Chapter 3 The Atomic Structure

Multiple Choice Questions (M.C.Qs)

Choose the correct answer for each from the given options:

01 This colour has the shortest wave length in the visible spectrum:
(a) red (b) violet (c) green (d) yellow

02 The n+l value of 4p orbital is:
(a) 2 (b) 5 (c) 7 (d) 4

03 This series of lines is produced when the electron jumps from the 4th orbital to the 2nd orbit:
(a) Lyman Series (b) Balmer Series (c) Paschen Series (d) Bracektt Series

04 The particle having a mass 1836 times that of the electron is:
(a) neutron (b) proton (c) meson (d) hyperon

05 'No two electrons of an atom can have all the four quantum numbers identical' is the statement of:
(a) Pauli's exclusion principle (b) Hund's rule
(c) Aufbau principle (d) (n+l) rule

06 The maximum number of unpaired electron in 3d energy level is:
(a) 5 (b) 6 (c) 7 (d) 8

07 An electron is said to be excited when it:
(a) loses energy (b) jumps to a lower orbit
(c) jumps to a higher orbit (d) enters the atom

08 The maximum number of electrons in the valence shell of any atom are:
(a) 4 (b) 6 (c) 8 (d) 32

09 The (n+l) value for 5d orbital is:
(a) 4 (b) 5 (c) 6 (d) 7

10 The colour of light depends upon its:
(a) wavelength (b) velocity (c) source (d) none

11 Which atomic orbital is always involved in sigma bonding?
(a) 's' orbital (b) 'p' orbital (c) 'd' orital (d) 'f' orbital

12 The total number of neutron in $_{12}Mg^{2+}$:
(a) 10 (b) 12 (c) 20 (d) 24

13 The smallest particle of matter is known as:
(a) electron (b) atom (c) molecule (d) proton

14 ***Nucleus of an atom contains:***
(a) electrons and protons *(b) protons and neutrons*
(c) electrons and neutrons *(d) all of these.*

15 ***The rays which were deflected towards positive plate of applied electrical field is called:***
(a) Positive rays *(b) Cathode rays* *(c) Gamma rays* *(d) None*

16 ***The e/m ratio of cathode rays resembles with that of :***
(a) γ-rays *(b) proton* *(c) electron* *(d) neutron*

17 ***When 6d orbital is complete, the entering electron goes into:***
(a) 7f *(b) 7p* *(c) 7s* *(d) 7d*

18 ***The mass of electron as determined by J. J. Thomas is:***
(a) 9.1×10^{-10} gm (b) 9.1×10^{-18} gm (c) 9.1×10^{-31} gm (d) 9.1×10^{-25} gm

19 ***Neutron is a fundamental particle which carries:***
(a) + 1 charge and unit mass *(b) no charge and no mass*
(c) no charge and unit mass *(d) -1 charge and unit mass*

20 ***Cathode rays are deflected by:***
(a) neither electric nor magnetic field *(b) electric field only*
(c) both electric and magnetic field *(d) magnetic field only*

21 ***The e/m ratio of an electron is:***
(a) 1.75×10^{9} C/kg (b) 1.75×10^{10} C/kg (c) 1.75×10^{8} C/kg (d) 1.75×10^{11} C/kg

22 ***The charge of an electron is:***
(a) 1.602×10^{-19} C (b) 1.602×10^{-16} C (c) 1.602×10^{-18} C (d) 1.602×10^{-16} C

23 ***The mass of an electron is:***
(a) 9.115×10^{-31} kg (b) 9.115×10^{-28} kg (c) 9.125×10^{-30} kg (d) 9.125×10^{-29} kg

24 ***The rays which are not affected by electric and magnetic field are:***
(a) α-rays *(b) γ-rays* *(c) β-rays* *(d) cathode rays*

25 ***The number of protons present in nucleus are also called:***
(a) Atomic mass *(b) Avogadro's number*
(c) Atomic number *(d) none*

26 ***The orbital having dumb-bell shape is called:***
(a) s-orbital *(b) p-orbital* *(c) d-orbital* *(d) f-orbital*

27 ***The capacity of 'd' orbital for filling electrons:***
(a) 2 *(b) 10* *(c) 6* *(d) 14*

28 ***Among the following which one is not a radioactive substance:***
(a) Thorium *(b) Radium* *(c) Uranium* *(d) Silicon*

29 ***The fast moving electrons like that of cathode rays are:***
(a) α-rays *(b) β-rays* *(c) γ-rays* *(d) None*

30 ***From the following radioactive rays which one has greater penetrating power?***
(a) β-rays *(b) γ-rays* *(c) α-rays* *(d) x-rays*

31 ***Neutron was discovered by:***
(a) Rutherford *(b) Chadwick* *(c) Bohr* *(d) Stoney*

32 ***Which of the following is found in visible region?:***
(a) Lyman series *(b) Balmer series* *(c) Paschen series* *(d) Pfund series*

33 ***Azimuthal quantum number describes the:***
(a) size of an atom (b) shape of an orbital
(c) spin of electrons (d) orientation of an orbital

34 ***Rays which are good ionizer of gases:***
(a) α (b) β (c) γ (d) x

35 ***The number of neutrons in Na^{23}_{11} is :***
(a) 10 (b) 11 (c) 12 (d) 13

36 ***Atomic size in periodic table increases along the:***
(a) groups (b) periods (c) groups & periods (d) none

37 ***In an atom the number of protons is equal to the number of:***
(a) positrons (b) electrons (c) neutrons (d) mesons

38 ***Value of Planck's constant is:***
(a) $h = 6.62\times10^{-28}Js$ (b) $h = 6.62\times10^{-30}Js$ (c) $h = 6.62\times10^{-32}Js$ (d) $h = 6.62\times10^{-34}Js$

39 ***Quantum number of 4d is:***
(a) $n = 1,\ l = 1$ (b) $n = 2,\ l = 2$ (c) $n = 3,\ l = 2$ (d) $n = 4,\ l = 2$

40 ***If the value of l = 0 the orbital is known as:***
(a) s (b) p (c) d (d) f

41 ***The energy required to remove electrons from the an atom in gaseous state is called:***
(a) ionization energy (b) electron affinity
(c) electronegativity (d) electropositivity

42 ***The power of an atom to attract a shared pair of electron is called:***
(a) electronegativity (b) ionization potential
(c) electron affinity (d) electropositivity

43 ***Which element has the highest value of electronegativity?***
(a) N (b) O (c) F (d) Cl

44 ***The angular momentum of electron is:***
(a) $\frac{nh}{2\pi}$ (b) $\frac{n^2h^2}{2\pi}$ (c) $\frac{nh^3}{3\pi}$ (d) $\frac{nh}{\pi}$

45 ***Which rule or principle is violated in this electronic configuration? $1s^2, 2s^2, 2p_x^2, 2p_y^1, 2p_z$***
(a) Hund's rule (b) Aufbau's principle
(c) (n+l) rule (d) Pauli's exclusion principle.

46 ***What happens during jumping up of electrons from lower to higher energy level?***
(a) Energy is released (b) Energy is absorbed
(c) Atomic size increases (d) Atomic size decreases

47 ***The formula to determine the number of electrons in an orbit:***
(a) $2n^2$ (b) $(n+l)$ (c) $(2l+1)$ (d) $3n^3$

48 ***Which quantum number determines the size of an orbital?***
(a) Magnetic (b) Azimuthal (c) Spin (d) Principal

49 ***The orbitals having same energy are called:***
(a) hybrid orbitals (b) atomic orbitals
(c) degenerated orbitals (d) d-orbitals

50 ***Energy packets are called:***
(a) heat (b) quanta (c) momentums (d) wave

51 ***Electromagnetic radiation produced from radioactive substances are called:***
(a) α-rays (b) β-rays (c) γ-rays (d) x-rays

52 ***The shape of an orbital is governed by:***
(a) magnetic quantum number (b) spin quantum number
(c) azimuthal quantum number (d) principal quantum number

53 ***In discharge tube, positive rays are produced:***
(a) from anode (b) from cathode
(c) by the ionization of gas (d) by passing electric current

54 ***Number of unpaired electrons in the electronic configuration: $1s^2$, $2s^2$, $2p^3$***
(a) 1 (b) 2 (c) 3 (d) 5

55 ***Radioactivity was first observed by:***
(a) Marie Curie (b) Pierre Curie (c) Henry Bequral (d) Rutherford

56 ***The spectrum in which the boundary lines are not distinct is called:***
(a) atomic spectrum (b) continuous spectrum
(c) line spectrum (d) None

57 ***An atom of sodium consists of 11 electrons, 11 protons and 12 neutrons, The mass number of this atom is:***
(a) 11 (b) 12 (c) 22 (d) 23

58 ***On the basis of Hund's rule the electronic configuration $1s^2$, $2s^2$, $2p_x^1$, $2p_y^1$, $2p_z^1$ represents:***
(a) Oxygen (b) Nitrogen (c) Fluorine (d) Bromine

59 ***The number of sub level in d-energy orbitals are:***
(a) 1 (b) 3 (c) 5 (d) 7

60 ***The three p-orbitals are situated as:***
(a) parallel to each other (b) perpendicular to each other
(c) in a straight line (d) none of the above

61 ***Which shell is nearest to the nucleus of an atom:***
(a) M (b) N (c) K (d) L

62 ***How many sub shells are present in M-shell?***
(a) 1 (b) 2 (c) 3 (d) 4

63 ***Atom which has no neutron is:***
(a) Hydrogen (b) Deutrium (c) Tritium (d) Helium

64 ***The wave length of the x-rays emitted decreases regularly with increasing:***
(a) atomic number (b) atomic mass (c) atomic radii (d) atomic size

65 ***The uncertainty principle is applied to:***
(a) aeroplane (b) cricket ball (c) electron (d) earth

66 ***Quantum number values for 3d orbital are:***
(a) $n = 3, \ell = 2$ (b) $n = 3, \ell = 3$ (c) $n = 2, \ell = 2$ (d) $n = 3, \ell = 0$

67 ***Most of the radiations coming out from Pitchblende were:***
(a) Protons (b) Electrons (c) Positrons (d) Neutrons

68 *The number of orbitals in each energy level is given by the formula:*
(a) $2n^2$ (b) $(2l+1)$ (c) $2(2l+1)$ (d) n^2

69 *The energy of each quantum of radiation is directly proportional to its:*
(a) wavelength (b) frequency (c) wave number (d) source of energy

Answer Key

1. (b)	11. (a)	21. (d)	31. (b)	41. (a)	51. (c)	61. (c)
2. (b)	12. (b)	22. (a)	32. (b)	42. (a)	52. (c)	62. (c)
3. (b)	13. (b)	23. (a)	33. (b)	43. (c)	53. (c)	63. (a)
4. (b)	14. (b)	24. (b)	34. (a)	44. (a)	54. (c)	64. (b)
5. (a)	15. (d)	25. (c)	35. (c)	45. (a)	55. (c)	65. (c)
6. (a)	16. (c)	26. (b)	36. (a)	46. (b)	56. (b)	66. (a)
7. (c)	17. (b)	27. (b)	37. (b)	47. (a)	57. (d)	67. (b)
8. (c)	18. (c)	28. (d)	38. (d)	48. (d)	58. (b)	68. (d)
9. (d)	19. (c)	29. (b)	39. (d)	49. (c)	59. (c)	69. (b)
10. (a)	20. (c)	30. (b)	40. (a)	50. (b)	60. (b)	

Short / Detailed Answer Questions

Q.01 ***Describe the theory of atomic structure.***

Ans. The theory that matter was made from small indivisible particles called atoms put forwarded by early philosophers Democrites and Leukiphes. In 1808 John Dalton put on this theory on scientific basis.
An atom is electrically neutral. It is a complex organization of matter and energy. Many particles have been discovered within the atom, these sub-atomic particles include electron, proton, neutron, positron, neutrino and several types of mesons and hyprons etc.

Q.02 ***Describe Faraday's experiment OR the passage of electricity through solution.***

Ans. Electrolytes are the chemical solutions through which electricity can be passed in the presence of two electrodes. The positive ions move towards cathode while negative ions move towards anode, here they gain or lose electrons and become neutral.
Faraday found:
(i) the charges of different ions, and *(ii)* the amount of element liberated.
It is found that there is some elementary unit of electric charge is associated with these ions. The ions were observed to carry some integral multiple of this charge. The basic unit of electric charge was later named by Stoney (1891) as 'electron'.

Q.03 ***What are cathode rays?***

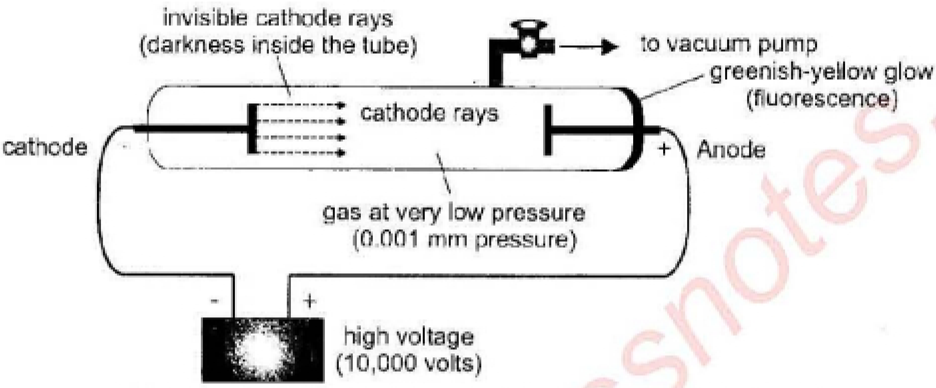
Ans. **Cathode Rays:** Rays emitted by cathode when electricity is passed through a gas taken in the discharge tube at a very low pressure are called cathode rays.

Q.04 ***Describe the discovery of electron through discharge tube experiment.***

Karachi Board: 2005, 2009, 2013, 2015

Ans. **<u>Discharge tube:</u>**

The discharge tube is a sealed glass tube that contains a gas at a greatly reduced pressure to that of normal air pressure at sea level. The tube has a positive electrode (anode) at one end and a negative electrode (cathode) at the other.



