CHEMISTRY

STUDY MATERIAL

STRUCTURE OF ATOM

AIEEE

THE NARAYANA GROUP

NARAYANA INSTITUTE OF CORRESPONDENCE COURSES

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PREFACE

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As you are planning to take various Engineering Entrance Examinations, we are sure that this **STUDY PACKAGE** is going to be of immense help to you.

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The salient features of this package include :

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- Sufficient number of solved examples in Physics, Chemistry & Mathematics in all the chapters to motivate the students attempt all the questions.
- All the chapters are followed by various types of exercises (Level-I, Level-II, Level-III and Questions asked in AIEEE and other Engineering Exams).

These exercises are followed by answers in the last section of the chapter. *This package will help you to know* what to study, how to study, time management, your weaknesses and improve your performance.

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We feel that there is always a scope for improvement. We would welcome your suggestions & feedback.

Wish you success in your future endeavours.

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ACKNOWLEDGEMENT

While preparing the study package, it has become a wonderful feeling for the NARAYANA TEAM to get the wholehearted support of our Staff Members including our Designers. They have made our job really easy through their untiring efforts and constant help at every stage.

We are thankful to all of them.

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STRUCTURE OF ATOM

AIEEE Syllabus

Constituents of the atom (discovery of electron, rutherford model of the atom). Electronic structure of atoms - nature of light and electromagnetic waves, atomic spectra, bohr's model of hydrogen, shortcomings of the bohr model Dual nature of matter and radiation. de-Broglie relation. The uncertainty principle, Quantum Mechanical Model of the atom, Orbitals and Quantum numbers. Shapes of orbitals. Aufbau principle, Pauli Exclusion Principle, Hund's Rule, Electronic Configuration of atoms.

CONTENTS

- *Dalton's atomic theory*
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INTRODUCTION

This chapter firstly deals with discovery of fundamental particles i.e. electron, proton and neutron. After this, it covers various atomic models and explanation of hydrogen spectrum from the most important feature of Bohr's model. Next we consider photo electric effect which leads us to the concept of dual nature of matter. Finally we show how this information is understood in terms of the quantum theory.

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1. 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.

1.1 LIMITATIONS

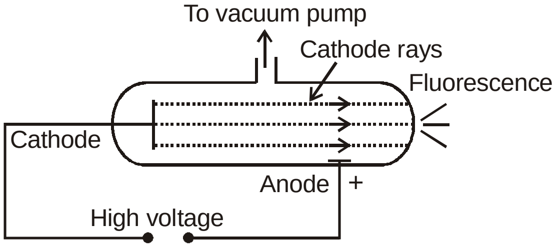
In 1833, Michael Faraday showed that there is a relationship between matter and electricity. This was the first major break-through to suggest that atom was not a simple indivisible particle of all matter but was made up of small particles. Discovery of electrons, protons and neutrons discarded the indivisible nature of the atom proposed by John Dalton.

2. FUNDAMENTAL PARTICLES

The complexity of the atom was further revealed when the discovery of cathode rays, positive rays, X-ray, radioactivity, isotopes, isobars and the new atomic model were made in subsequent years.

2.1 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

2.1.1 Properties of Cathode Rays

(a) 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.
(c) Cathode rays get deviated in electrical and magnetic fields. This proves that they are composed of charged particles. Their deviation 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.
(d) Cathode rays produce green fluorescence on the walls of the glass tube.
(e) Cathode rays produce incandescence at thin metal foil.

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(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.

(j) The nature of the cathode rays is independent of (i) the nature of the cathode and (ii) the gas in the discharge tube.

2.1.2 Charge/Mass Ratio of Electron

The ratio of negative charge (e) and mass (m) for cathode ray particle (electrons) is a constant. 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.

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

2.1.3 Charge on the Electron

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.

2.1.4 Mass of Electrons

(a) 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.

$$\frac{e}{e/m} = \frac{1.6012\times10^{-19}\text{coulomb}}{1.76\times10^{8}\text{coulomb / gram}} = 9.1091 \times 10^{-28} \text{ gram} = 0.0005486 \text{ a.m.u.}$$

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

(b) 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.

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

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

(c) 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 $\frac{1.008}{5.483\times10^{-4}} = 1837$. In other words, an atom of hydrogen (or a proton) is 1837 times heavier than electron.

$$\frac{\text{Mass of H atom}}{\text{Mass of electron}} = \frac{1.67\times10^{-24}}{5.483\times10^{-28}} = 1837$$

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

(e) 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.

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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.

2.2 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.

2.2.1 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.

2.2.2 Charge/Mass Ratio of Proton

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.

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.

2.2.3 Charge on the Proton

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

2.2.4 Mass of Proton

(a) Mass of proton (m) $= 1.6725 \times 10^{-24}$ gram ; $= 1.6725 \times 10^{-27}$ kg

$= 1.6725 \times 10^{-29}$ quintal ; = 1837 times that of electron

= 1.00757 a.m.u. ; = Mass of hydrogen atom

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

(b) 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)

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2.3 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 electroneutral 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, $_0n^1$.

2.3.1 Properties of Neutron

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

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

(c) 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 = 1.6748×10^{-24} gram = Approximately mass of a proton

(d) Neutron is the all of the three fundamental particles heaviest.

(e) 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 $1.6748 \times 10^{-24} \times 6.23 \times 10^{23} = 1.00893$.

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

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

(h) 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.

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

2.4 CHARACTERISTICS OF FUNDAMENTAL PARTICLES

	Electron	**Proton**	**Neutron**
Symbol	e or e^{-1}	P	n
Approximate relative mass	1/1836	1	1
Approximate relative charge	–1	+1	No charge
Mass in kg	9.109×10^{-31}	1.673×10^{-27}	1.675×10^{-27}
Mass in amu	5.485×10^{-4}	1.007	1.008
Actual charge (coulomb)	1.602×10^{-19}	1.602×10^{-19}	0
Actual charge (e.s.u.)	4.8×10^{-10}	4.8×10^{-10}	0

2.5 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.

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

3. ATOMIC MODELS

3.1 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.

3.2 RUTHERFORD'S MODEL OF AN ATOM

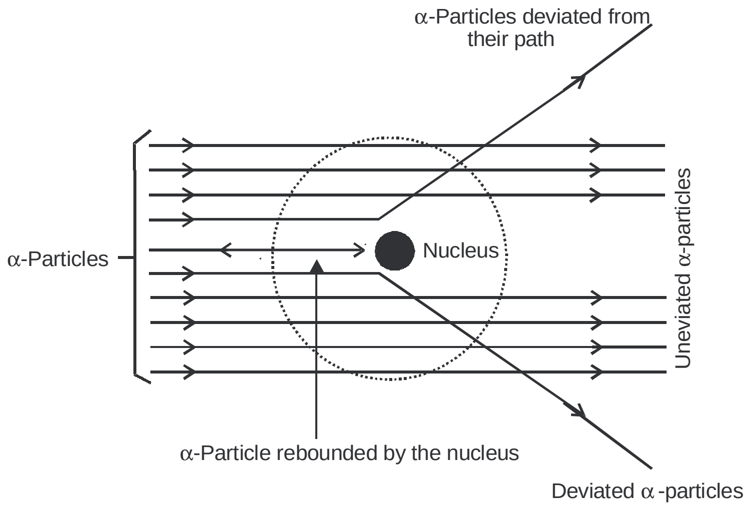
Ernest Rutherford in 1911put forward the **"nuclear model"** of atom on the basis of α particle scattering experiment. In this experiment, **Rutherford** showered α–particles (Helium nuclei, He^{+2}) on a 0.01mm thin gold foil.

3.2.1 Rutherford Observed

(a) Most of the a-particle passed through the gold film without any appreciable deviation from their path.

(b) Some particles got deviated from their original path of movement.

(c) A few of them rebounded towards own source.



3.2.2 Inferences

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

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(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 = $\dfrac{\text{Radius of atom}}{\text{Radius of atomic nucleus}} = \dfrac{10^{-8}\,\text{cm}}{10^{-13}\,\text{cm}} = 10^5$

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

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

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

3.2.3 Model

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

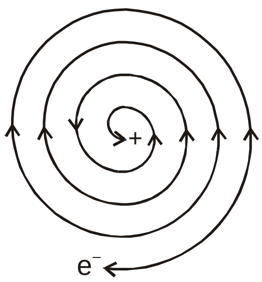
(ii) whole of the positive charge and almost all the mass of atom is confined to a very minute part at the centre of the atom, called the nucleus of the atom. The radius of nucleus is about 10^{-13} to 10^{-12} cm (or 10^{-15} to 10^{-14} meter), while the radius of atom is in the order of 10^{-8} cm.

(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.

3.2.4 Demerits

(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 H-atom.

Example - 1 Calculate the density of fluorine nucleus supposing that the shape of the nucleus is spherical and its radius is 5×10^{-13}. (Mass of F = 19 amu)

Solution : $\therefore$ Mass of the nucleus of F atom $= 19 \times 1.66 \times 10^{-24}$ gm

(1 amu = 1.660×10^{-24} gm)

Volume of the nucleus of F atom $= \frac{4}{3} \pi r^3 = \frac{4}{3} \times 3.14\ (5 \times 10^{-13})^3$

$= 525 \times 10^{-39}\ cm^3$

Density of the nucleus of F atom $= \frac{\text{Mass}}{\text{Volume}} = \frac{19 \times 1.66 \times 10^{-24} \text{gm}}{525 \times 10^{-39} \text{cm}^3}$

$= 6.0 \times 10^{13}\ gm\ cm^{-3}$

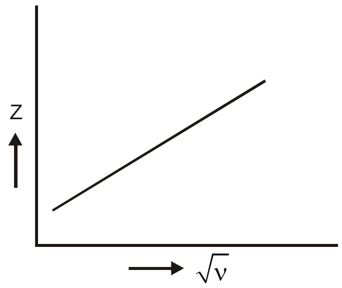
3.3 SOME IMPORTANT TERMS

3.3.1 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.

$$\sqrt{\nu} = a(z - b); \quad \frac{1}{a}\sqrt{\nu} = z - b$$

$$\therefore \quad Z = \frac{1}{a}\sqrt{\nu} + b \quad [y = mx + c]$$



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

3.3.2 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.

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(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.

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.

3.3.3 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

3.3.4 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.

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	$_8O^{16}$	$_8O^{17}$	$_8O^{18}$
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 $_{17}Cl^{37}$

3.3.5 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

3.3.6 Isotones

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

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

3.3.7 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		

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Example - 2 If the atomic weight of Zn is 70 and its atomic number is 30, then what will be the atomic weight of Zn^{+2} ?

(1) 70 (2) 68

(3) 72 (4) 74 **Ans. (1)**

Solution : Two electrons are removed in the formation of Zn^{+2} from Zn. The numbers of protons and neutrons remain unchanged.

Example - 3 The mass of one mole electrons should be -

(1) 0.55 mg (2) 1.0008 gm

(3) 1.000 gm (4) 0.184 gm **Ans. (1)**

Solution : The number of electrons in one mole = Avogadro number = 6.023×10^{23}

Mass of one electron = 9.1×10^{-28} gm

Therefore Mass of 6.023×10^{23} electrons = $9.1 \times 10^{-28} \times 6.023 \times 10^{23}$ = 0.55 mg

Example - 4 The number of atoms present in 20 grams of calcium will be equal to the number of atoms present in

(20 gm Ca = $\frac{1}{2}$ mole Ca); Ca $= \frac{6.023 \times 10^{23}}{2} = 3.012 \times 10^{23}$

(1) 12 gm C (2) 12.15 gm Mg

(3) 24.0 gm C (4) 24.3 gm Mg **Ans. (2)**

Solution : 24.3 gm Mg = 1 mole, therefore 12.15 gm = $\frac{1}{2}$ mole

Example - 5 Which of the following pairs consists of molecules having same mass number ?

(1) H_2O and D_2O (2) H_2O and HTO

(3) D_2O and HTO (4) D_2O and HCl **Ans. (3)**

Solution : Mass number of H_2O = 18; Mass number of D_2O = 20

Mass number of HTO = 20; Mass number of HCl = 36.5

Example - 6 The mass number of three isotopes of an element are 11, 12 and 13 units. Their percentage abundance is 80, 15 and 5 respectively. What should be the atomic weight of the element ?

