Chapter

2 Structure of Atom

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SUB-ATOMIC PARTICLES

Many different kinds of sub-atomic particles were discovered in the twentieth century. Three of these are electron, proton and neutron.

Electron

Electrons are the basic constituent of all atoms. They were discovered by cathode ray discharge tube experiment. Cathode rays consist of negatively charged particles called electrons.

The charge to mass ratio of electron is given by

$$\frac{e}{m_e} = 1.758820 \times 10^{11}\ \text{C kg}^{-1}$$

where $m_e \rightarrow$ mass of e^-

$e \rightarrow$ magnitude of charge on electron.

Charge on electron is -1.6×10^{-19} C and the mass of electron is 9.1×10^{-31} kg.

Proton

The smallest and lightest positive ion is proton. These positively charged particles are called canal rays. The credit for the discovery of proton goes to **Goldstein. Thomson** and **Wein** estimated the value of e/m as 9.578×10^4 coulomb per gram for the positively charged particle proton. Proton is obtained when the only one electron present in hydrogen atom is removed. Hydrogen atom consists of one electron, one proton and no neutron.

Neutron

The electrically neutral particles present in the atom are neutrons.It was discovered by **James Chadwick** in 1932. Neutron is relatively heavier out of the three fundamental particles of atom.It is assumed that a neutron is a result of joining together of an electron and a proton. A neutron, being unstable, decays as follows :

$${}^{1}_{0}n \longrightarrow {}^{1}_{+1}H + {}^{0}_{-1}e + {}^{0}_{0}q \text{ (antineutrino)}$$

Its half-life is 20 minutes.

Charge and Mass of Fundamental Subatomic Particles

Name	Symbol	Absolute charge/c	Relative charge	Mass(kg)	Mass(u)	Approx. mass(u)
Electron	e	-1.6×10^{-19}	-1	9.1×10^{-31}	0.00054	0
Proton	p	$+1.6\times10^{-19}$	$+1$	1.67×10^{-27}	1.00727	1
Neutron	n	0	0	1.67×10^{-27}	1.00867	1

ATOMIC MODELS

Thomson Model of Atom

It is also called plum pudding, raisin pudding or water melon model. It can be visualised as a pudding or watermelon of positive charge with plums or seeds (e^{-1}s) embedded into it. An important feature of this model is that the mass of atom is assumed to be uniformly distributed over the atom. The model was able to explain the overall neutrality of atom but was not consistent with results of later experiments.

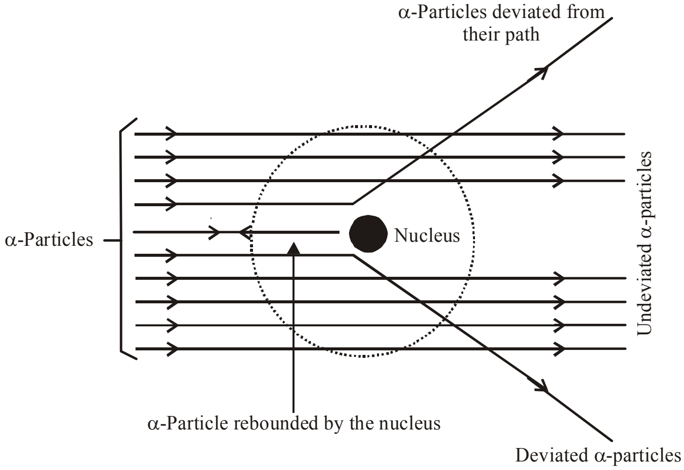
Rutherford's Model of Atom

Rutherford performed the α-particle scattering experiment. A stream of high energy α-particles from a radioactive source was directed at a thin foil of gold metal. It was observed that:

(i) Most of α-particles passed through gold foil undeflected

(ii) A small fraction of α-particles was deflected by small angles.

(iii) A few α-particles bounced back, i.e., were deflected by nearly 180°.



On the basis of the observations, following conclusions were drawn.

(i) Most of the space in the atom is empty.

(ii) The deflection of particles was due to the positive charge which is concentrated in a very small volume

(iii) The volume occupied by the nucleus is small as compared to the total volume of the atom.

Thus, according to this model,

(i) The positive charge and most of the mass of the atom was densely concentrated in a small region called the nucleus of atom.

(ii) The nucleus is surrounded by e^{-1}s that move around the nucleus with a very high speed in circular paths called orbits.

(iii) e^{-1}s and nucleus are held together by electrostatic forces of attraction.

Drawback of Rutherford model: (i) It cannot explain the stability of atom. (ii) It does not explain the electronic structure of atom.