Production of cathode rays in a discharge tube

<u>Discharge Tube Experiment:</u>

William Crooks, a British scientist, studied the passage of electricity through gases taken at different pressures of the gas in a discharge tube. He observed that air taken in the gas discharge tube at ordinary pressure did not allow the electricity to flow, even when a source of high potential of about 5000 volt was used. However, when pressure of air was reduced by removing the most of air from the discharge tube then it allowed the current to flow and emitted light when pressure was reduced further to about 0.01 torr, then emission of light by air ceased but the current still flows between the electrodes and produced fluorescent on striking the glass walls opposite the cathode. Rays emitted by cathode when electricity is passed through a gas taken in the discharge tube at a very low pressure are called cathode rays.

Q.05 ***Describe the properties of cathode rays.*** *Karachi Board: 2005, 2007, 2013*

Ans. ***The following are the properties of cathode rays :***

1. ***<u>Cathode rays are negatively charged.</u>***

In 1897, J.J. Thomson established their electric charge by the application of electric field, this proved that they carry a negative charge.

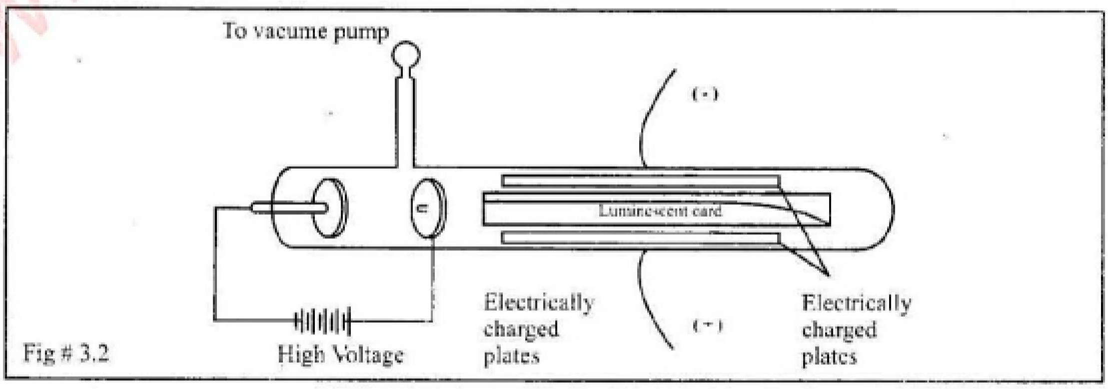
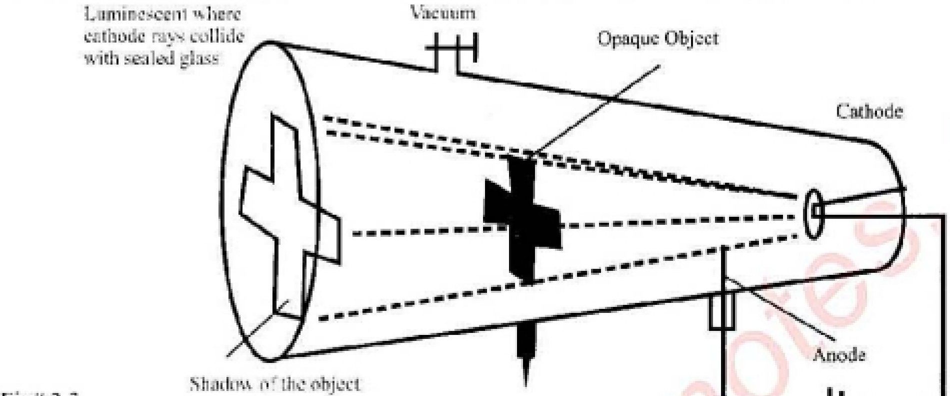


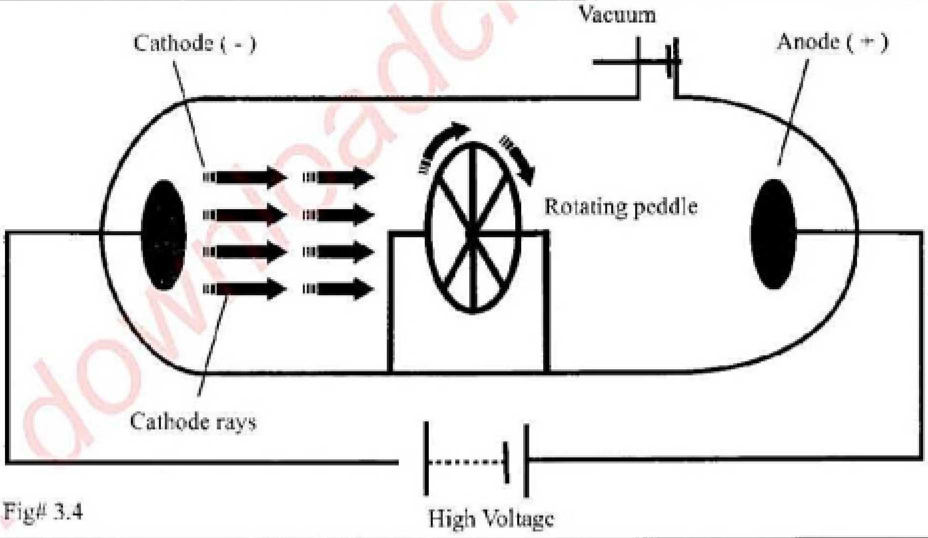
Fig # 3.2

2. ***Cathode rays cast a shadow.***
Cathode rays cast a shadow when an opaque object is placed in their path. This proves that cathode rays travel in a straight line, perpendicular to the surface of cathode.



Fig# 3.3

3. These rays can drive a small paddle wheel placed in their path (see fig 3.4) this shows that these rays possess momentum. From this observation, it is proved that cathode rays are not rays but material particles having a definite mass and velocity and possess kinetic energy.



Fig# 3.4

4. Cathode rays can produce x-rays.
5. Cathode rays can produce heat when they fall on matter e.g. when cathode rays from a concave cathode are focused on platinum foil, it begins to glow.
6. Cathode rays can ionize gases.
7. The rays were seen neither to depend upon the material of which the electrodes made nor upon the gas which is filled in the tube.
8. These rays penetrate small thickness of matter like aluminum or gold foil without producing any perforations in the foil.
9. These rays consist of particles now called electrons carrying a fixed unit of charge and a fixed mass.

Q.06 ***What conclusions were drawn from the properties of cathode rays?***

Ans. ***Conclusions:*** These rays consist of particles now called Electrons carrying a fixed unit of charge and a fixed mass. Different discharge tubes with different electrodes and residual gases were tried by a number of workers. All the experiments gave the same value for charge and mass ratio (e/m). This shows that electrons could be produced from any kind of matter and hence perhaps were constituent of all matter.

Q.07 ***Describe the discovery of protons (positive rays).*** ***Karachi Board: 2007, 2009, 2015***

Ans. As the atom was considered neutral therefore after the discovery of negatively charged particles i.e. electrons, scientists tried to discover a positively charged particle of atom.

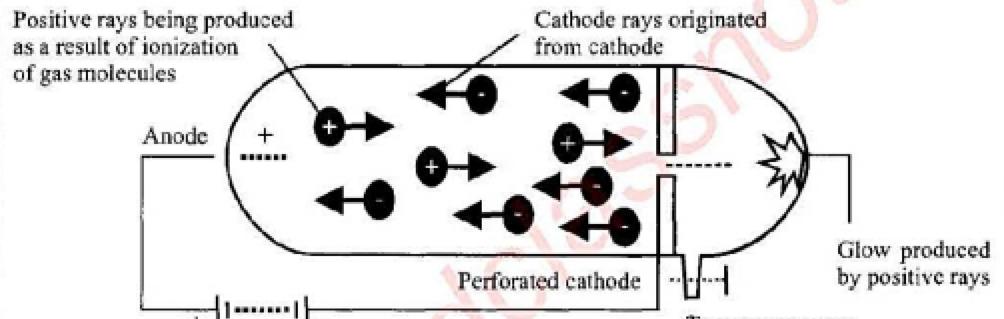


Fig # 3.5

In 1886, Eugene Goldstein used a discharge tube with perforated cathode. He observed a glow in the region behind the cathode. This glow was the result of rays, which were striking the glass wall after passing through the holes of cathode, these rays were named as canal rays, because they pass through the canal or opening into the cathode, these rays are known as the positive rays.

Q.08 ***Give the reason for production of the positive rays.***

Ans. Canal rays or positive rays are the cations, which move towards the cathode and are obtained when cathode rays knock out electrons from the atoms of the gas taken in the discharge tube at low pressures.

Q.09 ***Write down the properties of positive rays (protons).***

Ans.
- *(i)* They are deflected by an electric or magnetic field, showing these are positively charged.
- *(ii)* These rays travel in straight line in a direction opposite to the cathode rays.
- *(iii)* They produce flashes on ZnS plate.
- *(iv)* The e/m value for the positive rays is always smaller than that of electrons and depends upon the nature of the gas used in the discharge tube.
- *(v)* The positive particle obtained from hydrogen is the lightest among all the other positive particle and is called proton (meaning the first particle) having mass 1836 times greater than that of electrons.

Q.10 ***Why the anode rays depend upon the nature of the gas?***

Ans. Anode rays are those particles formed after the removal of one electron from the atoms of the gas, therefore, the mass of every anode rays particle depends upon the nature for the gas. The anode ray for all gaseous substances are different.

Q.11 ***The e/m value of positive rays for different gases are different, but those for cathode rays the e/m value are same. Justify the difference.***

Ans. The nature of positive rays in a discharge tube depends upon the nature of the gas because the nucleus of every gas has its own number of protons and neutrons. Greater the particles in the nucleus, smaller the e/m values of positive rays. But in the case of cathode rays which are always electrons, e/m values remain the same.

Q.12 ***Why it is necessary to decrease the pressure in the discharges tube to get the cathode rays?***

Ans. It is necessary to decrease pressure in the discharge tube because if the gas was at normal pressure there would be too much gas in the way and it would stop being a beam. The electrons in the beam can be deflected if they hit molecules of gas. If too many of them are deflected either none will reach the target or the spot on the target. Lowering the pressure means that there are less molecules to collide with so everything works nicely.

Q.13 ***What is radioactivity? Describe the three types of radiation obtained from radioactive source. How were they discovered?*** ***Karachi Board: 2006-2010-2012-2017-2019***

Ans. ***Radioactivity :*** The spontaneous emission of nuclear radiation from certain elements (U, Ra, Ac) due to disintegration of their nucleus is called radioactivity. Radioactivity is classified into two groups.

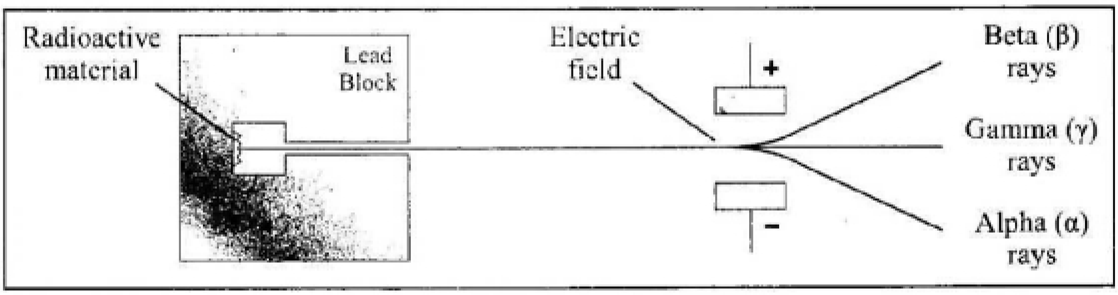
(a) Natural radioactivity *(b)* Artificial radioactivity

Prof. Henry Becqueral in 1895 discovered that there are certain substances (like uranium, thorium, actinium) whose nucleus are heavier and possess large number of neutrons give out three types of rays. These produce fogging on photographic plate and ionizes the gases and produces fluorescence with ZnS, such substances are called radioactive substances and this property is called radioactivity or natural radioactivity.

Discovery of Radioactive rays :

These are discovered by Madam Curie. These rays were found to contains three types of rays:

(i) α - rays
(ii) β - rays
(iii) γ - rays



The rays emitted by the radioactive substance were allowed to pass through two oppositely charges i.e negative and positive plates.
The rays bending towards negatively charged plate carry positive charge and are called alpha rays (α - rays). Those bending towards positively charged plate carry negative charge and are called beta rays (β - rays). While those rays which remain undeflected i.e moving straight and carry no charge are called gamma rays (γ - rays).
All these rays are obtained by placing a photographic plate in their path. Fogging of photographic plate confirm the three types of radioactive rays.

Q.14 ***Differentiate among alpha, beta and gamma rays.*** *Karachi Board: 2007-2010-2012*
OR
Write the properties of the alpha particles. *Karachi Board: 2017-2019*

Ans. The following are the different properties of alpha, beta and gamma rays:

No.	α - Rays	β - Rays	γ - Rays
1	They are helium nuclei (atomic mass = 4 a.m.u.)	They are fast moving electrons.	They are electromagnetic radiations.
2	They have double +ve charge.	They have -ve charge.	They have no charge.
3	Their velocity is about 1/10th of velocity of light.	Their velocity is about the velocity of light.	They travel with velocity of light.
4	Their penetrating power is low.	Their penetrating power is 100 times that of α-particles.	Their penetrating power is more than α & β particles.
5	They are good ionizer of gases.	Their power of ionization is much less than a-particles.	Power of ionization of gases is weak.
6	They are deflected by electric and magnetic fields.	They are also deflected by electric & magnetic fields.	They are not deflected by electric and magnetic fields.
7	They affect photographic plate.	Their e/m ratio resembles that of electrons.	Wavelength of γ rays is shorter than x-rays.

Q.15 ***Describe briefly about the discovery of neutron.***

Ans. Chadwick discovered neutron in 1932 and was awarded Noble prize in Physics in 1935.
Experiment: When Be is bombarded with a-particles, highly penetrating radiations are released. Chadwick made this experiment and found these observations:
- *(i)* These radiations are composed of material particles.
- *(ii)* They carry no charges.
- *(iii)* These particles come out on disintegration of atom by the bombarded a-particles.
- *(iv)* These particles were given the name neutrons and their existence can be indicated by the following equation.

$$_4Be^9 + {}_2He^4 \longrightarrow {}_6C^{12} + {}_0n^1$$

Q.16 Compare the masses and charges of atomic particles.