(1) 11.25 (2) 20

(3) 16 (4) 10 **Ans. (1)**

Solution : 80 : 15 : 5

Thus the ratio is 16 : 3 : 1

Total = 16 + 3 + 1 = 20

Average weight $= \frac{11 \times 16 + 12 \times 3 + 13 \times 1}{20} = 11.25$

Example - 7 If two neutrons are added to an element X, then it will get converted to its

(1) isotope (2) isotone

(3) isobar (4) None of the above **Ans. (1)**

Solution : The number of neutrons are different in the isotopes of the same element.

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Example - 8 In two elements ${}_{Z_1}A^{M_1}$ and ${}_{Z_2}B^{M_2}$, $M_1 \neq M_2$ and $Z_1 \neq Z_2$ but $M_1 - Z_1 = M_2 - Z_2$. These elements are

(1) isotonic (2) isotopic
(3) isobaric (4) isoprotonic **Ans (1)**

Solution : M_1 = Atomic weight Z_1 = Atomic number

In isobars $M_1 = M_2$ and in isotopes $Z_1 = Z_2$

In isotones (isoneutronic elements) $M_1 - Z_1 = M_2 - Z_2$

Example - 9 Two nuclides A and B are isoneutronic. Their mass numbers are 76 and 77 respectively. If atomic number of A is 32, then the atomic number of B will be

(1) 33 (2) 34
(3) 32 (4) 30 **Ans. (1)**

Solution :

	${}_{32}A^{76}$		${}_{P}B^{77}$
Protons	= 32	Protons + Neutrons	= 77
Protons + Neutrons	= 76	Neutrons	= 44
Neutrons	= 44	Protons	= 33

Example - 10 The isoelectronic pair of 32 electron is

(1) BO_3^{-3} and CO_3^{-2} (2) PO_4^{-3} and CO_3^{-2}
(3) N_2 and CO (4) All of the above **Ans. (1)**

Solution : BO_3^{-3} CO_3^{-2}

$5 + 24 + 3 = 32$ $6 + 24 + 2 = 32$

Example - 11 The pair $NH_3 + BH_3$ is isoelectronic with

(1) B_2H_6 (2) C_2H_6
(3) C_2H_4 (4) CO_2 **Ans. (2)**

Solution : $NH_3 + BH_3$ C_2H_6

$7 + 3 + 5 + 3 = 18$ $6 \times 2 + 6 = 18$

Example - 12 Which of the following is a one-electron species ?

(1) He (2) N
(3) H_2 (4) N_2 **Ans. (4)**

Solution : There is only one electron in H_2^+

Example - 13 The molecular weight of an oxide of nitrogen is 30. What should be the number of electrons in it ?

(1) 15 (2) 30
(3) 45 (4) 20 **Ans. (1)**

Solution : The molecular weight of NO is 30. It will have 15 electrons.

Example - 14 A diapositive ion has 16 protons. What should be the number of electrons in its tetrapositive ion.

(1)16 (2) 14
(3) 12 (4) 10 **Ans. (3)**

Solution : X^{+2} has 16 protons, then In X – 16 protons and 16 electrons

In X^{+2} – 16 protons and 14 electrons In X^{+4} – 16 protons and 12 electrons

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Example - 15 If atomic weights of C and Si are 12 and 28 respectively, then what is the ratio of numbers of neutrons in them

(1) 1 : 2 (2) 2 : 3
(3) 3 : 4 (4) 3 : 7 **Ans. (4)**

Solution : Number of neutrons in $_6C^{12}$ = 12 – 6 = 6; Number of neutrons in $_{14}Si^{28}$ = 28 – 14 = 14

The ratio of number of neutrons in C and Si is 6 : 14 or 3 : 7.

3.4 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) Velocity (c or v)

The above said properties of a wave have the following relationship

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

3.4.1 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).

3.4.2 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.

3.4.3 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 = 3×10^8 m/second or 3×10^{10} cm/second.

3.4.4 Amplitude

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

3.4.5 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}

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

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Example - 16 What should be the wavelength of an ultraviolet wave, if its frequency is 12×10^{16} cycles per seond and $c = 3 \times 10^8$ m/second ?

(1) 25 Å (2) 2.5 Å

(3) 0.25 Å (4) 0.025 Å **Ans. (1)**

Solution : $\because c = \nu\lambda$

Therefore $\lambda = \dfrac{c}{\nu} = \dfrac{3\times10^8}{12\times10^{16}} = 0.25 \times 10^{-8}$ m

or $\lambda = 2.5 \times 10^{-9}$ m or 25×10^{-10} m

$\lambda = 25$ Å

3.5 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 two types viz.

(i) Emission spectrum
(ii) Absorption spectrum and

(i) 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.

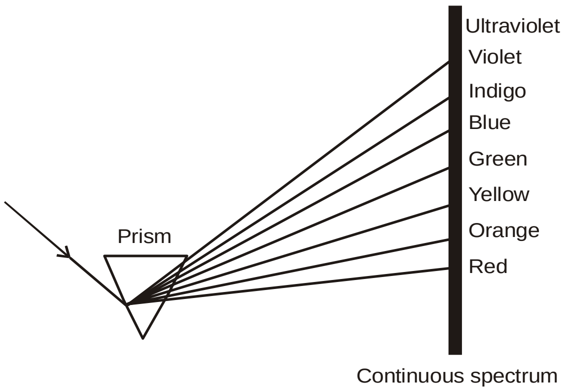
(c) By passing electric discharge through a metallic filament.

Emission spectra are of the following two types.

(X) Continuous spectrum and (Y) line spectrum or atomic spectrum

(X) 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.



Continuous spectrum of white light

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(Y) 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.

(ii) Absorption Spectrum

When white light emitted by glowing heat fluorescent substance is passed through another substance like sodium. 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.

3.5.1 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.

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

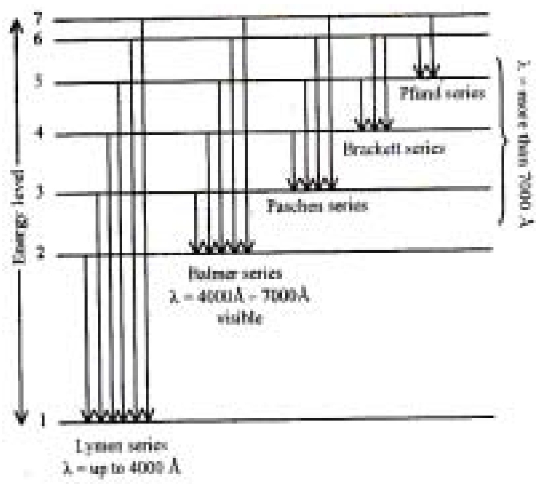
3.5.2 Series of Lines in Hydrogen Spectrum

(i) Lyman Series

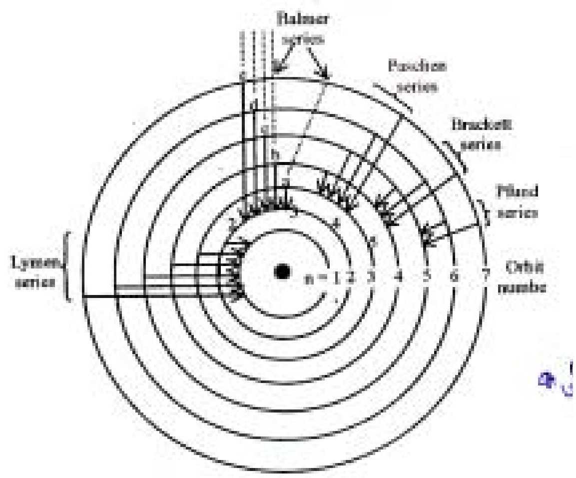
When an electron undergoes transition from a higher energy level (n_2), e.g. 2, 3, 4, 5, ... ∞ to ground state or lower energy level, the spectrum is said to belong to Lymen series. For this, $n_1 = 1$ and $n_2 = 2$, 3, 4, 5, 6, 7, 8 .... ∞.

(ii) 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

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(iii) 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 $n_2 = 4, 5, 6, 7, 8 \ldots\ldots\ldots \infty$.

(iv) 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 $n_2 = 5, 6, 7, 8 \ldots\ldots \infty$.

(v) 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 $n_2 = 6, 7, 8, 9, 10 \ldots\ldots\ldots \infty$.

(vi) Humphrey Series

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

S.No.	Series of lines	n_1	n_2	Spectral region	Wavelength
1.	Lyman 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 expression $\frac{n(n-1)}{2}$.

Example - 17 How many emission spectral lines in all should be visible, if an electron is present in the third orbit of hydrogen atom ?

(1) 6 (2) 3

(3) 5 (4) 15 **Ans. (2)**

Solution : The expression of maximum number is $\frac{n(n-1)}{2} = \frac{3(3-1)}{2} = \frac{6}{2} = 3$

Example - 18 Which of the following should be the expression for the last line of Paschen series ?

(1) $\frac{1}{\lambda}=R\left(\frac{1}{9}-\frac{1}{\infty^2}\right)$ (2) $\frac{1}{\lambda}=R\left(\frac{1}{4}-\frac{1}{9}\right)$

(3) $\frac{1}{\lambda}=R\left(\frac{1}{9}-\frac{1}{16}\right)$ (4) $\frac{1}{\lambda}=R\left(\frac{1}{16}-\frac{1}{\infty}\right)$ **Ans. (1)**

Solution : $\bar{\nu} = \frac{1}{\lambda} = R\left(\frac{1}{9}-\frac{1}{\infty}\right)$

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3.6 PLANCK'S QUANTUM THEORY OF RADIATION

In order to explain the spectral distribution of energy radiated by a back body. Max Planck (1901) put forward quantum theory of radiation. According to this theory.

(i) Radiant energy is emitted or absorbed discontinuously in the form of tiny bundles or packets known as quanta or photon.

(ii) The energy of a quantum is directly proportional to the frequency of the radiation

$$E \propto v$$
$$E = hv$$

v = frequency of radiation

h = planck's constant

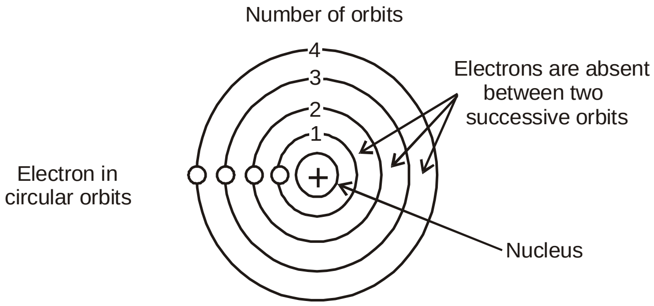
(iii) A body can emit or absorb energy only in whole number multiples of quantum i.e 1hv, 2hv, etc.

Energy in fractions of a quantum cannot be lost or absorbed. this is known as quantization of energy.

3.7 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.

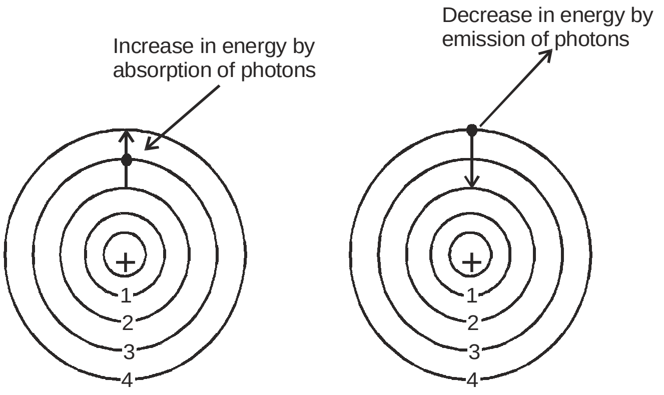


Electrons moving around in circular orbits

(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 electron 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.



