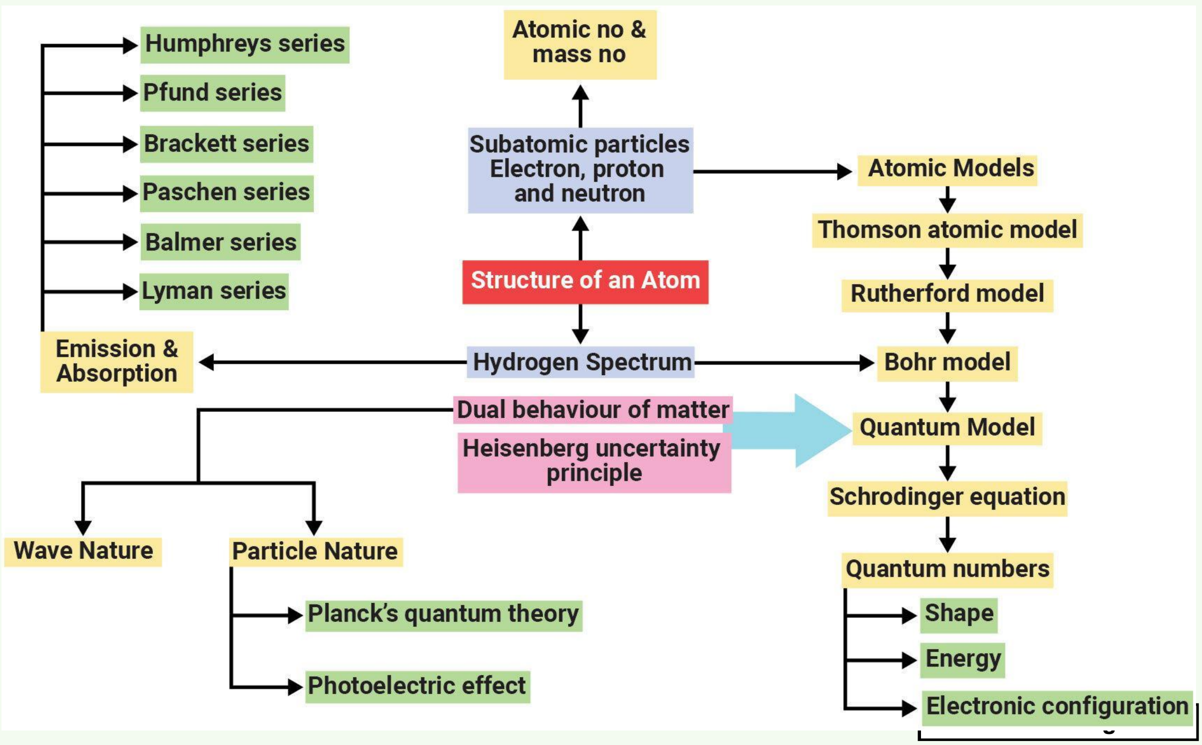
Chapter
1

ATOMIC STRUCTURE

Neils Bohr was a Danish Physicist who made foundational contributions to understanding atomic structure and quantum theory, for which he received the Nobel Prize in Physics in 1922. Bohr developed a model of the atom, in which he proposed that energy levels of electrons are discrete. Although the Bohr model has been supplanted by other models, its underlying principles remain valid.

CONCEPT MAP



CONCEPT 1.1

Introduction: (Recapitulation)

- We have learnt that atoms are the fundamental building blocks of matter.
- In 600 B.C.'s Indian Philosopher and saint **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).
- Around 400 B.C, Greek philosopher **Democritus** proposed that all matter is composed of tiny, discrete, indivisible particles that he called 'atoms', meaning indivisible. His ideas based entirely on philosophical speculation rather than experimental evidence.
- But experiments in late 1800s and early 1900s have shown that atoms are composed of three principal kinds of subatomic particles: **electrons**, **protons** and **neutrons**. Experiments also revealed that at the centre of the atom, there exists a very tiny, extremely dense core called the nucleus, which is containing protons and neutrons collectively called **nucleons**. The electrons surround the nucleus and fill the remaining volume of the atom.

Subatomic Particles:

- Electron: a negatively charged particle (-1.602×10^{-19} C) with negligible mass (9.11×10^{-28} g).
- Proton: a positively charged particle ($+1.602\times10^{-19}$ C) with one unit mass (1.67×10^{-24} g).
- Neutron: a sub-atomic particle with no charge and with one unit mass (1.67×10^{-24} g).
- Protons and neutrons present in the nucleus and the 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**.

Electromagnetic Wave Theory:

This theory was put forward by **James Clark Maxwell** in 1864. According to this theory:

- The energy emitted from any source (like the heated metal rod or the filament of bulb through which electric current is passed) continuously in the form of radiation and it is called electromagnetic radiation.

- 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:

  i. Wavelength (λ)　　ii. Frequency (ϑ)

  iii. Velocity (c)　　iv. Wave number ($\bar{\vartheta}$)

  v. Amplitude (a)　　vi. Time period (T)
- Some relations between characteristics of electromagnetic radiations can be taken as following for calculation purposes:

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

- Electromagnetic wave theory could not explain the phenomena of photoelectric effect and black body radiation.

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:

1. The radiant energy absorbed or emitted by the substances is not continuously but discontinuously in the form of small packets.
2. The smallest packet of energy is called quantum. For light, quantum energy is termed as photon.
3. The radiation is propagated in the form of waves. The energy of a quantum is directly proportional to the frequency of the radiation. $E \propto \vartheta$.
4. 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.

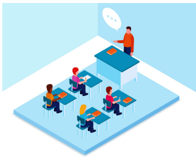
   E = Energy in ergs or Jouls,

   c = Velocity of light = 3×10^{10} cm/sec = 3×10^{8} m/sec,

   ϑ = Frequency of radiation, λ = Wave length, $\bar{\vartheta}$ = Wave number.

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

   i.e. $E = n(h\vartheta)$



CLASSROOM DISCUSSION QUESTIONS

CDQ 01

1. Who proposed the idea that matter is made up of tiny particles called 'paramanus' (atoms)?

(A) Democritus

(B) Maharshi Kanad

(C) J.J. Thomson

(D) James Clark Maxwell

2. What are the three principal kinds of subatomic particles found in atoms?

(A) Electrons, neutrons, photons

(B) Protons, electrons, molecules

(C) Electrons, protons, neutrons

(D) Electrons, positrons, neutrons

3. Which scientist proposed the plum pudding model of the atom?

(A) Democritus

(B) J.J. Thomson

(C) Max Planck

(D) James Clark Maxwell

4. According to electromagnetic wave theory, which property characterizes the energy emitted from any source?

(A) Mass (B) Velocity

(C) Charge (D) Wavelength

5. What is the smallest packet of energy called according to Planck's quantum theory?

(A) Photon (B) Quantum

(C) Electron (D) Neutron

6. What property of electromagnetic radiations is directly proportional to the energy of a quantum according to Planck's quantum theory?

(A) Wavelength (B) Frequency

(C) Amplitude (D) Velocity

7. What is the numerical value of Planck's constant?

(A) 6.625×10^{27} erg-sec

(B) 6.625×10^{34} Joule-sec

(C) 3×10^{8} cm/sec

(D) 3×10^{10} m/sec

8. What phenomena could electromagnetic wave theory not explain, leading to the development of Planck's quantum theory?

(A) Photoelectric effect and black body radiation

(B) Atomic structure and nuclear decay

(C) Radioactive decay and nuclear fusion

(D) Ionization and nuclear fission

9. According to Planck's quantum theory, what is the energy of a quantum expressed as?

(A) $E = hc/\lambda$ (B) $E = h\lambda$

(C) $E = n(h\lambda)$ (D) $E = h\lambda$

10. In Planck's quantum theory, what term is used for the smallest packet of energy for light?

(A) Quark (B) Photon

(C) Neutrino (D) Boson

MARK YOUR ANSWERS WITH PEN ONLY. Time Taken in Minutes

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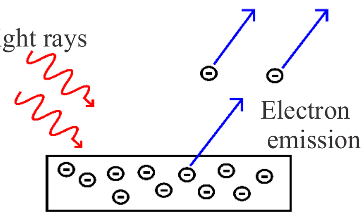
CONCEPT 1.2

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**.

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, caesium 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

The results observed in this experiment were:

1. The electrons are ejected from the clean metal surface as soon as the beam of light strikes the surface, i.e. there is no time lag between the striking of light beam and the ejection of electrons from the metal surface.
2. The number of electrons ejected is proportional to the intensity or brightness of light.
3. For each metal, there is a characteristic minimum frequency, ϑ_0 (also known as **Threshold frequency**) below which photoelectric effect is not observed. At a frequency $\vartheta > \vartheta_0$ the ejected electrons come out with certain kinetic energy. The kinetic energies of these electrons increase with the increase of frequency of the light used.

