2

STRUCTURE OF ATOM

ATOM & MOLECULES

(a) The smallest particle of a matter that takes part in a chemical reaction is called an atom. The atom of all gases except those of noble gases, cannot exist in free state. These exist in molecular form. The molecules of hydrogen, nitrogen, oxygen and halogens are diatomic (H_2, N_2). Phosphorus and sulphur molecules are in P_4 & S_8 state.

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

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

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 :

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

(b) Atoms cannot be further subdivided.

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

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

(e) Atom can neither be created nor destroyed.

MODERN CONCEPT

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

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

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

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

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

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

(f) Atoms of elements combines to form molecules.

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

(h) Atoms participate in chemical reactions.

PROTON

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

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

(c) Proton bears one unit positive charge.

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

$$\frac{e}{e/m} = \frac{1.6012\times10^{-19}\,\text{coulomb}}{1.76\times10^{8}\,\text{coulomb / gram}}$$

$= 9.1091 \times 10^{-28}$ gram = 0.0005486 a.m.u.

= 1/1837th of H atom

(g) An atom of hydrogen (or a proton) is 1837 times heavier than electron.

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

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

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

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

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

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

DISCOVERY OF NEUTRON

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

Neutron (${}_0n^1$) :

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

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

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

(d) The mass of a neutron is almost equal to that of a proton. Actually it is a little bit heavier than proton. Its mass (m) is as fallows : Mass (m) of a neutron = 1.6748×10^{-24} gram

= Approximately mass of a proton

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

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

OTHER PARTICLES OF ATOM

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

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

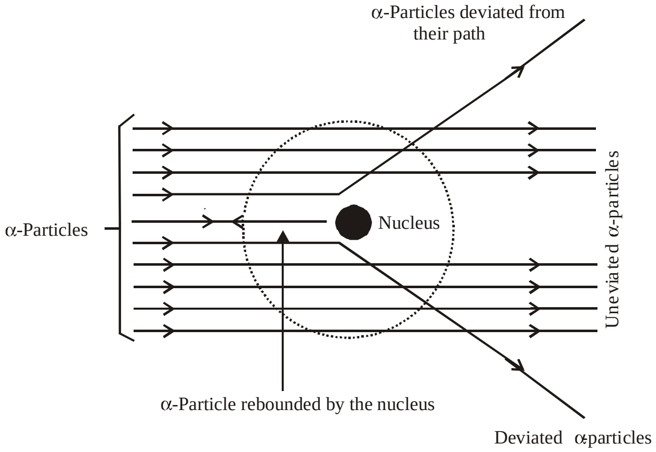
(c) **Neutrino :** Pauling discovered these particles in 1927. They do not bear any charge, i.e. they are electro neutral particle.

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

NUCLEUS

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

- **(a)** Most of the α-particle passed through the gold film without deviating from their path.
- **(b)** Some particles got deviated from their path on colliding with the gold film.
- **(c)** A very small number of particles rebounded after colliding with the gold film.



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

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

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

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

(d) Radius of nucleus : $r_n = r_0 \times A^{1/3}$,
where r_0 is a constant and A is mass number.

(e) Nuclear 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 = $\frac{4}{3}\pi r^3$

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)

Sol. ∴ 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 = $\frac{4}{3}\pi r^3$

$$= \frac{4}{3} \times 3.14\,(5 \times 10^{-13})^3 = 525 \times 10^{-39}\ \text{cm}^3$$

Density of the nucleus of F atom

$$= \frac{\text{Mass}}{\text{Volume}} = \frac{19 \times 1.66 \times 10^{-24}\ \text{gm}}{525 \times 10^{-39}\ \text{cm}^3} = 6.0 \times 10^{13}\ \text{gm cm}^{-3}$$

NUCLEAR CHARGE AND ATOMIC NUMBER

Positive charge on the nucleus of an atom is equal to the atomic number of that atom. A scientist named Mosley studied the frequency of X-rays emitted by showering high velocity electrons on a metal and established the following relationship. $\sqrt{\nu} = a(z-b)$, where, ν = frequency of X-rays z = atomic number or nuclear charge, a and b are constants. Thus nuclear charge of an atom is equal to the atomic number of that atom. Since an atom is electro neutral, the number of positively charged protons in its nucleus is equal to the negatively charged electrons moving around the nucleus in the atom. Thus

Atomic number = number of protons in the atom
or number of electrons in the atom

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 = No. of protons(Z) + No. of neutrons (n)
= Atomic number + Number of neutrons

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

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

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

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

(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.

Example 2 :

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

Sol. $_8O^{16}$

Initial weight – final weight

Protons	8p	→	8p
Neutrons	8n	→	4n
Weight	16	→	12

Thus decrease in mass number = 25%

Example 3 :

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

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

ISOTOPES

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

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

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

(c) Hydrogen has the following three isotopes.

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

(d) Chlorine has the following two isotopes.

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

ISODIAPHERES

The elements which have same value of (n – p).

e.g	$_7N^{14}$	$_8O^{16}$
Values of (n – p)	0	0

ISOSTERS

They are the molecules which have the same number of atoms and electrons.

Example 4 :

Which of the following pairs are isosters –

(1) CO_2 and NO_2 (2) CaO and KF
(3) OF_2 and HClO (4) All of these

Sol. **(4).** CO_2 :
Atoms = 1 + 2 = 3, Electrons = 6 + 8 × 2 = 22 e^-
NO_2 : Atoms = 1 + 2 = 3, Electrons = 7 × 2 + 8 = 22 e^-
CaO : Atoms = 2, Electrons = 20 + 8 = 28 e^-
KF : Atoms = 2, Electrons = 19 + 9 = 28 e^-
OF_2 : Atoms = 3, Electrons = 8 + 18 = 26 e^-
HClO : Atoms = 3, Electrons = 1 + 17 + 8 = 17 + 9 = 26 e^-

Example 5 :

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

Sol. 80 : 15 : 5
Thus the ratio is 16 : 3 : 1
Total = 16 + 3 + 1 = 20

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

ISOBARS

Isobars are the atoms of different elements having same atomic weight. Isobars have different numbers of protons as well as neutrons. 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	

ISOTONES

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

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

Example 6 :

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

Sol. 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) $M_1 - Z_1 = M_2 - Z_2$

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		

Example 7 :

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

Sol. **(1).** BO_3^{-3} CO_3^{-2}
5 + 24 + 3 = 32 6 + 24 + 2 = 32

Example 8 :

The pair $NH_3 + BH_3$ is isoelectronic with

(1) B_2H_6 (2) C_2H_6
(3) C_2H_4 (4) CO_2

Sol. **(2).** $NH_3 + BH_3$ C_2H_6
7 + 3 + 5 + 3 = 18 6 × 2 + 6 = 18

Example 9 :

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

Sol. **(4).** 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.

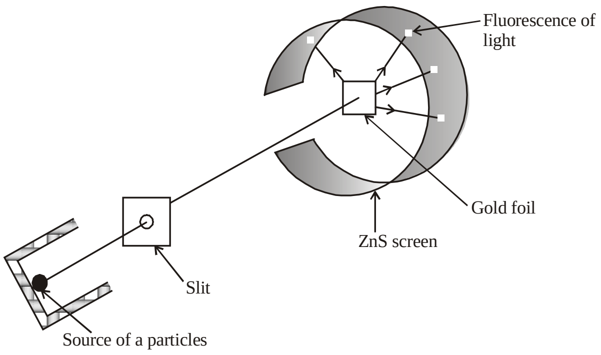
ATOMIC MODEL

Thomson's Model of an Atom

Atom is a very minute, spherical, electro neutral particle that consists of positively and negatively charged matter. 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". Thomson's model of an atom failed to explain the production of the atomic spectrum. It cannot explain Rutherford's α particle scattering experiment also.

Rutherford's Model of an Atom :

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



Rutherford's a particle scattering experiment

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

(a) Positively charged nucleus and

(b) a vast extranuclear space in which electrons are present.

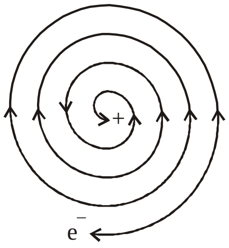
(ii) whole of the positive charge and almost all the mass of atom is confined to a very minute part at the centre of the atom, called the nucleus of the atom. The radius of nucleus is about 10^{-13} to 10^{-12} cm (or 10^{-15} to 10^{-14} meter), while the radius of atom is in the order of 10^{-8} cm.

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

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

Demerits of Rutherford's Model of an Atom

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



An electron emitting energy and pluging into nucleus

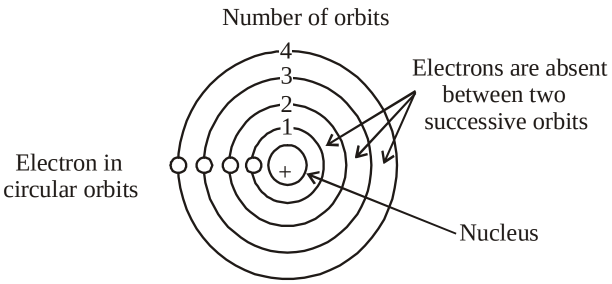
(ii) Plunging of an electron into the nucleus would definitely mean destruction of the atom or end of the existence of the atom. But we know that it never happens. Atom is a stable system. Therefore Rutherford model failed in explaining the stability of an atomic system.

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

BOHR'S MODEL OF AN ATOM

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

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

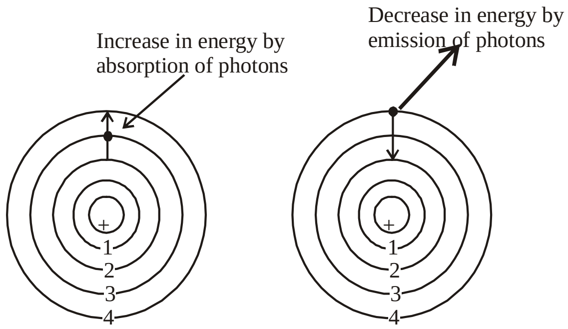


Electrons moving around in circular orbits

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

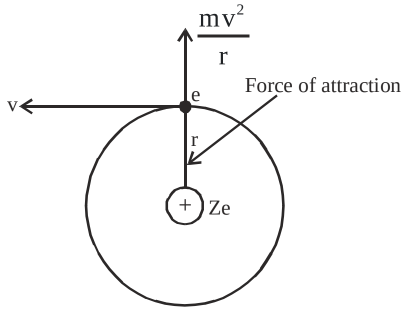
(iii) When energy is provided to an atom, its electrons get excited by absorption of energy and jumps to the orbits of higher energy.

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



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

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



The electron moving in an orbit by various forces

Note : (a) $mvr = \frac{nh}{2\pi}$ ....(1)

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

(c) $E_{n_2} - E_{n_1} = h\nu$ ....(3)

Example 10 :

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) $4 \to 1$	(2) $4 \to 2, 2 \to 1$
(3) $4 \to 3, 3 \to 2, 2 \to 1$	(4) All of the above

Sol. (4). Electron can undergo transition from higher state to all lower states by loss of energy.

Example 11 :

How much total energy will be released when an electron present in hydrogen atom undergoes the following sequence of transition ? $n = 4 \to n = 2 \to n = 1$

Sol. 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 12 :

Which of the following is a fundamental particle

(1) Nucleus of He	(2) Nucleus of H
(3) A positive atom	(4) None of these

Sol. (2). Fundamental particle proton is the nucleus of H.

VELOCITY OF THE ELECTRON OF BOHR'S ORBIT

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

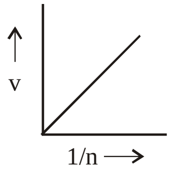
From Bohr's postulate, $mvr = \frac{nh}{2\pi}$ ....(2)

Eq. (1) divided by (2), $v = \frac{2\pi Ze^2}{nh}$ or $v = K\frac{z}{n}$

Here π , e and h are constants, therefore

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

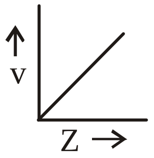
or $v = \frac{Z}{n} \times 2.188 \times 10^8$ cm / second



(a) If Z is a constant, then

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

Therefore, velocity goes on decreasing with increase in the number of orbits. Thus $\frac{v_1}{v_2} = \frac{n_2}{n_1}$



(b) If n is a constant, then $v \propto Z$
Therefore, velocity goes on increasing with increase in the atomic number.

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

(c) Time period

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

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

Example 13 :

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

Sol. (2). Z is a constant, therefore $v \propto \frac{1}{n}$ i.e. $v_4 < v_3 < v_2 < v_1$

Example 14 :

Find the ratio of velocities of electrons present in Na^{+10} and H.

Sol. $Na^{11} \to Na^{+10} \to Is^1$

Thus, n is a constant. Therefore $\frac{v_1}{v_2} = \frac{Z_1}{Z_2} = \frac{11}{1}$

RADIUS OF nth BOHR'S ORBIT

According to Bohr's hypothesis, put the value of v in

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

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

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

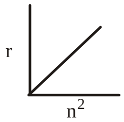
or $r = \frac{n^2}{Z} \times 0.529$ Å

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

If Z is a constant, then $r \propto n^2$

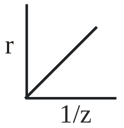
Thus, the radius of atoms goes on increasing as the number (n) of energy levels in the atoms goes on increasing as shown

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



(b) If n is a constant, then

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



Example 15 :

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

Sol. (2). $r \propto n^2$; $r_1 < r_2 < r_3 < r_4$

Example 16 :

Find the ratio of radius of the fifth orbits of He^+ and Li^+.

Sol. Here n is a constant, therefore, $\frac{r_1}{r_2} = \frac{Z_2}{Z_1} = \frac{3}{2} = 3:2$

Example 17 :

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

Sol. (4). 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 $\sqrt{4h^2}\,\pi a$

ENERGY OF ELECTRON IN BOHR'S nth ORBIT

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

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

(c) Stability would increase as the electron in an atom moves from the infinity distance to a distance r from the nucleus, resulting in the value of the potential energy becoming negative. This is because of the fact that when two opposite charges attract each other, there is a decrease in the potential energy, as attractive forces = $\frac{Ze^2}{r^2}$

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

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

(f) **Energies are of two types.**

Kinetic Energy (E_K) : This energy is produced due to the velocity of electron. If mass is m, velocity is v and radius is r then kinetic energy = $\frac{1}{2}mv^2 = \frac{1}{2}\frac{Ze^2}{r}$

Potential Energy (E_P) : This energy is produced due to electrostatic attractive forces between electron and proton, and its value is negative. If atomic number is Z, charge is e and radius is r, then Potential energy = $\frac{-Ze^2}{r}$

Total Energy (E_T) :

Total energy = Kinetic energy + potential energy

$$E_T = E_K + E_P$$

$$\frac{1}{2}mv^2 + \frac{-Ze^2}{r} \text{ ; Total energy } E = -\frac{1}{2}\frac{Ze^2}{r}$$

Formula

(i) Total energy = – Kinetic energy ($E_T = -E_K$)
(ii) Potential energy = 2 × Total energy) ($E_P = 2E_T$)

Example 18 :

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

Sol. (a) Total energy = $\frac{\text{Potential energy}}{2} = \frac{-5.02}{2}$

Total energy = – 2.51 eV

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

Calculation of energy :

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

$$E_T = -\frac{Z^2}{n^2} \times \frac{2\pi^2 me^4}{h^2} = -K\frac{Z^2}{n^2}, \text{ where } K = \frac{2\pi^2 me^4}{h^2} = A$$

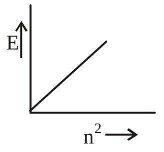
constant, whose values can be depicted as follows

(a) = 13.60 eV per atom
(b) = 2.179×10^{-11} ergs per atom
(c) = 313.6 kilocalories per mole

(d) = 21.79×10^{-19} joules per atom
(e) = 1312.1 kilojoules per atom
Note : Units – (a) 1 erg = 10^{-7} joule
(b) 1 erg = 6.2419×10^{11} eV
(c) 1 eV = 23.06 kilocalories
(d) 1 eV = 1.602×10^{-12} ergs
(e) 1 joule = 6.2419×10^{18} eV
(f) 1 kilocalorie = 4.184 kilojoule

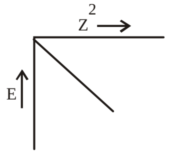
(i) $E = -\frac{Z^2}{n^2} \times 13.6$ eV ; If Z is a constant, then $E \propto -\frac{1}{n^2}$

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



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

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



$$\frac{E_1}{E_2} = \frac{Z_1^2}{Z_2^2}\ ;\ E = \frac{Z^2}{n^2} \times Rhc$$

(1) Kinetic energy = $\frac{Z^2}{n^2} \times Rhc$

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

Example 19 :

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) $E_2 > E_3 < E_4 < E_1$

Sol. **(3).** $E \propto -\frac{1}{n^2}$

Example 20 :

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

Sol. Total energy of the third orbit of H atom

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

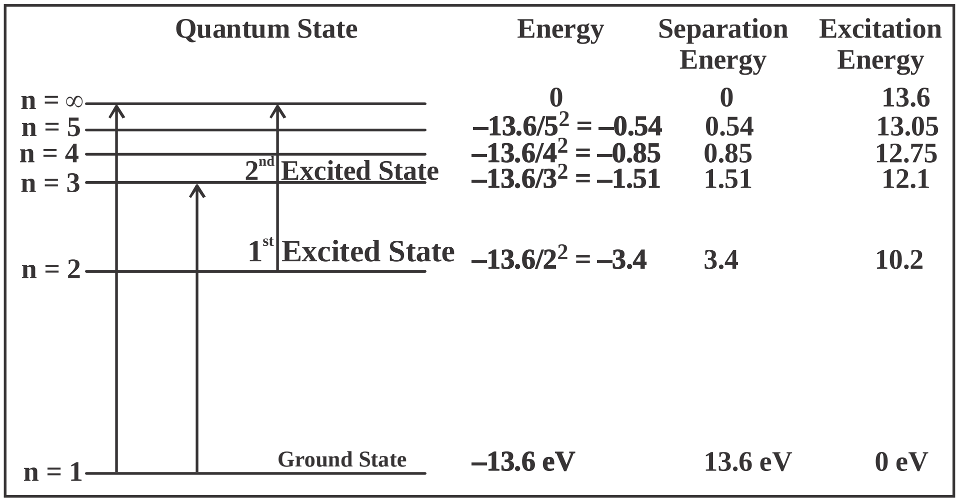
(1) Kinetic energy = – Total energy = – (–1.5) = + 1.5 eV
(2) Potential energy = 2 × Total energy = 2 × –1.5 = –3.0 eV

Example 21 :

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

Sol. Here n is a constant, therefore $\frac{E_1}{E_2} = \frac{Z_1^2}{Z_1^2} = \frac{(11)^2}{(1)^2} = 121 : 1$

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	$-13.6/4^2 = -0.85$	0.85	12.75
n = 3 (2nd Excited State)	$-13.6/3^2 = -1.51$	1.51	12.1
n = 2 (1st Excited State)	$-13.6/2^2 = -3.4$	3.4	10.2
n = 1 (Ground State)	–13.6 eV	13.6 eV	0 eV

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

Example 22 :

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

Sol. **(1).** An electron in ground state is in n = 1 orbit. Therefore the energy of the electron = – 13.6 eV

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.

Example 23 :

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

Sol. Second excited state n = 3

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

Excitation Potential :

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

Example 24 :

How much minimum energy should be absorbed by a hydrogen atom in ground state to reach excites state ?

Sol. 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

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.

Example 25 :

Find the maximum energy absorbed by hydrogen atom in its ground state.

Sol. $E_\infty - E_1$; $0 - (-13.6) = 13.6$ eV

Example 26 :

Find the energy required in the process $He^{+2} \rightarrow He^{+3}$.

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

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.

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.

$$\bar{\nu} = \frac{1}{\lambda} = R \times Z^2 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right), \text{ where } \frac{1}{\lambda} \text{ is wave number}$$

R = Rydberg constant,
n_1 = Number of lower energy level
n_2 = Number of higher energy level

Example 27 :

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 ?

Sol. The electron jumps from M orbit (n = 3) to K orbit (n = 1).

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

$$= R\left(\frac{1}{1} - \frac{1}{3^2}\right) = R\left(\frac{1}{1} - \frac{1}{9}\right) = R\left(\frac{9-1}{9}\right) = R \times \frac{8}{9}$$

Example 28 :

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

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

$$E = h\nu = \frac{hc}{\lambda} = \frac{6.62 \times 10^{-34} \times 3 \times 10^8}{4 \times 10^{-7}}$$

$$= 4.96 \times 10^{-19} \text{ joule}$$

EMISSION SPECTRUM AND ABSORPTION SPECTRUM

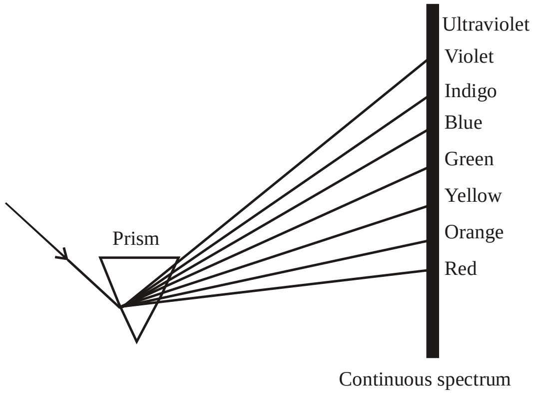
When a beam of white light passes through a slit or an aperture and then falls on a prism, it gets spilt into many coloured bands. The image of colours so obtained is known as a spectrum. A spectrum is of mainly three types viz.
(i) Emission spectrum (ii) absorption spectrum and
(iii) molecular spectrum

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

(a) By heating the substance at high temperature
(b) By passing electric current through a discharge tube having gaseous substance at very low pressure.
(c) By passing electric discharge through a metallic filament.

Emission spectra are of the following two types.
(i) Continuous spectrum and
(ii) line spectrum or atomic spectrum.

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



Continuous spectrum of white light

Line spectrum or Atomic spectrum : When atoms of a substance is excited, it emits radiations. These radiations are analyzed with the help of a spectroscope, then many fine bright lines of specific colours in a sequence are seen in the spectrum, which is not continuous, i.e. there is dark zone in between any two lines. Such a spectrum is called a line spectrum or atomic spectrum.

For example, neon single lamp, sodium vapour lamp, mercury vapour lamp, etc. emit light of different colours and they give specific line spectra.

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

Molecular Spectrum :

Molecular spectrum is given by molecules and it is also known as band spectrum. Three types of energy transitions are found in molecules. These are as follows : (i) electronic transitions, (ii) vibrational transitions and (iii) rotational transitions. Therefore, bands are obtained in the spectrum, which are actually groups of lines.

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

Series of Lines in Hydrogen Spectrum

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 $n_2 = 2, 3, 4, 5, 6, 7, 8 ...\infty$.

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.

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 $n_2 = 4, 5, 6, 7, 8 ...... \infty$.

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

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

Humphry 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 $n_1 = 6$ & $n_2 = 7, 8, 9, 10, 11.... \infty$.