Ans. The following are the properties of three fundamental particles:

Particles	Charges (C)	Relative Charges	Mass (kg)	Mass (a.m.u.)
Proton	$+1.6022 \times 10^{-19}$	+1	1.6726×10^{-27}	1.0073
Electron	-1.6022×10^{-19}	-1	9.1095×10^{-31}	5.4858×10^{-4}
Neutron	0	0	1.6750×10^{-27}	1.0087

Q.17 Find out the number of protons, neutrons and electrons in the following elements:
Cu_{63}^{29}, Mg_{24}^{12}, Cl_{35}^{-17}, K_{39}^{19}, O_{16}^{8}, F_{19}^{9}, Li_{7}^{+3}, O^{2}, Mg^{2+} *Karachi Board: 2014*

Ans.
No. of Neutrons = Mass number - No. of proton
No. of Protons = No. of Electrons (in neutral atom)

Cu_{63}^{29} :	No. of Neutrons = 63 - 29 = 34,	No. of protons = 29,	No. of electrons = 29
Mg_{24}^{12} :	No. of Neutrons = 24 - 12 = 12,	No. of protons = 12,	No. of electrons = 12
Cl_{35}^{-17} :	No. of Neutrons = 35 - 17 = 18,	No. of protons = 17,	No. of electrons = 18
K_{39}^{19} :	No. of Neutrons = 39 - 19 = 20,	No. of protons = 19,	No. of electrons = 19
O_{16}^{8} :	No. of Neutrons = 16 - 8 = 8,	No. of protons = 8,	No. of electrons = 8
F_{19}^{9} :	No. of Neutrons = 19 - 9 = 10,	No. of protons = 9,	No. of electrons = 9
Li_{7}^{+3} :	No. of Neutrons = 7 - 3 = 4,	No. of protons = 3,	No. of electrons = 0
O^{-2} :	No. of Neutrons = 16 - 8 = 8,	No. of protons = 8,	No. of electrons = 10
Mg^{2+} :	No. of Neutrons = 24 - 12 = 12,	No. of protons = 12,	No. of electrons = 10

Q.18 Describe the mains postulates of Planck's quantum theory? *Karachi Board: 2009, 2008*

Ans. ***Introduction:*** In 1900, Max Planck proposed the quantum theory to explain the emission and absorption of radiations. These radiations are given off by heated objects.

Quanta: According to his revolutionary hypothesis energy travels in a discontinuous manner and is composed of large number of tiny discrete units called quanta.

Main Postulates: The main points of his theory are:

1. Energy is not emitted or absorbed continuously. Each wave packet or quanta is associated with a definite amount of energy. In case of light, the quanta of energy is often called photon.
2. The amount of energy associated with a quanta of radiation is proportional to the frequency of the radiation.
$$E \propto \upsilon$$
or
$$E = h\upsilon$$
where "h" is a constant known as Planck's constant and is equal to
6.625×10^{-34} joule. sec = 6.625×10^{-27} erg sec.
3. A body can emit or absorb energy only in terms of integral multiples of a quanta.
$$E = nh\upsilon \qquad \text{where} \qquad n = 1, 2, 3$$
The frequency 'υ' is related to the wavelength of the photon as
$$\upsilon = \frac{c}{\lambda} \qquad \text{so,} \qquad E = \frac{hc}{\lambda}$$

Q.19 Define spectrum. Name the types of spectrums.

Ans. ***Definition:*** A visual display or dispersion of components of white light when it is passed through a prism is called a spectrum.

Types of spectrum:

1. Continuous spectrum.
2. Line spectrum (Atomic spectrum)

Atomic Spectrum.

- Atomic Emission spectrum
- Atomic Absorption Spectrum

Q.20 Briefly describe about continuous spectrum.

Ans. When white sunlight is allowed to pass through the prism, it deviates and split into seven colours

VIBGYOR: Violet, Indigo, Blue, Green, Yellow, Orange, and Red.

It shows all the colours from red to violet without any dark space as shown in fig # 3.7. The colours diffuse into each other.

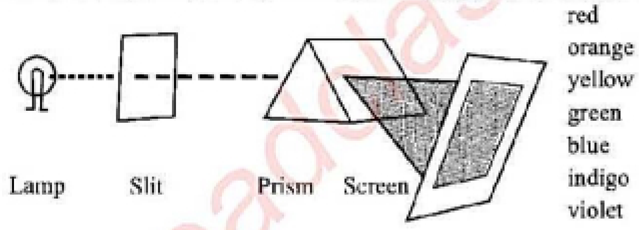


Fig # 3.7 Continuous spectrum

The best example of continuous spectrum is rainbow. It is obtained from the light emitted by the sun or incandescent (electric light) solids. The colour of light depends on its wavelength, violet has shortest wave length (about 4000 Å) and red light the longest (about 7000 Å).

Q.21 Briefly describe about atomic or line spectrum.

Ans. When an element or its compound is volatilized on a flame and the light emitted is seen through a spectrometer, we see distinct lines separated by dark spaces. This type of spectrum is called line spectrum or atomic spectrum. Line spectrum can also be observed when elements in gaseous state are heated at high temperature or subjected to an electric discharge. This is a characteristic of an atom, the number of lines and the distance between them depend upon the element volatilized.

Example:

Line spectrum of sodium contains two yellow coloured lines separated by a definite distance. Similarly, the spectrum of hydrogen consists of a number of lines of different colours having different distances from each other.

The following are the observations from line spectrum:

1. Samples of the same element always emit the same wavelength of radiation.
2. Under the right conditions only certain wavelengths are emitted by any one element.
   This leads to the fact that electrons are arranged around the nucleus in definite energy level E_1 and when they are excited they go to definite excited levels E_2.

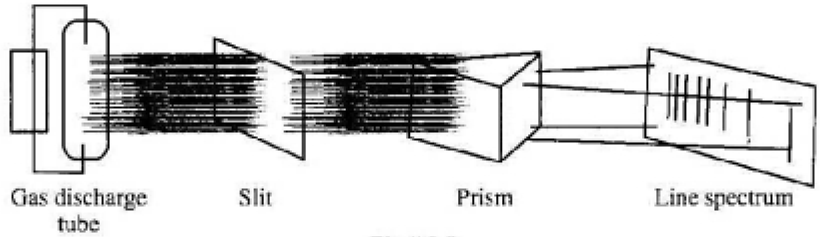


Fig # 3.8

The energies emitted by a given excited element always have the same wavelengths.

Q.22 ***Describe the Rutherford's atomic model about the discovery of nucleus.***

Karachi Board: 2006, 2012, 2014, **2018**

Ans. Lord Rutherford in 1911, performed a classic experiment. He studied the scattering of high speed α-particles which were emitted form a radioactive metal (Ra or Po).

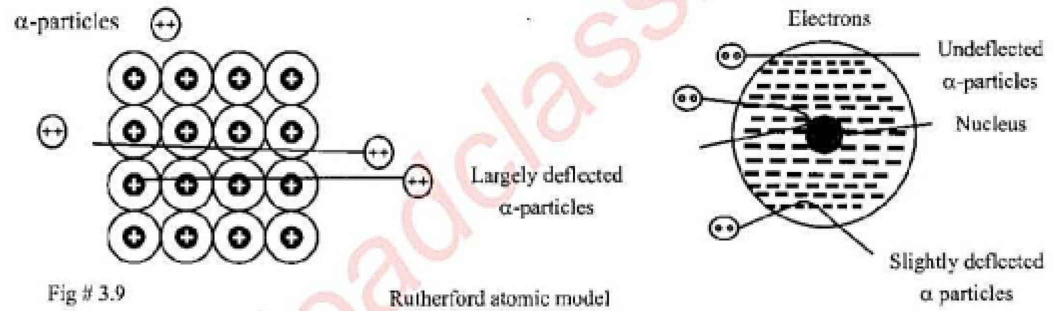


Fig # 3.9 Rutherford atomic model

Experiment: A beam of a-particles emitted form polonium bombarded on thin gold foil through a pinhole in lead plate.

Observations: A screen coated with zinc sulphide was used as a detector, whenever an α-particle struck the screen a flash of light was produced at that point. It was observed that most of the α-particles went through the foil un-deflected. Some were deflected at fairly large angles and a few were deflecting backward. Rutherford proposed that the rebounding particles must have collided with the central heavy portion of the atom which he called as nucleus.

On the basis of the experimental observations, Rutherford proposed the planetary model (similar to the solar system) for an atom in which a tiny nucleus is surrounded by an appropriate number of electrons. Atom as a whole being neutral therefore the nucleus must be having the same number of protons as there are a number of electrons surrounding the nucleus.

Q.23 ***Give the conclusion of Rutherford's experiment.*** *Karachi Board: 2014*, **2018**

Ans. In Rutherford's model for the structure of an atom, the outer electrons could not be stationary, If they were, they would gradually be attracted by the nucleus till they ultimately fall into it. Therefore, to have a stable atomic structure the electrons were supposed to be moving around the nucleus in closed orbits. The centrifugal force kept the nucleus force of attraction balanced.

Q.24 ***Write down the defects of Rutherford's model.*** *Karachi Board: 2006, 2008, 2012, 2015*

Ans. 1. The Rutherford's planet like picture of atom was defective and unsatisfactory because the moving electron must be attracted towards the nucleus, therefore, the radius of the orbiting electron should become smaller and smaller and should fall into nucleus, but we know that this does not happen.

2. If the revolving electron radiates energy continuously, then a continuous spectrum should be produced. On the contrary formation of line spectrum is observed.

Thus according to the classical principle of Physics the Rutherford's model of the atom could not exist.

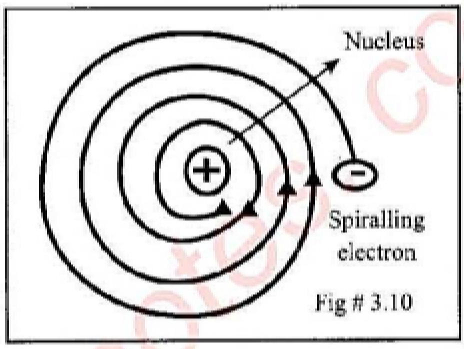


Fig # 3.10

Q.25 ***What is the relationship between x-ray and atomic number?*** *Karachi Board: 2007, 2016*

Ans. W. Roentgen in 1895 discovered that when electrons from cathode collide with the anode in Crook's discharge tube, highly penetrating rays are produced, these rays were named as x-rays and also called Roentgen rays.

Moseley obtained x-rays of different wave length by using anodes made up of different metals.

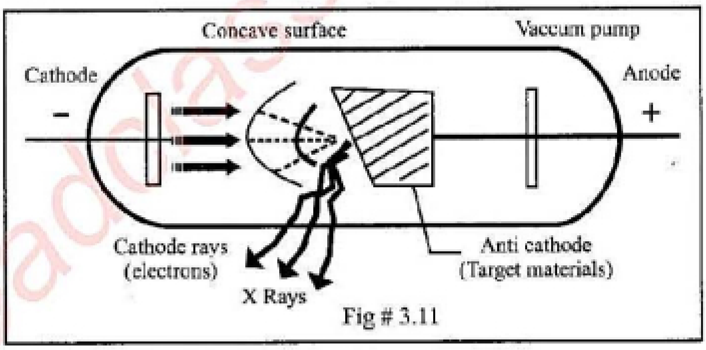


Fig # 3.11

He found that a wave of x-rays decreases regularly with increasing atomic mass. He also found that number of positive charges in nucleus increasing form atom to atom by single electron unit. He called the number of positive charges, the atomic number (Z). Thus atomic number of an element is the number of protons persent in its nucleus.

Q.26 ***Explain the atomic structure on the basis of x-rays.***

Ans.
1. Positively charged nucleus contains protons and neutrons.
2. The electrons are surrounding the nucleus and their number is equal to the number of protons.
3. Atomic number of an atom (Z) is equal to the number of protons.
4. Masses of fundamental particles of C_{12} are given below:

Proton = 1.008, Electron = 0.00055, Neutron = 1.009

Q.27 ***Write down the main postulates of Bohr's model of an atom.***

Karachi Board: 2005, 2008, 2011, 2013, 2015, 2017, 2018

Ans. Bohr put forward a new atomic model by joining the Planck's quantum theory and solar idea of Rutherford.

Main Postulates:

The main postulates of Bohr's atomic theory are:

1. Electrons revolve in circular orbits outside the nucleus. Each orbit has a fixed energy and a quantum number is assigned to it.
2. The electron present in particular orbit do not radiate energy, the energy is emitted or absorbed only when an electron jumps from an orbit to another.
3. When an electron jumps, the energy change, ΔE is given by the Planck's equation:

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

Energy ΔE is the energy difference of any two orbits with energies E_1 and E_2. Energy is absorbed by the electron when it jumps form an inner orbit to an outer and vice versa.

4. The electrons revolve in these orbits with a fixed angular momentum which depends on the quantum number of the orbit. The angular momentum (mvr) is an integral multiple of the factor $h/2\pi$, i.e.,

$$mvr = nh/2\pi$$

where $n = 1, 2, 3, \ldots\ldots\ldots$

The permitted values of angular momenta are therefore

$$\frac{h}{2\pi}, \quad \frac{2h}{2\pi}, \quad \frac{3h}{2\pi} \ldots\ldots\ldots$$

The electron is bound to remain one of these orbit and not in between them, so angular momentum is quantized.

Q.28 Describe the defects of Bohr's atomic model.

Ans. ***The following are the defects of Bohr's atomic model:***

1. It fails to explain spectra of multi-electron atoms
2. It fails to explain the wave nature of electrons.
3. Electrons are moving in three dimensional space and not in fixed orbit as suggested by Bohr.

Q.29 Derive the equation for the radius of the n^{th} orbit of the hydrogen using Bohr's postulates

Karachi Board: 2004, 2009, 2013, 2016

Ans. Let 'm' be the mass of an electron which is revolving with a velocity 'v' and 'r' is the radius of the orbit. The charge on the nucleus is Ze.

Z is the atomic number and 'e' is the charge, then:

(i) Centrifugal Force $= mv^2/r$

(ii) Force of Attraction $= Ze^2/r^2$

Now two types of forces are acting on the revolving electron simultaneously.