Atomic Number and Mass Number

(i) Atomic number (Z) = number of protons in nucleus of atom
= number of e^{-1}s in a neutral atom

(ii) Mass number (A) = number of protons (Z) + number of neutrons (n)

For example,	$^{16}_{8}O$	$^{14}_{7}N$	$^{23}_{11}Na$
Protons	8	7	11
Neutrons	8	7	12
Mass number	16	14	23

Isobars and Isotopes

(i) Isobars are the atoms with same mass number but different atomic number. For ex:

$^{40}_{18}Ar$ and $^{40}_{19}K$

(ii) Isotopes are atoms of the same element with same atomic number but different mass number. For ex: Hydrogen has 3 isotopes

$^{1}_{1}H$, $^{2}_{1}H$, $^{3}_{1}H$

Isosters

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

Ex.		CO_2	N_2O
Atoms	=	1 + 2	2 + 1
	=	3	= 3
Electrons	=	6 + 8 × 2	7 × 2 + 8
	=	22	= 22

DEVELOPMENTS LEADING TO BOHR'S MODEL OF ATOM

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

(i) Dual character of electromagnetic radiation

(ii) Atomic spectra which is explained by assuming quantized electronic energy levels in atoms.

Dual Nature of Electromagnetic Radiation

It means that the radiations possess both wave-like and particle–like properties, i.e.,

(a) Wave nature : Electromagnetic radiations are waves which are transmitted when electrically charged particle moves under acceleration, producing alternate electric and magnetic fields. Electromagnetic radiations are of many types, differing from one another in wavelength (or frequency). They constitute the electromagnetic spectrum.

These radiations are characterised by frequency (ν) and wavelength (λ). SI unit of ν is Hertz (Hz) or s^{-1}

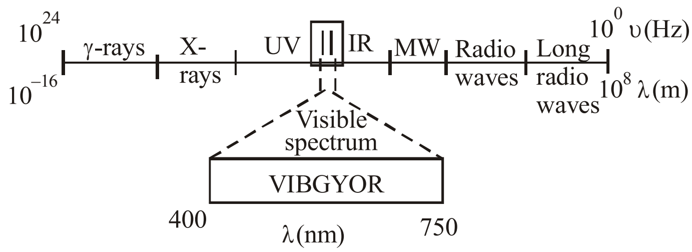
SI unit of λ is m.

ν, λ and c (speed of light) are related as $c = \nu\lambda$

$\bar{\nu}$ = wave number = $\frac{1}{\lambda}$

Electromagnetic spectrum: It is defined as the arrangement of various types of electromagnetic radiation in terms of increasing (or decreasing) wave lengths (or frequency). The complete range of electromagnetic waves is called electromagnetic spectrum. The wavelength of various waves increases in the following order

Cosmic rays < γ-rays < X-rays < UV rays < Visible < IR rays < Micro waves < Radio waves



(b) Particle nature

According to Planck, atoms and molecules emit energy only in discrete quantities and gave the name quantum to the smallest quantity of energy that can be emitted or absorbed

The energy of quantum of radiation is given by:

$E = h\nu$

where $\nu \rightarrow$ frequency

$h \rightarrow$ Planck's constant $(6.6 \times 10^{-34}$ Js)

Photoelectric effect: when a beam of light strikes the surface of a metal, e^{-1}s are ejected from it and the phenomenon is called photoelectric effect. For each metal, there is a minimum frequency (ν_0) called threshold frequency below which photoelectric effect is not observed.

Work function : A part of the photons energy that is absorbed by the metal surface to release the e^{-1}s known as work function of the surface. The remaining part of energy goes into K.E.

At $\nu > \nu_0$, ejected e^{-1}s come out with some K.E.

$\therefore h\nu = h\nu_0 + \frac{1}{2}mv^2$ i.e. $E = E_0 + KE$

Where, E : energy of radiation

E_0 : minimum energy

Electromagnetic wave theory could explain phenomenon like interference, diffraction, etc., but it could not explain some other phenomenon like black body radiation, photoelectric effect, etc. These phenomenon could be explained only if em waves are supposed to have particle nature. It was explained by **Max Planck** and is called the **Planck's Quantum Theory**.

ATOMIC SPECTRA

Spectrum is a series of coloured bands which are formed when white light passes through a prism. A spectrum is of mainly three types :

(i) Emission spectrum
(ii) Absorption spectrum
(iii) Molecular spectrum

Emission Spectra

The spectrum of radiation emitted by a substance that has absorbed energy is called emission spectrum. Emission spectra are of the following two types

(i) Continuous spectrum and
(ii) line spectrum or atomic spectrum

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

(ii) Line Spectrum : In the gas phase, the emission spectra of atoms is not continuous, rather they emit light only at specific wavelengths with dark spaces between them. Such spectra are called line spectra or atomic spectra.

Absorption Spectra

It is observed when a continuum of radiation is passed through a sample which absorbs radiation of certain wavelengths. The missing wavelength leaves dark spaces in the bright continuum spectrum.

Molecular Spectrum

Molecular spectrum is given by molecules and it is also known as band spectrum.

LINE SPECTRUM OF HYDROGEN

The spectrum of hydrogen consists of several lines named after their discover.