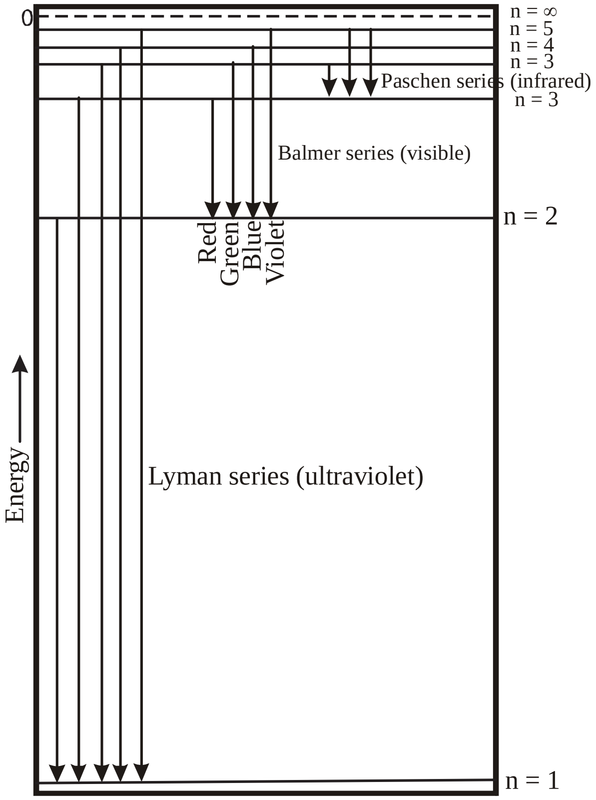


Figure : Transitions of the electron in the hydrogen atom (The diagram shows the Lyman, Balmer and Paschen series of transitions)

Example 29 :

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

Sol. The expression of maximum number is

$$\frac{n(n-1)}{2}=\frac{3(3-1)}{2}=\frac{6}{2}=3$$

Example 30 :

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) $\frac{1}{\lambda}=R\left(\frac{1}{16}-\frac{1}{\infty}\right)$

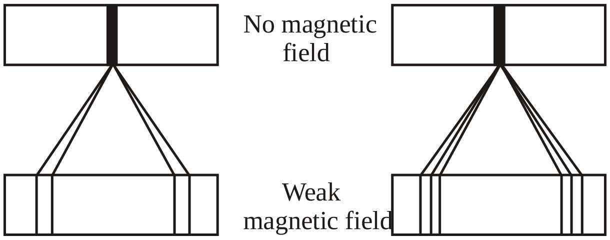
Sol. (1). $\bar{\nu}=\frac{1}{\lambda}=R\left(\frac{1}{9}-\frac{1}{\infty}\right)$

FAILURES OF BOHR'S ATOMIC MODEL

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

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

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



Representation of Zeeman effect

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

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

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

TRY IT YOURSELF-1

Q.1 Rutherford's experiment, which established the nuclear model of the atom, used a beam of –

(A) β-particle, which impinged on a metal foil & got absorbed

(B) γ-rays, which impinged on a metal foil & ejected electrons

(C) helium atoms, which impinged on a metal foil and got scattered

(D) helium nuclei, which impinged on a meal foil and got scattered

Q.2 Calculate total number of spectral lines and total spectral lines in Balmer series if electron undergoes a transition from 7th orbit to ground state.

Q.3 The energy of an electron in the first Bohr orbit of H atom is –13.6 eV. Possible energy value(s) of the excited state(s) for the electrons in Bohr orbits of hydrogen is (are)

(A) –3.4 eV (B) –4.2 eV

(C) –6.8 eV (D) +6.8 eV

Q.4 Calculate the wavelength of radiation emitted by electron during transition from second orbit to 4th orbit in He^+.

Q.5 Calculate the energy and the radius of the second orbit of He^+ ion.

Q.6 Calculate the wavelength of radiation emitted when an electron in hydrogen atom makes a transition from an energy level n = 3 to a level with n = 2.

Q.7 The energy of the electron in the first orbit of He^+ is -871.6×10^{-20} J. The energy of the electron in the first orbit of hydrogen would be –

(A) -871.6×10^{-20} J (B) -435×10^{-20} J

(C) -217.9×10^{-20} J (D) -108.9×10^{-20} J

Q.8 Wavelength of high energy transition of H-atoms is 91.2nm. Calculate the corresponding wavelength of He atoms.

Q.9 Calculate ionization energy of He^+ if electron is present in 2nd orbit.

Q.10 The kinetic energy of an electron in the second Bohr orbit of a hydrogen atom is [a_0 is Bohr radius]

(A) $\frac{h^2}{4\pi^2 ma_0^2}$ (B) $\frac{h^2}{16\pi^2 ma_0^2}$

(C) $\frac{h^2}{32\pi^2 ma_0^2}$ (D) $\frac{h^2}{64\pi^2 ma_0^2}$

ANSWERS

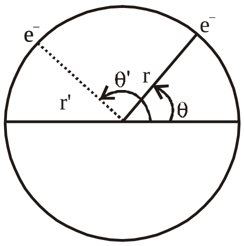
(1) (D)	**(2)** 21, 5	**(3)** (A)
(4) 1.216×10^{-7} m	**(5)** -2.18×10^{-18} J/atm, 1.058 Å	
(6) 656.64 nm	**(7)** (C)	**(8)** 22.85 nm
(9) 2.18×10^{-18} J/atom		**(10)** (C)

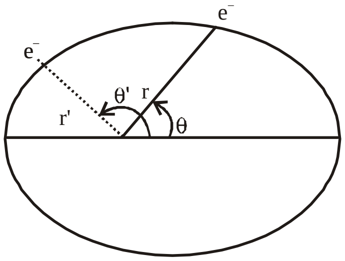
SOMMERFELD'S EXPANSION OF BOHR'S MODEL

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

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

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





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

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

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

(i) towards the radius, which is called radial component P_r, and (ii) in the perpendicular direction to radius, which is called azimuthal component P_k.

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

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

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) Frequency (ν)
(4) Amplitude (A) (5) Wave velocity (c or v)

The aforesaid properties of a wave have the following relationship $\nu = \frac{1}{T}$ and $c = \frac{\lambda}{T}$ or $c = \nu\lambda$

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

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

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

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

Wave Number : The reciprocal of wavelength is called wave number. It is represented by $\bar{\upsilon}$.

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

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

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

ELECTROMAGNETIC SPECTRUM

If all the components of Electromagnetic Radiation (EMR) are arranged in order of decreasing or increasing wavelengths or frequencies, the pattern obtained is known as Electromagnetic Spectrum.

S.N.	Name	Wavelength(Å)	Source
1.	Radio wave	$3\times10^{14} - 3\times10^{7}$	Alternating current of high frequency
2.	Microwave	$3\times10^{7} - 6\times10^{6}$	Klystron tube
3.	Infrared (IR)	$6\times10^{6} - 7600$	Incandescent objects
4.	Visible	7600–3800	Electric bulbs, sun rays
5.	Ultraviolet (UV)	3800–150	Sun rays, arc lamps with mercury vapours
6.	X-Rays	150–0.1	Cathode rays striking metal plate
7.	γ-Rays	0.1–0.01	Secondary effect of radioactive decay
8.	Cosmic Rays	0.01–zero	Outer space

Example 31 :

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

Sol. $\because c = \nu\lambda$

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

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

PLANCK'S QUANTUM THEORY

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

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

THE DUAL NATURE OF MATTER

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

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

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

$$\lambda \propto \frac{1}{p} \text{ or } \lambda = \frac{h}{p}$$

Here h = Planck's constant, p = Momentum of electron

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

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

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

But according to Einstein's equation

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

BOHR'S THEORY AND DE BROGLIE CONCEPT

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

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

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

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

or $2\pi r = \frac{nh}{mv}$ ($\because$ mv = p momentum)

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

(e) $\because 2\pi r = \frac{nh}{mv}$ or $mvr = \frac{nh}{2\pi}$ $\because$ mvr = Angular momentum

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

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

Example 32 :

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/s?

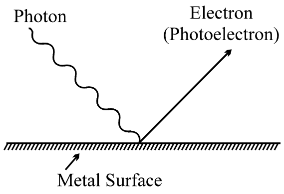
Sol. $\lambda = \frac{h}{m \times c}$ or $\frac{h}{c\lambda}$ ($\because \lambda = 5894$ Å or 5894×10^{-10} m)

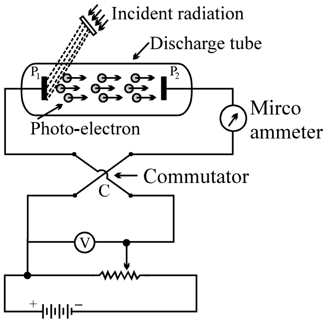
$$m = \frac{6.652 \times 10^{-34}}{3 \times 10^8 \times 5894 \times 10^{-10}}$$

or $\frac{6.652}{17682} \times 10^{-32} = 3.746 \times 10^{-36}$ Kg

PHOTOELECTRIC EFFECT

When light of particular wavelength of frequency falls on a metal, electrons are emitted form it. This phenomenon of emission of electrons from a metal surface is called photoelectric effect. The electrons emitted are called photoelectrons.





Laws of Photoelectric Effect :

(a) The number of photoelectrons emitted from a metal surface in a unit time (the rate of emission of photoelectrons) is proportional to the intensity of incident light.

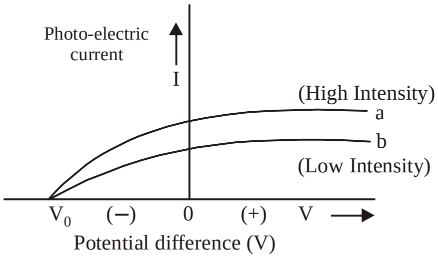
(b) The maximum kinetic energy of the emitted photoelectrons increases with increasing frequency or decreases with increasing wavelength of the incident light. The kinetic energy of the electrons does not depend upon the intensity of the incident light.

(c) The emission of photoelectrons occurs upto a definite minimum frequency (maximum wavelength) of the incident light. This minimum frequency is called threshold frequency and the maximum wavelength is called threshold wavelength. Threshold wavelength depends upon the maximum wavelength is called threshold wavelength.
Threshold wavelength depends upon the nature of the substance and it is different for different metals.

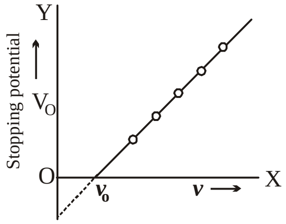
(d) Within the limits of accuracy (~ 10^{-9} second) there is no time lag between the incidence of light at the metal surface and the emission of electrons from the metal surface, whatever is the intensity of incident light.

Photoelectric current and stopping potential :

(a) The negative potential of the plate relative to the electron emitter, i.e., cathode at which photoelectric current becomes zero is called stopping potential. It measures the maximum kinetic energy of the photoelectrons. $E_{max} = \frac{1}{2}mv^2_{max} = eV_0$



(b) The stopping potential does not depend upon the intensity of incident light.



Work function : The minimum energy per electron given to the free electrons of the metal which enables them to cross the potential barrier present at the surface of the metal, is called work function.

$$W_0 = h\nu_0$$

Einstein's photoelectric equation :

(a) $\frac{1}{2}mv^2_{max} = eV_0 = h(\nu - \nu_0)$

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

(b) Maximum velocity of emitted electrons :

$$v_{max} = \sqrt{\frac{2h(\nu - \nu_0)}{m}} = \sqrt{\frac{2hc(\lambda_0 - \lambda)}{m\lambda\lambda_0}}$$

(c) $v_{max} = \sqrt{\frac{2eV_0}{m}}$

Example 33:

A photon of wavelength 5000 Å strikes a metal surface, the work function of the metal being 2.20 eV. Calculate

(i) The energy of the photon in eV.

(ii) The kinetic energy of the emitted photo electron.

(iii) The velocity of the photo electron.

Sol. (i) Energy of the photon,

$$E = h\nu = \frac{hc}{\lambda} = \frac{(6.6 \times 10^{-34}\,Js)\,(3 \times 10^8\,ms^{-1})}{5 \times 10^{-7}\,m}$$

$$= 3.96 \times 10^{-19}\,J;\ 1eV = 1.6 \times 10^{-19}\,J$$

$$\text{Therefore } E = \frac{3.96 \times 10^{-19}\,J}{1.6 \times 10^{-19}\,J/eV} = 2.475\,eV$$

(ii) Kinetic energy of the emitted photon electron work function = 2.20 eV

Therefore, KE = 2.475 – 2.20 = 0.275 eV = 4.4× 10^{-20} J

(iii) Velocity of the photon electron KE

$$= \frac{1}{2}mv^2 = 4.4 \times 10^{-20} = 0.275\,eV = 4.4 \times 10^{-20}\,J$$

$$\text{Therefore, velocity } (v) = \sqrt{\frac{2 \text{ x } 4.4 \text{ x } 10^{-20}}{9.1 \text{ x } 10^{-31}}}$$

$$= 3.11 \times 10^5\,ms^{-1}$$

QUANTUM MECHANICAL THEORY OF ATOM

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

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

(i) Heisenberg's uncertainty principal and

(ii) Schrodinger's wave equation

Heisenberg's Uncertainty Principle

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

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

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

Here Δx is uncertainty of position,
Δp is uncertainty of momentum and
h is Planck's constant

Example 34 :

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 ?

Sol. Given that : Δp = 1 × 10^{-2} gm cm/sec. = 1 × 10^{-7} Kg m/sec.
h = 6.6252 × 10^{-34} Js

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

$$\text{or} \quad \Delta x = \frac{6.6252 \times 10^{-34}}{2 \times 3.14 \times 10^{-7}}\ \text{m} = 1.054 \times 10^{-27}\text{m}$$

SCHRONDINGER'S WAVE EQUATION

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

$$\nabla^2 \psi + \frac{8\pi^2 m}{h^2}(E - v)\,\phi = 0\ ; \quad \nabla^2 = \frac{d^2}{dx^2} + \frac{d^2}{dy^2} + \frac{d^2}{dz^2}$$

where m = Mass of electron, h = Planck constant,
E = Total energy of electron, v=Potential energy of electron,
ψ = Wave function, ∇ = Laplacian Operator

Probability distributive curve : An atomic orbital is a one electron wave function ψ (r, θ, φ) obtained from the solution of schrondinger wave equation this am be represented by ψ (r, θ, φ) = R(r) Θ (θ) ϕ (φ), where R = Radial function, which gives the dependence of orbital upon distance r.

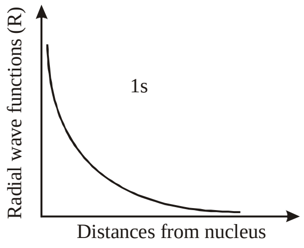
Θ (θ) & θ (φ) $\Rightarrow$ Angular function giving the angular dependence of orbital on θ and φ respectively.

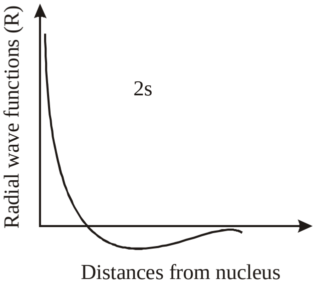
The orbital wave function ψ does not have any physical significance but the square of (ψ) provides the information regarding the probability of electron at a point in an atom. To draw the representation of variation of ψ^2 in space we required following functions :

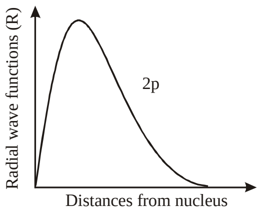
(a) Radial wave function
(b) Radial probability density (R^2)
(c) Radial probability function ($4\pi r^2R^2$)

(a) Radial wave functions (R) : By drawing these curve we can find the node in 2s, radial function. At node the value of radial function changes from positive to negative.

By drawing these we can get the information that how the radial wave function changes with distance r.

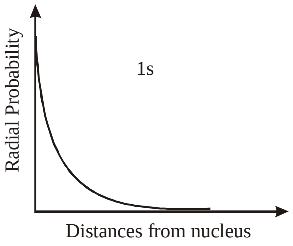


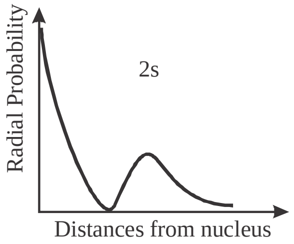


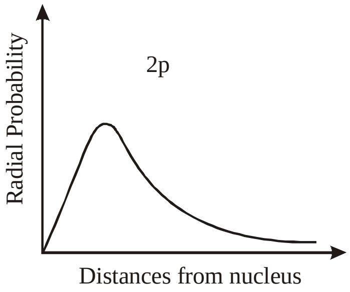


(b) Radial probability density (R^2) : The square of radial wave function R^2 for an orbital give the radial density & this radial density give the probability density of finding the electron.

By drawing the curve we can obtain the useful information about probability density or relative electron density at a point as a function of radius.

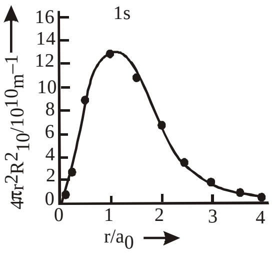


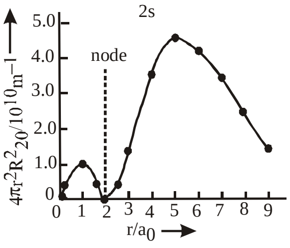


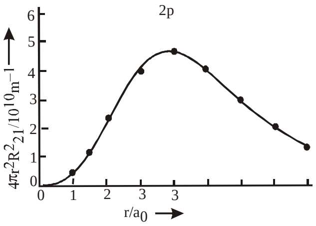


(c) Radial probability function (4®r^2R^2) : The shape of an atom is assumed to be spherical so it better to discuss the probability of finding the electron in a spherical shell between the radius (r + dr) and r. This probability which is independent of direction is called radial probability & equal to ($4\pi r^2R^2$).

By the drawing the curve we can obtain the information regarding the variation of radial probability function ($4\pi r^2R^2$) with distance r.







TRY IT YOURSELF-2

Q.1 Calcualte the wavelength, frequency and wave number of a light whose period is 4×10^{-8} s.

Q.2 Calculate the momentum of a particle which has wavelength of 2Å.

Q.3 A 100 watt bulb is emitting monochromatic light having wavelength 6000 Å. Calculate the number of photons emitted by bulb in five minute.

Q.4 Calculate the uncertainty in position of an electron if the uncertainty in its velocity is 5.7×10^{5} m/s.

Q.5 The threshold frequency for the photoemission of electrons from potassium metal is $5.3 \times 10^{14}\ s^{-1}$. Will the photon of a radiation having energy 3.3×10^{-19} J cause photoelectric effect ?

Q.6 Calculate the de-Broglie wavelength of an electron that has been accelerated from rest through a potential difference of 2 kV.

Q.7 Two lines, one violet and the other red, in a spectrum have wavelengths equal to 4000 Å and 6000 Å, respectively. Calculate the energies of photons transmitted by them.
(Planck's constant = 6.626×10^{-34} Joule sec)

Q.8 Calculate the product of uncertainities of displacement and velocity of a moving electron of mass 9.1×10^{-28}g.

Q.9 Light of wavelength 400nm strikes on cesium metal which has a photoelectric work function of 2.13 eV. Find out the maximum kinetic energy of the photoelectrons.

Q.10 An electron is moving with a kinetic energy of 4.55×10^{-25} J. Calculate de-Broglie wavelength for it.

ANSWERS

(1) 12m, $2.5 \times 10^{7}\ s^{-1}$, $8.3 \times 10^{-2}\ m^{-1}$
(2) 3.3×10^{-24} kg ms^{-1} **(3)** 9.0×10^{22} photons
(4) 1.0×10^{-10} m **(5)** No
(6) 1.93×10^{-11} m
(7) 4.970×10^{-19} joule, 3.313×10^{-19} joule
(8) $5.77 \times 10^{-5}\ m^{2}\ s^{-1}$ **(9)** 1.56×10^{-19} J
(10) 7.25×10^{-7} m

QUANTUM NUMBERS

The position of any electron in any atom can be ascertained with the help of quantum numbers. In an atom, the shell consists of sub-shells and the sub-shell consists of orbital can accommodate only two electrons, which are in opposite spins.

Principal Quantum Number (n) :

(a) Principal quantum number indicates the shell or energy level or orbit.
(b) An atoms has K, L, M, N, O, P, Q, etc. shells.
(c) Principal quantum number also gives information about the radius of size.
(d) Principal quantum number also gives information about the distance of an electron from the nucleus in an atom
(e) Principal quantum number also given information about the energy of an electron.
(f) Principal quantum number also gives information about the velocity of an electron.
(g) In any orbit, the number of orbitals is given by n^2 and number of electrons is given by $2n^2$. This is called Bohr-Bury rule.

Example 35 :

Find the principal quantum number for the last electron of $_{11}Na$?

Sol. $_{11}Na = 1s^2, 2s^2, 2p^6, 3s^1$; n = 3

Example 36 :

Which of the following should have greater size ?

(1) 1s (2) 2s
(3) 3s (4) 4s

Sol. **(4).** n = 4 for 4s

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 $2\ell + 1$ and the number of electrons is determined by the expression $2(2\ell + 1)$.
(g) $\ell = 0 \rightarrow$ s Sub-shell $\rightarrow$ Spherical
$\ell = 1 \rightarrow$ p Sub-shell $\rightarrow$ Dumb-bell
$\ell = 2 \rightarrow$ d Sub-shell $\rightarrow$ Double dumb-bell
$\ell = 3 \rightarrow$ 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 $\mu_i = \sqrt{\ell(\ell+1)} \times \frac{h}{2\pi}$
ℓ = Azimuthal quantum number and h = Planck's constant

Example 37 :

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

Sol. **(1).** $n + \ell = 7$
$7 + 0 = 7s$; $6 + 1 = 6p$; $5 + 2 = 5d$; $4 + 3 = 4f$

Example 38 :

The sub-shell 2d is not possible because

(1) $n \neq \ell$ (2) $\ell > n$
(3) $n < \ell$ (4) None of these

Sol. (1). For sub-shell 2d, n = 2 and ℓ = 2 and the values of n and ℓ can never be equal.

Example 39 :

Which of the following orbitals should be nearest to the nucleus ?

(1) 5s (2) 6p

(3) 3d (4) 4d

Sol. (3). n = 3 will be nearest to the nucleus.

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 – ℓ to + ℓ.

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

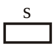
m = 2ℓ + 1, where m is total value of magnetic quantum number and l is the value of azimuthal quantum number.

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

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

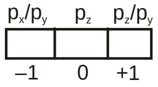
(iii) For d sub-shell, ℓ = 2. Thus, m = 2 × 2 + 1 = 5 and therefore d sub-shell consists of five orbitals called d_{xy}, d_{yz}, d_{z^2}, d_{xz} and $d_{x^2-y^2}$ orbitals.

* For s sublevel, ℓ = 0., for s orbital, the value of m is 0.

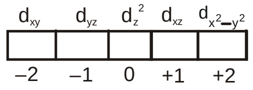


m = 0

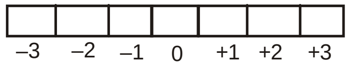
* For p sub-level, ℓ=1. Thus, the values of m for p orbitals



* For d sub-level, ℓ= 2. Thus, the values of m for d orbitals



* For f sub-level, ℓ = 3. Thus, the values of m for f orbitals



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

Example 40 :

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

Sol. (1). 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 41 :

Orbital having n = 6, ℓ = 2 and m = 0 will be designated as

(1) 6 d_{z^2} (2) 6 $d_{x^2-y^2}$

(3) 6d_{xy} (4) 6p_z

Sol. (1). For 6th of P energy level, ℓ = 2 is for d sub-level, and m = 0 for d_{z^2} orbital.

Example 42 :

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

(1) $2p_z$ (2) $2p_x$

(3) $2p_y$ (4) 3 d_{z^2}

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

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 + 1/2 or a sign [↑]. Anticlockwise spin of the electron is denoted by s = – 1/2 or [↓].

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

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

(f) The angular momentum of an electron is not only due its motion around the nucleus in an energy level but also due to its rotation along its own axis. The angular momentum that arises due to rotation of an electron along its axis, is called spin angular momentum and is depicted by the symbol µs. The value of µs can be found out with the help of the following expression. $\mu s = \sqrt{s(s+1)} \times \frac{h}{2\pi}$ where s is spin quantum number. In this expression the value of s is always taken as 1/2 and not – 1/2.

