJ/mol

Example 2.2

Calculate the energy of an electron in L-shell of hydrogen atom (the value of K is 2.178×10^{-18} J/atom)

Solution:

L-shell correspond to second energy level (n = 2), Bohr's formula for energy of electron is given as

$$E = \frac{-K}{n^2}$$

$$E = \frac{-2.18 \times 10^{-18}}{(2)^2}$$

$$E = -5.43 \times 10^{-19} \text{ joule}$$

Calculation of Frequency of radiations emitted from an electron:

When an electron jumps from higher energy to lower energy orbit, a definite amount of energy is emitted. Let E_1 be the energy of n_1 orbit and E_2 is for n_2 orbit, the emitted energy is written as

$\Delta E = E_2 - E_1$

Considering eq. (viii), the energies of electron in n_1 and n_2 orbit of hydrogen atom are written as

$$E_1 = \frac{-me^4}{8\varepsilon_o^2 h^2}\left(\frac{1}{n_1^2}\right) \text{ and } E_2 = \frac{-me^4}{8\varepsilon_o^2 h^2}\left(\frac{1}{n_2^2}\right)$$

Where ε_o (the vaccume permittivity constant)

Now

$$\Delta E = \left[\frac{-me^4}{8\varepsilon_o^2 h^2}\left(\frac{1}{n_2^2}\right)\right] - \left[\frac{-me^4}{8\varepsilon_o^2 h^2}\left(\frac{1}{n_1^2}\right)\right]$$

$$\Delta E = \left[\frac{-me^4}{8\varepsilon_o^2 h^2}\left(\frac{1}{n_1^2}\right)\right] - \left[\frac{-me^4}{8\varepsilon_o^2 h^2}\left(\frac{1}{n_2^2}\right)\right]$$

$$\Delta E = \frac{-me^4}{8\varepsilon_o^2 h^2}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) \text{---------} (x)$$

$$\Delta E = 2.18 \times 10^{-18}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) \text{joule/atom---}(xi)$$

This equation is used for determining the emission or absorption of energy when electron jumps from one orbit to another.

If in equation (x), 'ΔE' is the energy difference appears as photon, it would have a frequency 'υ' and thus according to plank quantum theory it should be $\Delta E = h\upsilon$.

Now equation (x) is written as

$$h\upsilon = \frac{me^4}{8\varepsilon_o^2 h^2}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

And

$$\upsilon = \frac{me^4}{8\varepsilon_o^2 h^3}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) \text{Hertz---------}(xii)$$

Self Assessment

Using Bohr model, determine the energy in Joule of a photon produced when an electron in hydrogen atom jump from an orbit n = 5 to n = 2.

Calculation of wave numbers of photons:

The wave numbers of absorbed or emitted photons are calculated by following equation.

$$\upsilon = \overline{V}C$$

By substituting in equation (xii) we get

$$\overline{V}C = \frac{me^4}{8\varepsilon_o^2 h^3}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$$\overline{V} = \frac{me^4}{8\varepsilon_o^2 h^3 C}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

$$\text{But } \frac{me^4}{8\varepsilon_o^2 h^3 C} = R_H$$

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

Where R_H is Rydberg constant and its value is 109678 cm^{-1}

It may also be written as

$R_H = 10967800\ m^{-1}$

$R_H = 1.09678 \times 10^7\ m^{-1}$

Example 2.3

Calculate the wave numbers of photon when electron of a hydrogen atom jumps from 4th orbit to 2nd orbit (value of $R_H = 1.09678 \times 10^7\ m^{-1}$).

Solution:

$n_1 = 2$

$n_2 = 4$

$Z = 1$

$\overline{V} = ?$

Bohr's equation for wave number of photon is given as

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

$$\overline{V} = \left(1.09678 \times 10^7\right)\left[\frac{1}{(2)^2} - \frac{1}{(4)^2}\right]$$

$$\overline{V} = \left(1.09678 \times 10^7\right) \times 0.187$$

$$\overline{V} = 2.051 \times 10^6 m^{-1}$$

Self Assessment

What is the wave number of a photon produced when an electron falls from n = 5 level to n = 3 level in hydrogen atom.

2.2.2 The defects of Bohr's theory or Bohr's atomic Model

Confusions arose in Rutherford atomic model were explained in Bohr's atomic theory comprehensively but there were also some deficiencies in his proposed model.

(i) Bohr's model is only applicable to Hydrogen and those species which have single revolving electron around the nucleus (He^{+}, Li^{2+}, Be^{3+}). It could not explain the spectra of multi electrons systems like He, Li, Be, B etc.

(ii) According to Bohr's concept, electron revolves around nucleus in circular orbits. Later on, it was proved that electron did not move in single plane but in three dimensional spaces.

(iii) According to Bohr's concept, the electron in an atom is located at a definite distance from the nucleus and revolves with a definite velocity which is against the Heisenberg uncertainty principle.

(iv) Bohr's theory explain only the particle nature of electron and did not explain the wave nature of electron (de Broglie's hypothesis).

2.2.3 Spectrum and Hydrogen spectrum

A light which is composed of one kind of rays or wavelengths is known as monochromatic light. However, lights like sunlight, bulb light etc are composed of more than one type of rays. These are known as Poly chromatic lights. When a beam of any poly chromatic light is passed through a glass prism (placed in spectrometer), it splits (disperse) into several colors in order of increasing or decreasing wavelengths. The band of colors is called spectrum.