The visible lines of the H-spectrum obey the following formula:

$$\bar{\nu} = 109,677\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

where $\bar{\nu}$ = wave number,

$109,677\ cm^{-1}$ = Rydberg's const (R_H)

Series	n_1	n_2	Spectral region
Lyman	1	2,3...	UV
Balmer	2	3,4,...	Visible
Paschen	3	4,5,...	IR
Brackett	4	5,6,...	IR
Pfund	5	6,7,...	IR

BOHR'S MODEL FOR H-ATOM

This model is based on following postulates :

(i) The electron in H-atom moves around the nucleus in circular paths called orbits, stationary states or energy states.

(ii) The energy of e^- in the orbit does not change. The energy change takes place only when an e^- jump from lower energy orbit to higher energy orbit or vice versa.

(iii) The frequency of radiation absorbed or emitted is given by:

$$\nu = \frac{\Delta E}{h} = \frac{E_2 - E_1}{h}$$

Where, E_1 = energy of lower state

E_2 = energy of higher state

(iv) The angular momentum of an electron in a given stationary state is $mvr = \frac{nh}{2\pi}$

(v) Frequency (ν) associated with the absorption and emission of photon is given as :

$$\nu = \frac{\Delta E}{h} = \frac{R_H}{h}\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$\bar{\nu} = \frac{\nu}{C} = \frac{R_H}{hc}\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$= 1.09677 \times 10^7\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) m^{-1}$$

Features of Bohr's Model

According to Bohr's theory of H-atom:

(i) The stationary states for e^- are numbered as 1, 2, 3, ... and are known as principal quantum numbers.

(ii) The radii of the stationary states are given by:

$r_n = n^2 a_0$ where $a_0 = 52.9$ pm

(iii) The energy of the stationary state is given by:

$$E_n = -R_H\left(\frac{1}{n^2}\right) \text{ where, } R_H \rightarrow \text{Rydberg constant}$$

(iv) Bohr's theory can be applied to ions containing only one e^- like He^+, Li^{2+}, Be^{3+} and so on. The energy of stationary states of these ions is given by:

$$E_n = -2.18 \times 10^{-18}\left(\frac{Z^2}{n^2}\right) J$$

and, radii by the expression

$$r_n = \frac{52.9\left(n^2\right)}{Z} pm$$

Limitations of the Bohr's Model

Bohr's model ,

(i) was unable to explain the spectrum of atoms other than hydrogen.

(ii) could not explain the ability of atoms to form molecules by chemical bonds.

(iii) was unable to explain splitting of spectral lines in magnetic field (Zeeman effect).

(iv) was unable to explain the splitting of lines in electrical field (Stark effect)

DUAL BEHAVIOUR OF MATTER

DeBroglie's concept was experimentally verified by **Davisson and Germer** by observing diffraction effect (a property shown by waves) with an electron beam.

On the basis of dual nature of matter, de Broglie gave the following relation betweem wavelength (λ) and momentum (p) of material particle, i.e.

$$\lambda = \frac{h}{mv} = \frac{h}{p} \qquad \text{where, h = Planck's constant}$$

HEISENBERG'S UNCERTAINTY PRINCIPLE

According to the principle, it is impossible to determine simultaneously, the exact position and exact momentum of an electron. Mathematically,

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

$$\Delta x.\Delta\left(mv_x\right) \geq \frac{h}{4\pi}$$

$$\Delta x.\Delta v_x \geq \frac{h}{4\pi m}$$

Here

Δx is uncertainty of position,

Δp is uncertainty of momentum and

h is Planck's constant

QUANTUM MECHANICAL MODEL OF ATOM

The branch of science that takes into account dual behaviour of matter is called quantum mechanics.

Schrodinger Wave Equation

Schrodinger (1927) gave a mathematical expression known as a **Schrodinger wave equation**. His theory is based on quantum mechanical model of atom in which the concept of probability of finding the electron at any position around the nucleus at any instant of time is considered.

Significance of wave function : An atomic orbital is the wave function ψ for an e^- in an atom. ψ^2 is known as probability density. ψ^2 gives the probability of finding the e^- around the nucleus.

QUANTUM NUMBERS

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

Quantum numbers are designated as **principal quantum number** (*n*), **azimuthal quantum number** (*l*), **magnetic quantum number** (*m*) and **spin quantum number**(s).The four quantum numbers provide the following information

(i) **n** defines the shell, determines the size of the orbital and the energy of the orbital. An atom has K, L, M, N, O, P, Q, etc. shell.

(ii) ℓ identifies the subshell and determines the shape of the orbital. There are $(2\ell + 1)$ orbitals of each type in a subshell, i.e., one s-orbital ($\ell = 0$); three p-orbitals ($\ell = 1$) and five d-orbitals ($\ell = 2$) per sub-shell.

