1 ATOMIC STRUCTURE - III

1. SUB ATOMIC PARTICLES & ATOMIC MODELS

	♦	SUB ATOMIC PARTICLES
	♦	ATOMIC MODELS
	♦	BOHR'S ATOMIC MODEL
	♦	ELECTROMAGNETIC RADIATIONS
	♦	PLANCK'S QUANTUM THEORY

INTRODUCTION

- We have learnt that atoms are the fundamental building blocks of matter.
- The Indian philosopher **Maharshi Kanad** came forward with the idea that matter is not continuous and made up of tiny particles named 'paramanus' (atoms). Kanad further said that two or more paramanus combine to form bigger particles called 'anus'(molecules).
- Greek philosopher **Democritus** proposed that matter is made up of extremely small particles 'atoms', meaning indivisible.
- In the **1**800s, chemist **John Dalton** considered previous theories. He proposed that atoms of a particular element were the same size and weight and unlike any other type of atom. He also believed that different atoms could combine in fixed ratio to make different chemical compounds.
- In **1911**, the experiment that contributed most to our knowledge of the structure of the atom was done by **Ernest Rutherford.** From his alpha ray scattering (gold foil) experiment, Rutherford concluded that the atoms were mostly empty space, however, small central region of the atom denser part which called the nucleus.

STRUCTURE OF AN ATOM:

An atom contains three sub-atomic particles. At the centre is a nucleus, which contains protons, positively charged particles, and neutrons, which are the particles with no charge. Surrounding the nucleus are moving electrons, which are negatively charged particles.

An exception to this structure is the simplest atom, hydrogen, which comprises a nucleus of one proton and no neutrons, surrounded by one electron.

- Electron: a negatively charged sub-atomic particle $\left(-1.602\times10^{-19}\,\mathrm{C}\right)$

- Proton: a positively charged sub-atomic particle $\left(+1.602\times10^{-19}\text{C}\right)$
- Neutron: a sub-atomic particle with no charge
- Protons and neutrons have about the same mass $\left(1.67\times10^{-24}\text{g}\right)$. Although smaller and of very little mass $\left(9.11\times10^{-31}\text{Kg}\right)$, electrons occupy the bulk of the space with their movement around the nucleus.

ATOMIC MODELS:

- The first model of atom was proposed by **J.J. Thomson.** This atomic model is called apple pie model or plum pudding model or watermelon model of atom.
- **Rutherford** proposed a model after his alpha ray scattering experiment. He compared the structure of atom with solar system. This model is also called solar system model.

DEVELOPMENTS LEADING TO THE BOHR'S MODEL OF ATOM :

Two important developments played a major role in the formulation of Bohr's model of atom. These were:

1. Dual character of electromagnetic radiations which means that radiations possess both wave like and particle like properties.
2. Quantization of electronic energy.

ELECTROMAGNETIC RADIATIONS:

The propagation of light occurs like wave which is just like the flow of water waves. The light waves are electromagnetic waves which are formed by electric and magnetic fields. The electromagnetic spectrum consists of the waves of all ranges of wavelength or frequency. It consists of the harmful rays like IR rays and also some useful rays like radio waves for communication, **X-**rays for medical purpose, microwaves for cooking etc. These electromagnetic radiations do not need any medium for their propagation.

ALL ELECTROMAGNETIC RADIATIONS ARE CHARACTERIZED BY THE FOLLOWING SIX PROPERTIES:

WAVELENGTH: The distance between two nearest crests or nearest troughs is called the wavelength. It is denoted by λ(Lamda) and is measured in terms of centimetre (cm), angstrom (A°), micrometre (μm) or nanometre (nm).

[1 A° = 10^{-10} m or 10^{-8} cm]

FREQUENCY: It is defined as the number of waves which pass through a point in one second. It is denoted by the symbol ϑ (nu) and is measured in terms of cycles (or waves) per second (cps) or hertz (Hz).

VELOCITY: It is defined as the distance covered in one second by the wave. It is denoted by the letter V. All electromagnetic waves travel with the same velocity, i.e. 3×10^{10} cm/sec.

WAVE NUMBER: This is the reciprocal of wavelength, i.e., the number of wavelengths per centimeter. It is denoted by the symbol $\bar{\vartheta}$ (nu bar). $\bar{\vartheta} = \frac{1}{\lambda}$ It is expressed in cm^{-1} or m^{-1}.

AMPLITUDE: It is defined as the height of the crest or depth of the trough of a wave. It is denoted by the letter 'a'. It determines the intensity of the radiation.

TIME PERIOD: Time taken by the wave for one complete cycle or vibration is called time period. It is denoted by T. $T = \frac{1}{\vartheta}$. **Unit**: Seconds per cycle.

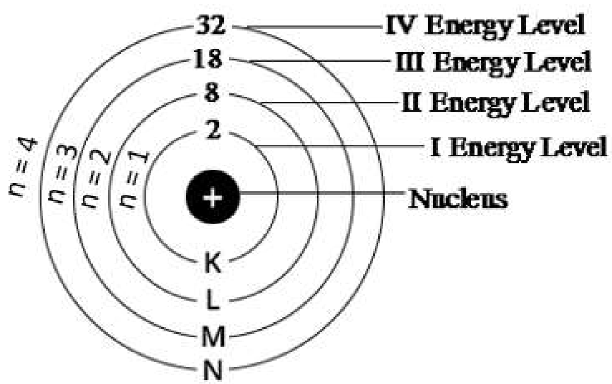
Note: Some relations between characteristics of electromagnetic radiations can be taken as following for calculation purposes:

$$\vartheta = \frac{c}{\lambda}; \bar{\vartheta} = \frac{1}{\lambda}; T = \frac{1}{\vartheta}$$

BOHR'S MODEL OF ATOM:

Bohr proposed his atomic model based on the quantum theory of radiation. According to Bohr's model,

- Electrons move around the nucleus in specified circular paths called orbits or shells or energy levels.
- Each orbit or shell is associated with a definite amount of energy. Hence these are also called energy levels and are designated as K, L, M, N shells respectively.



- The energy associated with a certain energy level increases with the increase of its distance from the nucleus. Hence if the energies associated with the K, L, M, N shells are E_1, E_2, E_3,.... respectively, then $E_1 < E_2 < E_3$.....etc.
- As long as the electron revolves in a particular orbit, the electron does not lose its energy. Therefore, these orbits are called stationary orbits and the electrons are said to be in stationary energy states.
- An electron jumps from a lower energy level to a higher energy level, by absorbing energy. It jumps from a higher energy level to a lower energy level, by emitting energy in the form of electromagnetic radiation.

The energy emitted or absorbed $\left(\Delta E\right)$ is given by Planck's equation i.e, $\Delta E = h\vartheta$

For example, If E_1 and E_2 are the energies of first and second orbits, the difference in energy is equal to $h\vartheta$. $E_2 - E_1 = h\vartheta$

- The electron can revolve only in an orbit in which the angular momentum of the electron (mvr) is a whole number multiple of $h/2\pi$. This is known as the principle of quantization of angular momentum.

Hence we can write angular momentum of the electron as, $mvr = \frac{nh}{2\pi}$,

Where n is an integer (n =1,2,3,4,......) and is called principal quantum number. m = mass of the electron, v = velocity of an electron in its orbit and r = distance of the electron from the nucleus.

By applying the concept of quantization of energy, Bohr calculated the radius and energy electron in the n^{th} orbit of any single electron species like Hydrogen atom (Z = 1) as follows:

Radius of the n^{th} orbit $r_n = \frac{0.529 \times n^2}{Z} A^o$

Energy of the electron in the n^{th} orbit $E_n = \frac{-13.6}{n^2} \times Z^2 eV$ per atom/ion

Where 'n' is the number of the orbit and ‘Z' is atomic number of the single electron species.

LIMITATIONS OF BOHR'S ATOMIC MODEL:

- Bohr could not explain the spectra of the multi electronic species satisfactorily.
- Bohr's model couldn't give a satisfactory justification for the assumption that electrons can revolve in those orbits where their angular momentum (mvr) is a whole number multiple of $h/2\pi$. i.e., he could not justify the quantization of angular momentum of the electron.
- Bohr assumed that an electron of an atom is located at a definite distance from the nucleus and revolves around the nucleus with a definite velocity. It is against to Heisenberg's uncertainty principle which proved that 'it is impossible to determine simultaneously and accurately the exact position and the momentum of the moving particle like electron'.
- The atomic spectral lines are found to split into a number of closely packed lines in the presence of a magnetic field and an electric field. These effects are called as **zeeman effect** and **stark effect** respectively Bohr failed to explain these effects.
- Each spectral line in the atomic spectrum of hydrogen consists of still finer lines when it is observed by using the instrument with high resolving power. Bohr's model failed to explain this observation.

BLACK BODY RADIATION:

When certain energy falls on the surface of a body, a part of energy is absorbed, a part of it is reflected and the remaining is transmitted. All the incident radiant energy is not absorbed completely by the body because ordinary bodies are not perfect absorbers of radiant energy. However, an ideal body is expected to absorb completely the radiant energy falling on it. Such a body is known as a black body.

PLANCK'S QUANTUM THEORY:

In order to explain black body radiation and photo electric effect, Max Planck in 1901 presented a new theory which is known as Planck's quantum theory of radiation.

ACCORDING TO THIS THEORY:

- Substances absorb or emit light discontinuously in the form of small packets.
- The smallest packet of energy is called quantum
- The radiation is propagated in the form of waves. The energy of a quantum is directly proportional to the frequency of the radiation. $E \propto \vartheta$.
- The energy of a quantum is $E = h\vartheta = \frac{hc}{\lambda} = hc\bar{\vartheta}$

Where h is a constant known as Planck's constant. Its numerical value is 6.625×10^{-27} erg-sec or 6.625×10^{-34} Joule-sec.

E = Energy in ergs or Jouls, c = Velocity of light = 3×10^{10} cm/sec = 3×10^{8} m/sec

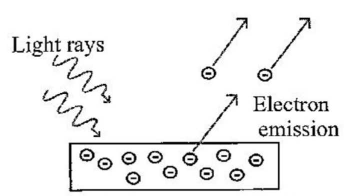
ϑ = Frequency of radiation, λ= wavelength, $\bar{\vartheta}$ = wave number.

A body can absorb or emit energy in whole numbers of quantum, i.e.

$E = n(h\vartheta)$

PHOTOELECTRIC EFFECT:

In 1887, **H. Hertz** performed a very interesting experiment in which electrons (or electric current) were ejected when certain metals (for example potassium, rubidium, cesium etc) were exposed to a beam of light as shown in fig. This phenomenon is called photoelectric effect, and these ejected electrons are called photo electrons.