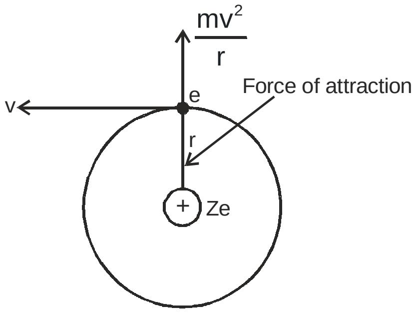
$$E_{n_2} - E_{n_1} = \Delta E = h\nu = \frac{hc}{\lambda}$$

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

$$\lambda \propto \frac{1}{\Delta E}$$

(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

$$\frac{mv^2}{r} = \frac{1}{4\pi\varepsilon_0}\cdot\frac{Ze^2}{r^2}$$

$$mv^2 = \frac{1}{4\pi\varepsilon_0}\cdot\frac{Ze^2}{r}$$

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)

3.7.1 Calculation of Velocity of the Electron of Bohr's Orbit

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

From Bohr's postulate

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

Eq. (1) divided by (2)

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

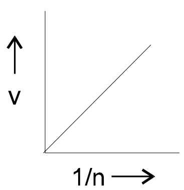
Here π , e and h are constants, therefore

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

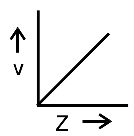
$$\boxed{v \propto \frac{1}{n}}$$



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

Thus

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



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(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 $T = \dfrac{2\pi r}{V} = \dfrac{2\pi \times n^2h^2}{4\pi^2 mZe^2} \times \dfrac{nh}{2\pi Ze^2} = \dfrac{n^3h^3}{4\pi^2 mZ^2e^4}$

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

3.7.2 Radius of nth Bohr's Orbit

According to Bohr's hypothesis

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

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

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

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

or $\boxed{r = \dfrac{n^2}{Z} \times 0.529 Å}$

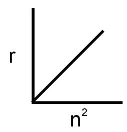
Note : (a) 1Å = 10^{-8} cm (b) 1Å = 10^{-10} m
(c) 1 nm = 10^{-9} m (d) 1 pm (picometer) = 10^{-10}cm

(a) If Z is a constant, then

$r \propto n^2$

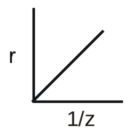
Thus, the radius of atoms goes on increasing as the number (n) of energy levels in the atoms goes on increasing as shown below.

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



(b) If n is a constant, then

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



Example - 19 An electron has been excited from the first to the fourth energy state in an atom. Which of the following transitions are possible when the electron comes back to the ground state ?

(1) $4 \rightarrow 1$
(2) $4 \rightarrow 2$, $2 \rightarrow 1$
(3) $4 \rightarrow 3$, $3 \rightarrow 2$, $2 \rightarrow 1$
(4) All of the above **Ans. (4)**

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Solution : Electron can undergo transition from higher state to all lower states by loss of energy.

Example - 20 How much total energy will be released when an electron present in hydrogen atom undergoes the following sequence of transition ?

$$n = 4 \rightarrow n = 2 \rightarrow n = 1$$

(1) One quantum (2) Two quantums

(3) Three quantums (4) Four quantums

Ans. (2)

Solution : One quantum of energy is released in each transition, i.e. one quantum in n = 4 to n = 2 and one quantum in n = 2 to n = 1 transition.

Example - 21 Which of the following is a fundamental particle

(1) Nucleus of He (2) Nucleus of H

(3) A positive atom (4) None of these **Ans. (2)**

Solution : Fundamental particle H^+ is the nucleus of H

Example - 22 The expression for calculation of velocity is

(1) $v = \left(\frac{Ze^2}{mr}\right)^{\frac{1}{2}}$ (2) $v = \frac{2\pi Ze^2}{nh}$

(3) $v = \frac{nh}{2\pi mr}$ (4) all of the above are correct **Ans. (4)**

Solution : (1) $v^2 = \frac{Ze^2}{mr}$ $v = \left(\frac{Ze^2}{mr}\right)^{\frac{1}{2}}$ (2) $v = \frac{2\pi Ze^2}{nh}$

(3) $mvr = \frac{nh}{2\pi}$ $v = \frac{nh}{2\pi mr}$

Example - 23 If the velocities of first, second, third and fourth orbits of hydrogen atom are v_1 , v_2 , v_3 and v_4 respectively, then which of the following should be their increasing order

(1) $v_1 > v_2 > v_3 > v_4$ (2) $v_4 < v_3 < v_2 < v_1$

(3) $v_1 > v_2 < v_3 > v_4$ (4) Equal for all **Ans. (2)**

Solution : Z is a constant, therefore $v \propto \frac{1}{n}$

i.e. $v_4 < v_3 < v_2 < v_1$

Example - 24 If the radius of first, second, third and fourth orbits of hydrogen atom are r_1, r_2, r_3 and r_4 respectively, then their correct increasing order will be

(1) $r_4 < r_3 < r_2 < r_1$ (2) $r_1 < r_2 < r_3 < r_4$

(3) $r_1 > r_2 > r_3 > r_4$ (4) Equal in all **Ans. (2)**

Solution : $r \propto n^2$

$\boxed{r_1 < r_2 < r_3 < r_4}$

Example - 25 The ratio of radius of the fifth orbits of He^+ and Li^+ will be

(1) 2 : 3 (2) 3 : 2

(3) 4 : 1 (4) 5 : 3 **Ans. (2)**

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Solution : Here n is a constant, therefore

$$\frac{r_1}{r_2} = \frac{Z_2}{Z_1} = \frac{3}{2} = 3:2$$

Example - 26 Which of the following orbits of hydrogen atom should have the values of their radii in the ratio of 1 : 4 ?

(1) K and L (2) L and N
(3) M and N (4) 1 and 2 both are correct **Ans. (4)**

Solution : (1) Ratio of radii of orbits K and L

$$\frac{r_1}{r_2} = \frac{n_1^2}{n_2^2} = \frac{1^2}{2^2} = 1:4$$

(2) Ratio of radii of orbits L and N

$$\frac{r_1}{r_2} = \frac{n_1^2}{n_2^2} = \frac{2^2}{4^2} = 4:16 \text{ or } 1:4$$

3.7.3 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.

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

(ii) 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 = $\frac{-Ze^2}{r}$

(iii) Total Energy (E_T)

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

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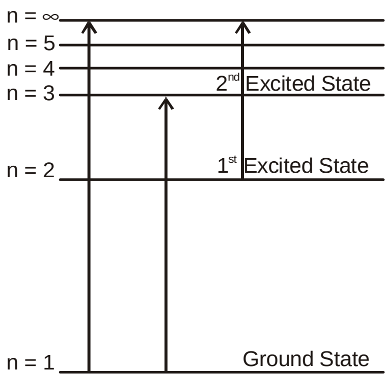
Total energy $E = -\frac{1}{2}\frac{Ze^2}{r}$

Formula

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

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

(iv) 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

(V) Ground state

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

(vi) 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.

(vii) Excitation Potential

(a) The energy required to excite an electron from ground state to any 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 = $E_4 - E_1$

(viii) 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.

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(ix) Separation Energy

The energy required to separate an electron from any 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.

3.7.4 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

$$E_{n_1} = -\frac{1}{2}\frac{Ze^2}{r_1} \quad (r_1 = \text{radius of the first orbit})$$

$$E_{n_2} = -\frac{1}{2}\frac{Ze^2}{r_2} \quad (r_2 = \text{radius of the second orbit})$$

$$E_{n_1} - E_{n_2} = -\frac{1}{2} Ze^2 \left(\frac{1}{r_1} - \frac{1}{r_2}\right)$$

According to Bohr hypothesis

$$E_{n_1} - E_{n_2} = h\nu;\ E_{n_1} - E_{n_2} = -h\nu$$

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

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

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

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

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

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

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Example - 27 What should be the kinetic energy and total energy of the electron present in hydrogen atom, if its potential energy is –5.02 eV

Solution : (a) Total energy = $\dfrac{\text{Potential energy}}{2} = \dfrac{-5.02}{2}$

Total energy = – 2.51 eV

(b) Kinetic energy = – Total energy = – (–2.51 eV) = + (2.51 eV)

Example - 28 What should be the order E_1, E_2, E_3 and E_4, if these are the respective energies of the first, second, third and fourth orbits of hydrogen atom ?

(1) $E_1 = E_2 = E_3 = E_4$ (2) $E_4 < E_3 < E_2 < E_1$

(3) $E_1 < E_2 < E_3 < E_4$ (4) $E_2 > E_3 < E_4 < E_1$ **Ans. (3)**

Solution : $E \propto -\dfrac{1}{n^2}$

Example - 29 What should be kinetic energy and potential energy, respectively, of the electron in the third orbit of hydrogen atom ?

(1) – 1.5 eV, 3.0 eV (2) 1.5 eV, – 3.0 eV

(3) 1.5 eV, 3.0 eV (4) 3.0 eV, – 3.0 eV **Ans. (2)**

Solution : Total energy of the third orbit of H atom

$$E = -\frac{Z^2}{n^2} \times 13.6 = -\frac{1}{9} \times 13.6 = -1.5\text{eV}$$

(1) Kinetic energy = – Total energy = – (–1.5) = + 1.5 eV

(2) Potential energy = 2 × Total energy = 2 × –1.5 = –3.0 eV

Example - 30 Which of the following should be the energy of an electron present in ground state of hydrogen atom ?

(1) – 13.6 eV (2) – 3.4 eV

(3) – 1.5 eV (4) – 0.85 eV **Ans. (1)**

Solution : An electron in ground state is in n = 1 orbit. Therefore the energy of the electron = – 13.6 eV

Example - 31 What should be the energy of the second excited state of Li^{+2} ?

(1) – 13.6 eV (2) – 30.6 eV

(3) – 3.4 eV (4) – 1.5 eV **Ans. (1)**

Solution : Second excited state n = 3

$$E_n = -13.6 \times \frac{Z^2}{n^2} = -13.6 \times \frac{3^2}{3^2} = -13.6 \text{ eV}$$

Example - 32 How much minimum energy should be absorbed by a hydrogen atom in ground state to reach excited state ?

(1) + 10.2 eV (2) + 13.4 eV

(3) + 3.4 eV (4) + 1.5 eV **Ans. (1)**

Solution : The electron has to go to the second orbit E_2 on excitation. Therefore

$E_2 - E_1 = -3.4 - (-13.6) = 13.6 - 3.4 = 10.2$ eV

Example - 33 The maximum energy absorbed by hydrogen atom in its ground state will be

(1) 13.6 eV (2) 3.4 eV

(3) 10.2 eV (4) 0 eV **Ans. (1)**

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Solution : $E_{\infty} - E_1$

$0 - (-13.6) = 13.6$ eV

Example - 34 The energy required in the process $He^{+2} \rightarrow He^{+3}$ will be

(1) 0 eV (2) + 13.6 eV
(3) + 3.4 eV (4) + 1.5 eV **Ans. (1)**

Solution : He^{+2} does not have any electron, therefore the ionisation energy will be 0.

Example - 35 What should be the value of wave number of emitted radiation with respect to R, when the electron present in hydrogen atom jumps from M orbit to K orbit ?

(1) $R \times \frac{8}{9}$ (2) $R \times \frac{5}{8}$

(3) $R \times \frac{3}{4}$ (4) $R \times \frac{5}{16}$ **Ans. (1)**

Solution : The electron jumps from M orbit (n = 3) to K orbit (n = 1). Therefore

$$\bar{\nu} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) = R\left(\frac{1}{1} - \frac{1}{3^2}\right) = R\left(\frac{1}{1} - \frac{1}{9}\right)$$

$$= R\left(\frac{9-1}{9}\right) = R \times \frac{8}{9}$$

3.7.5 Limitation of Bohr's Model

(a) It does not explain the spectra of multi-electron atoms.

(b) When a high resolving power spectroscope is used, it is observed that a spectral line in the hydrogen spectrum is not a simple line but a collection of several lines which are very close to one another. This is known as fine spectrum. Bohr's theory does not explain the fine spectra of even hydrogen atom.

(c) It does not explain the splitting of spectra lines into a group of finer lines under the influence of magnetic field (Zeeman effect) and electric field (Stark effect).

(d) Bohr's theory is not in agreement with Heisenberg's uncertainity principle.

4. PHOTOELECTRIC EFFECT

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

Work function or threshold energy may be defined as the minimum amount of energy required to eject electrons from a metal surface.

According to Einstien,

Maximum kinetic energy of the ejected electron = absorbed energy – work function

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

where v_0 and λ_0 are threshold frequency and threshold wavelength.