Example 43 :

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

(1) ℓ_x (2) $n\ell^x$

(3) nm^x (4) None of these

Sol. (2). 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.

Example 44 :

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

Sol. (2). All the energy states in which n is greater than 1 are called excited states.

AUFBAU PRINCIPLE

Aufbau is a German word that means to build up. Therefore, electrons are filled up in accordance with this principle. 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. 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. The electrons are filled in a sub-shell according in n + ℓ rule.

PAULI'S EXCLUSION PRINCIPLE

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

For example :

(a) $6s^1$ and $6s^2$

$6s^1$		$6s^2$
n = 6		n = 6
ℓ = 0		ℓ = 0
m = 0		m = 0
s = + 1/2		s = – 1/2

(b) $4p^2$ and $4p^5$

$4p^2$		$4p^5$
n = 4		n = 4
ℓ = 1		ℓ = 1
m = 0		m = 0
s = + 1/2		s = – 1/2

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

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

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

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

Example

$1s^1$	and	$1s^2$
n = 1		n = 1
ℓ = 0		ℓ = 0
m = 0		m = 0
s = + 1/2		s = – 1/2

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

Example 45 :

Pauli exclusion principle applies to

(1) H (2) H^+

(3) H^- (4) None of the above

Sol. (3). Since, H has one electron and H^+ has no electron, therefore Paulie principal does not apply to them. However, H^- has two electrons, hence this principle applies on it.

Example 46 :

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

Sol. (4). In (1) & (2) two electrons have same set of quantum nos.

Example 47 :

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) $1s^1, 2s^1, 2p^1$

Sol. (4). 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 occupier in a orbital.

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 member of electrons and the degenerate orbitals should be more than one

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

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

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

Example 48 :

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

(1) $C^6 = 1s^2, 2s^2$ $\boxed{\uparrow\uparrow}\boxed{}\boxed{}$

(2) $O^8 = 1s^2, 2s^2$ $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{}$

(3) $N^7 = 1s^2, 2s^2,$ $\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{}$

(4) $F^9 = 1s^2, 2s^2$ $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}$

Sol. (4). Configuration of C^6 should be $2p_x^1\ 2p_y^1$ instead of $2p_x^2$

Configuration of O^8 should be $2p_x^2\ 2p_y^1\ 2p_z^1$ instead of $2p_x^2\ 2p_y^2$

Configuration of N^7 should be $2p_x^1\ 2p_y^1\ 2p_z^1$ instead of $2p_x^2\ 2p_y^1$

Configuration of F^9 $2p_x^2\ 2p_y^2\ 2p_z^1$ is correct because two out of the three degenerate p orbitals are fully-filled, one is half-filled and there is no unfilled p orbital.

n + ℓ RULE

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

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

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

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

Example 49 :

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

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

(3) 7s > 6p < 5d < 4p (4) 4f < 5d < 6p > 7s

Sol. (1). n + ℓ = 7

7 + 0 = 7s	Order of energy
6 + 1 = 6p	4f < 5d < 6p < 7s
5 + 2 = 5d	
4 + 3 = 4f	

Example 50 :

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

Sol. (4). The electron goes to 4p after filling up to 3d.

Example 51 :

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

Sol. (1). 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.

Increasing order of energy

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

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

Exceptions to n + ℓ Rule

There are mainly two exceptions of n + *l* rule.

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

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

Specific Electronic Configuration

Element		=Atomic number	Expected configuration	Atcual 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.

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

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

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

MODE OF FILLING UP TO ELECTRONS

Writing the configuration of ions

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

Example 52 :

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

Sol. **(1).** $(3-2)s^2\ (3-1)s^2\ p^6\ 3s^2\ 3p^2$

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

The atomic number of the atom is 14

Example 53 :

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

Sol. **(1)** X^{-3} has 14 protons, i.e. X also has 14 protons and therefore 14 electrons.

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

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

Example 54 :

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

Sol. **(2).** 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 55 :

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

Sol. **(2).** Electronic configuration of P in H_3PO_4 is $[Ne]\ 3s^2\ 3p^6$

DIFFERENCE BETWEEN ORBIT AND ORBITAL

S.N.	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

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 is an orbital is 95% to 98%.

s-Orbital

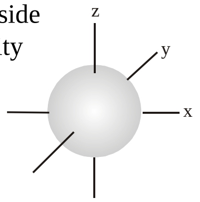
(a) Only one s-orbital is possible in an orbit because $\ell = 0$ and m = 0 for it.

(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)



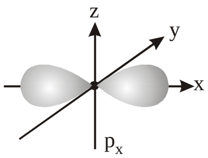
p-Orbital

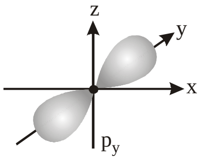
(a) For p orbitals, $\ell = 1$ and m = – 1, 0, + 1. Thus, it can have three configurations, which are distributed in x, y and z axes. Therefore, there are three p-orbitals, which are dumbbell, shaped.

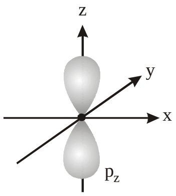
(b) Each p-orbital has two lobes and the probability of finding electron inside these two lobes is equal. The plane perpendicular to the axis two lobes and passing through the point where these two lobes join, is the nodal plane of p-orbital, because the probability of finding electron in this plane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

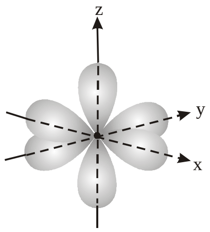
(c) The value of nodal planes for each of the p_x, p_y and p_z orbitals is same and these nodal planes are present in xy, yz and xz planes, respectively.

(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.









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 :

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

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

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

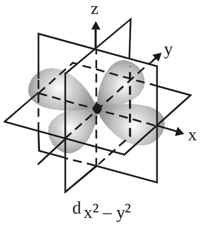
(iv) The double dumbbell of $d_{x^2-y^2}$ orbital are directed on x and y axes

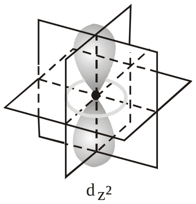
(v) d_{z^2} orbital is composed of one dumbbell and one ring. The dumbbell is situated on z axis and the ring is present on its middle part.

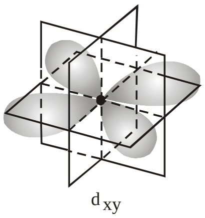
The aforesaid five d orbitals can be classified into the following two categories.

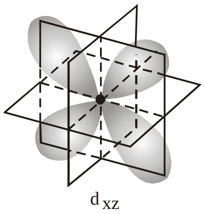
(i) t_{2g} Orbitals (d_{xy}, d_{xz} and d_{yz}) – In these, the electron density is concentrated in-between the axes. These are also called grade orbitals.

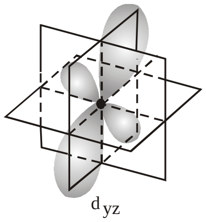
(ii) e_g Orbitals ($d_{x^2-y^2}$ and d_{z^2}) – In these, the electron density is concentrated on the axes.











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.

TRY IT YOURSELF-3

Q.1 Find the number of radial node(s) and angular node (s) in 4p orbital?

Q.2 An electron is in a 3d orbital. What possible values for the quantum numbers n, l, m and s can it have ?

Q.3 The number of radial nodes of 3s and 2p orbitals are respectively

(A) 2,0 (B) 0, 2

(C) 1, 2 (D) 2, 1

Q.4 What designation is given to an orbital having

(i) $n = 2, l = 0$ (ii) $n = 3, l = 1$

(iii) $n = 5, l = 2$ (iv) $n = 6, l = 3$

Q.5 Ground state electronic configuration of nitrogen atom can be represent by –

1. ↑↓ ↑↓ ↑ ↑ ↑

2. ↑↓ ↑↓ ↑ ↓ ↑

3. ↑↓ ↑↓ ↑ ↓ ↓

4. ↑↓ ↑↓ ↓ ↓ ↓

(A) 1 only (B) 1,2 only

(C) 1, 4 only (D) 2, 3 only

Q.6 The quantum numbers + 1/2 and –1/2 for the electron spin represents :

(A) rotation of the electron in clockwise and anticlockwise direction respectively.

(B) rotation of the electron in anticlockwise and clockwise direction respectively.

(C) magnetic momentum of electron pointing up and down respectively.

(D) two quantum mechanical spin states which have no classical analogues.

Q.7 An atom of an element has 2K, 8L and 5M electrons. Find out the following from the data:
(i) Atomic number
(ii) Total number of s- and p-electrons
(iii) Total number of protons in the nucleus.
(iv) Valency of the element.

Q.8 Which of the following set of quantum numbers is not permitted?
(i) n = 3, l = 2, m = 0, s = –1/2
(ii) n = 3, l = 2, m = +2, s = –1/2
(iii) n = 3, l = 3, m = –3, s = +1/2
(iv) n = 3, l = 4, m = +4, s = +1/2

Q.9 Write the electronic configurations for the elements with atomic numbers 11, 15, 24, 28, 31.

Q.10 In an atom, the total number of electrons having quantum numbers n = 4, $|m_\ell| = 1$ and $m_s = -1/2$ is

Q.11 The number of nodal planes in p_x orbital is –
(A) one (B) two
(C) three (D) zero

Q.12 The electrons identified by quantum number n and ℓ (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 –
(A) (iv) < (ii) < (iii) <(i) (B) (ii) < (iv) < (i) < (iii)
(C) (i) < (iii) < (ii) < (iv) (D) (iiii) < (i) < (iv) < (ii)

Q.13 Give the electronic configurations of the following ions:-
(i) H^- (ii) F^-
(iii) Mg^{2+} (iv) O^{2-}

ANSWERS

(1) 2, 1
(2) n = 3, l = 2, m = –2, –1, 0, + 1, +2, s = +1/2 or –1/2
(3) (A) **(4)** (i) 2s orbital., (ii) 3p orbital, (iii) 5d orbital. (iv) 6f orbital. **(5)** (C)
(6) (D) **(7)** (i) 15, (ii) 15, (iii) 15, (iv) 4
(8) (i) permitted, (ii) permitted, (iii) not permitted (iv) not permitted.
(9) (i) (Z = 11) $1s^2\,2s^2\,2p^6\,3s^1$
(ii) (Z = 15) $1s^2\,2s^2\,2p^6\,3s^2\,3p_x^{\,1}\,3p_y^{\,1}\,3p_z^{\,1}$.
(iii) (Z = 24) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^5\,4s^1$
(iv) (Z = 28) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^8\,4s^2$
(v) (Z = 31) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^{10}\,4s^2\,4p^1$
(10) 6 **(11)** (A) **(12)** (A)
(13) (i) $1s^2$ (ii) $1s^2\,2s^2\,2p_x^{\,2}\,2p_y^{\,2}\,2p_z^{\,2}$.
(iii) $1s^2\,2s^2\,2p_x^{\,2}\,2p_y^{\,2}\,2p_z^{\,2}$ (iv) $1s^2\,2s^2\,2p_x^{\,2}\,2p_y^{\,2}\,2p_z^{\,2}$.

IMPORTANT POINTS

*

Type of orbital	No. of peaks or maxima (n – l)	No. of spherical nodes (n – l – 1)	No. of angular node or nodal plane (l)	No. of nodes (n – 1)
s	n	n – 1	0	n – 1
p	n – 1	n – 2	1	n – 1
d	n – 2	n – 3	2	n – 1
f	n – 3	n – 4	3	n – 1

S.No.	Radius	Velocity	Energy	Wavelength
1.	$r = \frac{n^2h^2}{4\pi^2mZe^2}$	$V = \frac{2\pi Ze^2}{nh}$	$E = -\frac{Ze^2}{2r}$	$R = \frac{2\pi^2me^4}{Ch^3}$
2.	$r = \frac{n^2}{Z} \times 0.529Å$	$V = \left(\frac{Ze^2}{rm}\right)^{1/2}$	$E = -\frac{2\pi^2mZ^2e^4}{n^2h^2}$	$\frac{1}{\lambda} = R \times Z^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$
3.	$r = \frac{n^2}{Z} \times 0.0529\,\text{nm}$	$V = \frac{nh}{2\pi mr}$	$E = -Rch \times \frac{z^2}{n^2}$	$E = h\upsilon$
4.	$r \propto n^2$ (Z const)	$V \propto \frac{1}{n}$ (Z const)	$E = -\frac{z^2}{n^2} \times 313.6\,\text{Kcal}$	$\lambda = \frac{h}{mc}$
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)	$c = \nu\lambda$
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)	$E = mc^2$

- * The maximum capacity of a main energy shell is equal to $2n^2$ electron.
- * The maximum capacity of a subshell is equal to $2(2\ell + 1)$ electrons.
- * Number of sub-shells in a main energy shell is equal to the value of n.
- * No. of orbitals in a main energy shell is equal to n^2.
- * One orbital cannot have more than two electrons.
- * Energy of 2s orbital of hydrogen atom is greater than that of 2s orbital of lithium and that of lithium is greater than that of sodium and so on, that is,

$$E_{2s}(H) > E_{2s}(Li) > E_{2s}(Na) > E_{2s}(K).$$

- * The nuclear density is nearly constant and is equal to 2.04×10^{17} kg/m^3
- * The maximum numbers of electrons that can be associated with $n = 3, \ell = 1$ and $m = -1$ is 2.
  [$n = 3, \ell = 1$ and $m = -1$ represent a 3p-orbital
  so maximum 2 electrons can be accommodate in it.]
- * The angular momentum of electron in 'd' orbital is equal to $\sqrt{6}\,\hbar$.
- * Values of quantum number ℓ and m for n = 1 to 4 principal shell

Principal shell	**Value of n**	**Value of *l* 0 to n–1**	**Value of m *-l* to *+l***	**Total no of electron**
K	1	$l = 0$ (s)	m = 0	2
L	2	$l = 0$ (s)	m = 0	
		$l = 1$ (p)	m = –1, 0, +1	8
M	3	$l = 0$ (s)	m = 0	
		$l = 1$ (p)	m = –1, 0, +1	18
		$l = 2$ (d)	m = –2, –1, 0, +1,+2	
N	4	$l = 0$ (s)	m = 0 (s)	
		$l = 1$ (p)	m = –1, 0, +1	
		$l = 2$ (d)	m = –2, –1, 0, +1,+2	32
		$l = 3$ (f)	m = –3, –2, –1, 0, +1,+2, +3	

ADDITIONAL EXAMPLES

Example 1 :

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

Sol. Total no. of molecules of $CO_2 = 12$

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

Similarly six molecules with C^{13} isotopes.

Example 2 :

Which of the following set of quantum nos. are not permitted

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

Sol. (a) This set of quantum number is permitted.
(b) This set of quantum number is not permitted as value of 's' cannot be zero.
(c) This set of quantum number is not permitted as the value of '*l*' cannot be equal to 'n'.
(d) This set of quantum number is not permitted as the value of 'm' cannot be greater than '*l*'.

Example 3 :

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

Sol. Kinetic energy of electron = $(1/2)\,mu^2$

Also from Bohr's concept K.E. = $\frac{1}{2}\frac{Ze^2}{r_n}$

$$\therefore \frac{1}{2}mu^2 = \frac{1}{2}\frac{Ze^2}{r_n} \;;\quad \nu = \sqrt{\left(\frac{Ze^2}{mr_n}\right)}$$

Example 4 :

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

Sol. Energy emitted by sodium lamp in one sec.

$$= \text{Watt} \times \text{sec} = 60 \times 1\ \text{J}$$

Energy of photon emitted

$$= \frac{hc}{\lambda} = \frac{6.626 \times 10^{-34} \times 3 \times 10^{8}}{5893 \times 10^{-10}} = 3.37 \times 10^{-19}\ \text{J}$$

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

$\therefore$ No. of photons emitted in 10 hours

$= 17.8 \times 10^{19} \times 10 \times 60 \times 60 = 6.41 \times 10^{24}$

Example 5 :

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

Sol. In M.K.S. system, $E_n = -\frac{2\pi^2 Z^2 me^2}{(4\pi\varepsilon_0)^2\, n^2\, h^2}$

$$= \frac{2 \times (3.14)^2 \times (1)^2 \times 9.108 \times 10^{-31} \times (1.602 \times 10^{-19})^4}{(1.11264 \times 10^{-10})^2 \times (2)^2 \times (6.625 \times 10^{-34})^2}$$

$= 5.443 \times 10^{-19}$ joule

Example 6 :

Find the shortest wave length in H spectrum of Lymen series when $R_H = 109678\ cm^{-1}$ is

Sol. For Lymen series $n_1 = 1$

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

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

$$= 109678 = 911.7 \times 10^{-8} = 911.7\ Å$$

Example 7 :

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

Sol. Energy associated with a photon of 242 nm

$$= \frac{6.625 \times 10^{-34} \times 3.0 \times 10^8}{242 \times 10^{-9}} = 8.21 \times 10^{-19} \text{ joule}$$

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

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

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

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

Example 8 :

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

(a) $n = 4, \ell = 2, m = 0$ (b) $n = 3$

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

Sol. (a) $\ell = 2$ means d-subshell and m = 0 refer to dz^2 orbital

$\therefore$ Number of electrons are 2.

(b) For $n = 3, \ell = 0, 1, 2$

$\ell = 0$	$m = 0$	2 electrons
$\ell = 1$	$m = -1$	6 electrons
$\ell = 2$	$m = -2, -1, 0, +1, +2$	10 electrons
Total electrons		18 electrons

Alternatively, number of electrons for any energy level is given by $2n^2$ i.e. $2 \times 3^2 = 18$ electrons

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

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

Example 9 :

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

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

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

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

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

Example 10 :

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

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

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

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

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

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

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

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

Example 11 :

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

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

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

Since λ is same

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

$$\therefore \quad \left[\frac{1}{1^2} - \frac{1}{2^2}\right] = \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] \ \therefore n_1 = 1 \ \text{ and } \ n_2 = 2$$

Example 12 :

Find the ionization energy of H-atom is 13.6 eV. Find the ionization energy of Li^{+2} ion

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

$= 13.6 \times 9 = 122.4$ eV

Example 13 :

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

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

$$\Delta x.\Delta u \geq \frac{h}{4\pi m} \quad \therefore \quad \Delta u = \frac{6.626 \times 10^{-27}}{4 \times 3.14 \times 9.108 \times 10^{-28} \times 10^{-13}}$$

$= 5.79 \times 10^{12}$ cm/sec

i.e., order of velocity of electron will be 100 times greater than the velocity of light which is impossible. Thus, possibility of electron to exist is nucleus is zero.

QUESTION BANK

CHAPTER 2 : STRUCTURE OF ATOM

EXERCISE - 1 [LEVEL-1]

Choose one correct response for each question.

PART-1 : SUB-ATOMIC PARTICLES

Q.1 The number of electrons which will together weight one gram is –

(A) 1.098×10^{27} electrons (B) 9.1096×10^{31} electrons
(C) 1 electron (D) 1×10^{4} electrons

Q.2 The ratio of charge to mass of an electron in coulombs per gram was determined by J.J.Thomson. He determined this ratio by measuring the deflection of cathode rays in electric and magnetic fields. What value did he find for this ratio?

(A) -1.76×10^{8} coulombs/g
(B) 1.76×10^{-8} coulombs/g
(C) -1.76×10^{10} coulombs/g
(D) -1.76×10^{-10} coulombs/g

Q.3 Which of the following statements is not correct about the characteristics of cathode rays?

(A) They start from the cathode and move towards the anode.
(B) They travel in straight line in the absence of an external electrical or magnetic field.
(C) Characteristics of cathode rays do not depend upon the material of electrodes in cathode ray tube.
(D) Characteristics of cathode rays depend upon the nature of gas present in the cathode ray tube.

Q.4 How many number of electrons are present in a particle which carries a charge of 5.5×10^{-16} C

(A) 3432 (B) 1560
(C) 8240 (D) 2432

Q.5 Which of the following statements about the electron is incorrect?

(A) It is a negatively charged particle.
(B) The mass of electron is equal to the mass of neutron.
(C) It is a basic constituent of all atoms.
(D) It is a constituent of cathode rays.

PART-2 : ATOMIC MODELS

Q.6 Which of the following observations was not correct during Rutherford's scattering experiment?

(A) Most of the α-particles passed through the gold foil undeflected.
(B) A small fraction of the α-particles was deflected by small angles.
(C) A large number of the α-particles were bounced back.
(D) A very few α-particles (~1 in 20,000) were bounced back.

Q.7 Which of the following conclusions could not be derived from Rutherford's α-particle scattering experiment?

(A) Most of the space in the atom is empty.
(B) The radius of the atom is about 10^{-10} m while that of nucleus is 10^{-15} m.
(C) Electrons move in a circular path of fixed energy called orbits.
(D) Electrons and the nucleus are held together by electrostatic forces of attraction.

Q.8 Two atoms are said to be isobars if

(A) they have same atomic number but different mass number.
(B) they have same number of electrons but different number of neutrons
(C) they have same number of neutrons but different number of electrons.
(D) sum of the number of protons and neutrons is same but the number of protons is different.

Q.9 How many neutrons are there in $^{88}_{38}\text{Sr}$?

(A) 38 (B) 50
(C) 126 (D) 88

Q.10 Chlorine exists in two isotopic forms, Cl-37 and Cl-35 but its atomic mass is 35.5. This indicates the ratio of Cl-37and Cl-35 is approximately

(A) 1 : 2 (B) 1 : 1
(C) 1 : 3 (D) 3 : 1

Q.11 Which of the following properties of atom could be explained correctly by Thomson model of atom?

(A) Overall neutrality of atom.
(B) Spectra of hydrogen atom.
(C) Position of electrons, protons and neutrons in atom.
(D) Stability of atom.

PART-3 : DEVELOPMENTS LEADING TO BOHR'S MODEL

Q.12 Calculate frequency of yellow radiation having wavelength 5800 Å.

(A) 5.712×10^{11} s^{-1} (B) 10.712×10^{14} s^{-1}
(C) 5.172×10^{14} s^{-1} (D) 5.712×10^{10} s^{-1}

Q.13 What will be the wavenumber of yellow radiation having wavelength 240 nm?

(A) 1.724×10^{4} m^{-1} (B) 4.16×10^{6} m^{-1}
(C) 4×10^{14} Hz (D) 219.3×10^{3} m^{-1}

Q.14 The Vividh Bharati station of All India Radio, Delhi, broadcasts on a frequency of 1,368 kHz (kilo hertz). Calculate the wavelength of the electromagnetic radiation emitted by transmitter.

(A) 219.3m (B) 119.3m
(C) 319.3m (D) 19.3m

Q.15 Electromagnetic radiation of wavelength 242 nm is just sufficient to ionize the sodium atom. What is the ionization energy of sodium per atom?

(A) 494.5×10^{-6} J/atom (B) 8169.5×10^{-10} J/atom
(C) 5.85×10^{-15} J/atom (D) 8.214×10^{-19} J/atom

Q.16 The energy difference between the ground state of an atom and its excited state is 3×10^{-19} J. What is the wavelength of the photon required for this transition?
(A) 6.6×10^{-34} m (B) 3×10^{-8} m
(C) 1.8×10^{-7} m (D) 6.6×10^{-7} m

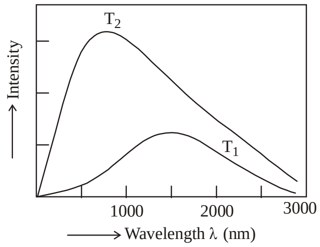
Q.17 The threshold frequency ν_0 for a metal is $7.0 \times 10^{14}\ s^{-1}$. Calculate the K.E. of an e^- emitted when radiation of frequency $\nu = 1.0 \times 10^{15} s^{-1}$ hits the metal.
(A) 1.988×10^{-19} J (B) 3.988×10^{-19} J
(C) 1.988×10^{-16} J (D) 1.988×10^{-21} J

Q.18 The spectrum of white light ranging from red to violet is called a continuous spectrum because –
(A) different colours are seen as different bands in the spectrum.
(B) the colours continuously absorb energy to form a spectrum.
(C) the violet colour merges into blue, blue into green, green into yellow and so on
(D) it is a continuous band of coloured and white light separating them.