Ejection of electrons from metal surface

1. SUB ATOMIC PARTICLES AND ATOMIC MODELS

WORK SHEET LEVEL–I MAINS CORNER

SINGLE CORRECT ANSWER TYPE QUESTIONS

SUB ATOMIC PARTICLES

1. The idea of tiniest unit of matter (anu and paramanu) was propounded by:
   1) Democritus 2) John Dalton
   3) William Crookes 4) Maharshi Kanada
2. Cathode rays are travel in:
   1) Straight line 2) Cross line 3) Curve line 4) Circle
3. The charge of electron is:
   1) 1.602×10^{-19}C 2)1.67×10^{-24}Kg 3) 9.11×10^{-28}Kg 4) -1.602×10^{-19}C
4. Anode rays are:
   1) Negative 2) Positive 3) Neutral 4) All of these
5. The e/m ratio is zero for:
   1) Electron 2) Proton 3) Neutron 4) All

ATOMIC MODELS

6. J.J. Thomson’s model of an atom commonly known as:
   1) Watermelon model 2) Planetary model
   3) Solar model 4) Bohr model
7. Thomson atomic model can explain only:
   1) Existence of Nucleus 2) Electrical neutrality
   3) Orbital concept 4) All the these
8. Alpha particles are:
   1) Positive 2) Negative 3) Neutral 4) All of these
9. Rutherford’s atomic model of an atom is also called as:
   1) Planetary model2) Solar model 3) Nuclear model 4) All of these

BOHR’S ATOMIC MODEL

10. Bohr explained the stability of an atom based on:
    1) stationary orbits 2) quantization of angular momentum
    3) Planck’s quantum theory 4) All of the above
11. Bohr model can explain:
    1) the spectrum of hydrogen atom only
    2) spectrum of an atom or ion containing one electron only
    3) the spectrum of hydrogen molecule
    4) the solar spectrum
12. According to Bohr’s atomic model mvr = _____.
    1) $\frac{nh}{2\pi}$ 2) $\frac{nh}{\pi}$ 3) $\frac{h}{4\pi}$ 4) $\frac{h}{2n\pi}$

ELECTRO MAGNETIC RADIATIONS

13. Which is not electromagnetic radiation?
    1) Infrared rays 2) X – rays 3) Cathode rays 4) γ - rays
14. All types of electromagnetic radiation possess same:
    1) Energy 2) Velocity 3) Frequency 4) Wavelength
15. The product of which of the following is equal to the velocity of light?
    1) Wavelength and wave number 2) Wavelength and frequency
    3) Frequency and wave number 4) Wavelength and amplitude

PLANCK'S QUNATUM THEORY

16. Quantum theory of radiation was proposed by:
    1) Heisenberg 2) Bohr 3) Planck 4) Einstein
17. An ideal body which is a perfect absorber and perfect emitter of radiation is called:
    1) Photoelectric effect 2) Black body radiation
    3) White body radiation 4) None
18. The electrons are ejected from surface of certain metals when light falls on it. This is known as:
    1) Stark effect 2) Zeeman effect
    3) Photoelectric effect 4) All of these

LEVEL–II

SUB ATOMIC PARTICLES

19. Which of the following are the characteristics of anode rays?
    i) The e/m ratio of positive rays depends upon the nature of the gas.
    ii) Anode rays are deflected both in electric and magnetic fields.
    iii) Anode rays cause mechanical motion.
    iv) Anode rays travel in Straight lines.
    v) Anode rays constitute positively charged particles.
    1) i, ii & iv 2) i, ii, iii, iv & v 3) i, ii & v 4) i, ii & iii
20. The mass of the electron is:
    1) 1.76×10^{-23} Kg 2) 1.67×10^{-24} Kg 3) 9.11×10^{-28} Kg 4) 9.11×10^{-31} Kg
21. The mass of protons is:
    1) 4.67×10^{-24}g 2) 3.67×10^{-24}g 3) 2.67×10^{-24}g 4) 1.67×10^{-24}g
22. Identify the correct statement:
    1) The electron has negligible mass.
    2) Neutron was discovered by James Chadwick by bombarding beryllium with particle.
    3) Neutron is the heaviest sub-atomic particle in the atom.
    4) All of these

ATOMIC MODELS

23. Nucleus of an atom consists of:
1) Proton and electron 2) Electron and neutron
3) Protons and Neutrons 4) Electrons protons neutrons

24. Rutherford identified the existence of protons at the centre of the atom in his experiment by:
1) The deflection of alpha particle 2) The absorption of alpha rays
3) The retention of alpha rays 4) None

25. Rutherford's experiment, which established the nuclear model of the atom, used a beam of:
1) β - particles, which impinged on a metal foil and got absorbed
2) γ - rays, which impinged on a metal foil and ejected electrons
3) Helium atoms, which impinged on a metal foil and got scattered
4) Helium nuclei, which impinged on a metal foil and got scattered

BOHR'S ATOMIC MODEL

26. According to Bohr's theory, angular momentum of an electron in fourth orbit is:
1) $\frac{h}{2\pi}$ 2) $\frac{h}{4\pi}$ 3) $\frac{2h}{\pi}$ 4) $\frac{4h}{\pi}$

27. Bohr's orbits are called stationary states because:
1) Electrons in them are stationary
2) There orbits have fixed radii
3) The electrons in them have fixed energy
4) The protons remain in the nuclei and are stationary.

ELECTRO MAGNETIC RADIATIONS

28. If the wavelength of green light is about 5000 A°, then the frequency of its wave is:
1) 16×10^{12} sec^{-1} 2) 16×10^{14} sec^{-1} 3) 6×10^{14} sec^{-1} 4) 6×10^{12} sec^{-1}

29. The frequency of light having wavelength 500 nm is:
1) 5×10^{15} Hz 2) 5×10^{10} MHz 3) 2×10^{-15} Hz 4) 6×10^{14} Hz

30. Light emitted from a sodium lamp has a wavelength (λ) of 580 nm. Calculate the wave number of the light.
1) $1.72\times10^{6}m^{-1}$ 2) $1.72\times10^{12}cm$
3) $1.72\times10^{-12}cm$ 4) $3.72\times10^{6}m^{-1}$

PLANCK'S QUNATUM THEORY

31. A 100 watt bulb emits mono chromatic light of wave length 400 nm. The no. of photons emitted per second by the bulb is:
1) $4.012\times10^{20}S^{-1}$ 2) $3.012\times10^{20}S^{-1}$ 3) $5.012\times10^{20}S^{-1}$ 4) $2.012\times10^{20}S^{-1}$

32. Calculate the energy of photons of radiation whose wavelength is 5000 A^0?

1) $3.97\times10^{-19}J$ 2) $3.97\times10^{19}J$ 3) $2.97\times10^{-19}J$ 4) $2.97\times10^{-19}J$

LEVEL-III ADVANCED CORNER

SINGLE CORRECT ANSWER TYPE QUESTIONS

33. The charge of electron in esu is:

1) $+4.8\times10^{-10}$ 2) -4.8×10^{-10} 3) $+1.6\times10^{-19}$ 4) -1.6×10^{-19}

34. An element occurs in two isotopic forms with atomic masses 10 and 11. What is the percentage of two isotopes in the sample having atomic mass 10.8?

1) 20,80 2) 50,50 3) 60,40 4) 25,75

35. Bohr's model of atom explains:

1) Zeeman effect 2) Photo electric effect
3) Stark effect 4) None of these

36. Calculate the wavelength, frequency and wave number of a light whose period is 2.0×10^{-10}s.

Wavelength	Frequency	Wave number
1) 3×10^{6} mm	$2\times10^{-3}\ s^{-1}$	$3.3\times10^{3}\ m^{-1}$
2) 6×10^{-2} m	$5\times10^{9}\ s^{-1}$	$16.66\ m^{-1}$
3) 6×10^{-2} m	$2\times10^{9}\ s^{-1}$	$16.66\ m^{-1}$

4) None of the above

37. The energy of photon of reddish light having wavelength 660 nm is ($h = 6.6\times10^{-34}$ J s).

1) 1×10^{-19} J 2) 3.0×10^{-18} J 3) 1×10^{19} J 4) 3.0×10^{-19} J

LEVEL–IV

STATEMENT TYPE QUESTIONS

38. Statement I: Bohr's model contradicts Heisenberg's uncertainty principle
Statement II: Bohr's model explain the spectra of atoms containing more than one electron

1) Both statements are true. 2) Both statements are false.
3) Statement I is true, statement II is false
4) Statement I is false, statement II is true

39. Statement I: On heating a solid for a longer time, radiation become white and then blue as the temperature becomes very high.
Statement II: As the temperature increases radiations emitted go from a lower frequency to higher frequency.

1) Both statements are true. 2) Both statements are false.
3) Statement I is true, statement II is false
4) Statement I is false, statement II is true

MULTI CORRECT ANSWER TYPE QUESTIONS

40. Positive rays are:
1) Electromagnetic waves 2) Electrons
3) Positively charged gaseous ions 4) Protons

41. Which of the following are the units of wavelength?
1) Angstrom 2) Nanometre 3) Picometre 4) Microns

42. The characteristics is / are associated with Planck's theory is:
1) Radiations are associated with energy
2) The magnitude of energy associated with a quantum is proportional to frequency
3) Radiation energy is neither emitted nor absorbed continuously
4) Radiation energy is either emitted or absorbed discontinuously

LEVEL–V

COMPREHENSION TYPE QUESTIONS

PASSAGE-1:

According to the classical laws of mechanics or dynamics of physics, any charge particle revolving around another charged particle should lose energy continuously. Hence electron revolving round the nucleus should lose energy and fall inside the nucleus. But nucleus is found to be stable. Thus, Rutherford's atomic model does not explain the stability of an atom. It could not explain the distribution of electrons around the nucleus and does not tell us anything about their energies. If the electron loses energy continuously, then the atomic spectra should be continuous, but it is discontinuous. Hence. It could not explain the line spectrum.

43. Rutherford atomic model does not obey:
1) Classical laws of electrodynamics 2) Laws of electrolysis
3) Both 1 and 2 4) None of these

44. Rutherford's atomic model could not explain.
1) Gaps present in the spectrum 2) Stability of the atom
3) Both 1 and 2 4) None of these

45. In Rutherford's atomic model one of the defect is:
1) Comparison of atomic model with solar system.
2) Comparison of atomic model with watermelon.
3) Comparison of atomic model with Custard apple.
4) None of these.

PASSAGE-2:

The amount of energy associated with a quantum of radiation is proportional to the frequency of radiation.

$E \propto \nu$

$E = h\nu = \dfrac{hc}{\lambda}$

46. Calculate the number of protons emitted in 10 hours by a 60W sodium lamp [λ of photon = 5893 A^0].

1) 6.41×10^{22} 2) 6.41×10^{24} 3) 6.41×10^{26} 4) 6.41×10^{28}

47. An electromagnetic radiation of wavelength 242 nm is just sufficient to ionize a sodium atom. Calculate the ionization energy of sodium in KJ / mol.

1) 494. 5 KJ mol^{-1} 2) 484.5 KJ mol^{-1}

3) 454.5 KJ mol^{-1} 4) 424.4 KJ mol^{-1}

48. A bulb emits light of wavelength 4500 A^0. The bulb is rated as 150 w and 8 % of the energy is emitted as light. How many photons are emitted by the bulb per second?

1) 2.715×10^{16} 2) 2.715×10^{17} 3) 2.715 x 10^{18} 4) 2.715 x 10^{19}

MATRIX MATCH TYPE QUESTIONS

49.

COLUMN-I	COLUMN-II
a) Electron	p) Not constant
b) Proton	q) 4.8×10^{-10}
c) e/m ratio of proton	r) Thomson's
d) charge of proton	s) Gold stein
	t) Constant

50.

COLUMN – I	COLUMN – II
a) K – shell	p) 18 electrons
b) L – shell	q) 32 electrons
c) M – shell	r) 2 electrons
d) N – shell	s) 8 electrons

2. SPECTRUM AND HYDROGEN SPECTRUM

	♦	**SPECTRUM**
	♦	**HYDROGEN SPECTRUM**

When an element is excited by some method such as by heating, by passing electric current or by passing electric discharge, the atoms of the element emit electromagnetic radiations of definite frequencies. The arrangement of these radiations in the order of increasing wavelength or decreasing frequencies is called emission spectrum of the element. Since the radiations in the spectrum are emitted due to energy changes taking place in the atoms, this spectrum is also known as atomic spectrum.

SPECTRUM: The phenomenon of splitting of a beam of light into radiations of different frequencies after passing through the prism is called dispersion and the pattern of radiations obtained after dispersion of beam is called spectrum.

The instrument used for obtaining a spectrum is called spectroscope or spectrograph.

