Inorganic

ATOMIC STRUCTURE

1. ATOM & MOLECULES

(a) The smallest particle of a matter that takes part in a chemical reaction is called an atom. The atom of all gases except those of noble gases, cannot exist in free state. These exist in molecular form. The molecules of hydrogen, nitrogen, oxygen and halogens are diatomic (H_2, N_2). Phosphorus molecule is tetratomic and that of sulphur is octa atomic.

(b) The smallest particle of a matter that can exist in free state in nature, is known as a molecule.

(c) Some molecules are composed of homoatomic atom, e.g., H_2, O_2, N_2, Cl_2, O_3 etc., while the molecules of compounds are made up of two or more heteroatomic atoms e.g., HCl, NaOH, HNO_3, $CaCO_3$, etc.

2. DALTON'S ATOMIC THEORY

The concepts put forward by John Dalton regarding the composition of matter are known as Dalton's atomic theory. Its important points are as follows.

(a) Every matter is composed of very minute particles, called atoms that take part in chemical reactions.

(b) Atoms cannot be further subdivided.

(c) The atoms of different elements differ from each other in their properties and masses, while the atoms of the same element are identical in all respects.

(d) The atoms of different elements can combine in simple ratio to form compounds. The masses of combining elements represent the masses of combining atoms.

(e) Atom can neither be created nor destroyed.

➢ **MODERN CONCEPT :**

Many of the concepts of Dalton's atomic theory cannot be explained. Therefore, foundation of modern atomic theory was laid down by the end of nineteenth century. The modern theory is substantiated by the existence of isotopes, radioactive disintegration, etc. The important points of the modern atomic theory are as follows.

(a) Prof. Henri Bacquerel discovered the phenomenon of radioactivity and found that an atom is divisible.

(b) An atom is mainly composed of three fundamental particles, viz. electron, proton and neutron.

(c) Apart from the aforesaid three fundamental particles, many others have also been identified, viz. positron, meson, neutrino, antiproton, etc.

(d) Soddy discovered the existence of isotopes, which were atom of the same element having different masses. For example, protium, deuterium and tritium are atoms of hydrogen having atomic masses 1, 2 and 3 a.m.u. respectively.

(e) Atoms having same mass may have different atomic numbers. These are known as isobars. For example, ${}^{40}_{18}Ar$ and ${}^{40}_{20}Ca$.

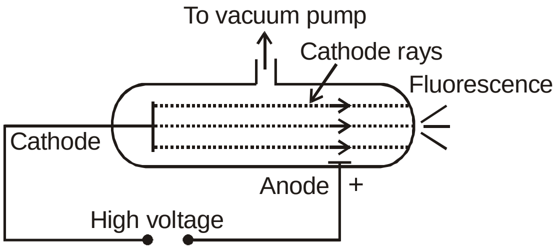
(f) Atoms of elements combines to form molecules.

(g) It is not necessary that the atoms should combine in simple ratio for the formation of compounds. The atoms in non-stoichiometric compounds are not present in simple ratio. For example, in ferrous sulphide crystals, iron and sulphur atoms are present in the ratio of 0.86 : 1.00.

(h) Atoms participate in chemical reactions.

3. CATHODE RAYS (DISCOVERY OF ELECTRON)

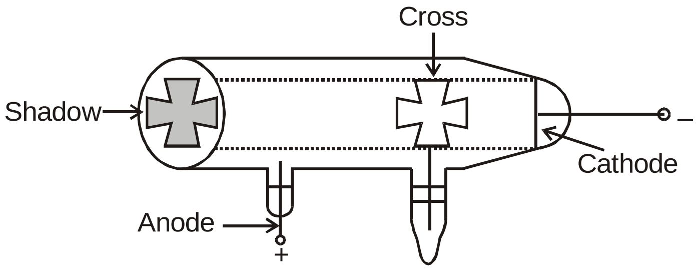
Dry gases are normally bad conductors of electricity. But under low pressure, i.e., 0.1 mm of mercury or lower, electric current can pass through the gases. Julius Plucker in 1859 found that a type of rays, called cathode rays, emit from the cathode when electricity is passed through a discharge tube. William Crookes (1879), J.J.Thomson and many other scientists studied the properties of cathode rays and came to the conclusion that the cathode rays of same properties are obtained using any gas or any cathode material.



Production of cathode rays

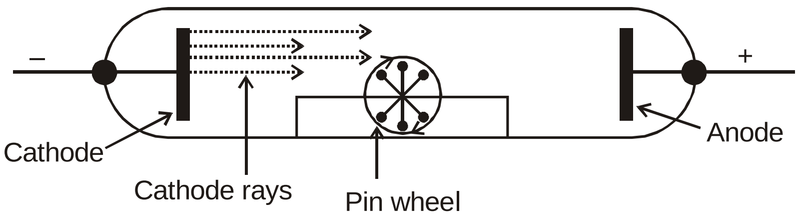
The salient features of cathode ray are as follows.

(a) Cathode rays travel in a straight line. This indicates that the formation of a shadow when an opaque object is placed in its path.



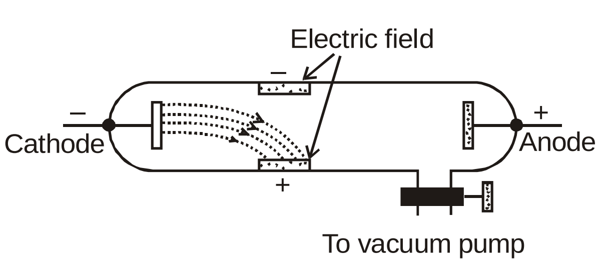
Cathode rays travel in a straight line

(b) If a light metal pinwheel is placed in the path of cathode rays, the wheel starts revolving. This proves that is cathode rays consist of tiny particles having momentum.

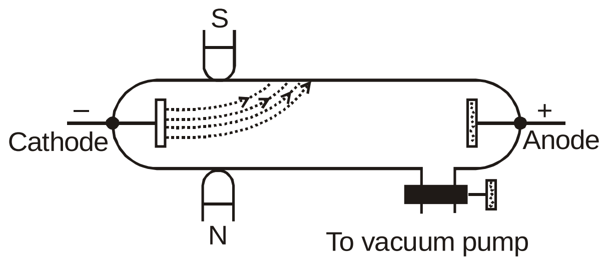


Cathode rays consist of tiny particles

(c) Cathode rays get deviated in electrical and magnetic fields. This proves that they are composed of charged particles. Their derivation towards anode indicates their negatively charged nature. The direction of their deviation in magnetic field depends on pole of the magnet which has been placed near the cathode ray tube.



(a) Electric Field



(b) Magnetic Field

(d) Cathode rays produce green fluorescence on the walls of the glass tube.

(e) Cathode rays produce incandescence in at thin metal foil.

(f) Cathode rays effect the photographic plate.

(g) Cathode rays ionize gases proving that they are charged.

(h) Cathode rays penetrates across a thin metal foil.

(i) Cathode rays produce X-rays when they hit a piece of tungsten or any other metal having high melting point.

➢ **NATURE OF CATHODE RAYS :**

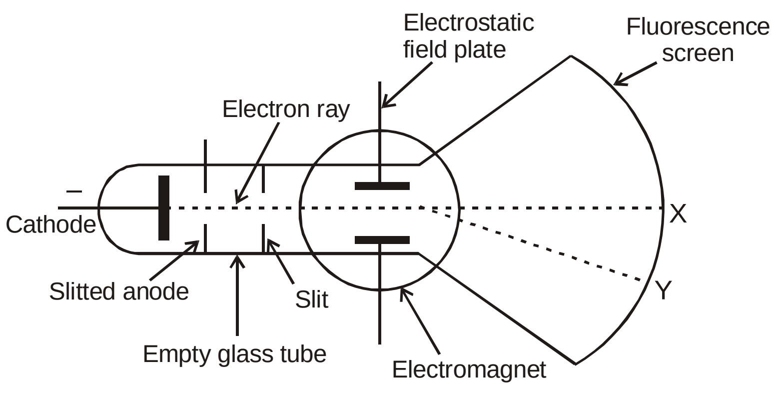
J.J. Thomson (1897) proved through experiments that.

(a) Cathode rays are composed of extremely tiny negatively charged particles (electrons).

(b) The ratio of negative charge (e) and mass (m) for cathode ray particle (electrons) is a constat. This ratio is independent of the material used in the preparation of the electrodes of the discharge tube or the gas filled in it. Thus, e/m of an electron is a universal constant.

$$\frac{\text{charge on electron}}{\text{mass of electron}} = \frac{e}{m} = 1.76 \times 10^{8} \text{ Coulomb/gm}$$

In addition to the above proofs, photoelectric effect, thermionic effect and emission of beta particles from radioactive elements also confirm that electron is an essential constituent of matter. These negatively charged tiny particles discovered by **Thomson**. It is denoted by e^{-} or ${}_{-1}e^{0}$.



Determination of $\frac{e}{m}$ of an electron

4. POSITIVE RAYS OR CANAL RAYS : DISCOVERY OF PROTON

Eugene Goldstein in 1886 found that a dim glow is visible behind the cathode when an electric discharge is passed through a perforated cathode in a discharge tube filled with a gas at low pressure. These new type of rays travel from anode to that cathode. Goldstein gave the name canal rays to these rays because these rays cross the canals of the cathode and reach the other side. **W.Wein** in 1897 proved through experiments that the canal rays consist of positively charged particles. **J.J. Thomson** gave the name **positive rays** to them because they are composed of positively charged particles.

- **PROPERTIES OF POSITIVE RAYS :**

(a) Positive rays travel in the direction opposite to that of cathode rays.

(b) Positive rays travel in straight line.

(c) Positive rays affect photographic plate.

(d) Positive rays are deviated in the electric and magnetic fields. The direction of their deviation proves the presence of positive charge on their particles.

(e) Positive rays pass across a very thin sheet of metal. But their penetrating power is less than that of cathode rays.

(f) Positive rays produce fluorescence and phosphorescence.

- **NATURE OF POSITIVE RAYS :**

Thomson and Wein studied the nature of positive rays and proved with the help of experiments that.

(a) Positive rays are composed of positively charged particles.

(b) The ratio (e/m), of positive charge (e) and mass (m) for the particles of positive rays depends on the nature of the gas filled in the discharge tube. The value of e/m for the particles of positive rays obtained from different gases is different. The e/m value for positive rays is not a universal constant. Thomson and Wein found out through experiments that the maximum value of e/m is for particles of positive rays of hydrogen gas.

(c) Experiments proved that for a positively charged particle (H^+) of the positive rays of hydrogen gas $\frac{e}{m} = 9.578 \times 10^4$ coulomb per gram. If we suppose that the charge (e) of this particle is 1.602×10^{-19} coulomb unit positive charge, the mass (m) of the particle will be 1.6725×10^{-24} gram. The particle (H^+) of the positive rays of hydrogen gas having 1.602×10^{-19} coulomb positive charge and 1.6725×10^{-24} gram mass is called a proton.

5. PROTON

(a) Proton is a fundamental particle of an atom. It is an essential constituent of every matter.

(b) The credit for the discovery of proton goes to **Goldstein**.

(c) Proton bears one unit positive charge.

(d) **Thomson** and **Wein** estimated the value of e/m as 9.578×10^4 coulomb per gram for the positively charged particle proton.

(e) The amount of positive charge (e) on proton is 1.602×10^{-19} coulomb or 4.8×10^{-10} e.s.u.

(f) Mass of proton (m) $= 1.6725 \times 10^{-24}$ gram ; $= 1.6775 \times 10^{-17}$ kilogram
$= 1.6725 \times 10^{-29}$ quintal ; $=$ 1837 times that of electron
$=$ 1.00757 a.m.u. ; $=$ Mass of hydrogen atom

Mass of proton (m) in a.m.u $= \dfrac{1.6725 \times 10^{-24}}{1.66 \times 10^{-24}} = 1.00757$ a.m.u.

(g) Mass of proton (m) multiplied by Avogadro number (6.023×10^{23}) gives molar mass of proton. Thus
Gram molecular mass of proton $= 1.6725 \times 10^{-24} \times 6.023 \times 10^{23} = 1.008$ (Approx)

(h) Proton is present in the nucleus of an atom.

(i) The number of electrons is equal to the number of protons in a neutral atom.

(j) The atomic number of an atom is equal to the number of protons present in the nucleus of that atom.

(k) Proton is the nucleus of protium i.e. the common hydrogen atom.

(l) Proton is ionized hydrogen atom, i.e. (H^+)

(m) Proton is obtained when the only one electron present in hydrogen atom is removed. Hydrogen atom consists of only one electron and one proton.

6. ELECTRON (e^- or $_{-1}e^0$)

(a) Electron is a fundamental particle of an atom, which is an essential constituent of every matter.

(b) The credit for discovery of cathode rays goes to **Sir William Crookes** while the credit for discovery of negatively charged electron goes to **J.J. Thomson**. The name 'electron' was first given by **Stony**.

(c) A unit negative charge is present on electron.

(d) The value of $\frac{e}{m}$ was found to be 1.76×10^8 coulomb/gram by **Thomson**.

(e) **R.A. Mulliken** calculated the charge on an electron by his famous **Oil Drop Experiment**. The value came out to be 1.6012×10^{-19} coulomb or 4.803×10^{-10} e.s.u.

(f) The value of e/m of an electron is known as its **specific charge**. With the help of this specific charge and the charge on the electron (determined by Mulliken), the mass of the electron could be calculated as follows.

$$\frac{e}{e/m} = \frac{1.6012 \times 10^{-19}\,\text{coulomb}}{1.76 \times 10^{8}\,\text{coulomb/gram}} \qquad = 9.1091 \times 10^{-28}\ \text{gram}$$

$$= 0.0005486\ \text{a.m.u.}$$

$$= 1/1837^{\text{th}}\ \text{of H atom}$$

(g) Molar mass of electron is obtained on multiplying mass of electron by Avogadro number (6.023×10^{23}). Therefore gram molecular mass of electron is as follows.

$$= 9.1091 \times 10^{-28} \times 6.023 \times 10^{23}$$

$$= 5.483 \times 10^{-4}$$

(h) Electron is very much lighter than an atom of the lightest element hydrogen. The gram molecular mass of hydrogen is 1.008. Therefore the ratio of gram molecular mass of hydrogen and that of electron is

$\frac{1.008}{5.483 \times 10^{-4}} = 1837$. In other words, an atom of hydrogen (or a proton) is 1837 times heavier than electron.

$$\frac{\text{Mass of H atom}}{\text{Mass of electron}} = \frac{1.67 \times 10^{-24}}{5.483 \times 10^{-28}} = 1837$$

(i) The mass of 1.1×10^{27} electrons is one gram.

(j) The mass of one mole of electrons is 0.5583 mg.

(k) The amount of charge on one mole of electrons is one faraday or 96500 coulomb.

(l) The mass of an electron at rest is called static electron mass and its value is 9.1091×10^{-28} gram.

(m) The mass of an electron in motion is calculated with the help of the following expression.

$$\text{Mass of electron in motion (m)} = \frac{\text{Rest mass of electron}}{\sqrt{\left[1 - \left(\frac{v}{c}\right)^2\right]}}$$

where v is velocity of electron and c is velocity of light.

When v = c, the mass of the electron in motion becomes infinity.

Therefore the mass of an electron increases with increase in its velocity due to which specific charge e/m on it decreases.

(n) Electron, being the fundamental particle of an atom, takes part in chemical combination.

(o) The physical and chemical properties of an element depend on the distribution of electrons in its outermost energy level.

7. DISCOVERY OF NEUTRON

Penetrating rays are emitted on bombarding α-particles on the elements like beryllium, boron and aluminium. **James Chadwick** in 1932 studied the nature of these radiation and came to the conclusion that these rays are composed of very tiny electro neutral particles. The mass of these particles is almost equal to that of the hydrogen atom. This particle is called neutron and is denoted by the symbol, ${}_0n^1$.

➢ **NEUTRON (${}_0n^1$)**

(a) It is a fundamental particle of atom that is present in the nuclei of all atoms except hydrogen or protium.

(b) It was discovered by James Chadwick in 1932.

(c) It is an electro neutral particle, i.e. it does not have any positive or negative charge on it.

(d) The mass of a neutron is almost equal to that of a proton. Actually it is a little bit heavier than proton. Its mass (m) is as fallows :

Mass (m) of a neutron $= 1.6748 \times 10^{-24}$ gram = Approximately mass of a proton

(e) Neutron is relatively heavier out of the three fundamental particles of an atom.

(f) Molar mass of a neutron is obtained by multiplying the mass (m) of a neutron with Avogadro number (6.023×10^{23}). Therefore the gram molecular mass of a neutron is $1.6748 \times 10^{-24} \times 6.23 \times 10^{23} = 1.00893$.

(g) The atomic mass is equal to the total mass of all the protons and neutrons present in the atom.