Q.19 Calculate energy of one mole of photons of radiation whose frequency is 5×10^{14} Hz.
(A) 99.51 kJ mol^{-1} (B) 199.51 kJ mol^{-1}
(C) 299.51 kJ mol^{-1} (D) 399.51 kJ mol^{-1}

Q.20 What will be the energy of a photon which corresponds to the wavelength of 0.50 Å?
(A) 3.98×10^{-15} J (B) 3×10^{15} J
(C) 3.9×10^{8} J (D) 3×10^{-34} J

Q.21 Wavelength-Intensity relationship is shown at two different temperature. Choose the correct relation between T_1 and T_2.



(A) $T_2 > T_1$ (B) $T_2 < T_1$
(C) $T_2 = T_1$ (D) None of these

PART-4 : BOHR'S MODEL FOR HYDROGEN ATOM

Q.22 Given below are the spectral lines for an atom of hydrogen. Mark the lines which are not correctly matched with the value of n_1 and n_2 ?

	Series	n_1	n_2	Region
(i)	Lyman	1	2, 3, ….	Ultraviolet
(ii)	Balmer	2	3, 4, ….	Infrared
(iii)	Paschen	3	4, 5, ….	Infrared
(iv)	Pfund	4	5, 6, ….	Infrared

(A) (i) and (ii) (B) (i) and (iii)
(C) (ii) and (iv) (D) (i) and (iv)

Q.23 An electron in excited hydrogen atom falls from fifth energy level to second energy level. In which of the following regions, the spectrum line will be observed and is part of which series of the atomic spectrum?
(A) Visible, Balmer (B) Ultraviolet, Lyman
(C) Infrared, Paschen (D) Infrared, Brackett

Q.24 What is the wavelength of a photon emitted during a transition from n = 5 state to the n= 2 state in the hydrogen atom?
(A) 234 nm (B) 334 nm
(C) 434 nm (D) 534 nm

Q.25 The third line of the Balmer series in the emission spectrum of the hydrogen atom is due to the transition from the –
(A) fourth Bohr orbit to the first Bohr orbit.
(B) fifth Bohr orbit to the second Bohr orbit.
(C) sixth Bohr orbit to the third Bohr orbit.
(D) seventh Bohr orbit to the third Bohr orbit.

Q.26 Calculate the energy associated with the first orbit of He^+.
(A) -8.72×10^{-18} J (B) -3.72×10^{-18} J
(C) -8.72×10^{-15} J (D) -8.72×10^{-21} J

Q.27 The frequency of radiation absorbed or emitted when transition occurs between two stationary states with energies E_1 (lower) and E_2 (higher) is given by
(A) $\nu = \frac{E_1 + E_2}{h}$ (B) $\nu = \frac{E_1 - E_2}{h}$
(C) $\nu = \frac{E_1 \times E_2}{h}$ (D) $\nu = \frac{E_2 - E_1}{h}$

Q.28 The radius of hydrogen atom in ground state is 0.53Å. What will be the radius of $_3Li^{2+}$ in the ground state?
(A) 1.06 Å (B) 0.265 Å
(C) 0.17 Å (D) 0.53 Å

Q.29 Calculate the radius of the first orbit of He^+.
(A) 0.03645 nm (B) 0.02645 nm
(C) 0.2545 nm (D) 0.002645 nm

Q.30 _____ series of lines are the only lines in hydrogen spectrum which appear in the visible region.
(A) Lyman (B) Balmer
(C) Paschen (D) Brackett

Q.31 What is the trend of energy of Bohr's orbits?
(A) Energy of the orbit increases as we move away from the nucleus.
(B) Energy of the orbit decreases as we move away from the nucleus.
(C) Energy remains same as we move away from the nucleus.
(D) Energy of Bohr's orbit cannot be calculated.

Q.32 Bohr's theory can also be applied to the ions like
(A) He^+ (B) Li^{2+}
(C) Be^{3+} (D) All of these.

Q.33 Which one is not in agreement with Bohr's model of the atom?
(A) Line spectra of hydrogen atom
(B) Pauli's exclusion principle

(C) Planck's theory
(D) Heisenberg’s uncertainty principle

PART-5 : TOWARDS QUANTUM MECHANICAL MODEL OF THE ATOM

Q.34 A body of mass 10g is moving with a velocity of 100 ms^{-1}. The wavelength associated with it is
(A) 6.626×10^{-7} m
(B) 6.626×10^{-34} m
(C) 6.626×10^{-4} m
(D) 6.626×10^{-35} m

Q.35 The mass of an electron is 9.1×10^{-31} kg. If its K.E. is 3.0×10^{-25} J, calculate its wavelength.
(A) 896.7 nm
(B) 696.7 nm
(C) 996.7 nm
(D) 796.7 nm

Q.36 What will be the mass of a particle if uncertainty in its position is 10^{-8} m and velocity is 5.26×10^{-25} m s^{-1}?
(A) 0.01 kg
(B) 0.1 kg
(C) 1 kg
(D) 10 kg

Q.37 A microscope using suitable photons is employed to locate an electron in an atom within a distance of 0.1 Å. What is the uncertainty involved in the measurement of its velocity?
(A) 3.79×10^{6} m s^{-1}
(B) 4.79×10^{6} m s^{-1}
(C) 5.79×10^{6} m s^{-1}
(D) 8.79×10^{6} m s^{-1}

Q.38 Which of the following is responsible to rule out the existence of definite paths or trajectories of electrons?
(A) Pauli's exclusion principle.
(B) Heisenberg's uncertainty principle.
(C) Hund's rule of maximum multiplicity
(D) Aufbau principle

Q.39 Calculate the mass of a photon with wavelength 3.6 Å.
(A) 3.135×10^{-29} kg
(B) 6.135×10^{-19} kg
(C) 4.135×10^{-29} kg
(D) 6.135×10^{-29} kg

Q.40 What will be the wavelength of an electron moving with 1/10th of velocity of light?
(A) 2.43×10^{-11} m
(B) 243×10^{-11} m
(C) 0.243 m
(D) 2.43×10^{-4} m

Q.41 If travelling at same speeds, which of the following matter waves have the shortest wavelength?
(A) Electron
(B) α particle (He^{2+})
(C) Neutron
(D) Proton

Q.42 What will be the uncertainty in velocity of an electron when the uncertainty in its position is 1000 Å ?
(A) 5.79×10^{2} m s^{-1}
(B) 5.79×10^{8} m s^{-1}
(C) 5.79×10^{4} m s^{-1}
(D) 5.79×10^{-10} m s^{-1}

Q.43 A golf ball has a mass of 40g, and a speed of 45 m/s. If the speed can be measured within accuracy of 2%, calculate the uncertainty in the position.
(A) 3.46×10^{-33} m
(B) 1.46×10^{-33} m
(C) 1.46×10^{-31} m
(D) 1.46×10^{-22} m

PART-6 : QUANTUM MECHANICAL MODEL OF ATOM

Q.44 How many orbitals and electrons are associated with n = 4?
(A) 32, 64
(B) 16, 32
(C) 4, 16
(D) 8, 16

Q.45 Which of the following configurations does not follow Hund’s rule of maximum multiplicity?
(A) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$
(B) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$
(C) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$
(D) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4\ 4s^2$

Q.46 The region where probability density function reduces to zero is called –
(A) probability density region
(B) nodal surfaces
(C) orientation surfaces
(D) wave function.

Q.47 What is the total number of orbitals associated with the principal quantum number n = 3 ?
(A) 4
(B) 3
(C) 5
(D) 9

Q.48 The pair of ions having same electronic configuration
(A) Cr^{3+}, Fe^{3+}
(B) Fe^{3+}, Mn^{2+}
(C) Fe^{3+}, Co^{3+}
(D) Sc^{3+}, Cr^{3+}

Q.49 Two electrons present in M shell will differ in
(A) principal quantum number.
(B) azimuthal quantum number.
(C) magnetic quantum number.
(D) spin quantum number.

Q.50 Few electrons have following quantum numbers,
(i) n = 4, $\ell = 1$
(ii) n = 4, $\ell = 0$
(iii) n = 3, $\ell = 2$
(iv) n = 3, $\ell = 1$
Arrange them in the order of increasing energy from lowest to highest.
(A) (iv) < (ii) < (iii) < (i)
(B) (ii) < (iv) < (i) < (iii)
(C) (i) < (iii) < (ii) < (iv)
(D) (iii)< (i) < (iv) < (ii)

Q.51 Describe the orbital with following quantum numbers.:
(i) n = 3, $\ell = 2$
(ii) n = 4, $\ell = 3$
(A) (i) 3p, (ii) 4f
(B) (i) 3d, (ii) 4d
(C) (i) 3f, (ii) 4f
(D) (i) 3d, (ii) 4f

Q.52 Which of the following options does not represent ground state electronic configuration of an atom?
(A) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^8\ 4s^2$
(B) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^9\ 4s^2$
(C) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$
(D) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^1$

Q.53 Choose the correct matching for the following quantum numbers

(a) n = 2, $\ell = 1$	(i) 2p
(b) n = 4, $\ell = 0$,	(ii) 3d
(c) n = 5, $\ell = 3$,	(iii) 5f
(d) n = 3, $\ell = 2$	(iv) 4s

(A) a-i, b-ii, c-iii, d-iv
(B) a-i, b-iv, c-iii, d-ii
(C) a-ii, b-iii, c-i, d-iv
(D) a-iv, b-iii, c-ii, d-i

Q.54 Effective nuclear charge (Z_{eff}) for a nucleus of an atom is defined as –
(A) shielding of the outermost shell electrons from the nucleus by the innermost shell electrons.
(B) the net positive charge experienced by electron from the nucleus.

(C) the attractive force experienced by the nucleus from electron.
(D) screening of positive charge on nucleus by innermost shell electrons.

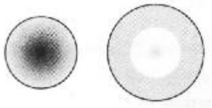
Q.55 For the electrons of oxygen atom, which of the following statements is correct?
(A) Z_{eff} for an electron in a 2s orbital is the same as Z_{eff} for an electron in a 2p orbital.
(B) An electron in the 2s orbital has the same energy as an electron in the 2p orbital.
(C) Z_{eff} for an electron in 1s orbital is the same as Z_{eff} for an electron in a 2s orbital.
(D) The two electrons present in the 2s orbital have spin quantum numbers, m_s but of opposite sign.

Q.56 The number of radial nodes for 3p orbital is
(A) 3 (B) 4
(C) 2 (D) 1

Q.57 Correct orbital picture for Neon atom is –
(A) [↑↓] [↑↓] [↑][↑][↑]
(B) [↑↓] [↑↓] [↑↓][↑][↑]
(C) [↑↓] [↑↓] [↑↓][↑↓][↑]
(D) [↑↓] [↑↓] [↑↓][↑↓][↑↓]

Q.58 Which of the sequences given below shows the correct increasing order of energy?
(A) 3s, 3p, 4s, 4p, 3d, 5s, 5p, 4d
(B) 3s, 3p, 3d, 4s, 4p, 4d, 5s, 5p
(C) 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p
(D) 3s, 3p, 4s, 4p, 5s, 3d, 4d, 5p

Q.59 The probability density plot of 1s and 2s atomic orbitals are given in figures



The density of dots in a region represents the probability density of finding electrons in the region.
Which of the statements is incorrect?
(A) 1s and 2s orbitals are spherical in shape.
(B) The probability of finding the electron is maximum near the nucleus.
(C) The probability of finding the electron at a given distance is equal in all directions.
(D) The probability density of electrons for 2s orbital decreases uniformly as distance from the nucleus increases.

Q.60 An electron can enter into the orbital when
(A) value of n is minimum
(B) value of ℓ is minimum
(C) value of $(n + \ell)$ is minimum
(D) value of $(n + m)$ is minimum.

Q.61 Total number of orbitals associated with third shell will be –
(A) 2 (B) 4
(C) 9 (D) 3

Q.62 In how many elements the last electron will have the following set of quantum numbers, $n = 3$, $\ell = 1$?
(A) 2 (B) 8
(C) 6 (D) 10

Q.63 Choose the correct statement
(A) In the ground state of the atoms, the orbitals are filled in order of their increasing energies.
(B) No two electrons in an atom can have the same set of four quantum numbers.
(C) the maximum number of electrons in the shell with principal quantum number n is equal to n^2.
(D) Both (A) and (B)

Q.64 Number of angular nodes for 4d orbital is –
(A) 4 (B) 3
(C) 2 (D) 1

Q.65 The configuration of the valence orbital of an element with atomic number 22 is
(A) $3d^5\ 4s^1$ (B) $3d^2\ 4s^2$
(C) $4s^1\ 4p^1$ (D) $3d^2\ 4s^1\ 4p^1$

Q.66 Orbital angular momentum depends on –
(A) ℓ (B) n and ℓ
(C) n and m (D) m and s

Q.67 Choose the correct statement
(A) the lower the value of $(n + \ell)$ for an orbital, the lower is its energy.
(B) Energies of the orbitals in the same subshell decrease with increase in the atomic number (Z_{eff}).
(C) The orbitals having the same energy are called degenerate.
(D) All of these

Q.68 Choose the correct statement –
(A) Size of the s orbital : $4s > 3s > 2s > 1s$
(B) Size of p orbitals : $4p > 3p > 2p$.
(C) Size of p orbitals : $4p < 3p < 2p$.
(D) Both (A) and (B)

Q.69 Which set of quantum numbers (n, ℓ, m_ℓ, m_s) is not possible?
(A) 1, 0, 0, 1/2 (B) 1, 1, 0, 1/2
(C) 1, 0, 0, –1/2 (D) 2, 1, –1, 1/2

Q.70 Correct orbital picture for oxygen atom is –
(A) [↑↓] [↑↓] [↑][↑][↑]
(B) [↑↓] [↑↓] [↑↓][↑][↑]
(C) [↑↓] [↑↓] [↑↓][↑↓][↑]
(D) [↑↓] [↑↓] [↑↓][↑↓][↑↓]

Q.71 Electron configuration of Cr
(A) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$
(B) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$
(C) $1s^2\ 2s^2\ 2p^8\ 3s^2\ 3p^6$
(D) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Q.72 Violates Pauli exclusion principle
(A) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$
(B) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$
(C) $1s^2\ 2s^2\ 2p^8\ 3s^2\ 3p^6$
(D) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

EXERCISE - 2 (LEVEL-2)

Choose one correct response for each question.

Q.1 Which of the following species is isoelectronic with CO?

(A) HF (B) N_2
(C) N_2^+ (D) O_2^-

Q.2 If the kinetic energy of the particle is increased to 16 times its previous value, the percentage change in the de-Broglie wavelength of the particle is X%. Find the value of (X – 70)

(A) 1 (B) 2
(C) 3 (D) 5

Q.3 When electromagnetic radiation of wavelength 300nm falls on the surface of sodium, electrons are emitted with a kinetic energy of 1.68×10^5 Jmol^{-1}. What is the minimum energy needed to remove an electron from sodium?

(A) 1.84×10^{-19} J (B) 3.84×10^{-15} J
(C) 3.84×10^{-21} J (D) 3.84×10^{-19} J

Q.4 Calculate the energy required for the process : $He^+(g) \rightarrow He^{2+}(g) + e^-$.
The ionisation energy for the H atom in the ground state is 2.18×10^{-18} J atom^{-1}.

(A) 8.72×10^{-18} J atom^{-1}
(B) 8.72×10^{-17} J atom^{-1}
(C) 2.18×10^{-18} J atom^{-1}
(D) 8.72×10^{-21} J atom^{-1}

Q.5 An electron is in one of the 3d-orbitals. What are the possible values of n, ℓ and m for this electron?

(A) $n = 3, \ell = 0, m_\ell = 0$
(B) $n = 3, \ell = 1, m_\ell = -1, 0, +1$
(C) $n = 3, \ell = 2, m_\ell = -2, -1, 0, +1, +2$
(D) $n = 3, \ell = 3, m_\ell = -3, -2, -1, 0, +1, +2, +3$

Q.6 Which of the following is true?

(A) The Bohr model applies to ^4He.
(B) The spectrum of hydrogen in the visible region of the electromagnetic spectrum is continuous.
(C) Because the electron has a more negative potential energy for n = 1 than it does for n = any integer> 1, the electron is most stable in its smallest allowed orbit.
(D) The electron in hydrogen may assume either a circular or elliptical orbit.

Q.7 An element with mass number 81 contains 31.7% more neutrons as compared to protons. Find the symbol of the atom.

(A) $^{81}_{34}Se$ (B) $^{81}_{35}Br$
(C) $^{81}_{36}Kr$ (D) $^{81}_{37}Rb$

Q.8 The 3d-orbitals having electron density in all the three axis is

(A) $3d_{xy}$ (B) $3d_{z^2}$
(C) $3d_{yz}$ (D) $3d_{zx}$

Q.9 The probability of finding out an electron at a point within an atom is proportional to the –

(A) square of the orbital wave function i.e., ψ^2
(B) orbital wave function i.e., ψ
(C) Hamiltonian operator i.e., H
(D) principal quantum number i.e., n

Q.10 With respect to the electromagnetic spectrum, which of the following is not correct ?

(A) Electromagnetic radiation of wavelength 700nm falls in the red region of the visible spectrum.
(B) A photon of wavelength 700 nm will have higher energy than a photon of infrared radiation.
(C) Electromagnetic radiation of wavelength 700nm has an identical speed to that of radiation of wavelength 400 nm.
(D) Electromagnetic radiation of wavelength 700nm has a higher frequency than that of X-rays.

Q.11 If the velocity of an electron in Bohr's first orbit is 2.19×10^6 ms^{-1}, what will be the de Broglie wavelength associated with it?

(A) 2.19×10^{-6} m (B) 4.38×10^{-6} m
(C) 3.32×10^{-10} m (D) 3.32×10^{-7} m

Q.12 The emission spectrum of hydrogen is found to satisfy the expression for the energy change ΔE (in joules) such that $\Delta E = 2.18 \times 10^{-18}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$ J where $n_1 = 1, 2, 3, \ldots$ and $n_2 = 2, 3, 4$. The spectral lines corresponds to Paschen series if –

(A) $n_1 = 1$ and $n_2 = 2, 3, 4$ (B) $n_1 = 3$ and $n_2 = 4, 5, 6$
(C) $n_1 = 1$ and $n_2 = 3, 4, 5$ (D) $n_1 = 2$ and $n_2 = 3, 4, 5$

Q.13 The order of increasing energies of the orbitals follows

(A) 3s, 3p, 4s, 3d, 4p (B) 3s, 3p, 3d, 4s, 4p
(C) 3s, 3p, 4s, 4p, 3d (D) 3s, 3p, 3d, 4p, 4s

Q.14 The de Broglie wavelength associated with a ball of mass 200g and moving at a speed of 5 m/hr, is of the order of ($h = 6.626 \times 10^{-34}$ J s).

(A) 10^{-15} m (B) 10^{-20} m
(C) 10^{-25} m (D) 10^{-30} m

Q.15 The number of radial nodes and angular nodes for d-orbital can be represented as

(A) (n – 2) radial nodes + 1 angular node = (n – 1) total nodes
(B) (n – 1) radial nodes + 1 angular node = (n – 1) total nodes
(C) (n – 3) radial nodes + 2 angular nodes = $(n - \ell - 1)$ total nodes
(D) (n – 3) radial nodes + 2 angular nodes = (n – 1) total nodes

Q.16 The radius of the stationary state which is also called Bohr radius is given by the expression $r_n = n^2 a_0$ where the value of a_0 is –

(A) 52.9 pm (B) 5.29 pm
(C) 529 pm (D) 0.529 pm

Q.17 If the energy of a photon corresponding to a wavelength of 6000 Å is 3.32×10^{-19} J, the photon energy for a wavelength of 4000 Å is
(A) 1.4 eV (B) 4.9 eV
(C) 3.1 eV (D) 1.6 eV

Q.18 In astronomical observations, signals observed from the distant stars are generally weak. If the photon detector receives a total of 3.15×10^{-18}J from the radiation of 600 nm, calculate the number of photons received by the detector.
(A) 10 (B) 20
(C) 25 (D) 35

Q.19 Few statements are given regarding nodes in the orbitals. Mark the statement which is not correct.
(A) In case of p_z-orbital, xy plane is a nodal plane.
(B) ns - orbital has (n + 1) nodes.
(C) The number of angular nodes is given by ℓ.
(D) The total number of nodes is given by (n – 1) i.e. sum of ℓ angular nodes and $(n - \ell - 1)$ radial nodes.

Q.20 Barium salts, when burned, emit a greenish light (often utilized in fireworks). One of the most common, $BaCl_2$, produces light with the wavelength of 511 nm. What would the corresponding frequency of this wavelength be?
(A) $5.82 \times 10^{10}\ s^{-1}$ (B) $5.87 \times 10^{14}\ s^{-1}$
(C) $8.48 \times 10^{16}\ s^{-1}$ (D) $2.32 \times 10^{18}\ s^{-1}$

Q.21 Which of the following quantum numbers are correct for the outermost electron of sodium atom?
(A) $n = 4, \ell = 0, m = 0, s = +1/2$
(B) $n = 3, \ell = 0, m = 0, s = -1/2$
(C) $n = 3, \ell = 1, m = +1, s = +1/2$
(D) $n = 3, \ell = 2, m = -1, s = -1/2$

Q.22 What does the negative electronic energy (negative sign for all values of energy) for hydrogen atom means?
(A) The energy of an electron in the atom is lower than the energy of a free electron at rest which is taken as zero.
(B) When the electron is free from the influence of nucleus it has a negative value which becomes more negative.
(C) When the electron is attracted by the nucleus the energy is absorbed which means a negative value.
(D) Energy is released by hydrogen atom in ground state.

Q.23 Which atom (X) is indicated by the following configuration? $X \rightarrow [Ne]\ 3s^2\ 3p^3$
(A) Nitrogen (B) Chlorine
(C) Phosphorus (D) Sulphur

Q.24 Nitrogen laser produces a radiations at a wavelength of 337.1 nm. If the number of photons emitted per second is 5.6×10^{24}, calculate the power of this laser.
(A) 1.3×10^3 kW (B) 4.3×10^3 kW
(C) 3.3×10^2 kW (D) 3.3×10^3 kW

Q.25 The wave-mechanical model of atom is based upon
(A) De Broglie concept of dual character of matter .
(B) Heisenberg's uncertainty principle
(C) Schrodinger wave equation
(D) All the above three

Q.26 Though the five d-orbitals are degenerate, the first four d-orbitals are similar to each other in shape whereas the fifth d-orbital is different from others. What is the name of the fifth orbital?
(A) $d_{x^2-y^2}$ (B) d_{z^2}
(C) d_{xz} (D) d_{xy}

Q.27 Two values of spin quantum numbers i.e., +1/2 and –1/2 represent
(A) up and down spin of the electrons respectively.
(B) two quantum mechanic spin states which refer to the orientation of spin of the electron.
(C) clockwise and anti-clockwise spin of the electrons respectively.
(D) anti-clockwise and clockwise spin of the electrons respectively.

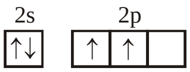
Q.28 Which of the following regarding the Rutherford experiment is not correct?
(A) Most alpha particles passed through the gold foil without being deflected since the nuclei of the gold atoms represent such a small portion of the total atomic volume.
(B) Coulomb's law, which states that like charges repel each other, accounts for the deflection of alpha particles passing close to gold atom nuclei.
(C) Most alpha particles passed through the gold foil without being deflected since the nuclei of the gold atoms represent such a small portion of the total atomic mass.
(D) Because of the electrons' small masses, they did not deflect the alpha particles.