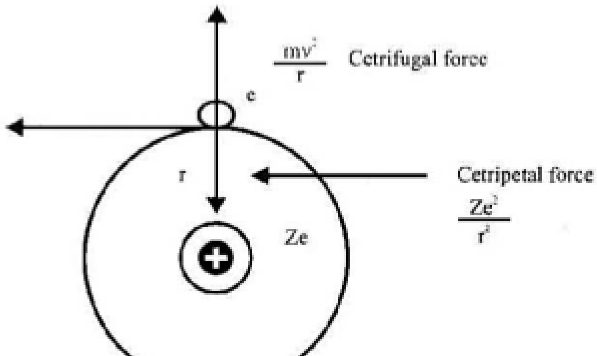


Fig # 3.12

(i) ***Electrostatic force of attraction:***

This force acts between nucleus and revolving electron and tends to attract electron towards the nucleus. According to Coulomb's law, this force is equal to

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

(ii) ***Centrifugal Force:***

This force tends to take the electron away from its orbit and is equal to

$$\frac{mv^2}{r}$$

Both the above forces act in opposite directions. To keep the electron in orbit, both forces must be equal to each other, i.e.

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

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

$$mv^2 = \frac{Ze^2}{r} \qquad \text{\_\_\_\_\_\_\_ (i)}$$

Angular moment is given as $mvr = \frac{nh}{2\pi}$ _______ (ii)

by squaring the equation (ii) we get,

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

$$m^2v^2 = \frac{n^2h^2}{4\pi^2r^2} \qquad \text{\_\_\_\_\_\_\_ (iii)}$$

By dividing equation (iii) by (i) we get,

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

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

$$r = \frac{n^2h^2}{4\pi^2 Ze^2 m} \qquad \text{\_\_\_\_\_\_\_ (iv)}$$

From equation (iv), the radius (r) of Bohr's orbit can be calculated. In the above equation:

$h = 6.625 \times 10^{-27}$ erg. sec or 6.625×10^{-34} js

$\pi = 3.14$

$m_e = 9.11 \times 10^{-28}$ g or 9.11×10^{-31} kg

$e = 4.802 \times 10^{-10}$ e.s.u or 1.602×10^{-19} C

$n = 1$

$Z = 1$ (for H)

For first orbit the radius is found to be 0.529×10^{-8} cm i,e., 0.529 A°.

Q.30 ***Bohr's radius is 0.529 Å. Find the radius of 2^{nd} orbit of Hydrogen atom.***

Karachi Board: 2006

Ans. **Data:**

Bohr's radius = a_o = 0.529 Å
Radius of 2^{nd} orbit = r_{2nd} = ?
Number of orbit = n = 2

Formula:

$$r_{2nd} = n^2 \times a_o$$
$$r_{2nd} = (2)^2 \times 0.529$$
$$r_{2nd} = 4 \times 0.529$$
$$r = 2.116 \text{ Å}$$ **Ans.**

Q.31 ***Bohr's radius is 0.529 Å. Find the radius of 3^{rd} orbit of hydrogen atom.***

Karachi Board: 2007, 2011

Ans. **Data:**

Bohr's radius = a_o = 0.529 Å
Number of orbit = n = 3
Radius of 1^{st} orbit = a_o = 0.529 Å
Radius of 3^{rd} orbit = r_n = ?

Solution:

$$r_n = n^2 \times a_o$$
$$r_n = (3)^2 \times 0.529 = 9 \times 0.529$$
$$r_n = 4.761$$
Radius of 3^{rd} orbit = 4.761 Å **Ans.**

Q.32 ***Derive an expression for the energy of electron in the orbit.***

Karachi Board: 2005, 2006, 2012 , 2018

Ans. The single electron of Hydrogen atom with mass 'm' revolves around the nucleus, it kinetic energy in given by $\frac{1}{2}$ mv^2 where 'v' is the velocity of electron in circular motion.

The potential energy possessed due to the position of isolated electron from the nucleus at a distance 'r' is given by the expression $-\frac{Ze^2}{r}$ (energy of electron at infinity is zero).

So that energy of the electron becomes

$$E\,(\text{energy}) = K.E + P.E$$
$$= \frac{1}{2}mv^2 + [-\frac{Ze^2}{r}]$$
$$= \frac{1}{2}mv^2 - \frac{Ze^2}{r} \qquad \text{\_\_\_\_\_\_(i)}$$

But we know that:

$$mv^2 = \frac{Ze^2}{r} \qquad [\frac{Ze^2}{r^2} = \frac{mv^2}{r}]$$

Put the value of mv^2 in equation (i):

Q.30 ***Bohr's radius is 0.529 Å. Find the radius of 2nd orbit of Hydrogen atom.***

Karachi Board: 2006

Ans. **Data:**

Bohr's radius = a_o = 0.529 Å
Radius of 2nd orbit = r_{2nd} = ?
Number of orbit = n = 2

Formula:

$$r_{2nd} = n^2 \times a_o$$
$$r_{2nd} = (2)^2 \times 0.529$$
$$r_{2nd} = 4 \times 0.529$$
$$r = 2.116 \text{ Å}$$ **Ans.**

Q.31 ***Bohr's radius is 0.529 Å. Find the radius of 3rd orbit of hydrogen atom.***

Karachi Board: 2007, 2011

Ans. **Data:**

Bohr's radius = a_o = 0.529 Å
Number of orbit = n = 3
Radius of 1st orbit = a_o = 0.529 Å
Radius of 3rd orbit = r_n = ?

Solution:

$$r_n = n^2 \times a_o$$
$$r_n = (3)^2 \times 0.529 = 9 \times 0.529$$
$$r_n = 4.761$$
Radius of 3rd orbit = 4.761 Å **Ans.**

Q.32 ***Derive an expression for the energy of electron in the orbit.***

Karachi Board: 2005, 2006, 2012 , 2018

Ans. The single electron of Hydrogen atom with mass 'm' revolves around the nucleus, it kinetic energy in given by $\frac{1}{2}$ mv^2 where 'v' is the velocity of electron in circular motion.

The potential energy possessed due to the position of isolated electron from the nucleus at a distance 'r' is given by the expression $-\frac{Ze^2}{r}$ (energy of electron at infinity is zero).

So that energy of the electron becomes

$$E\,(\text{energy}) = K.E + P.E$$
$$= \frac{1}{2}mv^2 + \left[-\frac{Ze^2}{r}\right]$$
$$= \frac{1}{2}mv^2 - \frac{Ze^2}{r} \quad \text{\_\_\_\_\_\_\_(i)}$$

But we know that:

$$mv^2 = \frac{Ze^2}{r} \qquad \left[\frac{Ze^2}{r^2} = \frac{mv^2}{r}\right]$$

Put the value of mv^2 in equation (i):

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

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

Now substitute the value of $r = \dfrac{n^2h^2}{4\pi^2 mZe^2}$ in equation (ii):

We get $$E = -\frac{Ze^2}{2\left[\dfrac{n^2h^2}{4\pi^2 mZe^2}\right]} = -\frac{Ze^2(2\pi^2 m Ze^2)}{n^2h^2}$$

or $$E = -\frac{Z^2e^4 2\pi^2 m}{n^2h^2} \qquad \text{(iii)}$$

The expression (iii) indicates the energy of electron in an orbit.

Q.33 ***Derive expression for the frequency and wave number of radiation when the electron jumps form higher orbit (n_2) to lower orbit (n_1), the energy of electron $E = \dfrac{-2\pi^2 mZ^2e^4}{n^2h^2}$***

Karachi Board: 2012-2010-2017-2019

Ans. ***(a) Energy of an electron in a lower orbit (n_1) and higher orbit (n_2) can be expressed as follows:***

$$E_1 = \frac{-2\pi^2 Z^2e^4 m}{n_1^2h^2} \quad \text{and} \quad E_2 = \frac{-2\pi^2 Z^2e^4 m}{n_2^2h^2}$$

$$E_2 - E_1 = \frac{-2\pi^2 Z^2e^4 m}{n_2^2h^2} - \left[\frac{-2\pi^2 Z^2e^4 m}{n_1^2h^2}\right]$$

Rearranged $$E_2 - E_1 = \frac{2\pi^2 mZ^2e^4}{n_1^2h^2} - \frac{2\pi^2 mZ^2e^4}{n_2^2h^2} \qquad \text{(i)}$$

i.e. $E_2 - E_1 = h\upsilon$

$$h\upsilon = \frac{2\pi^2 mZ^2e^4}{h^2}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

$$\upsilon = \frac{2\pi^2 mZ^2e^4}{h^3}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] \qquad \text{(ii)}$$

This is the expression for frequency of emitted wave or photon.

(b) Calculation of wave number:

Wave number is defined as the number of waves per unit distance i.e

$\upsilon = c\bar{\upsilon}$ ($\bar{\upsilon}$ = wave number, υ = frequency, c = velocity of light)

Now by substituting the value of frequency from equation (ii) we get:

So. $$c\bar{\upsilon} = \frac{2\pi^2 mZ^2e^4}{h^3}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

or $$\bar{\upsilon} = \frac{2\pi^2 mZ^2e^4}{ch^3}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] \qquad \text{(iii)}$$

$$R_H = \frac{2\pi^2 mZ^2e^4}{ch^3} \qquad \text{(constant)}$$

$$\bar{\upsilon} = R_H Z^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] \qquad \text{(iv)}$$

$Z = 1$ and R_H = Rydberg constant = $109678\ cm^{-1}$

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

$$E = -\frac{Ze^2}{2r} \qquad \text{\_\_\_\_\_\_\_\_ (ii)}$$

Now substitute the value of $r = \frac{n^2h^2}{4\pi^2 mZe^2}$ in equation (ii):

We get $$E = -\frac{Ze^2}{2\left[\frac{n^2h^2}{4\pi^2 mZe^2}\right]} = -\frac{Ze^2(2\pi^2 m Ze^2)}{n^2h^2}$$

or $$E = -\frac{Z^2e^4 2\pi^2 m}{n^2h^2} \qquad \text{\_\_\_\_\_\_\_\_ (iii)}$$

The expression (iii) indicates the energy of electron in an orbit.

Q.33 ***Derive expression for the frequency and wave number of radiation when the electron jumps form higher orbit (n_2) to lower orbit (n_1), the energy of electron $E = \frac{-2\pi^2 mZ^2 e^4}{n^2h^2}$***

Karachi Board: 2012-2010-2017-2019

Ans. ***(a) Energy of an electron in a lower orbit (n_1) and higher orbit (n_2) can be expressed as follows:***

$$E_1 = \frac{-2\pi^2Z^2e^4m}{n_1^2h^2} \quad \text{and } E_2 = \frac{-2\pi^2Z^2e^4m}{n_2^2h^2}$$

$$E_2 - E_1 = \frac{-2\pi^2Z^2e^4m}{n_2^2h^2} - \left[\frac{-2\pi^2Z^2e^4m}{n_1^2h^2}\right]$$

Rearranged $$E_2 - E_1 = \frac{2\pi^2mZ^2e^4}{n_1^2h^2} - \frac{2\pi^2mZ^2e^4}{n_2^2h^2} \qquad \text{\_\_\_\_\_\_\_\_ (i)}$$

i.e. $E_2 - E_1 = h\upsilon$

$$h\upsilon = \frac{2\pi^2mZ^2e^4}{h^2}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

$$\upsilon = \frac{2\pi^2mZ^2e^4}{h^3}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] \qquad \text{\_\_\_\_\_\_\_\_ (ii)}$$

This is the expression for frequency of emitted wave or photon.

(b) Calculation of wave number:

Wave number is defined as the number of waves per unit distance i.e

$\upsilon = c\bar{\upsilon}$ ($\bar{\upsilon}$ = wave number, υ = frequency, c = velocity of light)

Now by substituting the value of frequency from equation (ii) we get:

So. $$c\bar{\upsilon} = \frac{2\pi^2mZ^2e^4}{h^3}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

or $$\bar{\upsilon} = \frac{2\pi^2mZ^2e^4}{ch^3}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] \qquad \text{\_\_\_\_\_\_\_\_ (iii)}$$

$$R_H = \frac{2\pi^2mZ^2e^4}{ch^3} \qquad \text{(constant)}$$

$$\bar{\upsilon} = R_H Z^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] \qquad \text{\_\_\_\_\_\_\_\_ (iv)}$$

$Z = 1$ and R_H = Rydberg constant = 109678 cm^{-1}

Q.34 *What are the assumptions of Bohr's theory?*

Ans. Bohr assumed that on the basis of the quantum theory, there exists the possibility that electron in certain orbits may not give out radiation and an electron revolving in any one of such orbits would be completely stable. Such orbits were called 'Stationary states.' Bohr envisioned the stationary states as circular orbits around the nucleus. He considered that an electron in a certain orbit has a certain energy and as long as it keeps revolving in that orbit, it neither absorbs nor radiates energy. If the electron absorbs energy equal to the energy difference between the two orbits, the electron is excited, i.e. it jumps to higher energy state. If it falls back to lower level, it must emit energy equal to the energy difference between the two orbits. (Fig: 3.11). If this energy is absorbed or emitted as light, a single photon (quantum) of absorbed or emitted light must account for the required energy differences, so that

$$h\upsilon = \Delta E$$

Where ΔE is the difference between the energies of the final and initial orbits, h = Planck's constant (6.625×10^{-34} J.S.) which has the dimensions of energy X time.

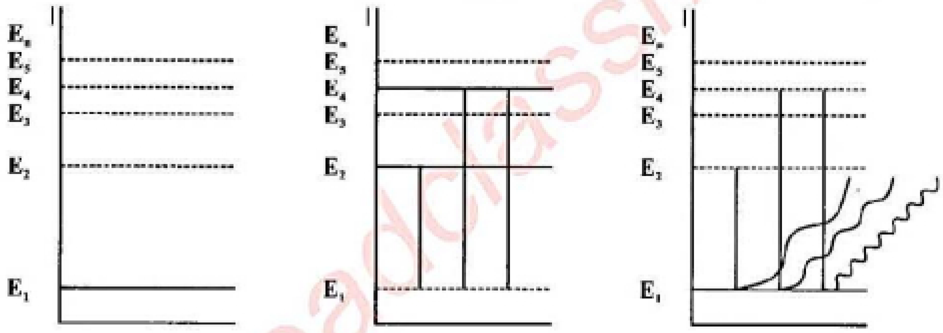


Fig: 3.11
Excitation of electrons from low energy levels to higher energy levels is a process that requires energy. When electrons tall from the excited levels. e.g., E_3. E_4. of E_1. radiant energy is given off, possibly as visible light.