CONTINUOUS SPECTRUM

- In case of dispersion of sunlight, the seven colours obtained change from violet to red without any discontinuity, which means each colour blends into the other. Such a spectrum is called continuous spectrum.
- The spectrum obtained by the continuous of radiation is called continuous spectrum.
- In a continuous spectrum each colour fades into the next colour as in rainbow.
- The spectrum of incandescent white light obtained by heating a solid to a very high temperature is a continuous spectrum.

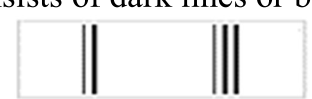
EMISSION SPECTRUM

Spectrum produced by the emitted radiation is known as emission spectrum. This spectrum corresponds to the radiation emitted (energy evolved) when an excited electron returns back to the ground state.

Emission spectrum consists of bright lines or bands on a dark background.

ABSORPTION SPECTRA:

Spectrum produced by the absorbed radiations is called absorption spectrum. Absorption spectrum consists of dark lines or bands on a bright background.



(Dark Lines with bright background)

Emission spectra:

(Bright Lines with dark background)

LINE SPECTRUM, BAND SPECTRUM

LINE SPECTRUM: Atomic spectra of most of the elements consist of a number of bright lines separated by dark bands. That is why atomic spectrum is also known as line spectrum. Atomic spectrum of an element can be used to identify the element and is sometimes called fingerprint of its atoms. Elements like rubidium (Rb), cesium (Cs), indium(In), scandium(Sc), etc. were discovered when their minerals were analyzed by spectroscopic methods. The spectrum obtained by the molecules is called band spectrum.

HYDROGEN SPECTRUM:

Hydrogen spectrum is the simplest of all atomic spectra. When an electric discharge is passed through hydrogen gas at low pressure, a bright light is emitted with a spectrometer and is found to comprise a series of lines of different wave lengths. Some of the lines are present in the visible region while the others in ultraviolet and infrared regions. The hydrogen spectrum consists of several series of lines named after their discoverers.

The first of these series was discovered by Balmer invisible region. Only this series is visible to the naked eye and hence it is called **Balmer series.** Balmer series is obtained when an excited electron is jumped from higher orbits to second orbit.

Balmer showed that the wave number of any line in the visible region can be expressed by a simple empirical equation which reads as

$$\overline{\vartheta} = R\left[\frac{1}{2^2}-\frac{1}{n^2}\right]cm^{-1}$$

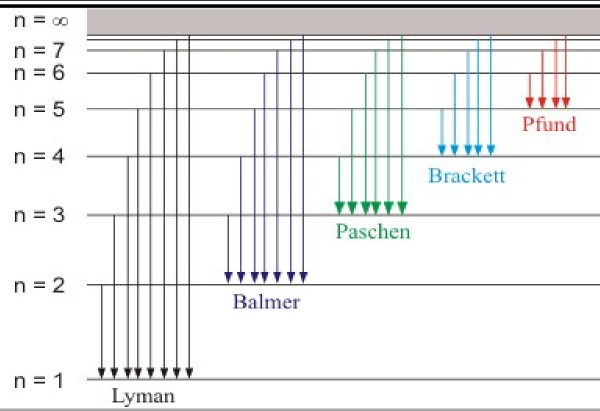
Where n is the number of the higher orbit from which electron jumped to second orbit i.e., n =3,4,5,6...., and R is a constant called **Rydberg constant.** The value of Rydberg constant for hydrogen, R_H= 109677 cm^{-1} or $1.09677 \times 10^7 m^{-1}$.

The wave numbers of different lines in the series are obtained by substituting n values.

Lyman observed same more spectral lines of the hydrogen emission in ultraviolet region.

Paschen, Brackett and Pfund observed separately, three different series of lines in infrared region

Different spectral lines of hydrogen emission are shown diagrammatically in figure and summarized in table.



DIFFERENT SERIES OF SPECTRAL LINES IN HYDROGEN SPECTRUM :

Name of the series	n_1	n_2	Spectral region	Equation for wave number
Lyman series	1	2,3,4,5,6,7,…	Ultraviolet	$\bar{\vartheta} = R\left[\frac{1}{1^2} - \frac{1}{n_2^2}\right]$
Balmer series	2	3,4,5,6,7,….	Visible	$\bar{\vartheta} = R\left[\frac{1}{2^2} - \frac{1}{n_2^2}\right]$
Paschen series	3	4,5,6,7,….	Near infrared	$\bar{\vartheta} = R\left[\frac{1}{3^2} - \frac{1}{n_2^2}\right]$
Brackett series	4	5,6,7,…..	Middle infrared	$\bar{\vartheta} = R\left[\frac{1}{4^2} - \frac{1}{n_2^2}\right]$
Pfund series	5	6,7,….	Far infrared	$\bar{\vartheta} = R\left[\frac{1}{5^2} - \frac{1}{n_2^2}\right]$

The wave numbers of all the lines in all the series can be calculated by the Rydberg equation

$$\bar{\vartheta} = \frac{1}{\lambda} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]cm^{-1}$$

Where n_1 and n_2 are whole numbers $n_2 > n_1$.n_1 is the number of the lower orbit to which electron jumped and n_2 is the number of higher orbit from which electron jumped.

For one electron species like He^+, Li^{2+} and Be^{3+}, the value of R is $109677 \times Z^2 cm^{-1}$, where Z is the atomic number of the species.

The wave number for any single electron species like He^+, Li^{2+} and Be^{3+} can be calculated:

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

where Z is atomic number of the species.

2. SPECTRUM AND HYDROGEN SPECTRUM

WORK SHEET LEVEL–I MAINS CORNER

SINGLE CORRECT ANSWER TYPE QUESTIONS

SPECTRUM

1. The pattern of radiations obtained after dispersion of beam is called:
   1) Spectrum 2) Hydrogen spectrum
   3) continuous spectrum 4) Emission spectrum
2. The spectrum obtained by the molecules is called:
   1) Line Spectrum 2) Hydrogen spectrum
   3) Continuous spectrum 4) Bond spectrum
3. The simplest of all atomic spectra is called:
   1) Spectrum 2) Hydrogen spectrum
   3) continuous spectrum 4) Emission spectrum
4. The spectrum produced by white light is:
   1) Emission spectrum 2) Continuous spectrum
   3) Absorption spectrum
   4) Both emission and continuous spectrum.
5. The spectrum of an atom is known as:
   1) Band spectrum 2) Emission spectrum
   3) Line Spectrum 4) Molecular spectrum
6. Band spectra is given by:
   1) Atoms 2) Molecules 3) Elements 4) None of these

HYDROGEN SPECTRUM

7. The spectral region of Bracket series is:
   1) Ultraviolet 2) Visible 3) Infrared 4) Near Infrared
8. Atomic spectrum is also known as:
   1) Band spectrum 2) Line spectrum
   3) Continuous spectrum 4) Emission spectrum
9. In hydrogen spectrum the following series of lines belongs to U.V region:
   1) Balmer series 2) Bracket series
   3) Paschen series 4) Lyman series

LEVEL–II

SPECTRUM

10. In hydrogen atom electron is present in N – shell. If its loss energy, a spectrum line may be observed in the region:
   1) Visible 2) Ultraviolet 3) Infrared 4) Any of these

11. Which of the following gives neither emission spectrum nor absorption spectrum?
1) He^{+} 2) H_2 3) H^{+} 4) He

12. The spectrum with all wavelengths may be:
1) Absorption spectrum 2) Emission spectrum
3) Continuous spectrum 4) Discontinuous spectrum

HYDROGEN SPECTRUM

13. When an electron in an excited hydrogen atom jumps from an energy level for which n=5 to a lower level for which n = 2, the spectral line is observed in the _____ series of the hydrogen spectrum.
1) Balmer 2) Lyman 3) Brackett 4) Pfund

14. The wavelength associated with a golf ball weighing 200 g and moving at a speed of 5mh-^{1} is of the order. Transition from n=2, 3, 4, 5 to n=——--- is called Lyman series.
1) 0 2)1 3) 2 4) 3

15. What are the values of n_1 and n_2 respectively for H_n line in the Lyman series of hydrogen atomic spectrum?
1) 3 and 5 2) 2 and 3 3) 1 and 3 4) 2 and 4

LEVEL-III ADVANCED CORNER

SINGLE CORRECT ANSWER TYPE QUESTIONS

16. Transition of electron from n= 3 to n = 1 level results in:
1) X - ray spectrum 2) Emission spectrum
3) Band spectrum 4) Infrared spectrum

17. In Bohr series of lines of hydrogen spectra, third line from the red and corresponds to which one of the following inner orbits jumps of electron for Bohr orbit in an atom of hydrogen:
1) $4 \to 1$ 2) $2 \to 5$ 3) $3 \to 2$ 4) $5 \to 2$

18. In which of the following cases, the wavelength emitted is minimum?
1) An electron jumps from 2nd to 1st level.
2) An electron jumps from 3rd to 2nd level.
3) An electron jumps from 4th to 3rd level.
4) An electron jumps from 5th to 4th level

19. Which of the following gives discrete emission spectrum?
1) Incandescent electric bulb 2) Sun
3) Mercury vapour lamp 4) Candle

20. Maximum difference in energy for hydrogen atom is found when:
1) $n_1=1$ & $n_2=2$ 2) $n_1=0$ & $n_2=0$ 3) $n_1=3$ & $n_2=1$ 4) $n_1=2$ & $n_2=1$

LEVEL–IV

STATEMENT TYPE QUESTIONS

21. Statement I: Limiting line in the Balmer series has a wavelength of 364.4 mm.
Statement II: Limiting line is obtained for a jump of electron from $n = \infty$.
1) Both statements are true 2) Both statements are false
3) Statement I is true, statement II is false
4) Statement I is false, Statement II is true

22. Statement I: It is essential that all the lines available in the emission spectrum will also be available in the absorption spectrum.
Statement II: The spectrum of hydrogen atom is only absorption spectrum.
1) Both statements are true. 2) Both statements are false.
3) Statement I is true, statement II is false.
4) Statement I is false, statement II is true.

MULTI CORRECT ANSWER TYPE QUESTIONS

23. The instrument used for spectrum is:
1) Spectroscope 2) Spectrograph 3) Seismograph 4) Stethoscope

24. Which of the following is / are correct statement?
1) Hydrogen spectrum is the simplest spectrum
2) Absorption spectrum consists of dark lines on bright background
3) Spectrum produced by the emitted radiation is called emission spectrum
4) In hydrogen spectrum wavelength are decreases from Lyman series to Pfund series

LEVEL–V

COMPREHENSION TYPE QUESTIONS

PASSAGE:

Hydrogen spectrum contains a number of groups of lines. They can be classified into various series.

25. Paschen series is due to the electronic transition from all higher levels to:
1) N – shell 2) M – shell 3) K – shell 4) L – shell

26. When the electron jumps from n = 5 to n = 4 level, the spectral line observe in the hydrogen spectrum belongs to:
1) Balmer series 2) Lyman series 3) Bracket series 4) Pfund series

27. The spectral region of Balmer series is:
1) UV region 2) Visible region 3) IR region 4) None of these

MATRIX MATCH TYPE QUESTIONS

28.

COLUMN-I	**COLUMN-II**
a) Absorptions spectra	p) Dark lines
b) Emission spectra	q) Atoms
c) Line spectrum	r) Dark background
d) Band spectrum	s) Molecules
	t) Bright background

3. QUANTUM MECHANICAL MODEL

	♦	DE-BROGLIE'S THEORY
	♦	HEISENBERG UNCERTAINTY PRINCIPLE
	♦	SCHRODINGER WAVE EQUATION

Quantum mechanical models are the modern concept to explain structure of atom. The main theories are:

DE-BROGLIE'S THEORY

In 1924, de-Broglie suggested that just as light exhibits wave and particles properties, all fundamental particles behave like a wave. Thus, according to de-Broglie, "all material particles in motion possess wave characteristics".