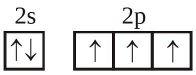
Q.29 Choose the correct statement –
(A) the wavenumber of radiation having wavelength 5000 Å is $2 \times 10^6\ m^{-1}$.
(B) The velocity of electron present in first Bohr orbit of hydrogen atom is 2.18×10^6 m/s.
(C) The wavenumber of radiation having wavelength 5000 Å is $2 \times 10^6\ cm^{-1}$.
(D) Both (A) and (B)

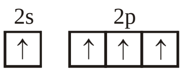
Q.30 What will be the orbital angular momentum of an electron in 2s-orbital?
(A) Zero (B) One
(C) Two (D) Three

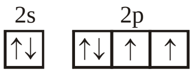
Q.31 The work function for cesium atom is 1.9 eV. Calculate the threshold wavelength
(A) 4.53×10^{-7} m (B) 3.53×10^{-7} m
(C) 6.53×10^{-7} m (D) 6.53×10^{-5} m

Q.32 The orbital diagram in which the Aufbau principle is violated is

(A) 

(B) 

(C) 

(D) 

Q.33 Read the following statements and mark the incorrect statement.

(A) No two electrons in an atom can have all the four quantum numbers same.

(B) All the orbitals in a subshell are first occupied singly with parallel spins.

(C) The outer electronic configuration of chromium atom is $3d^4\ 4s^2$.

(D) Lyman series of hydrogen spectrum lies in ultraviolet region.

Q.34 Arrange the following particles in increasing order of values of e/m ratio : Electron (e), proton (p), neutron(n) and α-particle (α) -

(A) n, p, e, α (B) n, α, p, e

(C) n, p, α, e (D) e, p, n, α

EXERCISE - 3 (NUMERICAL VALUE BASED QUESTIONS)

NOTE : The answer to each question is a NUMERICAL VALUE.

Q.1 The ionization energy of the Be^{3+} ion with the electron in its ground state is 2.10×10^{A} kJ/mol. Find the value of A.

Q.2 The maximum number of electrons that can have principal quantum number, n = 3, and spin quantum number, $m_s = -1/2$, is

Q.3 The work function (ϕ) of some metals is listed below. The number of metals which will show photoelectric effect when light of 300 nm wavelength falls on the metal is

Metal	Li	Na	K	Mg	Cu
ϕ (eV)	2.4	2.3	2.2	3.7	4.8
Metal	Ag	Fe	Pt	W	
ϕ (eV)	4.3	4.7	6.3	4.75	

Q.4 The atomic masses of He and Ne are 4 and 20 a.m.u., respectively. The value of the de Broglie wavelength of He gas at –73°C is "M" times that of the de Broglie wavelength of Ne at 727°C. M is –

Q.5 In an atom, the total number of electrons having quantum numbers n = 4, $|m_\ell| = 1$ and $m_s = -1/2$ is

Q.6 Determine the number of unpaired electrons in an atom of tellurium, Te.

EXERCISE - 4 [PREVIOUS YEARS AIEEE / JEE MAIN QUESTIONS]

Q.1 An atom has a mass of 0.02 kg & uncertainity in its velocity is 9.218×10^{-6} m/s then uncertainity in position is
($h = 6.626 \times 10^{-34}$ J - s) **[AIEEE- 2002]**
(A) 2.86×10^{-28} m (B) 2.86×10^{-32} cm
(C) 1.5×10^{-27} m (D) 3.9×10^{-10} m

Q.2 Energy of H–atom in the ground state is –13.6 eV , Hence energy in the second excited state is – **[AIEEE- 2002]**
(A) –6.8 eV (B) –3.4 eV
(C) –1.51 eV (D) –4.3 eV

Q.3 Unertainty in position of a particle of 25 g in space is 10^{-5} m. Hence uncertainty in velocity (ms^{-1}) is
(Planck's constant $h = 6.6 \times 10^{-34}$ Js) **[AIEEE- 2002]**
(A) 2.1×10^{-28} (B) 2.1×10^{-34}
(C) 0.5×10^{-34} (D) 5.0×20^{-24}

Q.4 The orbital angular momentum for an electron revolving in an orbit is given by $\sqrt{\ell(\ell+1)}.\frac{h}{2\pi}$. This momentus for an s-electron will be given by - **[AIEEE- 2003]**
(A) $\frac{h}{2\pi}$ (B) $\sqrt{2}.\frac{h}{2\pi}$
(C) $+\frac{1}{2}.\frac{h}{2\pi}$ (D) zero

Q.5 In Bohr series of lines of hydrogen spectrum, third line from the red end corresponds to where one of the following inter-orbit jumps of electron for Bohr orbits in an atom of hydrogen. **[AIEEE- 2003]**
(A) $4 \rightarrow 1$ (B) $2 \rightarrow 5$
(C) $3 \rightarrow 2$ (D) $5 \rightarrow 2$

Q.6 The de Broglie wavelength of a tennis ball mass 60 g moving with a velocity of 10 mt. per second is approximately - **[AIEEE- 2003]**
(A) 10^{-16} metres (B) 10^{-25} metres
(C) 10^{-33} metres (D) 10^{-31}metres

Q.7 Which of the following sets of quantum numbers is correct for an electron in 4f orbital ? **[AIEEE- 2004]**
(A) $n = 4, \ell = 3, m = +4, s = +1/2$
(B) $n = 4, \ell = 4, m = -4, s = -1/2$
(C) $n = 4, \ell = 3, m = +1, s = +1/2$
(D) $n = 4, \ell = 3, m = -2, s = +1/2$

Q.8 Consider the ground state of Cr atom (Z = 24). The numbers of electrons with the azimuthal quantum numbers, $\ell = 1$ and 2 are, respectively **[AIEEE- 2004]**
(A) 12 and 4 (B) 12 and 5
(C) 16 and 4 (D) 16 and 5

Q.9 The wavelength of the radiation emitted, when in a hydrogen atom electron falls from infinity to stationary state 1, would be (Rydberg constant = 1.097×10^{7} m^{-1})
(A) 91 nm (B) 192 nm **[AIEEE- 2004]**
(C) 406 nm (D) 9.1×10^{-8} nm

Q.10 Which one of the following sets of ions represents the collection of isoelectronic species ? **[AIEEE- 2004]**
(A) $K^+, Ca^{2+}, Sc^{3+}, Cl^-$ (B) $Na^+, Ca^{2+}, Sc^{3+}, F^-$
(C) $K^+, Cl^-, Mg^{2+}, Sc^{3+}$ (D) $Na^+, Mg^{2+}, Al^{3+}, Cl^-$
(Atomic nos.: F = 9, Cl = 17, Na = 11, Mg = 12, Al = 13, K = 19, Ca = 20 , Sc = 21)

Q.11 In a multi-electron atom, which of the following orbitals described by the three quantum members will have the same energy in the absence of magnetic and electric fields **[AIEEE- 2005]**
(a) $n = 1, \ell = 0, m = 0$ (b) $n = 2, \ell = 0, m = 0$
(c) $n = 2, \ell = 1, m = 1$ (d) $n = 3, \ell = 2, m = 1$
(e) $n = 3, \ell = 2, m = 0$
(A) (b) and (c) (B) (a) and (b)
(C) (d) and (e) (D) (c) and (d)

Q.12 Of the following sets which one does NOT contain isoelectronic species ? **[AIEEE- 2005]**
(A) CN^-, N_2, C_2^{2-} (B) $PO_4^{3-}, SO_4^{2-}, ClO_4^-$
(C) $BO_3^{3-}, CO_3^{2-}, NO_3^-$ (D) $SO_3^{2-}, CO_3^{2-}, NO_3^-$

Q.13 According to Bohr's theory, the angular momentum of an electron in 5th orbit is - **[AIEEE 2006]**
(A) 1.0 h/π (B) 10 h/π
(C) 2.5 h/π (D) 25 h/π

Q.14 Uncertainty in the position of an electron (mass = 9.1×10^{-31} kg) moving with a velocity 300 m/s, accurate upto 0.001 %, will be **[AIEEE 2006]**
(A) 5.76×10^{-2} m (B) 1.92×10^{-2} m
(C) 3.84×10^{-2} m (D) 19.2×10^{-2} m

Q.15 Which of the following sets of quantum numbers represents the highest energy of an atom **[AIEEE 2007]**
(A) $n = 3, \ell = 1, m = 1, s = +1/2$
(B) $n=3, \ell = 2, m = 1, s = +1/2$
(C) $n = 4, \ell = 0, m = 0, s = +1/2$
(D) $n = 3, \ell = 0, m = 0, s = +1/2$

Q.16 The ionization enthalpy of hydrogen atom is 1.312×10^{6} J mol^{-1}. The energy required to excite the electron in the atom from n = 1 to n = 2 is **[AIEEE 2008]**
(A) 6.56×10^{5} J mol^{-1} (B) 7.56×10^{5} J mol^{-1}
(C) 9.84×10^{5} J mol^{-1} (D) 8.51×10^{5} J mol^{-1}

Q.17 Calculate the wavelength (in nanometer) associated with a proton moving at 1.0×10^{3} ms^{-1} (Mass of proton = 1.67×10^{-27} kg and $h = 6.63 \times 10^{-34}$ Js):
(A) 14.0 nm (B) 0.032 nm **[AIEEE 2009]**
(C) 0.40 nm (D) 2.5 nm

Q.18 In an atom, an electron is moving with a speed of 600 m/s with an accuracy of 0.005%. Certainity with which the position of the electron can be located is
($h = 6.6 \times 10^{-34}$ kg m^2 s^{-1},
mass of electron, $e_m = 9.1 \times 10^{-31}$ kg): **[AIEEE 2009]**
(A) 3.84×10^{-3} m (B) 1.52 × 10–4 m
(C) 5.10×10^{-3} m (D) 1.92×10^{-3} m

Q.19 Ionisation energy of He^+ is 19.6×10^{-18} J $atom^{-1}$. The energy of the first stationary state (n = 1) of Li^{2+} is – **[AIEEE 2010]**
(A) 4.41×10^{-16} J $atom^{-1}$ (B) -4.41×10^{-17}J$atom^{-1}$
(C) -2.2×10^{-15} J $atom^{-1}$ (D) 8.82×10^{-17} J $atom^{-1}$

Q.20 A gas absorbs a photon of 355 nm and emits at two wavelengths. If one of the emission is at 680 nm, the other is at **[AIEEE 2011]**

(A) 1035 nm (B) 325 nm
(C) 743 nm (D) 518 nm

Q.21 The electrons identified by quantum numbers n and ℓ : **[AIEEE 2012]**

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

can be placed in order of increasing energy as :

(A) (c) < (d) < (b) < (a) (B) (d) < (b) < (c) < (a)
(C) (b) < (d) < (a) < (c) (D) (a) < (c) < (b) < (d)

Q.22 The increasing order of the ionic radii of the given isoelectronic species is : **[AIEEE 2012]**

(A) $Cl^-, Ca^{2+}, K^+, S^{2-}$ (B) $S^{2-}, Cl^-, Ca^{2+}, K^+$
(C) $Ca^{2+}, K^+, Cl^-, S^{2-}$ (D) $K^+, S^{2-}, Ca^{2+}, Cl^-$

Q.23 Energy of an electron is given by

$E = -2.178 \times 10^{-18}\text{J}\left(\frac{Z^2}{n^2}\right)$. Wavelength of light required to excite an electron in an hydrogen atom from level n = 1 to n = 2 will be –

($h = 6.62 \times 10^{-34}$ Js and $c = 3.0 \times 10^8\ ms^{-1}$) **[JEE MAIN 2013]**

(A) 1.214×10^{-7} m (B) 2.816×10^{-7} m
(C) 6.500×10^{-7} m (D) 8.500×10^{-7} m

Q.24 The first ionisation potential of Na is 5.1 eV. The value of electron gain enthalpy of Na^+ will be : **[JEE MAIN 2013]**

(A) –2.55 eV (B) –5.1 eV
(C) –10.2 eV (D) +2.55 eV

Q.25 The correct set of four quantum numbers for the valence electrons of rubidium atom (Z = 37) is: **[JEE MAIN 2014]**

(A) 5, 1, 1, +1/2 (B) 5, 0, 1, +1/2
(C) 5, 0, 0, +1/2 (D) 5, 1, 0, +1/2

Q.26 Which of the following is the energy of a possible excited state of hydrogen? **[JEE MAIN 2015]**

(A) – 6.8 eV (B) – 3.4 eV
(C) +6.8 eV (D) +13.6 eV

Q.27 A stream of electrons from a heated filament was passed between two charged plates kept at a potential difference V esu. If e and m are charge and mass of an electron, respectively, then the value of h/λ (where λ is wavelength associated with electron wave) is given by **[JEE MAIN 2016]**

(A) 2 meV (B) $\sqrt{meV}$
(C) $\sqrt{2meV}$ (D) meV

Q.28 The radius of the second Bohr orbit for hydrogen atom is: (Planck's constant, $h = 6.6262 \times 10^{-34}$ Js; mass of electron $= 9.1091 \times 10^{-31}$ kg; charge of electron $e = 1.60210 \times 10^{-19}$ C; permittivity of vacuum $\epsilon_0 = 8.854185 \times 10^{-12}\ kg^{-1}\ m^{-3}\ A^2$) **[JEE MAIN 2017]**

(A) 2.12Å (B) 1.65Å
(C) 4.76Å (D) 0.529Å

Q.29 For emission line of atomic hydrogen from $n_i = 8$ to $n_f = n$, the plot of wave number $(\bar{\nu})$ against $(1/n^2)$ will be (The Rydberg constant, R_H is in wave number unit). **[JEE MAIN 2019 (JAN)]**

(A) Linear with slope - R_H (B) Linear with intercept - R_H
(C) Non linear (D) Linear with slope R_H

Q.30 The quantum number of four electrons are given below

I. $n = 4, \ell = 2, m_\ell = -2, m_s = -½$
II. $n = 3, \ell = 2, m_\ell = 1, m_s = +½$
III. $n = 4, \ell = 1, m_\ell = 0, m_s = +½$
IV. $n = 3, \ell = 1, m_\ell = 1, m_s = -½$

The correct order of their increasing energies will be - **[JEE MAIN 2019 (APRIL)]**

(A) IV < III < II < I (B) IV < II < III < I
(C) I < II < III < IV (D) I < III < II < IV

Q.31 The size of the iso-electronic species Cl^-, Ar and Ca^{2+} is affected by - **[JEE MAIN 2019 (APRIL)]**

(A) Principal quantum number of valence shell.
(B) Nuclear charge.
(C) Azimuthal qunatum number of valence shell.
(D) Electron-electron interaction in the outer orbitals

Q.32 If p is the momentum of the fastest electron ejected from a metal surface after the irradiation of light having wavelength λ, then for 1.5 p momentum of the photoelectron, the wavelength of the light should be: (Assume kinetic energy of ejected photoelectron to be very high in comparison to work function) **[JEE MAIN 2019 (APRIL)]**

(A) (1/2) λ (B) (3/4) λ
(C) (2/3) λ (D) (4/9) λ

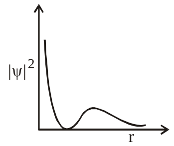
Q.33 The element having greatest difference between its first and second ionization energies, is : **[JEE MAIN 2019 (APRIL)]**

(A) Ca (B) K
(C) Ba (D) Sc

Q.34 Which one of the following about an electron occupying the 1s orbital in a hydrogen atom is incorrect ? (The Bohr radius is represented by a_0) **[JEE MAIN 2019 (APRIL)]**

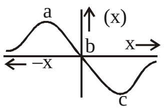
(A) The electron can be found at a distance $2a_0$ from the nucleus.
(B) The probability density of finding the electron is maximum at the nucleus.
(C) The magnitude of potential energy is double that of its kinetic energy on an average.
(D) The total energy of the electron is maximum when it is at a distance a_0 from the nucleus.

Q.35 The graph betweeen $|\psi|^2$ and r (radial distance) is shown below. This represents : **[JEE MAIN 2019 (APRIL)]**



(A) 3s orbital (B) 1s orbital
(C) 2p orbital (D) 2s orbital

Q.36 The electrons are more likely to be found :



[JEE MAIN 2019 (APRIL)]

(A) in the region a and b
(B) in the region a and c
(C) only in the region c
(D) only in the region a

Q.37 n = 5, $m_s = +1/2$. How many orbitals are possible: **[JEE MAIN 2020 (JAN)]**

(A) 25
(B) 30
(C) 50
(D) 35

Q.38 Given for H-atom $\overline{\nu} = R_H\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$

Select the correct options regarding this formula for Balmer series. **[JEE MAIN 2020 (JAN)]**

(a) $n_1 = 2$
(b) Ionization energy of H atom can be calculated from above formula.
(c) $\lambda_{maximum}$ is for $n_2 = 3$.
(d) If λ decreases then spectrum lines will converse.

(A) a, b
(B) c, d
(C) a & c
(D) a, b, c & d

Q.39 The radius of the second Bohr orbit, in terms of the Bohr radius, a_0, in Li^{2+} is : **[JEE MAIN 2020 (JAN)]**

(A) $\frac{3a_0}{4}$
(B) $\frac{4a_0}{3}$
(C) $\frac{a_0}{3}$
(D) $\frac{16a_0}{9}$

Q.40 The de Broglie wavelength of an electron in the 4th Bohr orbit is : **[JEE MAIN 2020 (JAN)]**

(A) $8\pi a_0$
(B) $2\pi a_0$
(C) $4\pi a_0$
(D) $6\pi a_0$

Q.41 The electronic configurations of bivalent europium and trivalent cerium are **[JEE MAIN 2020 (JAN)]**

(Atomic number : Xe = 54, Ce = 58, Eu = 63)

(A) $[Xe]\ 4f^4$ and $[Xe]\ 4f^9$
(B) $[Xe]\ 4f^7$ and $[Xe]\ 4f^1$
(C) $[Xe]\ 4f^7\ 6s^2$ and $[Xe]\ 4f^2\ 6s^2$
(D) $[Xe]\ 4f^2$ and $[Xe]\ 4f^7$

EXERCISE - 5 (PREVIOUS YEARS AIPMT/NEET EXAM QUESTIONS)

Choose one correct response for each question.

Q.1 The energy of second Bohr orbit of the hydrogen atom is –328 kJ mol^{-1}, hence the energy of fourth Bohr orbit would be – **[AIPMT 2005]**
(A) – 41 kJ mol^{-1} (B) – 82 kJ mol^{-1}
(C) – 164 kJ mol^{-1} (D) – 1312 kJ mol^{-1}

Q.2 Given : The mass of electron is 9.11×10^{-31} kg , Planck constant is 6.626×10^{-34} Js, the uncertainity involved in the measurement of velocity within a distance of 0.1Å is – **[AIPMT 2006]**
(A) 5.79×10^{7} m/s (B) 5.79×10^{8} m/s
(C) 5.79×10^{5} m/s (D) 5.79×10^{6} m/s

Q.3 The orientation of an atomic orbital is governed by
(A) Spin quantum number **[AIPMT 2006]**
(B) Magnetic quantum number
(C) Principal quantum number
(D) Azimuthal quantum number

Q.4 Consider the following sets of quantum numbers : **[AIPMT 2007]**

	n	ℓ	m	s
(i)	3	0	0	+1/2
(ii)	2	2	1	+1/2
(iii)	4	3	–2	–1/2
(iv)	1	0	–1	–1/2
(v)	3	2	3	+1/2

Which of the following sets of quantum number is not possible –
(A) (i), (ii), (iii) and (iv) (B) (ii), (iv) and (v)
(C) (i) and (iii) (D) (ii), (iii) and (iv)

Q.5 If uncertainty in position and momentum are equal, then uncertainty in velocity is : **[AIPMT 2008]**
(A) $\sqrt{\frac{h}{\pi}}$ (B) $\frac{1}{2m}\sqrt{\frac{h}{\pi}}$
(C) $\sqrt{\frac{h}{2\pi}}$ (D) $\frac{1}{m}\sqrt{\frac{h}{\pi}}$

Q.6 The measurement of the electron position is associated with an uncertainty in momentum, which is equal to 1×10^{-18} g cms^{-1}. The uncertainty in electron velocity is, (mass of an electron is 9×10^{-28}g) **[AIPMT 2008]**
(A) 1×10^{11} cms^{-1} (B) 1×10^{9} cms^{-1}
(C) 1×10^{6} cms^{-1} (D) 1×10^{5} cms^{-1}

Q.7 The energy absorbed by each molecule (A_2) of a substance is 4.4×10^{-19} J and bond energy per molecule is 4.0×10^{-19} J. The kinetic energy of the molecule per atom will be: **[AIPMT 2009]**
(A) 2.2×10^{-19} J (B) 2.0×10^{-19} J
(C) 4.0×10^{-20} J (D) 2.0×10^{-20} J

Q.8 Maximum number of electrons in a subshell of an atom is determined by the following: **[AIPMT 2009]**
(A) $2\ell + 1$ (B) $4\ell - 2$
(C) $2n^2$ (D) $4\ell + 2$

Q.9 Which of the following is not permissible arrangement of electrons in an atom? **[AIPMT 2009]**
(A) n = 5, ℓ = 3, m = 0, s = + 1/2
(B) n = 3, ℓ = 2, m = –3, s = – 1/2
(C) n = 3, ℓ = 2, m = –2, s = –1/2
(D) n = 4, ℓ = 0, m = 0, s = – 1/2

Q.10 A 0.66 kg ball is moving with a speed of 100 m/s. The associated wavelength will be : (h = 6.6×10^{-34} Js) **[AIPMT 2010]**
(A) 6.6×10^{-32} m (B) 6.6×10^{-34} m
(C) 1.0×10^{-35} m (D) 1.0×10^{-32} m

Q.11 The total number of atomic orbitals in fourth energy level of an atom is – **[AIPMT 2010]**
(A) 4 (B) 8
(C) 16 (D) 32

Q.12 The energies E_1 and E_2 of two radiations are 25eV and 50eV respectively. The relation between their wavelengths i.e. λ_1 and λ_2 will be **[AIPMT 2011]**
(A) $\lambda_1 = \frac{1}{2}\lambda_2$ (B) $\lambda_1 = \lambda_2$
(C) $\lambda_1 = 2\lambda_2$ (D) $\lambda_1 = 4\lambda_2$

Q.13 If n = 6, the correct sequence of filling of electrons will be **[AIPMT 2011]**
(A) ns → np (n – 1) d → (n – 2) f
(B) ns → n (n – 2) f → (n – 1) d → np
(C) ns → n (n – 1) d → (n – 2) f → np
(D) ns → (n – 2) f → np → (n – 1) f

Q.14 According to the Bohr Theory, which of the following transitions in the hydrogen atom will give rise to the least energetic photon **[AIPMT 2011]**
(A) n = 6 to n = 1 (B) n = 5 to n = 4
(C) n = 6 to n = 5 (D) n = 5 to n = 3

Q.15 Maximum number of electrons in a subshell with ℓ = 3 and n = 4 is : **[AIPMT 2012]**
(A) 14 (B) 16
(C) 10 (D) 12

Q.16 The correct set of four quantum numbers for the valence elecron of rubidium atom (Z = 37) is :**[AIPMT 2012]**
(A) 5, 1, + 1/2 (B) 6, 0, 0 + 1/2
(C) 5, 0, 0 + 1/2 (D) 5, 1, 0 + 1/2

Q.17 The orbital angular momentum of a p-electron is given as **[AIPMT 2012]**
(A) $\frac{h}{\sqrt{2}\pi}$ (B) $\sqrt{3}\frac{h}{2\pi}$ (C) $\sqrt{\frac{3}{2}}\frac{h}{\pi}$ (D) $\sqrt{6}\frac{h}{2\pi}$

Q.18 The value of Planck's constant is 6.63×10^{-34} Js. The speed of light is 3×10^{17} nm s^{-1}. Which value is closest to the wavelength in nanometer of a quantum of light with frequency of 6×10^{15} s^{-1} ? **[NEET 2013]**
(A) 75 (B) 10
(C) 25 (D) 50

Q.19 Based on equation E = -2.178×10^{-18} J (Z^2/n^2) certain conclusions are written. Which of them is not correct ? **[NEET 2013]**
(A) For n = 1, the electron has a more negative energy than it does for n = 6 which means that the electron is more loosely bound in the smallest allowed orbit.