(h) Isotopes are formed as a result of difference in the number of only neutrons in the nuclei of atoms.

(i) It is assumed that a neutron is a result to joining together of an electron and a proton. A neutron, being unstable, decays as fallows :

$${}_0n^1 \longrightarrow {}_{+1}P^{1-} + {}_{-1}e^0 + {}_0q^0 \text{ (antineutrino)}$$

Its half-life is 20 minutes.

(j) The density of neutrons is of the order of 1×10^{12} Kg/c.c.

8. OTHER PARTICLES OF ATOM

(a) **Positron :** It was discovered by **C.D. Anderson** in 1932. It beards a unit positive charge and its mass is equal to that of an electron. Thus its mass regarded as negligible. It merges with an electron and emit electromagnetic radiations. It is denoted by e^+.

(b) **Meson : Yukawa** in 1935 discovered this particle. Different types of meson particles are possible in the atom. These are called meson family.

(c) **Neutrino : Pauling** discovered these particles in 1927. They do not bear any charge, i.e. they are electro neutral particle.

(d) **Antiproton : Segre** discovered this particle in 1956. It bears a unit negative charge and its mass is equal to that of a proton.

9. CLASSIFICATION OF ATOMIC PARTICLES

➢ STABLE PARTICLES

Properties of Stable Fundamental Particles

	Particle	Symbol	Charge	Mass*	Mass**	Spin***
1.	Proton	p	+	1.00758	1, 836	1/2
2.	Electron	e^-, β^-	–	0.0005486	1	1/2
3.	Positron	e^+, β^+	+	0.0005486	1	1/2
4.	Neutrino	v	0	0.000022	0.04	1/2
5.	Antiproton	p^-	–	1.00758	1, 836	1/2
6.	Graviton	G	0	0	0	2
7.	Photon	γ	0	0	0	1

*Physical atomic weight unit ${}_8O^{16}$ = 16,00,000

**Mass with respect to e, where e = 9.11×10^{-28} gram

*** $\frac{h}{2\pi}$ unit

➢ UNSTABLE PARTICLES

Properties of Some Unstable Fundamental Particles

	Particle	Symbol	Charge	Mass*	Mass**	Spin***
1.	Neutron	n	0	1.00893	1, 836	1/2
2	.Negative μ meson	μ^-	–	0.1152	210	1/2
3.	Positive μ meson	μ^+	+	0.1152	210	1/2
4.	Neutral π meson	π^0	0	0.1454	265	0
5.	Negative π meson	π^-	–	0.1514	276	0
6.	Positive π meson	π^+	+	0.1514	276	0

➢ **Distinction of α, β and γ Rays**

	Property	α Ray	β Ray	γ Ray
1.	Velocity	2×10^9 cm/sec	2.8×10^{10} cm/sec	Equal to velocity of light. 3×10^{10} cm/sec
2.	Penetration power	Very low	About 10 times to that of α rays	About 1000 times to that that of α rays
3.	Charge and mass	2 unit positive charge and 4 unit mass	1 unit negative charge and zero mass	Magnetic radiations of very high frequency
4.	Effect of ZnS plate	Produce fluoresence	No effect	No effect
5.	Nature	He^{+2} or helium nuclei, denoted by the symbol $_2He^4$.	Denoted by the symbol $_1e^0$ or electron	denoted by the symbol $_0\gamma^0$
6.	Megnetic field	Deviation towards cathode	Deviation towards anode	No effect
7.	Nature of the product	$_4A^9 \xrightarrow{\alpha^-} {_2A^5}$	$_4A^9 \xrightarrow{\beta^-} {_5A^9}$	$_4A^9 \xrightarrow{\gamma^-} {_4A^9}$

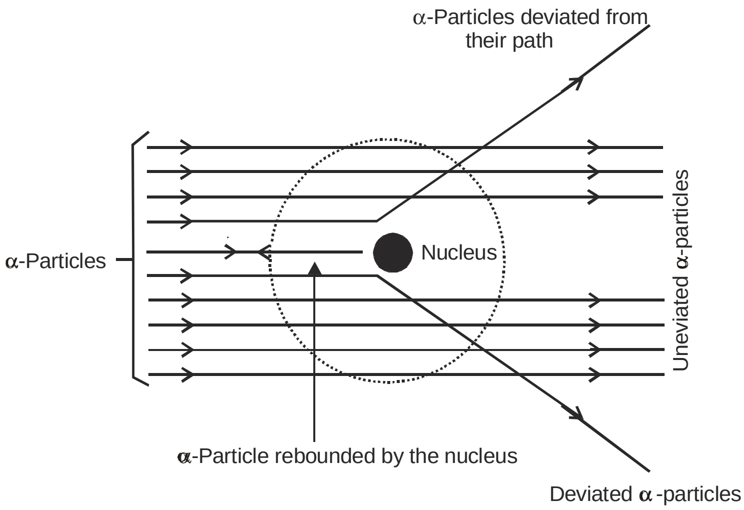
10. NUCLEUS

(i) **Rutherford** discovered the nucleus in an atom by **α-particle scattering experiment**. He showered α-particles, $_2He^4$ (obtained from radium) on a 0.01 mm thin gold film and allowed them to collide with a screen coated with zinc sulphide and placed behind the gold film. He observed fluorescence on the screen.

(a) Most of the α-particle passed through the gold film without deviating from their path.

(b) Some particles got deviated from their path on colliding with the gold film.

(c) A very small number of particles rebounded after colliding with the gold film.



(ii) The following are the inferences derived from the above experiment.

(a) Most of the α-particles pass through the gold foil without deviation in their path, showing that most of the part if an atom is vacant.

(b) Whole of the mass of an atom is confined to its nucleus, which consists of positively charged protons and neutral neutrons. These together are termed as nucleons.

(c) It has been found on the basis of calculation that the radius of the atomic nucleus is 1×10^{-13} to 1×10^{-12} cm or 1×10^{-15} to 1×10^{-14} meter, while radius of an atom is 1×10^{-8} cm.

(d) Magnitude of atomic nucleus = $\frac{\text{Radius of atom}}{\text{Radius of atomic nucleus}}$

(e) Nuclear density Density (D) = $\frac{\text{Mass(M)}}{\text{Volume (V)}}$

Since, the shape of atom is regarded as spherical, therefore, if radius of the nucleus is r, then

$$\text{Volume of nucleus} = \frac{4}{3}\pi r^3$$

11. NUCLEAR CHARGE AND ATOMIC NUMBER

Positive charge on the nucleus of an atom is equal to the atomic number of that atom. A scientist named **Mosley** studied the frequency of X-rays emitted by showering high velocity electrons on a metal and established the following relationship.

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

where ν = frequency of X-rays

z = atomic number or nuclear charge

a and b are constants.

Thus nuclear charge of an atom is equal to the atomic number of that atom. Since an atom is electro neutral, the number of positively charged protons in its nucleus is equal to the negatively charged electrons moving around the nucleus in the atom. Thus

Atomic number = number of protons in the atom or

number of electrons in the atom

12. ATOMIC WEIGHT OR MASS NUMBER

The value of mass number of an atom (in amu) is always a whole number.

Mass number of an atom is the sum of number of protons and number of neutrons present in that atom.

Mass number = Number of protons (Z) + Number of neutrons (n)

= Atomic number + Number of neutrons

For example	$_8O^{16}$	$_7N^{14}$	$_{11}Na^{23}$
Protons	8	7	11
Neutrons	8	7	12
Atomic weight	16	14	23

(a) The protons and neutrons present in the nucleus are known as nucleons.

(b) The weight of electrons is neglected during calculation of the atomic weight, because the mass of an electron is negligible in comparison to that of a proton or a neutron.

(c) In the nucleus of an electro neutral atom, the number of positively charged protons is equal to that of negatively charged electrons.

Particle	$_8O^{16}$	$_7N^{14}$	$_{11}Na^{23}$	$_6C^{12}$	$_9F^{17}$
Protons	8	7	11	6	9
Neutrons	8	7	12	6	8
Atomic weight	16	14	23	12	17
Electrons	8	7	11	6	9

(d) The number of protons present in an atom is called atomic number of that atom.

For example	O	F	Ne
Protons	8	9	10
Atomic number	8	9	10

(e) **Kernel :** The group of all the electrons except those of the outermost energy level, is called that kernel of that atom and the electrons present in the kernel are known as electron of the kernel.

For example, if the electronic configuration of an atom is 2, 6, then the number of kernel electrons is 2.

If the electronic configuration of an atom is 2, 8, 8, then the number of kernel electrons is 10.

If the electronic configuration of an atom is 2, 8, 8, 8, then the number of kernel electrons is 18.

13. IONS

When an atom loses electron, it is converted into a cation, while it is converted into an anion on gaining electron.

(a) Number of electrons in a cation = Number of protons - charge present on the cation

(b) Number of electrons in a anion = Number of protons + Charge present on the anion

For example	Na^+	Mg^{+2}	Al^{+3}
Protons	11	12	13
Electrons	10	10	10
	Cl^-	O^{-2}	F^-
Protons	17	8	9
Electrons	18	10	10

14. ISOTOPES

(a) The atoms of the same element having same atomic number but different atomic weights, are called isotopes.

(b) Isotopes of an element have same number of protons but different number of neutrons in their atoms. Hence their atomic weight are different. For example, oxygen has the following three isotopes.

	$_8O^{16}$	$_8O^{17}$	$_8O^{18}$
Protons	8	8	8
Neutrons	8	9	10
Atomic weights	16	17	18

(c) Hydrogen has the following three isotopes.

	$_1H^1$ (Protium)	$_1D^2$ (Deuterium)	$_1T^3$ (Tritium)
Protons	1	1	1
Neutrons	0	1	2
Atomic weights	1	2	3

(d) Chlorine has the following two isotopes.

$_{17}C^{35}$ and $_{17}Cl^{37}$

15. ISOBARS

(a) Isobars are the atoms of different elements having same atomic weight.

(b) Isobars have different numbers of protons as well as neutrons.

(c) The sum of number of protons and neutrons in isobars is same. For example

Atomic weight of three elements $_{18}Ar^{40}$, $_{19}K^{40}$ and $_{20}Ca^{40}$ is 40.

(i)	Ar^{40}	K^{40}	Ca^{40}
Protons	18	19	20
Neutrons	22	21	20

(ii)	$_{32}Ge^{76}$	$_{34}Se^{76}$
Protons	32	34
Neutrons	44	42

16. ISOTONES

The atoms having same number of neutrons are called isoneutronic or isotones. For example

	$_{14}Si^{30}$	$_{15}P^{31}$	$_{14}S^{32}$
Protons	14	15	16
Neutrons	16	16	16
Atomic weight	30	31	32

17. ISOELECTRONIC

The chemical species in which number of electrons is same are called isoelectronic. For example

(a)	Li^+	Be^{+2}	B^{+3}		
Electrons	2	2	2		
(b)	Na^+	Mg^{+2}	Al^{+3}	F^-	O^{-2}
Electrons	10	10	10	10	10
(c)	K^+	Ca^{+2}	Ar		
Electrons	18	18	18		

18. ATOMIC MODEL

➢ **THOMSON'S MODEL OF AN ATOM**

(i) Atom is a very minute, spherical, electro neutral particle that consists of positively and negatively charged matter.

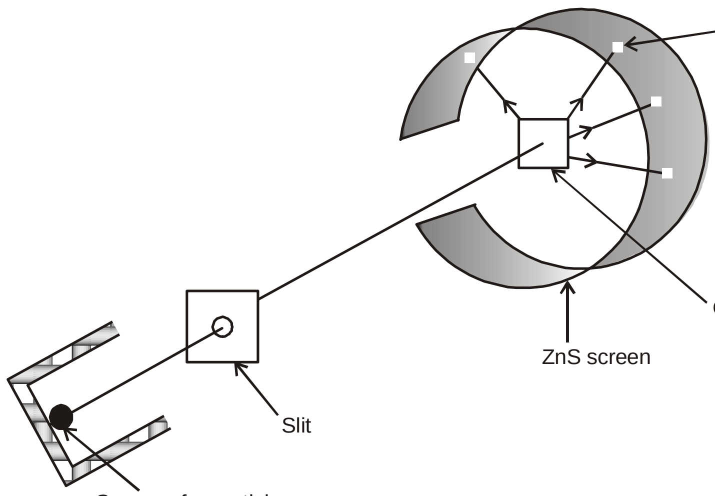
(ii) The positively charged matter is uniformly distributed in the atom and the negatively charged electrons are embedded in it just as the seeds in water melon. Therefore, Thomson model of an atom is also called "water melon model".

(iii) Thomson's model of an atom failed to explain the production of the atomic spectrum. It cannot explain Rutherford's α particle scattering experiment also.

➢ RUTHERFORD'S MODEL OF AN ATOM

Ernest Rutherford in 1911put forward the **"nuclear model"** of atom on the basis of α particle scattering experiment. In this experiment, **Rutherford** showered α–particles (Helium nuclei, He^{+2}) on a thin gold foil and observed that most of the α-particles travelled straight without deviation in the direction of their path, some of them deviate from their path by different angles, while very few get rebounded after colliding with the foil.

Rutherford gave the following nuclear model on the basis of the experiment.



Rutherford's α particle scattering experiment

(i) Atom is a very minute, spherical, electro neutral particle composed of the following two parts :

(a) Positively charged nucleus and **(b)** a vast extranuclear space in which electrons are present.

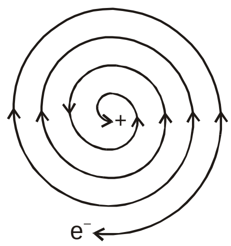
(ii) whole of the positive charge and almost all the mass of atom is confined to a very minute part at the centre of the atom, called the nucleus of the atom. The radius of nucleus is about 10^{-13} to 10^{-12} cm (or 10^{-15} to 10^{-14} meter), while the radius of atom is in the order of 10^{-8} cm.

(iii) The number of electrons in an atom is equal to the number of protons present in the nucleus. That is why an atom is electroneutral.

(iv) This model of an atom is also called "solar model" of "planetary model". This is because, the movement of electrons around the nucleus in this model has been compared to that of planets moving around the sun in the solar system.

- **Demerits of Rutherford's Model of an Atom**

(i) According to Clark Maxwell's theory of electrodynamics, an electrically charged particle in motion continuously emits energy. This results in regular decrease in the energy of that particle. On the basis of this principle, it can be concluded that an electron moving around the nucleus will continuously emit the energy. This will result in decrease in the radius of the electron orbit, due to which the electron would ultimately plunge into the nucleus.



An electron emitting energy and pluging into nucleus

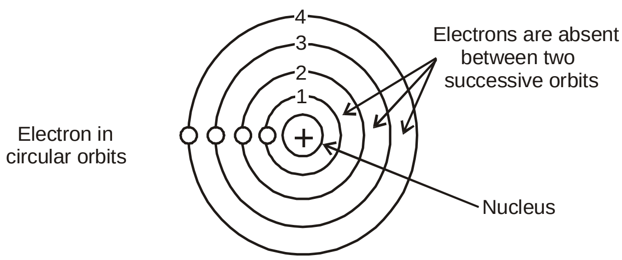
(ii) Plunging of an electron into the nucleus would definitely mean destruction of the atom or end of the existence of the atom. But we know that it never happens. Atom is a stable system. Therefore Rutherford model failed in explaining the stability of an atomic system.

(iii) If an electron moving around the nucleus continuously emits energy, then the atomic spectrum must be continuous, i.e. the spectrum should not have lines of definite frequency. However, the atomic spectrum is actually not continuous and possesses so many lines of definite frequency. Therefore, Rutherford model failed to explain the line spectrum of an atom.

➢ **BOHR'S MODEL OF AN ATOM**

Neil Bohr in 1913 presented a quantum mechanical model of atomic structure.

(i) An electron moves around the nucleus in constant circular orbits.

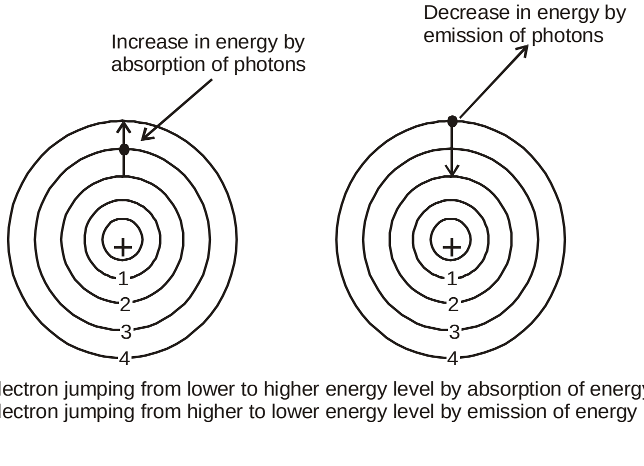


Electrons moving around in circular orbits

(ii) The electrons moving around the nucleus in only those circular orbits for which their angular momentum (mvr) is integral multiple of $\frac{h}{2\pi}$. This is called the condition of quantization. The angular momentum (mvr) of an electron is $\frac{nh}{2\pi}$ where m is the mass of electron. r is radius of its circular orbit, v is the velocity of electron, h is Planck's constant; n is a whole number whose value may be 1, 2, 3, 4 etc, : n is called principal quantum number.