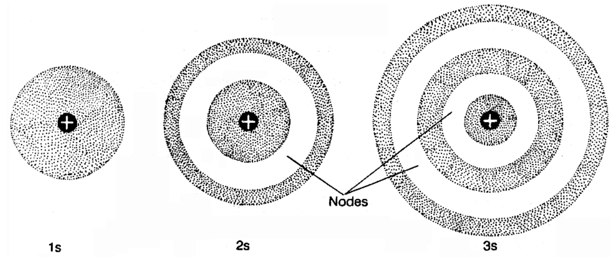
(iii) $\mathbf{m_l}$ designates the orientation of the orbital. For a given value of ℓ, m_l has $(2\ell + 1)$ values.

(iv) $\mathbf{m_s}$ refers to orientation of the spin of the e^-. m_s can have two values $+\frac{1}{2}$ or $-\frac{1}{2}$.

SHAPES OF ATOMIC ORBITALS

(i) s-orbital:

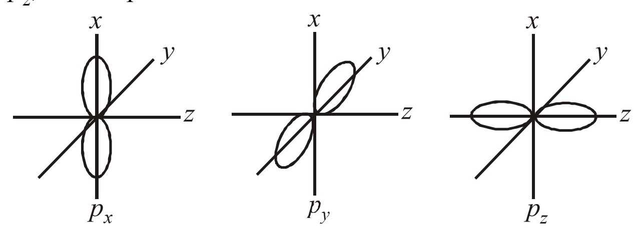
For *s*-orbitials $l = 0$, so there is only one value of m i.e. $m = 0$. Thus, *s*-orbital can have only one orientation i.e. the probability of finding electron is same in all directions at a given distance from the nucleus. Hence *s*-orbital is symmetrical around the nucleus and thus has spherical shape.



Shapes of 1s, 2s and 3s orbitals

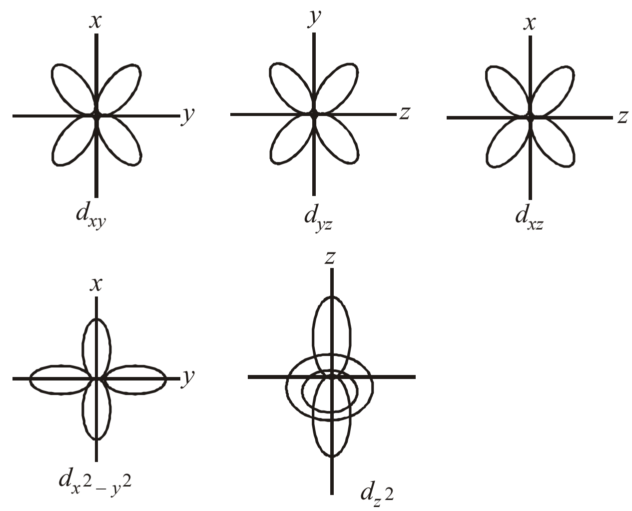
(ii) p-orbital:

For p-orbitals $l = 1$, the permissible values of m are +1, 0 and –1. Thus, there are three p-orbitals, designated as p_x, p_y and p_z, in each p-subshell.



(iii) d-orbital

For d-orbitals $l = 2$ i.e. the permissible values of m are –2, –1,0, +1, +2. This indicates for the five d-orbitals designated as $d_{xy}, d_{yz}, d_{zx}, d_{x^2-y^2}$ and d_z^2.



NODES OF ORBITALS

The region where probability density function reduces to zero is called nodes.

Generally ns – orbital has (n – 1) nodes

Number of spherical/radial nodes in any orbital = $n - \ell - 1$

Number of planer nodes in any orbital = ℓ

$\therefore$ Total number of nodes in any orbital = n – 1

ENERGIES OF ORBITALS

The energy of orbitals for H-atom depends only on n and increases as follows:

$$1s < 2s = 2p < 3s = 3p = 3d < 4s = 4p = 4d = 4f$$

The orbitals having same energy are called degenerate. For multi e^- atoms, energy of an e^- depends on n and ℓ. The lower the value of $(n + \ell)$ for an orbital, the lower is its energy. If two orbitals have same value of $(n + \ell)$, the orbital with lower value of n will have lower energy.

Filling of Orbitals in Atom

(i) Aufbau principle: It states that in the ground state of atoms, orbitals are filled in order of their increasing energies.

The order in which energies of orbitals increase is:

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

(ii) Pauli Exclusion principle: It states that no two e^{-1}s in an atom can have same set of four quantum numbers.

(iii) Hund's rule of maximum multiplicity: It states that pairing of e^{-1}s in orbitals belonging to same subshell (*p*, *d* or *f*) does not take place until each orbital belonging to that subshell has got one e^- each, i.e., it is singly occupied.

Writing The Configuration of Atoms/ Ions

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

Stability of Completely Filled or Half-filled Sub-shells

The completely filled or exactly half-filled sub-shells are stable due to following reasons.

(i) Orbitals which are either half filled or fully filled are more symmetrical and therefore possess lower energy i.e. extra stability.