(B) The negative sign in equation simply means that the energy of electron bound to the nucleus is lower than it would be if the electrons were at the infinite distance from the nucleus.
(C) Larger the value of n, the larger is the orbit radius.
(D) Equation can be used to calculate the change in energy when the electron change orbit

Q.20 What is the maximum numbers of electrons that can be associated with the following set of quantum numbers ?
$n = 3$, $\ell = 1$ and $m = -1$ **[NEET 2013]**
(A) 2
(B) 10
(C) 6
(D) 4

Q.21 What is the maximum number of orbitals that can be identified with the following quantum numbers?
$n = 3$, $\ell = 1$, $m_\ell = 0$ **[AIPMT 2014]**
(A) 1
(B) 2
(C) 3
(D) 4

Q.22 Calculate the energy in joule corresponding to light of wavelength 45 nm : (Planck's constant $h = 6.63 \times 10^{-34}$ Js; speed of light $c = 3 \times 10^{8}$ ms^{-1}) **[AIPMT 2014]**
(A) 6.67×10^{15}
(B) 6.67×10^{11}
(C) 4.42×10^{-15}
(D) 4.42×10^{-18}

Q.23 Be^{2+} is isoelectronic with which of the following ions? **[AIPMT 2015]**
(A) H^+
(B) Li^+
(C) Na^+
(D) Mg^{2+}

Q.24 Angular momentum of electron in 'd' orbital is equal to **[AIPMT 2015]**
(A) $\sqrt{2}\hbar$
(B) $2\sqrt{3}\,\hbar$
(C) $0\,\hbar$
(D) $\sqrt{6}\,\hbar$

Q.25 The number of d-electrons in Fe^{2+} (Z=26) is not equal to the number of electrons in which one of the following? **[AIPMT 2015]**
(A) p-electrons in Cl (Z = 17)
(B) d-electrons in Fe (Z = 26)
(C) p-electrons in Ne (Z = 10)
(D) s-electrons in Mg (Z = 12)

Q.26 Gadolinium belongs to 4f series. It's atomic number is 64. Which of the following is the correct electronic configuration of gadolinium ? **[RE-AIPMT 2015]**
(A) $[Xe]\,4f^7 5d^1 6s^2$
(B) $[Xe]\,4f^6 5d^2 6s^2$
(C) $[Xe]\,4f^8 6d^2$
(D) $[Xe]\,4f^9 5s^1$

Q.27 Which is the correct order of increasing energy of the listed orbitals in the atom of titanium ?
(At. no. Z = 22) **[RE-AIPMT 2015]**
(A) 3s 3p 3d 4s
(B) 3s 3p 4s 3d
(C) 3s 4s 3p 3d
(D) 4s 3s 3p 3d

Q.28 The electronic configurations of Eu (Atomic no. 63), Gd (Atomic No. 64) and Tb (Atomic No 65) are **[NEET 2016 PHASE 1]**
(A) $[Xe]\,4f^7 6s^2$, $[Xe]\,4f^8 6s^2$ and $[Xe]\,4f^8 5d^1 6s^2$
(B) $[Xe]\,4f^6 5d^1 6s^2$, $[Xe]\,4f^7 5f^1$ and $[Xe]\,4f^9 6s^2$
(C) $[Xe]\,4f^6 5d^1 6s^2$, $[Xe]\,4f^7 5d^1 6s^2$ and $[Xe]\,4f^8 5d^1 6s^2$
(D) $[Xe] 4f^7 6s^2$, $[Xe]\,4f^7 5d^1 6s^2$ and $[Xe]\,4f^9 6s^2$

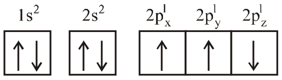
Q.29 Two electrons occupying the same orbital are distinguished by **[NEET 2016 PHASE 1]**
(A) Principal quantum number
(B) Magnetic quantum number
(C) Azimuthal quantum number
(D) Spin quantum number

Q.30 How many electrons can fit in the orbital for which $n = 3$ and $\ell = 1$? **[NEET 2016 PHASE 2]**
(A) 2
(B) 6
(C) 10
(D) 14

Q.31 Which of the following pairs of d-orbitals will have electron density along the axes? **[NEET 2016 PHASE 2]**
(A) d_{z^2}, d_{xz} (B) d_{xz}, d_{yz} (C) $d_{z^2}, d_{x^2-y^2}$ (D) $d_{xy}, d_{x^2-y^2}$

Q.32 Which one is the wrong statement ? **[NEET 2017]**
(A) The uncertainty principle is $\Delta E \times \Delta t \geq h/4\pi$
(B) Half filled and fully filled orbitals have greater stability due to greater exchange energy, greater symmetry and more balanced arrangement.
(C) The energy of 2s orbital is less than the energy of 2p orbital in case of Hydrogen like atoms
(D) de-Broglies's wavelength is given by $\lambda = \frac{h}{mv}$, where m = mass of the particle, n = group velocity of the particle

Q.33 Which one is a wrong statement? **[NEET 2018]**
(A) The electronic configuration of N atom is



(B) An orbital is designated by three quantum numbers while an electron in an atom is designated by four quantum numbers.
(C) Total orbital angular momentum of electron in 's' orbital is equal to zero.
(D) The value of m for d_z2 is zero

Q.34 4d, 5p, 5f and 6p orbitals are arranged in the order of decreasing energy. The correct option is **[NEET 2019]**
(A) 5f > 6p > 5p > 4d
(B) 6p > 5f > 5p > 4d
(C) 6p > 5f > 4d > 5p
(D) 5f > 6p > 4d > 5p

Q.35 Which of the following series of transitions in the spectrum of hydrogen atom fall in visible region? **[NEET 2019]**
(A) Lyman series
(B) Balmer series
(C) Paschen series
(D) Brackett series

ANSWER KEY

EXERCISE - 1

Q	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20	21	22	23	24	25
A	A	A	D	A	B	C	C	D	B	C	A	C	B	A	D	D	A	C	B	A	A	C	A	C	B
Q	26	27	28	29	30	31	32	33	34	35	36	37	38	39	40	41	42	43	44	45	46	47	48	49	50
A	A	D	C	B	B	A	D	D	B	A	A	C	B	D	A	B	A	B	B	D	B	D	B	D	A
Q	51	52	53	54	55	56	57	58	59	60	61	62	63	64	65	66	67	68	69	70	71	72			
A	D	B	B	B	D	D	D	C	D	C	C	C	D	C	B	A	D	D	B	B	B	C			

EXERCISE - 2

Q	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20	21	22	23	24	25
A	B	D	D	A	C	C	B	B	A	D	C	B	A	D	D	A	C	A	B	B	B	A	C	D	D
Q	26	27	28	29	30	31	32	33	34																
A	B	B	C	D	A	C	C	C	B																

EXERCISE - 3

Q	1	2	3	4	5	6
A	4	9	4	5	6	2

EXERCISE - 4

Q	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20	21	22	23	24	25
A	A	C	A	D	D	C	CD	B	A	A	C	D	C	B	B	C	C	D	B	C	B	C	A	B	C
Q	26	27	28	29	30	31	32	33	34	35	36	37	38	39	40	41									
A	B	C	A	D	B	B	D	B	D	D	B	A	C	B	A	B									

EXERCISE - 5

Q	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20	21	22	23	24	25
A	B	D	B	B	B	B	D	D	B	C	C	C	B	C	A	C	A	D	A	A	A	D	B	D	A
Q	26	27	28	29	30	31	32	33	34	35															
A	A	B	D	D	A	C	C	A	A	B															

STRUCTURE OF ATOM

TRY IT YOURSELF-1

(1) **(D).** Rutherford's experiment was actually α-particle scattering experiment. α-particle is doubly positively charged helium ion i.e., He-nucleus.

(2) Total number of spectral lines = $\frac{7\,(7-1)}{2} = 21$

Total number of lines in Balmer series = 7 – 2 = 5

(3) **(A).** Energy of electron in H-atom is determined by the expression $E_n = -\frac{13.6}{n^2}\text{eV}$, where n = 1, 2, 3, ...

In excited states, $E_2 = -\frac{13.6}{4} = -3.4\text{ eV}$

$E_3 = -\frac{13.6}{9} = -1.51\text{ eV}$, etc.

(4) $\frac{1}{\lambda} = 109678\ Z^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]\text{cm}^{-1}$

$= 109678 \times 4\left[\frac{1}{4} - \frac{1}{16}\right] = 109678 \times \frac{3}{4}\text{cm}^{-1}$

$\lambda = 91.2 \times \frac{4}{3}\text{nm} = 121.6\text{ nm or } 1.216 \times 10^{-7}\text{ m}$

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

$E_2 = -2.18 \times 10^{-18}\frac{(2)^2}{(2)^2}\text{J / atom} = -2.18 \times 10^{-18}\text{ J/atm}$

$r_n = (0.529\text{Å}) \times \frac{n^2}{Z}$; $r_2 = \frac{0.529 \times 2^2}{2}\text{Å} = 1.058\text{ Å}$

(6) $\frac{1}{\lambda} = 109678\ Z^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]\text{cm}^{-1}$

$= 109678 \times 1 \times \left[\frac{1}{4} - \frac{1}{9}\right]\text{cm}^{-1} = 109678 \times \frac{5}{36}\text{cm}^{-1}$

$\lambda = 91.2 \times \frac{36}{5}\text{cm}^{-1} = 656.64\text{nm}$

(7) **(C).** $E \propto Z^2$; $E_{\text{Hydrogen}} = \frac{-871.6}{4} \times 10^{-20}\text{J}$

(8) **22.85 nm**

For maximum energy, $n_1 = 1$ and $n_2 = \infty$

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

Since R_H is a constant and transition remains the same

$\frac{1}{\lambda} \propto Z^2$; $\frac{\lambda_{He}}{\lambda_H} = \frac{Z_H^2}{Z_{He}^2} = \frac{1}{4}$

Hence, $\lambda_{He} = \frac{1}{4} \times 91.2 = 22.8\text{nm}$

(9) $\Delta E = 2.18 \times 10^{-18} \times Z^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$

$\Delta E = 2.18 \times 10^{-18} \times 4\left(\frac{1}{2^2} - \frac{1}{\infty^2}\right) = 2.18 \times 10^{-18}\text{ J/atom}$

(10) **(C).** $v = \frac{nh}{2\pi mr}$

$\text{K.E.} = \frac{1}{2}m\frac{n^2h^2}{4\pi^2m^2r^2}$; For second orbit $\text{K.E.} = \frac{4h^2}{8\pi^2m^2r^2}$

Since, $r_1 = a_0$, $r_2 = 4a_0$.

$\text{K.E.} = \frac{h^2}{32\pi^2 m a_0^2}$

TRY IT YOURSELF-2

(1) Frequency = $\frac{1}{\text{period}} = \frac{1}{4 \times 10^{-8}} = 2.5 \times 10^7\text{ s}^{-1}$

$\lambda = \frac{c}{\nu} = \frac{3 \times 10^8}{2.5 \times 10^7} = 12\text{m}$

$\bar{\nu} = \frac{1}{\lambda} = \frac{1}{12} = 0.083\text{ m}^{-1} = 8.3 \times 10^{-2}\text{ m}^{-1}$

(2) $\lambda = \frac{h}{p} \Rightarrow p = \frac{h}{\lambda} = \frac{6.6 \times 10^{-34}}{2 \times 10^{-10}} = 3.3 \times 10^{-24}$

Momentum of particle = $3.3 \times 10^{-24}\text{ kg ms}^{-1}$

(3) Energy emitted in 5 minutes = 100 × 5 × 60 = 30,000 joule

$E = nh\nu = \frac{nhc}{\lambda}$

$n = \frac{E\lambda}{hc} = \frac{30000 \times 6000 \times 10^{-10}}{6.6 \times 10^{-34} \times 3 \times 10^8}$

$= 0.9 \times 10^{23} = 9.0 \times 10^{22}$ photons

(4) $\Delta x \times m\Delta v = \frac{h}{4\pi}$

$$\Rightarrow \quad \Delta x = \frac{h}{4\pi \times m\Delta v} = \frac{6.6\times10^{-34}}{9.1\times10^{-31}\times5.7\times10^{5}\times3.14\times4}$$

$$\Rightarrow \quad \Delta x = 1.0 \times 10^{-10}\ \text{m}$$

(5) Threshold energy for photoemission = $h\nu°$

$= 6.6 \times 10^{-34} \times 5.3 \times 10^{14} = 3.49 \times 10^{-19}$J

Energy of photon = 3.3×10^{-19}J

As the energy provided by the striking photon is less, photoelectrons will not be ejected.

(6) Energy acquired by the electron (as kinetic energy) after being accelerated by a potential difference of 2kV (i.e., 2000V) is 2000 eV.

$= 2000 \times 1.602 \times 10^{-19}$ J $= 2 \times 1.602 \times 10^{-16}$ J

i.e., Kinetic energy = $2 \times 1.602 \times 10^{-16}$ J

or $\frac{1}{2} \times 9.1 \times 10^{-31}\, v^2 = 2 \times 1.602 \times 10^{-16}$ J

or $v^2 = \frac{2\times1.602\times10^{-16}\times2}{9.1\times10^{-31}}$ or $v = 2.66 \times 10^{7}\ \text{ms}^{-1}$

$$\therefore \quad \lambda = \frac{h}{mv} = \frac{6.626\times10^{-34}}{9.1\times10^{-31}\times2.66\times10^{7}}$$

$= 1.93 \times 10^{-11}$ m

(7) Energy of photon (E) is given by

$E = h\nu = hc/\lambda$

For violet light, $\lambda = 4000$ Å $= 4000 \times 10^{-10}$ m

$c = 3 \times 10^{8}$ m/s

$$E = \frac{6.626\times10^{-34}\times3\times10^{8}}{4000\times10^{-10}} = 4.970 \times 10^{-19}\ \text{joule}$$

For red light, $\lambda = 6000$ Å $= 6000 \times 10^{-10}$ m

$$E = \frac{6.626\times10^{-34}\times3\times10^{8}}{6000\times10^{-10}} = 3.313 \times 10^{-19}\ \text{joule}$$

(8) $\Delta x \times m\Delta v = \frac{h}{4\pi} \Rightarrow \Delta x \times \Delta v = \frac{h}{4\pi m} = \frac{6.6\times10^{-34}}{9.1\times10^{-31}\times4\times3.14}$

$= 5.77 \times 10^{-5}\ \text{m}^2\,\text{s}^{-1}$

Product of uncertainities = $5.77 \times 10^{-5}\ \text{m}^2\,\text{s}^{-1}$

(9) Energy required to eject the electron,

$E° = 2.13 \times 1.6 \times 10^{-19}$ J $= 3.41 \times 10^{-19}$ J

Energy of incident photon,

$$E = \frac{hc}{\lambda} = \frac{6.63\times10^{-34}\times3\times10^{8}}{400\times10^{-9}} = 4.97 \times 10^{-19}\ \text{J}$$

Maximum kinetic energy = $E - E°$

$= 4.97 \times 10^{-19}$ J $- 3.41 \times 10^{-19}$ J $= 1.56 \times 10^{-19}$ J

(10) $KE = \frac{1}{2}mv^2 \Rightarrow 4.55 \times 10^{-25} = \frac{1}{2}\times9.1\times10^{-31}\times v^2$

$$\Rightarrow \quad v^2 = \frac{2\times4.55\times10^{-25}}{9.1\times10^{-31}} = 10^{6} \Rightarrow v = 10^{3}\ \text{ms}^{-1}$$

$$\lambda = \frac{h}{mv} = \frac{6.6\times10^{-34}}{9.1\times10^{-31}\times10^{3}} = 0.725\times10^{-6}$$

Wavelength = 7.25×10^{-7}

TRY IT YOURSELF-3

(1) 2, 1

(2) As the electron is in 3d orbital, the value of the n = 3, further, as the electron is in d orbital, therefore, *l* = 2.

Now, when *l* = 2, the possible values of m can be

–2, –1, 0, +1, +2.

Lastly, for each value of m, s can have $+\frac{1}{2}$ or $-\frac{1}{2}$.

(3) **(A).** Number of radial nodes = $n - \ell - 1$

So, for 3s: 3 – 0 – 1 = 2

For 2p: 2 – 1 – 1 = 0

(4) (i) n = 2, *l* = 0 means 2s orbital.

(ii) n = 3, *l* = 1 means 3p orbital.

(iii) n = 5, *l* = 2 means 5d orbital.

(iv) n = 6, *l* = 3 means 6f orbital.

(5) **(C).** Both (A) and (D) are correct. Three electrons in the 2p orbitals must have same spin, no matter up spin or down spin.

(6) **(D).** The plus and minus signs indicate that electrons are rotating in opposite directions. They are two quantum mechanical states which have no classical analog.

(7) (i) Since the atomic number of an element is the number of electrons present in all the shells, therefore, the atomic number of the element = 2 + 8 + 5 = 15

(ii) The orbital electronic configuration of the element with atomic number 15 is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p_x^{\ 1}\ 3p_y^{\ 1}\ 3p_z^{\ 1}$.

(iii) Total number of protons in the nucleus = Atomic no. of the element = 15.

(iv) The valency of the element under consideration has three half-filled atomic orbitals, therefore, the valency of the element = 3.

(8) (i) This set of quantum numbers is permitted.

(ii) This set of quantum numbers is also permitted.

(iii) This set of quantum numbers is not permitted because for n = 3 the value of *l* cannot be 3 but can be at the maximum 2.

(iv) This set of quantum numbers is also not permitted as above.

(9) (i) (Z = 11) $1s^2\ 2s^2\ 2p^6\ 3s^1$

(ii) (Z = 15) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p_x^1\ 3p_y^1\ 3p_z^1$.

(iii) (Z = 24) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^1$

(iv) (Z = 28) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^8\ 4s^2$

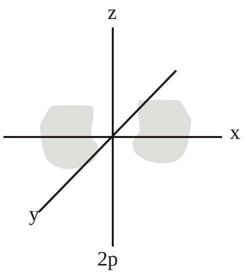
(v) (Z = 31) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^1$

(10) **6.** $|m_\ell| = 1$

means m_ℓ can be +1 and –1

So, for n = 4 six orbitals are possible and bear six electrons with s = –1/2.

(11) **(A).** p_x orbital being dumbell shaped, number of nodal planes = 1, in yz plane.



(12) **(A).** (i) n = 4, *l* = 1 ⇒ 4p orbital

(ii) n = 4, *l* = 0 ⇒ 4s orbital

(iii) n = 3, *l* = 2 ⇒ 3d orbital

(iv) n = 3, *l* = 1 ⇒ 3p orbital

According to aufbau principle, energies of above mentioned orbitals are in the order of

(iv) 3p < (ii) 4s < (iii) 3d < (i) 4p

(13) (i) Since H^- is obtained by adding one electron to H atom

$H + e^- \rightarrow H^-$

Therefore, H^- must have two electrons.

Hence, its electronic configuration is $1s^2$.

(ii) Since F^- is obtained as follows :

$F + e^- \rightarrow F^-$

Therefore, F^- must have one electron more than F atom.

Hence, it must have 10 electrons and its electronic configuration is $1s^2\ 2s^2\ 2p_x^2\ 2p_y^2\ 2p_z^2$.

(iii) Mg^{2+} is obtained by removing two electrons from Mg

$Mg \rightarrow Mg^{2+} + 2e^-$

Therefore, Mg^{2+} must have the following electronic configuration with 10 electrons

$1s^2\ 2s^2\ 2p_x^2\ 2p_y^2\ 2p_z^2$.

(iv) O^{2-} is obtained by adding two electrons to oxygen atom. $O + 2e^- \rightarrow O^{2-}$

Hence, O^{2-} must have the 10 electrons and its electronic configuration must be

$1s^2\ 2s^2\ 2p_x^2\ 2p_y^2\ 2p_z^2$.

CHAPTER-2 : STRUCTURE OF ATOM

EXERCISE-1

(1) **(A).** Mass of an electron = 9.1096×10^{-31} kg

$$1\text{g or } 10^{-3}\text{ kg} = \frac{1}{9.1096 \times 10^{-31}} \times 10^{-3}$$
$$= 1.098 \times 10^{27} \text{ electrons}$$

(2) **(A).** $\frac{q}{m} = \frac{-1.6 \times 10^{-19}}{9.1 \times 10^{-28}} = -1.76 \times 10^{8}$ coulombs / g

(3) **(D).** The characteristics of cathode rays (electrons) do not depend upon the material of electrodes and the nature of the gas present in the cathode ray tube.

(4) **(A).** Charge carried by one electron = 1.6022×10^{-19} C
Electrons present in 5.5×10^{-16} C

$$= \frac{5.5 \times 10^{-16}}{1.6022 \times 10^{-19}} = 3432$$

(5) **(B).** The mass of electron (i.e., 9.10939×10^{-31} kg) is very small as compared to the mass of neutron (i.e., 1.67493×10^{-27} kg).

(6) **(C).** Only a few α-particles (1 in 20,000) bounced back.

(7) **(C).** The conclusion "Electrons move in a circular path of fixed energy called orbits" was given by Bohr's model, not by Rutherford's α-particle scattering experiment.

(8) **(D).** Isobars are the atoms with same mass number (i.e., sum of protons and neutrons) but different atomic number (i.e., no . of protons).

(9) **(B).** ${}^{88}_{38}\text{Sr}$. Atomic number = No. of protons
= No. of electrons = 38
Atomic mass = 88
Number of neutrons = 88 – 38 = 50

(10) **(C).** Let relative abundance of Cl-37 = x %
then relative abundance of Cl-35 = (100 – x)%
Average atomic mass

$$= \frac{x \times 37 + (100 - x)\,35}{100} = 35.5$$

$\Rightarrow\ 37x + 3500 - 35x = 3550 \Rightarrow x = 25 \Rightarrow 100 - x = 75$
Thus, the ratio of Cl-37 and Cl-35 is
$x : (l00 - x) = 25 : 75 = 1 : 3$

(11) **(A).** Only overall neutrality of an atom could be explained correctly by Thomson model of atom.

(12) **(C).** Calculation of the frequency (ν)

$$\nu = \frac{c}{\lambda} = \frac{3 \times 10^{8}\,\text{ms}^{-1}}{5800 \times 10^{-10}\,\text{m}} = 5.172 \times 10^{14}\,\text{s}^{-1}$$

(13) **(B).** Wavelength of yellow light = 240 nm = 240×10^{-9} m

$$\text{Wavenumber } (\overline{\nu}) = \frac{1}{\lambda} = \frac{1}{240 \times 10^{-9}\,\text{m}} = 4.16 \times 10^{6}\,\text{m}^{-1}$$

(14) **(A).** The wavelength, λ, is equal to c/ν , where c is the speed of electromagnetic radiation in vacuum and ν is the frequency. Substituting the given values, we have,

$$\lambda = \frac{c}{\nu} = \frac{3.00 \times 10^{8}\,\text{ms}^{-1}}{1368\text{ kHz}} = \frac{3.00 \times 10^{8}\,\text{ms}^{-1}}{1368 \times 10^{3}\,\text{s}^{-1}}$$
$$= 219.3\text{m}$$

This is a characteristic radiowave wavelength.