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4.1 LAWS OF PHOTOELECTRIC EFFECT

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

(ii) The maximum kinetic energy of photoelectrons is directly proportional to the frequency of incident radiation; moreover, it is independent of the intensity of light used.

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

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

Example - 36 Photo - electrons are liberated by ultraviolet light of wavelength 3000Å from a metallic surface for which the photoelectric threshold is 4000Å. Calculate de-Broglie wavelength of electrons emitted with maximum kinetic energy

(1) 1.2×10^{-9} m (2) 1.4×10^{-9} m

(3) 1.6×10^{-10} m (4) 1.2×10^{-12} m **Ans. (1)**

Solution: As, $\nu = h\nu^\circ + KE \Rightarrow \frac{hc}{\lambda} = \frac{hc}{\lambda^o} + K.E$

$\therefore$ $E. = hc\left(\frac{1}{\lambda} - \frac{1}{\lambda^o}\right) = \frac{hc(\lambda - \lambda^o)}{\lambda \times \lambda} = \frac{6.625\times10^{-34}\times3\times10^{8}(4000\times10^{-10} - 3000\times10^{-10})}{3000\times10^{-10}\times4000\times10^{-10}}$

6565×10^{-19} Joule

Also $\frac{1}{2}$ $mv^2 = 1.6565 \times 10^{-19}$

$\Rightarrow$ $m^2v^2 = 2 \times 1.65665 \times 10^{-19} \times m$

$\Rightarrow$ $v^2 = 2 \times 1.6565 \times 10^{-19} \times 9.109 \times 10^{-31} \Rightarrow m\nu = 5.49 \times 10^{-25}$

$= \frac{h}{mv} = \frac{6.625\times10^{-34}}{5.49\times10^{-25}} = 1.2 \times 10^{-9}$ m

5. PARTICLE AND WAVE NATURE OF ELECTRON

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

de Broglie derived an expression for calculating the wavelength of the wave associated with the electron.

According to Planck's equation

$$E = hv = h \cdot \frac{c}{\lambda} \quad \text{...(i)}$$

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

$$E = mc^2 \quad \text{...(ii)}$$

where c is the velocity of the electron.

Equating both the equations, we get

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

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

Momentum of the moving electron is inversely proportional to its wavelength.

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Let kinetic energy of the particle of mass ‘m’ is E.

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

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

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

$$\lambda = \frac{h}{p} = \frac{h}{\sqrt{2Em}}$$

where c is the velocity of the electron.
Equating both the equations, we get

5.1 DAVISSON AND GERMER MADE THE FOLLOWING MODIFICATION IN DE BROGLIE EQUATION

Let a charged particle say an electron be accelerated with a potential of V; then the kinetic energy may be given as :

$$\tfrac{1}{2}mv^2 = eV$$

$$m^2v^2 = 2eVm$$

$$mv = \sqrt{2eVm} = p$$

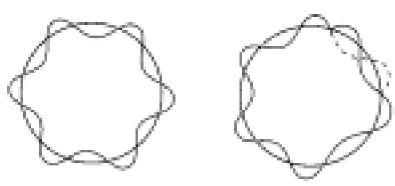
$$\lambda = \frac{h}{\sqrt{2eVm}}$$

de Broglie waves are not radiated into space, i.e. they are always associated with electron.The wavelength decreases if the value of mass (m) increases, i.e., in the case of heavier particles, the wavelength is too small to be measured. de Broglie equation is applicable in the case of smaller particles like electron and has no significance for larger particles.

5.2 BOHR’S THEORY VERSUS DE BROGLIE EQUATION

One of the postulates of Bohr’s theory is that angular momentum of an electron is an integral multiple of $\frac{h}{2\pi}$. This postulate can be derived with the help of de Broglie concept of wave nature of electron.

Consider an electron moving in a circular orbit around nucleus. The wave train would be associated with the circular orbit as shown in figure.



(a) (b)

(a) The number of waves around the orbit
is a whole number (6) times the wavelength.
(b) The number of waves around the orbit is not
a whole number multiple of the wavelength.
When the wave goes around the orbit many times,
it cancles itself out.

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If the two ends of the electron wave meet to give a regular series of crests and troughs, the electron wave is said to be in phase, i.e., the circumference of Bohr's orbit is equal to whole number multiple of the wavelength of the electron wave.

So, $$2\pi r = n\lambda$$

or $$\lambda = \frac{2\pi r}{n}$$

From de Broglie equation,

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

Thus, $$\frac{h}{mv} = \frac{2\pi r}{n}$$

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

[v = velocity of electron and r = radii of the orbit]

i.e., Angular momentum = $n \cdot \frac{h}{2\pi}$

This proves that the de Broglie and Bohr–concepts are in perfect agreement with each other.

Example - 37 What is the mass of a photon of sodium light with a wavelength of 5890 Å.

(1) 6.552×10^{-30} g (2) 5.892×10^{-30} g

(3) 9.582×10^{-28} g (4) 3.752×10^{-33} g **Ans. (4)**

Solution: $$\lambda = \frac{h}{mc}$$

or $$m = \frac{h}{\lambda c}$$

So, $$m = \frac{6.63 \times 10^{-27}}{5890 \times 10^{-8} \times 3 \times 10^{10}} = 3.752 \times 10^{-33} \text{ g}$$

Example - 38 What should be the mass of the photon of sodium light if its wavelength is 5894 Å, the velocity of light is 3×10^8 metre/second and the value of h is 6.6252×10^{-34} Kg m^2/sec ?

(1) 3.746×10^{-26} (2) 3.746×10^{-30}

(3) 3.746×10^{-34} (4) 3.746×10^{-36} **Ans. (4)**

Solution : $\lambda = \frac{h}{m \times c}$ or $\frac{h}{c\lambda}$

($\because \lambda = 5894$ Å or 5894×10^{-10} m)

$$m = \frac{6.652 \times 10^{-34}}{3 \times 10^8 \times 5894 \times 10^{-10}} \text{ or } \frac{6.652}{17682} \times 10^{-32} = 3.746 \times 10^{-36} \text{ Kg}$$

6. 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.

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

Here Δx is uncertainty of position,

Δp is uncertainty of momentum and

h is Planck's constant

Example - 39 What should be the uncertainty in position if uncertainty in momentum is 1×10^{-2} gm cm/sec and value of h is 6.6253×10^{-34} Js ?

(1) 1.054×10^{-22}m (2) 1.054×10^{-25}m

(3) 1.054×10^{-27} m (4) 1.054×10^{-32} m **Ans. (3)**

Solution : Given that

$\Delta p = 1 \times 10^{-2}$ gm cm/sec. $= 1 \times 10^{-7}$ Kg m/sec.

$h = 6.6252 \times 10^{-34}$ Js

$$\Delta x \times \Delta p = \frac{h}{2\pi} \qquad \therefore \Delta x = \frac{h}{2\pi \times \Delta p}$$

$$\text{or} \quad \Delta x = \frac{6.6252 \times 10^{-34}}{2 \times 3.14 \times 10^{-7}}$$

m $= 1.054 \times 10^{-27}$m

7. 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

8. 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.

8.1 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 gives 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.

8.2 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

8.3 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.

(1) 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.

(2) 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.

(3) 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.

8.4 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) 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 μ_s. 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 $-\frac{1}{2}$.

Example - 40 Which of the following is the principal quantum number for the last electron of $_{11}Na$?

(1) 3 (2) 2

(3) 4 (4) 1 **Ans. (1)**

Solution : $_{11}Na = 1s^2, 2s^2, 2p^6, 3s^1$

n = 3

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Example - 41 Which of the following should have greater size ?

(1) 1s (2) 2s
(3) 3s (4) 4s **Ans. (4)**

Solution : n = 4 for 4s

Example - 42 Which of the following should be the possible sub-shells, for $n + \ell = 7$?

(1) 7s, 6p, 5d, 4f (2) 4f, 5p, 6s, 4d
(3) 7s, 6p, 5d, 6d (4) 4s, 5d, 6p, 7s **Ans. (1)**

Solution : $n + l = 7$

7 + 0 = 7s; 6 + 1 = 6p; 5 + 2 = 5d; 4 + 3 = 4f

Example - 43 What should be the maximum number of electron in the possible sub-shells, for $n + l = 4$?

(1) 8 (2) 6
(3) 12 (4) 16 **Ans. (1)**

Solution : $n + l = 4$ Maximum number of electrons

4 + 0 = 4s $\rightarrow$ 2
3 + 1 = 3p $\rightarrow$ 6
8

Example - 44 The sub-shell 2d is not possible because

(1) $n \neq l$ (2) $l > n$
(3) $n < l$ (4) None of these **Ans. (1)**

Solution : For sub-shell 2d, n = 2 and l = 2 and the values of n and l can never be equal.

Example - 45 What should be the maximum number of elements, if the elements above n = 4 do not exist in nature ?

(1) 40 (2) 60
(3) 44 (4) 108 **Ans. (2)**

Solution : Since, n = 1, 2, 3 and 4, therefore $\frac{1s}{2}$, $\frac{2s, 2p}{8}$, $\frac{3s, 3p, 3d}{18}$, $\frac{4s, 4p, 4d, 4f}{32}$

Thus, total number of existent elements = 2 + 8 + 18 + 32 = 60

Example - 46 The orbital having n = 6, l = 2 and m = 0 will be designated as

(1) $6d_{z^2}$ (2) $6d_{x^2-y^2}$
(3) $6d_{xy}$ (4) $6p_z$ **Ans. (1)**

Solution : For 6th of P energy level, l = 2 is for d sub-level, and m = 0 for d_{z^2} orbital

Example - 47 The orbital having n = 2, l = 1 and m = 0 is designated as

(1) $2p_z$ (2) $2p_x$
(3) $2p_y$ (4) $3d_{z^2}$ **Ans. (1)**

Solution : In the second or L energy level (n = 2) , l = 1 for p orbital, m = 0 for z axis, Hence, the orbital will be designated as $2p_z$.

Example - 48 If x is the number of electron in an atom, the configuration should be expressed as :

(1) l_x (2) nl^x
(3) nm^x (4) None of these **Ans. (2)**

Solution : The electronic configuration of an atom is expressed by first writing principal quantum number (n), followed by azimuthal quantum number (l) and then writing number of electrons (x) as superscript.

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Example - 49 What should be the atomic number of an element, if the quantum numbers of the highest energy electron of the element in ground state are n = 4, l = 1, m = – 1 s = + 1/2 ?

(1) 31 (2) 35

(3) 30 (4) 32 **Ans. (1)**

Solution : The electronic configuration of the element will be $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $4s^2$, $3d^{10}$, $4p^1$. Thus, the total number of electrons is 31 and hence the atomic number will be 31.

Example - 50 The orbital having m = – 2 should not be present in the following sub-shell

(1) d (2) f

(3) g (4) p **Ans. (4)**

Solution : For p sub-shell, m = – 1, 0, + 1. Therefore, m = – 2 orbital will not be present in p sub-shell.

Example - 51 What should be the value of spin quantum number of the last electron for d^9 configuration ?

(1) 0 (2) – 1/2

(3) 1/2 (4) 1 **Ans. (2)**

Solution : The value of spin quantum number (s) can be + 1/2 or – 1/2, because an electron can rotate along its axis either in clockwise or in anticlockwise direction. But one quantum number depicts one electrons and thus its value will be – 1/2 for d^9 configuration.

Example - 52 The all energy levels are called excited states when the value of principal quantum number is :

(1) n = 1 (2) n > 1

(3) n < 1 (4) n > – 1 **Ans. (2)**

Solution : All the energy states in which n is greater than 1 are called excited states.

9. RULES FOR FILLING ELECTRONS

9.1 AUFBAU PRINCIPLE

Aufbau is a German word that means building 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.

9.2 PAULI'S EXCLUSION PRINCIPLE

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

For Examples

(a)	$6s^1$	and	$6s^2$
	n = 6		n = 6
	l = 0		l = 0
	m = 0		m = 0
	s = + 1/2		s = – 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 respective 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)

Examples

$1s^1$ 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.

9.3 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) $\boxed{\uparrow\downarrow}\quad \boxed{\uparrow\downarrow}\boxed{\ \ }\boxed{\ \ }$

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

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

(1) The orbitals should be degenerate

(2) The number 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 $_5B$ also, because there is only one electrons in p orbital. Therefore, this rule is applicable from $_6C$ onwards.