(ii) When orbitals are half filled or fully filled, the exchange of electrons between the orbitals is maximum. Such exchanges leads to greater stability of electrons in the orbitals, because low exchange energy results in to higher stabilization.

CONCEPT MAP

STRUCTURE OF ATOM

Discovery of substomic particles

Discovery of Electron:
(William crookes 1879)
cathode rays consist of negatively charged particles called electrons.
$e/m = 1.76 \times 10^8$ C/g
$m = 9.1 \times 10^{-31}$ kg

Discovery of Proton:
(Gold stein 1886)
Anode rays are positively charged

Discovery of Neutron:
(C Hadwick 1932)
Neutral particles with mass highly greater than proton present in the nucleus

Developments leading to Bohr's Model

Wave nature of e.m waves consist of electric and magnetic fields oscillating perpendicular to each other
$c = v \times \lambda$
$\upsilon \propto 1/\lambda$

Particle nature of e.m. waves
$E = nh\nu$
$K.E = h(\nu - \nu_o)$
No. of photoelectrons α intensity of light

Models of Atom

J.J. Thomson's Model:
also known as Plum-pudding Model

Rutherford's Model:
Fast moving α-particles are bombarded on a thin foil of metals like Ag, Au, Cu, Pt etc
Atom consists of two parts
(a) Nucleus
(b) Extranuclear

Atomic number (Z)
= No. of protons
= no. of electrons
Mass number (A)
= No. of protons + No. of neutrons
Isotopes
= Same atomic no. but different mass no.
Isobars
= Same mass no. but different atomic no.

Bohr's Model of Atom

$$E_n = \frac{E_2}{n^2} = -R_H\left(\frac{1}{n^2}\right) = \frac{-1312}{n^2} \text{ kg/mole}$$

$$r_n = 0.529 \times n^2 A, \; mvr = \frac{nj}{2\pi}$$

$$\nu = \frac{\Delta E}{h}, \; \bar{\nu} = \frac{R_H}{Hc}\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

Quantum Mechanical Model of Atom

$$\lambda = \frac{h}{mV} = \frac{h}{P} \quad \Delta x \times \Delta v \geq \frac{h}{4\pi m}$$

Principal quantum number (n):
Denotes the energy level of electron
No. of electron in any shell = $2n^2$
Azimuthal quantum number (l) =
Tell about the no. of subshells within a given principal energy shell. For n, l = 0 to n – 1)
Magnetic quantum number m)
(Behaviour of e^- in magnetic field)
For l, m = – ℓ through 0 to +*l*
Spin quantum number(s)
(spinning orientation of the electron)
$$s = +\frac{1}{2} \text{ or } -\frac{1}{2}$$

Shapes of orbitals

S-orbital: l = 0, i.e., m = 0 (Symmertical with spherical shape) p-orbital: l = 1 i.e., m = +1, 0 and Y (there are three dumbell shaped p-orbitals, designated as p_x, p_y

Filling of orbitals

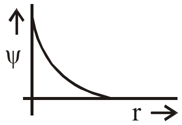
- orbital with lower value of (n + l) will be filled first. In case of equal (n + l) values value, of n should be low
- Aufbau's rule; orbitals are filled in order
- Pauli's exclusion principle: No two electrons can have equal values for all the four quantum numbers
- Hund's rule: Electron pairing in degenerate orbitals take place if each orbital is singly occupied.

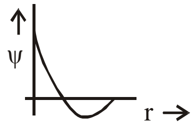
EXERCISE - 1
Conceptual Questions

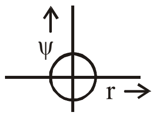
1. The ion that is isoelectronic with CO is
(a) CN^- (b) O_2^+ (c) O_2^- (d) N_2^+

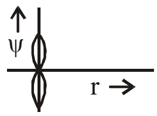
2. Set of isoelectronic species is
(a) N_2, CO_2, CN^-, O_2 (b) N_2, H_2S, CO
(c) N_2, CO, CN^-, O_2^{+2} (d) Ca, Mg, Cl

3. The wave-function (ψ) for a 2s– electron as a function of its distance from the nucleus (r) can be sketched as:

(a) 

(b) 

(c) 

(d) 

4. The combination of three 2p orbitals of one atom and three 2p orbitals of another atom can give rise to :
(a) three molecular orbitals
(b) six (π 2p) molecular orbitals
(c) four (π 2p) and two (σ 2p) molecular orbitals
(d) two molecular orbital which are symmetrical around the bond axis