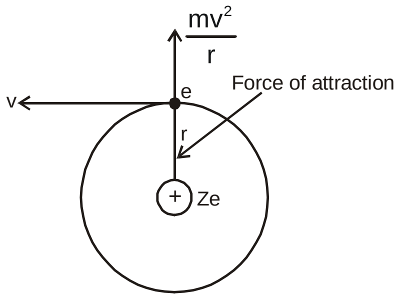
(iii) When energy is provided to an atom, its electrons get excited by absorption of energy and jumps to the orbits of higher energy.

(iv) When an electron in an atom falls from higher energy level to lower energy level, spectral lines are formed.



(a) Electron jumping from lower to higher energy level by absorption of energy
(b) Electron jumping from higher to lower energy level by emission of energy

(v) The force of attraction on electron by the nucleus is equal to the centrifugal force of that electron.



The electron moving in an orbit by various forces

Note : (a) $mvr = \frac{nh}{2\pi}$ ....(1) (b) $\frac{mv^2}{r} = \frac{Ze^2}{r^2}$ ....(2) (c) $E_{n_2} - E_{n_1} = h\nu$...........(3)

19. CALCULATION OF VELOCITY OF THE ELECTRON OF BOHR'S ORBIT

$\frac{mv^2}{r} = \frac{Ze^2}{r^2}$ **.............(1)**

From Bohr's postulate

$mvr = \frac{nh}{2\pi}$ **...............(2)**

Eq. (1) divided by (2)

$\boxed{v = \frac{2\pi Ze^2}{nh}}$ or $v = K\frac{z}{n}$

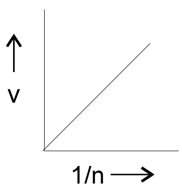
Here π , e and h are constants, therefore

Here $K = \frac{2\pi e^2}{h}$ = 2.188 × 10^8 cm/second

or $\boxed{v = \frac{Z}{n} \times 2.188 \times 10^8 \text{ cm/second}}$

(a) If Z is a constant, then

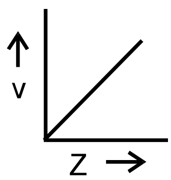
$\boxed{v \propto \frac{1}{n}}$



Therefore, velocity goes on decreasing with increase in the number of orbits.

Thus

$\boxed{\frac{v_1}{v_2} = \frac{n_2}{n_1}}$



(b) If n is a constant, then

$$v \propto Z$$

Therefore, velocity goes on increasing with increase in the atomic number.

$$\frac{v_1}{v_2} = \frac{z_1}{z_2}$$

(c) Time period T = $\frac{2\pi r}{V}$

$$= \frac{2\pi \times n^2h^2}{4\pi^2 mZe^2} \times \frac{nh}{2\pi Ze^2} = \frac{n^3h^3}{4\pi^2 mZ^2e^4}$$

(d) Frequency $\frac{1}{T} = \frac{V}{2\pi r}$

20. RADIUS OF nth BOHR'S ORBIT

According to Bohr's hypothesis

put the value of v in mvr = $\frac{nh}{2\pi}$

$$\boxed{r = \frac{n^2h^2}{4\pi^2 mZe^2}} \quad \text{or} \quad \boxed{r = K\frac{n^2}{z}}$$

In the above expression h, π, m and e, all are constants. therefore

$$\left(K = \frac{h^2}{4\pi^2 me^2} = \text{constant} = 0.529\text{Å}\right)$$

or $\boxed{r = \frac{n^2}{Z} \times 0.529\,\text{Å}}$

Note : (a) 1Å = 10^{-8} cm

(b) 1Å = 10^{-10} m

(c) 1 nm = 10^{-9} m

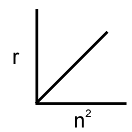
(d) 1 pm (picometer) = 10^{-10}cm

If Z is a constant, then

$$r \propto n^2$$

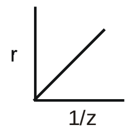
Thus, the radius of atoms goes on increasing as the number (n) of energy levels in the atoms goes on increasing as shown below.

$$\boxed{\frac{r_1}{r_2} = \frac{n_1^2}{n_2^2}}$$



(b) If n is a constant, then

$$\boxed{\frac{r_1}{r_2} = \frac{Z_2}{Z_1}}$$



21. ENERGY OF ELECTRON IN BOHR'S n^{th} ORBIT

(a) The energy of an electron is negative because according to Bohr's hypothesis, the maximum energy of an electron at infinity is zero. Therefore, value of energy should be negative on moving towards lower side from infinity.

(b) The energy of electron at infinity is zero because attractive force between electron and the nucleus is minimum.

(c) Stability would increase as the electron in an atom moves from the infinity distance to a distance r from the nucleus, resulting in the value of the potential energy becoming negative. This is because of the fact that when two opposite charges attract each other, there is a decrease in the potential energy, as attractive forces

$= \frac{Ze^2}{r^2}$

(d) Potential energy of electron is negative while kinetic energy is positive.

(e) Total energy is negative and the negative value shows that attractive forces are working between electron and nucleus. Therefore, work is to be done to remove the electron from this equilibrium state.

(f) Energies are of two types.

➢ **KINETIC ENERGY (E_K)**

This energy is produced due to the velocity of electron. If mass is m, velocity is v and radius is r then

Kinetic energy = $\frac{1}{2}mv^2 = \frac{1}{2}\frac{Ze^2}{r}$

➢ **POTENTIAL ENERGY (E_P)**

This energy is produced due to electrostatic attractive forces between electron and proton, and its value is negative. If atomic number is Z. charge is e and radius is r, then

Potential energy = $\frac{-Ze^2}{r}$

➢ **TOTAL ENERGY (E_T)**

Total energy = Kinetic energy + potential energy $\boxed{E_T = E_K + E_P}$ $\frac{1}{2}mv^2 + \frac{-Ze^2}{r}$

Total energy E = $-\frac{1}{2}\frac{Ze^2}{r}$

Formula

(i) Total energy = – Kinetic energy ($E_T = -E_K$)

(ii) Potential energy = 2 × Total energy) ($E_P = 2E_T$)

➢ **CALCULATION OF ENERGY**

Formula :- ∵ $E = -\frac{1}{2}\frac{Ze^2}{r}$ put the value of r

$E_T = -\frac{Z^2}{n^2} \times \frac{2\pi^2 me^4}{h^2}$ or $= -K\frac{Z^2}{n^2}$

where $K = \frac{2\pi^2 me^4}{h^2}$ = A constant, whose values can be depicted as follows

(a) = 13.60 eV per atom

(b) = 2.179×10^{-11} ergs per atom

(c) = 313.6 kilocalories per mole

(d) = 21.79×10^{-19} joules per atom

(e) = 1312.1 kilojoules per atom

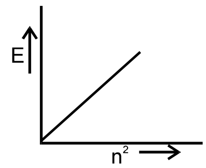
Note : Units – (a) 1 erg = 10^{-7} joule
(b) 1 erg = 6.2419×10^{11} eV
(c) 1 eV = 23.06 kilocalories
(d) 1 eV = 1.602×10^{-12} ergs
(e) 1 joule = 6.2419×10^{18} eV
(f) 1 kilocalorie = 4.184 kilojoule

(i) $E = -\dfrac{Z^2}{n^2} \times 13.6$ eV

If Z is a constant, then $E \propto -\dfrac{1}{n^2}$

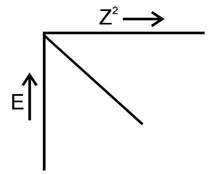
Therefore, the energy of electron goes on increasing as the number of orbits increases.

$$\frac{E_1}{E_2} = \frac{n_2^{\,2}}{n_1^{\,2}}$$



(ii) If n is a constant, then $E \propto -Z^3$

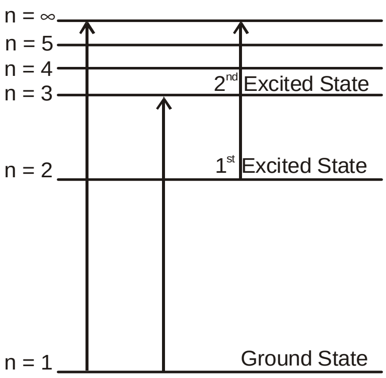
$$\frac{E_1}{E_2} = \frac{Z_1^{\,2}}{Z_2^{\,2}}$$



$E = \dfrac{Z^2}{n^2} \times Rhc$

(1) Kinetic energy = $\dfrac{Z^2}{n^2} \times Rhc$ (2) Potential energy = $2\left(-\dfrac{Z^2}{n^2} \times Rhc\right)$

➢ QUANTIZATION OF ELECTRONIC ENERGY LEVELS



Quantum State	Energy	Separation Energy	Excitation Energy
n = ∞	0	0	13.6
n = 5	$-13.6/5^2 = -0.54$	0.54	13.05
n = 4	$-13.6/4^2 = -0.65$	0.85	12.75
n = 3	$-13.6/3^2 = 1.51$	1.51	12.1
n = 2	$-13.6/2^2 = -3.4$	3.4	10.2
n = 1	–13.6 eV	13.6 eV	0 eV

Electronic energy levels of hydrogen atoms

➢ GROUND STATE

Ans atom in its lowest energy state or initial state is said to be in ground state. This is the most stable of an atom.

- **EXCITED STATE**

The states of higher energy than the ground state are said to be in excited state. For example, the electron of hydrogen atom in ground state is present in n = 1 orbit.

(a) Electron in n = 2 orbit is in first excited state

(b) Electron in n = 3 orbit is in second excited state

(c) Electron in n = 4 orbit is in third excited state

This means that the energy of n + 1 orbit is in first excited state, of n + 2 orbit in second excited state and of n + 3 orbit in third excited state, where n = the energy in ground state.

- **EXCITATION POTENTIAL**

(a) The energy required to excite an electron from ground state to given excited state is known as excitation potential.

(b) Excitation potential has a positive value. For example,

First excitation potential of hydrogen atom = $E_2 - E_1$

Second excitation potential of hydrogen atom = $E_3 - E_1$

Third excitation potential of hydrogen atom = $E_4 - E_1$

- **IONISATION ENERGY OR IONISATION POTENTIAL**

The energy required to remove an electron from the outermost orbit of a gaseous atom in ground state is called ionisation energy or ionisation potential. Its value is positive.

- **SEPARATION ENERGY**

The energy required to separate an electron from given excitation state of an atom is known as separation energy. For example, the first separation energy, i.e. the energy required to remove an electron from the first excited state in hydrogen is + 3.4 eV.

22. SPECTRAL EVIDENCE FOR QUANTIZATION IN BOHR'S THEORY

(a) When an electron undergoes transition from lower to higher orbit, there is absorption of energy and the spectrum obtained thereby is called absorption spectrum.

(b) When an electron undergoes transition from higher to lower orbit, there is emission of energy and the spectrum obtained thereby is called emission spectrum.

(c) A hydrogen atom has only one electron, yet a very large number of lines are visible in its spectrum.

(d) The wave number of spectrum can be find out using the following expression.

$$\bar{\nu} = \frac{1}{\lambda} = R \times Z^2 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

where $\frac{1}{\lambda}$ is wave number

R = Rydberg constant,

n_1 = Number of lower energy level

n_2 = Number of higher energy level

Calculation of formula

$En_1 = -\frac{1}{2}\frac{Ze^2}{r_1}$ (r_1 = radius of the first orbit)

$En_2 = -\frac{1}{2}\frac{Ze^2}{r_2}$ (r_2 = radius of the second orbit)

$En_1 - En_2 = -\frac{1}{2} Ze^2 \left(\frac{1}{r_1} - \frac{1}{r_2}\right)$

According to Bohr hypothesis

$En_1 - En_2 = h\nu$

$En_1 - En_2 = -h\nu$

Therefore, $h\nu = \dfrac{2\pi^2 mZ^2e^4}{h^2}\left(\dfrac{1}{n_1^2} - \dfrac{1}{n_2^2}\right)$

$$\nu = \frac{2\pi^2 mZ^2e^4}{h^3}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

Here $\dfrac{2\pi^2 mZ^2e^4}{ch^3}$ is a constant, because for hydrogen atom Z = 1

Thus $R = \dfrac{2\pi^2 me^4}{ch^3}$ Value of R = 109678 cm^{-1}

If calculation, this value is 109700 cm^{-1}.

Formula = $\dfrac{1}{\lambda} = RZ^2\left(\dfrac{1}{n_1^2} - \dfrac{1}{n_2^2}\right)$

23. EMISSION SPECTRUM AND ABSORPTION SPECTRUM

When a beam of white light passes through a slit or an aperture and then falls on a prism, it gets spilt into many coloured bands. The image of colours so obtained is known as a spectrum. A spectrum is of mainly three types viz.

(i) Emission spectrum

(ii) absorption spectrum

and (iii) molecular spectrum

➢ EMISSION SPECTRUM

When energy is provided to any substance, it starts emitting radiations. These radiations are passed through a spectroscope, they get split up into spectral lines producing emission spectrum. Normally a substance can be excited by any of the following ways.

(a) By heating the substance at high temperature

(b) By passing electric current through a discharge tube having gaseous substance at very low pressure.

(c) By passing electric discharge through a metallic filament.

Emission spectra are of the following two types.

(i) Continuous spectrum and (ii) line spectrum or atomic spectrum

- **Continuous Spectrum**

When sunlight or a glowing heat fluorescent substance like tungsten wire present in an electric bulb, is analysed with the help of a spectroscope, the spectrum obtained on a screen is observed as divided into bands of seven colours, which are in a continuous sequence. Such a spectrum is called a continuous spectrum.

- **Line spectrum or Atomic spectrum**

  When atoms of a substance is excited, it emits radiations. These radiations are analyzed with the help of a spectroscope, then many fine bright lines of specific colours in a sequence are seen in the spectrum, which is not continuous, i.e. there is dark zone in between any two lines. Such a spectrum is called a line spectrum or atomic spectrum. For example, neon single lamp, sodium vapour lamp, mercury vapour lamp, etc. emit light of different colours and they give specific line spectra.

➢ **ABSORPTION SPECTRUM**

When white light emitted by glowing heat fluorescent substance is passed through another substance lime sodium substance. This results in appearance of some black lines in the spectrum. These are present at those places where the line spectrum of the substance i.e. sodium vapour is formed. The spectrum so formed is known as absorption spectrum.

➢ **MOLECULAR SPECTRUM**

Molecular spectrum is given by molecules and it is also known as band spectrum. Three types of energy transitions are found in molecules. These are as follows.

(i) electronic transitions, (ii) vibrational transitions and (iii) rotational transitions.

Therefore, bands are obtained in the spectrum, which are actually groups of lines.

24. HYDROGEN SPECTRUM

Hydrogen atom gives line spectrum. When hydrogen gas is filled at low pressure in a discharge tube and electric discharge is passed through it, a pink coloured is produced in the visible region due to the formation of hydrogen atoms. On studying this light with the help of a spectroscope, series of lines of various wavelengths are obtained in the spectrum.

The frequency of spectral lines in the form of wave number can be calculated with the help of the following expression.

or = R ×

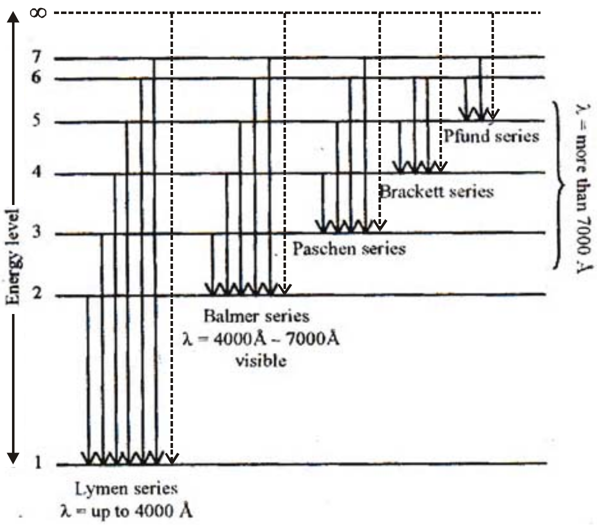
➢ **SERIES OF LINES IN HYDROGEN SPECTRUM**

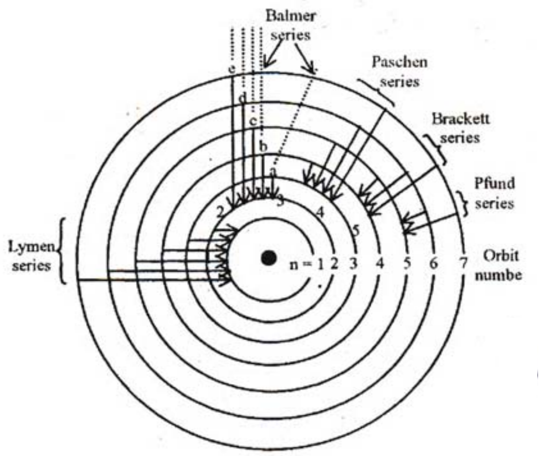
- **Lymen Series**

  When an electron undergoes transition from a higher energy level (n_2), e.g. 2, 3, 4, 5, ... ∞ to ground state or lower energy level, the spectrum is said to belong to Lymen series. For this, n_1 = 1 and n_2 = 2, 3, 4, 5, 6, 7, 8 .... ∞.