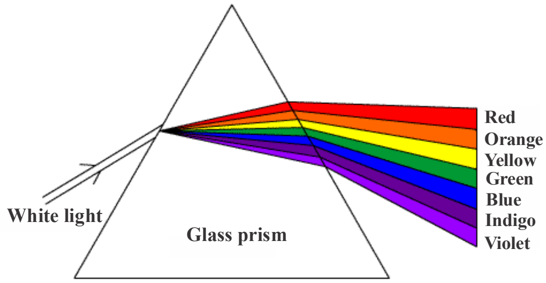
If the striking photon has energy equal to $h\vartheta$ and the minimum energy required to eject the electron is $h\vartheta_0$ (also called **work function** W_0), then the difference in energy $(h\vartheta - h\vartheta_0)$ is transferred as the kinetic energy of the photoelectron. Following the conservation of energy principle, the kinetic energy of the ejected electron is given by the equation $h\vartheta = h\vartheta_0 + \frac{1}{2}m_e v^2$ where m_e is the mass of the electron and v is the velocity associated with the ejected electron.

Note:

Cesium is the best used metal in photoelectric cells as it has very less ionisation energy.

Spectrum:

The light that we see from sun or from an incandescent filament lamp is white light. A white light is composed of seven component colours namely violet, indigo, blue, green, yellow, orange and red (VIBGYOR) as shown in fig.



The splitting of white light into its component seven colours is known as dispersion.

The array of seven colours extending from red at one end to violet at the other end similar to rainbow colours is called **spectrum.**

Atoms, molecules or ions that have absorbed radiation (energy) are said to be excited. The radiation emitted by these substances form a spectrum. The light emitted by a sample of excited atoms is called **atomic spectrum.**

A sample of a gas contains an enormous number of atoms. Although a single atom can be in only one excited state at a time, the collection of atoms contains all possible excited states. The light emitted as these atoms fall to lower energy states is responsible for the 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 in visible 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

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

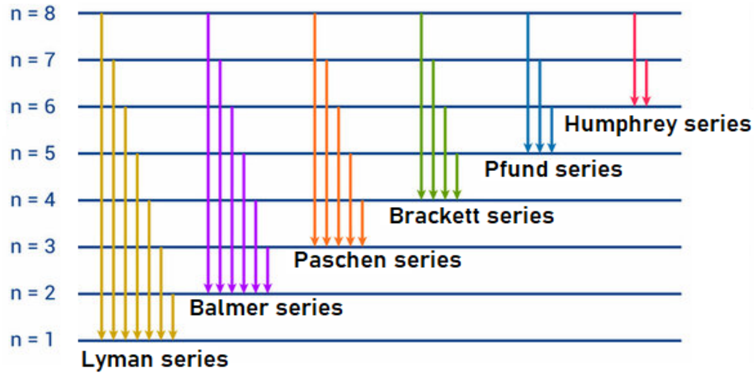
Where n is the number of the higher orbit from which electron jumped to second orbit i.e., n = 3,4,5,6...., and R is a constant called **Rydberg constant**. The value of Rydberg constant for hydrogen, R_H = 109677 cm^{-1} or 1.09677×10^7 m^{-1}.

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

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

Paschen, Brackett and Fund 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.



Rydberg gave a more general expression which can be applied to all the series of the hydrogen spectrum. This expression is called Rydberg's formula.

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

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 higher orbit from which electron jumped.

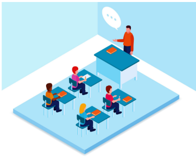
For one electron species like He^+, Li^{2+} and Be^{3+}, the value of R is $109677 \times Z^2$ cm^{-1}, where Z is the atomic number of the species.

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

$$\bar{\vartheta} = Z^2 R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) \text{ cm}^{-1}\text{, where Z is atomic number of the species.}$$

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	$\bar{\vartheta} = R\left[\frac{1}{5^2} - \frac{1}{n_2^2}\right]$



CLASSROOM DISCUSSION QUESTIONS

CDQ 02

1. What is a black body?

(A) A body that reflects all incident radiation

(B) A body that transmits all incident radiation

(C) A body that absorbs all incident radiation

(D) A body that emits all incident radiation

2. Who performed the experiment leading to the discovery of the photoelectric effect?

(A) J.J. Thomson

(B) James Clerk Maxwell

(C) H. Hertz

(D) Max Planck

3. Which metal is commonly used in photoelectric cells due to its low ionization energy?

(A) Gold (B) Silver

(C) Cesium (D) Iron

4. What is the characteristic minimum frequency required for the photoelectric effect to occur called?

(A) Ionization frequency

(B) Threshold frequency

(C) Emission frequency

(D) Work function frequency

5. What is the phenomenon called when white light is separated into its component colors?

(A) Dispersion

(B) Absorption

(C) Reflection

(D) Diffraction

6. What is emitted by excited atoms and forms a spectrum?

(A) Electrons (B) Protons

(C) Photons (D) Neutrons

7. What is the simplest atomic spectrum?

(A) Helium spectrum

(B) Oxygen spectrum

(C) Hydrogen spectrum

(D) Nitrogen spectrum

8. Who discovered the series of lines in the hydrogen spectrum named after him?

(A) Balmer (B) Lyman

(C) Paschen (D) Brackett

9. What is the expression for the wave number in Rydberg's formula for hydrogen spectrum?

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

(B) $\frac{1}{\lambda} = R\left(1 - \frac{1}{n^2}\right)$

(C) $\frac{1}{\lambda} = \frac{R}{n^2}$

(D) $\frac{1}{\lambda} = R\left(\frac{1}{n_2} - \frac{1}{n_1}\right)$

10. What does the value of R represent in Rydberg's formula for hydrogen spectrum?

(A) Speed of light

(B) Atomic number

(C) Rydberg constant

(D) Work function

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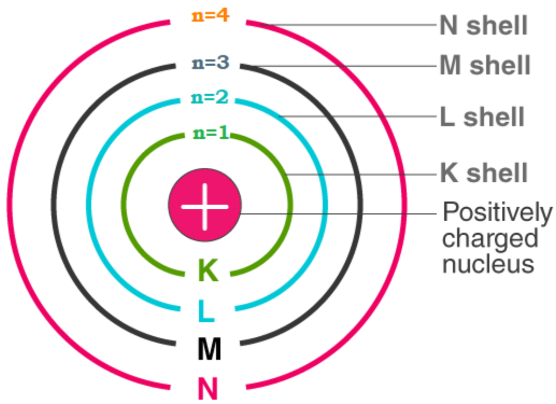
CONCEPT 1.3

Bohr's Model of Atom:



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

1. Electrons move around the nucleus in specified circular paths called orbits or shells or energy levels.

2. 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.

3. 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.

4. 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.

5. 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 (ΔE) 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 $h\vartheta$. $E_2 - E_1 = h\vartheta$

6. 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 of the electron in the n^{th} orbit of any single electron species like Hydrogen atom (Z = 1) as follows:

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

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

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

Limitations of Bohr's atomic model:

1. Bohr's model could not explain the spectra of atoms containing more than one electron.

2. It could not explain the **zeeman effect**. In presence of magnetic field, each spectral line gets split up into fine lines, the phenomenon, is known as zeeman effect.

3. It could not explain the **stark effect**. In presence of electric field, each spectral line gets split up into fine lines, the phenomenon, is known as stark effect.

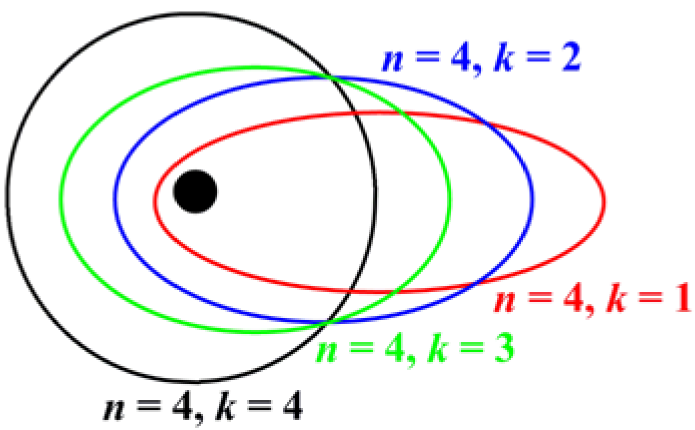
4. The main objection to Bohr's model was raised by Heisenberg's uncertainty principle. The Bohr atomic model theory considers electrons to have both a known radius and orbit i.e. known position and momentum at the same time, which is impossible according to Heisenberg.

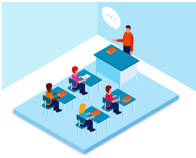
Sommerfeld's Extension of Bohr's Theory:

Arnold Sommerfeld, in 1915, extended Bohr theory and gave his postulates. According to him, the stationary orbits in which electrons are revolving around the nucleus in the atom are not circular but elliptical in shape. It is due to the influence of the centrally located nucleus.