5. The species isoelectronic with C_2H_4 is
(a) O_2^+ (b) CN^-
(c) N_2^+ (d) O_2

6. Cathode rays are deflected by
(a) an electric field only (b) magnetic field only
(c) by both (d) by none

7. The number of neutrons in dipositive zinc ion with mass number 70 is
(a) 34 (b) 36 (c) 38 (d) 40

8. Which has the highest *e*/*m* ratio ?
(a) He^{2+} (b) H^+ (c) He^+ (d) D^+

9. One would expect proton to have very large :
(a) ionization potential (b) radius
(c) charge (d) hydration energy

10. Which is correct statement about proton ?
(a) Proton is nucleus of deuterium
(b) Proton is α-particle
(c) Proton is ionized hydrogen molecule
(d) Proton is ionized hydrogen atom

11. Which one of the following is expected to have largest size?
(a) F^- (b) O^{2-} (c) Al^{3+} (d) N^{3-}

12. The maximum number of electrons in p-orbital with n = 6; m = 0 is
(a) 6 (b) 2 (c) 14 (d) 10

13. The compound in which cation is isoelectronic with anion is :
(a) NaCl (b) CsF (c) NaI (d) K_2S

14. What is the ratio of mass of an electron to the mass of a proton?
(a) 1 : 2 (b) 1 : 1 (c) 1 : 1837 (d) 1 : 3

15. The increasing order for the values of e/m (charge/mass) is
(a) e, p, n, α (b) n, p, e, α (c) n, p, α, e (d) n, α, p, e

16. If the mass number of an element is *W* and its atomic number is *N*, then
(a) Number of $e^{-1} = W - N$ (b) Number of $_1H^1 = W - N$
(c) Number of $_0n^1 = W - N$ (d) Number of $_0n^1 = N$

17. According to Bohr's theory the energy required for an electron in the Li^{2+} ion to be emitted from n = 2 state is (given that the ground state ionization energy of hydrogen atom is 13.6 eV)
(a) 61.2 eV (b) 13.6 eV (c) 30.6 eV (d) 10.2 eV

18. Which one is correct for the de-Broglie equation?
(a) $E = \lambda/p$ (b) $p = h/\lambda$
(c) $E = h/p$ (d) none of these

19. Which of the following pairs is correctly matched ?
(a) Isoelectronics N^{3-}, O^{2-}, Cr^{3-}
(b) Isotones $_{14}Si^{30}, _{15}P^{31}, _{16}S^{32}$
(c) Isotopes $_{20}Ca^{40}, _{19}K^{40}$
(d) Isobars $_8O^{16}, _8O^{17}, _8O^{18}$

20. The spectrum of He is expected to be similar to that of
(a) H (b) Li^+ (c) Na (d) He^+

21. Which of the following statements do not form a part of Bohr's model of hydrogen atom ?
(a) Energy of the electrons in the orbits are quantized
(b) The electron(s) in the orbit nearest to the nucleus has the lowest energy
(c) Electrons revolve in different orbits around the nucleus
(d) The position and velocity of the electrons in the orbit cannot be determined simultaneously.

22. The radius of hydrogen atom in the ground state is 0.53 Å. The radius of Li^{2+} ion (atomic number = 3) in a similar state is
(a) 0.17 Å (b) 0.265 Å (c) 0.53 Å (d) 1.06 Å

23. The Bohr orbit radius for the hydrogen atom (n = 1) is approximately 0.530 Å. The radius for the first excited state (n = 2) orbit is (in Å)
(a) 0.13 (b) 1.06 (c) 4.77 (d) 2.12

24. MO configuration of He_2^- is
(a) $(\sigma 1s)^2, (\sigma^* 1s)^2, (\sigma 2s)^1$ (b) $(\sigma 1s)^2, (\sigma^* 1s)^2, (\sigma^* 2s)^1$
(c) $(\sigma 1s)^2, (\sigma^* 1s)^1, (\sigma 2s)^2$ (d) $(\sigma 1s)^2, (\sigma^* 1s)^1, (\sigma^* 2s)^2$

25. The quantum number which is responsible for the size of electron cloud is
(a) spin (b) azimuthal
(c) principal (d) magnetic

26. "Electrons are filled in energy orbitals, in increasing order of energy." This statement is related to
(a) Planck's rule (b) Hund's rule
(c) Pauli's rule (d) Aufbau principle

27. The energy of second Bohr orbit of the hydrogen atom is -328 kJ mol^{-1}; hence the energy of fourth Bohr orbit would be:
(a) -41 kJ mol^{-1} (b) -82 kJ mol^{-1}
(c) -164 kJ mol^{-1} (d) -1312 kJ mol^{-1}

28. The orbital angular momentum for an electron revolving in an orbit is given by $\sqrt{l(l+1)}.\frac{h}{2\pi}$. This momentum for an *s*-electron will be given by
(a) zero (b) $\frac{h}{2\pi}$ (c) $\sqrt{2}.\frac{h}{2\pi}$ (d) $+\frac{1}{2}.\frac{h}{2\pi}$

29. The wavelength of the radiation emitted, when in a hydrogen atom electron falls from infinity to stationary state 1, would be (Rydberg constant = 1.097×10^7 m^{-1})
(a) 406 nm (b) 192 nm
(c) 91 nm (d) 9.1×10^{-8} nm