According to de-Broglie, the wavelength associated with a particle of mass m, moving with velocity v is given by the relation:

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

Where h is Planck's constant, v is the velocity and $p(=mv)$ is momentum of the particles. The waves associated with material particles are called matter waves.

DERIVATION OF DE-BROGLIE RELATIONSHIP

The relationship may be derived by combining the mass-energy relationships proposed by Max Planck and Einstein.

According to Planck, photon of light having energy E is associated with a wave of frequency v as:

$$E = hv.......(i)$$

According to Einstein, mass and energy are related as:

$$E = mc^2........(ii)$$ (c is velocity of light)

Combining the above two relations in eq. (i) and (ii), we get:

$$hv = mc^2$$

Since $v = \frac{c}{\lambda}$ $\quad$ $or\ h\frac{c}{\lambda} = mc^2$

Or $\frac{h}{\lambda} = mc$ $\quad$ $or\ \lambda = \frac{h}{mc}$

The equation is valid for a photon. De-Broglie suggested that on substituting the mass of the particle (m) and its velocity (v) in place of velocity of light (c), the equation can also be applied to material particles.

Thus, the wavelength of material particle, λ is:

$$\lambda = \frac{h}{mv} \text{ (where } mv = p = \text{momentum)}$$

This equation is known as de-Broglie's equation.

Since h is constant,

$$\lambda \propto \frac{1}{\text{momentum}}$$

The equation is known as de-Broglie relation.

Ex. Calculate the wavelength of a body of mass 1mg moving with a velocity of 10 m sec^{-1}

Sol. We know, $\lambda = \frac{h}{mv}$

substituting the values, m = 1mg = 10^{-6} kg, v = 10 $msec^{-1}$

$$h = 6.625 \times 10^{-34}\, kgm^2 s^{-1},$$

$$\lambda = \frac{6.625 \times 10^{-34}}{10^{-6}} = 6.625 \times 10^{-29}\, m$$

Ex. Calculate the momentum of a moving particle which has a de Broglie wavelength of 200 pm.

Sol.

$$\text{we know, } \lambda = \frac{h}{mv} \text{ or } mv = \frac{h}{\lambda}$$

$$h = 6.62 \times 10^{-34}\, Js = 6.62 \times 10^{-34}\, kg\ m^2 s^{-1}$$

$$\lambda = 200pm = 2 \times 10^{-10}\, m$$

$$\text{momentum, } mv = \frac{h}{\lambda} = \frac{6.62 \times 10^{-34}\, kgm^2 s^{-1}}{2 \times 10^{-10}\, m} = 3.31 \times 10^{-24}\, kg\ ms^{-1}$$

HEISENBERG UNCERTAINTY PRINCIPLE

The subatomic particles have significant wave character. Since the wave spread out in space, hence it is not possible to determine accurately the position and velocity of such particles at the same time. This physical limit was suggested by Werner Heisenberg (1927) in his uncertainty principle. This principle commonly known as the Heisenberg uncertainty principle. It states that "it is not possible to measure simultaneously both the position and momentum (or velocity) of a microscopic particle, with absolute accuracy.

Mathematically, this law may be expressed as:

$$\Delta x \times \Delta p \geq \frac{h}{4\pi}$$

Where Δx = uncertainty in position

$\Delta p =$ uncertainty in momentum

The sign ≥ means that the product of Δx and Δp can be either greater than or equal to $\frac{h}{4\pi}$. It can never be less than $\frac{h}{4\pi}$. The sign equality refers to minimum uncertainty and is equal to $\frac{h}{4\pi}$. Thus,

(i) If Δx is small i.e., the position of particle is measured accurately, Δp would be large, i.e., there would be large uncertainty in its momentum.

(ii) On the other hand, if Δp is small, the momentum of the particle is measured more accurately, Δx would be large i.e., there would be large uncertainty with regard to the position of the particle.

Ex. The uncertainty in the position of an electron (mass = 9.1×10^{-28}g) moving with a velocity of 3.0×10^4 cm s^{-1} accurate upto 0.011% will be:

Sol. $\Delta x . m\Delta v = \frac{h}{4\pi}$

$$\Delta x = \frac{6.625\times10^{-27}}{4\times3.14\times9\times10^{-28}\times3\times10^{4}} = 0.0193\times10^{-3} = 1.92\,cm$$

SCHRODINGER WAVE EQUATION

Quantum mechanics, as developed by Erwin Schrodinger in 1926, is based on the wave motion associated with the particles. For the wave motion of the electron in three-dimensional space around the nucleus, he put forward an equation, known after his name as Schrodinger wave equation, which plays an important role in quantum mechanics. The equation is:

$$\frac{\delta^2\Psi}{\delta x^2}+\frac{\delta^2\Psi}{\delta y^2}+\frac{\delta^2\Psi}{\delta z^2}+\frac{8\pi^2 m}{h^2}(E-V)\Psi=0$$

Where Ψ is the amplitude of the wave where the coordinates of the electron are (x, y, z), E is the total energy of the electron, V is its potential energy, m is the mass of the electron and h is Planck's constant. $\delta^2\Psi / \delta x^2$ represents second derivative of Ψ with respect to x and so on.

In short, Schrodinger wave equation is written as:

$\hat{H}\Psi = E\Psi$

Where $\hat{H}$ is a mathematical operator, called Hamiltonian operator.

The solution of Schrodinger wave equation for an electron in an atom gives the value of E and Ψ. The values of E represents the quantized values of energy which the electrons in the atom can have. The corresponding values of Ψ are called wave functions.

3. QUANTUM MECHANICAL MODEL

WORK SHEET **LEVEL–I** **MAINS CORNER**

SINGLE CORRECT ANSWER TYPE QUESTIONS

DE-BROGLIE'S THEORY

1. The wave nature of an electron was first given by:
   1) De-Broglie 2) Heisenberg 3) Mosley 4) Somerfield
2. Which one of the following expressions give the de-Broglie relationship?
   1) $h = \frac{\lambda}{mv}$ 2) $\lambda = \frac{h}{mv}$ 3) $\lambda = \frac{m}{hv}$ 4) $\lambda = \frac{\upsilon}{mv}$
3. The de-Broglie equation applies:
   1) To electrons only
   2) To neutrons only
   3) To protons only
   4) All the material objects in motion
4. Minimum de-Broglie wavelength is associated with:
   1) Electron 2) Proton 3) CO_2 molecule 4) SO_2 molecule

HEISENBERG UNCERTAINTY PRINCIPLE

5. Heisenberg's uncertainty principle rules out the exact simultaneously measurement of:
   1) Probability and Intensity
   2) Energy and velocity
   3) Charge density and radius
   4) Position and momentum
6. If the uncertainty in the position of an electron is zero, the uncertainty in its momentum would be:
   1) Zero 2) Greater than $\frac{h}{4\pi}$
   3) less than $\frac{h}{4\pi}$ 4) infinite
7. The uncertainty found from the uncertainty principle $\left(\Delta x \times \Delta p = \frac{h}{4\pi}\right)$ is:
   1) The minimum value. 2) The maximum value.
   3) The exact value. 4) Only an approximate value.

SCHRODINGER WAVE EQUATION

8. The correct Schrodinger's wave equation for an electron with E as total energy and V as potential energy is:

1) $\frac{\partial^2\psi}{\partial x^2}+\frac{\partial^2\psi}{\partial y^2}+\frac{\partial^2\psi}{\partial z^2}+\frac{8\pi m}{h^2}(E-V)\psi=0$

2) $\frac{\partial^2\psi}{\partial x^2}+\frac{\partial^2\psi}{\partial y^2}+\frac{\partial^2\psi}{\partial z^2}+\frac{8\pi^2}{mh^2}(E-V)\psi=0$

3) $\frac{\partial^2\psi}{\partial x^2}+\frac{\partial^2\psi}{\partial y^2}+\frac{\partial^2\psi}{\partial z^2}+\frac{8\pi^2 m}{h^2}(E-V)\psi=0$

4) $\frac{\partial^2\psi}{\partial x^2}+\frac{\partial^2\psi}{\partial y^2}+\frac{\partial^2\psi}{\partial z^2}+\frac{8\pi^2 m}{h^2}(E-V)x=0$

9. The probability of finding the electron in the orbital is:
1) 20% 2) 90–95% 3) 70–80% 4) 50–60%

10. ψ^2, the wave function represents the probability of finding electron. Its value depends:
1) Inside the nucleus 2) Far from the nucleus
3) Near the nucleus 4) Upon the type of orbital

LEVEL–II

DE-BROGLIE'S THEORY

11. The de-Broglie wavelength of a particle with mass 1g and velocity $100 m/sec$ is:
1) $6.63\times10^{-33}m$ 2) $6.63\times10^{-34}m$ 3) $6.63\times10^{-35}m$ 4) $6.65\times10^{-35}m$

12. An electron has kinetic energy $2.8\times10^{-23}J$. de-Broglie wavelength will be nearly: $(m_e=9.1\times10^{-31}kg)$
1) $9.28\times10^{-4}m$ 2) $9.28\times10^{-7}m$ 3) $9.28\times10^{-8}m$ 4) $9.28\times10^{-10}m$

13. What will be de-Broglie wavelength of an electron moving with a velocity of $1.2\times10^5 ms^{-1}$?
1) 6.066×10^{-9} 2) 3.133×10^{-37} 3) 6.5×10^{-9} 4) 6.018×10^{-7}

HEISENBERG UNCERTAINTY PRINCIPLE

14. The uncertainty in momentum of an electron is $1\times10^{-5}kg-m/s$. The uncertainty in its position will be ($h=6.62\times10^{-34}kg-m^2/s$):
1) $1.05\times10^{-28}m$ 2) $1.05\times10^{-26}m$
3) $5.27\times10^{-30}m$ 4) $5.25\times10^{-28}m$

15. The uncertainty in the position of a moving bullet of mass 10 g is 10^{-5} m. Calculate the uncertainty in its velocity.
1) 5.2×10^{-28} m / sec 2) 3.0×10^{-28} m / sec
3) 5.2×10^{-22} m / sec 4) 3×10^{-22} m / sec

16. Uncertainty in position of a 0.25 g particle is 10^{-5}. Uncertainty of velocity is $(h = 6.6 \times 10^{-34}$ Js$)$:
1) 1.2×10^{34} 2) 2.1×10^{-29} 3) 1.6×10^{-20} 4) 1.7×10^{-9}

SCHRODINGER WAVE EQUATION

17. The solutions of Schrodinger's wave equation are called:
1) Schrodinger solution 2) Hamiltonian operator
3) Wave function 4) All

18. Each permitted solution or wave function corresponds to a definite energy state called:
1) Orbit 2) Sub-level 3) Orbital 4) All

19. The probability of finding an electron at a point within an atom is proportional to:
1) ψ 2) ψ^2 3) ψ^3 4) ψ^4

LEVEL-III ADVANCED CORNER

SINGLE CORRECT ANSWER TYPE QUESTIONS

20. The de-Broglie wavelength associated with a particle of mass 10^{-6} kg moving with a velocity of $10\ ms^{-1}$, is:
1) 6.63×10^{-22} m 2) 6.63×10^{-29} m 3) 6.63×10^{-31} m 4) 6.63×10^{-34} m

21. If the velocity of hydrogen molecule is 5×10^4 cm sec^{-1}, then its de-Broglie wavelength is:
1) 2 A^0 2) 4 A^0 3) 8 A^0 4) 100 A^0

22. A cricket ball of 0.5 kg is moving with a velocity of 100 m / sec. The wavelength associated with its motion is:
1) 1 / 100cm 2) 6.6×10^{-34} m 3) 1.32×10^{-35} m 4) 6.6×10^{-28} m

23. The uncertainty in momentum of an electron is 1×10^{-5} kg m / s. The uncertainity in its position will be $(h = 6.63 \times 10^{-34}$ Js$)$:
1) 5.28×10^{-30} m 2) 5.25×10^{-28} m 3) 1.05×10^{-26} m 4) 2.715×10^{-30} m

24. According to Heisenberg's uncertainty principle, the product of uncertainties in position and velocities for an electron of mass $9.1\times10^{-31}\text{kg}$ is:
1) $2.8\times10^{-3}\text{m}^2\text{s}^{-1}$ 2) $3.8\times10^{-5}\text{m}^2\text{s}^{-1}$ 3) $5.8\times10^{-5}\text{m}^2\text{s}^{-1}$ 4) $6.8\times10^{-6}\text{m}^2\text{s}^{-1}\text{a}$

LEVEL–IV

STATEMENT TYPE QUESTIONS

25. Statement I: Wave character of electrons was experimentally verified by Davisson and Germer.
Statement II: Formula of sulphur dioxide is SO_3.
1) Both statements are true. 2) Both statements are false.
3) Statement I is true, statement II is false
4) Statement I is false, statement II is true.