The electron revolves in elliptical path with nucleus at one of its foci. So, there will be a major and a minor axis of the path.

He concluded that, the first orbit is circular in nature. The second orbit has two sub orbits, one is elliptical, and the other is circular in nature. Third orbit has three sub orbits, two are elliptical and one is circular in nature.





CLASSROOM DISCUSSION QUESTIONS

CDQ 03

1. According to Bohr's model of the atom, electrons move around the nucleus in circular paths called:
(A) Orbits
(B) Shells
(C) Energy levels
(D) All of the above

2. Which principle states that the electron can only revolve in an orbit where its angular momentum is quantized?
(A) Planck's principle
(B) Uncertainty principle
(C) Quantization of energy
(D) Quantization of angular momentum

3. In Bohr's model, the energy associated with each energy level increases with:
(A) Decrease in distance from the nucleus
(B) Increase in distance from the nucleus
(C) The size of the nucleus
(D) The number of electrons in the atom

4. What equation describes the energy emitted or absorbed by an electron when it jumps between energy levels in Bohr's model?
(A) Einstein's equation
(B) Planck's equation
(C) Newton's equation
(D) Maxwell's equation

5. According to Bohr's model, the radius of the nth orbit of a single electron species like hydrogen is proportional to:
(A) n (B) n^2
(C) 1/n (D) n^3

6. What is the energy of an electron in the nth orbit of a single electron species like hydrogen according to Bohr's model?
(A) $\frac{13.6}{n}$eV (B) 13.6eV
(C) 13.6eV×n^2 (D) $\frac{13.6\times n^2}{Z}$eV

7. What phenomenon splits spectral lines into fine lines in the presence of a magnetic field?
(A) Zeeman effect
(B) Stark effect
(C) Bohr effect
(D) Photoelectric effect

8. What is one limitation of Bohr's atomic model?
(A) Inability to explain the spectra of multi-electron atoms
(B) Inability to explain the photoelectric effect
(C) Inability to describe the dual nature of electrons
(D) Inability to predict the behavior of light

9. Sommerfeld's extension of Bohr's theory suggested that orbits in which electrons revolve are:
(A) Circular (B) Elliptical
(C) Spiral (D) Parabolic

10. In Sommerfeld's extension, which orbit(s) are elliptical in shape?
(A) First orbit (B) Second orbit
(C) Third orbit (D) Both (B)&(C)

MARK YOUR ANSWERS WITH PEN ONLY. Time Taken in Minutes

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CONCEPT 1.4

Dual Nature (Particle and Wave Nature) of Electron:

De-Broglie's Wave Equation:

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

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

According to Planck's quantum theory, energy of photon is given by

$$E = h\vartheta = h\frac{c}{\lambda} \qquad \text{........................ (1)}$$

According to Einstein's mass - energy relationship, energy of a photon is given by $E = mc^2$ .... (2)

On equating both the equations (1) & (2), we get $h\frac{c}{\lambda} = mc^2$

$$\Rightarrow \quad \lambda = \frac{h}{mc} = \frac{h}{p} \quad [\because \text{mass(m)} \times \text{velocity(c)} = \text{momentum(p)}]$$

Where h = Planck's constant, m = mass of the electron, c = velocity of the electron and p = momentum of the electron.

Heisenberg's Uncertainty Principle:

In 1927, **Werner Heisenberg** presented a principle known as Heisenberg's uncertainty principle which states that: "It is impossible to measure simultaneously, the exact position and exact momentum of a body as small as an electron. If the uncertainty (error) in the measurement of position of a particle moving with high velocity is 'Δx', and the uncertainty in the measurement of momentum is 'Δp' or 'Δmv', then these two are related according to Heisenberg's relationship as:
$\Delta x.\Delta p \geq h/4\pi$ or

$$\Delta x.m\Delta v \geq h/4\pi \qquad (\because p = mv)$$

Where, h is Planck's' constant

Quantum Numbers:

As we know, to search a particular person in this world, four things are needed:

1. The country to which the person belongs.
2. The city in that country to which the person belongs.
3. The street in that city where the person is residing.
4. The house number.

Similarly, four identification numbers are required to locate a particular electron in an atom. These set of identification numbers are called **quantum numbers.** The four quantum numbers are discussed below.

I. Principal Quantum Number:

It was given by **Bohr,** it is denoted by '**n**'. It represents the name, size and energy of the shell to which the electron belongs. The value of 'n' lies between 1 to ∞. $n = 1, 2, 3, 4, \ldots\ldots \infty$

Value of n	=	1	2	3	4	5	6	7
Designation of shell	=	K	L	M	N	O	P	Q

Higher is the value of 'n', greater is the distance of the shell from the nucleus, greater is the magnitude of energy of the electron in that orbit.

Energy separation between two shells decreases on moving away from the nucleus.

$(E_2 - E_1) > (E_3 - E_2) > (E_4 - E_3) > (E_5 - E_4)\ldots$

Maximum number of electrons in a shell = $2n^2$

II. Azimuthal Quantum Number:

It was given by **Sommerfeld.** It is also called angular quantum number or secondary quantum number**.** It is denoted by 'l'. Its values lie between 0, 1, 2, ... , (n – 1) where n is principle quantum number.

It describes the shape and name of the sub shell associated with the main shell. $l = 0 \rightarrow$s–sub shell; $l = 1 \rightarrow$p–sub shell; $l = 2 \rightarrow$d–sub shell;

$l = 3 \rightarrow$f– sub shell; $l = 4 \rightarrow$g – sub shell.

s, p, d, f and g are spectral terms and signify sharp, principal, diffused, fundamental and generalized respectively. The energies of the various sub shells in the same shell are in the order of $s < p < d < f < g$ (increasing order).

Maximum electrons present in a sub shell = $2(2l + 1)$

s – sub shell $\rightarrow$ 2 electrons $\quad$ p – sub shell $\rightarrow$ 6 electrons

d – sub shell $\rightarrow$ 10 electrons $\quad$ f – sub shell $\rightarrow$ 14 electrons

g – sub shell $\rightarrow$ 18 electrons

III. Magnetic Quantum Number:

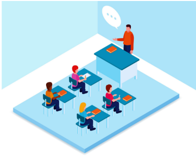
It was given by **Lande** and is designated by the symbol '**m**'. It describes the orientation or distribution of electron cloud. For each value of 'l', the magnetic quantum number 'm' may assume all integral values from '$-l$' to '$+l$' including zero, i.e., total $(2l + 1)$ values.

Thus, when $l = 0$, $m = 0$ (only one value) *i.e.*, only one orientation.

When $l = 1$, $m = -1, 0, +1$ (three values) *i.e.*, three orientations.

Each orientation corresponds to one orbital.

Thus, the number of 'm' values indicates the number of orbitals in a subshell with a particular 'l' value.



CLASSROOM DISCUSSION QUESTIONS

CDQ 04

1. De Broglie's wave equation proposed that electrons behave both as particles and as waves, and he derived an expression for the wavelength of the wave associated with the electron using:

(A) Einstein's equation
(B) Planck's equation
(C) Newton's equation
(D) Heisenberg's equation

2. According to Heisenberg's uncertainty principle, if the uncertainty in the measurement of position of a particle is '?x', and the uncertainty in the measurement of momentum is '?p', then these uncertainties are related by:

(A) $\Delta x.\Delta p = \frac{h}{4\pi}$ (B) $\Delta x.\Delta p = \frac{h}{2\pi}$
(C) $\Delta x + \Delta p = \frac{h}{4\pi}$ (D) $\Delta x - \Delta p = \frac{h}{2\pi}$

3. Quantum numbers are required to locate a particular electron in an atom. How many quantum numbers are there?

(A) Two (B) Three
(C) Four (D) Five

4. Which quantum number represents the size and energy of the shell to which the electron belongs?

(A) Azimuthal quantum number
(B) Magnetic quantum number
(C) Principal quantum number
(D) Angular quantum number

5. The maximum number of electrons that can be accommodated in a shell is given by:

(A) $2n^2$ (B) $2(n+1)^2$
(C) n^2 (D) n+2

6. What is the maximum number of electrons present in the 'p' subshell?

(A) 2 electrons
(B) 6 electrons
(C) 10 electrons
(D) 14 electrons

7. How many possible orientations does the magnetic quantum number 'm' have when the azimuthal quantum number 'l' is equal to 1?