(15) **(D).** λ = 242 nm = 242×10^{-9} m
Energy required to ionise one atom of Na

$$(E) = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34}\,\text{Js}) \times (3 \times 10^{8}\,\text{m/s})}{242 \times 10^{-9}\,\text{m}}$$
$$= 8.214 \times 10^{-19} \text{ J/atom}$$

(16) **(D).** $\Delta E = \frac{hc}{\lambda}$

$$\lambda = \frac{6.6 \times 10^{-34}\,\text{Js} \times 3 \times 10^{8}\,\text{m/s}}{3 \times 10^{-19}\,\text{J}} = 6.6 \times 10^{-7}\,\text{m}$$

(17) **(A).** According to Einstein's equation

$$\text{Kinetic energy} = \frac{1}{2} m_e v^2 = h\,(\nu - \nu_0)$$
$$= (6.626 \times 10^{-34}\text{ J s})\,(1.0 \times 10^{15}\,\text{s}^{-1} - 7.0 \times 10^{14}\,\text{s}^{-1})$$
$$= (6.626 \times 10^{-34}\text{ J s})\,(10.0 \times 10^{14}\,\text{s}^{-1} - 7.0 \times 10^{14}\,\text{s}^{-1})$$
$$= (6.626 \times 10^{-34}\text{ J s})\,(3.0 \times 10^{14}\,\text{s}^{-1}) = 1.988 \times 10^{-19}\text{ J}$$

(18) **(C).** In a continuous spectrum, the colours merge into each other in a continuous pattern.

(19) **(B).** Energy (E) of one photon is given by the expression,
E = h ν
h = 6.626×10^{-34} J s, $\nu = 5 \times 10^{14}\,\text{s}^{-1}$ (given)
E = $(6.626 \times 10^{-34}\text{ J s})\,(5 \times 10^{14}\,\text{s}^{-1}) = 3.313 \times 10^{-19}$ J
Energy of one mole of photons
= $(3.313 \times 10^{-19}\text{ J})\,(6.022 \times 10^{23}\,\text{mol}^{-1}) = 199.51\text{ kJ mol}^{-1}$

(20) **(A).** $E = h\nu = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34}\,\text{Js})\,(3 \times 10^{8}\,\text{m/s})}{0.5 \times 10^{-10}\,\text{m}}$

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

(21) **(A).** The frequency of emitted radiation goes from a lower frequency to a higher frequency as the temperature increases.

(22) **(C).** (ii) Balmer, $n_1 = 2$; $n_2 = 3, 4$; region-visible
(iv) Pfund, $n_1 = 5$; $n_2 = 6, 7$; region - infrared

(23) **(A).** For Balmer series, $n_1 = 2$, $n_2 = 3, 4, 5, ...$
The spectral lines are seen in visible region.

(24) **(C).** Since $n_i = 5$ and $n_f = 2$, this transition gives rise to a spectral line in the visible region of the Balmer series.

$$\Delta E = 2.18 \times 10^{-18}\,\text{J} \left[\frac{1}{5^2} - \frac{1}{2^2}\right] = -4.58 \times 10^{-19}\,\text{J}$$

It is an emission energy.
The frequency of the photon (taking energy in terms of magnitude) is given by

$$\nu = \frac{\Delta E}{h} = \frac{4.58 \times 10^{-19}\,\text{J}}{6.626 \times 10^{-34}\,\text{Js}} = 6.91 \times 10^{14}\,\text{Hz}$$

$$\lambda = \frac{c}{\nu} = \frac{3.0 \times 10^{8}\,\text{ms}^{-1}}{6.91 \times 10^{14}\,\text{Hz}} = 434\text{ nm}$$

(25) **(B).** Rydberg's formula for calculation of wavenumbers of lines in hydrogen spectrum is $\overline{\nu} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$

For Balmer series, $n_1 = 2$ and $n_2 = 3, 4, 5$, so third line of the Balmer series is due to transition from fifth Bohr orbit to second Bohr orbit.

(26) **(A).** $E_n = -\dfrac{(2.18\times10^{-18}\,\text{J})\,Z^2}{n^2}\ \text{atom}^{-1}$

For He^+, n = 1, Z = 2

$$E_1 = -\frac{(2.18\times10^{-18}\,\text{J})\,(2^2)}{1^2} = -8.72\times10^{-18}\ \text{J}$$

(27) **(D).** $\nu = \dfrac{\Delta E}{h}$ or $\dfrac{E_2 - E_1}{h}$

E_1 and E_2 are the energies of the lower and higher energy states respectively.

(28) **(C).** Radius of nth orbit is given by $r_n = \dfrac{r_0 \times n^2}{Z}$

For $_3Li^{+2}$, $r = \dfrac{r_0}{3} = \dfrac{0.53}{3} = 0.176$ Å

(29) **(B).** The radius of the orbit is $r_n = \dfrac{(0.0529\ \text{nm})\,n^2}{Z}$

Since n = 1, and Z = 2

$$r_n = \frac{(0.0529\ \text{nm})\,1^2}{2} = 0.02645\ \text{nm}$$

(30) **(B).** In Balmer series, the lines appear in visible region.

(31) **(A).** Energy of the orbit increases as we move away from the nucleus or as the value of n increases.

(32) **(D).** Bohr's atomic theory is applicable to hydrogen and monoatomic ions like He^+, Li^{2+}, Be^{3+}, etc.

(33) **(D).** Bohr's model is not in agreement with Heisenberg's uncertainty principle.

(34) **(B).** $\lambda = \dfrac{h}{mv} = \dfrac{6.626\times10^{-34}\ \text{Js}}{0.01\ \text{kg}\times100\ \text{ms}^{-1}} = 6.626\times10^{-34}\ \text{m}$

(35) **(A).** Since $KE = \dfrac{1}{2}mv^2$

$$v = \left(\frac{2KE}{m}\right)^{1/2} = \left(\frac{2\times3.0\times10^{-25}\ \text{kg m}^2\ \text{s}^{-2}}{9.1\times10^{-31}\ \text{kg}}\right)^{1/2}$$

= 812 m/s

$$\lambda = \frac{h}{mv} = \frac{6.626\times10^{-34}\ \text{Js}}{(9.1\times10^{-31}\ \text{kg})\,(812\ \text{ms}^{-1})}$$

$= 8967\times10^{-10}$ m = 896.7 nm

(36) **(A).** $\Delta x\cdot m\Delta v = \dfrac{h}{4\pi}$

$$m = \frac{6.626\times10^{-34}\ \text{Js}}{4\times3.14\times10^{-8}\ \text{m}\times5.26\times10^{-25}\ \text{m/s}} = 0.01\ \text{kg}$$

(37) **(C).** $\Delta x\cdot\Delta p = \dfrac{h}{4\pi}$ or $\Delta x\ m\Delta v = \dfrac{h}{4\pi}$; $\Delta v = \dfrac{h}{4\pi\ \Delta x\ m}$

$$\Delta v = \frac{6.626\times10^{-34}\ \text{Js}}{4\times3.14\times0.1\times10^{-10}\ \text{m}\times9.11\times10^{-31}\ \text{kg}}$$

$= 0.579\times10^{7}\ \text{m s}^{-1}$ ($1\ \text{J} = 1\ \text{kg m}^2\ \text{s}^{-2}$)

$= 5.79\times10^{6}\ \text{m s}^{-1}$

(38) **(B).** Heisenberg's uncertainty principle rules out the existence of definite paths or trajectories of electrons and other similar particles.

(39) **(D).** $\lambda = 3.6$ Å $= 3.6\times10^{-10}$ m

Velocity of photon = velocity of light

$$m = \frac{h}{\lambda v} = \frac{6.626\times10^{-34}\ \text{Js}}{(3.6\times10^{-10}\ \text{m})\,(3\times10^{8}\ \text{ms}^{-1})}$$

$= 6.135\times10^{-29}$ kg

(40) **(A).** $v = \dfrac{3\times10^8}{10} = 3\times10^7\ \text{m/s}$

$$\lambda = \frac{h}{mv} = \frac{6.626\times10^{-34}\ \text{Js}}{9.1\times10^{-34}\ \text{kg}\times3\times10^{7}\ \text{m/s}}$$

$= 2.43\times10^{-11}$ m

(41) **(B).** $\lambda = h/mv$, if v is same, then higher the value of m means lower will be the value of λ.

Name	Electron	Proton	Neutron	α-particle (He^{2+})
Mass / u	0.00054	1.00727	1.00867	4.0026

Thus, α-particle has the largest mass i.e., shortest wavelength.

(42) **(A).** $\Delta x\cdot m\Delta v = \dfrac{h}{4\pi}$

$\Delta x = 1000$ Å $= 1000\times10^{-10}$ m or 10^{-7} m

$$\Delta v = \frac{6.626\times10^{-34}\ \text{Js}}{4\times3.14\times9.1\times10^{-31}\ \text{kg}\times10^{-7}\ \text{m}}$$

$= 5.79\times10^{2}\ \text{m s}^{-1}$

(43) **(B).** The uncertainty in the speed is 2%, i.e.,

$$45\times\frac{2}{100} = 0.9\ \text{ms}^{-1}$$

$$\Delta x = \frac{h}{4\pi m\Delta v} = \frac{6.626\times10^{-34}\ \text{Js}}{4\times3.14\times40\text{g}\times10^{-3}\ \text{kg g}^{-1}\ (0.9\ \text{ms}^{-1})}$$

$= 1.46\times10^{-33}$ m

This is nearly ~ 10^{18} times smaller than the diameter of a typical atomic nucleus. For large particles, the uncertainty principle sets no meaningful limit to the precision of measurements.

(44) **(B).** When n = 4; ℓ = 0, 1, 2, 3

Number of orbitals = 1 + 3 + 5 + 7 = 16

No. of electrons = 2 + 6 + 10 + 14 = 32

(45) **(D).** The configuration does not follow Hund's rule of maximum multiplicity because 3p will be fully filled before the electrons go to 4s.

(46) **(B).** The region where probability density function reduces to zero is called nodal surfaces.

(47) **(D).** For n = 3, the possible values of ℓ are 0, 1 and 2. Thus there is one 3s orbital (n = 3, $\ell = 0$ & $m_\ell = 0$); there are three 3p orbitals (n = 3, l = 1 and $m_\ell = -1, 0, +1$); there are five 3d orbitals (n = 3, $\ell = 2$ and $m_\ell = -2, -1, 0, +1, +2$). Therefore, the total number of orbitals is 1 + 3 + 5 = 9. The same value can also be obtained by using the relation; number of orbitals = n^2, i.e. $3^2 = 9$.

(48) **(B).** Cr (24) : $3d^5 4s^1$, Cr^{3+} : $3d^3$
Fe(26) : $3d^6 4s^2$, Fe^{3+} : $3d^5$
Mn (25) : $3d^5 4s^2$, Mn^{2+} : $3d^5$
Co (27) : $3d^7 4s^2$, Co^{3+} : $3d^6$
Sc (21) : $3d^1 4s^2$, Sc^{3+} : $3d^0$
Thus, Fe^{3+} and Mn^{2+} have same electronic configuration.

(49) **(D).** For electrons present in M shell the value of other quantum numbers are same. But, the value of spin quantum number will be different.

(50) **(A).** (i) 4p, (ii) 4s, (iii) 3d, (iv) 3p
The order of increasing energy (iv) < (ii) < (iii) < (i)

(51) **(D).** (i) n = 3, $\ell = 2 \Rightarrow 3d$ (ii) n = 4, $\ell = 3 \Rightarrow 4f$

(52) **(B).** The correct representation of the ground state electronic configuration of Cu(29) is
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$

(53) **(B).** (a) n = 2, $\ell = 1$, orbital = 2p
(b) n = 4, $\ell = 0$, orbital = 4s
(c) n = 5, $\ell = 3$, orbital = 5f
(d) n = 3, $\ell = 2$, orbital = 3d

(54) **(B).** Effective nuclear charge is the net positive charge experienced by the electrons from the nucleus. Due to presence of electrons in the innermost shells the electrons in the outermost shell do not experience the full positive charge from the nucleus. This is known as shielding of the outermost shell electrons from the nucleus by the innermost shell electrons.

(55) **(D).** (A) Z_{eff} for an electron in a 2s orbital is greater than that in a 2p orbital (s orbital is more tightly bound to the nucleus than p orbital).
(B) Energy of 2s < 2p, lower the value of n+ℓ, lower is the energy.
(C) Z_{eff} for an electron in 1s orbital is greater than Z_{eff} for an electron in a 2s orbital.
(D) The two electrons present in any orbital have spin quantum numbers with opposite sign
(i.e., $m_s = +1/2$ and $-1/2$).

(56) **(D).** No. of radial nodes = $n - \ell - 1$
For 3p orbital, n = 3, $\ell = 1$
∴ No. of radial nodes for 3p orbital
$= 3 - 1 - 1 = 3 - 2 = 1$

(57) **(D).** Neon (Ne, $1s^2\ 2s^2\ 2p^6$)

(58) **(C).** 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p is the correct order of orbitals with increasing energy.

(59) **(D).** For 1s orbital, the probability density is maximum at the nucleus and it decreases sharply as we move away from it. On the other hand, for 2s orbital, the probability density first decreases sharply to zero and again starts increasing.

(60) **(C).** The electron will enter into an orbital with minimum value of n + ℓ.

(61) **(C).** Total no. of orbitals in n^{th} shell = n^2
∴ Total no. of orbitals in 3rd shell (n = 3) = $3^2 = 9$

(62) **(C).** n = 3, $\ell = 1$ represents 3p orbital. Since p has three orientations p_x, p_y and p_z, 6 electrons will show same quantum number values of n and ℓ.

(63) **(D).**
(A) In the ground state of the atoms, the orbitals are filled in order of their increasing energies.
(B) No two electrons in an atom can have the same set of four quantum numbers.
(C) The maximum number of electrons in the shell with principal quantum number n is equal to $2n^2$.

(64) **(C).** No. of angular nodes = ℓ
For 4d orbital, n = 4, $\ell = 2$
∴ No. of angular nodes for 4d orbital = 2

(65) **(B).** Atomic no. 22 has the following configuration:
${}_{22}X = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^2$

(66) **(A).** Orbital angular momentum = $\sqrt{\ell(\ell+1)}\,\dfrac{h}{2\pi}$
Thus, it depends on 'ℓ' only.

(67) **(D).**
(A) the lower the value of (n + ℓ) for an orbital, the lower is its energy. If two orbitals have the same value of (n + ℓ), the orbital with lower value of n will have the lower energy.
(B) Energies of the orbitals in the same subshell decrease with increase in the atomic number (Z_{eff}).
(C) The orbitals having the same energy are called degenerate.

(68) **(D).**
(A) The size of the s orbital increases with increase in n, that is, 4s > 3s > 2s > 1s
(B) The size of various p orbitals is
4p > 3p > 2p.

(69) **(B).** Choice (B) is an impossible configuration because of the value of the ℓ quantum number. The value of ℓ is only permitted to go from n – 1 to 0. Therefore, ℓ cannot have the same value as n.

(70) **(B).** Oxygen (O, $1s^2\ 2s^2\ 2p^4$)

(71) **(B).** (B) is the better answer because of chromium's exceptional electron configuration. The chromium atom gains increased stability by promoting one of the 4s electrons to the vacant 3d orbital. This promotion creates half-filled orbitals, which increase stability.

(72) **(C).** By placing 8 electrons in the 2p orbitals, (C) is in violation of the Pauli exclusion principle, which states that each orbital may only contain 2 electrons and that those must have opposite spins. Eight electrons exceed the maximum six.

EXERCISE-2

(1) **(B).** Isoelectronic means same number of electrons.
In CO ; it is = 6 + 8 = 14
In N_2 ; it is = 7 + 7 = 14

(2) **(D).** $\lambda = \frac{h}{p} = \frac{h}{\sqrt{2mE}}$ $(\because p = \sqrt{2mE})$

$$\lambda' = \frac{h}{\sqrt{2m\,(16E)}} = \frac{\lambda}{4} = 0.25\lambda$$

% change = –75%

(3) **(D).** The energy(E) of a 300 nm photon is given by

$$h\nu = \frac{hc}{\lambda} = \frac{6.626\times10^{-34}\,\text{Js}\times3.0\times10^{8}\,\text{ms}^{-1}}{300\times10^{-9}\,\text{m}}$$

$= 6.626 \times 10^{-19}$ J

The energy of one mole of photons

$= 6.626 \times10^{-19}\text{ J} \times 6.022 \times10^{23}\text{ mol}^{-1}$

$= 3.99 \times 10^{5}\text{ J mol}^{-1}$

The minimum energy needed to remove one mole of electrons from sodium

$= (3.99 – 1.68)\,10^{5}\text{ J mol}^{-1} = 2.31 \times 10^{5}\text{ J mol}^{-1}$

The minimum energy for one electron

$$= \frac{2.31\times10^{5}\text{ J mol}^{-1}}{6.022\times10^{23}\text{ electron mol}^{-1}} = 3.84 \times10^{-19}\text{ J}$$

(4) **(A).** Ionisation energy for the ground state H atom

$= 2.18 \times 10^{-18}\text{ J atom}^{-1}$

The nuclear charge of helium is 2.

So, $E_{He^{+}\to He^{2+}} = Z^2 \times E_{H\to H^{+}}$

Energy required for removing one electron from He^{+}

$= 2^2 \times (2.18 \times 10^{-18}\text{ J atom}^{-1})$

$= 4 \times 2.18 \times 10^{-18}\text{ J atom}^{-1}$

$= 8.72 \times 10^{-18}\text{ J atom}^{-1}$

(5) **(C).** For 3d, n = 3, ℓ = 2,

m_ℓ = –2, –1, 0, +1, +2

(6) **(C).** An electron at n = 1 has more negative energy than an electron at any other value of n, implying that the electron is more stable and most tightly bound in its smallest orbit. Choice (A) is false since helium has two electrons and the Bohr model therefore cannot be solved exactly. Choice (B) is false since unlike a rainbow, the hydrogen line spectrum contains only a few discrete wavelengths, even for those energy transitions corresponding to the visible region of the electromagnetic spectrum. Choice (D) is incorrect since the Bohr model assumes that all orbits are circular.

(7) **(B).** Mass number of the element = 81 i.e., p + n = 81

Let the number of protons be x.

$$\text{Number of neutrons} = x + \frac{31.7}{100}\times x = 1.317x$$

$\therefore\ x + 1.317\,x = 81$ or $2.317\,x = 81$

or $x = \frac{81}{2.317} = 35$

Symbol of the element = ${}^{81}_{35}\text{Br}$

(8) **(B).** d_{z^2} has electron density in all three axis.

(9) **(A).** ψ^2 is known as probability density and is always positive. From the value of ψ^2 at different points within an atom it is possible to predict the region around the nucleus where electron will most probably be found.

(10) **(D).** The electromagnetic spectrum is organized according to increasing wavelength, decreasing frequency, and decreasing energy per photon. Thus, X-rays, which lie to the left of visible light on the electromagnetic spectrum, have a shorter wavelength, higher frequency, and greater energy per photon than visible light. Therefore, choice (D) is incorrect.

Choice (A) is correct since red light is in the visible region of the electromagnetic spectrum and has a wavelength in the region of 700 nm.

Choice (B) is correct since infrared radiation occurs above (i.e., to the right of) the visible spectrum and therefore has a longer wavelength, a lower frequency, and less energy per photon than visible light. Choice (C) is correct since electromagnetic radiation travels at the speed of light, or 3.00×10^{8} m/s.

(11) **(C).** $\lambda = \frac{h}{mv} = \frac{6.626\times10^{-34}\,\text{Js}}{9.11\times10^{-31}\text{kg}\times2.19\times10^{6}\,\text{m/s}}$

$= 3.32 \times 10^{-10}$ m

(12) **(B).** For Paschen series, n_1 = 3 and n_2 = 4, 5, 6 ...

(13) **(A).** Follow (n + ℓ) rule

(14) **(D).** According to de Broglie's equation

$$\lambda = \frac{h}{mv} = \frac{6.626\times10^{-34}}{\frac{200}{1000}\times\frac{5}{3600}}$$

$$\lambda = \frac{6.626\times10^{-34}\times10^{5}\times36}{10^{3}} = 2.385\times10^{-30}\text{ m}$$

(15) **(D).** Total number of nodes = n – 1

For d-orbital, radial nodes = n – 3 and there are 2 angular nodes. The number of angular nodes is given by ℓ. i.e., for p : 1 angular node, for d : 2 angular nodes and so on.

(16) **(A).** $r_n = \frac{52.9\times n^2}{Z}$ pm

(17) **(C).** $E = \frac{hc}{\lambda} \Rightarrow \frac{E_1}{E_2} = \frac{\lambda_1}{\lambda_2} \Rightarrow \frac{3.32\times10^{-19}}{E_2} = \frac{4000}{6000}$

$\Rightarrow E_2 = 4.98\times10^{-19}\,\text{J} = 3.1\text{eV}.$

(18) **(A).** Total energy = 3.15×10^{-18} J

Wavelength, λ = 600 nm = 600×10^{-9} m

No. of photons received by the detector=?

Energy per photon

$$= \frac{hc}{\lambda} = \frac{6.626\times10^{-34}\,\text{Js}\times(3\times10^{8}\,\text{ms}^{-1})}{600\times10^{-9}\,\text{m}}$$

$= 3.313 \times 10^{-19}$ J

So, no. of photons received by the detector

$$= \frac{3.15\times10^{-18}\text{ J}}{3.313\times10^{-19}\text{ J}} = 10$$

(19) **(B).** ns orbital has (n – 1) nodes.

(20) **(B).** The question asks us to relate wavelength and frequency. We should start by rearranging the equation: $\nu\lambda = c$ or $\nu = c/\lambda$

We then need to put everything in the correct units:

$$5.11\times10^{2}\text{nm}\times\left(\frac{1\text{m}}{10^{9}\text{nm}}\right) = 5.11\times10^{-7}\text{ m}$$

$$\nu = \frac{c}{\lambda} = \frac{(3.00\times10^{8}\text{ m/s})}{(5.11\times10^{-7}\text{ m})} = 5.87\times10^{14}\text{ s}^{-1}$$

(21) **(B).** $_{11}\text{Na} = 1s^2\ 2s^2\ 2p^6\ 3s^1$

For $3s^1$, n = 3, $\ell = 0$, $m_\ell = 0$, s = –1/2

(22) **(A).** The negative sign means that the energy of an electron in an atom is lower than the energy of a free electron at rest. When electron is free from the influence of nucleus, the energy is taken as zero. When the electron is attracted by the nucleus, the energy is emitted and lowered.

(23) **(C).** Phosphorus has atomic number 15.

Electronic configuration :

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$ or $[\text{Ne}]\ 3s^2\ 3p^3$.

(24) **(D).** Wavelength of radiation,

$\lambda = 337.1\text{ nm} = 337.1\times10^{-9}\text{ m}$

No. of photons emitted = 5.6×10^{24}

Energy of photon

$$= \frac{hc}{\lambda} = \frac{6.626\times10^{-34}\text{ Js}\times(3\times10^{8}\text{ ms}^{-1})}{337.1\times10^{-9}\text{ m}} = 5.9\times10^{-19}\text{ J}$$

Total energy produced

$= 5.6\times10^{24}\text{ s}^{-1}\times(5.9\times10^{-19}\text{ J}) = 3.3\times10^{6}\text{ J s}^{-1}$

So, power of the laser = $3.3\times10^{6}\text{ J s}^{-1}$

$= 3.3\times10^{6}\text{ W} = 3.3\times10^{3}\text{ kW}$

(25) **(D).** Wave mechanical model of atom includes :

(i) De Broglie concept of dual character of matter,

(ii) Heisenberg's uncertainty principle

(iii) Schrodinger wave equation.