26. Statement I: For objects of ordinary size, Heisenberg's principle has no significance.
Statement II: Heisenberg's principle has no significance for a moving cricket ball.
1) Both statements are true. 2) Both statements are false.
3) Statement I is true, statement II is false
4) Statement I is false, statement II is true.

MULTI CORRECT ANSWER TYPE QUESTIONS

27. Heisenberg uncertainity principle is not valid for:
1) Moving electron 2) Motor car
3) Stationary particle 4) None of these

28. The De – Broglie wavelength for a proton with a velocity 15 % of the speed of light is:
1) $0.088 \text{ x } 10^{-13}$m 2) $8.8 \text{ x } 10^{-15}$ m 3) $0.088 \text{ x } 10^{-11}$ cm4) $8.8 \text{ x } 10^{-13}$ cm

LEVEL–V

COMPREHENSION TYPE QUESTIONS

PASSAGE:

De – Broglie derived a relationship for the calculation of wavelength (λ) of the wave associated with a particle of mass (m) moving with velocity (v) as given below:

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

29. The momentum of radiation of wavelength of 0.33 nm is _____ kgmsec^{-1}.
1) 2×10^{-24} 2) 2×10^{-12} 3) 2×10^{-6} 4) 2×10^{-48}

30. Calculate the wavelength (in nm) associated with a proton moving at $1.0\times10^{3}\,m/s$. The mass of proton is $1.67\times10^{-27}kg$ and h is 6.63×10^{-34} J sec.
1) 0.032 nm 2) 2.5 nm 3) 14 nm 4) 0.4 nm

31. The velocity of an electro with de – Broglie wavelength of 1.0 x 10^{2} nm is:
1) $7.2\times10^{5}cm/\sec$ 2) $7.2\times10^{3}cm/\sec$
3) $7.2\times10^{4}cm/\sec$ 4) $3.6\times10^{5}cm/\sec$

MATRIX MATCH TYPE QUESTIONS

32.

COLUMN-I	COLUMN-II
a) Planck's law	p) Determines the probability of an electron
b) de Broglie equation	q) $E = h\nu$
c) Wave function	r) It refers to amplitude of an electron wave
d) ψ^2	s) $\lambda = \frac{h}{mv}$
	t) ψ

4. SHAPES OF ORBITALS

	♦	s-ORBITAL
	♦	p-ORBITAL
	♦	d-ORBITAL
	♦	NODAL PLANES

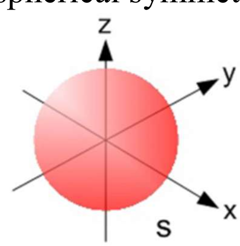
ORBITAL

The three-dimensional space around the nucleus in an atom where the probability of finding an electron is maximum is called an orbital.

SHAPE OF S-ORBITAL

The orbital in which the electrons with the quantum numbers $n=1$ and $\ell=0$ are present is called 1s orbital. Similarly, the electrons having quantum numbers $n=2$ and $\ell=0$ are present in 2s orbital. Thus all the s-orbitals have $\ell=0$ while 'n' can have values 1,2,3,4,…

s-orbital is spherical. It has spherical symmetry.



SHAPE OF P-ORBITAL

The orbital in which the electrons have quantum number $n=2$ and $\ell=1$ are present, is called 2p orbital. Similarly, the orbital with quantum numbers $n=3$ and $\ell=1$ is called 3p orbital.

For all p-orbitals $\ell=1$ and $n=2,3,.....$

In each principal quantum number (except first orbit) there are three p-orbitals and they are mutually perpendicular to one another and oriented along the three axes.

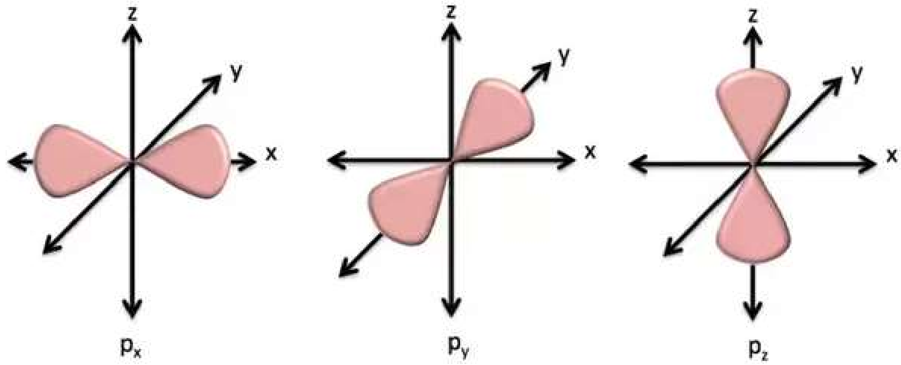
Each p-orbital has a dumbbell shape

The probability of finding the electron at the nucleus is zero. At a certain distance from the nucleus the probability is maximum.

A p-orbital has one nodal plane, and it consists of two lobes on either side as the nodal plane extending along the axis.

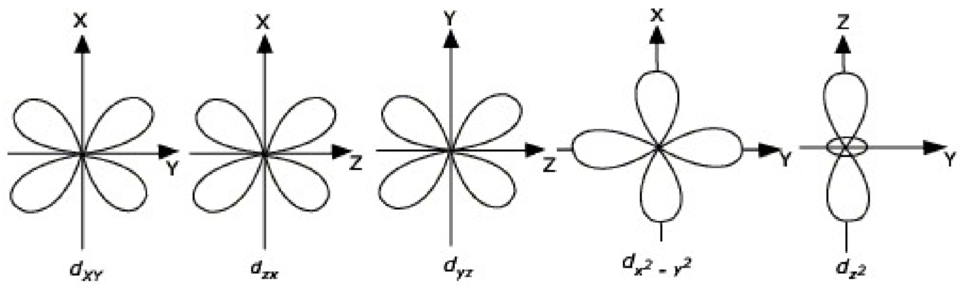
For the p_x orbital the nodal plane is YZ plane.

For the p_y & p_z orbitals, the nodal planes are XZ plane and XY plane respectively.



SHAPES OF d-ORBITALS

The orbitals having electrons with quantum numbers $n = 3$ and $l = 2$ are called 3d orbitals.



There are five d-orbitals. They are $d_{xy}, d_{yz}, d_{xz}, d_{x^2-y^2}, d_{z^2}$. The first four are of double dumbbell in shape. Each has four lobes. The shape of d_{z^2} is doughnut shape.

d-orbital	Plane
d_{xy}	Between X and Y axes
d_{yz}	Between Y & Z axes
d_{zx}	Between Z and X axes
$d_{x^2-y^2}$	Along X and Y axes
d_{z^2}	Along Z axis with a dumb-bell shape. It contains a ring (torus, collar or tyre) along xy-plane

NODAL PLANES

The plane where the probability of finding the electrons is zero $\left(\phi^2 = 0\right)$ is called a nodal plane.

- Number of nodal planes in an orbital $= \ell$
- When the number of nodal planes increases, the energy of the orbital increases. So, the energy order of the orbital is $s < p < d < f$
- Number of radial nodes $= n - \ell - 1$
  Where n = principal quantum number, $\ell =$ azimuthal quantum number.

4. SHAPES OF ORBITALS

WORK SHEET LEVEL–I MAINS CORNER

SINGLE CORRECT ANSWER TYPE QUESTIONS

s-ORBITAL

1. The three-dimensional space around the nucleus in an atom where the probability of finding an electron is maximum is called:
1) an orbital 2) s-orbital 3) p-orbital 4) d-orbital
2. The orbital in which the electrons with the quantum numbers n = 1 & l = 0 are called:
1) an orbital 2) s-orbital 3) p-orbital 4) d-orbital
3. The shape of s-orbital is:
1) Pyramidal 2) Spherical 3) Tetrahedral 4) Dumb-bell shaped
4. The space within an atom, where there is maximum probability of finding an electron at any instant is:
1) An orbit 2) An orbital
3) A stationary orbit 4) Shell
5. An example of non-directional orbital is:
1) 3s 2) 2p 3) 3d 4) 4f

p-ORBITAL

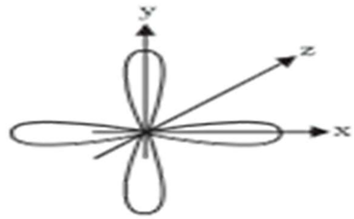
6. The orbital in which the electrons have quantum numbers n = 2 & l = 1 are called:
1) an orbital 2) s-orbital 3) p-orbital 4) d-orbital
7. The shape of p-orbital is:
1) Elliptical 2) Spherical
3) Dumb-bell 4) Complex geometrical
8. The number of orbitals in 2p sub-shell is:
1) 6 2) 2 3) 3 4) 4
9. Which of the following is correct with respect to p orbitals?
1) Spherical 2) Strong directional character
3) Fivefold degenerate 4) No directional character

d-ORBITAL

10. The orbitals having electrons with quantum numbers n = 3 & l = 2 are called:
1) an orbital 2) s-orbital 3) p-orbital 4) d-orbital
11. The shape of d_{xy} orbital will be:
1) Circular 2) Dumb-bell
3) Double dumb-bell 4) Trigonal
12. The number of orbitals in d sub-shell is:
1) 1 2) 3 3) 5 4) 7

13. Maximum number of electrons that d orbitals can accommodate is:
1) 6 2) 10 3) 2 4) 18

14. The figure given below is a representation of the shape of:



1) $3d_{xy}$ orbital 2) $3d_{z^2}$ orbital 3) $2p_z$ orbital 4) $3d_{x^2-y^2}$ orbital

NODAL PLANES

15. The plane where the probability of finding the electrons is zero $\left(\phi^2 = 0\right)$ is called:
1) nodal planes 2) s-orbital 3) p-orbital 4) d-orbital

16. Energy of atomic orbitals in a particular shell is in order of:
1) s< p < d < f 2) s > p > d > f 3) p < d < f < s 4) f > d > s > p