Continuous Spectrum

The spectrum which contains a continuous band of different colors is known as continuous spectrum. In this spectrum all colors are diffused into each other and the boundary line between the colours is not marked. When sun light passes through prism, it deviates and splits into continuous band of seven colors (VIBGYOR). Red color is least deviated and has high wavelength (7000Å) while violet color has maximum deviation and low wavelength (4000 Å). In rainy season, rainbow in sky is common example of continuous spectrum.

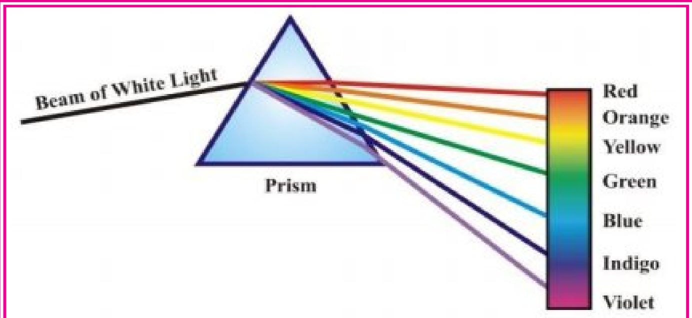


Fig 2.3: Continuous spectrum

Line Spectrum

When light emitted from electrically or thermally excited atoms is passed through prism, certain distinct lines separated by dark space are obtained, this is known as line spectrum.

A gas is excited by strongly heating or passing through electric discharge tube at low pressure. Each element emit light of specific wavelength therefore the number of lines and the distance between them depends upon the nature of element, so line spectra is used as 'Finger Print' for the identification of elements.

> **Do You Know?**
> The word spectroscopy is derived from Latin word spectrum, which means image and Greek word skopia, which means observation. It is the study of the absorbtion and emission of light by matter.

For example, line spectrum of sodium contains two yellow colored lines separated by a definite distance.

Line spectra of the elements give the information that electrons around nucleus have definite amount of energy and are arranged in definite energy levels E_1. After absorbing energy, they jump to an appropriate energy level E_2 and then return back. The difference in emitted energy of electrons, E_2-E_1 is equal to the energy absorbed.

Line spectrum can be seen by the following two ways.

(i) Absorption line spectrum

(ii) Emission line spectrum

Absorption Line Spectrum

When a beam of light is passed through an absorbing material like gaseous sample of an element, the element absorbs certain wavelengths while rest of wavelengths pass through it. The spectrum obtained consists of a series of dark lines with a bright background and known as absorption line spectrum.

Emission Line Spectrum

Sodium vapours lamp (street light), mercury vapours lamp, electrical discharge tube, hot solids or elements emits radiations of certain wavelengths. The spectrum which is formed from such radiations is called emission line spectrum.

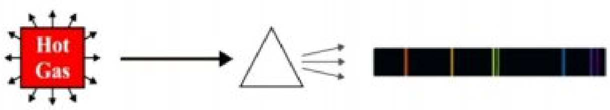


Fig 2.4: (a) Emission spectrum

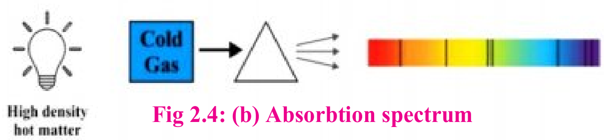


Fig 2.4: (b) Absorbtion spectrum

Hydrogen Spectrum

Hydrogen is a simplest atom. It has one electron revolving around mono positively charged nucleus. According to Bohr's theory, at ordinary temperature the electron in hydrogen atom resides in lowest energy level i.e in first orbit (ground state). When electric discharge is passed through hydrogen (Crook's tube) molecular hydrogen breaks up into atomic state. These atoms absorb energy from electric spark and come into excited state. Different electrons of different hydrogen atoms absorb different amount of energy and migrate to an appropriate different high energy levels (excited states). When electrons begin to revolve in higher energy level, hydrogen atoms become unstable. Then, electrons fall back to the original first orbit directly or to some other level by the emission of energies (photons). The emitted energies are equal to the difference of energies between the two levels. These radiations when pass through prism, a line emission spectrum of hydrogen is obtained. When hydrogen spectrum is viewed through high resolution spectrometer, several sharp fine lines are seen in the spectrum. The wavelengths of these lines lie into the ultraviolet, visible and infrared regions. These spectral lines are classified into five spectral series and named after their discoverers.

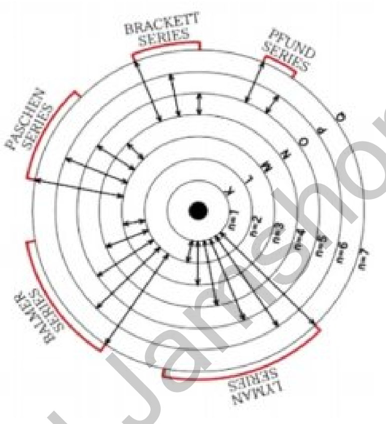


Fig 2.5: Transition of electron for five spectral lines of hydrogen spectrum

The wave numbers of each series of spectral lines is determined by the following equation.