(A) 1 (B) 2 (C) 3 (D) 4

8. Which quantum number describes the orientation or distribution of electron cloud?

(A) Principal quantum number
(B) Azimuthal quantum number
(C) Magnetic quantum number
(D) Secondary quantum number

9. The energies of the various subshells in the same shell are in the order of:

(A) s > p > d > f > g
(B) g > f > d > p > s
(C) s < p < d < f < g
(D) f < d < p < s < g

10. For each value of 'l', the magnetic quantum number 'm' may assume all integral values from:

(A) 0 to n (B) –l to +l
(C) –n to +n (D) 0 to +l

MARK YOUR ANSWERS WITH PEN ONLY. Time Taken in Minutes

1 Ⓐ Ⓑ Ⓒ Ⓓ	2 Ⓐ Ⓑ Ⓒ Ⓓ	3 Ⓐ Ⓑ Ⓒ Ⓓ	4 Ⓐ Ⓑ Ⓒ Ⓓ	5 Ⓐ Ⓑ Ⓒ Ⓓ
6 Ⓐ Ⓑ Ⓒ Ⓓ	7 Ⓐ Ⓑ Ⓒ Ⓓ	8 Ⓐ Ⓑ Ⓒ Ⓓ	9 Ⓐ Ⓑ Ⓒ Ⓓ	10 Ⓐ Ⓑ Ⓒ Ⓓ

CONCEPT 1.5

Orbital:

The place or the region where there is finite probability of finding electron around the nucleus is known as orbital.

When $l = 0$, m = 0, i.e., one value implies that 's' sub shell has only one space orientation and hence, it can be arranged in space only in one way along x, y or z axes. Thus, ***s***- orbital has a symmetrical **spherical** shape.

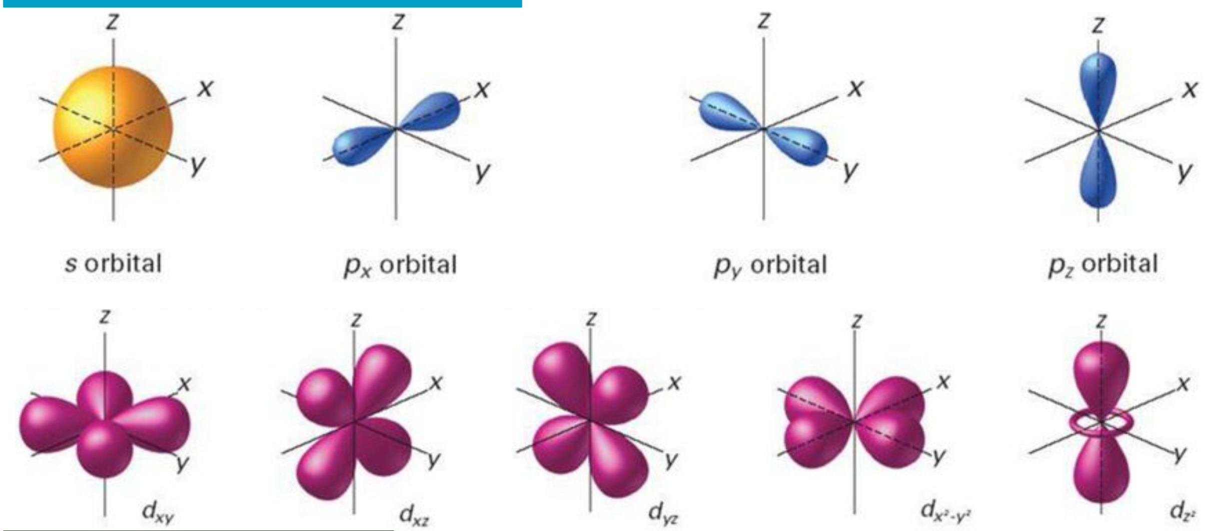
When $l = 1$, 'm' has three values –1, 0, +1. It implies that 'p' sub shell of any energy shell has three space orientations, i.e., three orbitals. Each ***p***–orbital has **dumbbell** shape. Each one is disposed symmetrically along one of the three axes. *p*–orbitals have directional character.

Orbital	p_z	p_x	p_y
m	0	±1	±1

When $l = 2$, 'm' has five values –2, –1, 0, +1, +2. It implies that d–sub shell of any energy shell has five orientations, i.e., five orbitals. They are $d_{xy}, d_{yz}, d_{zx}, d_{x^2-y^2}$ and d_{z^2}. Each *d*-orbital has **double dumbbell** shape.

Orbital	d_{xy}	d_{yz}	d_{zx}	d_{x2-y2}	d_{z2}
m	± 2	± 1	± 1	± 2	0

Outlines of Atomic Orbitals:



Degenerate Orbitals:

Orbitals which are having same energy are called degenerated orbitals.

For example, p_x, p_y and p_z are degenerate orbitals.

$d_{xy}, d_{yz}, d_{zx}, d_{x^2-y^2}$ and d_{z^2} are degenerate orbitals.

All the degenerate orbitals of the same sub shell have same energy but differ in the direction of their space orientation.

Note: Total number of orbitals in a main energy shell is equal to n^2 (but not more than 16 in any of the main shells of the known elements).

Ex: n = 1 No. of orbitals = $(1)^2 = 1$ (1s)

n = 2 No. of orbitals = $(2)^2 = 4$ ($2s, 2p_x, 2p_y, 2p_z$)

n = 3 No. of orbitals = $(3)^2 = 9$ ($3s, 3p_x, 3p_y, 3p_z, 3d_{xy}, 3d_{yz}, 3d_{zx}, 3d_{x^2-y^2}, 3d_{z^2}$)

IV. Spin Quantum Number:

It is denoted by 's' and it was given by "**Uhlenbeck** and **Goudsmit**".

Spin quantum number represents the direction of electron spin around its own axis.

For clockwise spin, $s = +1/2\left(\uparrow \text{arrow respentation}\right)$

For anti-clockwise spin, $s = -1/2\left(\downarrow \text{arrow respentation}\right)$

Electronic configurations of atoms:

The filling of orbitals in an atom is a hypothetical process in which the atom is built up by feeding electrons in orbitals, one at a time and by placing each new electron in the lowest available energy orbital. The systematic distribution of electrons in different orbitals is known as **electronic configuration** of the atoms.

n, l, x notation: This characterizes each electron in an atom.

$n\,l^{x}$

No. of electrons (e.g, 1,2,3...)

Principle Quantum Numbers (e.g, 1,2,.....)

Orbital (e,g, s,p,d, or f)

For example, hydrogen atom having atomic number (Z) = 1, the number of electrons is one, then the electronic configuration is $1s^1$.

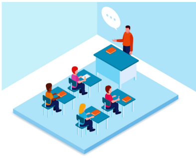
Principle quantum number $\rightarrow 1s^{1} \leftarrow$ No. of electrons in the orbital

Orbital (Azimuthal quantum nuber)

The electronic configuration can be represented by showing the spin of the electrons in boxes. For the electron in H-atom, as we have seen, the set of quantum numbers is:

$n = 1,\ l = 0,\ m = 0,\ s = +1/2.$ $\boxed{\uparrow}$

For many electrons system we must know the electronic configuration of the atom.



CLASSROOM DISCUSSION QUESTIONS

CDQ 05

1. What is the shape of the 's' orbital?
(A) Cubical
(B) Dumb-bell
(C) Tetrahedral
(D) Spherical

2. When 'l' equals 1, how many space orientations does the 'p' subshell have?
(A) One
(B) Two
(C) Three
(D) Four

3. Which of the following orbitals are degenerate?
(A) 1s, 2s, 3s
(B) 2px, 2py, 2pz
(C) 2s, 2p, 3s
(D) 3px, 3py, 3pz

4. How many degenerate orbitals are there when 'l' equals 2?
(A) Two (B) Three
(C) Four (D) Five

5. What is the maximum number of orbitals in a main energy shell?
(A) n (B) n + 1
(C) n^2 (D) $2n^2$

6. What does the spin quantum number 's' represent?
(A) Direction of electron spin around the nucleus
(B) Number of electrons in an orbital
(C) Shape of the orbital
(D) Energy level of the electron

7. How is the clockwise spin represented in terms of the spin quantum number?
(A) s = +1/2
(B) s = –1/2
(C) s = +1
(D) s = –1

8. What does the electronic configuration of an atom represent?
(A) The shape of the atom
(B) The arrangement of electrons in different orbitals
(C) The number of protons in the nucleus
(D) The energy levels of the electrons

9. What does the notation 'n, l, x' represent in electronic configurations?
(A) Number of protons, neutrons, and electrons
(B) Principle quantum number, orbital, and number of electrons
(C) Atomic number, mass number, and energy level
(D) Spin quantum number, azimuthal quantum number, and orbital

10. If the electronic configuration of a hydrogen atom is $1s^1$, what does it represent?
(A) One electron in the s orbital
(B) One electron in the p orbital
(C) Two electrons in the s orbital
(D) Two electrons in the p orbital