LEVEL–II

s-ORBITAL

17. Which of the sub-shell is circular?
1) 4s 2) 4f 3) 4p 4) 4d

18. For which of the following orbital, a plane never exists with φ^2 value zero:
1) S - orbital 2) p – orbital 3) d – orbital 4) f – orbital

p-ORBITAL

19. The shape of 2p orbital is:
1) Spherical 2) Ellipsoidal 3) Dumb-bell 4) Pyramidal

20. p_x orbital can accommodate:
1) 4 electrons 2) 6 electrons
3) 2 electrons with parallel spins 4) 2 electrons with opposite spins

d-ORBITAL

21. Which of the following orbitals has appearance like a baby soother?
1) d_{xy} 2) d_{yz} 3) $d_{x^2-y^2}$ 4) d_{z^2}

22. Which of the following orbitals does not exist?
1) $d_{x^2-y^2}$ 2) $d_{x^2-z^2}$ 3) d_{xy} 4) d_{xz}

23. The d - orbital with the orientation along x and y - axis is called as:
1) d_{xy} 2) d_{yz} 3) d_{xz} 4) $d_{x^2-y^2}$

NODAL PLANES

24. The no. of radial nodes for 5p orbital:
1) 2 2) 3 3) 4 4) 5
25. The orbital that has equal number of nodal spaces and nodal planes:
1) 2p 2) 3p 3) 4p 4) 5p

LEVEL-III ADVANCED CORNER

SINGLE CORRECT ANSWER TYPE QUESTIONS

26. There is no difference between a 2p and a 3p orbital regarding
1) Size 2) Value of n 3) Energy 4) Shape
27. Which one of the following atomic orbitals is not directed along the axis?
1) p_x 2) $d_{x^2-y^2}$ 3) d_{xy} 4) d_{z^2}
28. The number of radial nodes, nodal planes for an orbital with n = 4, $l=1$ is:
1) 3, 1 2) 2, 1 3) 2, 0 4) 4, 0
29. The probability of finding an electron in YZ plane for a p_x orbital is:
1) Zero 2) Negative 3) Maximum 4) Not determined
30. The electron density between ‘1s’ and ‘2s’ is:
1) High 2) Low 3) Zero 4) Abnormal

LEVEL–IV

STATEMENT TYPE QUESTIONS

31. Statement I: An orbital cannot have more than two electrons.
Statement II: The two electrons in an orbital create opposite magnetic field.
1) Both statements are true 2) Both statements are false
3) Statement I is true, Statement II is false
4) Statement I is false, Statement II is true
32. Statement I: The P_x orbital has maximum electron density along the X – axis and its nodal plane is YZ plane.
Statement II: For a given atom, for all values of n, the p – orbitals have the same shape, but the overall size increase as ‘n’ increase.
1) Both statements are true 2) Both statements are false
3) Statement I is true, Statement II is false
4) Statement I is false, Statement II is true

MULTI CORRECT ANSWER TYPE QUESTIONS

33. The correct statement is / are:
    1) Electron density in the xy – plane in $3d_{x^2-y^2}$ orbital is zero
    2) Electron density in the xy – plane in $3d_{z^2}$ orbital is zero
    3) 2s – orbital has one nodal surface
    4) For $2p_z$ orbital yz is the nodal plane
34. Which is correct with respect to 'd' orbital?
    1) They are double dumbbell shape 2) They are five fold degenerate
    3) Degenerate orbitals have same energy
    4) They have a directional character

LEVEL–V

COMPREHENSION TYPE QUESTIONS

PASSAGE:

In an atom, a large number of orbitals are possible. These orbitals differ in their size, shape and orientation in space around the nucleus. The state of an electron in any atom is described by its location with respect to the nucleus and by its energy.

35. Number of electrons in p, d subshells respectively are:
    1) 2, 6 2) 6, 10 3) 10, 14 4) 3, 5
36. A neutral atom of an element has two 'K', eight 'L', nine 'M' and two 'N' electrons. The total number of electrons present in the subshell having *l* value "1" is:
    1) 6 2) 8 3) 10 4) 12
37. f – orbital consist of _________ subshells.
    1) 1 2) 3 3) 5 4) 7

MATRIX MATCH TYPE QUESTIONS

38.

COLUMN – I	**COLUMN – II**
a) s – orbital	p) Double dumb bell
b) p – orbital	q) Spherical
c) d – orbital	r) Complex
d) f – orbital	s) Dumb bell

5. QUANTUM NUMBERS

	♦	PRINCIPAL QUANTUM NUMBER
	♦	AZIMUTHAL QUANTUM NUMBER
	♦	MAGNETIC QUANTUM NUMBER
	♦	SPIN QUANTUM NUMBER

- A set of numbers used to provide a complete description of an electron in an atom is called quantum numbers.

  Four quantum numbers required for a complete explanation of electrons in an atom. They are

  1. Principal quantum number 2. Azimuthal quantum number
  3. Magnetic quantum number 4. Spin quantum number

PRINCIPAL QUANTUM NUMBER

- It was proposed by Niels Bohr
- It is denoted by 'n'
- The values of n=1,2,3,4…. or K, L, M, N………respectively
- It indicates the size and energy of the orbit.
- The maximum number of electrons in an orbit = $2n^2$

  Total number of orbital's $=n^2$, where n = no. of the orbit
- Angular momentum of an electron in an orbit $= \dfrac{nh}{2\pi}$

AZIMUTHAL QUANTUM NUMBER

- It was proposed by Somerfield
- It is denoted by ***'l'***
- The values of l = 0,1, 2….(n-1)
- The values of l represents various subshells, when l =0,1,2,3….etc are called s,p,d,f…..sub shells respectively
- Energies are in the order of s<p<d<f
- It indicates the shape of orbit or orbital and angular momentum of electron.
- Number of sub shells in an energy level = n

  Where n = no. of the orbit
- Angular momentum of the electron in an orbital

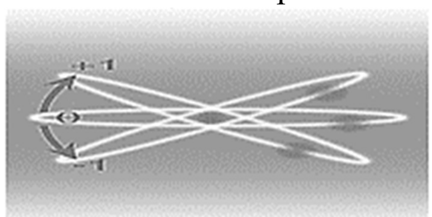
  $$= \frac{h}{2\pi}\sqrt{\ell(\ell+1)} = \bar{h}\sqrt{\ell(\ell+1)} \left(\bar{h} = \frac{h}{2\pi}\right)$$

  Where h = Planck's constant l = Azimuthal quantum number

MAGNETIC QUANTUM NUMBER

- It was proposed by Lande.

- It is denoted by **'m'**
- The values of m= –l….0…..+l
- The total m values = $2l + 1$
- The total number of m values indicate the total number of orbitals in the sub shell.
- The number of orbitals in s, p, d and f sub shells are 1,3,5 and 7 respectively.
- It indicates the orientation of orbitals in space.



- The number of orbitals in a sub shell =*2l+1*
- Maximum number of electrons in a sub shell=*2(2l+1)* where *l*=azimuthal quantum number.

SPIN QUANTUM NUMBER

- It was proposed by Goudsmith and Uhlenbeck.
- The values of s = $+\frac{1}{2}$ and - $\frac{1}{2}$.
- The clock wise**(↑)** direction spin is represented by+ $\frac{1}{2}$ and antic lock wise**(↓)** direction spin is represented by – $\frac{1}{2}$.
- For each value of m, there can be two 's' values.
- It indicates the direction of the spin of the electron.
- Maximum number of electrons in an orbital=2.
- The maximum number of electrons present in s, p, d and f shells are 2,6,10 and 14 respectively.

If 2 e^- are present both are moving in same direction are called parallel spin.

If the 2 e^- are present both are moving in opposite direction is called anti parallel spin.

Ex. Orbital angular momentum for a d-electron is:

Sol. We know that for d-electron l = 2.

$$\mu = \sqrt{l(l+1)}\frac{h}{2\pi}; \mu = \sqrt{2(2+1)}\frac{h}{2\pi} = \sqrt{2(2+1)}\frac{h}{2\pi}; \mu = \sqrt{6}\frac{h}{2\pi}$$

Ex. The set of quantum numbers for the outermost electron for copper in its ground state is:

Sol: Electronic configuration of Cu

29 Cu → $ls^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$ outermost electron is in 4s subshell

For 4s n = 4, l = 0, m = 0, s = + 1/2 or - ½.

5. QUANTUM NUMBERS

WORK SHEET LEVEL–I MAINS CORNER

SINGLE CORRECT ANSWER TYPE QUESTIONS

PRINCIPAL QUANTUM NUMBER

1. Principal quantum number of an atom represents:
   1) Size of the orbital 2) Spin angular momentum
   3) Orbital angular momentum 4) Space orientation of the orbital
2. The total number of orbitals in an energy level designated by principal quantum number n is equal to:
   1) $2n$ 2) $2n^2$ 3) n 4) n^2
3. The number of orbitals in the fourth principal quantum number will be:
   1) 4 2) 8 3) 12 4) 16

AZIMUTHAL QUANTUM NUMBER

4. The shape of an orbital is given by the quantum number:
   1) n 2) l 3) m 4) s
5. The angular momentum of an electron depends on:
   1) Principal quantum number 2) Azimuthal quantum number
   3) Magnetic quantum number 4) All of these
6. A sub-shell $l=2$ can take how many electrons:
   1) 3 2) 10 3) 5 4) 6
7. When the azimuthal quantum number has a value of $l = 0$, the shape of the orbital is:
   1) Rectangular 2) Spherical 3) Dumbbell 4) Unsymmetrical
8. When the azimuthal quantum number has a value of $l = 1$, the shape of the orbital is:
   1) Unsymmetrical 2) Spherically symmetrical
   3) Dumb-bell 4) Complicated

MAGNETIC QUANTUM NUMBER

9. The magnetic quantum number specifies:
   1) Size of orbitals 2) Shape of orbitals
   3) Orientation of orbitals 4) Nuclear stability
10. For a given value of quantum number l, the number of allowed values of m is given by:
    1) $l+2$ 2) $2l+2$ 3) $2l+1$ 4) $l+1$
11. For p-orbital, the magnetic quantum number has value:
    1) 2 2) 4, – 4 3) – 1, 0, +1 4) 0
12. The magnetic quantum number for valency electrons of sodium is:
    1) 3 2) 2 3) 1 4) 0

SPIN QUANTUM NUMBER

13. A set of numbers used to provide a complete description of an electron in an atom is called:
1) Quantum number 2) Azimuthal quantum number
3) Spin quantum number 4) Principal quantum number

14. The quantum number not obtained from the Schrodinger's wave equation is:
1) n 2) l 3) m 4) s

LEVEL–II

PRINCIPAL QUANTUM NUMBER

15. The quantum number which specifies the location of an electron as well as energy is:
1) Principal quantum number 2) Azimuthal quantum number
3) Spin quantum number 4) Magnetic quantum number

16. The maximum number of electrons that can be accommodated in the M shell is:
1)2 2)8 3)18 4)32

17. For $n = 3$ energy level, the number of possible orbitals (all kinds) are:
1) 1 2) 3 3) 4 4) 9

AZIMUTHAL QUANTUM NUMBER

18. 2*p* orbitals have:
1) $n = 1, l = 2$ 2) $n = 1, l = 0$ 3) $n = 2, l = 1$ 4) $n = 2, l = 0$

19. If $n = 3$, then the value of 'l' which is incorrect?
1) 0 2) 1 3) 2 4) 3

20. For the dumb-bell shaped orbital, the value of l is:
1) 3 2) 1 3) 0 4) 2

21. How many electrons can be accommodated in a sub-shell for which n=3, l=1?
1) 8 2) 6 3) 18 4) 32

MAGNETIC QUANTUM NUMBER

22. For ns orbital, the magnetic quantum number has value:
1) 2 2) 4 3) – 1 4) 0

23. If value of azimuthal quantum number l is 2, then total possible values of magnetic quantum number will be:
1) 7 2) 5 3) 3 4) 2