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

Here R_H is Rydberg constant
$R_H = 1.09678 \times 10^7\ m^{-1}$
n_1 = lower energy level
n_2 = higher energy level

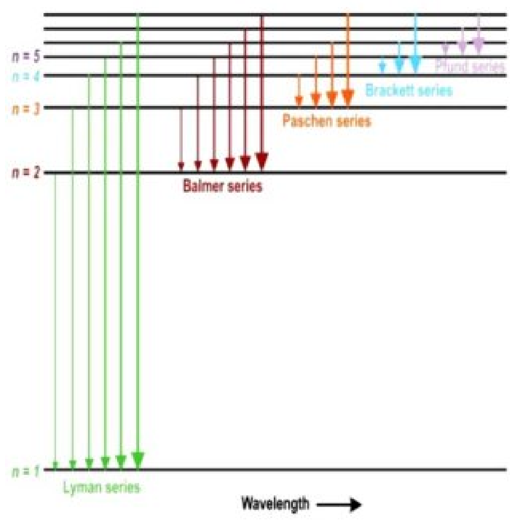


Fig 2.6: Origin of spectral lines of hydrogen spectrum as described by Bohr

Lyman Series: These spectral lines are produced when electrons fall back from $n_2 = 2,3,4,5,6,7 \cdots \infty$ to $n_1 = 1$ level. The range of wave number of this series lie in the ultraviolet region of the hydrogen spectrum (wave length less than 4000 Å).

Balmer Series: This series includes the spectral lines emitted by the transition of electrons occur from $n_2 = 3,4,5,6,7 \cdots \infty$ to $n_1 = 2$ level. All the wave numbers of this series are in the visible region (wave length 4000-7000 Å).

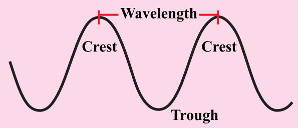
Paschen Series: In this series, transition of electron takes place from $n_2 = 4, 5, 6, 7 \ldots \infty$ to $n_1 = 3$ level. The wave number values of these spectral lines lie in near infrared region (wave length above 7000 Å).

Bracket Series: If the migration of electrons occur from $n_2 = 5, 6, 7, 8 \ldots \infty$ to $n_1 = 4$ level Bracket series is obtained. The wave number values of these spectral lines lie in mid infrared region.

P-fund Series: These spectral lines are produced when electrons fall back from $n_2 = 6, 7 \ldots \infty$ to $n_1 = 5$ level. The wave number values of these spectral lines lie in far infrared region.

Do You Know?

Wavelength is the distance between two identical points (Crest to crest or Trough to trough) in a wave. Its unit is cm, mm, m etc



2.3 PLANCK'S QUANTUM THEORY

This theory was given by German physicist Max Plank in 1900 A.D. to describe the emission and absorption of radiations. For this quantum theory, he was awarded Nobel Prize in 1918 A.D.

Postulates of plank's quantum theory

1) Atoms cannot absorb or emit energy continuously.
2) The absorption or emission of energy takes place in specified amounts called quanta (packets of energy). Quantum is a smallest unit of radiation energy, which can exist independently. A quantum of light energy is often called photon.
3) The energy of quantum (photon) is not fixed.
4) The amount of energy of quantum is directly proportional to the frequency of the radiations emitted or absorbed by the body.

$$E \propto \upsilon \quad \text{or} \quad E = h\upsilon$$

This is called as plank equation where h is called Planck's constant. Its value in SI units is 6.625×10^{-34} Js.

2.4 X-RAYS

In 1895 A.D, W. Roentgen accidently observed that when fast moving electrons (cathode rays) collide with metal anode in discharge tube, highly penetrating short wavelength radiations are produced. Initially, these rays were named as Roentgen rays but later on, these were called X-rays. These rays can penetrate through paper, glass, rubber, metal and human flesh.

2.4.1 Types of X-rays

Moseley used different anodes and analyzed the intensity of X-rays. On the basis of wavelengths there are two main types of spectral line of X-rays.

(a) **K-series**: These are produced by the elements (anodes) having large atomic number. These spectral lines have short wavelengths and high energy because transition of electrons occur from high energy level to low energy level (means there is big difference between two energy levels)

(b) **L-series**: Anodes having small atomic number produce these long wavelength spectral lines. These rays have low energy because there is small difference between two energy levels.

2.4.2 Production, Properties and Uses of X-rays

Production of X-rays

X-rays can be produced in a special type of discharge tube in which cathode is a heated filament. Under high voltage (5000 volt) and very low pressure (0.001mm) cathode rays are emitted from cathode and travel towards anode where they strike with high speed. The transition of electron occurs in the anode atoms which cause the production of X-rays photon.

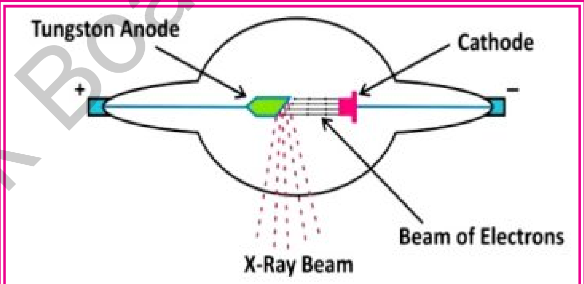


Fig 2.7: Production of X-rays in a discharge tube

Properties of X-rays

1. These are short wave length and high energy invisible electromagnetic radiations.
2. The range of wavelength lies between 0.1-10Å.
3. They travel with the speed of light.
4. Penetration power increases as energy of x-rays increases.
5. Like cathode rays, X-rays travel in straight line.
6. These rays are unaffected by electric or magnetic field.
7. They affect the photographic film.
8. X-rays possess enough energy to ionize gases, they damage and destroy the living cells.

Uses of X-rays

Initially, X-rays were used to assist in the setting of broken arm of a person. Then, uses of X-rays are increased with the passage of time.

1. These are used for the analysis of metallic substance or bullets in flesh.
2. Dentists use them to examine the defective or damaged teeth.
3. These are used for destroying the cancer cells.
4. At air ports, these are used for checking the baggage containing metallic knife, blade or weapons, transport of illegal goods etc.
5. In crystallography, these are used for the determination of structures of crystals. Thus, x-ray diffraction technique was developed.