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

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9.4 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) If 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) $_{57}La – 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6, 6s^2, 5d^1$

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

Specific Electronic Configuration

	Element	Atomic number	Expected Configuration	Atcual Configuration
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.

9.5 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.

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(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.

Example - 53 Pauli exclusion principle applies to

(1) H (2) H^+

(3) H^- (4) None of the above **Ans. (3)**

Solution : Since, H has one electron and H^+ has no electron, therefore Paulic principal does not apply to them. However, H^- has two electrons, hence this principle applies on it.

Example - 54 Which of the following statements is true ?

(1) One orbit can accommodate a maximum of two electrons

(2) One sub-shell can accommodate a maximum of two electrons

(3) One orbital can accommodate a maximum of two electrons

(4) None of the above **Ans. (3)**

Solution : It is an orbital that can accommodate a maximum of two electrons having opposite spins

Example - 55 Which of the following is not according to Pauli exclusion principle ?

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

(3) $\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{}$ (4) 1 and 2 both **Ans. (4)**

Solution : The set of four quantum numbers are not same for the three electrons in answer 3. In answer 1 both of the electrons have same set of quantum numbers, while in answer 2 the first and third electrons have same set of quantum numbers.

Example - 56 Supposing that Pauli exclusion principle is not correct, if one orbital can accommodate three electrons, what should be the respective atomic number of the second member of alkali metal family and the first member of halogen family

(1) 16, 14 (2) 11, 9

(3) 16, 9 (4) 34, 17 **Ans. (1)**

Solution : (a) Sodium is the second member of alkali metal family

$Na^{11} = 1s^2, 2s^2, 2p^6, 3s^1$

We know that the inner orbitals of sodium are fully filled and the outer most orbit has one electron. If inner orbitals can accommodate three electrons each, the configuration will be as follows.

$$1s^3, 2s^3, 2p^9, 3s^1$$

Therefore, three will be 16 electrons in all. Hence the atomic number will be 16.

(b) The first member of halogen family is fluorine, F^9 whose configuration is $1s^2, 2s^2, 2p^5$

Halogen has one electron less than the next inert of noble gas. If inner orbitals can accommodate three electron each, the configuration will be as follows :

$$1s^3, 2s^3, 2p^8$$

Therefore, total number of electrons will be 14 and thus the atomic number will also be 14.

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Example - 57 Which of the following should be correct according to Hund's rule ?

(1) $_6C = 1s^2, 2s^2$ $\boxed{\uparrow\uparrow}\boxed{\ }\boxed{\ }$ (2) $_8O = 1s^2, 2s^2$ $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\ }$

(3) $_7N = 1s^2, 2s^2,$ $\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\ }$ (4) $_9F = 1s^2, 2s^2$ $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}$ **Ans. (4)**

Solution: Configuration of $_6C$ should be $2p_x^1\ 2p_y^1$ instead of $2p_x^2$

Configuration of $_8O$ should be $2p_x^2\ 2p_y^1\ 2p_z^1$ instead of $2p_x^2\ 2p_y^2$

Configuration of $_7N$ should be $2p_x^1\ 2p_y^1\ 2p_z^1$ instead of $2p_x^2\ 2p_y^1$

Configuration of $_9F$ $2p_x^2\ 2p_y^2\ 2p_z^1$ is correct because two out of the three degenerate p orbtials are fully-filled, one is half-filled and there is no unfilled p orbital.

Example - 58 Which of the following should be the basis of entry of an electron in 4s orbital before 3d orbital?

(1) Energy level diagram (2) Hund's rule

(3) Pauli's principle (4) Shielding constant **Ans. (1)**

Solution : n + l of 4s = 4 + 0 = 4 and that of 3d is 3 + 2 = 5. Therefore, energy of 4s is lower than that of 3d.

Example - 59 Which of the following should be the number of electrons present in X^{+2} on the basis of electronic configuration, if the ion X^{-3} has 14 protons ?

(1) 12 (2) 14

(3) 16 (4) 18 **Ans. (1)**

Solution : X^{-3} has 14 protons, i.e. X also has 14 protons and therefore 14 electrons.

$$X = 14 = 1s^2, 2s^2\ 2p^6, 3s^3\ 3p^2$$

$$X^{+2} = 1s^2, 2s^2\ 2p^6, 3s^2 = 12 \text{ electrons}$$

Example - 60 Which of the following should be the electronic configuration of P in H_3PO_4 ?

(1) [Ne] (2) [Ne] $3s^2\ 3p^6$

(3) [Ne] $3s^2$ (4) [Ne] $3s^2\ 3p^1$ **Ans. (2)**

10. ORBITAL

(a) The space around the nucleus where probability of finding an electron is maximum, is called an orbital.

(b) An electron cloud is negatively charged and the nucleus is positively charged. Therefore, the probability of finding an electron is maximum around the nucleus.

(c) The probability of finding an electron in an orbital is 95% to 98%.

(d) he place where probability of finding an electron is zero is known as node and a plane passing through node is known as nodal plane.

(e) Total number of nodal planes = l.

(f) There are two types of nodes :

(i) Radial node : These are the points at some distance from the nucleus where there is zero probability of finding the elecetrons.

(ii) Angular node : These are directional in nature so these are associated with p and d orbitals.

(g) For a particular quantum number 'n'

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Total nodes = n – 1
Radial nodes = n - l - 1
Angular nodes = l

10.1 s-ORBITAL

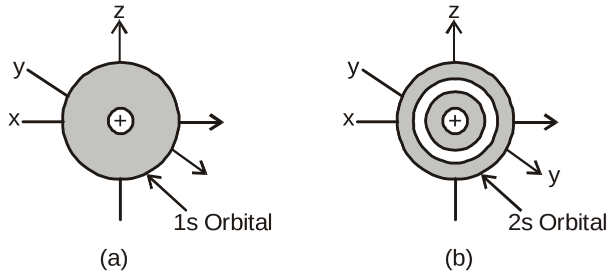
(a) Only one s-orbital is possible in an orbit because l = 0 and m = 0 for it.

(b) It is spherical in shape and thus the electron density is uniform in all directions.

(c) The size increases with increase in the value of n. There is vacant space between 1s orbital and 2s orbital, where the probability of finding electron is minimum, it is known as **nodal surface**.

(d) The nodal surface is missing inside 1s orbital because of its proximity with the nucleus.

(e) The number of nodal surfaces in an orbit is equal to (n – 1)



The shapes of s-orbital

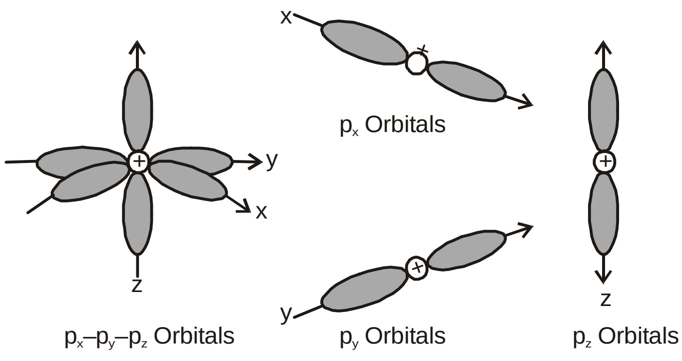
10.2 p-ORBITAL

(a) For p orbitals, l = 1 and m = – 1, 0, + 1. Thus, it can have three configurations, which are distributed in x, y and z axes. Therefore, there are three p-orbitals, which are dumbbell, shaped.

(b) Each p-orbital has two lobes and the probability of finding electron inside these two lobes is equal. The plane perpendicular to the axis two lobes and passing through the point where these two lobes join, is the nodal plane of p-orbital, because the probability of finding electron in this plane is negligible or minimum.

(c) The value of nodal planes for each of the p_x, p_y and p_z orbitals is same and these nodal planes are present in xy, yz and xz planes, respectively.

(d) The three p-orbitals of a particular orbit (p_x, p_y and p_z) have equal energy and therefore these are called degenerated orbitals.



Shape of p orbitals

10.3 d-ORBITALS

(a) For d orbitals, l = 2 and m = – 2, – 1, 0, +1, + 2. Therefore, there are five orientations and thus five d-orbitals.

(b) Its shape is like a double dumbbell.

(c) The five orientations of d-orbitals are as follows :

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(1) The double dumbbell of d_{xy} orbital are situated between x and y axes.

(2) The double dumbbell of d_{yz} orbital are situated between y and z axes

(3) The double dumbbell of d_{xz} orbital are situated between x and z axes.

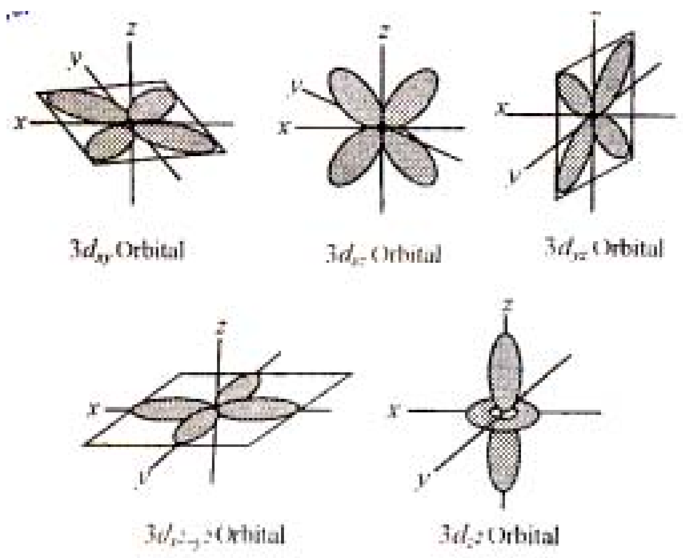
(4) The double dumbbell of $d_{x^2-y^2}$ orbital are directed on x and y axes

(5) d_{z^2} orbital is composed of one dumbbell and one ring. The dumbbell is situated on z axis and the ring is present on its middle part.

Therefore on the above said five d orbtials can be classified into the following two categories.

(a) t_{2g} Orbitals (d_{xy}, d_{xz} and d_{yz}) – In these, the electron density is concentrated in-between the axes. These are also called grade orbitals.

(b) e_g Orbitals ($d_{x^2-y^2}$ and d_{z^2}) – In these, the electron density is concentrated on the axes.



Shape of d Orbitals

10.4 f-ORBITALS

(i) They have complex shapes,

(ii) For these, l = 3 and m = – 3, –2, –1, 0, +1, +2, +3

(iii) These have seven orientations.

10.5 DIFFERENCE BETWEEN ORBIT AND ORBITAL

S.No.	Orbit	Orbital
1.	It is depicted by n.	It is depicted 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

11. CONTRIBUTION OF SOME SCIENTISTS

(1) Atom

John Dalton in 1880 said that matter is composed of very minute and indivisible particles, called atoms.

(2) Electrons

(a) Discoverer – J.J. Thomoson

(b) Weight – $\frac{1}{1837}$ th of the weight of H atom

(c) Mass – 9.1×10^{-28} gram

(d) Amount of charge – (a) 4.8×10^{-10} e.s.u. (electrostatic unit) (b) 1.6012×10^{-19} coulomb

(e) Rest mass – Mass of electron = 9.1×10^{-28} grams

1 a.m.u. = 1.6×10^{-24} gm

$$\text{Rest mass} = \frac{9.1\times10^{-28}}{1.6\times10^{-24}} = 5.51 \times 10^{-4} \text{ a.m.u.}$$

The mass of an electron 9.1×10^{-28} gram is also called its rest mass.

(f) Discovery – With the help of cathode rays.

(3) Proton

(a) Discoverer – Goldstein

(b) Discovery – With the help of anode rays

(c) Mass – 1.6748×10^{-24} gram or 1.00757 a.m.u.