- **Balmer Series**

  When an electron undergoes transition from a higher energy level (n_2), e.g. 3, 4, 5, 6, 7, 8 .... ∞ to the second energy level n_1 = 2, the spectrum is said to belong to Balmer series.





Regions of line spectrum of hydrogen atom

Explanation of Balmer series (line spectrum) on the basis of Bohr model

- **Paschen Series**

  When an electron falls from a higher energy level to third orbit (n = 3). It gives a spectrum that is associated with Paschen series. For this $n_1 = 3$ and $n_2 = 4, 5, 6, 7, 8 \ldots\ldots \infty$.

- **Brackett Series**

  When an electron falls from a higher energy level to the fourth orbit (n = 4), the spectrum obtained is associated with Brackett series. For this $n_1 = 4$ and $n_2 = 5, 6, 7, 8 \ldots \infty$.

- **Pfund Series**

  When an electron falls from a higher energy level to the fifth orbit (n = 5), the spectrum obtained is associated with Pfund series. For this $n_1 = 5$ and $n_2 = 6, 7, 8, 9, 10 \ldots\ldots \infty$.

- **Humphry Series**

  When an electron falls from a higher energy level to the sixth orbit (n = 6), Humphrey series of the spectrum is obtained. For this $n_1 = 6$ and $n_2 = 7, 8, 9, 10, 11 \ldots \infty$.

1.	Lymen series	1	2, 3, 4, 5 ...... ∞	Ultraviolet	< 4000Å
2.	Balmer series	2	3, 4, 5, 6 ...... ∞	Visible	4000Å to 7000Å
3.	Paschen series	3	4, 5, 6, 7 ...... ∞	Near infrared	> 7000Å
4.	Brackett series	4	5, 6, 7, 8 ..... ∞	Infrared	> 7000Å
5.	Pfund series	5	6, 7, 8, 9 ..... ∞	Far infrared	> 7000Å
6.	Humphrey series	6	7, 8, 9, 10 .... ∞	Far infrared	> 7000Å

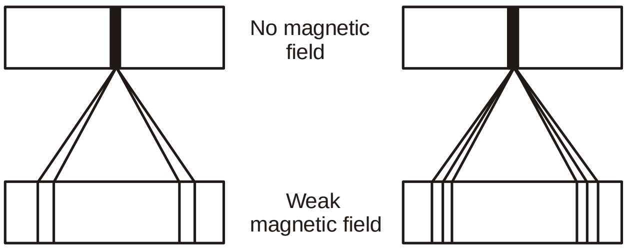
For the given value of n (principal quantum number), the total number of spectral lines can be calculate by the by the expression $\frac{n(n-1)}{2}$.

25. FAILURES OF BOHR'S ATOMIC MODEL

(a) Bohr model cannot explain the elements having more than one electron. Only one-electron species, like hydrogen atom, He^{+1} ion, Li^{+2} ion, Be^{+3} ion, etc. can be explained with the help of Bohr model.

(b) Bohr model can explain only circular orbits in the atom and not the elliptical ones.

(c) Bohr model cannot explain splitting of spectral lines into finer lines in a magnetic field, which is known as **Zeeman effect**.



Representation of Zeeman effect

(d) Bohr model fails to explain the splitting of spectral lines into finer lines in an electric field, which is known as **Stark effect**.

(e) Bohr model fails to explain Hiesenberg uncertainty principle and it cannot be applied for giving any basis to classification of elements and periodicity in their properties.

(f) Bohr model cannot be used for explaining finer structure of spectrum and calculating intensity of spectral lines.

26. SOMMERFELD'S EXPANSION OF BOHR'S MODEL

(a) The aforesaid discovery proved that each principal quantum number (n) is composed of many suborbits.

(b) Sommerfeld suggested that electrons moves around the nucleus not only in circular orbits but also in elliptical orbits.

(c) When an electron travels in an elliptical orbit, its distance (r) from the nucleus and its angle of rotation both will change.

(d) In circular orbit, the distance r remains constant but angle of rotation will change. In elliptical orbit, the nucleus is regarded as situated at the focal point.

(e) A circular orbit is a particular situation of an elliptical orbit, in which the lengths of major axis is equal to that of the minor axis.

(f) In elliptical orbits, the orbital angular momentum is a sum of the following two vector number.

(i) towards the radius, which is called radial component P_r, and

(ii) in the perpendicular direction to radius, which is called azimuthal component P_k.

(g) The above two momenta are separately quantized, i.e. both are multiple of $\frac{h}{2\pi}$.

(h) Sommerfeld suggested that Bohr quantum number n is a sum of two quantum numbers, of which one is radial quantum number n_r and the other is azimuthal quantum number K, i.e. $n = n_r + K$

27. THE WAVE THEORY OF LIGHT

Light, X-rays and radiation produced by a radioactive substance are some of the examples of radiation energy. In 1856 Clark Maxwell showed that energy of radiation is of wave nature, i.e. the energy is emitted in the form of a wave. Therefore, he called the emitted energy as electromagnetic wave or electromagnetic radiation. Since energy is a sort of wave, it is explained as wave motion. Following are the salient features of this wave motion.

(1) Wavelength (λ)

(2) Period (T)

(3) Fequcney (ν)

(4) Amplitude (A)

(5) Wave velocity (c or v)

The aforesaid properties of a wave have the following relationship

$$\nu = \frac{1}{T} \text{ and } c = \frac{\lambda}{T} \text{ or } c = \nu\lambda$$

- **WAVE LENGTH**

The distance between any two successive crests (or troughs) is known as wavelength. This is expressed as λ (Lambda). The range of the wavelength associated with spectrum line is 10^8 to 10^6 cm. Its common units are as follows. Angstrom (Å).

- **FREQUENCY**

The number of vibrations produced in a unit time is called frequency. Here, the time is taken in seconds. The number of wavelengths passing forward in one second from a fixed point is called frequency.

- **VELOCITY OF LIGHT**

The distance traveled by a light wave in a unit time (second) is called the velocity of that wave. It is represented by c and its unit is normally cm/second or m/second. Its value is definite. For example, for a light wave, the velocity $c = 3 \times 10^8$ m/second or 3×10^{10} cm/second.

- **AMPLITUDE**

The maximum deviation of a wave from its equilibrium point is known as its amplitude.

- **WAVE NUMBER**

The reciprocal of wavelength is called wave number. It is represented by $\overline{\upsilon}$.

$$\overline{\upsilon} = \frac{1}{\lambda}$$

Therefore, the unit of wave number is cm^{-1} or m^{-1}

$$\because \; c = \upsilon\lambda \text{ or } \lambda = \frac{c}{\upsilon} \text{ or } \upsilon = \frac{c}{\lambda} \text{ or } \upsilon = c\overline{\upsilon} \text{ or } \overline{\upsilon} = \frac{\upsilon}{c}$$

28. PLANCK'S QUANTUM THEORY

If a substance emits or absorbs energy, it does not do so continuously but does but does it in the form of discrete series of small packet or bundles, called quanta. This energy could be any of the quantum numbers 1, 2, 3, 4, 5 .... n but not in the form of fractional quantum number.

$$\because \nu = \frac{c}{\lambda} \qquad \text{Therefore } E = h \times \frac{c}{\lambda}$$

29. THE DUEL NATURE OF MATTER (THE WAVE NATURE OF ELECTRON)

(a) In 1924, a French physicist, Luis De Breoglie suggested that if the nature of light is both that of a particle and of a wave, then this dual behaviour should be true for the matter also.

(b) The wave nature of light rays and X-rays is provided on the basis of their interference and diffraction and, many facts related to radiations can only be explained when the beam of light rays is regarded as composed of energy corpuscles or photons whose velocity is 3×10^{10} cm/second.

(c) According to ***De Broglie***, the wavelength λ of an electron is inversely proportional to its momentum p.

$\lambda \propto \frac{1}{p}$ or $\lambda = \frac{h}{p}$ Here h = Planck's constant

p = Momentum of electron

$\therefore$ Momentum (p) = Mass (m) × Velocity (c)

Therefore $\lambda = \frac{h}{mc}$ This is called De-Broglie equation

(d) The above relation can be confirmed as follows by using Einstein's equation, Planck's quantum theory and wave theory of light.

But according to Einstein's equation

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

30. BOHR'S THEORY AND DE BROGLIE CONCEPT

(a) According to De Broglie, the nature of an electron moving around the nucleus is like a wave that flows in circular orbits around the nucleus.

(b) If an electron is regarded as a wave, the quantum condition as given by Bohr in his theory is readily fulfilled.

(c) If the radius of a circular orbit is r, its circumference will be $2\pi r$.

(d) We known that according to Bohr theory, $mvr = \frac{nh}{2\pi}$

or $2\pi r = \frac{nh}{mv}$ ($\because$ mv = p momentum) or $2\pi r = \frac{nh}{p}$ ($\because \frac{h}{p} = \lambda$ De Broglie equation)

(e) $\because 2\pi r = \frac{nh}{mv}$

or $mvr = \frac{nh}{2\pi}$ $\because$ mvr = Angular momentum

Thus mvr = Angular momentum, which is a integral multiple of $\frac{h}{2\pi}$

(f) It is clear from the above description that according to De Broglie there is similarity between wave theory and Bohr theory.

31. QUANTUM MECHANICAL THEORY OF ATOM

(a) The dual nature (particle and wave) of electron led to the use of a new system of mechanics called quantum mechanics. This system was first put forward by an Austrian physicist E. Schrodinger and a German physicist W. Heisenberg.

(b) The two fundamental principles of quantum mechanics are given below :

(i) Heisenberg's uncertainty principal and

(ii) Schrodinger's wave equation

➢ **HEISENBERG'S UNCERTAINTY PRINCIPLE**

(a) According to this principle, it is impossible to experimentally determine together both exact position and actual momentum of a minute particle like an electron.

(b) This principal can be depicted mathematically as follows.

$$\Delta x \times \Delta p \geq \frac{h}{4\pi} \qquad \text{or} \qquad \Delta x \times m \times \Delta v \geq \frac{h}{4\pi}$$

Here Δx is uncertainty of position,

Δp is uncertainty of momentum and

h is Planck's constant

➢ **SCHRONDINGER'S WAVE EQUATION**

Schrondinger regarded electron as having wave nature and put forward the following complex differential equation.

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

$$\nabla^2 = \frac{d^2}{dx^2} + \frac{d^2}{dy^2} + \frac{d^2}{dz^2}$$

where m = Mass of electron, h = Planck constant,

E = Total energy of electron, v = Potential energy of electron,

ψ = Wave function, ∇ = Laplacian Operator

32. QUANTUM NUMBERS

(a) The position of any electron in any atom can be ascertained with the help of quantum numbers.

(b) In an atom, the shell consists of sub-shells and the sub-shell consists of orbital can accommodate only two electrons, which are in opposite spins.

➢ **PRINCIPAL QUANTUM NUMBER (n)**

(a) Principal quantum number indicates the shell or energy level or orbit.

(b) An atoms has K, L, M, N, O, P, Q, etc. shells.

(c) Principal quantum number also gives information about the radius of size.

(d) Principal quantum number also gives information about the distance of an electron from the nucleus in an atom

(e) Principal quantum number also given information about the energy of an electron.

(f) Principal quantum number also gives information about the velocity of an electron.

(g) In any orbit, the number of orbitals is given by n^2 and number of electrons is given by $2n^2$. This is called Bohr-Bury rule.

➢ **AZIMUTHAL QUANTUM NUMBER (*l*)**

(a) Azimuthal quantum number gives information that a particular electron belongs to which sub-shell.

(b) In an atom the shells consist or sub-shells, which are indicated as *s*, *p*, *d* and *f*.

(c) Azimuthal quantum number determines the shape of an orbital.

(d) The value of n starts from 1, while that of l starts from 0. Therefore, the maximum value of l is $n - 1$.

(e) The values of n and l can never be equal.

Sub shell	s	p	d	f
l	0	1	2	3

(f) The number of orbitals in any sub orbit is determined by the expression $2l + 1$ and the number of electrons is determined by the expression $2(2l + 1)$.

(g) $l = 0 \rightarrow$ s Sub-shell $\rightarrow$ Spherical

$l = 1 \rightarrow$ p Sub-shell $\rightarrow$ Dumb-bell

$l = 2 \rightarrow$ d Sub-shell $\rightarrow$ Double dumb-bell

$l = 3 \rightarrow$ f Sub-shell $\rightarrow$ Complex

(h) The order of energy of various sub-shells present in any shell is s < p < d < f < g .... and so on.

(i) The value of orbital angular momentum, μ_i, of an electron can be determined with the help of azimuthal quantum number

$$\mu_i = \sqrt{l(l+1)} \times \frac{h}{2\pi}$$

Here ℓ = Azimuthal quantum number and h = Planck's constant

➢ **MAGNETIC QUANTUM NUMBER (m)**

(a) Magnetic quantum number gives information about an orbital. It is depicted by the symbol m.

(b) Magnetic quantum number gives information about orientation of orbitals.

(c) The value of m ranges from $-\ell$ to $+\ell$.

(d) The total number of orbitals present in a sublevel is equal to the total values of magnetic quantum number. This can be find out by the following expression.

$m = 2l = 1$

where m is total value of magnetic quantum number and l is the value of azimuthal quantum number.

(i) For s sub-shell, l = 0. Thus, m = 2 × 0 + 1 = 1 and therefore s sub-shell consists of only one orbital called s orbital.

(ii) For p sub-shell, l = 1. Thus, m = 2 × 1 + 1 = 3 and therefore p sub-shell consists of three orbitals called p_x, p_y and p_z orbitals.

(iii) For d sub-shell, l = 2. Thus, m = 2 × 2 + 1 = 5 and therefore d sub-shell consists of five orbitals called d_{xy}, d_{yz}, d_{z^2}, d_{xz} and $d_{x^2-y^2}$ orbitals.

(i) For s sublevel, l = 0. Thus, for s orbital, the value of m is 0.

s	
	m = 0

(ii) For p sub-level, l = 1. Thus, the values of m for p orbitals are as follows.

p_x/p_y	p_z	p_z/p_y
−1	0	+1

(iii) For d sub-level, l = 2. Thus, the values of m for d orbitals are as follows.

d_{xy}	d_{yz}	d_{z^2}	d_{xz}	$d_{x^2-y^2}$
−2	−1	0	+1	+2

(iv) For f sub-level, l = 3. Thus, the values of m for f orbitals are as follows.

−3	−2	−1	0	+1	+2	+3

(e) The total number of orbitals present in an energy level is determined by the formula n^2 where n is principal quantum number.

➢ **SPIN QUANTUM NUMBER (s)**

(a) Spin quantum number gives information about the spin of an electron.

(b) The value of s is 1/2 which depicts the direction of spin of the electron.

(c) If the electron spins in clockwise direction, s is denoted by $+\frac{1}{2}$ or a sign [↑]. Anticlockwise spin of the electron is denoted by $s = -\frac{1}{2}$ or [↓].

(d) One orbital can accommodate only two electrons, with opposite spins.

(e) One orbital can accommodate only two electrons, with opposite spins.

(f) The angular momentum of an electron is not only due its motion around the nucleus in an energy level but also due to its rotation along its own axis. The angular momentum that arises due to rotation of an electron along its axis, is called spin angular momentum and is depicted by the symbol μs. The value of μs can be found out with the help of the following expression.

$\mu s = \sqrt{s(s+1)} \times \frac{h}{2\pi}$ where s is spin quantum number. In this expression the value of s is always taken as $\frac{1}{2}$ and not $-\frac{1}{2}$.

33. AUFBAU PRINCIPLE

Aufbau is a German word that means to build up. Therefore, electrons are filled up in accordance with this principle.

(a) Pauli's exclusion principle should be followed during filling up of electrons, i.e. no two electrons should have same set of four quantum numbers. This means that maximum number of electrons to be filled in various sub-shells are 2 in s, 6 in p, 10 in d and 14 in f.

(b) Hund's rule should be followed during filling up of electrons i.e. the electrons are to be filled in the degenerate orbitals first in unpaired state.

(c) The electrons are filled in a sub-shell according in $n + l$ rule.