30. The frequency of radiation emitted when the electron falls from $n=4$ to $n=1$ in a hydrogen atom will be (Given : ionization energy of H=2.18×10^{-18}J atom^{-1}and $h = 6.625\times10^{-34}$ J s)
(a) 1.54×10^{15} s^{-1} (b) 1.03×10^{15} s^{-1}
(c) 3.08×10^{15} s^{-1} (d) 2.00×10^{15} s^{-1}

31. According to Bohr's theory, the angular momentum of an electron in 5th orbit is
(a) 10 h/π (b) 2.5 h/π (c) 25 h/π (d) 1.0 h/π

32. Which of the following transitions of electrons in the hydrogen atom will emit maximum energy?
(a) $n_5 \rightarrow n_4$ (b) $n_4 \rightarrow n_3$
(c) $n_3 \rightarrow n_2$ (d) all will emit same energy

33. The allowed values of m for l = 2 is
(a) 0, $\pm$1 (b) $\pm$1, $\pm$2
(c) 0, $\pm$1, $\pm$2 (d) 0, $\pm$2

34. Pick out the isoelectronic structure from the following:
(i) $\overset{+}{C}H_3$ (ii) $H_3\overset{+}{O}$
(iii) NH_3 (iv) $\overset{-}{C}H_3$
(a) i and ii (b) iii and iv
(c) ii, iii and iv (d) iii, iv and i

35. The first emission line of hydrogen atomic spectrum in the Balmer series appears is (R = Rydberg constant)
(a) $\frac{5}{36}R$ cm^{-1} (b) $\frac{3}{4}R$ cm^{-1}
(c) $\frac{7}{144}R$ cm^{-1} (d) $\frac{9}{400}R$ cm^{-1}

36. Identify the correct statement:
(a) s and p orbitals are degenerate
(b) sp and sp^2 orbitals are degenerate
(c) p and sp^3 orbitals are degenerate
(d) All sp^3 orbitals are degenerate

37. Correct set of four quantum numbers for valence electron of rubidium (Z = 37)is
(a) $5,0,0,+\frac{1}{2}$ (b) $5,1,0,+\frac{1}{2}$
(c) $5,1,1,+\frac{1}{2}$ (d) $6,0,0,+\frac{1}{2}$

38. Zeeman effect refers to the :
(a) splitting up of the lines in an emission spectrum in the presence of an external electrostatic field
(b) random scattering of light by colloidal particles
(c) splitting up of the lines in an emission spectrum in a magnetic field
(d) emission of electrons from metals when light falls upon them

39. The energy of electron in first energy level is -21.79×10^{-12} erg per atom. The energy of electron in second energy level is :
(a) -54.47×10^{-12} erg atom^{-1}
(b) -5.447×10^{-12} erg atom^{-1}
(c) -0.5447×10^{-12} erg atom^{-1}
(d) -0.05447×10^{-12} erg atom^{-1}

40. Which of the following statements about the electron is incorrect?
(a) It is negatively charged particle
(b) The mass of electron is equal to the mass of neutron.
(c) It is a basic constituent of all atoms.
(d) It is a constituent of cathode rays.

41. In a Bohr model of an atom, when an electron jumps from n = 3 to n = 1, how much energy will be emitted?
(a) 2.15×10^{-11} ergs (b) 2.389×10^{-12} ergs
(c) 0.239×10^{-10} ergs (d) 0.1936×10^{-10} ergs

42. When atoms are bombarded with alpha particles, only, a few in million suffer deflection, others pass out undeflected. This is because
(a) the force of repulsion on the moving alpha particle is small
(b) the force of attraction between alpha particle and oppositely charged electrons is very small
(c) there is only one nucleus and large number of electrons
(d) the nucleus occupies much smaller volume compared to the volume of the atom

43. The ionisation potential of a hydrogen atom is -13.6 eV. What will be the energy of the atom corresponding to n = 2.
(a) -3.4 eV (b) -6.8 eV (c) -1.7 eV (d) -2.7 eV

44. If the energy of a photon is given as : = 3.03×10^{-19} J then, the wavelength (λ) of the photon is :
(a) 6.56 nm (b) 65.6 nm (c) 656 nm (d) 0.656 nm

45. In the photo-electron emission, the energy of the emitted electron is
(a) greater than the incident photon
(b) same as than of the incident photon
(c) smaller than the incident photon
(d) proportional to the intensity of incident photon

46. Uncertainty in position of a n electron (mass = 9.1×10^{-28} g) moving with a velocity of 3×10^4 cm/s accurate upto 0.001% will be (use $h/4\pi$) in uncertainty expression where $h = 6.626 \times 10^{-27}$ erg-second).
(a) 1.93 cm (b) 3.84 cm (c) 5.76 cm (d) 7.68 cm