(d) Charge – Unit positive charge

(4) Neutron

(a) Discoverer – James Chadwick (1932)

(b) Charge – Zero (i.e. a neutral particle)

(c) Mass – 1.67×10^{-24} gram or 1.6×10^{-27} Kg

(d) Density – 10^{-12} Kg/cm^3

(5) Nucleus

(a) Discoverer – Rutherford

(b) Size of nucleus – 10^{-13} to 10^{-12} cm i.e. 10^{-15} to 10^{-14} metre

(c) Size of atoms – 10^{-8} cm

(d) Atomic radius – 10^5 × Radius of the nucleus

(6) Positron

(a) Discoverer – C.D. Anderson (1932)

(b) Symbol – e^{+1}

(c) Charge – Unit positive charge

(d) Mass – Negligible (like electron)

(e) Stable particle

(7) Meson

(a) Discoverer – Ukawa

(b) Charge – Positive, negative or zero

(c) Mass – In between proton and electron

(d) Unstable particle

(8) Neutrino

(a) Discoverer – Pauling

(b) Charge – Zero

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(c) **Mass** – Negligible (less than that of electron)

(d) **Stable particle**

(9) **Antiproton**

(a) **Discoverer** – Segre

(b) **Charge** – Unit negative charge

(c) **Mass** – Equal to that of proton

(d) **Stable particle**

(10) **Volume of atom** $= \frac{4}{3}\pi r^3 = \frac{4}{3}\pi\, 10^{-24}$ cm

(11) **Volume of nucleus** $= \frac{4}{3}\pi r^3 = \frac{4}{3}\pi\, 10^{-39}$ cm $= \frac{10^{-39}}{10^{-24}} = 10^{-15}$ cm

Thus, the nucleus of an atom occupies 10^{-15} part of an atom

(12) **Some discoverers**

(a) Positive rays – **Goldstein**

(b) Cathode rays – **William Crookes**

(c) Atomic number – **Mosley**

(d) e/m of proton – **J.J. Thomson**

(e) e/m of electron – **R.A. Milliken**

(f) Radioactivity – **Henry Becquerel**

12. IMPORTANT POINTS

(a) Number of sub-shells in the shell = n

(b) Number of orbitals in the shell = n^2

(c) Number of electrons in the shell = $2n^2$

(d) No. of orbitals in the sub-shell = $2l + 1$

(e) Number of elliptical orbits according to Sommerfeld = n – 1

(f) Maximum number of spectral lines = $\frac{n(n-1)}{2}$

(g) Number of nodal surfaces = n – 1

(h) K. E. = – T.E.

(i) T.E. = P.E./2

(j) **Nodal Point :** The nucleus of an atom called **Nodal Point**.

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

(l) **Isomorphous :** The two different type of compound which contain same crystalline structure called **isomorphous** and this property called **isomorphism**.

(m) Substance which have same number of electron and atoms called **Isosteres**.

(n) **Core :** The outer most shell of an atom called **Core** and the number of electron present to that shell is called **Core electron**.

(o) **Promotion :** The transfer of electron between subshells in an orbit is called promotion. While the transfer of one energy level to another is called transition. After the completion of promotion the transition process is occurred.

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S.No.	Radius	Velocity	Energy	Wavelength
1.	$r = \frac{n^2h^2}{4\pi^2 mZe^2}$	$V = \frac{2\pi Ze^2}{nh}$	$E = -\frac{Ze^2}{2r}$	$R = \frac{2\pi^2 me^4}{Ch^3}$
2.	$r = \frac{n^2}{Z} \times 0.529Å$	$V = \left(\frac{Ze^2}{rm}\right)^{1/2}$	$E = -\frac{2\pi^2 mZ^2e^4}{n^2h^2}$	$\frac{1}{\lambda} = R \times Z^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$
3.	$r = \frac{n^2}{Z} \times 0.0529$ nm	$V = \frac{nh}{2\pi mr}$	$E = -Rch \times \frac{z^2}{n^2}$	$E = h\upsilon$
4.	$r \propto n^2$ (Z const)	$V \propto \frac{1}{n}$ (Z const)	$E = -\frac{z^2}{n^2} \times 313.6$ Kcal	$\lambda = \frac{h}{mc}$
5.	$\frac{r_1}{r_2} = \frac{n_1^2}{n_2^2}$ (Z const)	$\frac{V_1}{V_2} = \frac{n_2}{n_1}$ (Z const)	$E \propto -Z^2$ (n const)	$c = v/t$
6.	$r \propto 1/Z$ (n const)	Time period $T = \frac{2\pi r}{V}$	$\frac{E_1}{E_2} = \frac{Z_1^2}{Z_2^2}$ (n const)	$E = mc^2$

SOLVED EXAMPLES

Example - 1 What difference will appear in the mass number if the number of neutrons is halved number of electrons is doubled in $_8O^{16}$.

(1) 25% decrease (2) 50% increase

(3) 150% increase (4) No difference **Ans.(1)**

Solution : $_8O^{16}$ Initial weight – final weight

Protons	8p	$\rightarrow$	8p
Neutrons	8n	$\rightarrow$	4n
Weight	16	$\rightarrow$	12

Thus decrease in mass number = 25%

Example - 2 What should be the percentage of deuterium in heavy water ?

(1) 20% (2) 80%

(3) 60% (4) 40% **Ans. (1)**

Solution : Deuterium in 20 parts of D_2O = 4 parts

Deuterium in 100 parts of $D_2O = \frac{4}{20} \times 100 = 20\%$

Example - 3 If $a = \frac{h}{4\pi^2 me^2}$, then the correct expression for calculation of the circumference of the first orbit of hydrogen atom should be

(1) $\sqrt{4h^2}\pi a$ (2) $2\pi r$

(3) $\sqrt{4}\,\pi ha$ (4) 1 and 3 both are correct **Ans. (4)**

Solution : Circumference = $2\pi r$

$2 \times \pi \times \frac{n^2h^2}{4\pi^2 mZe^2}$, n = 1 , Z = 1 and $\frac{h}{4\pi^2 me^2} = a$

Thus $2 \times \pi \times h \times a$ or $\sqrt{4}\pi ha$ or $\sqrt{4h^2}\,\pi a$

Example - 4 What should be the ratio of energies of the electrons of the first orbits of Na^{+10} and H ?

(1) 11:1 (2) 121 : 1

(3) 1 : 121 (4) 1 : 11 **Ans. (2)**

Solution : Here n is a constant, therefore

Example - 5 What should be the energy of a photon whose wavelength is 4000 Å ?

(1) 4.06×10^{-19} joule (2) 4.96×10^{-19} joule

(3) 3.0×10^{-12} joule (4) 2.4×10^{-19} joule **Ans. (2)**

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Solution : λ = 4000 Å i.e. 4000 × 10^{-8} cm 4 × 10^{-7} meter

$$E = h\nu = \frac{hc}{\lambda} = \frac{6.62\times10^{-34}\times3\times10^{8}}{4\times10^{-7}} = 4.96\times10^{-19}\text{ joule}$$

$$\frac{E_1}{E_2} = \frac{Z_1^{\,2}}{Z_1^{\,2}} = \frac{(11)^2}{(1)^2} = 121 : 1$$

Example - 6 Which of the following orbitals should be nearest to the nucleus ?

(1) 5s (2) 6p

(3) 3d (4) 4d **Ans. (3)**

Solution : n = 3 will be nearest to the nucleus.

Example - 7 What should be the total numbers of orbitals and electrons for m = 0, if there are 30 protons in an atom ?

(1) 7 orbitals, 14 electrons (2) 6 orbitals, 12 electrons

(3) 5 orbitals, 10 electrons (4) 3 orbitals, 6 electrons **Ans. (1)**

Solution : The configuration of the atom of atomic number 30 is $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^{6,}$ $3d^{10}$, $4s^2$. This will have 7 orbitals of m = 0.

Example - 8 Supposing that Pauli exclusion principle is nonexistent, which of the following should be the most unacceptable configuration of Li in ground state ?

(1) $1s^2$, $2s^1$ (2) $1s^3$

(3) $1s^1$, $2s^2$ (4) $1s^1$, $2s^1$, $2p^1$ **Ans. (4)**

Soluton : As a matter of fact, the configuration given in 2, 3 and 4 are wrong, but configuration given in 4 is most unacceptable because there is one electron in each of the three orbitals and according to Paulis exclusion principle maximum two electrons can be occupied in a orbital.

Example - 9 If the value of n + l = 7, then what should be the increasing order of energy of the possible sub-shells ?

(1) 4*f* < 5d < 6p < 7s (2) 7s < 6p < 5d < 4f

(3) 7s > 6p < 5d < 4p (4) 4f < 5d < 6p > 7s **Ans. (1)**

Solution : n + l = 7

7 + 0 = 7s Order of energy

6 + 1 = 6p 4f < 5d < 6p < 7s

5 + 2 = 5d

4 + 3 = 4f

Example - 10 Which of the following sub-shells will be filled by the electron after complete filling up of the orbital of the third principal shell ?

(1) 4s (2) 4f

(3) 4d (4) 4p **Ans. (4)**

Solution : The electron goes to 4p after filling up to 3d.

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Example - 11 Which of the following should be the atomic number of an atom if its electronic configuration is $(n-2)s^2$, $(n-1)\ s^a\ p^b\ ns^a\ p^2$ where n = 3, a = 2 and b = 6 ?

(1) 14 (2) 12

(3) 16 (4) 15 **Ans. (1)**

Solution : $(3-2)s^2\ (3-1)s^2\ p^6\ 3s^2\ 3p^2$; $1s^2, 2s^2\ 2p^6, 3s^2, 3p^2 = 14$

The atomic number of the atom is 14

Example - 12 Which of the following should be the electronic configuration of an atom in its first excited state if that atom is isoelectronic with O_2 ?

(1) [Ne] $3s^2\ 3p^4$ (2) [Ne] $3s^2\ 3p^3\ 3d^1$

(3) [Ne] $3s^1\ 3p^5$ (4) None of the above **Ans. (2)**

Solution : 16 electrons = $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^4$ (Two unpaired electrons)

Excited state = [Ne] $3s^2, 3p^3, 3d^1$ (Four unpaired electrons)

Example - 13 The mass charge ratio for A^+ ion is 1.97×10^{-7} kg C^{-1}. Calculate the mass of A atom.

(1) 9.02×10^{-31} kg (2) 3.16×10^{-26} kg

(3) 11.16×10^{-28} kg (4) 6.40×10^{-30} kg **Ans. (2)**

Solution : Given $\frac{m}{e} = 1.97 \times 10^{-7}$

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

m = 3.16×10^{-26} kg

Example - 14 AIR service on Vividh Bharati is transmitted on 219 m band. What is its transmission frequency in Hertz ?

(1) 1.37×10^6 Hz (2) 3.16×10^8 Hz

(3) 1.12×10^{10} Hz (4) 2.9×10^6 Hz **Ans. (1)**

Solution : Given $\lambda = 219$ m Thus, $\nu = \frac{c}{\lambda}$

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

Example - 15 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)**

Solution : 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^+ = $19.6 \times 10^{-18} \times \frac{9}{4}$ **= 44.10×10^{-18} J/atom**

Example - 16 Which of the following set of quantum numbers is permitted

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

(3) n = 2, l = 2, m = +1, s = – 1/2 (4) n = 2, l = 2, m = +1, s = – 1/2 **Ans. (1)**

Solution : (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’.

Example - 17 Calculate the number of proton emitted in 10 hours by a 60 W sodium lamp (λ or photon = 5893 Å)

(1) 17.8×10^{29} (2) 3.37×10^{20}

(3) 11.9×10^{19} (4) 6.41×10^{24} **Ans. (4)**

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

Energy of photon emitted = $\dfrac{hc}{\lambda} = \dfrac{6.626 \times 10^{-34} \times 3 \times 10^{8}}{5893 \times 10^{-10}} = 3.37 \times 10^{-19}$ J

$\therefore$ No of photons emitted per sec. = $\dfrac{60}{3.37 \times 10^{-19}}$

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

Example - 18 Find out the energy of H atom in first excitation state.The value of permittivity factor $4\pi\varepsilon_0 = 1.11264 \times 10^{-10}\ C^2\ N^{-1}\ m^{-2}$.

(1) 9.3×10^{-11} J (2) 6.62×10^{-34} J

(3) 5.443×10^{-19} J (4) 2.44×10^{-10} J **Ans. (3)**

Solution : 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**

Example - 19 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)**

Solution : 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. $n_2 = \infty$) $\dfrac{1}{\lambda} = R_H\left[\dfrac{1}{1^2} - \dfrac{1}{\infty^2}\right]$

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

Example - 20 Electromagnetic radiations of wavelength 242 nm is just sufficient to ionise sodium atom. Calculate the ionisation energy of sodium in kJ mol^{-1}.