MARK YOUR ANSWERS WITH PEN ONLY. Time Taken in Minutes

1 Ⓐ Ⓑ Ⓒ Ⓓ | 2 Ⓐ Ⓑ Ⓒ Ⓓ | 3 Ⓐ Ⓑ Ⓒ Ⓓ | 4 Ⓐ Ⓑ Ⓒ Ⓓ | 5 Ⓐ Ⓑ Ⓒ Ⓓ
6 Ⓐ Ⓑ Ⓒ Ⓓ | 7 Ⓐ Ⓑ Ⓒ Ⓓ | 8 Ⓐ Ⓑ Ⓒ Ⓓ | 9 Ⓐ Ⓑ Ⓒ Ⓓ | 10 Ⓐ Ⓑ Ⓒ Ⓓ

CONCEPT 1.6

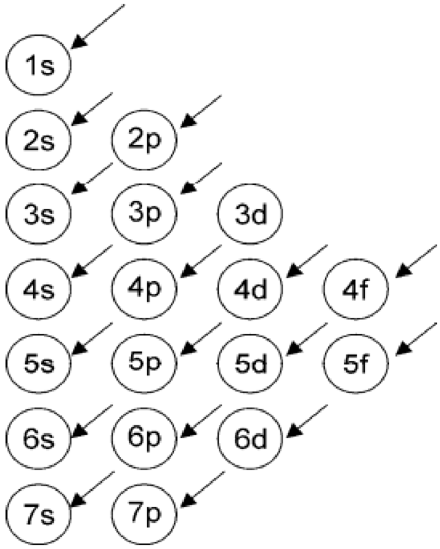
The filling of electrons into different orbitals is governed by the following rules:

Aufbau's Principle:

The word 'Aufbau' in German means 'building up'. Thus, this is also known as building up principle. The building up of orbitals means the filling up or orbitals with electrons. The principle states as follows:

"The electrons in an atom are so arranged that they occupy orbitals in the order of their increasing energy".

In other words, 'electrons first occupy the lowest energy orbitals available to them and enter into higher energy orbitals only when the lowest orbitals are filled'. Since, the energy of an orbital in the absence of any magnetic field, depends upon the principal quantum number (*n*) and the azimuthal quantum number (*l*), hence the order of filling orbitals with electrons may be obtained from the following generalizations:



a) The orbital for which (n + *l*) value is the lowest is filled first.

b) When the orbitals have the same values of (n+*l*) the orbital having the lower value of *n* is filled first.

The sequence in which the electrons occupy various orbitals can be easily remembered with the help of **Moeller's diagram**.

The order of filling of various orbitals with electrons obtained by this rule is given below:

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

Pauli's Exclusion Principle:

In 1925, Wolfgang Pauli proposed exclusion principle. This principle is very useful in construction of the electronic configuration of atom. According to this principle:

"No two electrons in an atom can have the same values for all the four quantum numbers".

For example, in 1s orbital of helium atom there are two electrons. According to the concept of quantum numbers and Pauli's rule, their quantum numbers are:

Electron	*n*	*l*	*m*	*s*
Electron 1	1	0	0	+1/2
Electron 2	1	0	0	-1/2

The '+' and '–' signs before 1/2 refers to the clockwise and anti-clockwise spins of the electrons respectively. Thus, the two electrons having the same values of *n*, *l* and *m* could have different values of *s*, that is, their spins are in opposite directions. This led to a very significant observation that "An orbital can have maximum two electrons and these must have opposite spins".

For example, two electrons in an orbital can be represented by,

Correct : $\boxed{\uparrow\downarrow}$ (i) or $\boxed{\downarrow\uparrow}$ (ii)

Incorrect : $\boxed{\uparrow\uparrow}$ (iii) or $\boxed{\downarrow\downarrow}$ (iv)

In (i) and (ii) representations, the two electrons (each indicated by an arrow) have opposite spins, that is, if one is revolving clockwise, the other is revolving anti clockwise or vice versa. In (iii) and (iv) representations, the two electrons have the same spin, that is, either clockwise or anti-clockwise. In the light of Pauli's exclusion principle, the presentation. (i) and (ii) are correct while (iii) and (iv) are incorrect.

Hund's Rule of Maximum Multiplicity:

This rule deals with the filling of electrons into degenerate (same energy) orbitals of the same subshell. This principle is very important in guiding the filling of electrons in *p, d* and *f* orbitals.

Hund's Rule:

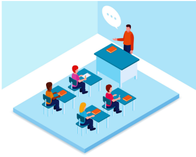
Hund's rule states that **"electron pairing takes place only after all the available degenerate orbitals are occupied by one electron each"**. For example, consider carbon atom (Z=6). It has six electrons. The first electron goes into the '1s' orbital of the K- shell. The second electron will be paired up with the first in the same '1s' orbital. Similarly, the third and fourth electrons occupy the '2s' orbital of the L-shell. The fifth electron goes into one of the three '2p' orbitals of the L-shell. Let it be $2p_x$. Since the three 'p' orbitals (viz. $2p_x, 2p_y, 2p_z$) are degenerate, the sixth electron goes into $2p_y$ or $2p_z$ but not $2p_x$. Thus the electronic configuration of carbon can be written as

$$1s^2 2s^2 2p_x^1 2p_y^1 \quad \text{or} \quad \underset{1s}{\boxed{\uparrow\downarrow}}\ \underset{2s}{\boxed{\uparrow\downarrow}}\ \underset{2p_x}{\boxed{\uparrow}}\underset{2p_y}{\boxed{\uparrow}}\underset{2p_z}{\boxed{}}$$

In the latter notation the arrows indicate electrons with spins + 1/2 and –1/2.

Hund's rule is also known as Hund's rule of maximum multiplicity. This is due to the fact that electrons being identical in charge, repel each other when present in the same orbital. This repulsion can, however, be minimised if two electrons move as far apart as possible by occupying different degenerate orbitals.

Further, all the singly occupied orbitals will have parallel spins, that is, in the same direction, either clockwise or anti-clockwise. This is due to the fact that two electrons with parallel spins (in different orbitals) will encounter less inter-electronic repulsions in space.



CLASSROOM DISCUSSION QUESTIONS CDQ 06

1. What is the principle behind Aufbau's Principle in electron configuration?
(A) Electrons occupy orbitals randomly
(B) Electrons occupy orbitals in order of decreasing energy
(C) Electrons occupy orbitals based on their mass
(D) Electrons occupy orbitals based on their charge

2. According to Aufbau's Principle, which orbital is filled first?
(A) The orbital with the highest value of (n + l)
(B) The orbital with the lowest value of (n + l)
(C) The orbital with the highest value of n
(D) The orbital with the lowest value of n

3. What is the significance of Pauli's Exclusion Principle?
(A) It determines the shape of orbitals
(B) It limits the number of electrons in each orbital
(C) It dictates the order of filling of orbitals
(D) It determines the energy levels of electrons

4. According to Pauli's Exclusion Principle, what cannot be the same for two electrons in an atom?
(A) Their position (B) Their charge
(C) Their spin (D) Their mass

5. How are the spins of two electrons in the same orbital represented according to Pauli's Exclusion Principle?
(A) Both clockwise
(B) Both counterclockwise
(C) One clockwise and one counterclockwise
(D) They cannot be in the same orbital

6. What does Hund's Rule of Maximum Multiplicity state?
(A) Electrons fill orbitals randomly
(B) Electron pairing occurs before orbital occupation
(C) Electron pairing occurs after all orbitals are singly occupied
(D) Electron pairing occurs in the order of increasing energy

7. In carbon (Z=6), how are the last two electrons accommodated according to Hund's Rule?
(A) Both go into the same orbital
(B) One goes into each of the three degenerate orbitals
(C) Both occupy the same subshell
(D) One fills the highest energy orbital first

8. Why is Hund's Rule also known as the Rule of Maximum Multiplicity?
(A) It maximizes the number of electrons in orbitals
(B) It minimizes electron repulsion
(C) It maximizes the energy of the electrons
(D) It determines the maximum number of orbitals

9. How do electrons behave in orbitals according to Hund's Rule?
(A) They repel each other
(B) They attract each other
(C) They occupy different orbitals before pairing up
(D) They fill orbitals randomly

10. What is the consequence of electrons following Hund's Rule in terms of their spins?
(A) They have parallel spins
(B) They have antiparallel spins
(C) They repel each other
(D) They occupy the same orbital

MARK YOUR ANSWERS WITH PEN ONLY. Time Taken in Minutes

1 Ⓐ Ⓑ Ⓒ Ⓓ	2 Ⓐ Ⓑ Ⓒ Ⓓ	3 Ⓐ Ⓑ Ⓒ Ⓓ	4 Ⓐ Ⓑ Ⓒ Ⓓ	5 Ⓐ Ⓑ Ⓒ Ⓓ
6 Ⓐ Ⓑ Ⓒ Ⓓ	7 Ⓐ Ⓑ Ⓒ Ⓓ	8 Ⓐ Ⓑ Ⓒ Ⓓ	9 Ⓐ Ⓑ Ⓒ Ⓓ	10 Ⓐ Ⓑ Ⓒ Ⓓ