➢ **PAULI'S EXCLUSION PRINCIPLE**

(i) According to Pauli exclusion principle, any two electron cannot have same set of four quantum numbers.

For example :

(a)	$6s^1$	and	$6s^2$
	$n = 6$		$n = 6$
	$l = 0$		$l = 0$
	$m = 0$		$m = 0$
	$s = +\frac{1}{2}$		$s = -\frac{1}{2}$
(b)	$4p^2$	and	$4p^5$
	$n = 4$		$n = 4$
	$l = 1$		$l = 1$
	$m = 0$		$m = 0$
	$s = +\frac{1}{2}$		$s = -\frac{1}{2}$

In the above illustrations, the respectively values of n, l and m are same but that of s is different.

(ii) Pauli exclusion principle can be stated in other words as that **"only two electrons can be accommodated in the same orbital only when their spin quantum number is different".**

(iii) If the third electron enters in an orbital, the set of four quantum numbers becomes same for any two electron.

(iv) According to this rule, for any two electrons, a set of maximum three quantum numbers can be same, but the fourth has to be different. For example, two electrons can have same (n, l and m) or (l, m or s) or (n, m or s)

Example	$1s^1$	and	$1s^2$
	$n = 1$		$n = 1$
	$l = 0$		$l = 0$
	$m = 0$		$m = 0$
	$s = +\frac{1}{2}$		$s = -\frac{1}{2}$

(v) This rule does not apply for hydrogen atom because it contains only one electron.

➢ **HUND'S RULE OF MAXIMUM MULTIPLICITY**

(a) **Degenerate orbitals**

The orbitals having same energy are called degenerate orbitals.

(b) s sub-shell consists of only one orbital. Thus, it cannot have degenerate orbital.

(c) According to Hund's rule, the degenerate orbitals get filled by electrons having parallel spin one by one to give an unpaired state.

(d) According to this rule, the degenerate orbitals are filled in such a way that there is a maximum number of unpaired electrons. For example, C^6 can possibly have the following two configurations of $2s^2$ $2p^2$.

2p

(1) $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}\boxed{}\boxed{}$ (2) $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}\boxed{\uparrow}\boxed{}$

(e) The following two conditions have to be fulfilled for Hund's rule.

(1) The orbitals should be degenerate

(2) The member of electrons and the degenerate orbitals should be more than one

(f) Hund's rule is not applicable for H, He, Li and Be, because electrons in them go to s sub-shell, which does not have any degenerate orbital.

(g) Hund's rule is not applicable for B^5 also, because there is only one electrons in p orbital. Therefore, this rule is applicable from C^6 onwards.

(h) Hund's rule is not important for elements belonging to groups IA, IIA and IIIA.

➢ **n + *l* RULE**

(a) n + l Rule gives information about the energy of various sub-shells.

(b) According to this rule, the sub-shells having higher value of n + l have higher energy.

(c) The sub-shells having lower value of n + l have lower energy.

(d) It two sub-shells have same value of n + l, then that sub-shell will have higher energy which has higher value of n.

Increasing order of energy

$$1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s < 5f < 6d < 7p$$

The maximum number of electrons that can be accommodated in s orbital is 2, that in p orbital is 6, that in d orbital is 10 and that in f orbital is 14.

Exceptions to n + *l* Rule

There are mainly two exceptions of n + l rule.

(a) La^{57} – $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $4s^2$, $3d^{10}$, $4p^6$, $5s^2$, $4d^{10}$, $5p^6$, $6s^2$, $5d^1$

(b) Ac^{89} – $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $4s^2$, $3d^{10}$, $4p^6$, $5s^2$, $4d^{10}$, $5p^6$, $6s^2$, $4f^{14}$, $5d^{10}$, $6s^2$, $7s^2$, $6d^1$

Specific Electronic Configuration

	Element	Atomic number	Expected configuration	Atcual configration
1.	Cr	24	$[Ar]^{18}\ 3d^4\ 4s^2$	$[Ar]^{18}\ 3d^5\ 4s^1$
2.	Cu	29	$[Ar]^{18}\ 3d^9\ 4s^2$	$[Ar]^{18}\ 3d^{10}\ 4s^1$
3.	Mo	42	$[Kr]^{36}\ 4d^4\ 5s^2$	$[Kr]^{36}\ 4d^5\ 5s^1$
4.	Pd	46	$[Kr]^{36}\ 4d^8\ 5s^2$	$[Kr]^{36}\ 4d^{10}\ 5s^0$
5.	Ag	47	$[Kr]^{36}\ 4d^9\ 5s^2$	$[Kr]^{36}\ 4d^{10}\ 5s^1$
6.	W	74	$[Xe]^{54}\ 4f^{14}\ 5d^4\ 6s^2$	$[Xe]^{54}\ 4f^{14}\ 5d^5\ 6s^1$
7.	Pt	78	$[Xe]^{54}\ 4f^{14}\ 5d^8\ 6s^2$	$[Xe]^{54}\ 4f^{14}\ 5d^9\ 6s^1$
8.	Au	79	$[Xe]^{54}\ 4f^{14}\ 5d^9\ 6s^2$	$[Xe]^{54}\ 4f^{14}\ 5d^{10}\ 6s^1$

Due to greater stability of half-filled and fully-filled orbitals, the configurations $d^5\ ns^1$ and $d^{10}\ ns^1$ are written in place of $d^4\ ns^2$ and $d^9\ ns^2$ respectively.

➢ **STABILITY OF HALF-FILLED AND FULLY-FILLED ORBITALS**

The stability of half-filled orbitals (p^3, d^5 and f^7) and fully-filled orbitals (p^6, d^{10} and f^{14}) is higher than that in other states. This is due the following reasons.

(a) When a sub-shell is half–filled or fully-filled, it means that the distribution of electrons is symmetrical in the orbitals of equal energy. Unsymmetrical distribution of electrons results in lower stability.

(b) The electrons present in orbtials of equal energy in an atom can interchange their position, in this process energy is released, resulting stable system. The possibility of interchange of positions is highest in half filled and fully-filled states. This provides greater stability to the system.

(c) The exchange energy for half-filled and fully-filled orbitals is maximum. As the number of electrons increases, electron start pairing resulting in spin coupling. The energy liberated in the process of coupling is called coupling energy.

(d) The spin of electrons in a fully-filled orbital are opposite to each other or antiparallel. The energy of the system decreases due to neutralization of opposite spins. So fully-filled orbitals are more stable.

34. MODE OF FILLING UP TO ELECTRONS

Writing the configuration of ions

First of all, the configuration of the atom is written. Then, appropriate number of electrons are deducted from the outermost shell for the configuration of the cation. Similarly, appropriate number of electrons are added to the outermost shell for the configuration of the anion.

35. DIFFERENCE BETWEEN ORBIT AND ORBITAL

S.No.	Orbit	Orbital
1.	It is represented by n.	It is represented by m.
2.	It has maximum electron capacity of $2n^2$	It has maximum electron capacity of 2 in accordance with Pauli's principle.
3.	It is bigger in size	It is smaller in size.
4.	Orbit consist of suborbits	Sub-orbit consists of orbitals
5.	The path of an electron around the nucleus is called an orbit	The space around the nucleus where probability of finding an electron is maximum, is called an orbital.

SOLVED EXAMPLES

Ex.1 Atomic radius is of the order of 10^{-8} cm and nuclear radius is of the order of 10^{-13} cm. Calculate what fraction of atom is occupied by nucleus ?

Sol. Volume of nucleus $= \frac{4}{3}\pi r^3 = \frac{4}{3}\pi (10^{-13})^3\ cm^3$

Volume of atom $= \frac{4}{3}\pi (10^{-8})^3\ cm^3$

$$\frac{V_N}{V_{Atom}} = \frac{10^{-39}}{10^{-24}} = 10^{-15}$$

$V_{Nucleus} = \mathbf{10^{-15} \times V_{Atom}}$

Ex.2 The atomic masses of two isotopes of O are 15.9936 and 17.0036. Calculate in each atom

[1] No. of neutrons [2] No. of protons [3] No. of electrons [4] Mass no. **Ans. [1]**

Sol.

	I isotope of O	**II** isotope of O
Atomic masses are	15.9936	17.0036
∴ Mass no. are	**16**	**17** (Integer values)
∴ No of neutrons	= 16 – 8 = **8**	17 – 8 = **9**
and no. of electrons	**= 8**	**= 8**

Mass no. – At no. = No. of neutrons

Ex.3 The mass charge ratio for A^+ ion is 1.97×10^{-7} kg C^{-1}. Calculate the mass of A atom.

Sol. Given $\frac{m}{e} = 1.97 \times 10^{-7}$

(since $e = 1.602 \times 10^{-19}$ C) $\therefore m = 1.97 \times 10^{-7} \times 1.602 \times 10^{-19}$ kg

$\mathbf{m = 3.16 \times 10^{-26}\ kg}$

Ex.4 AIR service on Vividh Bharati is transmitted on 219 m band. What is its transmission frequency in Hertz ?

Sol. Given $\lambda = 219$ m Thus, $\nu = \frac{c}{\lambda}$

or $\nu = \frac{3.0\times10^{8}}{219} = \mathbf{1.37 \times 10^{6}\ Hz}$

Ex.5 Write down the numerical value of h and its unit

Sol. $h = 6.625 \times 10^{-27}$ erg sec $= 6.625 \times 10^{-34}$ joule sec

The unit of h = joule sec. or erg sec. $\left(\begin{array}{l} \because h\nu = E \\ \therefore h = \frac{E}{\nu} = \frac{erg}{sec^{-1}} \end{array}\right)$

Ex.6 Calculate the number of proton emitted in 10 hours by a 60 W sodium lamp (λ or photon = 5893 Å)

Sol. Energy emitted by sodium lamp in one sec. = Watt × sec = 60 × 1 J

Energy of photon emitted $= \frac{hc}{\lambda} = \frac{6.626\times10^{-34}\times3\times10^{8}}{5893\times10^{-10}} = 3.37 \times 10^{-19}$ J

$\therefore$ No of photons emitted per sec. $= \frac{60}{3.37\times10^{-19}}$

$\therefore$ No. of photons emitted in 10 hours $= 17.8 \times 10^{19} \times 10 \times 60 \times 60 = \mathbf{6.41 \times 10^{24}}$

Ex.7 Calculate the longest wavelength which can remove the electron from I Bohr's orbit. Given E_1 = 13.6 eV.

Sol. The photon capable of removing electron from I Bohr's orbit must possess energy

$$= 13.6 \text{ eV}$$

$$= 13.6 \times 1.602 \times 10^{-19} \text{ J} = 21.787 \times 10^{-19} \text{ J}$$

$$\because \quad E = \frac{hc}{\lambda} \; ; \; 21.787 \times 10^{-19} = \frac{6.625 \times 10^{-34} \times 3.0 \times 10^{8}}{\lambda}$$

$$\therefore \quad \lambda = 912.24 \times 10^{-10} \text{ m} = \mathbf{912.24 \ \text{Å}}$$

Ex.8 Calculate momentum of radiations of wavelength 0.33 nm.

Sol. We have $\lambda = \frac{h}{mu}$ $\quad \therefore \; mu = \frac{h}{\lambda}$

$$= \frac{6.625 \times 10^{-34}}{0.33 \times 10^{-9}} = 2.01 \times 10^{-24} \text{ kgmsec}^{-1}$$

Ex.9 The ionization energy of He^+ is 19.6×10^{-18} J $atom^{-1}$. The energy of the first stationary state of Li^{+2} will be :

[1] 84.2×10^{-18} J/atom [2] 44.10×10^{-18} J/atom [3] 63.2×10^{-18} J/atom [4] 21.2×10^{-18} J/atom **Ans. [2]**

Sol. E_1 for Li^{+2} = E_1 for H × Z^2 = E_1 for H × 9

E_1 for He^+ = E_1 for H × Z^2_{He} = E_1 for H × 4

or E_1 for $Li^{+2} = \frac{9}{4}$ E_1 for He^+ = $19.6 \times 10^{-18} \times \frac{9}{4}$ **= $\mathbf{44.10 \times 10^{-18}}$ J/atom**

Ex.10 The ionization energy of H-atom is 13.6 eV. The ionization energy of Li^{+2} ion will be

[1] 54.4 eV [2] 122.4 eV [3] 13.6 eV [4] 27.2 eV **Ans. [2]**

Sol. E_1 for Li^{+2} = E_1 for H × Z^2 [for Li, Z = 3]

= 13.6 × 9 = **122.4 eV**

Ex.11 Which of the following set of quantum numbers are not permitted

(a) n = 3, *l* = 2, m = – 2, s = +1/2 (b) n = 3, *l* = 2, m = – 1, s = 0

(c) n = 2, *l* = 2, m = +1, s = – 1/2 (d) n = 2, *l* = 2, m = +1, s = – 1/2

Sol. (a) This set of quantum number is permitted.

(b) This set of quantum number is not permitted as value of 's' cannot be zero.

(c) This set of quantum number is not permitted as the value of '*l*' cannot be equal to 'n'.

(d) This set of quantum number is not permitted as the value of 'm' cannot be greater than '*l*'.

Ex.12 Find out the energy of H atom in first excitation state. The value of permittivity factor $4\pi\varepsilon_0 = 1.11264 \times 10^{-10} \ C^2 N^{-1} m^{-2}$.

Sol. In M.K.S. system

$$E_n = -\frac{2\pi^2 Z^2 me^2}{(4\pi\varepsilon_0)^2 n^2 h^2} = \frac{2 \times (3.14)^2 \times (1)^2 \times 9.108 \times 10^{-31} \times (1.602 \times 10^{-19})^4}{(1.11264 \times 10^{-10})^2 \times (2)^2 \times (6.625 \times 10^{-34})^2} = \mathbf{5.443 \times 10^{-19} \ joule}$$

Ex.13 The shortest wave length in H spectrum of Lymen series when R_H = 109678 cm^{-1} is

[1] 1002.7 Å [2] 1215.67 Å [3] 1127.30 Å [4] 911.7 Å **Ans. [4]**

Sol. For Lymen series n_1 = 1

For shortest 'l' or Lymen series the energy difference in two levels showing transition should be maximum

(i.e. $n_2 = \infty$) $\quad \frac{1}{\lambda} = R_H \left[\frac{1}{1^2} - \frac{1}{\infty^2}\right]$

= 109678 = 911.7 × 10^{-8} = **911.7 Å**

Ex.14 Electromagnetic radiations of wavelength 242 nm is just sufficient to ionise sodium atom. Calculate the ionisation energy of sodium in kJ mol^{-1}.

Sol. Energy associated with a photon of 242 nm $= \dfrac{6.625\times10^{-34}\times3.0\times10^{8}}{242\times10^{-9}}$ = 8.21 × 10^{-19} joule

$\because$ 1 atom of Na for ionisation requires = 8.21 × 10^{-19} J

$\therefore$ 6.023 × 10^{23} atoms of Na for ionisation requires = 8.21 × 10^{-19} × 6.023 × 10^{23}

= 49.45 × 10^4 J = **494.5 kJ mol^{-1}**

Ex.15 How many electrons in a given atom can have the following quantum numbers

(a) n = 4, l = 2, m = 0 (b) n = 3

(c) n = 2, l = 1, m = – 1, s = + 1/2 (d) n = 4, l = 1

Sol. (a) l = 2 means d-subshell and m = 0 refer to dz^2 orbital $\therefore$ Number of electrons are **2**.

(b) For n = 3, l = 0, 1, 2

l = 0	m = 0	2 electrons
l= 1	m = –1	6 electrons
l = 2	m = –2, –1, 0, +1, +2	10 electrons
	Total electrons	18 electrons

Alternatively, number of electrons for any energy level is given by

$2n^2$ i.e. 2 × 3^2 = 18 electrons

(c) l = 1 refers to p-subshell, m = – 1 refers to p_x or p_y orbital whereas, s = + 1/2 indicate for only 1 electron.

(d) l = 1 refers to p-subshell which has three orbitals (p_x, p_y and p_z) each having two electrons. Therefore , total number of electrons are **6**.

Ex.16 What transition in the hydrogen spectrum would have the same wavelength as the Balmer transition n = 4 to n = 2 of He^+ spectrum ?

Sol. For He^+, $\dfrac{1}{\lambda} = R_H Z^2 \left[\dfrac{1}{2^2}-\dfrac{1}{4^2}\right]$

For H, $\dfrac{1}{\lambda} = R_H \left[\dfrac{1}{n_1^2}-\dfrac{1}{n_2^2}\right]$

Since λ is same

$\therefore$ $Z^2\left[\dfrac{1}{2^2}-\dfrac{1}{4^2}\right] = \left[\dfrac{1}{n_1^2}-\dfrac{1}{n_2^2}\right]$ $\because$ Z = 2

$\therefore$ $\left[\dfrac{1}{1^2}-\dfrac{1}{2^2}\right] = \left[\dfrac{1}{n_1^2}-\dfrac{1}{n_2^2}\right]$ $\therefore$ n_1 = **1** and n_2 = **2**

Ex.17 On the basis of Heisenberg's uncertainty principle, show that the electron cannot exist within the nucleus.