(1) 806.3 (2) 80.63

(3) 49.45 (4) 494.5 **Ans. (4)**

Solution : Energy associated with a photon of 242 nm = $\dfrac{6.625 \times 10^{-34} \times 3.0 \times 10^{8}}{242 \times 10^{-9}} = 8.21 \times 10^{-19}$ 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

= $8.21 \times 10^{-19} \times 6.023 \times 10^{23}$

= 49.45×10^{4} J = $\mathbf{494.5\ kJ\ mol^{-1}}$

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Example - 21 What is the longest wavelength which can remove the electron from I Bohr's orbit. Given E_1 = 13.6 eV.

(1) 912.24Å (2) 21.78Å

(3) 13.34Å (4) 264.92Å **Ans. (1)**

Solution : The photon capable of removing electron from I Bohr's orbit must possess energy

= 13.6 eV

= 13.6 × 1.602 × 10^{-19} J = 21.787 × 10^{-19} J

$$\because \quad E = \frac{hc}{\lambda}\ ;\ 21.787 \times 10^{-19} = \frac{6.625\times10^{-34}\times3.0\times10^{8}}{\lambda}$$

$\therefore$ λ = 912.24 × 10^{-10} m **= 912.24 Å**

Example - 22 Naturally occurring boron consists of two isotopes A and B whose atomic weights are 10.01 and 11.01 respectively. The atomic weight of natural boron is 10.81. What is the percentage of isotope B in natural boron

(1) 20 (2) 80

(3) 70 (4) 30 **Ans. (2)**

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

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

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

$$\text{or Average atomic wt.} = \frac{x\times10.01+(100-x)\times11.01}{100}$$

$$10.81 = \frac{x\times10.01+(100-x)\times11.01}{100} = 20$$

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

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

Example - 23 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)**

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

= 13.6 × 9

= **122.4 eV**

Example - 24 What will be the momentum of radiations of wavelength 0.33 nm.

(1) 2.01 ×10^{-24} kg ms^{-1} (2) 4.91 ×10^{-24} kg ms^{-1}

(3) 6.41 ×10^{-28} kg ms^{-1} (4) 14.1 ×10^{-28} kg ms^{-1} **Ans. (1)**

Solution : We have $\lambda = \dfrac{h}{mv}$ $\quad\therefore$ $mv = \dfrac{h}{\lambda}$

$$= \frac{6.625\times10^{-34}}{0.33\times10^{-9}} = 2.01 \times 10^{-24}\ \text{kgmsec}^{-1}$$

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Example - 25 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)**

Solution :

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

Mass no. – At no. = No. of neutrons

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EXERCISES

LEVEL - I

1. Which of the following species will produce the shortest wavelength for the transition n = 2 to n = 1?

(1) H-atom (2) He^+ -ion

(3) H_1^2 (4) Li^{+2} - ion

2. The difference in angular momentum associated with the electron in two successive orbits of hydrogen atom is:

(1) $\frac{h}{\pi}$ (2) $\frac{h}{2\pi}$

(3) $\frac{h}{2}$ (4) $(n-1)\frac{h}{2\pi}$

3. Suppose 10^{-17} Jouls of light energy is required by the interior of the human eye to see on object. How many photons of green light (λ = 550 nm) are needed to generate this minimum amount of energy?

(1) 20 (2) 28

(3) 22 (4) 30

4. Non-directional orbital is

(1) 3s (2) 4f

(3) 4d (4) 4p

5. When the frequency of light incident on a metallic plate is doubled, the KE of the emitted photo electrons will be:

(1) doubled (2) halved

(3) increased but more than double of the previous KE (4) unchanged

6. In an atom two eletrons move around the nuclens in circular orbit of R and 4R. The ratio of the time taken by them to complete one revolution is :

(1) 1 : 4 (2) 4 : 1

(3) 1 : 8 (4) 8 : 7

7. A 3p-orbital has

(1) two non-spherical nods (2) two spherical nods

(3) one spherical and one non-spherical node (4) one spherical and two nonspherical nods

8. Calculate no. of photons emitted in 10 hrs by a 60 W sodium 1 amp (λ = 5893Å)

(1) 6.4×10^{24} (2) 64×10^{24}

(3) 6.4×10^{20} (4) None

9. How many times does the electrons go around the first Bohr Orbit H-atom in one second?

(1) 6.57×10^{15} (2) 65.7×10^{15}

(3) 6.57×10^{16} (4) None

10. What is the speed of an electron whose de-Broglie wavelength is 0.1 nm?

(1) $7.2 \times 10^6\ ms^{-1}$ (2) $72 \times 10^6\ ms^{-1}$

(3) $72 \times 10^2\ ms^{-1}$ (4) $7.2 \times 10^4\ ms^{-1}$

11. Find the product of uncertainty in position and velocity for an electron of mass 9.1×10^{-31} kg?

(1) $5.8 \times 10^{-5}\ m^2s^{-1}$ (2) $58 \times 10^{-5}\ m^2s^{-1}$

(3) $5.8 \times 10^{-3}\ m^2s^{-1}$ (4) $5.8 \times 10^{-2}\ m^2s^{-1}$

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12. How many photons of light having a wavelnegth of 4000Å are required to provide 1 Joule of energy?
(1) 2×10^{18} (2) 0.2×10^{20}
(3) 2×10^{15} (4) None

13. Which one is the heaviest fundamental particle
(1) Electron (2) Proton
(3) Neutron (4) None

14. The no. of waves made by a Bohr electron in an Orbit of maximum magnetic quantic number 3 is
(1) 3 (2) 4
(3) 2 (4) 1

15. Which orbital notation does not have spherical node;
(1) n = 2, l = 0 (2) n = 2, l = 1
(3) n = 3, l = 0 (4) n = 1, l = 0

16. The distance between 3rd and 2nd orbits in the H-atom is;
(1) 2.116×10^{-8} cm (2) 2.646×10^{-8} cm
(3) 0.529×10^{-8} cm (4) 1.058×10^{-8} cm

17. 13.5 gm of Al when change to Al^{+3} ion in solution, will lose;
(1) 18×10^{23} electron (2) 6.022×10^{23} electrons
(3) 3.01×10^{23} electrons (4) 9.1×10^{23} electrons

18. In Bohr's model of the H-atom the ratio between the period of revolution of an electron in the orbit n = 1 to the period of revolution of the electron in the orbit n == 2 is
(1) 1 : 2 (2) 2 : 1
(3) 1 : 4 (4) 4 : 1

19. Ionisation potential of H-atom is 13.6 eV. If hydrogen atom in the ground state is excited by monochromatic light of energy 12.1 eV then the spectral line emitted according to Bohr's theory will be ;
(1) 1 (2) 2
(3) 3 (4) 4

20. The wave length of a certain line in Balmer series is observed to be 4341Å. To what value of "n" does this correspond?
(1) 5 (2) 4
(3) 4 (4) 6

21. Which one is not possible;
(1) n = 2, l = 2 (2) n = 2, l = 0
(3) n = 3, l = 2 (4) n = 2, l = 1

22. The two unpaired electrons of carbon differs in terms of their;
(1) n (2) l
(3) m (4) s

23. Isotopes are possible due to differene no of
(1) Electrons (2) Protons
(3) Neutrons (4) None

24. An electron can revolve only in those orbit in which angular momentum is integral multiple of
(1) $\frac{h}{\pi}$ (2) $\frac{h}{2\pi}$
(3) $\frac{h}{3\pi}$ (4) $\frac{h}{4\pi}$

25. Maximum no. of electrons in; n = 1, l = 0, m = 0, s = +1/2 is;
(1) 2 (2) 1
(3) 3 (4) 4

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LEVEL - II

1. The molecular weight of an oxide of nitrogen is 30. The number of electrons present in one molecule of this compound is :

(1) 15 (2) 30

(3) $6.02 \times 10^{23} \times 15$ (4) $6.02 \times 10^{23} \times 30$

2. 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^+

3. Which of the following statements is false ?

(1) $(n + \ell)$ rule arranges the orbitals in increasing order of energy

(2) Wavelength of a particle is inversely proportional to its momentum

(3) Aufbau's principle was given a scientist named Aufbau

(4) Velocity of all types of electromagnetic radiation is same

4. Electron density in the region between 1s and 2s- orbital is :

(1) high (2) low

(3) zero (4) None of these

5. If the radius of first orbit of H-atom is 5 pm, the radius of third orbit Li^{2+} will be :

(1) 106 pm (2) 23 pm

(3) 32 pm (4) 15 pm

6. The K.E. of an electron in first Bohr's orbit of H-atom is 13.6 eV. Total energy of first orbit is :

(1) $-\frac{1}{2} \times 13.6$ eV (2) – 13.6 eV

(3) 2×13.6 eV (4) $\frac{1}{2} \times 13.6$ eV

7. The shape of p-orbital is :

(1) Elliptical (2) spherical

(3) dumb-bell (4) None of these

8. ${}_6C^{11}$ and ${}_5B^{11}$ are called :

(1) Nuclear isomers (2) Isobars

(3) Isotopes (4) Fission products

9. No two electrons of an atom can have same :

(1) principle quantum number

(2) azimuthal quantum number

(3) set of four quantum numbers

(4) magnetic quantum number

10. Correct set of four quantum numbers for the valence (outermost) electron of rubidium (Z = 37) is :

(1) 5, 0, 0 + 1/2 (2) 5, 1, 0, + 1/2

(3) 5, 1, 1, + 1/2 (4) 6, 0, 0, + 1/2

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11. Which of the following formula represents the K.E. of an electron in n^{th} Bohr's orbit of H-atom ?

(1) $\frac{Rhc}{n^2}$ (2) $-\frac{Rhc}{n^2}$

(3) $-\frac{2Rhc}{n^2}$ (4) $\frac{2Rhc}{n^2}$

12. The energy of electron in excited H-atom is – 3.4 eV. What is the angular momentum of electron ?

(1) $\frac{h}{\pi}$ (2) $\frac{h}{2\pi}$

(3) $\frac{2h}{\pi}$ (4) $\frac{3h}{\pi}$

13. What is the frequency of the electron in an orbit of radius r, if its velocity is v ?

(1) $\frac{2\pi r}{v}$ (2) $2\pi rv$

(3) $\frac{vr}{2\pi}$ (4) $\frac{v}{2\pi r}$

14. How many spectral lines will be obtained by the various transitions when an electron comes from excited state n = 5 to its original state ?

(1) 20 (2) 5

(3) 4 (4) 10

15. Principal, azimuthal and magnetic quantum numbers are respectively related to :

(1) size, shape and orientation (2) shape, size and orientation

(3) size, orientation and shape (4) None of these

16. Bohr's model of the atom can explain :

(1) The spectrum of H-atom only

(2) The spectrum of hydrogen molecule

(3) The spectrum of atom or ion containing one electron only

(4) The solar spectrum

17. The wavelength of first line of Balmer series of H-atom is – (R = Rydberg's constant)

(1) $\frac{36}{5R}$ (2) $\frac{36R}{5}$

(3) $\frac{5R}{36}$ (4) $\frac{5}{36R}$

18. The excitation energy of an electron from second orbit to third orbit of an atom with + Ze nuclear charge is 47.2 eV. If the energy of H-atom in lowest energy state is – 13.6 eV. What will be the value of Z ?

(1) 4 (2) 5

(3) 6 (4) 7

19. The statement 'It is not possible to estimate accurately the position and momentum of an electron simultaneously is associated with :

(1) Heisenberg's uncertainty principle (2) De-Broglie's principle

(3) Pauli's uncertainty principle (4) Aufbau principle

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20. The electron of H-atom transit from n = 1 to n = 4 by absorbing energy. If the energy of n = 1 state is – 21.8 × 10^{-19} Joule then its energy in n = 4 state will be :

(1) – 21.8 × 10^{-19} Joule
(2) – 5.45 × 10^{-19} Joule
(3) – 2.725 × 10^{-19} Joule
(4) – 1.362 × 10^{-19} Joule

21. The wavelength of first line of Lymen series is 1216 Å. What is the wave length of last line ?

(1) 3648 Å
(2) 608 Å
(3) 912 Å
(4) 2432 Å

22. 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

23. de' Broglie equation tells about :

(1) the relation between electron and nucleus
(2) the relation between electron and proton
(3) the relation between electron and neutron
(4) electrons' dual nature of wave and particle

24. The circumference of first Bohr's orbit of hydrogen atom is how many times the circumference of second Bohr's orbit of He^+ ?