CONCEPT 1.7

Electronic configuration of elements up to zinc:

Element	Atomic number	Electronic configuration	
H	1	$1s^1$	$1s^1$
He	2	$1s^1$	$1s^2$
Li	3	$1s^2 2s^1$	$[He]2s^1$
Be	4	$1s^2 2s^2$	$[He]2s^2$
B	5	$1s^2 2s^2 3p^1$	$[He]2s^2 2p^1$
C	6	$1s^2 2s^2 2p^2$	$[He]2s^2 2p^2$
N	7	$1s^2 2s^2 2p^3$	$[He]2s^2 2p^3$
O	8	$1s^2 2s^2 2p^4$	$[He]2s^2 2p^4$
F	9	$1s^2 2s^2 2p^5$	$[He]2s^2 2p^5$
Ne	10	$1s^2 2s^2 2p^6$	$[He]2s^2 2p^6$
Na	11	$1s^2 2s^2 2p^6 3s^1$	$[Ne]3s^1$
Mg	12	$1s^2 2s^2 2p^6 3s^2$	$[Ne]3s^2$
A*l*	13	$1s^2 2s^2 2p^6 3s^2 3p^1$	$[Ne]3s^2 3p^1$
Si	14	$1s^2 2s^2 2p^6 3s^2 3p^2$	$[Ne]3s^2 3p^2$
P	15	$1s^2 2s^2 2p^6 3s^2 3p^3$	$[Ne]3s^2 3p^3$
S	16	$1s^2 2s^2 2p^6 3s^2 3p^4$	$[Ne]3s^2 3p^4$
C*l*	17	$1s^2 2s^2 2p^6 3s^2 3p^5$	$[Ne]3s^2 3p^5$
Ar	18	$1s^2 2s^2 2p^6 3s^2 3p^6$	$[Ne]3s^2 3p^6$
K	19	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$	$[Ar]4s^1$
Ca	20	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$	$[Ar]4s^2$
Sc	21	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$	$[Ar]4s^2 3d^1$
Ti	22	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$	$[Ar]4s^2 3d^2$
V	23	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$	$[Ar]4s^2 3d^3$
Cr	24	$1s^2 2s^2 2p^6 3s^2 3p^6 \mathbf{4s^1 3d^5}$	$[Ar]\mathbf{4s^1 3d^5}$
Mn	25	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$	$[Ar]4s^2 3d^5$
Fe	26	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$	$[Ar]4s^2 3d^6$
Co	27	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7$	$[Ar]4s^2 3d^7$
Ni	28	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$	$[Ar]4s^2 3d^8$
Cu	29	$1s^2 2s^2 2p^6 3s^2 3p^6 \mathbf{4s^1 3d^{10}}$	$[Ar]\mathbf{4s^1 3d^{10}}$
Zn	30	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$	$[Ar]4s^2 3d^{10}$

Some important observations noted from the above table are given below:

1. Electronic configuration of an element can be written in short form by using nearby noble gas symbol. For example the electronic configuration of Li (Z=3) which is $1s^2\ 2s^1$ can be written as [He] $2s^1$. Here [He] corresponds to the electronic configuration of helium i.e., $1s^2$.
2. Consider the electronic configuration of potassium (Z=19). The 19 electrons in potassium have the configuration [Ar] $4s^1$. Here [Ar] represents $1s^2 2s^2 2p^6 3s^2 3p^6$.
3. Consider the electronic configuration of vanadium (V) and chromium (Cr)

$$V(Z = 23) = [Ar]\ 4s^2 3d^3.$$

The electronic configuration of Cr is expected to be Cr (Z=24) = $[Ar]4s^2 3d^4$

But experimentally it is found to be Cr (Z=24) = [Ar] $4s^1 3d^5$

The valence electronic configuration of Cr can be represented as

4s	3d				
↑	↑	↑	↑	↑	↑

This is due to the extra stability obtained when degenerate orbitals are half - filled or completely filled. In this example $3d^5$ is half filled because the d-orbitals can take 10 electrons.

Thus, $[Ar]4s^1\ 3d^5$ configuration is more stable. Similarly consider the electronic configuration of Ni (Z=28) and Cu (Z=29).

$$Ni\ (Z = 28) = [Ar]\ 4s^2 3d^8$$

The electronic configuration of Cu is expected to be Cu = [Ar] $4s^2 3d^9$.

But experimentally it is found to be Cu=$[Ar]4s^1 3d^{10}$. The valence electronic configuration of Cu can be represented as

4s	3d				
↑	↑↓	↑↓	↑↓	↑↓	↑↓

This is because $3d^{10}$ is completely filled configuration and hence more stable.

Note:

1. In case of Cr (Z=24) = [Ar] $4s^1 3d^5$ and Cu = $[Ar]4s^1 3d^{10}$, the actual rules are slightly violated to attain extra stability.

So, these are considered as **'anomalous electronic configurations'**.

2. The electronic configuration of heavier elements does not strictly follow the above rules.

Solved Examples:

On Electromagnetic radiations:

1. A radio station broadcasts at a frequency of 300 MHz, calculate wavelength of the wave?

Sol: Given: ϑ = 300 MHz

Speed of light c = 3×10^8 m/s,

Using the wavelength formula, we have, $\vartheta = \frac{c}{\lambda}$

$$\Rightarrow \quad \lambda = \frac{c}{\vartheta}$$

$$= \frac{3\times10^8}{300\times10^6} = 1\,m$$

On Photoelectric effect:

2. An ultraviolet radiation with a wavelength of 250 nm is incident on a silver foil (work function Wo= 6.9 × 10-19 J). What is the maximum kinetic energy of the emitted electrons?

Sol: Given:

Work function of silver $W_o = 6.9 \times 10^{-19}$ J

Wavelength of UV radiation λ = 250 nm

$= 250\times10^{-9}$m

$= 2.50\times10^{-7}$m

Planck's constant h = 6.625×10^{-34}J-s

Speed of light c = $3.0 \times 10^8 m.s^{-1}$

Energy of UV radiation:

$$\Rightarrow \quad E = \frac{hc}{\lambda}$$

$$= \frac{6.625\times10^{-34}\times3\times10^8}{2.5\times10^{-7}} = 7.95\times10^{-19}\,J$$

$= 7.95 \times 10^{-19}$ J

We have E = W_o + KE

$\Rightarrow$ KE = E – W_o

$= 7.95 \times 10^{-19} - 6.9 \times 10^{-19}$

$= 1.05 \times 10^{-19}$ J

$\therefore$ The maximum kinetic energy of the emitted electron will be 1.05×10^{-19} J.

On Hydrogen Spectrum:

3. Calculate the wave number of the first spectral line in the lyman series of He^+ spectrum.

Sol: For first line in Lyman series,

n_1 = lower energy = 1, n_2 = higher energy = 2

wave number $(\bar{\vartheta})$ is given by the equation

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

$$= 4 \times 109677 \times \frac{3}{4}$$

$$= 329031 \ cm^{-1}$$

$\therefore$ Wave number of first line in Lyman series of He^+ is $3.29 \times 10^5 \, cm^{-1}$.

On Bohr's model:

4. Calculate the radius of 1st orbit of hydrogen atom.

Sol: Given n = 1, Z = 1

We have, radius of n^{th} orbit $r_n = \frac{0.529 \times n^2}{Z} A^o$

$$= 0.529 \times \frac{1}{1} = 0.529 \ A^o$$

5. What is the energy of the electron in its first exited state in He+ ion?

Sol: Given n = 2 (first exited state), Z = 2

We have, energy of the electron in n^{th} orbit $E_n = \frac{-13.6}{n^2} \times Z^2 eV / ion$

$$= \frac{-13.6}{4} \times 4 = -13.6 \ eV / ion$$

On Aufbau's Principle:

6. Following Aufbau's principle write the electronic configurations of elements from boron (Z = 5) to neon (Z = 10).

Sol: B (Z=5) configuration $1s^2 2s^2 2p^1$

C (Z=6) configuration $1s^2 2s^2 2p^2$

N (Z=7) configuration $1s^2 2s^2 2p^3$

O (Z=8) configuration $1s^2 2s^2 2p^4$

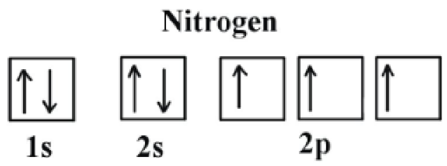
F (Z=9) configuration $1s^2 2s^2 2p^5$

Ne (Z=10) configuration $1s^2 2s^2 2p^6$

On Hund's Rule:

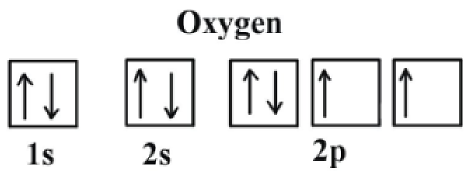
7. Explain Hund's rule by taking example elements nitrogen (Z = 7) to oxygen (Z = 8).