Bohr assumed that the transitions that electrons make between two orbits yield a single unique spectral line.

Bohr further assumed that the stationary states were only those orbits in which the product momentum (mb) x circumference ($2\pi r$), sometimes called 'the action,' was equal to the Planck's constant 'h' or some integral multiple of 'h' therefore for the first possible orbit.

momentum x circumference = h

$$mv \times 2\pi r = h$$

or for any other orbit,

$$mv \times 2\pi r = nh$$

were 'n' was a simple integer, n = 1 for first orbit, n = 2 for the second and so on.

This equation could be re-written as,

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

Here 'mvr' becomes the angular momentum of the electron. Thus Bohr's first condition defining the stationary states could be stated as,

"Only those orbits were possible in which the angular momentum of the electrons would be an integral multiple of $h/2\pi$. These stationary stats correspond to energy levels in the atom."

Q.35 Describe briefly the Bohr's theory and hydrogen spectrum. *Karachi Board: 2010, 2012*

Ans. According to the Bohr's theory " at ordinary temperature, the electron in hydrogen atom resides in lower energy level i.e. first orbit or ground state. When electrically it is heated at low pressure in a discharge tube, the electron of different hydrogen atoms absorb different amount of energy and jump to an appropriate high energy level. They are now said to be in excited state.

All the atoms in excited states, due to their high energy, are unstable. So now the electrons jump back to the original first orbit directly or to some other level of low energy. In doing so they emit energy equal to the difference of energy of two levels. These energetic waves are separated by a prism according to their wavelength and thus hydrogen spectrum is obtained. (Fig # 3.13)

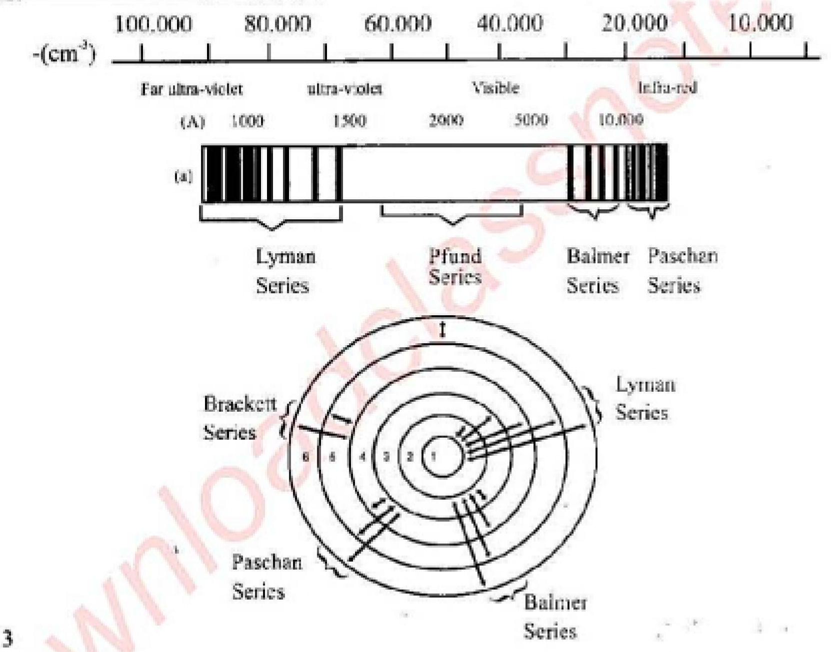


Fig # 3.13
Diagram hydrogen atom spectrum

Q.36 Describe about the spectral lines obtained in hydrogen spectrum.

Ans. Hydrogen spectrum is an example of atomic spectrum. Hydrogen is filled in a discharge tube at a very low pressure, a bluish light is emitted form the discharge tube. This light when viewed through a spectrometer shows several isolated sharp lines. These are called spectral lines. The 9wavelength of these lines lie in the visible, ultraviolet and infrared regions these spectral lines are classified into five groups called spectral series. These series are named after their discoverers as shown below:

1	Lyman series	(Ultra-violet region)
2	Balmer series	(Visible region)
3	Paschen series	(Infra-red region)
4	Brackett Series	(Infra-red region)
5	Pfund series	(Infra-red region)

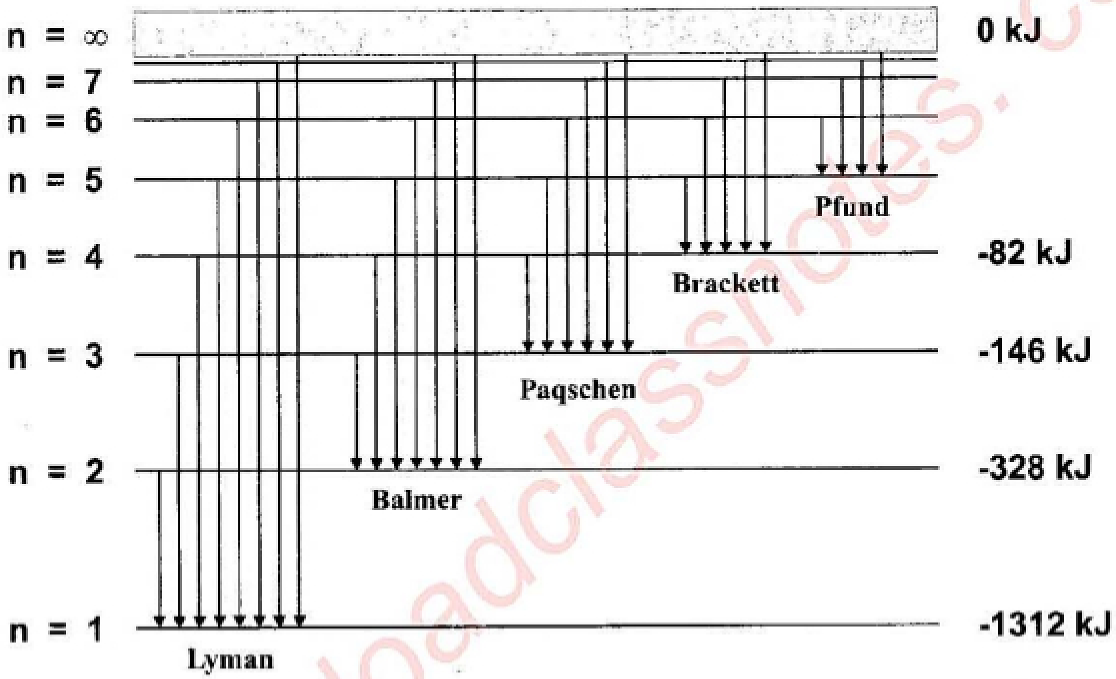
(as shown in fig # 3.14)

Q.37 ***Give detail account of spectral series obtained from hydrogen spectrum.***

Karachi Board: 2015

Ans. According to Bohr's atomic theory the amount of energy in terms of wave number is given by the equation.

$$\bar{\upsilon} = R_H \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$



The frequency of each lines in different series in hydrogen spectrum can be calculated by the following formula:

1. ***Lyman Series:*** Lyman identified a series of lines in the ultraviolet region of spectrum. It is represented by following equation:

$$\bar{\upsilon} = R_H \left(\frac{1}{1^2} - \frac{1}{n_2^2}\right)$$

Hence $n_1 = 1$ and $n_2 = 2,3,4......$ the series is obtained when electron jumps from 2^{nd}, 3^{rd}, 4^{th} orbit to the first orbit.

2 ***Balmer Series:*** It is the series of lines in the visible region of the spectrum. The series is obtained when electron jumps form 3^{rd}, 4^{th}, or 5^{th} energy levels to 2^{nd} energy level. The wave number can be obtained by the following formula:

$$\bar{\upsilon} = R_H \left(\frac{1}{2^2} - \frac{1}{n_2^2}\right)$$

Hence $n_1 = 2$ and $n_2 = 3^{rd}$, 4^{th} or 5^{th} etc.

3 ***Paschen Series:*** It is the series of lines in the infrared region of the spectrum. This series is obtained when electron jumps for 4^{th}, 5^{th} or 6^{th} energy level to 3^{rd} energy level.

$$\bar{\upsilon} = R_H \left(\frac{1}{3^2} - \frac{1}{n_2^2}\right)$$

Hence $n_1 = 3^{rd}$ and $n_2 = 4^{th}$, 5^{th} or 6^{th} etc.

4 ***Brackett Series:*** It is obtained when electron jumps form 5^{th} or 6^{th} or 7^{th} levels to 4^{th} energy level.

5 ***Pfund Series:*** This series in obtained when electron jumps form 6^{th}, 7^{th} or 8^{th} energy to 5^{th} energy level.

Q.38 ***Calculate the wave number of the line in Lymen Series when an electron jumps from 3^{rd} orbit to the 1^{st} orbit ($R_H = 109678\ cm^{1}$).***

Solution:

Karachi Board: 2012, 2008

Higher orbit = n_2 = 3
Lower orbit = n_1 = 1
Rydberg's constant = R_H = 109678 cm^{1}
Wave number = $\bar{\upsilon}$ = ?

$$\bar{\upsilon} = R_H \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$$\bar{\upsilon} = 109678 \left(\frac{1}{(1)^2} - \frac{1}{(3)^2}\right)$$

$$\bar{\upsilon} = 109678 \left(\frac{1}{1} - \frac{1}{9}\right)$$

$$\bar{\upsilon} = 109678 \left(\frac{9 - 1}{9}\right)$$

$$\bar{\upsilon} = 109678 \left(\frac{8}{9}\right)$$

$$\bar{\upsilon} = 109678 \times 0.8888$$

$$\bar{\upsilon} = \mathbf{97491.55}$$ **Ans.**

Q.39 ***Calculate the wave number of spectral lines of hydrogen gas when the electron jumps form $n = 4$ to $n = 2$. ($R_H = 109{,}678\ cm^{-1}$)***

Solution:

Karachi Board: 2011

$n_1 = 4$, $n_2 = 2$
R_H – 109,678 cm^{-1}

$$\bar{\upsilon} = 109678 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$$= 109678 \left(\frac{1}{(4)^2} - \frac{1}{(2)^2}\right) = 109678 \left(\frac{1}{16} - \frac{1}{4}\right)$$

$$= 109678 \left(\frac{1 - 4}{16}\right) = 109678 \left(\frac{(-3)}{16}\right)$$

$$- \frac{109678\ (-3)}{16} = \frac{329034}{16}$$

$$\bar{\upsilon} = \mathbf{20564.63}$$ **Ans.**

Q.40 ***Calculate the wave number of an electron when it jumps from an orbit $n = 5$ to an orbit $n = 1$ ($R_H = 109678\ cm^{-1}$).***

Solution:

Karachi Board: 2008

The general expression (formula) to calculate wave number of the line of any series is given below:

$$\bar{\vartheta} = R_H \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

whereas $n_1 = 1$, $n_2 = 5$ and $R_H = 109678$

Therefore,

$$\bar{\vartheta} = 109678 \left(\frac{1}{1^2} - \frac{1}{5^2}\right) = 109678 \left(\frac{1}{1} - \frac{1}{25}\right)$$

$$\bar{\vartheta} = 109678 \left(\frac{24}{25}\right) = 109678\ (0.96)$$

$$\bar{\vartheta} = \mathbf{105290.88\ cm^{-1}}$$ **Ans.**

Q.41 ***Derive expressions for frequency (υ) and wave number (υ) of energy emitted when an electron jumps form higher energy state (E_2) to lower energy sate (E_1).***

Solution:

Karachi Board: 2010

Energy of an electron in a lower orbit (n_1) and higher orbit (n_2) can be expressed as:

$$E_1 = -\frac{2\pi^2 mZ^2e^4}{n_1^2h^2}$$

$$E_2 = -\frac{2\pi^2 mZ^2e^4}{n_1^2h^2}$$

Energy emitted $\Delta E = E_2 - E_1$

$$E_2 - E_1 = -\frac{2\pi^2 mZ^2e^4}{n_2^2h^2} - \frac{2\pi^2 mZ^2e^4}{n_1^2h^2}$$

$$E_2 - E_1 = -\frac{2\pi^2 mZ^2e^4}{n_2^2h^2} + \frac{2\pi^2 mZ^2e^4}{n_1^2h^2}$$

$$E_2 - E_1 = \frac{2\pi^2 mZ^2e^4}{h^2} \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

Accordingly to Bohr's postulate:

$$E_2 - E_1 = h\upsilon$$

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

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

Expression for wave number:

$$\upsilon = c\bar{\upsilon}$$

Where $\bar{\upsilon}$ is wave number:

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

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

as

$$R_H = \frac{2\pi^2 mZ^2e^4}{ch^3} \left(\frac{1}{n^2} - \frac{1}{n_2^2}\right)$$

$$\bar{\upsilon} = R_H Z^2 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

Q.42 ***Complete the following nuclear reaction.*** ***Karachi Board: 2006***

Ans.

1. $Al_{13}^{27} + He_2^4 \longrightarrow P_{15}^{30} + n_0^1$
2. $N_7^{14} + He_2^4 \longrightarrow O_8^{17} + H_1^1$
3. $Li_3^7 + H_1^1 \longrightarrow He_2^4 + He_2^4$

Q.43 ***Describe Heisenberg's uncertainty principle:*** ***Karachi Board: 2008***

Ans. According to Bohr's theory, an electron is a material particle and it's position as well as momentum can be determined with great accuracy. But with the advent of the concept of wave nature of electron, it is not possible for us to measure the exact position and momentum of electron simultaneously. This was suggested by Heisenberg in 1927.

Principle: Suppose that Δx is the uncertainty in the measurement of position and Δp is the uncertainty of momentum of an electron,

Then $$\Delta x . \Delta p = \frac{h}{2\pi}$$

This relationship is called uncertainty principle. This equation shows that if Δx is small then Δp will be large and vice versa.