(1) two times
(2) half
(3) equal
(4) none of these

25. The mass of a cricket ball is 0.21 kg. If the order of uncertainty in position is 100 pm then uncertainty in its velocity will be :

(1) 3.5 × 10^{-24} m/sec
(2) 6.02 × 10^{23} m/sec
(3) 6.602 × 10^{-27} m/sec
(4) 2.5 × 10^{-24} m/sec

LEVEL - III

1. The orbital cylindrically symmetrical about x-axis is :

(1) p_z (2) p_y

(3) p_x (4) d_{xz}

2. Which of the d-orbital lies in the xy-plane :

(1) d_{xz} only (2) d_{xy} only

(3) $d_{x^2-y^2}$ only (4) d_{xy} & $d_{x^2-y^2}$ only

3. The probability of finding an electron residing in a p_x orbital is zero in the :

(1) xy plane (2) yz plane

(3) y direction (4) z direction

4. If the series limit of wavelength of the Lyman series for the hydrogen atoms is 912Å, then the series limit of wavelength for the Balmer series of the hydrogen atom is :

(1) 912 Å (2) 912 × 2 Å

(3) 912 × 4 Å (4) 912/2 Å

5. An element of atomic weight Z consist of two isotopes of mass number Z – 1 and Z + 2. Percentage of abundanc of the heavier isotope is :

(1) 25 (2) $33\frac{1}{3}$

(3) $66\frac{2}{3}$ (4) 75

6. When an electron of charge e and mass m moves with velocity v about the nuclear charge Ze in the circular orbit of radius r, then the potential energy of the electron in given by :

(1) Ze^2/r (2) $-Ze^2/r$

(3) Ze^2/r^2 (4) mv^2/r

7. If uncertainty in position of electron is zero, then the uncertainty in its momentum would be :

(1) Zero (2) $h/2\pi$

(3) $3h/2\pi$ (4) Infinity

8. It is known that atoms contain protons, neutrons and electrons. If the mass of neutron is assumed to be half of its original value whereas that of electron is assumed to be twice to this original value. The atomic mass of ${}_6C^{12}$ will be :

(1) Twice (2) 75% less

(3) 25% less (4) One half of its original value

9. The speed of a proton is one hundredth of the speed of light in vacuum. What is its de-Broglie wavelength? Assume that one mole of protons has a mass equal to one gram [h = 6.626×10^{-27} erg sec] :

(1) 13.31×10^{-3} Å (2) 1.33×10^{-3} Å

(3) 13.13×10^{-2} Å (4) 1.31×10^{-2} Å

10. The ratio of ionization energy of H and Be^{+3} is :

(1) 1 : 1 (2) 1 : 3

(3) 1 : 9 (4) 1 : 16

11. Hydrogen spectrum consists of :

(1) An intense line (2) Six series of lines

(3) Three series of lines (4) Four series of lines

12. Which set of quantum numbers is possible for the last electron of Mg^+ ion :

(1) $n = 3, l = 2, m = 0, s = +\frac{1}{2}$ (2) $n = 2, l = 3, m = 0, s = +\frac{1}{2}$

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

13. In hydrogen atom, If an electron jumps from n = 6 to n = 2, how many possible spectral lines are obtained:

(1) 15 (2) 10

(3) 6 (4) 12

14. The speed of the electron in the 1st orbit of the hydrogen atom in the ground state is–

(1) c/1.37 (2) c/1370

(3) c/13.7 (4) c/137

15. Difference between nth and (n + 1)th Boh's radius of 'H' atom is equal to it's (n –1)th Bohr's radius. The value of n is :

(1) 1 (2) 2

(3) 3 (4) 4

16. The potential energy of the electron present in the ground state of Li^{2+} ion is represent by :

(1) $+\frac{3e^2}{4\pi\varepsilon_0 r}$ (2) $-\frac{3e}{4\pi\varepsilon_0 r}$

(3) $-\frac{3e^2}{4\pi\varepsilon_0 r}$ (4) None of these

17. Assume that the nucleus of the F-atom is a sphere of radius 5×10^{-13} cm. What is the density of matter in F–nucleus :

(1) 6.02×10^{11} g/ml (2) 6.02×10^{13} g/ml

(3) 6.02×10^{18} g/ml (4) None

18. Assuming the nucleus and an atom to be spherical, the radius of the nucleus of mass number a is given by $1.25 \times 10^{-13} \times A^{1/3}$ cm. The atomic radius of atom is 1Å. If the mass number is 64, the fraction of the atomic volume that is occupied by the nucleus is :

(1) 1.25×10^{-13} (2) 2.50×10^{-13}

(3) 5×10^{-5} (4) None

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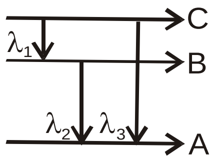
19. An electron in a hydrogen atom in its ground state absorbs 1.50 times as much energy as the minimum required for it to escape from the atom. What is the wavelength of the emitted electron :

(1) 4.70 Å (2) 4.70 nm

(3) 9.4 Å (4) 9.40 nm

20. Energy levels A, B, C of a certain atom corresponds to increasing values of energy, i.e., $E_A < E_B < E_C$. If λ_1, λ_2 and λ_3 are the wavelengths of radiations corresponding to the transitions C to B, B to A and C to A respectively, which of the following statement is correct :



(1) $\lambda_3 = \lambda_1 + \lambda_2$ (2) $\lambda_3 = \dfrac{\lambda_1\lambda_2}{\lambda_1 + \lambda_2}$

(3) $\lambda_1 + \lambda_2 + \lambda_3 = 0$ (4) $\lambda_3^2 = \lambda_1^2 + \lambda_2^2$

21. The uncertainty in the position of an electron (mass 9.1×10^{-28} gm) moving with a velocity of 3×10^4 cm sec^{-1}, Accurate upto 0.011% will be :

(1) 1.92 cm (2) 7.68 cm

(3) 0.175 cm (4) 3.84 cm

22. Two particles A and B are in motion. It the wavelength associated with the particle A is 5.0×10^{-8} m, the wavelength of particle B having momentum half of A is :

(1) 2.5×10^{-8} m (2) 1.25×10^{-8} m

(3) 1.0×10^{-7} m (4) 1.0×10^{-8} m

23. A particle A moving with a certain velocity has de Broglie wavelength of 1 Å. If particle B has mass 25% of that A and velocity 75% of that of A, the de Broglie wavelength of B will be approximately :

(1) 1 Å (2) 5.3 Å

(3) 3 Å (4) 0.2 Å

24. Energy required to pull out an electron from 1st orbit of hydrogen atom to infinity is 100 units. The amount of energy needed to pull out the electron from 2nd orbit to infinity is :

(1) 50 units (2) 100 units

(3) 25 units (4) Zero

25. If each hydrogen atom is excited by giving 8.4eV energy, then the number of spectral lines emitted is equal to:

(1) none (2) 2

(3) 3 (4) 4

PROBLEMS ASKED IN COMPETITIVE EXAMINATIONS

1. Which of the following sets of quantum numbers represents the highest energy of an atom ? **[AIEEE 2007]**

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

2. Uncertainty in the position of an electron (mass = 9.1×10^{-31} kg) moving with a velocity 300 ms^{-1}, accurte upto 0.001%, will be **[AIEEE 2006]**

(1) 1.92×10^{-2} m (2) 3.84×10^{-2} m

(3) 19.2×10^{-2} m (4) 5.76×10^{-2} m

(h = 6.63×10^{-34} Js)

3. According to Bohr's theory, the angular momentum of electron in 5th orbit is **[AIEEE 2006]**

(1) 10 h/π (2) 2.5 h/π

(3) 25 h/π (4) 1.0 h/π

4. Electrons will first enter into the orbital with set of quantum numbers **[DCE 2005]**

(1) n = 5, l = 0 (2) n = 4, l = 1

(3) n = 3, l = 2 (4) any of these

5. The H-spectrum shows **[DCE 2005]**

(1) Heisenberg's uncertainty principle (2) diffraction

(3) polarisation (4) presence of quantized energy levels

6. The number of radial nodes in 3s and 2p respectively are **[IIT 2005]**

(1) 2 and 0 (2) 1 and 2

(3) 0 and 2 (4) 2 and 1

7 The wavelength of the electron 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) 91 nm (2) 192 nm

(3) 406 nm (4) 9.1×10^{-8} nm

8. If n = 3, l = 0, m = 0, then atomic number is **[Bihar CECE 2004]**

(1) 12, 13 (2) 13, 14

(3) 10, 11 (4) 11, 12

9. The one electron species having ionization energy of 54.4 eV is **[Kerala C.E.E. 2004]**

(1) H (2) He^{+}

(3) B^{4+} (4) Li^{2+}

10. Consider the ground state of Cr atom (Z = 24). The number of electrons with the azimuthal quatum number l = 1 and 2 are respectively **[AIEEE 2004]**

(1) 12 and 4 (2) 12 and 5

(3) 16 and 4 (4) 16 and 5

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11. The radius of which of the which of the following orbit is same as that of the first Bohr's orbit of hydrogen atom ? **[IIT 2004]**

(1) He^+ (n = 2)
(2) Li^{2+} (n = 2)
(3) Li^{2+} (n = 3)
(4) Be^{3+} (n = 2)

12. 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) $\frac{h}{2\pi}$

13. In the Bohr's orbit, what is the ratio of total kinetic energy and total energy of electron : **[RPET 2002]**

(1) – 1
(2) – 2
(3) 1
(4) + 2

14. Which set of quantum number for an electron of an atom is not possible : **[RPET 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 = $+\frac{1}{2}$

15. 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 = $+\frac{1}{2}$

16. 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^{-35} m

17. 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)

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18. 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

19. Heaviest particle is : **[MP PET 1999]**

(1) Meson (2) Neutron

(3) Proton (4) Electron

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

(1) $n = 2, \ell = 1, m = -1, s = -1/2$ (2) $n = 3, \ell = 0, m = 0, s = +1/2$

(3) $n = 3, \ell = 2, m = -2, s = -1/2$ (4) $n = 3, \ell = 2, m = 2, s = +1/2$

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

(1) 3 : 2 (2) 4 : 9

(3) 9 : 4 (4) 9 : 1

22. 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

23. 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}

24. 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

25. 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

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ANSWERS

EXERCISES

LEVEL - I

1. (4)	2. (2)	3. (2)	4. (1)	5. (3)
6. (3)	7. (3)	8. (1)	9. (1)	10. (1)
11. (1)	12. (1)	13. (3)	14. (2)	15. (2)
16. (2)	17. (4)	18. (4)	19. (3)	20. (1)
21. (1)	22. (3)	23. (3)	24. (2)	25. (1)

LEVEL - II

1. (1)	2. (1)	3. (3)	4. (3)	5. (4)
6. (4)	7. (3)	8. (2)	9. (3)	10. (1)
11. (1)	12. (1)	13. (4)	14. (4)	15. (1)
16. (3)	17. (1)	18. (2)	19. (1)	20. (4)
21. (3)	22. (4)	23. (4)	24. (2)	25. (4)

LEVEL - III

1. (3)	2. (3)	3. (2)	4. (3)	5. (2)
6. (2)	7. (4)	8. (3)	9. (3)	10. (4)
11. (2)	12. (4)	13. (2)	14. (4)	15. (4)
16. (3)	17. (2)	18. (1)	19. (1)	25. (2)
21. (3)	22. (3)	23. (2)	24. (3)	25. (1)

PROBLEMS ASKED IN COMPETITIVE EXAMINATIONS

1. (4)	2. (1)	3. (3)	4. (3)	5. (4)
6. (1)	7. (1)	8. (4)	9. (2)	10. (2)
11. (4)	12. (3)	13. (1)	14. (2)	15. (4)
16. (1)	17. (1)	18. (3)	19. (2)	20. (2)
21. (3)	22. (4)	23. (4)	24. (3)	25. (1)

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