Sol: If we look at the correct electron configuration of the Nitrogen (Z = 7) atom, a very important element in the biology of plants: $1s^2 2s^2 2p^3$



We can clearly see that p orbitals are half-filled as there are three electrons and three p orbitals. This is because Hund's Rule states that the three electrons in the 2p subshell will fill all the empty orbitals first before filling orbitals with electrons in them.

If we look at the element after nitrogen in the same period, oxygen (Z=8) its electron configuration is: $1s^2 2s^2 2p^4$.

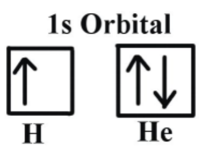


Oxygen has one more electron than nitrogen and as the orbitals are all half-filled the electron must pair up.

On Pauli's Exclusion Principle:

8. Explain Pauli's Exclusion Principle by taking example elements hydrogen (Z = 1) to helium (Z = 2).

Sol: The first three quantum numbers of an electron are n=1, l=0, m=0. Only two electrons can correspond to these, which would be either s=-1/2 or s=+1/2. As we already know from our studies of quantum numbers and electron orbitals, we can conclude that these four quantum numbers refer to the 1s subshell. If only one of the s values are given, then we would have $1s^1$ (denoting hydrogen) if both are given we would have $1s^2$ (denoting helium). Visually, this is be represented as:



As shown, the 1s subshell can hold only two electrons and, when filled, the electrons have opposite spins.

On de-Broglie's Wave Equation:

9. Calculate the wavelength associated with an electron moving with a velocity of 1010 cm /sec.

Sol: Mass of the electron = 9.10×10^{-28} g

Velocity of electron = 10^{10} cm / sec

h = 6.625×10^{-27} erg-sec

According to de Broglie's wave equation,

$$\lambda = \frac{h}{mv} = \frac{6.62 \times 10^{-27}}{9.10 \times 10^{-28} \times 10^{10}}$$

$$= 0.727 \times 10^{-9} \text{ cm}$$

$$= 0.0727 \text{ A}^{o} \qquad (\because 10^{-8} \text{ cm} = 1\text{A}^{o})$$

$\therefore$ The wavelength associated with an electron moving with a velocity of 10^{10} cm /sec is 0.0727 A^{o}

On Heisenberg's Uncertainty Principle:

10. Calculate the uncertainty in the position of a particle when the uncertainty in momentum is:

(a) 1×10^{-3} g cm sec^{-1} (b) Zero

Sol: (a) Given

$\Delta p = 1 \times 10^{-3} \text{g cm sec}^{-1}$

h = 6.62×10^{-27} erg-sec and

π = 3.142

According to Heisenberg's uncertainty principle,

$$\Delta x . \Delta p = \frac{h}{4\pi}$$

So, $$\Delta x = \frac{h}{4\pi} \times \frac{1}{\Delta p}$$

$$= \frac{6.62 \times 10^{-27}}{4 \times 3.142} \times \frac{1}{10^{-3}}$$

$$= 0.527 \times 10^{-24} \text{ cm}$$

(b) When the value of $\Delta p = 0$, the value of Δx will be infinity.

C.D.F. CONCEPTS, DEFINITIONS AND FORMULAE

1. The idea of 'anu' and 'paramanu' was proposed by Indian philosopher **Maharshi Kanad.**
2. Electromagnetic radiations are energy radiations which do not need any medium for propagation.
3. Quantum theory of radiation was presented by **Max Planck.**
4. The ejection of electrons from a clean metal surface when a beam of light strikes the surface is known as **Photo Electric Effect.**
5. Rydberg equation is $\bar{\vartheta} = Z^2 R_H \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right) cm^{-1}$

   Rydberg constant R = 109677 cm^{-1} or $1.09677 \times 10^7\ m^{-1}$.
6. Electrons move around the nucleus in specified circular paths called **Orbits** or **Shells** or energy levels.
7. The energy emitted or absorbed (ΔE) is given by Planck's equation i.e., $E = h\vartheta$
8. Angular momentum of the electron $mvr = \frac{nh}{2\pi}$.
9. Radius of 'n[th]' orbit is given by $r_n = 0.529 \frac{n^2}{Z} A^o$
10. Energy expression is $E_n = \frac{-13.6}{n^2} \times Z^2 eV$ per atom / ion
11. The concept of elliptical orbits was proposed by **Sommerfeld.**
12. Dual nature micro particles like electrons was proposed by **De Broglie.** De Broglie's wave equation is $\lambda = \frac{h}{mc} = \frac{h}{p}$.
13. **Heisenberg's** uncertainty principle; It is impossible to determine accurately and simultaneously the position and momentum of a particle in an atom.
14. Heisenberg's uncertainty equation is: $\Delta x . \Delta p \geq \frac{h}{4\pi}$ (or) $\Delta x . m\Delta v \geq \frac{h}{4\pi}$
15. A set of numbers used to provide a complete description of an electron in an atom are called **Quantum Numbers.**
16. **Aufbau's Principle:** Electron filling follows energy ranking. 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 4f, 5d, 6p, 7s ....
17. **Hund's rule:** Pairing of electrons in an orbital starts only after all the orbitals of a subshell are half filled or singly filled.
18. **Pauli's Exclusion Principle:** No two electrons in the same atom can have same values for all the four quantum numbers.

ADVANCED WORKSHEET

LEVEL 1

Single Correct Answer Type (S.C.A.T)

1. The idea of smallest unit of matter (anu and paramanu) was proposed by:

(A) Democritus
(B) John Dalton
(C) William Crookes
(D) Maharshi Kanad

2. Quantum theory of radiation was proposed by:

(A) John Dalton
(B) J.J.Thomson
(C) Max Planck
(D) Rutherford

3. Which of the following relations is correct?

(A) $\vartheta = \frac{c}{\lambda}$
(B) $E = h\vartheta$
(C) $\bar{\vartheta} = \frac{1}{\lambda}$
(D) All of these

4. Work function $h\vartheta_0$ =

(A) $h\vartheta - \frac{1}{2}m_e v^2$
(B) $h\vartheta + \frac{1}{2}m_e v^2$
(C) $\frac{1}{2}m_e v^2 - h\vartheta$
(D) $\frac{1}{2}m_e v^2$

5. In hydrogen spectrum which one of the following series lies in the ultraviolet region?

(A) Balmer series
(B) Pfund series
(C) Bracket series
(D) Lyman series

6. The metal best used in photoelectric cells is:

(A) Na (B) Mg
(C) A*l* (D) Cs

7. In Zeeman effect the splitting of spectral lines is under the influence of:

(A) Electric field
(B) Magnetic field
(C) Both (A) and (B)
(D) None

8. The concept of elliptical orbits was introduced by:

(A) Bohr
(B) Rutherford
(C) Sommerfeld
(D) Max Planck

9. As the velocity of moving particle increases the De Broglie's wavelength:

(A) Decreases

(B) Increases

(C) Constant

(D) Can not say

10. Uncertainty principle was proposed by:

(A) De Broglie

(B) Bohr

(C) Heisenberg

(D) Sommerfeld

11. Who proposed magnetic quantum number?

(A) Bohr

(B) Sommerfeld

(C) Lande

(D) Uhlenbeck

12. Which of the following has same (n+l) value as that of 3p sub shell?

(A) 4s

(B) 3d

(C) 4f

(D) 3s

13. Filling of electrons in various sub shells present in main shell is governed by:

(A) Aufbau's principle

(B) Pauli's principle

(C) Hund's rule

(D) All the above

14. According to Pauli's principle an orbital can have only ____ electrons.

(A) Three

(B) Two

(C) One

(D) Four

15. The electronic configuration of phosphorous is:

(A) [Ne] $3s^2 3p_x^2 3p_y^1 3p_z^0$

(B) [Ne] $3s^2 3p_x^1 3p_y^1 3p_z^1$

(C) [Ne] $3s^2 3p_y^2 3p_x^1 3p_z^0$

(D) [Ne] $3s^2 3p_x^0 3p_y^1 3p_z^2$

16. The value of Max Planck's constant in C.G.S units is:

(A) 6.625×10^{-27}

(B) 6.625×10^{-34}

(C) 3×10^{8}

(D) 3×10^{10}

17. The ratio of the energy of photon of 2000A° wave length to that of a photon of 4000Ao wave length is:

(A) 1:4

(B) 4:1

(C) 1:2

(D) 2:1

18. Calculate the energy associated with a photon of light whose frequency is 3×10^{15} Hz.

($h = 6.625 \times 10^{-34}$ Js)