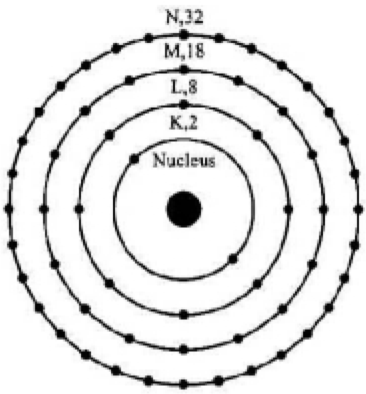
Q.44 ***Describe orbit and orbital.***

Ans. ***Shell / Orbit / Energy Level :***

1st	shell	(K)	=	$2e^-$
2nd	shell	(L)	=	$8e^-$
3rd	shell	(M)	=	$18e^-$

No of electrons (e) in any shell determined by the formula = $2n^2$, where n = No. of Shell For 2nd shell:

No. of e^- = $2(2)^2 = 2(4) = 8e^-$ and so on.

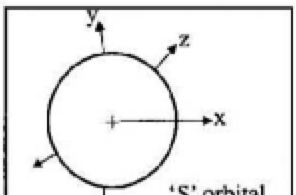


Subshell / orbital / sub energy level :

1st	sub shell	(s)	=	2e
2nd	sub shell	(p)	=	6e
3rd	sub shell	(d)	=	10e
4th	sub shell	(f)	=	14e

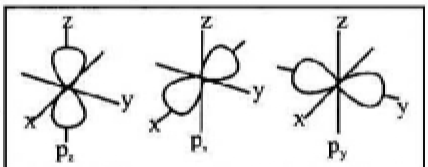
Q.45 ***Write down the shapes of orbitals.*** ***Karachi Board: 2013***

Ans. ***Shape of s-orbital:*** s-orbital has a spherical shape (see fig # 3.15 (a) and is usually represented by a circle. With the increase of value of principle quantum number (n) the size of s-orbital increases. 2s orbital is larger in size than 1s orbital. 2s orbital is further away from the nucleus.



'S' orbital
3.15 (a)

Shape of p-orbitals: There are three values of magnetic quantum number for p-subshell, so p-orbital has three orientations in space i.e. along x, y and z axis. All the these p-orbitals namely p_x , p_y and p_z have dumb bell shape. Fig # 3.15 (b)



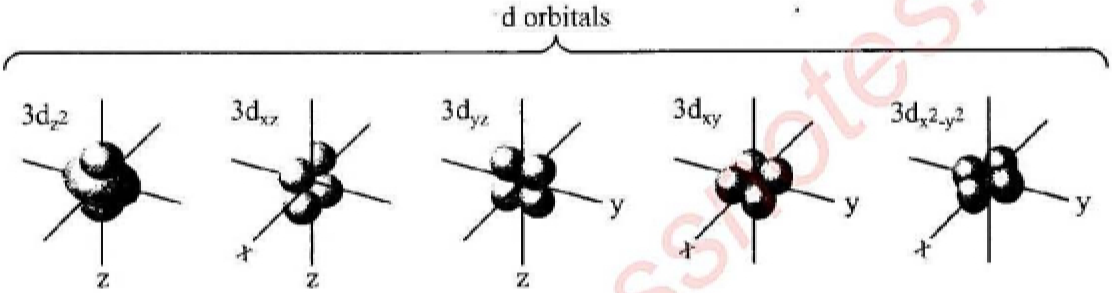
Fig# 3.15(b) Shapes of 'p' orbitals

So p- orbitals have directional character which determines the geometry of molecules. All the p-orbitals of all the energy levels have similar shape but with the increase of principle quantum number of their shell, their size are increased.

Shapes of d-orbitals: For the d sub-shell there are five values of magnetic quantum number. So there are five space orientation and it can be arranged in space in five different ways along x, y and z axis (as shown is fig #. 3.16) they are designated.

d_{xy}, d_{yz}, d_{xz}, $d_{x^2y^2}$, d_{z^2}.

They are not identical in shape. Four d-orbital out of these five contain four lobes while the fifth orbital d_{z^2} consists of two lobes.



Q.46 ***What is the shape of the orbitals for the second energy level.*** *Karachi Board: 2013*

Ans. Spherical and dumb bell.

Q.47 ***Define quantum numbers and name its types.*** *Karachi Board: 2010*

Ans. The set of numerical values which specify the location of electrons in an atom describing their behaviors such as energy, shape of orbital, orientation etc. is called quantum number.

Types:

The quantum number are as under:

(i) Principal Quantum Number (n)
(ii) Azimuthal Quantum Number (ℓ)
(iii) Magnetic quantum Number (m)
(iv) Spin Quantum Number (s)

Q.48 ***Write down the values of all four quantum number for each of the two electron for helium?***

Ans. *Karachi Board: 2009, 2008*

Solution:

Four quantum numbers for the two electrons of helium given as under:

For 1st electron	*For 2nd electron*
n = 1	n = 1
ℓ = 0	ℓ = 0
m = 0	m = 0
s = + ½	s = - ½

Q.49 ***What are quantum numbers ? Discuss their significance.***

Ans. 1. ***Principal Quantum Number (n):***

It tells us about distance of an electron form nucleus, size of shell and energy of the electron. It represents shell (energy level) in which the electron revolves around nucleus.

Mathematical value :

$$n = 1, 2, 3, \ldots\ldots$$

Number of shell is determined by number of electrons. and number of electrons in shell are determined by Bohr's formula "$2n^2$".

Number of shells are represented by:

$n = 1$	,	K shell
$n = 2$	,	L shell
$n = 3$	,	M shell
$n = 4$	,	N shell

2. ***Azimuthal Quantum Number (ℓ)***

It determines the shape of orbital of electrons s-subshell (spherical), p-subshell (dumb-bell) d-subshell (complicated) and f-subshell (more complicated).

Mathematical value: $\ell = 0, 1, 2, 3, \ldots\ldots\ldots (n-l)$

Shapes of Sub - Shell :

Sub-shell are represented by :

1^{st} sub-shell	$= s$,	$(\ell = 0)$	sharp spherical shape
2^{nd} sub-shell	$= p$,	$(\ell = 1)$	principal dumb bell shape
3^{rd} sub-shell	$= d$,	$(\ell = 2)$	diffused orbital, complicated shape
4^{th} sub-shell	$= f$,	$(\ell = 3)$	fundamental, more complicated shape

Number of electrons in a sub-shell are determined by : $2(2\ell+1)$

3. ***Magnetic Quantum number (m)***

This quantum number represents the orientation of an orbital in applied magnetic filed. It tells us about the number of degenerate orbitals.

Mathematical value: $m = 0 \pm 1 \pm 2 \pm 3, \ldots\ldots\ldots\ldots$

number of orientations of sub-shell are represented by:

(i)	$\ell = 0$	(s orbital)	$m = 0$ (one orbital)
(ii)	$\ell = 1$	(p orbital)	$m = -1, 0, +1$ (three degenerate orbitals)
(iii)	$\ell = 2$	(d orbital)	$m = -1, -2, 0, +2, +1$ (5 degenerate orbitals)
(iv)	$\ell = 3$	(f orbital)	$m = -1, -2, -3, 0, +3, +2, +1$ (7 degenerate orbitals)

4. ***Spin Quantum number (s)***

Quantum number which represents clockwise or anti-clockwise rotation of electron in an orbital. It tells us about doublet line structure in spectrum.

Mathematical Value : $s = +\frac{1}{2}, \quad -\frac{1}{2}$

Q.50 ***State Pauli's exclusion principle and explain it giving the example of helium atom.***

Karachi Board: 2005, 2012, 2018

Ans. ***Wolfgang Pauli in 1925 stated that:***

"In an atom no two electrons can have the same set of four quantum numbers"

Consider an electron 'A' present in 1s orbital. For this electron, $n=1$, $\ell=0$ and $m=0$. Suppose the spin of this electron is $s=+\frac{1}{2}$ which will be symbolized by an upward (↑) now if another electron 'B' put in the same orbital (1s) for that electron $n=1$, $\ell=0$, $m=0$. It can occupy this orbital only if the direction of its spin is opposite to that of the first electron so $s=-\frac{1}{2}$ which is symbolized by downward arrow (↓).

Consider an example of He.

$$He = 1s^2$$

For 1st e^-	*For 2nd e^-*
$n = 1$	$n = 1$
$\ell = 0$	$\ell = 0$
$m = 0$	$m = 0$
$s = -\frac{1}{2}$	$s = -\frac{1}{2}$

As per Pauli's principle spin of both electrons should be opposite.

Q.51 ***Describe electronic configuration.***

Ans. The electronic configuration of an element refers to the arrangement of electrons in shells, sub shells or orbitals in their ground state. The electrons in the atoms are at the lowest energy level.

The following rules are used to determine the ground state electronic configuration of an element.

(i) Pauli's Exclusion Principle
(ii) Aufbau Principle
(iii) $(n-\ell)$ rule
(iv) Hund's rule

Q.52 ***Briefly describe Aufbau principle.***

Ans. The meaning of Aufbau is "building up". This principle gives us a sequence in which various orbitals are filled with electrons.

This principle states that the electrons in their ground states occupy orbitals in order of the orbitals energy levels, the lowest energy orbitals are always filled first with electrons and only then the orbitals of higher energy are filled

$1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d\ldots$

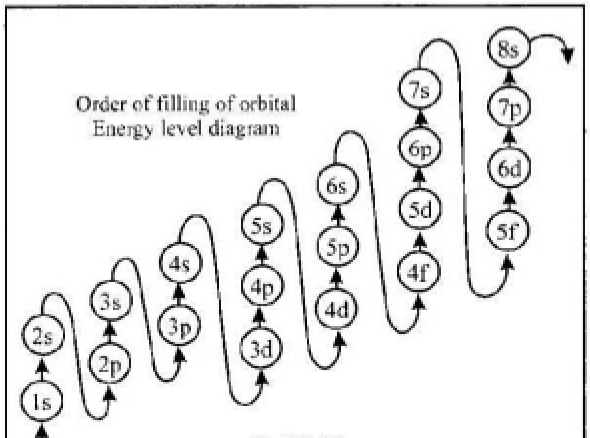


Fig # 3.17

Q.50 ***State Pauli's exclusion principle and explain it giving the example of helium atom.***

Karachi Board: 2005, 2012 , 2018

Ans. ***Wolfgang Pauli in 1925 stated that:***

"In an atom no two electrons can have the same set of four quantum numbers"

Consider an electron 'A' present in 1s orbital. For this electron, $n=1$, $\ell=0$ and $m=0$. Suppose the spin of this electron is $s=+\frac{1}{2}$ which will be symbolized by an upward (↑) now if another electron 'B' put in the same orbital (1s) for that electron $n = 1$, $\ell = 0$, $m = 0$. It can occupy this orbital only if the direction of its spin is opposite to that of the first electron so $s = -\frac{1}{2}$ which is symbolized by downward arrow (↓).

Consider an example of He.

$$He = 1s^2$$

For 1st e^-	*For 2nd e^-*
$n = 1$	$n = 1$
$\ell = 0$	$\ell = 0$
$m = 0$	$m = 0$
$s = +\frac{1}{2}$	$s = -\frac{1}{2}$

As per Pauli's principle spin of both electrons should be opposite.

Q.51 ***Describe electronic configuration.***

Ans. The electronic configuration of an element refers to the arrangement of electrons in shells, sub shells or orbitals in their ground state. The electrons in the atoms are at the lowest energy level.

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(i) Pauli's Exclusion Principle
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This principle states that the electrons in their ground states occupy orbitals in order of the orbitals energy levels, the lowest energy orbitals are always filled first with electrons and only then the orbitals of higher energy are filled

1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d...

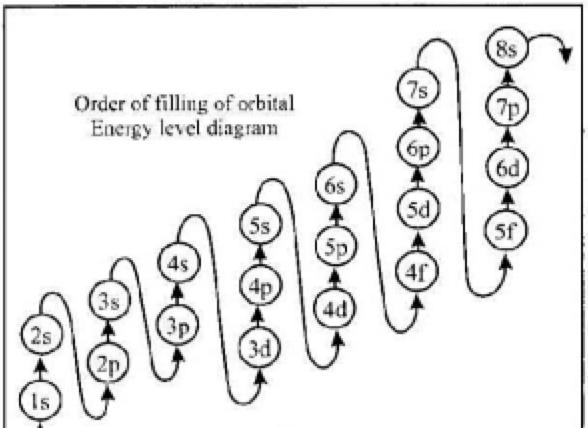


Fig # 3.17

Q.53 ***Describe (n + ℓ) rule with example.*** *Karachi Board: 2019*

Ans. This rule is also called Wiswesser's rule .

"In building up the electronic configuration of elements the orbital with the lowest value of $(n+\ell)$ fill first. When two orbitals have the same value of $(n+\ell)$ the orbital with the lower value of 'n' fill first."

Examples:

(a) Which orbital 3d or 4s will fill first ?

(i) The values for 3d sub-shell, apply — $n = 3, \quad \ell = 2$; $(n+\ell) = 3 + 2 = 5$

(ii) The values for 4s sub-shell apply — $n = 4, \quad \ell = 0$; $(n+\ell) = 4 + 0 = 4$

Since $(n+\ell)$ value of 4s sub-shell is less than 3d, the 4s sub-shell has lower energy and is filled first.

(b) Which orbital 3d or 4p will fill first ?

(i) For 3d sub-shell, — $n = 3, \quad \ell = 2$; $(n+\ell) = 3 + 2 = 5$

(ii) For 4p sub-shell — $n = 4, \quad \ell = 1$; $(n+\ell) = 4 + 1 = 5$

In both cases the value of $(n+\ell)$ is the same i.e, 5, hence electrons will enter into that level which has lower "n" value i.e. 3d.