(26) **(B).** The five d-orbitals are degenerate and have equal energy. The shape of first four orbitals is similar which are d_{xy}, d_{yz}, d_{zx}, $d_{x^2-y^2}$ but the fifth one, d_{z^2} has a different shape.

(27) **(B).** Spin angular momentum of the electron, a vector quantity can have two orientations relative to the chosen axis. These two orientations take the values of + 1/2 or

–1/2 and are called two spin states of the electrons.

(28) **(C).** The statement is incorrect since the nucleus of an atom, consisting of protons and neutrons, represents the majority of atomic mass. The other statements regarding the Rutherford experiment are true.

(29) **(D).**

(A) 5000 Å = 5000×10^{-10} m or 5×10^{-7} m

$$\bar{\nu} = \frac{1}{\lambda} = \frac{1}{5\times10^{-7}\text{ m}} = 2\times10^{6}\text{ m}^{-1}$$

(B) From Bohr's theory, $\nu = 2.18\times10^{6}\times\frac{Z}{n}$ m/s

$$= 2.18\times10^{6}\times\frac{1}{1} = 2.18\times10^{6}\text{ m/s}$$

(30) **(A).** For 2s-orbital, n = 2, $\ell = 0$

(31) **(C).** Work function of Cs atom, $W_0 = 1.9$ eV

We know, $W_0 = h\nu_0 = hc/\lambda_0$

Threshold wavelength,

$$\lambda_0 = \frac{hc}{W_0} = \frac{6.626\times10^{-34}\text{ Js}\times3\times10^{8}\text{ ms}^{-1}}{1.9\text{eV}}$$

$(1\text{ eV} = 1.6021\times10^{-19}\text{ J})$

$$= \frac{1.9878\times10^{-25}\text{ Jm}}{3.044\times10^{-19}\text{ J}} = 6.53\times10^{-7}\text{ m}$$

(32) **(C).** The Aufbau principle states that in the ground state of an atom, an electron enters the orbitals of lowest energy first and subsequent electrons are filled in the order of increasing energies. In the given orbital diagram, option (C) violates the Aufbau principle as electrons switch to p orbitals without completely filling the s orbital which has comparatively lower energy.

(33) **(C).** The outer electronic configuration of chromium atom is $3d^5\ 4s^1$.

(34) **(B).**

	Electron	Proton	Neutron	α-particle
e	1 unit	1 unit	zero	2 units
m	1/1837 unit	1 unit	1 unit	4-units
e/m	1837	1	zero	1/2

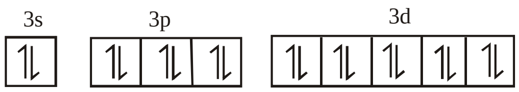
EXERCISE-3

(1) **4.** The Be^{3+} ion has one electron in the 1s orbital (n = 1) in its ground state. Since beryllium has four protons in its nucleus (Z = 4), the energy of the electron in the 1s orbital of Be^{3+} can be calculated.

$$E = -\left(\frac{Z^2}{n^2}\right)(1310)\text{ kJ/mole} = -\left(\frac{4^2}{1^2}\right)(1310)\text{ kJ/mole}$$

$= 2.10\times10^{4}$ kJ/mol

(2) **9.**



So, electrons with spin quantum number = – 1/2 will be 1 + 3 + 5 = 9.

(3) **4.** $E_{photon} = \frac{12400}{3000} = 4.13$ eV

Photoelectric effect can take place only if $E_{photon} \geq \phi$

Thus, Li, Na, K, Mg can show photoelectric effect.

(4) **5.** Since, $\lambda = \frac{h}{mV} = \frac{h}{\sqrt{2M\,K.E.}}$ (Since K.E. $\propto$ T)

$$\Rightarrow \lambda \propto \frac{1}{\sqrt{MT}}$$

For two gases, $\frac{\lambda_{He}}{\lambda_{Ne}} = \sqrt{\frac{M_{Ne}T_{Ne}}{M_{He}T_{He}}} = \sqrt{\frac{20}{4}\times\frac{1000}{200}}$

(5) **6.** $|m_\ell| = 1$

means m_ℓ can be +1 and –1

So, for n = 4 six orbitals are possible and bear six electrons with s = –1/2.

(6) **2.** Te is in group VIA in the periodic table, which tells us that its configuration is s^2p^4. All other shells are completely filled, so they contain only paired electrons. We need only to find out how many unpaired electrons are represented by s^2p^4.

The notation s^2p^4 is a short representation for

$s\underline{\uparrow\downarrow}\ p\underline{\uparrow\downarrow}\ \underline{\uparrow}\ \underline{\uparrow}$.

This shows that an atom of Te contains two unpaired electrons.

EXERCISE-4

(1) **(A).** We know that according to Heisenberg's uncertainty principle, $\Delta x.m.\Delta v = \frac{h}{4\pi}$. Therefore,

$$\Delta x = \frac{h}{\Delta v.m.4\pi} = \frac{6.6\times10^{-34}}{9.218\times10^{-6}\times0.02\times4\times3.14}$$

$= 2.87 \times 10^{-28}$ m

(2) **(C).** From the formula, $E_n = \frac{E_1}{n^2}$

Given values E_1 = – 13.6 eV, n = 3 (Because IInd excited state means third energy level)

$$\text{So, } E_3 = -\frac{13.6}{(3)^2} = -\frac{13.6}{9} = -1.51\ \text{eV}$$

(3) **(A).** From formula, $\Delta x.m.\Delta v = \frac{h}{4\pi}$

$$\Delta v = \frac{h}{\Delta x.m.4\pi} = \frac{6.6\times10^{-34}}{25\times10^{3}\times10^{-5}\times4\times3.14}$$

$= 0.2 \times 10^{-27} = 2.0 \times 10^{-28}$

(4) **(D).** Orbital angular momentum $= \sqrt{\ell(\ell+1)}\,\frac{h}{2\pi}$

For s-e, $\ell = 0$ So, $\sqrt{0\,(0+1)}\,\frac{h}{2\pi}$ = Zero

(5) **(D).** In Bohr series of lines of hydrogen spectrum third line from the red end e^- jump from = 5 to n = 2.

(6) **(C).** From de-Broglie equation,

$$\lambda = \frac{h}{mv} = \frac{6.6\times10^{-34}}{60\times10^{-3}\times10} = 0.01\times10^{-31} = 1.0\times10^{-33}\ \text{m}$$

(7) **(CD).** For 4f orbital correct orbital of quantum number

n = 4, ℓ = 3, m = +1, s = +1/2

or n = 4, ℓ = 3, m = –2, s = +1/2

(8) **(B).** $${}_{24}Cr = 1s^2\ 2s^2\ \underset{\ell=1}{2p^6}\ 3s^2\ \underset{\ell=1}{3p^6}\ 4s^1\ \underset{\ell=2}{3d^5}$$

So, no. of e^- in $\ell = 1 \Rightarrow 6 + 6 = 12$ (p-subshell)

No. of e^- in $\ell = 2 \Rightarrow 5$ (d-subshell)

(9) **(A).** From the formula,

$$\frac{1}{\lambda} = R\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right) = 1.097\times10^{7}\left(\frac{1}{(1)^2}-\frac{1}{(\infty)^2}\right)$$

$$\frac{1}{\lambda} = 1.097\times10^{7}\ \text{m}$$

$$\lambda = \frac{1}{1.097\times10^{7}}\ \text{m} = \frac{10^9}{1.097\times10^{7}}\ \text{nm} = 91\text{nm}$$

(10) **(A).** Isoelectronic speices $\underset{18e^-}{K^+},\ \underset{18e^-}{Ca^{2+}},\ \underset{18e^-}{Sc^{3+}},\ \underset{18e^-}{Cl^-}$

(11) **(C).** From the positive we can find that (d) and (e) having n = 3 and ℓ = 2 (d-subshell) having same energy. So, we can describe it in absence of magnetic and electrical field.

(12) **(D).** Isoelectronic → $\underset{42e^-}{SO_3^{2-}},\ \underset{32e^-}{CO_3^{2-}},\ \underset{32e^-}{NO_3^{-}}$

(13) **(C).** According to Bohr, the angular momentum of an e^- in n^{th} orbit $= \frac{nh}{2\pi}$ So, for 5^{th} orbit $= \frac{5h}{2\pi} = \frac{2.5h}{\pi}$

(14) **(B).** Given value, mass of e^- m = 9.1×10^{-31} kg

Uncertainty in velocity $\Delta v = 300\times\frac{0.001}{100} = 0.003$

$= 3\times10^{-3}$ m/sec

h = 6.63×10^{-34} Js, Δx = ? From the formula,

$$\Delta x = \frac{h}{4\pi.m.\Delta v} = \frac{6.63\times10^{-34}}{4\times3.14\times9.1\times10^{-31}\times3\times10^{-3}}$$

$= 1.93\times10^{-2}$ m

(15) **(B).** n = 3 ℓ = 2, m = 1 s = +1/2

3 d······Highest energy of an atom

(16) **(C).** From the formula, $IP = KZ^2$

For H_2, Z = 1

IP = 1.312×10^{6} J/mol (given value)

$$K = \frac{IP}{Z^2} = \frac{1.312\times10^{6}\ J/mol}{(1)^2} = 1.312\times10^{6}\ \text{J/mol}$$

$$\text{For } n = 1,\ E_1 = \frac{KZ^2}{n^2} = -\frac{1.312\times10^6\times(1)^2}{(1)^2}$$

$$= -1.312\times10^6 \text{ J/mol}$$

$$\text{For } n = 2,\ E_2 = -\frac{1.312\times10^6\times(1)^2}{(2)^2} = -\frac{1.312\times10^6}{4}$$

$$= -0.328\times10^6 \text{ J/mol}$$

Required energy $\Delta E = E_2 - E_1$

$$= [-0.328-(-1.312)]\times10^6 = 9.8\times10^5 \text{ J mol}$$

(17) **(C).** $\lambda = \frac{h}{mv} = \frac{6.63\times10^{-34}\,\text{Js}}{1.67\times10^{-27}\,\text{kg}\times1\times10^3\,\text{m/s}}$

$= 3.97\times10^{-10}\text{m} = 0.397\text{nm}$

(18) **(D).** $\Delta v = 600\times5\times10^{-5}$ m/s ;

$$\Delta x \ge \frac{h}{4\pi m\Delta v} \ge \frac{6.6\times10^{-34}}{4\times3.14\times9.1\times10^{-31}\times600\times5\times10^{-5}}$$

$\approx 1.92\times10^{-3}$ m

(19) **(B).** $IE_{He^+} = 13.6\,Z^2_{He^+}\left[\frac{1}{1^2}-\frac{1}{\infty}\right] = 13.6\,Z^2_{He^+}$

where $Z_{He^+} = 2$, Hence, $13.6\times Z^2_{He^+} = 19.6\times10^{-18}\text{ J atom}^{-1}$

$$(E_1)_{Li^{+2}} = -13.6\times Z^2_{Li^+}\times\frac{1}{1^2}$$

$= -19.6\times10^{-18}\times(9/4) = -4.41\times10^{-17}\text{ J atom}^{-1}$

(20) **(C).** $E = E_1 + E_2$

$$\frac{hc}{\lambda} = \frac{hc}{\lambda_1}+\frac{hc}{\lambda_2}\ ;\ \frac{1}{\lambda} = \frac{1}{\lambda_1}+\frac{1}{\lambda_2}$$

$$\frac{1}{355} = \frac{1}{680}+\frac{1}{\lambda_2}\ ;\ \lambda_2 = 742.76\text{ nm.}$$

(21) **(B).** (a) 4 p (b) 4 s (c) 3 d (d) 3 p
Acc. to $(n + \ell)$ rule, increasing order of energy
(d) < (b) < (c) < (a).

(22) **(C).** Order of ionic radii $Ca^{2+} < K^+ < Cl^- < S^{2-}$
in isoelectronic species as Z/e increases size decreases.

(23) **(A).** $\Delta E = 2.178\times10^{-18}\left(\frac{1}{1^2}-\frac{1}{2^2}\right) = \frac{hC}{\lambda}$

$$2.178\times10^{-18}\left(\frac{1}{1^2}-\frac{1}{2^2}\right) = \frac{6.62\times10^{-34}\times3.0\times10^8}{\lambda}$$

$\therefore\ \lambda \approx 1.214\times10^{-7}$ m

(24) **(B).** $Na \to Na^+ + e^-$ Ist I.E.
$Na^+ + e^- \to Na$
Electron gain enthalpy of Na^+
Because reaction is reverse so then $\Delta H_{eg} = -5.1$ eV.

(25) **(C).** $_{37}Rb = [Kr]\,5s^1$
$n = 5, \ell = 0, m = 0, s = +1/2$

(26) **(B).** Energy of excited state is negative and correspond to n > 1.

$$n = \sqrt{\frac{-13.6}{E_{\text{excited state}}}} = \sqrt{\frac{-13.6}{-3.4}} = \sqrt{4} = 2$$

(27) **(C).** de-Broglie wavelength, $\lambda = \frac{h}{p} \Rightarrow P = \frac{h}{\lambda}$

and $eV = \frac{1}{2}\frac{P^2}{m} \Rightarrow P = \sqrt{2meV} = \frac{h}{\lambda}$

(28) **(A).** $r_n = 0.529\times\frac{n^2}{Z}$; $r_2 = 0.529\times4 = 2.116$ Å

(29) **(D).** $\frac{1}{\lambda} = \bar{\nu} = R_H Z^2\left(\frac{1}{\eta_1^2}-\frac{1}{\eta_2^2}\right)$

$$\bar{\nu} = R_H\times\left(\frac{1}{\eta_1^2}-\frac{1}{8^2}\right) = R_H\times\frac{1}{\eta_1^2}-\frac{R_H}{8^2}$$

$$\bar{\nu} = R_H\times\frac{1}{\eta^2}-\frac{R_H}{64}\ ;\ m = R_H$$

Linear with slope R_H

(30) **(B).** According to $(n + \ell)$ rule : 3p < 3d < 4p < 4d

(31) **(B).** For isoelectronic species the size is compared by nuclear charge.

(32) **(D).** $h\nu - \phi = KE$

$$\Rightarrow \left(\frac{hc}{\lambda}\right)_{\text{incident}} = KE + \phi \Rightarrow \left(\frac{hc}{\lambda}\right)_{\text{incident}} = KE$$

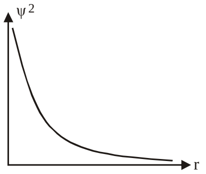
$$KE = \frac{p^2}{2m} = \frac{hc}{\lambda_{\text{incident}}} = \frac{hc}{\lambda} \qquad ...(1)$$

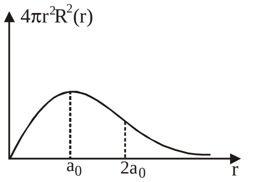
$$\Rightarrow \frac{p^2\times(1.5)^2}{2m} = \frac{hc}{\lambda'} \qquad ...(2)$$

Divide (1) and (2), $(1.5)^2 = \frac{\lambda}{\lambda'} \Rightarrow \lambda' = \frac{4\lambda}{9}$

(33) **(B).** K = 2, 8, 8 ,1
After removal of one electron, second electron we have to remove from another shell, hence there is large difference between first and second ionization energies.

(34) **(D).**





(35) **(D).** Graph of $|\psi^2|$ v/s r, touches r axis at 1 point so it has one radial node and since at r = 0, it has some value so it should be for 's' orbital.

$\therefore\ n - \ell - 1 = 1$ where $\ell = 0 \Rightarrow n - 1 = 1$

$\therefore\ n = 2 \Rightarrow$ '2s' orbital

(36) **(B).** $P(x) = 4\pi x^2 \times [\Psi(x)]^2$
Probability will be maximum at a and c.

(37) **(A).** No. of orbitals = $n^2 = 5^2 = 25$
For n = 5, no. of orbitals = $n^2 = 25$
Total number of orbitals is equal to no. of electrons having $m_s = 1/2$.

(38) **(C).**

(39) **(B).** $r = \frac{a_0 n^2}{Z}$. For Li^{2+}, $r = \frac{a_0 (2)^2}{3} = \frac{4a_0}{3}$

(40) **(A).** $2\pi r = n\lambda$
For n = 1, $r = a_0$
$n = 4, r = 16a_0$
So, $2\pi \times 16a_0 = 4 \times \lambda$
$\lambda = 8\pi a_0$

(41) **(B).** $Eu_{63} \Rightarrow [Xe]\, 4f^7\, 5d^0\, 6s^2$

$Eu^{2\oplus} \Rightarrow [Xe]\, 4f^7$

$Ce_{58} \Rightarrow [Xe]\, 4f^1\, 5d^1\, 6s^2$

$Ce^{3\oplus} \Rightarrow [Xe]\, 4f^1$

EXERCISE-5

(1) **(B).** We know that $E_n = E_1 \times \left(\frac{z}{n}\right)^2$

Given $E_2 = \frac{E_1}{2^2} = -328 \text{ kJ mol}^{-1}$

$E_4 = \frac{E_2}{2^2} = \frac{-328}{4} \text{ kJ mol}^{-1} = -82 \text{ kJ mol}^{-1}$

(2) **(D).** We know that $\Delta x.\Delta p \geq \frac{h}{4\pi}$

$\Delta x.m\Delta v > \frac{h}{4\pi}$; $\Delta v > \frac{h}{4\pi \Delta xm}$

$\Delta v = \frac{6.626 \times 10^{-34}}{4\pi \times 0.1 \times 10^{-10} \times 9.11 \times 10^{-31}} = 5.79 \times 10^6$ m/s

(3) **(B).** Magnetic quantum no. represents the orientation of atomic orbitals in an atom. For example, p_x, p_y & p_z have orientation along X-axis, Y-axis and Z-axis.

(4) **(B).**
(ii) is not possible because for any value of n, ℓ varies from 0 to (n – 1) thus for n = 2, ℓ can be only 0, 1, 2.
(iv) is not possible because for ℓ = 0, m = 0.
(v) is not possible because for ℓ = 2, m varies from – 2 to +2.

(5) **(B).** $\Delta x.\Delta p = \frac{h}{4\pi}$; $\Delta p^2 = \frac{h}{4\pi}$

$\Delta p = \frac{1}{2}\sqrt{\frac{h}{\pi}}$; $\Delta v = \frac{1}{2m}\sqrt{\frac{h}{\pi}}$

(6) **(B).** $\Delta P = m\Delta V$
$1 \times 10^{-18} = 9 \times 10^{-28} \times \Delta V$
$\Delta V = 10^{10}/9 = 1.1 \times 10^9$ cm/s

(7) **(D).** K.E per atom

$$= \frac{(4.4 \times 10^{-19}) - (4.0 \times 10^{-19})}{2}$$

$$= \frac{0.4 \times 10^{-19}}{2} = 2.0 \times 10^{-20}$$

(8) **(D).** The number of sub shell is $(2\ell + 1)$.
The maximum number of electrons in the sub shell is $2(2\ell + 1) = (4\ell + 2)$.

(9) **(B).** For ℓ = 2, m cannot have – 3 value

(10) **(C).** According to de-Broglie equation, $\lambda = \frac{h}{mv}$
Given, $h = 6.6 \times 10^{-34}$ Js, m = 0.66 kg, v = 100 m/s

$\therefore \lambda = \frac{6.6 \times 10^{-34}}{0.66 \times 100} = 1 \times 10^{-35}$ m

(11) **(C).** Number of atomic orbitals in an orbit = $n^2 = 4^2 = 16$

(12) **(C).** $E = \frac{hc}{\lambda}$; $\frac{E_1}{E_2} = \frac{\lambda_2}{\lambda_1} = \frac{25}{50}$; $\lambda_1 = 2\lambda_2$

(13) **(B).** 6s → 4f → 5d → 6p for n = 6

(14) **(C).** Energy of photon obtained from the transition n = 6 to n = 5 will have least energy.

$$\Delta E = 13.6 Z^2 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

(15) **(A)** (n = 4, ℓ = 3) ⇒ 4f subshell
So, total No. of electron in subshell
$= 2(2\ell + 1) = 2(2 \times 3 + 1) = 14$ electron

(16) **(C).** Electronic configuration = $[Kr]\, 5s^1$
Set of quantum numbers ⇒ n = 5, ℓ = 0, m = 0, s = 1/2

(17) **(A).** Orbital angular momentum = $\frac{h}{2\pi}\sqrt{\ell(\ell+1)}$;

$\ell = 1 = \frac{h}{2\pi}\sqrt{2} = \frac{h}{\sqrt{2}\pi}$

(18) **(D).** $\lambda = \frac{C}{\nu} = \frac{3 \times 10^{17} \text{ nm s}^{-1}}{6 \times 10^{15} \text{s}^{-1}} = 50\text{nm}$

(19) **(A).** Correct answer is, for n = 1 the electron has more negative energy than it does for n = 6 which means that the electron is less loosely bound in the smallest allowed orbit.

(20) **(A).** n = 3, ℓ = 1 and m = –1 represent a 3p-orbital, so maximum 2 electrons can be accomodate in it.

(21) **(A).** Orbital is $3p_z$.

(22) **(D).** $E = \frac{hc}{\lambda} = \frac{6.626 \times 10^{-34} \times 3 \times 10^8}{45 \times 10^{-9}} = 4.42 \times 10^{-18}$ J

(23) **(B).** Both Be^{2+} and Li^{+1} have two electrons.

(24) **(D).** Angular momentum = $\sqrt{\ell(\ell+1)}\dfrac{h}{2\pi} = \sqrt{\ell(\ell+1)}\,\hbar$

For d orbital $\ell = 2$

Angular momentum = $\sqrt{2(2+1)}\,\hbar = \sqrt{6}\,\hbar$

(25) **(A).** $Fe^{2+} = 3d^6 4s^0$ six d-electrons, p-electrons in chlorine ($1s^2 2s^2 2p^6 3s^2 3p^5$) are=11 as p-electrons = 6+5 =11

(26) **(A).** ${}_{64}Gd = {}_{54}Xe\ 6s^2\ 4f^7\ 5d^1$

(27) **(B).** Ti (22) = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$

Order of energy is 3s 3p 4s 3d

(28) **(D).** ${}_{63}Eu \rightarrow [Xe]\ 4f^7\ 6s^2$

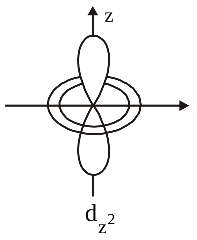
${}_{64}Gd \rightarrow [Xe]\ 4f^7\ 5d^1\ 6s^2$

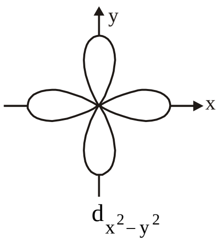
${}_{65}Tb \rightarrow [Xe]\ 4f^9\ 5d^0\ 6s^2$

(29) **(D).** Same orbital can have two different values of spin of e^- of +½ and –½ (spin quantum number).

(30) **(A).** An orbital can accommodate maximum of 2 electrons with anti-parallel spins.

(31) **(C).**

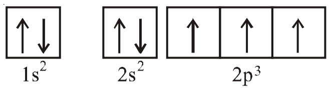




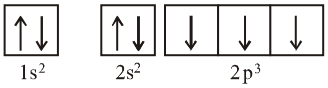
(32) **(C).** In H-like atom energy of 2s = 2p orbital

Incorrect statement is (C).

(33) **(A).** According to Hund's Rule of maximum multiplicity, the correct electronic configuration of N-atom is



OR



Option (A) violates Hund's Rule.

(34) **(A).** (n + ℓ) values for,

4d = 4 + 2 = 6
5p = 5 + 1 = 6
5f = 5 + 3 = 8
6p = 6 + 1 = 7

Correct order of energy would be

5f > 6p > 5p > 4d

(35) **(B).** In H-spectrum, Balmer series transitions fall in visible region.