2.4.3 Moseley's Law and Atomic Number

Moseley in 1913 A.D., comprehensively studied the different wavelengths of x-rays produced from the anodes of thirty eight (38) different elements from Aluminum to Gold. During his work he noticed that the wavelengths of x-rays decreased regularly with the increase of atomic masses of anode metals. He further observed that the frequencies of these radiations are directly proportional to the number of proton in the nucleus. Mosely law states that the square root of the frequency $\left(\sqrt{\upsilon}\right)$ is directly proportional to the atomic number (Z) of an element.

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

Where a and b are constant

2.5 RADIOACTIVITY

The phenomenon of radioactivity was given by French scientist Henry Becquerel in 1896 A.D. while working on uranium mineral called pitchblende. It was observed that there was continuous emission of some invisible rays. These rays were producing bright spots on photographic plates, ionizing gases, penetrating through thin metal sheets and producing fluorescence on zinc sulphide screen. Further, it was also observed that cooling, heating and compression did not affect these radiations. However, wrapping of pitchblende in lead sheet could stop the emission of these rays. Initially, these rays were named as Becquerel rays but later on, Marrie Curie coined the term radioactive rays. The elements, which emit these radiations spontaneously, are known as natural radioactive elements and the phenomenon is termed as natural radioactivity.

Do You Know?

Radiocarbon dating is a method for determining the age of an object containing organic material by using the properties of radiocarbon, a radioactive isotope of carbon. It is usually used to predict date ancient objects.

Marrie Curie and her husband, Pierre Curie isolated the radioactive components of the pitchblende mineral and separated two new radioactive elements Polonium and Radium.

Natural radioactive elements after spontaneous emission of rays, break down to more stable elements. This emission of radiations continues till the stable element lead (Pb-82) is formed.

$$ {}_{92}U^{238} \rightarrow {}_{90}Th^{234} + {}_{2}He^{4} $$

(Uranium) (Thronium) (α−rays)

Nuclear radiations are also produced when a stable element is bombarded or struck by a nuclear particle, this is known as artificial radioactivity.

Types of radiations:

Rutherford placed radioactive material in lead box and the radiations emitted from it were passed through an electric field. He observed three types of radiations- (α, β and γ). The separation of radioactive rays was also made by magnetic field. In magnetic field, alpha rays are deflected towards south pole and beta towards north pole.

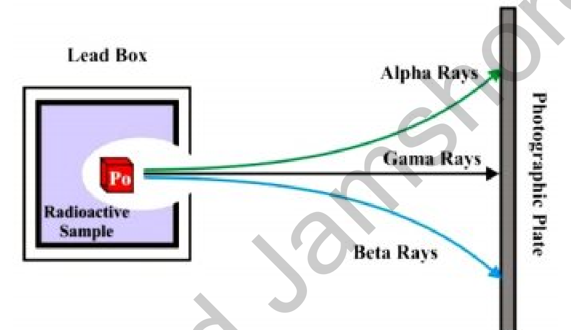


Fig 2.8
The production and separation of radioactive rays.

Table 2.2 **Comparison between the properties of alpha, beta and gamma rays.**

Property	Alpha, (∝)	Beta, (β)	Gamma, (γ)
Mass	4 amu	1/1836 amu	Zero
Composition	∝ particle is composed of two proton and two neutron like the composition of Helium nuclei	Beta particle is composed of electrons	These are electromagnetic radiations.
Deflection	They are deflected by electric and magnetic fields.	They are deflected by electric and magnetic fields but to the opposite direction of α-rays	They pass without any deflection in electric and magnetic field
Charge	+2 charge.	−1 charge.	Neutral
Ionization power	They have ionization power 100 times greater than β-rays and 10,000 times greater than γ-rays	They have lower ionization power than α- rays.	They have lower ionization power than α- rays and β-rays.
Penetration power	They have least penetration power.	They have moderate penetration power.	They have highest penetration power.
Speed	They travel 1/10 to 1/20th of light velocity.	They travel 9/10th of light velocity.	They travel as fast as velocity of light.

Self Assessment

(i) What is an $\propto$ particle? What is its approximate molar mass?
(ii) Compare the penetrating power and ionization power of $\propto$, β and γ rays.

2.5.1 Uses of Nuclear Radiations

Following are the important uses of nuclear radiations.

1. In medical field, nuclear radiations are used to diagnose, monitor and treat various diseases. These radiations are used to study the bone formation in mammals. Radiotherapy is most commonly used for the treatment of cancer. Radioisotopes has immense role in the growing field called nuclear medicine.
2. In agriculture field, radioisotopes are used to treat the seeds in the production of new varieties of crops.
3. They are used for the production of energy. Nuclear power stations like Karachi Nuclear Power Plant produce electrical energy.
4. In industries they are used to monitor the quality of products. Radioisotopes are also used to measure the density of metals and thickness of plastics.
5. In the field of geology, these are used to study the rocks.
6. Carbon-14 isotope is used to measure the age of fossils and artifacts (archeology).

2.6 QUANTUM NUMBERS AND ORBITALS

According to Bohr, an electron is a particle and form circular path around the nucleus. de Broglie considered the motion of electron as wave. Schrodinger described the movement of an electron as wave in three dimensional space around the nucleus. A region of three dimensional space around the nucleus where maximum probability of finding of electrons take place is known as orbital.

Schrodinger through his wave equation calculated some mathematical integers which are called quantum numbers. Quantum numbers describe energy levels, sub-levels and orbitals available for electron.

There are four quantum numbers.

2.6.1 Principal quantum number:

It is represented by (n). It has any positive value: 1,2,3 ….∞. The energy of an electron in an atom depends on this quantum number. Principal quantum number describes the size and energy of orbit and the distance of electron from the nucleus. Larger the value of **n**, the greater is the distance between electron and nucleus.