Q.54 ***Arrange the orbitals in order of ascending energy according to (n+l) rule.***

Karachi Board: 2009, 2011, 2012, 2016

Ans. *(i)* ***3d, 4s, 4p, 4d, 5s, 6s, 5p***

Solution: 4s, 3d, 4p, 5s, 4d, 5p, 6s

(ii) ***4d, 7s, 4f***

Solution: 4d, 4*f*, 7s

Q.55 ***State Hund's rule and justify it with examples .*** *Karachi Board: 2017*

Ans. If more than one degenerating orbitals are available the electron will prefer to occupy different orbitals with same spins rather to occupy same orbital and with opposite spin. (↑↓) ( )

Example:

		$1s^2$	$2s^2$	$2p_x^1$	$2p_y^2$	$2p_z$
Carbon	$Z = 6$	↑↓	↑↓	↑	↑	
Nitrogen	$Z = 7$	↑↓	↑↓	↑	↑	↑

Q.56 ***Write electronic configuration of the following and also give the number of protons and electrons in each.***
(a) Cu (Z = 29) (b) Mg^{+2} (Z = 12) (c) Cl^- (Z = 17)

Karachi Board: 2012, 2017

Ans. (a) Cu^{29} = $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^9$
No of protons = 29 No. of electrons = 29

Karachi Board: 2019

(b) Mg^{+2} = $1s^2, 2s^2, 2p^6, 3s^0$
No. of protons = 12 No. of electrons = 10

(c) Cl^- = $1s^2, 2s^2, 2p^6, 3s^2, 3p^6$
No. of protons = 17 No. of electrons = 18

Q.57 ***What rules and principles are violated in the following electronic configuration?***
(a) $1s^2, 2s^3$, (b) $1s^2, 2p^2$, (c) $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^2, 4s^1$ *Karachi Board: 2011*

Ans.
(a) Aufbau rule
(b) Pauli's Exclusion Principle
(c) $(n + \ell)$ rule

Q.58 ***Which rule is violated in the following electronic configurations?***
(i) $2s^2 2p^6 3s^2$ (ii) $1s^2 2s^2 3s^2 2p^6$ (iii) $1s^2 2s^2 2p_x^{\ 2}$ *Karachi Board: 2015*

Ans.
(i) Aufbau and $n + \ell$ rules are violated.
(ii) Aufbau and $n + \ell$ rules are violated.
(iii) Hund's rule is violated.

Q.59 ***Write the electronic configuration of the following:***
(i) Cu^+ (Z = 29) (ii) Br^- (Z = 35)

Ans. (i) Cu^- = $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^0, 3d^{10}$
No of protons = 29 No. of electrons = 28

(b) Br^- = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$
No. of protons = 35 No. of electrons = 36

Q.60 ***Write the electronic configuration of Na (Z = 11) and B (Z = 5). What are the values of n and ℓ for the last sub-shell of each element? What are their shapes?***

Karachi Board: 2010

Ans.

Electronic configuration of Na (Z = 11)
$1s^2, 2s^2, 2p^6, 3s^2$
Electronic configuration of B (Z = 5)
$1s^2, 2s^2, 2p^1$
Value of $(n + \ell)$ for 3s orbital of Na:
n = 3, ℓ = 0; $n + \ell$ = 3 spherical shape
Value of $(n + \ell)$ for 2p orbital of B:
n = 2, ℓ = 1; $n + \ell$ = 3 dumb bell shape

Q.56 ***Write electronic configuration of the following and also give the number of protons and electrons in each.***
(a) Cu (Z = 29) ***(b) Mg^{+2} (Z = 12)*** ***(c) Cl^- (Z = 17)***

Karachi Board: 2012, 2017
Karachi Board: 2019

Ans. ***(a)*** $Cu^{29} = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^9$
No of protons = 29 No. of electrons = 29

(b) $Mg^{+2} = 1s^2, 2s^2, 2p^6, 3s^0$
No. of protons = 12 No. of electrons = 10

(c) $Cl^- = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6$
No. of protons = 17 No. of electrons = 18

Q.57 ***What rules and principles are violated in the following electronic configuration?***
(a) $1s^2, 2s^3$, (b) $1s^2, 2p^2$, (c) $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^2, 4s^2$ *Karachi Board: 2011*

Ans.
(a) Aufbau rule
(b) Pauli's Exclusion Principle
(c) $(n + \ell)$ rule

Q.58 ***Which rule is violated in the following electronic configurations?***
(i) $2s^2 2p^6 3s^2$ (ii) $1s^2 2s^2 3s^2 2p^6$ (iii) $1s^2 2s^2 2p_x^2$ *Karachi Board: 2015*

Ans.
(i) Aufbau and $n + \ell$ rules are violated.
(ii) Aufbau and $n + \ell$ rules are violated.
(iii) Hund's rule is violated.

Q.59 ***Write the electronic configuration of the following:***
(i) Cu^+ (Z = 29) (ii) Br^- (Z = 35)

Ans. ***(i)*** $Cu^- = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^0, 3d^{10}$
No of protons = 29 No. of electrons = 28

(b) $Br^- = 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$
No. of protons = 35 No. of electrons = 36

Q.60 ***Write the electronic configuration of Na (Z = 11) and B (Z = 5). What are the values of n and ℓ for the last sub-shell of each element? What are their shapes?***

Karachi Board: 2010

Ans.

Electronic configuration of Na (Z = 11)
$1s^2, 2s^2, 2p^6, 3s^2$
Electronic configuration of B (Z = 5)
$1s^2, 2s^2, 2p^1$
Value of $(n + \ell)$ for 3s orbital of Na:
n = 3, $\ell = 0$; $n + \ell = 3$ spherical shape
Value of $(n + \ell)$ for 2p orbital of B:
n = 2, $\ell = 1$; $n + \ell = 3$ dumb bell shape

Q.61 ***Write the ground state and excited state configuration of carbon atom?***

Karachi Board: 2010

Ans.

C^6 $1s^2, 2s^2, 2p^2$ (ground sate)

or $1s^2$, $2s^2$, $2p_x1$, $2p_y1$, $2p_z$

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}$ $\boxed{\uparrow}$ $\boxed{}$

$1s^2, 2s^1, 2p^3$ (excited sate)

or $1s^2$, $2s^2$, $2p_x1$, $2p_y1$, $2p_z1$

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}$ $\boxed{\uparrow}$ $\boxed{\uparrow}$

Q.62 ***Which principle is violated in the following configuration?***
(i) $1s^2, 2s^1, 2p^5$ (ii) $1s^2, 2s^1, 2p^5$

Ans. ***Aufbau principle is violated:***
It should be (i) $1s^2, 2s^2, 2p^4$ (ii) $1s^2, 2s^2, 2p^5$

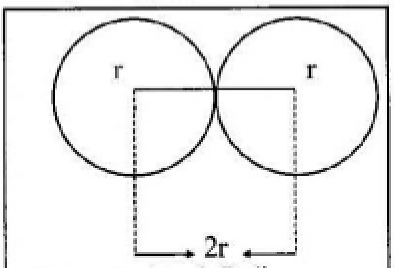
Q.63 ***Which rule or principle is violated in writing the following electronic configuration?***
$1s^2, 2s^2, 2p^2_x, 2p^1_y, 2p^0_z$

Karachi Board: 2005

Ans. ***Hand's rule violated it should be as follows:***
As per Hand's rule, it should be: $1s^2, 2s^2, 2p^1_x, 2p^1_y, 2p^1_z$

Q.64 ***Define atomic radius and why it can not be determined precisely?***

Ans. ***Atomic radii:*** "The average distance between the nucleus of a atom and its outer most electronic shell is called atomic radius." (Fig#3.18)



(Fig # 3.18) Atomic Radius

Reasons for non precise determination of atomic size:
The size of an atom cannot be determined precisely due to following reasons:

(i) There is no sharp boundary of an atom because the probability of finding electrons never become exactly zero even at large distances from the nucleus.

(ii) The electronic probability distribution is affected by the neighbouring atoms which have the size of an atom. It may change from one compound to another.

Due to the above reasons, the sizes of atoms are expressed in terns of atomic, ionic and covalent radii which depend on the type of the compound used for its measurement.

Measurement of atomic radius:
Atomic radius can be determined by measuring the distance between the centers of adjacent atoms with the help of x-rays or by spectroscopic measurement, for example, the distance between two atoms of carbon in diamond is 1.54 Å.

The radius of carbon atom would be as 0.77 Å ($1 Å = 10^{-8}$ cm)

The radius of $S = 2.06/2 = 1.03$ Å.

In case of heterogeneous diatomic molecules like AB, the band length is $(r_A + r_B)$ and if any one of the radii is known, the other can be calculated.

Q.65 Define and explain the term ionic radius.

Ans. Radius of an ion, which is spherical in shape is called ionic radius.

(i) When an electron is removed form a neutral atom. The positively charged ion is obtained. This is called cation.

$$\underset{\text{atom}}{M} \longrightarrow \underset{\text{cation}}{M^{+}} - e^{-}$$

(ii) When an atom gains one electron, negatively charged ion is produced. It is called anion.

$$\underset{\text{atom}}{M + e^{-}} \longrightarrow \underset{\text{anion}}{M^{-}}$$

Isoelectronic ions: The isoelectronic ions are the ions having same electronic configuration, for example, Na^{+}, Mg^{++} and Al^{+++} ($1s^2, 2s^2, 2p^6$). In these ions the radii decrease with the increase in nuclear charge i.e.

ION	Na^{+}	Mg^{++}	Al^{+++}
Atomic Number	11	12	13
Ionic radius	0.95 Å	0.65 Å	0.50 Å

With the increase in atomic number, nuclear charge is increasing while valence shell remains the same. Hence large nuclear charges attract valence shell. As a result radii decrease with the increase of nuclear charge.

Q.66 Why a cation is smaller in size and an anion is larger in size from its atoms?

Ans. *(i)* A cation is formed when an atom loses electron the cationic radius decreases due to the increase in the effective nuclear charge on the ion. Thus the nucleus attracts lesser number of electrons with greater force as a result, the size of cation decreases.

e.g.	Na	$\longrightarrow$	$Na^{+} + e^{-}$
	$11e^{-}$		$10e^{-}$
	11p		11p
	1.57 Å		0.95 Å

(ii) The ionic radius of an anion is greater than the atomic radius of its parent atom. It is due to an excess of negative charge results in greater electron repulsion.

e.g.	$Cl + e^{-}$	$\longrightarrow$	Cl^{-}
	$17e^{-}$		$18e^{-}$
	17p		17p
	0.99 Å		1.81 Å

Q.67 Give definition, unit and examples of ionization energy.

Ans. ***Ionization energy:*** *"The energy required for the removal of the least tightly bound electron from its isolated atom in the gaseous state is called ionization energy or ionization potential of the element".*

Unit of I.E: It in expressed in $KJ\ mol^{-1}$

Example:

$M_{(gas)} \longrightarrow M^{+}_{(gas)} - e^{-}$ ΔE in the first ionization energy

$M^{+}_{(gas)} \longrightarrow M^{++}_{(gas)} + e^{-}$ ΔE in the second ionization energy

$M^{++}_{(gas)} \longrightarrow M^{+++}_{(gas)} + e^{-}$ ΔE in the third ionization energy

The ionization potential is determined from spectroscopic methods or by measurement of current passing through a discharge tube. The ionization potential depends upon the distance of the electron form the nucleus. Farther the electron is from the nucleus, more easily it is removed. When electron is removed, the remaining electrons on tightly binds with the nucleus. It is therefore, becomes, more and more difficult to remove the second and then the third electron. It is easier to remove an electron from 'f' shell than a 'd' shell and removal from a 'd' shell is easier than a 'p' shell due to their large distance.

Q.68 ***Write down the definition, unit and examples of electron affinity.***

Ans. ***Definition:*** The amount of energy released when an electron is absorbed in the lowest energy state of an isolated atom in the gaseous state to form an anion is called electron affinity of the element.

Unit: It is expressed in $KJ\ mol^{-}$.

$$Cl + e^{-} \longrightarrow Cl^{-} \qquad \Delta H = -348\ KJ/mol$$

When a second electron is absorbed by a uni-negative ion the incoming electron is repelled by the negative ion and energy is absorbed in this process.

e.g.

$$O^{(g)} + e^{-} \longrightarrow O^{-(g)} \qquad \Delta H = -141\ KJ/mol^{-}$$

$$O^{(g)} + e^{-} \longrightarrow O^{2-(g)} \qquad \Delta H = +844\ KJ/mol^{-}$$

Q.69 ***Describe the factors on which electron affinity depends.***

Ans. Factors on which electron affinity depends:

(i) ***Nuclear charge:*** Greater the nuclear charge, greater would be the electron affinity.

(ii) ***Atomic Size:*** Smaller the atomic size, greater would be the electron affinity.

(iii) ***Shielding affect:*** Greater the shielding effect, smaller would be the electron affinity.

(iv) ***Electronic Configuration:*** The atoms with stable configuration have no tendency to gain an electron. An atom has stable configuration if it has:

(a) inert gas configuration
(b) fully filled orbitals
(c) half filled orbitals

Q.70 ***Write short notes on electronegativity.*** ***Karachi Board: 2013***

Ans. ***Definition:*** "The power of an atom to attract the shared pair of electrons of a chemical bond towards itself is called electronegativity".

Measurement of Electronegativity values:

Pauling assigned number on an arbitrary scales from 0 to 4 as a measure of electronegativities of atoms.

Determination of Electronegativity values:

Electronegativity values are determined from bond energy data. The covalent bond formed between dissimilar atoms is polar i.e. the one atom is partially negative and the other is partially positive.

e.g: Fluorine = 4.0 E.N (Highest)
Francium = 0.7 E.N (Least)

Q.71 ***Name the factors an which electronegativity depends.***

Ans. Factors on which electronegativity depends, are:

- *(i)* nuclear charge
- *(ii)* atomic size.
- *(iii)* electron affinity
- *(iv)* ionization energy

Q.72 ***Write the electronic configuration of Na (Z=11) and B (Z=5). What are the values of n and ℓ for the last sub-shell of each element? What are their shapes?***

Ans. **Solution:**

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

For last sub-shell, the value of n = 3

For last sub-shell, the value of ℓ = 0

Shape = Spherical

$^5B = 1s^2\ 2s^2\ 2p^1$

For last sub-shell, the value of n = 2

For last sub-shell, the value of ℓ = 1

Shape = dumb-bell

Q.73 ***Explain sp^3 hybridization in carbon in detail.***

Ans. ***sp^3 Hybridization:***

A carbon atom contains six electrons which occupy the following electron configuration: $(1s)^2\ (2s)^2\ (2p)^2$. In the ground state there are two unpaired electrons in the outer shell, so that one could assume the ability to bind only two additional molecules. In contrast, a binding ability for four electrons is detected in experiments. The reason is the small energy difference between the 2s and the 2p-state, so that it is easily possible to excite one electron from the 2s-state into the 2p-state. In the presence of a external perturbation, such as a nearby by hydrogen, the energy difference is overcome. This result in a mixed state formed cut of on s-orbital and three p-orbitals, namely p_x, p_y and p_z, is produced. Four new hybrid orbitals are formed. The directions of the orbitals and also the center of mass are datermned by the specific contributions of the p-orbitals and the s-orbital. A combination of the hybrid orbitals produces a tetrahedral assembly with the center of masses in the ecrners (see fig. 1.). The characteristic angle between the hybrid orbitals in sp^3-configuration is 109.5 degree.