24. If the value of azimuthal quantum number is 3, the possible values of magnetic quantum number would be:
1) 0, 1, 2, 3 2) 0, – 1, – 2, – 3 3) 0, ± 1, ± 2, $\pm$ 3 4) ± 1, ± 2, ± 3

SPIN QUANTUM NUMBER

25. The two electrons in K sub-shell will differ in:
1) Principal quantum number 2) Azimuthal quantum number
3) Magnetic quantum number 4) Spin quantum number

26. If an electron has spin quantum number of $+\frac{1}{2}$ and a magnetic quantum number of -1, it cannot be presented in an:
1) d-orbital 2) f-orbital 3) p-orbital 4) s-orbital

LEVEL-III ADVANCED CORNER

SINGLE CORRECT ANSWER TYPE QUESTIONS

27. Principal, azimuthal and magnetic quantum numbers are respectively related to:
1) Size, shape and orientation
2) Shape, size and orientation
3) Size, orientation and shape
4) None of the above

28. The quantum numbers for the outermost electron of an element are given below as $n = 2, l = 0, m = 0, s = +\frac{1}{2}$. The atoms is:
1) Lithium 2) Beryllium 3) Hydrogen 4) Boron

29. Which of the following sets of quantum numbers represent an impossible arrangement?

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

30. The magnetic quantum number for an electron when the value of principal quantum number is 2 can have:
1) 3 values 2) 2 values 3) 9 values 4) 6 values

31. If magnetic quantum number of a given atom represented by –3, then what will be its principal quantum number?
1) 2 2) 3 3) 4 4) 5

LEVEL–IV

STATEMENT TYPE QUESTIONS

32. Statement I: Total number of orbitals associated with principal quantum number n=3 is 6.
Statement II: Number of orbitals in a shell equals to n^2.
1) Both statements are true.
2) Both statements are false.
3) Statement I is true, statement II is false
4) Statement I is false, statement II is true

33. Statement I: Energy of the orbitals increases as 1s < 2s = 2p < 3s = 3p < 3d = 4s = 4p = 4d = 4f < ....
Statement II: Energy of the electron depends completely on principle quantum number.
1) Both statements are true.
2) Both statements are false.
3) Statement I is true, statement II is false
4) Statement I is false, statement II is true

MULTI CORRECT ANSWER TYPE QUESTIONS

34. Which of the following is correct regarding M – shell?
1) 3 subshells 2) 18 electrons 3) s, p, d, f 4) 8 electrons

35. Each electron is designated by:
1) n 2) l 3) m 4) s

LEVEL–V

COMPREHENSION TYPE QUESTIONS

PASSAGE:

Four quantum numbers are required for the complete explanation of electrons in an atom.

36. How many sets of four quantum numbers are possible for electrons present in He^{+2} anion?
1) 2 2) 4 3) 5 4) 7

37. For the azimuthal quantum number 'l' the total number of magnetic quantum numbers is given by:

1) $l=\frac{m+1}{2}$ 2) $l=\frac{m-1}{2}$ 3) $l=\frac{2m+1}{2}$ 4) $l=\frac{2m-1}{2}$

38. The lowest orbital in which an electron with azimuthal quantum no. value 3 is:

1) 4 2) 5 3) 1 4) 6

MATRIX MATCH TYPE QUESTIONS

39.

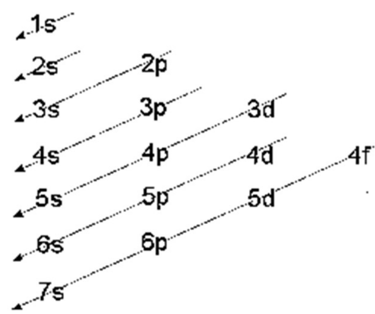
COLUMN-I	COLUMN-II
a) Principal quantum number proposed by	p) Bohr
b) Azimuthal quantum number proposed by	q) Somerfield
c) Maximum number of electrons in an orbit=	r) $2n^2$
d) Shapes of subshells	s) Azimuthal quantum number
	t) Size of an atom

6. ARRANGEMENT OF ELECTRONS IN AN ORBITALS

	♦	**AUFBAU PRINCIPLE**
	♦	**PAULI'S EXCLUSION PRINCIPLE**
	♦	**HUND'S RULE**
	♦	**ELECTRONIC CONFIGURATION**

AUFBAU PRINCIPLE

In German language Aufbau means building up (Building up of orbital's). According to Aufbau principle electrons first enter into lower energy orbitals.



Increasing order of energy orbitals: 1s < 2s < 2p < 3s < 3p < 4s < 3d < 4P < 5s < 4d…………

Filling up of electrons in orbital's based on "n+*l*" values.

Case 1: An electron first enters into the sub shell which is having least "n+*l*" value.

2s	2p
n = 2	n = 2
$l = 0$	$l = 1$
$n + l = 2 + 0 = 2$	$n + l = 2 + 1 = 3$

$\therefore$ First electrons enters into "2s" orbital.

Case-2: "n + *l*" value for the orbitals is same then electron enters into sub shell which is having least "n" value.

3s	2p
n = 3	n = 2
$l = 0$	$l = 1$
$n + l = 3 + 0 = 3$	$n + l = 2 + 1 = 3$

$\therefore$ Then electrons first enters into "2p"

PAULI'S EXCLUSION PRINCIPLE

➢ It was stated by W. Pauli.

- Pauli's principle states that no two electrons in an atom can have all the four quantum numbers identical.
- The significance of Pauli's principle - there is room for only two electrons in an orbital and they should have opposite spins.
- The capacity of an orbital is restricted to two because of Pauli's principle.
- In an atom, if electrons have same n, l and m values, they must differ in spin quantum number.
- This means in an atom if any two electrons have any three quantum numbers the same, then certainly those two electrons differ in fourth quantum number.
- The capacity of any sub-level is determined on the basis of Pauli's principle.
- s-sub level - one orbital so, 2 electrons

  p-sub level - 3 orbitals so, 6 electrons.

  d-sub level - 5 orbitals so, 10 electrons.

  f-sub level - 7 orbitals so, 14 electrons.

  g-sub level - 9 orbitals so, 18 electrons.

HUND'S RULE OF MAXIMUM MULTIPLICITY:

- Pairing of electrons take place in a sub-shell only after all the orbitals are singly filled.
- In p sub - level pairing of electrons take place with 4^{th} electron according to Hund's rule.
- Similarly pairing of electrons starts with 6^{th} electron and 8^{th} electron in d and f sub - levels respectively.
- Nitrogen, phosphorus, arsenic, antimony and bismuth atoms have 3 unpaired electrons in their p orbitals because of Hund's rule.
- When all the orbitals of the given sub- level are filled with one electron each (half filled) or two electrons each (completely filled) that electronic configuration gets extra stability due to symmetry.
- The order of filling the different orbitals in a sub-energy level is governed by Hund's rule.

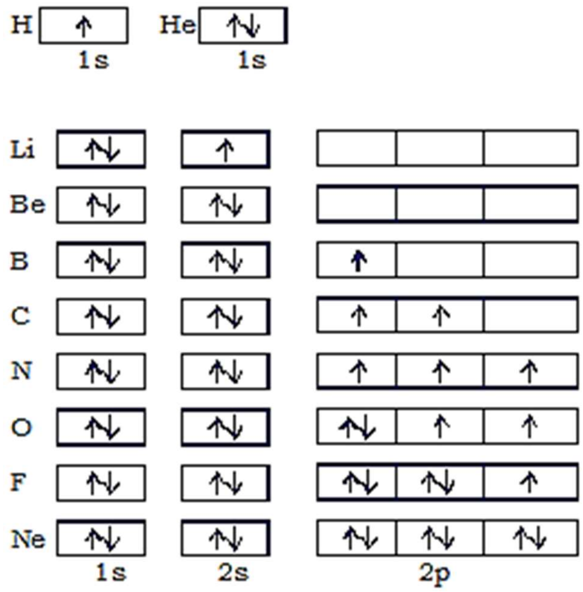
Ex. The electronic configuration of carbon can be written as $1s^2, 2s^2, 2p_x^1, 2p_y^1$

In the later notation the arrows indicate electrons with spins +1/2 and -1/2. Similarly, consider the nitrogen atom. It has 7 electrons. The first six electrons have the same arrangement as that of carbon atom $1s^2$, $2s^2$, $2p_x^1, 2p_y^1$.

The seventh electron occupies the vacant $2p_Z$ orbital but not paired in $2p_X$ or $2p_Y$ orbital. Thus, its electronic configuration is $1s^2$, $2s^2$, $2p_x^1, 2p_y^1, 2p_z^1$

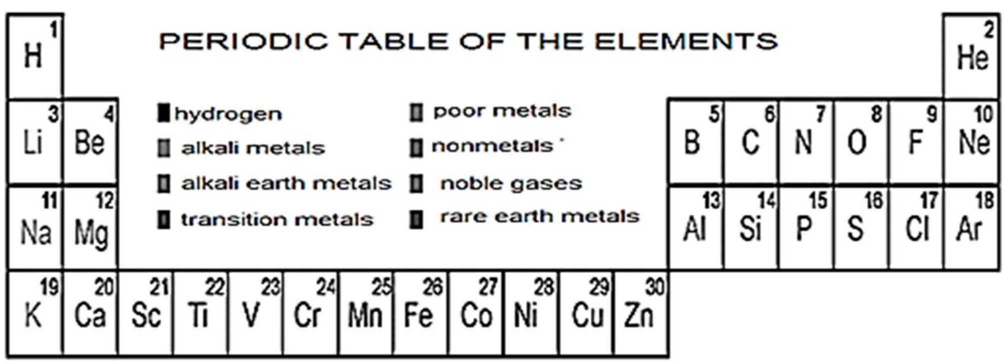
Now, consider the oxygen atom (Z=8). The eighth electron will be paired with one of the three electrons present in 2p orbitals. Its electronic configuration is $1s^2$,$2s^2$, $2p_x^1, 2p_y^1, 2p_z^1$

The advantage of second notation over the first is that it represents all the four quantum numbers. The electronic configuration of elements from hydrogen ($1s^1$), helium ($1s^2$), lithium ($1s^2\ 2s^1$), beryllium ($1s^2\ 2s^2$), boron ($1s^2, 2s^2, 2p^1$), carbon ($1s^2\ 2s^2\ 2p^2$), nitrogen ($1s^2\ 2s^2\ 2p^3$), oxygen ($1s^2\ 2s^2\ 2p^4$), fluorine ($1s^2\ 2s^2\ 2p^5$), neon ($1s^2\ 2s^2\ 2p^6$) is as follows :



ELECTRONIC CONFIGURATION

The distribution of electrons into different levels, sublevels and orbitals of an atom is called electronic configuration.



Generally, it is denoted by nl^x notation,

Where $n \rightarrow$ principal quantum number

$l \rightarrow$ Azimuthal quantum number

$x \rightarrow$ No. of electrons in the orbital

Group number can be predicated by the following rule:

- If the last shell contains 1 or 2 electrons, then the group number is 1 or 2 respectively.
- If the last shell contains 3 or more than 3 electrons, then the group number is the total number of electrons in the last shell plus 10.