Sol. Radius of the nucleus is of the order of 10^{-13} cm and thus uncertainty in position of electron, i.e., (Δx), if it is within the nucleus will be 10^{-13} cm.

Now, $\Delta x.\Delta u \geq \dfrac{h}{4\pi m}$

$\therefore$ $\Delta u = \dfrac{6.626\times10^{-27}}{4\times3.14\times9.108\times10^{-28}\times10^{-13}}$ = 5.79 × 10^{12} cm/sec i.e., order of velocity of electron will be 100 times greater than the velocity of light which is impossible. Thus, possibility of electron to exist is nucleus is zero.

Ex.18 Oxygen consists of isotopes of O^{16}, O^{17} and O^{18} and carbon consists of isotopes of C^{12} and C^{13}. How many types of CO_2 molecule can be formed ? Also report their mol, wt.

Sol. Total no. of molecules of CO_2 = 12

[1] $C^{12}O^{16}O^{16}$	Mol. wt. = 44
[2] $C^{12}O^{17}O^{17}$	= 46
[3] $C^{12}O^{18}O^{18}$	= 48
[4] $C^{12}O^{16}O^{17}$	= 45
[5] $C^{12}O^{16}O^{18}$	= 46
[6] $C^{12}O^{17}O^{18}$	= 47

Similarly six molecules with C^{13} isotopes.

Ex.19 Naturally occurring boron consists of two isotopes whose atomic weights are 10.01 and 11.01. The atomic weight of natural boron is 10.81. Calculate the percentage of each isotope in natural boron

Sol. Let the percentage of isotope with atomic wt. 10.01 = x

$\therefore$ Percentage of isotope with atomic wt. 11.01 = 100 – x

Average atomic wt. = $\dfrac{m_1x_1 + m_2X_2}{x_1 + x_2}$

or Average atomic wt. = $\dfrac{x \times 10.01 + (100 - x) \times 11.01}{100}$

10.81 = $\dfrac{x \times 10.01 + (100 - x) \times 11.01}{100}$ = 20

$\therefore$ % of isotope with atomic wt. 10.01 = **20**

% of isotope with atomic wt. 11.01 = 100 – x = **80**

Ex.20 Prove that $u_n = \sqrt{\left(\dfrac{Ze^2}{mr_n}\right)}$ where u is velocity of electron in a one electron atom of at. no. Z at a distance r_n from the nucleus, m and e are mass and charge of electron.

Sol. Kinetic energy of electron = $\dfrac{1}{2}mu^2$

Also from Bohr's concept K.E. = $\dfrac{1}{2}\dfrac{Ze^2}{r_n}$ $\qquad \therefore \dfrac{1}{2}mu^2 = \dfrac{1}{2}\dfrac{Ze^2}{r_n}$

$$\nu = \sqrt{\left(\dfrac{Ze^2}{mr_n}\right)}$$

EXERCISE - 1

Discovery and Properties of Anode, Cathode Rays, Neutron and Nuclear Structure

Q.1 The size of nucleus is of the order of

[1] 10^{-12} m [2] 10^{-8} m [3] 10^{-15} m [4] 10^{-10} m

Q.2 The ratio of charge and mass would be greater for

[1] Proton [2] Electron [3] Neutron [4] Alpha

Q.3 The radius of an atom is of the order of

[1] 10^{-10} cm [2] 10^{-13} cm [3] 10^{-15} cm [4] 10^{-8} cm

Q.4 Which one of the following pairs is not correctly matched

[1] Rutherford-Proton [2] J.J.Thomson-Electron [3] J.H.Chadwick-Neutron [4] Bohr-Isotope

Q.5 Proton was discovered by

[1] Chadwick [2] Thomson [3] Goldstein [4] Bohr

Q.6 The average distance of an electron in an atom from its nucleus is of the order of

[1] 10^{6} m [2] 10^{-6} m [3] 10^{-10} m [4] 10^{-15} m

Q.7 The ratio of specific charge of a proton and an α-particle is

[1] 2 : 1 [2] 1 : 2 [3] 1 : 4 [4] 1 : 1

Atomic Number, Mass Number, Atomic Species

Q.8 Which of the following are isoelectronic with one another

[1] Na^+ and Ne [2] K^+ and O [3] Ne and O [4] Na^+ and K^+

Q.9 The number of electrons in one molecule of CO_2 are

[1] 22 [2] 44 [3] 66 [4] 88

Q.10 An atom has 26 electrons and its atomic weight is 56. The number of neutrons in the nucleus of the atom will be

[1] 26 [2] 30 [3] 36 [4] 56

Q.11 An atom has the electronic configuration of $1s^2$, $2s^2\ 2p^6$,$3s^2\ 3p^6\ 3d^{10}$, $4s^2\ 4p^5$. Its atomic weight is 80. Its atomic number and the number of neutrons in its nucleus shall be

[1] 35 and 45 [2] 45 and 35 [3] 40 and 40 [4] 30 and 50

Q.12 Which of the following particles has more electrons than neutrons

[1] C [2] F^- [3] O^{-2} [4] Al^{+3}

Q.13 Na^+ ion is isoelectronic with

[1] Li^+ [2] Mg^{+2} [3] Ca^{+2} [4] Ba^{+2}

Q.14 Which of the following oxides of nitrogen is isoelectronic with CO_2

[1] NO_2 [2] N_2O [3] NO [4] N_2O_2

Q.15 The hydride ions (H^-) are isoelectronic with

[1] Li [2] He^+ [3] He [4] Be

Q.16 Which of the following are isoelectronic species I = CH_3^+, II = NH_2, III = NH_4^+, IV = NH_3

[1] I, II, III [2] II, III, IV [3] I, II, IV [4] I and II

Q.17 Number of unpaired electrons in inert gas is

[1] 0 [2] 8 [3] 4 [4] 18

Q.18 Which one of the following is not isoelectronic with O^{2-}

[1] N^{3-} [2] F^{-} [3] Tl^{+} [4] Na^{+}

Q.19 If molecular mass and atomic mass of sulphur are 256 and 32 respectively, its atomicity is

[1] 2 [2] 8 [3] 4 [4] 16

Q.20 The number of neutron in tritium is

[1] 1 [2] 2 [3] 3 [4] 0

Q.21 The atomic mass of an element is 12.00719 amu. If there are 6 neutrons in the nucleus of the atom of the element, the binding energy per nucleon of the nucleus will be

[1] 7.64 MeV [2] 76.4 MeV [3] 764 MeV [4] 0.764 MeV

(e = 0.00055 amu, p = 1.00814 amu, n = 1.00893 amu)

Atomic models and Planck's quantum theory

Q.22 Bohr's model can explain

[1] The spectrum of hydrogen atom only

[2] Spectrum of atom or ion containing one electron only

[3] The spectrum of hydrogen molecule

[4] The solar spectrum

Q.23 When an electron jumps from L to K shell

[1] Energy is absorbed

[2] Energy is released

[3] Energy is sometimes absorbed and sometimes released

[4] Energy is neither absorbed nor released

Q.24 When an electron jumps from lower to higher orbit, its energy

[1] Increases [2] Decreases [3] Remains the same [4] None of these

Q.25 The energy of second Bohr orbit of the hydrogen atom is – 328 kJ mol^{-1}, hence the energy of fourth Bohr orbit would be

[1] – 41 kJ mol^{-1} [2] – 1312 kJ mol^{-1} [3] –164 kJ mol^{-1} [4] – 82 kJ mol^{-1}

Q.26 If wavelength of photon is 2.2×10^{-11} m, h = 6.6×10^{-34} J-sec, then momentum of photon is

[1] 3×10^{-23} kg ms^{-1} [2] 3.33×10^{22} kg ms^{-1} [3] 1.452×10^{-44} kg ms^{-1} [4] 6.89×10^{43} kg ms^{-1}

Q.27 If electron falls from n = 3 to n = 2, then emitted energy is

[1] 10.2 eV [2] 12.09 eV [3] 1.9 eV [4] 0.65 eV

Q.28 The specific charge of proton is 9.6×10^{6} C kg^{-1} then for an α-particle it will be

[1] 38.4×10^{7} C kg^{-1} [2] 19.2×10^{7} C kg^{-1} [3] 2.4×10^{7} C kg^{-1} [4] 4.8×10^{7} C kg^{-1}

Q.29 Radius of the first Bohr's orbit of hydrogen atom is

[1] 1.06 Å [2] 0.22 Å [3] 0.28 Å [4] 0.53 Å

Q.30 Energy of electron of hydrogen atom in second Bohr orbit is

[1] -5.44×10^{-19} J [2] -5.44×10^{-19} kJ [3] -5.44×10^{-19} cal [4] -5.44×10^{-19} eV

Q.31 The radius of first Bohr's orbit for hydrogen is 0.53 Å. The radius of third Bohr's orbit would be

[1] 0.79 Å [2] 1.59 Å [3] 3.18 Å [4] 4.77 Å

Q.32 The value of the energy for the first excited state of hydrogen atom will be

[1] –13.6 eV [2] – 3.40 eV [3] – 1.51 eV [4] – 0.85 eV

Q.33 Which of the following is not true in Rutherford's nuclear model of atom

[1] Protons and neutrons are present inside nucleus

[2] Volume of nucleus is very small as compared to volume of atom

[3] The number of protons and neutrons are always equal

[4] The number of electrons and protons are always equal

Q.34 Ratio of radii of second and first Bohr orbits of H atom

[1] 2 [2] 4 [3] 3 [4] 5

Q.35 The energy of the electron in the first orbit of He^+ is – 871.6 × $10^{–20}$ J. The energy of the electron in the first orbit of hydrogen would be

[1] – 871.6 × $10^{–20}$ J [2] – 435.8 × $10^{–20}$ J [3] – 217.9 × $10^{–20}$ J [4] – 108.9 × $10^{–20}$ J

Q.36 The Bohr orbit radius for the hydrogen atom (n = 1) is approximately 0.530 Å. The radius for the first excited state (n = 2) orbit is

[1] 0.13 Å [2] 1.06 Å [3] 4.77 Å [4] 2.12 Å

Q.37 In case of hydrogen atom when electron falls from higher level to M shell, the corresponding spectral line will form the part of

[1] Balmer series [2] Lyman series [3] Paschen series [4] Pfund series

Q.38 The velocity of an electron on 1, 2 and 3rd orbit of hydrogen atom is given respectively as V_1, V_2, V_3. They maintain the order

[1] $V_1 > V_2 > V_3$ [2] $V_1 < V_2 < V_3$ [3] $V_1 = V_2 = V_3$ [4] $V_1 > V_2 = V_3$

Q.39 I.P. of hydrogen atom is equal to 13.6 eV. What is the energy required for the process :

$$He^+ + \text{energy} \longrightarrow He^{+2} + e$$

[1] 2 × 13.6 eV [2] 1 × 13.6 eV [3] 4 × 13.6 eV [4] None of these

Q.40 E_1, E_2, E_3 and E_4 respectively are the energies of the first line of Lyman, Balmer, Paschen and Brackett series. The decreasing order of their energy is

[1] E_1, E_2, E_3, E_4 [2] E_4, E_3, E_2, E_1 [3] E_1, E_3, E_2, E_4 [4] E_2, E_1, E_4, E_3

Q.41 If x is the velocity of an electron in first Bhor's orbit. What would be the velocity of the electron in third Bhor's orbit ?

[1] x/9 [2] x/3 [3] 3x [4] 9x

Q.42 What should be the maximum number of lines obtained in the spectrum, if total number of energy levels are four ?

[1] 4 [2] 6 [3] 2 [4] 3

Q.43 Which of the following transitions in hydrogen atom should emit a radiation of highest frequency ?

[1] n = 5 to n = 2 Balmer [2] n = 3 to n = 2 Balmer

[3] n = 4 to n = 2 Balmer [4] n = 3 to n = 1 Lymen

Q.44 Which of the following should be the correct value of the wave number of first line in Balmer Series of hydrogen atom ?

[1] 5R/36 [2] –5R/36 [3] R/9 [4] –R/9

Q.45 What should be the value of wave number of the radiation produced when electron jumps from the second orbit to first orbit in hydrogen atom ?

[1] 82276 $cm^{–1}$ [2] 3200 $cm^{–1}$ [3] 52276 $cm^{–1}$ [4] 83376 $cm^{–1}$

Q.46 What should be the quantum number of the highest energy state when an electron falls from the highest energy state to Lymen Series and for this transition the wave number will be = 97492.2 cm^{-1}?

[1] 2 [2] 3 [3] 4 [4] 5

Dual nature of electron, Uncertainty principle and Schrodinger wave equation

Q.47 What is the de-broglie wavelength associated with the hydrogen electron in its third orbit

[1] 9.96×10^{-10} cm [2] 9.96×10^{-8} cm [3] 9.96×10^{4} cm [4] 9.96×10^{8} cm

Q.48 Which one is not the correct relation in the following

[1] $h = E/\upsilon$ [2] $E = mc^2$ [3] $\Delta x \times \Delta p = h/4\pi$ [4] $\lambda = h/m\upsilon$

Q.49 The uncertainty in momentum of an electron is 1×10^{-5} kg – m/s. The uncertainty in its position will be ($h = 6.62 \times 10^{-34}$ kg –m^2/s)

[1] 1.05×10^{-28} m [2] 1.05×10^{-26} m [3] 5.27×10^{-30} m [4] 5.25×10^{-28} m

Q.50 The uncertainty in the position of a moving bullet of mass 10 gm is 10^{-5} m. Calculate the uncertainty in its velocity

[1] 5.2×10^{-28} m/sec [2] 3.0×10^{-28} m/sec [3] 5.2×10^{-22} m/sec [4] 3×10^{-22} m/sec

Q.51 The ratio of energy of photons having wavelength 2000 Å and 4000Å respectively would be

[1] 1/4 [2] 4 [3] 1/2 [4] 2

Q.52 Helium atom is twice as heavy as hydrogen molecule. At 298 K the average kinetic energy of helium atom would be

[1] Equal to that of hydrogen molecule [2] Four times that of hydrogen molecule

[3] Twice that of hydrogen molecule [4] Thrice that of hydrogen molecule

Q.53 K.E. of the electron is 4.55×10^{-25} J. Its de Broglie wave length is

[1] 4700 Å [2] 8300 Å [3] 7200 Å [4] 7400 Å

Quantum Number, Electronic Configuration and Shape of Orbitals

Q.54 Nitrogen has the electronic configuration $1s^2$, $2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$ and not $1s^2$, $2s^2\ 2p_x^2\ 2p_y^1\ 2p_z^0$ which is determined by

[1] Aufbau's principle [2] Pauli's exclusion principle

[3] Hund's rule [4] uncertainty principle

Q.55 2p orbitals have

[1] n = 1, l = 2 [2] n = 1, l = 0 [3] n = 2, l = 1 [4] n = 2, l = 0

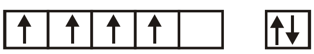
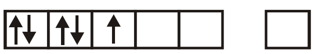
Q.56 An element has the electronic configuration $1s^2$, $2s^2\ 2p^6$, $3s^2\ 3p^2$. Its valency electrons are

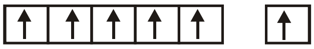
[1] 6 [2] 2 [3] 3 [4] 4

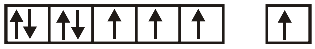
Q.57 Which of the following sets of quantum numbers represent an impossible arrangement

	n	l	m	m_s		n	l	m	m_s
[1]	3	2	–2	$(+)\frac{1}{2}$	[2]	4	0	0	$(-)\frac{1}{2}$
[3]	3	2	–3	$(+)\frac{1}{2}$	[4]	5	3	0	$(-)\frac{1}{2}$

Q.58 Which one is the correct outer configuration of chromium

[1]  [2] 

[3] 

[4] 

Q.59 The electronic configuration of calcium ion (Ca^{2+}) is

[1] $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6, 4s^2$ [2] $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6, 4s^1$

[3] $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ 3d^2$ [4] $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6, 4s^0$

Q.60 The structure of external most shell of inert gases is

[1] $s^2\ p^3$ [2] s^2p^6 [3] s^1p^2 [4] $d^{10}s^2$

Q.61 The number of orbitals in the fourth principal quantum number will be

[1] 4 [2] 8 [3] 12 [4] 16

Q.62 Which set of quantum numbers are not possible from the following

[1] n = 3, l = 2, m = 0, s = $-\frac{1}{2}$ [2] n = 3, l = 2, m = –2, s = $-\frac{1}{2}$

[3] n = 3, l = 3, m = –3, s = $-\frac{1}{2}$ [4] n = 3, l = 0, m = 0, s = $-\frac{1}{2}$

Q.63 The four quantum number for the valence shell electron or last electron of sodium (Z = 11) is