(A) 6.625×10^{-19} J

(B) 1.987×10^{-19} J

(C) 1.987×10^{-18} J

(D) 5.387×10^{-20} J

19. The wavelength of a radiation is 97540 cm. Calculate its frequency.

(A) $3.07 \times 10^{5}\ s^{-1}$

(B) $3.07 \times 10^{15}\ s^{-1}$

(C) $3.07 \times 10^{2}\ s^{-1}$

(D) $3.07 \times 10^{20}\ s^{-1}$

20. Light of any electromagnetic radiation travels in vacuum or air with a speed of:

(A) 3×10^{8} m/s

(B) 3×10^{2} m/s

(C) 2×10^{8} m/s

(D) 1×10^{8} m/s

21. The equation corresponding to the wave number of spectral line in P-fund series is:

(A) $R\left[\frac{1}{4^2}-\frac{1}{5^2}\right]$

(B) $R\left[\frac{1}{3^2}-\frac{1}{4^2}\right]$

(C) $R\left[\frac{1}{2^2}-\frac{1}{3^2}\right]$

(D) $R\left[\frac{1}{5^2}-\frac{1}{6^2}\right]$

22. Radius of the nth orbit in single electron species is given by:

(A) $r = \frac{0.529\times Z}{n^2}$ A°

(B) $r = \frac{0.529\times n^2}{Z}$ A°

(C) $r = \frac{Z\times n^2}{0.529}$ A°

(D) $r = \frac{Z}{0.529\times n^2}$ A°

23. The angular momentum of an electron revolving in Bohr's second orbit is:

(A) $h/2\pi$

(B) h/π

(C) $h/3\pi$

(D) $3h/2\pi$

24. Momentum of a photon of wavelength 'λ' is:

(A) h/λ

(B) Zero

(C) $h\lambda/c^2$

(D) $h\lambda/c$

25. Calculate the momentum of a particle which has a de-Broglie wavelength of 2.5×10^{-10}m.

($h = 6.625 \times 10^{-34}$ kg m^{2}s^{-1})

(A) 5.32×10^{-23} Kg m sec^{-1}

(B) 2.65×10^{-24} Kg m sec^{-1}

(C) 1.38×10^{-19} Kg m sec^{-1}

(D) 9.2×10^{-20} Kg m sec^{-1}

26. Calculate the uncertainty in velocity of a cricket ball of mass 150 g if the uncertainty in its position is of the order of 1 A°.

($h = 6.6 \times 10^{-34}$ kg m^2 s^{-1}).

(A) 4.261×10^{-20} ms^{-1}

(B) 8.356×10^{-24} ms^{-1}

(C) 3.50×10^{-24} ms^{-1}

(D) 1.624×10^{-24} ms^{-1}

27. If the electron is spinning in clockwise direction, its spin value is:

(A) $+\frac{1}{2}$ (B) $-\frac{1}{2}$

(C) 0 (D) $\pm\frac{1}{2}$

28. Which of the following options does not represent ground state electronic configuration of an atom?

(A) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^8 4s^2$

(B) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^9 4s^2$

(C) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$

(D) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$

29. The pair of ions having same electronic configuration is ___.

(A) Cr^{3+}, Fe^{3+}

(B) Fe^{3+}, Mn^{2+}

(C) Fe^{3+}, Co^{3+}

(D) Sc^{3+}, Cr^{3+}

Multi Corrrect Questions (M.C.Q)

30. Bohr's theory is not applicable to:

(A) He (B) Li^{2+}

(C) He^{2+} (D) H^+

31. Heisenberg's uncertainty principle is not valid for:

(A) Moving electrons

(B) Motor car

(C) Stationary particles

(D) Revolving planet

32. Which of the following elements are having more number of electrons with spin value +1/2 than electrons with spin value –1/2?

(A) Na (B) B

(C) C (D) He

33. Which of the following configurations violates Hund's rule?

(A) $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}\boxed{\downarrow}\boxed{\uparrow}$

(B) $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$

(C) $\boxed{\uparrow\downarrow}$ $\boxed{\downarrow}\boxed{\downarrow}\boxed{\downarrow}$

(D) $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{}$

34. Which of the following orbitals have same (n+l) value?

(A) 3d (B) 4p

(C) 5s (D) 5p

Comprehension Passage (C.P.T.)

The set of numbers used to locate an electron in an atom are known as quantum numbers.

There are four quantum numbers, they are:

Principal quantum number (n), Azimuthal quantum number (l), Magnetic quantum number (m) and Spin quantum number (s).

35. Which of the following sets of quantum numbers is not possible?

(A) n = 2, l = 1, m = 0

(B) n = 2, l = 0, m = –1

(C) n = 3, l = 0, m = 0

(D) n = 3, l = 1, m = –1

36. Which of the following sets of quantum numbers is correct for an electron in 4f orbital?

(A) n = 4, l = 3, m = +4, s = + 1/2

(B) n = 4, l = 4, m = –4, s = – 1/2

(C) n = 4, l = 3, m = +1, s = + 1/2

(D) n = 3, l = 2, m = –2, s = + 1/2

37. The size of the orbit is given by the quantum number:

(A) n

(B) l

(C) m

(D) s

Matrix Matching Type (M.M.T.)

Column - I

38. Elliptical orbits

39. Stationary orbits

40. Uncertainty principle

41. Dual nature of electron

Column - II

(A) Bohr

(B) de Broglie

(C) Sommerfeld

(D) Heisenberg

Assertion Reason Type (A.R.T.)

(A) Assertion and Reason are true and Reason is the correct explanation of Assertion

(B) Assertion and Reason are true but Reason is not the correct explanation of Assertion

(C) Assertion is true but Reason is false

(D) Assertion is false but Reason is true

42. **Assertion:** For n = 3, l = 0, 1 & 2, and m = –2, –1, 0, +1, & +2

Reason: For a given value of n, the values of l are all integers from 0 to n–1 and for a given value of l, the values of m are all integers from $-l$ to $+l$ including 0.

43. **Assertion:** The main shell with principal quantum number n = 2 has four orbitals present in it.

Reason: Number of orbitals present in a shell is given by n^2.

Statement Type (S.T.)

(A) Both statements are correct

(B) Both statements are incorrect

(C) Statement I is correct statement II is incorrect

(D) Statement I is incorrect Statement II is correct

44. **Statement I:** Energy levels are also called as stationary orbits.

Statement II: Energy of an electron remains constant so long as it revolves in a given orbit.

45. **Statement I:** Bohr's orbits are called stationary orbits.

Statement II: The angular momentum of the electron revolving in a stationary orbit is equal to integral multiples of $\frac{h}{2\pi}$.

46. **Statement I:** The principal quantum number determines the orientation and energy of the orbital.

Statement II: Angular quantum number determines the three dimensional shape of the orbital.

47. **Statement I:** Two electrons of an atom may have identical values of n, *l* and m.

Statement II: Two electrons of equal energy occupying p-orbitals of an atom may have parallel spin

Integer Type Questions (I.T.Q.)

48. The N-shell can accommodate a maximum of how many electrons?

49. If an atom has 2 electrons in its K-shell and 5 electrons in its L-shell, what is the atomic number of the element?

Analytical Approach Type (A.A.T.)

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

Metal	Li	Na	K	Mg	Cu	Ag	Fe	Pt	W
W_0(eV)	2.4	2.3	2.2	3.7	4.8	4.3	4.7	6.3	4.75

(A) 4

(B) 6

(C) 8

(D) 12

51. **Consider the ground state of Cr atom (Z = 24). The numbers of electrons with the azimuthal quantum numbers, l = 1 and 2 are, respectively:**

(A) 12 and 4

(B) 12 and 5

(C) 16 and 4

(D) 16 and 5

52. **The de Broglie wavelength of a tennis ball of mass 60 g moving with a velocity of 10 metres per second is approximately :**

(A) 10^{-33} metre

(B) 10^{-31} metre

(C) 10^{-16} metre

(D) 10^{-25} metre

53. **Consider the ground state of Cr atom (Z = 24). The number of electrons with the azimuthal quantum numbers, l = 1 and 2 are, respectively:**

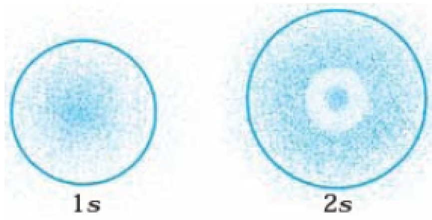
(A) 12 and 4

(B) 12 and 5

(C) 16 and 4

(D) 16 and 5

54. The probability density plots of 1s and 2s orbitals are given in figure:



The density of dots in a region represents the probability density of finding electrons in the region.

On the basis of above diagram which of the following statements is incorrect?

(A) 1s and 2s orbitals are spherical in shape.

(B) The probability of finding the electron is maximum near the nucleus.

(C) The probability of finding the electron at a given distance is equal in all directions.

(D) The probability density of electrons for 2s orbital decreases uniformly as distance from the nucleus increases.