For n = 1, 2, 3…….., shells are indicated by K, L, M……..

The maximum number of orbitals in an orbit can be calculated by the formula $\mathbf{n^2}$ and the maximum number of electrons in an orbit are determined by using the formula $2n^2$.

2.6.2 Azimuthal quantum number:

It is also known as subsidiary quantum number. It is denoted by (ℓ). It governs the shape of orbitals. This can have values $\ell = 0$ to $(n - 1)$. When value of $\ell = 0$, the orbitals is called **s** orbital, when $\ell = 1$, it is **p** orbitals, when $\ell = 2$, **d** orbitals and $\ell = 3$, it is **f** orbital.

The maximum number of orbitals in an orbit are determined by the formula **$(2\ell + 1)$** whereas maximum number of electrons in a given sub-shell are calculated by **$2(2\ell + 1)$.**

2.6.3 Magnetic quantum number:

It is denoted by (m). It tells about the different orientations or directions of an orbital in space when subjected to magnetic field. Orbitals of same sub-shell have different orientation but same energy (degenerated orbitals) and same shapes.

The value of **m** depends upon the values of (ℓ).

Sub-shell	$(2\ell + 1)$ value	Orientations.($m = -\ell$ to 0 to $+\ell$)
S	One	1
P	Three	(–1, 0, +1)
d	Five	(–2, –1, 0, +1, +2)
f	Seven	(–3, –2, –1, 0, +1, +2, +3)

2.6.4 Spin quantum number:

It is denoted by (s). An electron while moving in an orbital around the nucleus also rotates or spins about its own axis. The spinning of an electron is either in clock-wise or anti clock wise. This spinning of electron is associated with magnetic field. In the same orbital one electron may spin clockwise (-1/2) or anti clockwise (+1/2), but in the same orbital, two electrons cannot have same spin quantum numbers.

Table 2.3 **A summary of quantum numbers, energy level, sub energy level and maximum number of electrons in sub energy level.**

Energy level (n)	Sub Energy levels (l)			Orientation of orbitals (m)
n	ℓ	Orbitals	Maximum electrons $2(2\ell + 1)$	$-\ell, 0, +\ell$
1	0	s	2	0
2	0 1	s p	2 6	0 -1, 0, +1
3	0 1 2	s p d	2 6 10	0 -1, 0, +1 -2, -1, 0, +1, +2
4	0 1 2 3	s p d f	2 6 10 14	0 -1, 0, +1 -2, -1, 0, +1, +2 -3, -2, -1, 0, +1, +2, +3

Self Assessment

In the following sets of quantum number, state which set is permissible? Explain why the other is not permissible?

(a) $n = 0, \ell = 0, m = 0, s = s = +\frac{1}{2}$ (b) $n = 1, \ell = 0, m = 0, s = -\frac{1}{2}$

2.6.5 Shapes of orbitals:

The regions of space around the nucleus where likelihood of finding an electron is maximum are called orbitals. Each orbital is associated with a particular size, shape and oriented around the nucleus.

s- orbital:

s-orbital has spherical shape in which the probability of finding the electron is uniformly distributed around the nucleus. It has only one possible orientation in space in the magnetic field because it spread over all the three axes uniformly. It has no nodel plane. Its size increases with the increase in value of **n**.

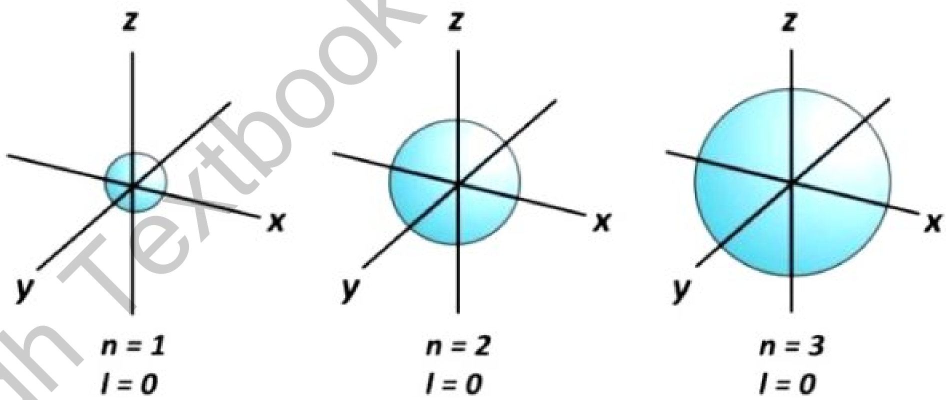


Fig 2.9 Shape of s-orbital

p-orbital:

The p orbitals are dumb-bell-shaped and they are oriented in space along the three mutually perpendicular axes (x, y, z), and are called $\mathbf{p_x}$, $\mathbf{p_y}$ and $\mathbf{p_z}$ orbitals. All the three **p** orbitals are perpendicular to each other. These are degenerate orbitals that are of equal energy.

Each **p** orbital has two lobes. One of which is labeled (+) and the other (-). Each lobe is like a pear. The point when the two lobes meet each other is usually referred to as the nodal plane along which the probability of finding the electron is zero.