[1] n = 2, l = 1, m = –1, s = $-\frac{1}{2}$ [2] n = 3, l = 0, m = 0, s = $+\frac{1}{2}$

[3] n = 3, l = 2, m = –2, s = $-\frac{1}{2}$ [4] n = 3, l = 2, m = 2, s = $+\frac{1}{2}$

Q.64 Which electronic configuration for oxygen is correct according to Hund's rule of multiplicity

[1] $1s^2, 2s^2\ 2p_x^2\ 2p_y^1\ 2p_z^1$ [2] $1s^2, 2s^2\ 2p_x^2\ 2p_y^2\ 2p_z^0$ [3] $1s^2, 2s^2\ 2p_x^3\ 2p_y^1\ 2p_z^0$ [4] None of these

Q.65 Which electronic configuration is not observing the (n + l) rule

[1] $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6 3d^1, 4s^2$ [2] $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6 3d^7, 4s^2$

[3] $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6 3d^5, 4s^1$ [4] $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6 3d^8, 4s^2$

Q.66 In a potassium atom, electronic energy levels are in the following order

[1] 4s > 3d [2] 4s > 4p [3] 4s < 3d [4] 4s < 3p

Q.67 How many electrons can be accommodated in a sub-shell for which n = 3, l = 1

[1] 8 [2] 6 [3] 18 [4] 32

Q.68 Azimuthal quantum number for last electron of Na atom is

[1] 1 [2] 2 [3] 3 [4] 0

Q.69 Which of the following set of quantum numbers belong to highest energy

[1] n = 4, l = 0, m = 0, s = $+\frac{1}{2}$ [2] n = 3, l = 0, m = 0, s = $+\frac{1}{2}$

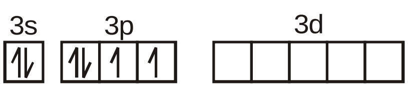
[3] n = 3, l = 1, m = 1, s = $+\frac{1}{2}$ [4] n = 3, l = 2, m = 1, s = $+\frac{1}{2}$

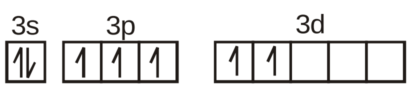
Q.70 Which of the following sets of quantum numbers is not allowed

[1] n = 1, l = 0, m = 0, s = $+\frac{1}{2}$ [2] n = 1, l = 1, m = 0, s = $-\frac{1}{2}$

[3] n = 2, l = 1, m = 1, s = $+\frac{1}{2}$ [4] n = 2, l = 1, m = 0, s = $-\frac{1}{2}$

Q.71 Which of the following has maximum energy

[1] 

[2] 

[3] 3s $\boxed{\uparrow\downarrow}$ 3p $\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$ 3d $\boxed{}\boxed{\uparrow}\boxed{}\boxed{}\boxed{}$

[4] 3s $\boxed{\uparrow\downarrow}$ 3p $\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$ 3d $\boxed{\uparrow}\boxed{}\boxed{}\boxed{}\boxed{}$

Q.72 Correct statement is

[1] K = $4s^1$, Cr = $3d^4\ 4s^2$, Cu = $3d^{10}\ 4s^2$
[2] K = $4s^2$, Cr = $3d^4\ 4s^2$, Cu = $3d^{10}\ 4s^2$
[3] K = $4s^2$, Cr = $3d^5\ 4s^1$, Cu = $3d^{10}\ 4s^2$
[4] K = $4s^1$, Cr = $3d^5\ 4s^1$, Cu = $3d^{10}\ 4s^1$

Q.73 The number of unpaired electrons in an O_2 molecule is

[1] 0 [2] 1 [3] 2 [4] 3

Q.74 The number of unpaired electrons in carbon atom in excited state is

[1] 1 [2] 2 [3] 3 [4] 4

Q.75 Which of the following has the least energy

[1] 2p [2] 3p [3] 2s [4] 4d

Q.76 The correct order of increasing energy of atomic orbitals is

[1] 5p < 4f < 6s < 5d [2] 5p < 6s < 4f < 5d [3] 4f < 5p < 5d < 6s [4] 5p < 5d < 4f < 6s

Q.77 The orbital angular momentum of an electron in 2s-orbital is

[1] $\frac{1}{2}\frac{h}{2\pi}$ [2] $\frac{h}{2\pi}$ [3] $\sqrt{2}\frac{h}{2\pi}$ [4] 0

Q.78 If the values of (n + l) is not > 3, then the maximum number of electrons in all the orbitals would be

[1] 12 [2] 10 [3] 2 [4] 6

Q.79 Given : Orbital – A Orbital – B

n = 3, l = 2 n = 5, l = 0

The order of their energies would be

[1] B > A [2] A > B [3] A = B [4] None of these

Q.80 Quantum numbers of 13th electron of silicon (Z = 14) are 3, 1, 1, +½. Which one of the following is true set of quantum numbers for the 14th electron ?

[1] 3, 1, 0, +½ [2] 3, 0, 1, +½ [3] 3, 1, 1, –½ [4] 3, 1, –1, +½

Q.81 Using hund's rule to decide which of the following is not the ground state electronic configuration

[1] [He] 2s $\boxed{\uparrow\downarrow}$ 2p $\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\uparrow}$
[2] [He] 2s $\boxed{\uparrow\downarrow}$ 2p $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}$
[3] [He] 2s $\boxed{\uparrow\downarrow}$ 2p $\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$
[4] [He] 2s $\boxed{\uparrow\downarrow}$ 2p $\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{}$

Q.82 What are the chances that the electron in a hydrogen atom has a spin quantum number of +½

[1] 25% [2] 100% [3] 50% [4] 0%

Q.83 Which species shows the electronic configuration $1s^2$, $2s^2\ 2p^5$, $3s^2$?

[1] Ne (Ground state) [2] Al^{+2} (Ground state) [3] Mg^+ (Excited state) [4] Na^+ (Excited state)

Q.84 If n and l are respectively the principal and azimuthal quantum numbers, then the expression for calculating the total number of electrons in any energy level is

[1] $\sum_{l=0}^{n=l} 2(2l+1)$ [2] $\sum_{l=1}^{l=n-1} 2(2l+1)$ [3] $\sum_{l=0}^{l=(n+1)} 2(2l+1)$ [4] $\sum_{l=0}^{l=n-1} 2(2l+1)$

ANSWER KEY

Qus.	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20
Ans.	3	2	4	4	3	3	1	1	1	2	1	3	2	2	3	2	1	3	2	2
Qus.	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36	37	38	39	40
Ans.	1	2	2	1	4	1	3	4	4	1	4	2	3	2	3	4	3	1	3	1
Qus.	41	42	43	44	45	46	47	48	49	50	51	52	53	54	55	56	57	58	59	60
Ans.	2	2	4	1	1	2	2	3	3	1	4	2	3	3	3	4	3	3	4	2
Qus.	61	62	63	64	65	66	67	68	69	70	71	72	73	74	75	76	77	78	79	80
Ans.	4	3	2	1	3	3	2	4	4	2	2	4	3	4	3	2	4	1	1	4
Qus.	81	82	83	84																
Ans.	4	3	3	4																

EXERCISE - 2

Q.1 From the given sets of quantum numbers the one that is inconsistent with the theory is : **[IIT-Scr. 1994]**

[1] n = 3, l = 2, m = – 3, s = + 1/2 [2] n = 4, l = 3, m = 3, s = + 1/2

[3] n = 2, l = 1, m = 0, s = – 1/2 [4] n = 4, l = 3, m = 2, s = + 1/2

Q.2 Calculate the momentum of a particle which has a de-Broglie's wavelength of 10^{-10} m (h = 6.6×10^{-34} kg m^2 sec^{-1}). [1] 6.6×10^{-25} kg ms^{-1} [2] 6.6×10^{-23} kg ms^{-1} [3] 6.6×10^{-24} kg ms^{-1} [4] None of these **[CPMT 1994]**

Q.3 The total number of electrons present in all the p-orbitals of bromine are : **[MP PET 1994]**

[1] Five [2] Eighteen [3] Seventeen [4] Thirty five

Q.4 When an electron revolves in a stationary orbit then : **[MP PET 1994]**

[1] It absorbs energy [2] It gains kinetic energy

[3] It emits radiation [4] Its energy remains constant

Q.5 The total number of valence electrons in 4.2 gm of N_3^- ion is (N_A is the Avogadro's number) : **[CBSE 1994]**

[1] 1.6 N_A [2] 3.2 N_A [3] 2.1 N_A [4] 4.2 N_A

Q.6 If n = 3, then the value of 'ℓ' which is incorrect : **[CPMT 1994]**

[1] 0 [2] 1 [3] 2 [4] 3

Q.7 Chlorine atom differs from chloride ion in the number of : **[MP PET 1995]**

[1] Proton [2] Neutron [3] Electrons [4] Protons and electrons

Q.8 The uncertainty in the position of an electron (mass = 9.1×10^{-28} g) moving with a velocity of 3.0×10^4 cm s^{-1} accurate upto 0.001% will be (use $\frac{h}{4\pi}$ in the uncertainty expression, where h = 6.62×10^{-27} erg-s)

[1] 1.92 cm [2] 7.68 cm [3] 5.76 cm [4] 3.84 cm **[CBSE 1995]**

Q.9 A 3p orbital has : **[IIT 1995]**

[1] Two spherical nodes [2] Two non-spherical nodes

[3] One spherical and one non-spherical nodes [4] One spherical and two non-spherical nodes

Q.10 Zeeman effect refers to the : **[AFMC 1995]**

[1] Splitting up of the lines in an emission spectrum in a magnetic field

[2] Splitting up to the lines in an emission spectrum in the presence of an external electrostatic field

[3] Emission of electrons from metals when light falls upon them

[4] Random scattering of light by colloidal particles

Q.11 For n = 3 energy level, the number of possible orbitals are : **[MP PMT 1995]**

[1] 1 [2] 3 [3] 4 [4] 9

Q.12 The orbital angular momentum of an electron in 2s orbital is : **[IIT 1996]**

[1] $+\frac{1}{2} \cdot \frac{h}{2\pi}$ [2] Zero [3] $\frac{h}{2\pi}$ [4] $\sqrt{2} \cdot \frac{h}{2\pi}$

Q.13 Which statement is not correct for n = 5, m = 3 : **[CPMT 1996]**

[1] l = 4 [2] l = 0, 1, 2, 3 ; s = + 1/2

[3] l = 3 [4] All are correct

Q.14 $1s^2$, $2s^2$ $2p^5$ $3s^2$ shows configuration of : **[CPMT 1996]**

[1] Al^{+3} in ground state [2] Ne in excited state
[3] Mg^{+1} in excited stated [4] All are correct

Q.15 In a Bohr's model of atom when an electron jumps from n = 1 to n = 3, how much energy will be emitted or absorbed : **[CBSE 1996]**

[1] 2.15×10^{-11} ergs [2] 0.1911×10^{-10} ergs [3] 2.389×10^{-12} ergs [4] 0.239×10^{-10} ergs

Q.16 The shape of an orbital is given by the quantum number : **[MP PMT 1996]**

[1] n [2] l [3] m [4] s

Q.17 Which of the following metal ions will have maximum number of unpaired electrons : **[CPMT 1996]**

[1] Fe^{+2} [2] Co^{+2} [3] Ni^{+2} [4] Mn^{+2}

Q.18 The maximum probability of finding an electron in the d_{xy} orbital is : **[MT PET 1996]**

[1] Along the x-axis [2] Along the y-axis
[3] At an angle of 45º from the x and y-axes [4] At an angle of 90º from the x and y-axes

Q.19 CO has same electrons as or the ion that is isoelectronic with CO is : **[CBSE 1997]**

[1] N_2^+ [2] CN^- [3] O_2^+ [4] O^-_2

Q.20 The total number of orbital in an energy level designated by principal quantum number n, is equal to :

[1] 2n [2] $2n^2$ [3] n [4] n^2 **[AIIMS 1997]**

Q.21 Which electronic level would allow the hydrogen atom to absorb a photon but not to emit a photon :

[1] 3s [2] 2p [3] 2s [4] 1s **[CPMT 1997]**

Q.22 An electron has principal quantum number 3. The number of its : **[MT PET 1997]**

(i) subshells and (ii) orbitals would be respectively

[1] 3 and 5 [2] 3 and 7 [3] 3 and 9 [4] 2 and 5

Q.23 Aufbau principle is not satisfied by : **[MP PMT 1997]**

[1] Cr and Cl [2] Cu and Ag [3] Cr and Mg [4] Cu and Na

Q.24 The first use of quantum theory to explain the structure of atom was made by : **[IIT 1997]**

[1] Heisenberg [2] Bohr [3] Planck [4] Einstein

Q.25 Five valence electrons of $_{15}P$ are labelled as $\underset{3s}{\boxed{AB}}$ $\underset{3p}{\boxed{X|Y|Z}}$. If the spin quantum number of B and Z is + 1/2, the group of electrons with three of the quantum number same are : **[JIPMER 1997]**

[1] Ab, XYZ, BY [2] AB [3] XYZ, AZ [4] AB, XYZ

Q.26 In an element going away from nucleus, the energy of particle : **[RPMT 1997]**

[1] Decreases [2] Unchanged [3] Increases [4] None of these

Q.27 In neutral atom, which particles are equivalent : **[RPMT 1997]**

[1] p^+, e^+ [2] e^-, e^+ [3] e^-, p^+ [4] p^+, n^0

Q.28 If n + l = 6, then total possible number of subshells would be : **[RPMT 1997]**

[1] 3 [2] 4 [3] 2 [4] 5

Q.29 The configuration $1s^2\ 2s^2\ 2p^5\ 3s^1$ shows : **[AIIMS 1997]**

[1] Ground state of fluorine atom [2] Excited state of fluorine atom
[3] Excited state of neon atom [4] Excited state of ion O_2^- ion

Q.30 The electron configuration of gadolinium (atomic no. 64) is : **[CBSE 1997]**

[1] [Xe] $4f^8\ 5d^9\ 6s^2$ [2] [Xe] $4f^7\ 5d^1\ 6s^2$ [3] [Xe] $4f^3\ 5d^5\ 6s^2$ [4] [Xe] $4f^6\ 5d^2\ 6s^2$

Q.31 If electron falls from n = 3 to n = 2, then emitted energy is : **[AFMC 1997]**

[1] 10.2 eV [2] 12.09 eV [3] 1.9 eV [4] 0.65 eV

Q.32 Number of protons, neutrons and electrons in the element $_{89}X^{231}$ is : **[AFMC 1997]**

[1] 89, 231, 89 [2] 89, 89, 242 [3] 89, 142, 89 [4] 89, 71, 89

Q.33 In the ground state configuration of Cr_{24} how many orbitals are present having paired and unpaired electrons :

[1] 10 [2] 12 [3] 15 [4] 16 **[RPMT 1997]**

Q.34 Discoverer of positron : **[RPMT 1997]**

[1] Paulling [2] Anderson [3] Yukawa [4] Segre

Q.35 Which of the following species not contains neutrons : **[RPMT 1997]**

[1] H [2] Li^{+2} [3] C [4] O

Q.36 The energy of an electron in the first Bohr orbit of H atom is – 13.6 eV. The possible energy value of the first excited state for electrons in Bohr orbits to hydrogen is (are) : **[IIT 1998]**

[1] – 3.4 eV [2] – 4.2 eV [3] – 6.8 eV [4] + 6.8 eV

Q.37 The energy of an electron in the first orbit of He^{+} is -871.6×10^{-20} J. The energy of the electron in the first orbit of hydrogen would be :
[Roorkee 1998]

[1] -871.6×10^{-20} J [2] -435.8×10^{-20} J [3] -217.9×10^{-20} J [4] -108.9×10^{-20} J

Q.38 The position of both an electron and a helium atom is known within 1.0 nm and the momentum of the electron is known within 50×10^{-26} kg ms^{-1}. The minimum uncertainty in the measurement of the momentum of the helium atom is : **[CBSE 1998]**

[1] 50 kg ms^{-1} [2] 60 kg ms^{-1} [3] 80×10^{-26} kg ms^{-1} [4] 50×10^{-26} kg ms^{-1}

Q.39 The Bohr orbit radius for the hydrogen atom (n = 1) is approximately 0.530 A. The radius for the first excited state (n = 2) orbit is : **[CBSE 1998]**

[1] 0.13 Å [2] 1.06 Å [3] 4.77 Å [4] 2.12 Å

Q.40 Which of the following explain the sequence of filling the electrons in different orbitals : **[AIIMS 1998]**

[1] Hund's rule [2] Octet rule [3] Aufbau principle [4] All of these

Q.41 Number of orbitals having paired electrons for gaseous Fe are : **[RPMT 1998]**