Assemblying many different sp^3-hybridised carbon atoms to one crystal, one achieves the typical diamond structure (see fig. 2).

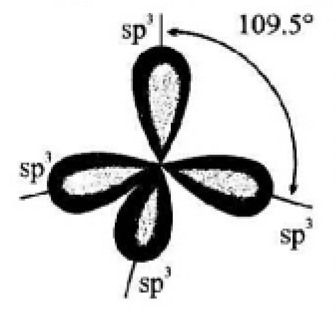


Fig: 1. (1) sp^3-hybrid orbital

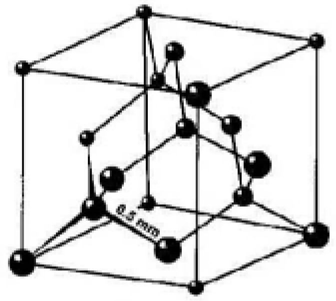


Fig. 2: (2, 3)
Crystal structure

Due to the three dimensional sp^3-structure the binding strength between neighboring carbon atoms is equal for each atom and very strong. Therefore, diamond is one of the hardest materials known, it is used as a cutting tool. The corresponding band structure reveals a large band gap of 5 eV, which corresponds to an in sulater. The electrical conductivity is very low. Furthermore, the transparent appearance of diamond corresponds to the fact that electrons cannot be excited out of the valence hand into the conduction band with a wavelength in the optical range.

Q.74 ***Write the electronic configuration of the following:***
(i) Mo (Z = 42) (ii) Ag (Z = 47) (iii) Cr (Z = 24) (iv) Rb^{-1} (Z = 37)

Karachi Board: 2017

Ans.

(i) Molybdenum **Symbol:** Mo **Atomic number :** 42
$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^5 5s^1$

(ii) Silver **Symbol:** Ag **Atomic number :** 47
$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^1$

Karachi Board: 2018

(iii) Chromium **Symbol:** Cr **Atomic number :** 24
$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^5, 4s^1$

(iv) Rubidium **Symbol:** Rb **Atomic number :** 37
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$

Differences

Q.01 ***Differentiate between the following:***
(i) Continuous and line spectrum (ii) Natural and artificial Radioactivity

Karachi Board: 2008-2013-2016-2019

S. No.	*Continuous Spectrum*	*Line Spectrum*
1.	This spectrum consists of bands of colours.	This spectrum consists of lines.
2.	There is no definite boundary between colours.	Lines are separated from each other.
3.	It consists of more than one colour.	The lines are of only one colour.
4.	It is produced by white or polychromatic light.	It is produced by monochromatic light.
5.	It is due to separation of components of light which passing through prism.	It is produced due to transition of electron from higher to lower energy level.
6.	It is only of one type.	It is of two types i.e emission and absorption spectrum.

Karachi Board: 2004

S. No.	*Natural Radioactivity*	*Artificial radioactivity*
1.	The spontaneous emissions of radioactive substance is called natural radioactivity.	When a non-radioactive element is made radioactive by bombardment of certain subatomic particles is called artificial radioactivity.
2.	It is a natural phenomena.	It is an artificial phenomena.
3.	It is exhibited by elements with atomic number greater than 82.	It can be produced in any element by bombardment with particles.
4.	Radiations are α, β and γ types.	New particles can also be produced.
5.	Example: ${}_{92}U^{238} \longrightarrow {}_{2}H^{4} + {}_{90}Th^{234}$	Example: ${}_{4}Be^{9} + {}_{2}He^{4} \longrightarrow {}_{6}C^{12} + {}_{0}n^{1}$

Q.02 ***Differentiate between the following:***
(i) Orbit and orbital (ii) Electronegativity and Electron Affinity

Karachi Board: 2011, 2014, 2015

S. No.	*Orbit*	*Orbital*
1.	It is well defined circular path around nucleus in which electrons revolve.	It is the three dimension space round the nucleus where the probability of finding the electrons is maximum.
2.	It is circular in shape (K, L, M, N).	They are of four types and different shapes (s, p, d, f).
3.	It is against Heisenberg un-certainity principle.	It is in accordance to Heisenberg un-certainity principle.
4.	Maximum number of electrons in an orbit is given by $2n^2$.	Maximum numbers of electrons is only 2. s^2, p_x^2, p_y^2, p_z^2

Karachi Board: 2008

S. No.	*Electronegativity*	*Electron Affinity*
1.	It is the power of an atom in a molecule to attract shared pair of covalent electrons to itself.	It is the power of an isolated atom to attract a single free electron to itself.
2.	It has no unit just arbitrary number 0-4.	Its unit is KJ / mole.
3.	It is a periodic property.	It is not a periodic property.
4.	It is a comparative value with 'F' as electronegativity atom as standard.	Its values are absolute.

Q.03 ***Differentiate between the following:***
(i) Alpha and Beta rays (ii) Balmer and Lyman Series

Karachi Board: 2007

S. No.	Alpha (α) rays	Beta (β) rays
1.	α Rays are fast moving helium neclei.	β Rays are fast moving electrons.
2.	Their velocity is equal to 1 / 10^{th} of light.	Their velocity almost equal to velocity of light.
3.	They are good ionizer of gases.	They are weak ionizer of gases.
4.	They can penetrate 1-2 cm in air. Penetrating power is very low.	They can penetrate 1-2 cm in air and can pass through small thickeness of matter.

Karachi Board: 2012

S. No.	Balmer Series	Lyman Series
1.	This spectrum is in visible region, $n_1 = 2$	This spectrum is in ultraviolet region, $n_1 = 1$
2.	Its energy is less than Lyman Series.	Its energy is greater than Balmer Series.
3.	Its wavelength is in between 3500-7000 Å.	Its wavelength is below 3500 Å.

Q.04 ***Differentiate between the following:***
(i) x-rays and gamma rays (ii) Principal and Azimuthal Quantum Numbers

S. No.	X-Rays	γ (gamma) Rays
1.	These are artificial rays.	These are natural rays.
2.	These are produced when fast moving electrons collided with a metallic plate.	These are produced naturally by nucleus disintegration of a radioactive substance.
3.	Extra nuclear in origin.	They have nuclear origin.
4.	They are less energetic & less penetrative.	They are more penetrating and energetic.

Karachi Board: 2005, 2017

S. No.	Principal Quantum Number (n)	Azimuthal Quantum Number (ℓ)
1.	It determines the distance between the electron of the nucleus.	It tells us about shape of orbitals.
2.	Number of shells are represented by: 1 = K, 2 = L, 3 = M, 4 = N,…..	Shape of orbitals are represented by: 0 (s), 1 (p), 2 (d), 3 (f)
3.	Mathematical Values: n = 1, 2, 3, ...	Mathematical value: ℓ = 0, 1, 2, 3 ....

Scientific Reasons

Q.01 ***Ionization potential of nitrogen is greater than that of oxygen.*** *Karachi Board: 2013*

Ans. Nitrogen has half filled valence shell electronic configuration, while oxygen has partial filled valence shell electronic configuration. Since half filled electronic configuration is more stable than partial filled configuration therefore I.P. of nitrogen is greater than oxygen.

Q.02 ***Na^+ ion is smaller in size than Na atom.*** *Karachi Board: 2012-2019*

Ans. A cation is formed by the removal of one or more electrons from the atom. Thus a cation has lesser number of electrons. This increased effective nuclear charge pull the electrons cloud inward nearer to the nucleus and thus makes the cation smaller in size than its parent neutral atom.

Q.03 ***Cl^- ion is bigger in size than Cl atom.***

Ans. As anion is formed by the addition of one or more electrons to the atom. Thus an anion has greater number of electrons. The magnitude of the screening constant decreases the effective nuclear charge. The decreased effective nuclear charge decreases the inward pull of nucleus and thus makes the anion bigger in size than its parent neutral atom.

Q.04 ***Electron affinity of noble gases is zero.*** *Karachi Board: 2016*

Ans. All noble gases have completely filled shell. The complete filled valence orbitals are stable and therefore have no tendency to gain or lose electrons. Thus noble gases have zero electron affinities.

Q.05 ***Mg^{+2} ion is smaller than Mg atom.*** *Karachi Board: 2014*

Ans. Mg^{+2} ion is produced when Mg atom loses two electrons. Decrease in electrons from outer shell makes nuclear charge more powerful in Mg^{+2} ion than in Mg atom. Because of increase in number of proton and decrease in number of electron in Mg^{+2} ion, electrons come closer to nucleus which makes the size of Mg^{+2} become smaller than size of Mg atom.

Q.06 ***Elements give line spectrum.*** *Karachi Board: 2017*

Ans. Elements give only line spectrum because the sample of the same element emits the same wavelength of radiation. Due to emission of radiation of same wavelength line spectrum is formed.

Solution of Textl ok Progress Test - 3

Q.01 ***What is the experimental evidence fr the presence of small nucleus containing most of the mass and all the positive charge in the atom ?***

Ans. See Q 22 on page 108.

Q.02 ***What information about the structure of atom is obtained from the experiment on the passage of electricity through gasses under low pressure?***

Ans. See Q 4, 5, 7 & 9 on page 101 to 103.

Q.03 ***Criticize the following statement "An electron moves about the nucleus of an atom in a manner analogous to the movement of a planet about the sun."***

Ans. See Q 24 on page 109.

Q.04 ***What information about the electron in the atom is obtained from the fact that the emission spectra of the elements are frequently discontinuous?***

Ans. See Q 21 on page 107.

Q.05 ***Explain quantized energy states in terms of Bohr's orbits and electron transition between the orbits.***

Ans. See Q 27 on page 109.

Q.06 ***What principles and rules are followed in the electronic configuration of atom?***

Ans. See Q 51 on page 112.

Q.07 ***Explain the origin of x-rays and relationship between their wavelength and nuclear change in the atoms from which they originate .***

Ans. See Q 25 on page 109.

Q.08 ***(a) State the restriction placed on our ability to know the position and momentum of an electron?***

Ans. See Q 43 on page 119.

(b) Explain the significance of Pauli's Exclusion principle in relation to the electronic structure of atoms.

Ans. For answer see Q 50 on page 122.

(c) ***Write down the electronic con ation for the ground states of each of the following:***
(i) Na (Z=11) (ii) Ca (Z=2 (iii) Sc (Z=21) (iv) Mg^{2+} (Z=12) (v) Cl^- (Z=17)

Ans. *(i)* Na^{11} = $1s^2, 2s^2, 2p^6, 3s^1$
(ii) Ca^{12} = $1s^2, 2s^2, 2p^6, 3s^2$
(iii) Sc^{21} = $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^1$
(iv) & (v) See Q. 56 on page 124.

Q.09 ***Explain the distinction between the following:***
(a) α & β particles (b) Orbit & orbital
(c) Ground and excited state (d) Continuous and discontinuous spectrums

Ans.

(a)	α & β rays.	See "Differences"	Q 3 on page 131.
(b)	Orbit & orbital	See "Differences"	Q 2 on page 130.
(c)	Ground and excited state	The first paragraph of	Q 35 on page 115.
(d)	Continuous and discontinuous spectrums	See "Differences"	Q. 1 on page 129.

Q.10 ***Describe contribution made by the following individuals to ward the structure of atom***
(a) Rutherford (b) Bohr (c) Mosely.

Ans. *(a)* ***Rutherford:*** Evidence for nucleus and arrangement of particles. For details see Q 22 on page 108.

(b) ***Bohr:*** Bohr assumed that all the transition that electrons make between two orbits, yield a single unique spectral line. For details see Q 27 on page 109.

(c) ***Mosely:*** He studied the different wavelengths of x-rays produced from anodes of different metals. He noticed that the wavelengths of the x-rays emitted decreased regularly with increasing atomic mass. One element can be distinguished from another on the basis of nuclear charge i.e. atomic number and not atomic mass.

Q.11 ***Give the number of protons neutrons and electrons in each of the following:***
(a) $_{39}^{19}K$ (b) $_{16}^{8}O$ (c) $_{7}^{3+}Li$ (d) $_{19}^{9-}F$ (e) $_{23}^{11}Na$

Ans. See Q 17 on page 106 & Q 56 on page 124.

Q.12 ***What do you understand by the terms:***
(a) Atomic radius (b) Ionization potential (c) Electron affinity (d) Electronegativity

Ans. *(a)* See Q 64 on page 125.
(b) See Q 67 on page 126.
(c) See Q 68 on page 127.
(d) See Q 70 on page 127.

Q.13 ***List the four quantum number that define the energy state of the electrons in a hydrogen atom along with their possible values and explain their significance in terms of orbitals.***

Ans. See Q 47 on page 120 for Quantum number.

$_1H^1 = 1s^1$ $\quad$ $n = 1$ (K shell) $\quad$ $\ell = 0$ (s orbital)

$m = 0$ $\quad$ $s = \pm \frac{1}{2}$

Shape of s orbital = spherical

Q.14 ***(a) The ionization potential of lithium and potassium are 5.4 and 4.3 ev. What do predict for the I.P of Na ?***

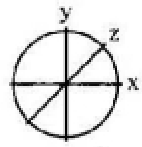
Ans. The atomic size of sodium is bigger than lithium but smaller than potassium. Shielding effect of sodium is greater than lithium but less than potassium. The attraction of nucleus in sodium is less than lithium but more than potassium. On the account of above facts, the ionization potential of sodium should be in between Li and K.

The approximate I.P. of Na $= \dfrac{5.4 + 4.3}{2} = 4.85$ ev

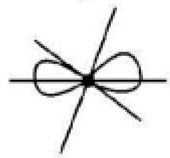
(b) What is the shape of orbital for which $\ell = 0$? Those for which $\ell = 1$?

Ans.

When $\ell = 0$ means 's' orbital and it is spherical in shape.



When $\ell = 1$ means 'p' orbital and it is dumb - bell in shape.



Q.15 ***(a) How is emission of radiation from an atom explained by Bohr's model ?***

Ans. See Q 35 on page 115.

(b) Why was the nucleus of hydrogen atom was assumed to be a fundamental particle?

Ans. The nucleus of hydrogen atom contains only one proton. Proton means first particle. Hydrogen is the smallest atom, therefore hydrogen atom was assumed to be a fundamental particle.