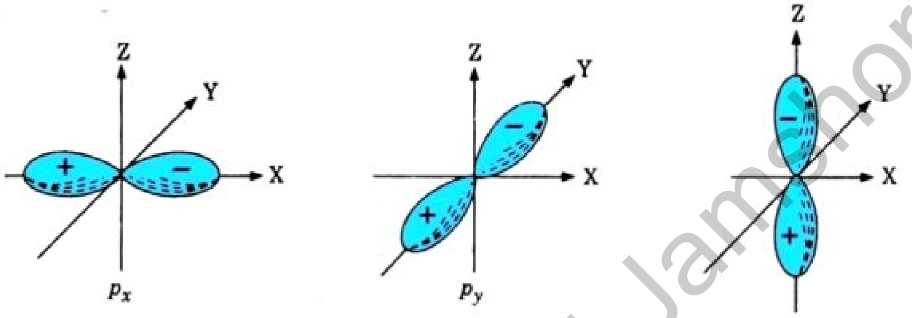


Fig 2.10 Shape and orientation of three p-orbitals

d-orbital:

d sub-shell is divided into five orbitals. Under the influence of magnetic field they have five directions dxy, dyz, dxz, dx^2y^2, dz^2 in space.

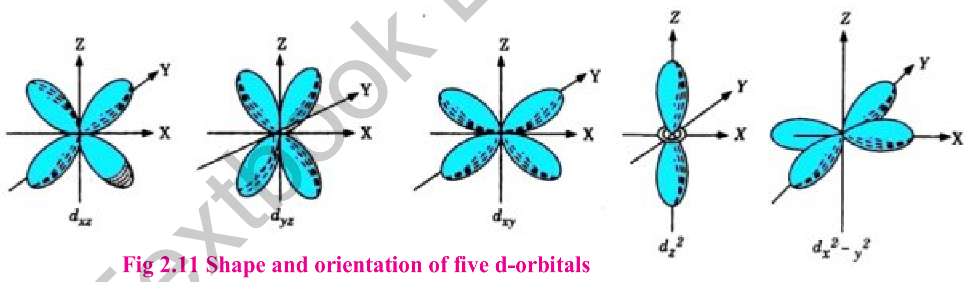


Fig 2.11 Shape and orientation of five d-orbitals

f-orbital:

These orbitals have seven directions when place in magnetic field.

2.7 ELECTRONIC CONFIGURATIONS

The distribution of electrons in the available sub-shells and orbitals (s, p, d, and f) is called electronic configuration. The superscript on sub-shells indicates the number of electrons and coefficient specifies the number of shell to which it belongs to. For example, the configuration of oxygen is O = $1s^2, 2s^2, 2p^4$. In $2p^4$, 2 is shell number and 4 number of electrons. Configuration of electrons in orbitals is indicated by a single arrow for one and double arrow for two electrons. Upward direction of an arrow (↿) shows the clock wise spin while downward direction (⇂) specify anti-clock wise spin of an electron.
Following are the rules for the electronic configuration.

(1) Pauli's Exclusion Principle

This principle was given by Wolfgang Pauli (1925 A.D.) and based on experimental observations. This principle states that, **"In an orbital of an atom, no two electrons can have the same set of four quantum numbers, at least one quantum number must be different"**.

According to this principle, two electrons in an orbital may have same values of three quantum numbers (n, ℓ , m) but the value of fourth quantum number (s) must be different. It means that if one electron of same orbital has clockwise spin then second electron must have anti-clockwise spin. Consider two electrons of helium which are lying in s-orbital of first shell ($1s^2$). The set of four quantum numbers will be written as:

Helium Atom	Quantum numbers			
	n	ℓ	m	s
First electron	1	0	0	$+\frac{1}{2}$
Second electron	1	0	0	$-\frac{1}{2}$

From the Pauli's principle, it is concluded that an orbital can accommodates only two electrons and these two electrons must have opposite spins (↿⇂).

(2) Aufbau principle

Aufbau is a German word, which means "building-up". Pauli named this principle as Aufbau Principle. **"According to this principle, electrons are filled progressively to the various sub-shells in the order of increasing energy, starting with the 1s sub-shell having lowest energy".**

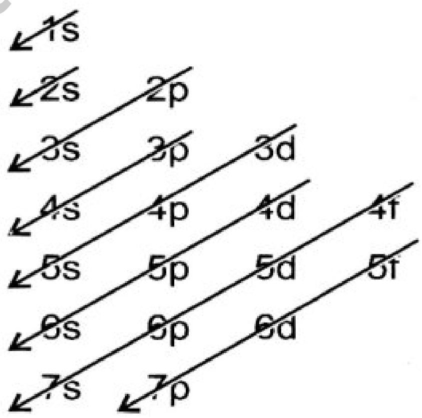


Fig 2.12: Systematic scheme of subshells in order of increasing energy.

Example 2.4

Write the electronic configuration of the following.
(a) S (Z=16) (b) Na^+ (Z=11) (c) Cl^- (Z=17)

Solution:

(a) S (Z=16) = $1s^2, 2s^2, 2p^6, 3s^2, 3p^4$

(b) Na^+ ion contains one electron less than Na atom, so its electronic configuration will be as:
Na^+ (Z=11) = $1s^2, 2s^2, 2p^6$

(c) Cl^- ion contains one electron greater than Cl atom, so its electronic configuration will be as:
Cl^- (Z=17) = $1s^2, 2s^2, 2p^6, 3s^2, 3p^6$

(3) (n + ℓ) Wiswesser rule:

This rule reminds us that the energy of an orbital depends upon the principal quantum number (n) and azimuthal quantum number (ℓ).

According to this rule, electrons are filled in various orbitals in the order of increasing (n + ℓ) value. **"Those orbitals which have lower (n+ ℓ) value are filled first. In case, if two orbitals which have same (n + ℓ) value, then orbital having lower 'n' value will be filled first".**

Do You Know?

The term iso electronic refers to those atoms or ions which have the same number of electrons and same electronic configuration.