[1] 4 [2] 11 [3] 15 [4] 19

Q.42 [Ar] $3d^6$ is the configuration of the following ion : **[RPMT 1998]**

[1] Fe^{+2} [2] Ti^{+3} [3] Co^{+2} [4] Cr^{+3}

Q.43 Which triad of quantum number [n, *l*, m] is not applicable for 3d-electron : **[RPMT 1998]**

[1] 3, 2, 0 [2] 3, 1, –1 [3] 3, 2, – 2 [4] 3, 2, +1

Q.44 Which of the following configuration follows the Hund's rule : **[RPMT 1998]**

[1] [He] 2s [↑↓] 2p [↑][↑][ ] [2] [He] 2s [↑↓] 2p [↑↓][↑][ ]

[3] [He] 2s [↑↓] 2p [↑][↑↓][ ] [4] [He] 2s [↑↓] 2p [↓][↑][ ]

Q.45 The ratio of radii of 3rd and 2nd Bohr's orbit of hydrogen atom is : **[RPMT 1998]**

[1] 3 : 2 [2] 4 : 9 [3] 9 : 4 [4] 9 : 1

Q.46 The four quantum number for the valence shell electron or last electron of sodium is : **[MP PMT 1999]**

[1] n = 2, ℓ = 1, m = – 1, s = – 1/2 [2] n = 3, ℓ = 0, m = 0, s = + 1/2
[3] n = 3, ℓ = 2, m = – 2, s = – 1/2 [4] n = 3, ℓ = 2, m = 2, s = + 1/2

Q.47 Heaviest particle is : **[MP PET 1999]**
[1] Meson [2] Neutron [3] Proton [4] Electron

Q.48 Which is correct statement about proton : **[MP PET 1999]**
[1] Proton is nucleus of deuterium [2] Proton is ionized hydrogen molecule
[3] Proton is ionized hydrogen atom [4] Proton is α-particle

Q.49 The energy of an electron in n^{th} orbit of hydrogen atom is : **[MP PET 1999]**
[1] $-\frac{13.6}{n^4}$ eV [2] $-\frac{13.6}{n^3}$ eV [3] $-\frac{13.6}{n^2}$ eV [4] $-\frac{13.6}{n}$ eV

Q.50 If wavelength of photon is 2.2×10^{-11} m, h = 6.8×10^{-34} Js, then momentum of photon is : **[MP PET 1999]**
[1] 3×10^{-23} Kg ms^{-1} [2] 3.33×10^{22} Kg ms^{-1}
[3] 1.452×10^{-44} kg ms^{-1} [4] 6.89×10^{43} kg ms^{-1}

Q.51 The electrons identified by quantum number n and l **[IIT 1999]**
(i) n = 4, l = 1 (ii) n = 4, l = 0 (iii) n = 3, l = 2 (iv) n = 3, l = 1
can be placed in order of increasing energy from the lowest to highest, as :
[1] (iv) < (ii) < (iii) < (i) [2] (ii) < (iv) < (i) < (iii)
[3] (i) < (iii) < (ii) < (iv) [4] (iii) < (i) < (iv) < (ii)

Q.52 Ground state electron configuration of nitrogen atom can be represent by : **[IIT 1999]**
[1] $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$ [2] $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\downarrow}\boxed{\uparrow}$ [3] $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\downarrow}\boxed{\downarrow}$ [4] None of these

Q.53 Which of the following has more unpaired d-electron : **[CBSE 1999]**
[1] Zn^+ [2] Fe^{2+} [3] Ni^{3+} [4] Cu^+

Q.54 The uncertainty in momentum of an electron is 1×10^{-5} kg-m/s. The uncertainty in its position will be : **[CBSE 1999]**
[1] 1.05×10^{-28} m [2] 1.05×10^{-26} m [3] 5.27×10^{-30} m [4] 5.25×10^{-28} m

Q.55 The de-Broglie wavelength of a particle with mass 1g and velocity 100 m/s is : **[CBSE 1999]**
[1] 6.63×10^{-33} m [2] 6.63×10^{-34} m [3] 6.63×10^{-35} m [4] 6.65×10^{-32} m

Q.56 Which of the following set of quantum numbers belong to highest energy : **[CPMT 1999]**
[1] n = 4, l = 0, m = 0, s = + $\frac{1}{2}$ [2] n = 3, l = 0, m = 0, s = + $\frac{1}{2}$
[3] n = 3, l = 1, m = 1, s = + $\frac{1}{2}$ [4] n = 3, l = 2, m = 1, s = + $\frac{1}{2}$

Q.57 Which of the following are isoelectronic species : **[CPMT 1999]**
I – CH^+_3, II – NH_2^-, III – NH^+_4, IV – NH_3
[1] I, II, III [2] II, III, IV [3] I, II, IV [4] I and II

Q.58 Which quantum number will determine the shape of the subshell : **[CPMT 1999]**

[1] Principal quantum number [2] Azimuthal quantum number

[3] Magnetic quantum number [4] Spin quantum number

Q.59 A metal in its di positive state has the electronic configuration 2, 8, 14 and has the atomic weight equal to 56. Number of neutrons in its nucleus would be : **[RPMT 1999]**

[1] 30 [2] 32 [3] 34 [4] 28

Q.60 Which set of quantum number for an electron of an atom is not possible : **[RPMT 1999]**

[1] $n = 1, l = 0, m = 0, s = +\frac{1}{2}$ [2] $n = 1, l = 1, m = 1, s = +\frac{1}{2}$

[3] $n = 1, l = 0, m = 0, s = -\frac{1}{2}$ [4] $n = 2, l = 1, m = -1, s = +\frac{1}{2}$

Q.61 Outer electronic configuration of the element of atomic number 24 is : **[RPMT 1999]**

[1] $3s^2\ 3p^6\ 3d^5\ 4s^1$ [2] $3s^2\ 3p^6\ 3d^4\ 4s^2$ [3] $3s^2\ 3p^6\ 3d^6$ [4] None

Q.62 The basis of three unpaired electrons in the configuration of nitrogen is : **[RPMT 1999]**

[1] Aufbau principle [2] Pauli's principle [3] Hund's principle [4] Uncertainty principle

Q.63 Correct order of size is : **[RPMT 1999]**

[1] $I > I^+ > I^-$ [2] $I > I^- > I^+$ [3] $I^- > I > I^+$ [4] $I^+ > I > I^-$

Q.64 Which is not electromagnetic radiation : **[RPMT 2000]**

[1] Infrared rays [2] X–ray [3] Cathode rays [4] Gamma rays

Q.65 Which of the following pair of orbitals posses two nodal planes : **[RPMT 2000]**

[1] $p_{xy}, d_{x^2-y^2}$ [2] d_{xy}, d_{zx} [3] p_{yz}, d_{zx} [4] $d_{z^2}, d_{x^2-y^2}$

Q.66 The mass of a neutron is of the order of : **[RPMT 2000]**

[1] 10^{-23} kg [2] 10^{-24} kg [3] 10^{-26} kg [4] 10^{-27} kg

Q.67 Smallest cation is : **[RPMT 2000]**

[1] Na^+ [2] Mg^{2+} [3] Ca^{2+} [4] Al^{3+}

Q.68 Electron enters the sub-shell for which $(n + l)$ value is minimum. This is enunciated as : **[RPMT 2000]**

[1] Hund's rule [2] Aufbau principle

[3] Heisenberg uncertainty principle [4] Pauli's exclusion principle

Q.69 The minimum real charge on any particle which can exist is : **[RPMT 2000]**

[1] 1.6×10^{-19} coulomb [2] 1.6×10^{-10} coulomb

[3] 4.8×10^{-10} coulomb [4] Zero

Q.70 Which sub-shell is not permissible : **[RPMT 2000]**

[1] 2d [2] 4f [3] 6p [4] 3s

Q.71 Sub-shell designated by azimuthal quantum number $l = 3$ can have maximum number of electrons :

[1] 14 [2] 6 [3] 10 [4] 0 **[RPMT 2000]**

Q.72 Quantum number $n = 3, l = 2, m = +2$ shows how many orbitals : **[CPMT 2001]**

[1] 1 [2] 2 [3] 3 [4] 4

Q.73 Which is isoelectronic with sulphide ion : **[RPMT 2001]**

[1] Cl [2] Ne [3] Ar [4] Kr

Q.74 Ground state electronic configuration of nitrogen is : **[RPMT 2001]**

[1] $1s^2, 2s^2, 2p_x^1, 2p_y^1, 2p_z^1$ [2] $1s^2, 2s^2, 2p_x^2, 2p_y^1$

[3] $1s^2, 2s^2, 2p_x^2, 2p_z^1$ [4] $1s^2, 2s^2, 2p_x^3$

Q.75 In the Bohr's orbit, what is the ratio of total kinetic energy and total energy of electron : **[RPMT 2002]**

[1] – 1 [2] – 2 [3] 1 [4] + 2

Q.76 Rutherford α-particle dispersion experiment concludes : **[RPMT 2002]**

[1] All positive ions are deposited at small part [2] All negative ions are deposited at small part

[3] Protons moves around the electrons [4] Neutrons are charged particles

Q.77 Which of the following element outermost orbit's last electron has magnetic quantum number m = 0 ?

[1] Na [2] O [3] Cl [4] N **[RPMT 2002]**

Q.78 The value of Planck's constant is 6.63×10^{-34} Js. The velocity of light is 3.0×10^8 m s^{-1}. Which value is closest to the wavelength in nanometers of a quantum of light with frequency of 8×10^{15} s^{-1} ? **[CPMT 2003]**

[1] 5×10^{-18} [2] 4×10^1 [3] 3×10^7 [4] 2×10^{-25}

Q.79 The orbital angular momentum for an electron revolving in an orbit is given by $\sqrt{l(l+1)}\cdot\frac{h}{2\pi}$. This momentum for an s-electron will be given by : **[AIEEE 2003]**

[1] $\sqrt{2}.\frac{h}{2\pi}$ [2] $+\frac{1}{2}.\frac{h}{2\pi}$ [3] zero [4] $\frac{h}{2\pi}$

Q.80 The number of d-electrons retained in Fe^{2+} (At. no. of Fe = 26) ion is : **[AIEEE 2003]**

[1] 6 [2] 3 [3] 4 [4] 5

Q.81 The de Broglie wavelength of a tennis ball of mass 60 g moving with a velocity of 10 metres per second is approximately : **[AIEEE 2003]**

[1] 10^{-25} metres [2] 10^{-33} metres [3] 10^{-31} metres [4] 10^{-16} metres

Q.82 In Bohr series of lines of hydrogen spectrum, the third line from the red end corresponds to which one of the following inter-orbit jumps of the electron for Bohr orbits in an atom of hydrogen ? **[AIEEE 2003]**

[1] $2 \rightarrow 5$ [2] $3 \rightarrow 2$ [3] $5 \rightarrow 2$ [4] $4 \rightarrow 1$

Q.83 Which of the following sets of quantum number is correct for an electron in 4f orbital ? **[AIEEE 2004]**

[1] n = 3, l = 2, m = – 2, s = $+\frac{1}{2}$ [2] n = 4, l = 4, m = – 4, s = $-\frac{1}{2}$

[3] n = 4, l = 3, m = + 1, s = $+\frac{1}{2}$ [4] n = 4, l = 3, m = + 4, s = $+\frac{1}{2}$

Q.84 consider the ground state of Cr atom (Z = 24). The numbers of electrons with the azimuthal quantum numbers, l = 1 and 2 are, respectively : **[AIEEE 2004]**

[1] 16 and 5 [2] 12 and 5 [3] 16 and 4 [4] 12 and 4

Q.85 The wavelength of the radiation emitted, when in a hydrogen atom electron falls from infinity to stationary state 1, would be (Rydberg constant = 1.097×10^7 m^{-1}) : **[AIEEE 2004]**

[1] 9.1×10^{-8} nm [2] 192 nm [3] 406 nm [4] 91 nm

Q.86 Which one of the following sets of ions represents the collection of isoelectronic species ? **[AIEEE 2004]**

[1] $Na^+, Mg^{2+}, Al^{3+}, Cl^-$ [2] $Na^+, Ca^{2+}, Sc^{3+}, F^-$ [3] $K^+, Cl^-, Mg^{2+}, Sc^{3+}$ [4] $K^+, Ca^{2+}, Sc^{3+}, Cl^-$

Q.87 Consider the following nuclear relations : **[AIEEE 2004]**

$$^{238}_{92}M \rightarrow ^{x}_{y}N + 2\,^{4}_{2}He \qquad\qquad ^{x}_{y}N \rightarrow ^{A}_{B}L + 2\beta^{+}$$

The number of neutrons in the element L is :

[1] 146 [2] 144 [3] 140 [4] 142

Q.88 Which of the following have same radius as hydrogen n = 1 : **[IIT Scr. 2004]**

[1] He^{+}, n = 2 [2] Be^{+3}, n = 2 [3] Li^{+2}, n =2 [4] Li^{+2}, n = 3

Q.89 In a multi-electron atom, which of the following orbitals described by the three quantum members will have the same energy in the absence of magnetic and electric fields ? **[AIEEE 2005]**

(A) n = 1, l = 0, m = 0 (B) n = 2, l = 0, m = 0 (C) n = 2, l = 1, m = 1 (D) n = 3, l = 2, m = 1

(E) n = 3, l = 2, m = 0

[1] (D) and (E) [2] (C) and (D) [3] (B) and (C) [4] (A) and (B)

Q.90 Of the following sets which one does NOT contain isoelectronic species ? **[AIEEE 2005]**

[1] $BO_3^{3-}, CO_3^{2-}, NO_3^{-}$ [2] $SO_3^{2-}, CO_3^{2-}, NO_3^{-}$ [3] CN^{-}, N_2, C_2^{2-} [4] $PO_4^{3-}, SO_4^{2-}, ClO_4^{-}$

Q.91 Which of the following statements in relation to the hydrogen atom is correct ? **[AIEEE 2005]**

[1] 3s, 3p and 3d orbitals all have the same energy

[2] 3s and 3p orbitals are of lower energy than 3d orbitals

[3] 3p orbital is lower in energy than 3d orbital

[4] 3s orbitals is lower in energy than 3p orbital

Q.92 Which of the following factors may be regarded as the main cause of lanthanide contraction ? **[AIEEE 2005]**

[1] Greater shielding of 5d electron by 4f electrons

[2] Poorer shielding of 5d electron by 4f electrons

[3] Effective shielding of one of 4f electrons by another in the subshell

[4] Poor shielding of one of 4f electron by another in the subshell

Q.93 Which of the following molecules/ions does not contain unpaired electrons ? **[AIEEE 2006]**

[1] B_2 [2] N_2^{+} [3] O_2 [4] O_2^{2-}

Q.94 According to Bohr's theory, the angular momentum of an electron in 5^{th} orbit is **[AIEEE 2006]**

[1] 1.0 h/π [2] 10 h/π [3] 2.5 h/π [4] 25 h/π

Q.95 Uncertainty in the position of an electron (mass = 9.1×10^{-31} Kg) moving with a velocity 300 ms^{-1}, accurate upto 0.001%, will be (h = 6.63×10^{-34} Js) **[AIEEE 2006]**

[1] 5.76×10^{-2} m [2] 1.92×10^{-2} m [3] 3.84×10^{-2} m [4] 19.2×10^{-2} m

Q.96 Which one of the following sets of ions represents a collection of isoelectronic species ? **[AIEEE 2006]**

[1] $Ba^{2+}, Sr^{2+}, K^{+}, Ca^{2+}$ [2] $N^{3-}, O^{2-}, F^{-}, S^{2-}$ [3] $Li^{+}, Na^{+}, Mg^{2+}, Ca^{2+}$ [4] $K^{+}, Cl^{-}, Ca^{2+}, Sc^{3+}$

Q.97 Lanthanoid contraction is caused due to **[AIEEE 2006]**

[1] the appreciable shielding on outer electrons by 5d electrons from the nuclear charge

[2] the same effective nuclear charge from Ce to Lu

[3] the imperfect shielding on outer electrons by 4f electrons from the nuclear charge

[4] the appreciable shielding on outer electrons by 4f electrons from the nuclear charge

ANSWER KEY

Qus.	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20
Ans.	1	3	3	4	1	4	3	1	3	1	4	2	2	3	2	2	4	3	2	4
Qus.	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36	37	38	39	40
Ans.	4	3	2	2	4	3	3	1	3	2	3	3	3	2	1	1	3	4	4	3
Qus.	41	42	43	44	45	46	47	48	49	50	51	52	53	54	55	56	57	58	59	60
Ans.	2	1	2	1	3	2	2	3	3	1	1	1	2	3	1	4	2	2	1	2
Qus.	61	62	63	64	65	66	67	68	69	70	71	72	73	74	75	76	77	78	79	80
Ans.	1	3	3	3	2	4	4	2	1	1	1	1	3	1	1	1	1	2	3	1
Qus.	81	82	83	84	85	86	87	88	89	90	91	92	93	94	95	96	97			
Ans.	2	3	3	2	4	4	2	2	1	2	1	4	4	3	2	4	3			