ELEMENT	At. No	SHELL CONFIGURATION	ELECTRONIC CONFIGURATION
H	1	1	$1s^1$
He	2	2	$1s^2$
Li	3	2,1	[He]$2s^1$
Be	4	2,2	[He]$2s^2$
B	5	2,3	[He]$2s^2 2p^1$
C	6	2,4	[He]$2s^2 2p^2$
N	7	2,5	[He]$2s^2 2p^3$
O	8	2,6	[He]$2s^2 2p^4$
F	9	2,7	[He]$2s^2 2p^5$
Ne	10	2,8	[He]$2s^2 2p^6$
Na	11	2,8,1	[Ne]$3s^1$
Mg	12	2,8,2	[Ne]$3s^2$
Al	13	2,8,3	[Ne]$3s^2 3p^1$
Si	14	2,8,4	[Ne]$3s^2 3p^2$
P	15	2,8,5	[Ne]$3s^2 3p^3$
S	16	2,8,6	[Ne]$3s^2 3p^4$
Cl	17	2,8,7	[Ne]$3s^2 3p^5$
Ar	18	2,8,8	[Ne]$3s^2 3p^6$
K	19	2,8,8,1	[Ar]$4s^1$
Ca	20	2,8,8,2	[Ar]$4s^2$
Sc	21	2,8,9,2	[Ar]$4s^2\ 3d^1$
Ti	22	2,8,10,2	[Ar]$4s^2\ 3d^2$
V	23	2,8,11,2	[Ar]$4s^2\ 3d^3$
Cr	24	2,8,13,1	[Ar]$4s^1 3d^5$(anomalous configuration)*
Mn	25	2,8,13,2	[Ar]$4s^2\ 3d^5$
Fe	26	2,8,14,2	[Ar]$4s^2\ 3d^6$
Co	27	2,8,15,2	[Ar]$4s^2\ 3d^7$
Ni	28	2,8,16,2	[Ar]$4s^2\ 3d^8$
Cu	29	2,8,18,1	[Ar]$4s^1 3d^{10}$(anomalous configuration)**
Zn	30	2,8,18,2	[Ar]$4s^2\ 3d^{10}$

ANOMALOUS CONFIGURATION:

Case-1: Cr has the valence configuration $4s^1 3d^5$ instead of $4s^2 3d^4$. This is because of the extra stability of a half-filled d-orbital. This is achieved by taking an electron from 4s-orbital.

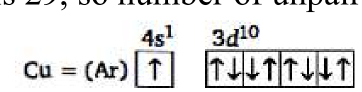
Case-2: In the case of Cu the valence configuration is $4s^1 3d^{10}$ instead of $4s^2 3d^9$. This is because of the extra stability of a fully filled d-orbital. This is achieved by taking an electron from 4s-orbital.

Ex. According to Aufbau's principle, which of the three 4d, 5p and 5s will be filled with electrons first?

Sol. According to the Aufbau's principle, electron will first enters in those orbital which have least energy. So, decreasing order of energy is 5p> 4d > 5s.

Ex. Total number of unpaired electrons in an atom of atomic number 29 is:

Sol. Atomic number of Cu is 29, so number of unpaired electrons is 1.



6. ARRANGEMENT OF ELECTRONS IN AN ORBITALS

WORK SHEET LEVEL–I MAINS CORNER

SINGLE CORRECT ANSWER TYPE QUESTIONS

AUFBAU PRINCIPLE

1. The ground state term symbol for an electronic state is governed by:
   1) Heisenberg's principle 2) Hund's rule
   3) Aufbau principle 4) Pauli exclusion principle
2. The increasing order of energy of the orbitals 1s, 2s and 2p is:
   1) 2p < 2s < 1s 2) 2s < 2p < 1s 3) 1s < 2s < 2p 4) 2p > 3d > 4s
3. According to Aufbau principle, the 19th electron in an atom goes into the:
   1) 4s-orbital. 2) 3d-orbital. 3) 4p-orbital. 4) 3p-orbital.

PAULI'S EXCLUSION PRINCIPLE

4. In a given atom no two electrons can have the same values for all the four quantum numbers. This is called:
   1) Hund's rule 2) Aufbau's principle
   3) Uncertainty principle 4) Pauli's exclusion principle
5. Quantum numbers of an atom can be defined on the basis of:
   1) Hund's rule. 2) Aufbau's principle.
   3) Pauli's exclusion principle. 4) Heisenberg's uncertainty principle

HUND'S RULE

6. Nitrogen has the electronic configuration $1s^2, 2s^2 2p_x^1 2p_y^1 2p_z^1$ and not $1s^2, 2s^2 2p_x^2 2p_y^1 2p_z^0$ which is determined by:
   1) Aufbau's principle 2) Pauli's exclusion principle
   3) Hund's rule 4) Uncertainty principle
7. Which electronic configuration for oxygen is correct according to Hund's rule of multiplicity?
   1) $1s^2, 2s^2 2p_x^2 2p_y^1 2p_z^1$ 2) $1s^2, 2s^2 2p_x^2 2p_y^2 2p_z^0$
   3) $1s^2, 2s^2 2p_x^3 2p_y^1 2p_z^0$ 4) None of these

ELCTRONIC CONFIGURATION

8. The distribution of electrons into different levels, sublevels and orbital's of an atom is called:
   1) Aufbau principle 2) Electronic configuration
   3) Hund's rule 4) Pauli's exclusion principle
9. The total number of unpaired electrons in d- orbitals of atoms of element of atomic number 29 is:
   1) 10 2) 1 3) 0 4) 5

10. The number of electrons in the valence shell of calcium is:
1) 6 2) 8 3) 2 4) 4
11. The electronic configuration of copper $({}_{29}Cu)$ is:
1) $1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^9, 4s^2$ 2) $1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^{10}, 4s^1$
3) $1s^2 . 2s^2 2p^6, 3s^2 3p^6, 4s^2 4p^6$ 4) $1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^{10}$
12. The electronic configuration of an element with atomic number 7 *i.e.* nitrogen atom is:
1) $1s^2, 2s^1, 2p_x^3$ 2) $1s^2, 2s^2 2p_x^2 2p_y^1$
3) $1s^2, 2s^2 2p_x^1 2p_y^1 2p_z^1$ 4) $1s^2, 2s^2 2p_x^1 2p_y^2$
13. Which one of the following configuration represents a noble gas:
1) $1s^2, 2s^2 2p^6, 3s^2$ 2) $1s^2, 2s^2 2p^6, 3s^1$
3) $1s^2, 2s^2 2p^6$ 4) $1s^2, 2s^2 2p^6, 3s^2 3p^6, 4s^2$

LEVEL–II

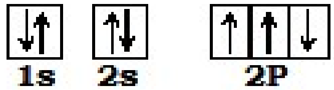
AUFBAU PRINCIPLE

14. The ground state term symbol for an electronic state is governed by:
1) Heisenberg's principle. 2) Hund's rule.
3) Aufbau principle. 4) Pauli's exclusion principle
15. According to Aufbau principle, the correct order of energy of 3d, 4s and 4p orbital is:
1) 4s < 3s < 4p 2) 4s < 3d < 4p 3) 3d < 4s < 4p 4) 4p > 4s < 3d
16. A new electron enters the orbital when:
1) $(n+l)$ is momentum 2) $(n+l)$ is minimum
3) (n + m) is maximum 4) (n + m) is minimum
17. After filling of 4d - orbital, an electron will enter in:
1) 4s 2) 4p 3) 4f 4) 5p

PAULI'S EXCLUSION PRINCIPLE

18. Pauli's exclusion principle states that:
1) Two electrons in the same atom can have the same energy
2) Two electrons in the same atom cannot have the same spin
3) The electrons tend to occupy different orbitals as far as possible
4) None of the above

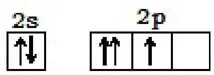
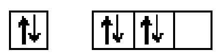
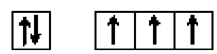
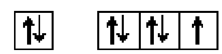
19. In the electronic configuration given below, which rule is violated?



1) Aufbau rule
2) Pauli's exclusion principle
3) Hund's rule
4) The configuration is correct

HUND'S RULE

20. The explanation for the presence of three unpaired electrons in the nitrogen atom can be given by:
1) Pauli's exclusion principle
2) Hund's rule
3) Aufbau's principle
4) Uncertainty principle

21. The orbital diagram, in which both Pauli's exclusion principle and Hund's rule are violated.
1)  2)  3)  4) 

22. Hund's rule deals with the distribution of electrons in:
1) A principal shell.
2) Different subshells.
3) Orbitals with slightly different energies.
4) Degenerate orbitals.

ELCTRONIC CONFIGURATION

23. Electronic configuration of H^- is:
1) $1s^0$
2) $1s^1$
3) $1s^2$
4) $1s^1 2s^1$

24. An element has the electronic configuration $1s^2, 2s^2 2p^6, 3s^2 3p^2$. Its valency electrons are:
1) 6
2) 2
3) 3
4) 4

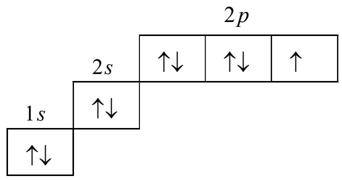
LEVEL-III ADVANCED CORNER

SINGLE CORRECT ANSWER TYPE QUESTIONS

25. Chromium has the electronic configuration $4s^1 3d^5$ rather than $4s^2 3d^4$ because:
1) $4s$ and $3d$ have the same energy
2) $4s$ has a higher energy than $3d$
3) $4s^1$ is more stable than $4s^2$
4) $4s^1 3d^5$ half-filled is more stable than $4s^2 3d^4$

26. Which electronic configuration is not observing the (n+*l*) rule?
1) $1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^1, 4s^2$
2) $1s^2, 2s^2 sp^6, 3s^2 3p^6 3d^7, 4s^2$
3) $1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^5, 4s^1$
4) $1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^8, 4s^2$

27. Which element is represented by the following electronic configuration?



1) Nitrogen 2) Oxygen 3) Fluorine 4) Neon

28. If n + *l* = 6, then total possible number of subshells would be:

1)3 2)4 3)2 4)5

29. The electronic configuration of an element is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$. This represents its:

1) Excited state 2) Ground state 3) Cationic form 4) Anionic form

LEVEL–IV

STATEMENT TYPE QUESTIONS

30. Statement I: 1s orbital possesses lower energy than 2s orbital.
Statement II: Pauli's exclusion principle states that an orbital can have maximum of two electrons, and these must have opposite spins.

1) Both statements are true. 2) Both statements are false.
3) Statement I is true, statement II is false
4) Statement I is false, statement II is true

31. Statement I: Energies of the orbital's increases in the order 1s,2s,2p,3s,3p,4s,3d and 4p.
Statement II: Maximum spin multiplicity means that the total spin of unpaired electrons is maximum.

1) Both statements are true. 2) Both statements are false.
3) Statement I is true, statement II is false.
4) Statement I is false, statement II is true

MULTI CORRECT ANSWER TYPE QUESTIONS

32. The number of d – electrons in Fe^{2+} is equal to that of the:

1) p – electrons in Ne 2) s – electrons in Na
3) d – electrons in Fe 4) s – electrons in Cl^- ion

33. Which of the following has spherical shell of electron?

1) Li 2) Be 3) He 4) B

LEVEL–V

COMPREHENSION TYPE QUESTIONS

PASSAGE:

Orbitals of the same kind should be half filled before electron pairing takes place. Orbitals having the same values for n and *l* are called degenerate orbitals

34. An atom 'Cr' has one 4s electron and five 3d electrons. How many unpaired electrons would be in Cr^{+3}?

1) 1 2) 2 3) 3 4) 4

35. In potassium, the order of energy levels is:

1) 3s > 3d 2) 4s < 3d 3) 4s < 4p 4) 4s = 3d

36. The electronic configuration of phosphorus is $1s^2 2s^2 2p^6 3s^2 3p_x^1 3p_y^1 3p_z^1$. This is accordance with:

1) Aufbau principle 2) Pauli's principle

3) Hund's rule 4) All of these

MATRIX MATCH TYPE QUESTIONS

37.

COLUMN – I	COLUMN – II $(n+l)$ value
a) 3s	p) 6
b) 4p	q) 8
c) 4d	r) 3
d) 5f	s) 5
	t) 4