For example Ne (Z = 10), Na^+(z = 11) and F^-(Z=9) have the same number of electrons (10) and same electronic configuration ($1s^2 2s^2 2p^6$).

To understand n + ℓ rule, let us consider the filling of electron in 3d and 4s orbitals. The n + ℓ value of 3d orbital (n = 3; ℓ = 2) is 5 but for 4s orbital (n = 4; ℓ = 0) is 4. Hence 4s orbital which has lower value of (n + ℓ) will fill first. Likewise, 4p orbital fills before 5s although the (n + ℓ) value for both is same, but 4p orbital has the lower value of the principal quantum number 'n'.

Self Assessment

Identify the orbital of higher energy in the following pairs.
(i) 4s and 3d (ii) 4f and 6p (iii) 5p, 6s (iv) 4s, 3d

(4) Hund's rule of Maximum Multiplicity:

Orbitals of same sub-shell possess same energy and are known as degenerated orbitals.
For example:
p-sub shell consists of three orbitals (p_x, p_y and p_z), all have same energy. Similarly, five orbitals of d-sub shell are also degenerated orbitals.
3p = 3px, 3py, 3pz
3d = 3dxy, 3dxz, 3dyz, $3dx^2$-y^2, $3dz^2$

In 1927 German Physicist Friedrich Hund gave his rule for filling of electrons in degenerated orbitals known as “Hund’s rule of Maximum Multiplicity”. According to this rule, **“In available degenerated orbitals (p, d and f), electrons are distributed in such a way that maximum number of half filled orbitals (single electron in orbital) are obtained”.** For example, if we have three electrons to fill the 2px, 2py, 2pz orbitals, we will fill single electron in each orbital $2px^1$, $2py^1$, $2pz^1$ rather than double electrons $2px^{\uparrow\downarrow}$, $2py^1$, 2pz. Unpaired electrons are more stable than paired electrons because paired electrons create repulsion. Electronic configurations of some elements considering Hund’s rule are given below:

$$C = 6 = 1s^{\uparrow\downarrow}\ 2s^{\uparrow\downarrow}\ 2px^1\ 2py^1\ 2pz$$

$$N = 7 = 1s^{\uparrow\downarrow}\ 2s^{\uparrow\downarrow}\ 2px^1\ 2py^1\ 2pz^1$$

$$O = 8 = 1s^{\uparrow\downarrow}\ 2s^{\uparrow\downarrow}\ 2px^{\uparrow\downarrow}\ 2py^1\ 2pz^1$$

$$F = 9 = 1s^{\uparrow\downarrow}\ 2s^{\uparrow\downarrow}\ 2px^{\uparrow\downarrow}\ 2py^{\uparrow\downarrow}\ 2pz^1$$

$$Ne = 10 = 1s^{\uparrow\downarrow}\ 2s^{\uparrow\downarrow}\ 2px^{\uparrow\downarrow}\ 2py^{\uparrow\downarrow}\ 2pz^{\uparrow\downarrow}$$

Society, Technology and Science

Firework Displays

We specially, see the colorful fascinating display of fireworks on the occasions of wedding ceremony, happy New Year, birth day parties, winning the elections or religious special days etc. These different colors of fireworks are due to the presence of different metallic compounds in burning material. The compounds of Copper give bluish green, Barium green and Strontium deep red colors.

In eighteenth century, scientists began to study these colors displayed during fireworks and used flame tests. From the light of flames line spectra were formed and analyzed the elements. These flame tests and spectra helped in the study elements and structures of atoms.

Activity

The purpose of this activity is to comprehend the phenomenon of line spectrum. You definitely have a table salt (NaCl) in your kitchen. Sodium, being an alkali metal, gives distinct line spectrum when the salt is strongly heated over the flame. You can perform this activity quite easily.

Paste a small amount of table salt on a glass rod and put it over the flame. When it gets sufficient heat, the electron (of sodium) in the lower energy level jumps to higher energy level thus becomes excited and when it gets back to lower energy level, it consequently produces yellow spectrum which can be observed easily.

SUMMARY with Key Terms

- **Atom** is a complex organization of matter and energy. It is the smallest particle of matter which may or may not exist free in nature and takes part in chemical reactions.
- **Bohr Atomic Theory** describes that an atom has fixed number of circular orbits in which electrons revolve. An electron requires energy when it jumps from lower orbit to higher orbit and loses energy when it returns to lower orbit.
- **Continuous spectrum** is produced when a beam of white light is passed through a prism. This spectrum consists of a band of seven colours (VIBGYOR). The colours are diffused into each other and have no distinct boundary line.
- **Absorbtion Line spectrum** consists of a series of dark lines with bright background. It is obtained when a beam of light is passed through an absorbing material.
- **Emission Line spectrum** consists of a series of bright lines with dark background. It is obtained when light emitted from a hot solid or discharge tube.
- **Lyman series** is formed when electrons fall back from $n_2 = 2, 3, 4, 5, 6, 7 \ldots\infty$ to $n_1 = 1$ level. This series lie in UV region.
- **Balmer series** is formed when electrons fall back from $n_2 = 3, 4, 5, 6, 7\ldots \infty$ to $n_1 = 2$ level. This series lie in visible region.
- **Paschen series** is formed when electrons transition takes place from $n_2 = 4, 5, 6, 7 \ldots\infty$ to $n_1 = 3$ level. This series lie in near IR region.
- **Bracket series** is formed when electrons jump down from $n_2 = 5, 6, 7, 8 \ldots\infty$ to $n_1 = 4$ level. This series lie in mid IR region.

- **Pfund series** is produced when electrons jump back from $n_2 = 6, 7 \ldots\infty$ to $n_1 = 5$ level. This series lie in far IR region.
- **X-rays** are short wavelength and high penetrating electromagnetic radiations. These are produced when fast moving electrons hit metal anode in a discharge tube.
- **Radioactivity** is a phenomenon in which nucleus of an atom is splitted and emission of rays takes place. Henry Becquerel gave the concept of radioactivity.
- **Alpha rays** consist of positive charge particles. They deflect toward negative pole in an electric field.
- **Beta rays** consist of negative charge particles. They deflect toward positive pole in an electric field.
- **Gamma rays** are shortest wave length rays which pass through electric and magnetic field without any deflection.
- **Principle quantum number** is indicated by 'n'. It describes the size and energy of an orbital.
- **Azimuthal quantum number** is indicated by 'ℓ'. It describes the shape of orbitals.
- **Magnetic quantum number** is indicated by 'm'. It describes the orientation of orbitals in space in an applied magnetic field.
- **Spin quantum number** is indicated by 's'. This quantum number describes the spin of electron in an orbital. Spin of electron may be clock wise $\left(-\frac{1}{2}\right)$ or anti clock wise $\left(+\frac{1}{2}\right)$.
- **Orbital** is a space around the nucleus where maximum (95%) finding of an electron takes place.
- **Electronic configuration** is the distribution of electrons in the available sub-shells and orbitals (s, p, d, and f).
- **Pauli Exclusion Principle** states that "In an atom no two electrons can have same set of four quantum numbers. At least one quantum number must be different".
- **Aufbau Principle** tells that electrons are filled progressively to various sub-shells in the order of increasing energy, starting with the 1s sub-shell having lowest energy.
- **$n + \ell$ rule** tells that, those orbitals are filled first which have lowest $(n + \ell)$ value. When two orbitals have same $(n + \ell)$ value, the orbitals having lower (n) value will be filled first.
- **Hund's rule** states that when degenerated orbitals are available and more than one electron are to be filled in them, they should be filled in separate orbitals in such a way so as to give maximum number of unpaired electron and have the same direction of spin.

EXERCISE

Multiple Choice Questions

1. Choose the correct answer

(i) Bohr's theory is not applicable to which of the following species.
(a) H (b) H^+
(c) He^{+1} (d) Li^{+2}

(ii) Nitrogen has the electronic configuration $1s^2 2s^2 2p_x^1 2p_y^1 2p_z^1$ and not $1s^2 2s^2 2p_x^2 2p_y^1$
This is determined by.
(a) Aufbau principle (b) Pauli's rule
(c) Hund's rule (d) $n + \ell$ rule

(iii) Quantum number values for 3s orbital are
(a) $n = 0, \ell = 1$ (b) $n = 1, \ell = 0$
(c) $n = 3, \ell = 1$ (d) $n = 3, \ell = 0$

(iv) The radius of first orbit of hydrogen atom is .....
(a) 529 Å (b) 52.9 Å
(c) 5.29 Å (d) 0.529 Å

(v) Line spectrum is used as a tool for the identification of ....
(a) Colors (b) Electrons
(c) Elements (d) Molecules

(vi) In 1935 A.D. James Chadwick was awarded Nobel Prize because ...
(a) He discovered proton
(b) He discovered neutron
(c) He determined the radius of hydrogen atom
(d) He gave the rules for electronic configuration

(vii) When 4d orbital is filled, the next electron enter into
(a) 5s (b) 5p
(c) 5d (d) 6s

(viii) Which of the following is not an iso electronic pair
(a) Na^+ and Ne (b) Na^+ and F^-
(c) Na and Ca (d) Na^+ and Mg^{+2}

(ix) Balmer series appears in the hydrogen spectrum if electron jump from any appropriate higher energy orbit to
(a) Second orbit (b) Third orbit
(c) Fourth orbit (d) Fifth orbit

(x) The shape of orbital for which $\ell = 0$ is
(a) Spherical (b) Dumbbell
(c) Double dumbbell (d) Complicated

Short Questions

1. Differentiate between Continuous and Line spectrum.
2. Give three properties of each $\propto$, β and γ rays.
3. What is the shape of orbitals for which $\ell = 0$ and $\ell = 1$.
4. How does an orbital different from orbit?
5. Explain why the filling of electron in 4s orbital takes place prior to 3d?
6. Mention the defects of Bohr's atomic model.
7. Write down the electronic configuration of the following.
   (i) Fe (Z= 26), (ii) Br^- (Z= 35), (iii) Ca^{+2} (Z= 20)

Descriptive Questions

1. (a) State the postulates of Bohr atomic theory.
   (b) Drive an expression for the frequency of radiation emitted from an electron. Given that

$$E = \frac{-me^4}{8\varepsilon_o^2 h^2 n^2}$$

2. What are X-rays? How are they produced? Give their properties and uses.
3. State and illustrate the following rules of electronic configuration.
   a) Pauli's exclusion rule b) Hund's rule of maximum multiplicity
4. Explain hydrogen spectrum in term of Bohr's theory.
5. Describe four quantum numbers needed to specify an electron in an atom. Write all possible values of ℓ, m and s for n = 2.

Numerical Questions

1. Calculate the radius of hydrogen in 3^{rd} orbit (Bohr constant for hydrogen is 0.529Å).
2. A photon of wave number $23 \times 10^5\ m^{-1}$ is emitted when electron undergoes a transition from a higher energy orbit to n = 2. Determine the orbit from which electron fall and also the spectral line appears in this transition of electron.
   (The value of Rydberg constant is $1.09678 \times 10^7\ m^{-1}$).