47. The configuration $1s^2, 2s^2\ 2p^5, 3s^1$ shows :
(a) excited state of O_2^-
(b) excited state of neon atom
(c) excited state of fluorine atom
(d) ground state of fluorine atom

48. Positron is :
(a) electron with positive charge
(b) a nucleus with one neutron and one proton
(c) a nucleus with two protons
(d) a helium nucleus

49. The position of both, an electron and a helium atom is known within 1.0 nm. Further the momentum of the electron is known within 5.0×10^{-26} kg ms^{-1}. The minimum uncertainty in the measurement of the momentum of the helium atom is
(a) 50 kg ms^{-1} (b) 80 kg ms^{-1}
(c) 8.0×10^{-26} kg ms^{-1} (d) 5.0×10^{-26} kg ms^{-1}

50. If electron, hydrogen, helium and neon nuclei are all moving with the velocity of light, then the wavelength associated with these particles are in the order
(a) Electron > hydrogen > helium > neon
(b) Electron > helium > hydrogen > neon
(c) Electron < hydrogen < helium < neon
(d) Neon < hydrogen < helium < electron

51. 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^{-31} metres (b) 10^{-16} metres
(c) 10^{-25} metres (d) 10^{-33} metres
Planck's constant, $h = 6.63 \times 10^{-34}$ Js

52. Which of the following is related with both wave nature and particle nature ?
(a) Interference (b) $E = mc^2$
(c) Diffraction (d) $E = h\nu$

53. If the radius of first Bohr orbit be a_0, then the radius of the third orbit would be
(a) $3 \times a_0$ (b) $6 \times a_0$
(c) $9 \times a_0$ (d) $1/9 \times a_0$

54. The ratio of the radius of the first Bohr orbit for the electron orbiting the hydrogen nucleus to that of the electron orbiting the deuterium nucleus (mass nearly twice that of the hydrogen nucleus) is approximately
(a) 2 : 1 (b) 1 : 1
(c) 1 : 2 (d) 4 : 1

55. The de-Broglie wavelength of an electron in the ground state of hydrogen atom is : [K.E. = 13.6 eV; 1eV = 1.602×10^{-19} J]
(a) 33.28 nm (b) 3.328 nm (c) 0.3328 nm (d) 0.0332 nm

56. The mass of a photon with a wavelength equal to 1.54×10^{-8} cm is
(a) 0.8268×10^{-34} kg (b) 1.2876×10^{-33} kg
(c) 1.4285×10^{-32} kg (d) 1.8884×10^{-32} kg

57. When a metal surface is exposed to solar radiations
(a) The emitted electrons have energy less than a maximum value of energy depending upon frequency of incident radiations
(b) The emitted electrons have energy less than maximum value of energy depending upon intensity of incident radiation
(c) The emitted electrons have zero energy
(d) The emitted electrons have energy equal to energy of photons of incident light

58. The wavelength of a moving electron having 4.55×10^{-25} J of kinetic energy is :
(a) 7.27×10^{-7} metre (b) 72.7×10^{-7} metre
(c) 7.27×10^{-9} metre (d) 72.7×10^{-9} metre

59. Which one of the following represents noble gas configuration :
(a) $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ 3d^{10}, 4s^2\ 4p^6\ 4d^{10}, 5s^2\ 5p^6\ 5d^6, 6s^2$
(b) $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ 3d^{10}, 4s^2\ 4p^6\ 4d^{10}, 5s^2\ 5p^6\ 5d^1, 6s^2$
(c) $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ 3d^{10}, 4s^2\ 4p^6\ 4d^{10}, 5s^2\ 5p^6$
(d) $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ 3d^{10}, 4s^2\ 4f^{14}, 5s^2\ 5p^6\ 5d^1$

60. Number of unpaired electrons in N^{2+} is
(a) 2 (b) 0 (c) 1 (d) 3

61. The total number of electrons that can be accommodated in all orbitals having principal quantum number 2 and azimuthal quantum number 1 is
(a) 2 (b) 4 (c) 6 (d) 8

62. An ion has 18 electrons in the outermost shell, it is
(a) Cu^+ (b) Th^{4+} (c) Cs^+ (d) K^+

63. For azimuthal quantum number $\ell = 3$, the maximum number of electrons will be
(a) 2 (b) 6 (c) 0 (d) 14

64. If magnetic quantum number of a given atom represented by -3, then what will be its principal quantum number?
(a) 2 (b) 3 (c) 4 (d) 5

65. For which one of the following sets of four quantum numbers, an electron will have the heighest energy?

	n	l	m	s
(a)	3	2	1	1/2
(b)	4	2	–1	1/2
(c)	4	1	0	–1/2
(d)	5	0	0	–1/2

66. Which of the following should be the possible sub-shells, for $n + \ell = 7$?
(a) 7s, 6p, 5d, 4f (b) 4f, 5p, 6s, 4d
(c) 7s, 6p, 5d, 6d (d) 4s, 5d, 6p, 7s

67. Wavelength associated with electron motion
(a) increases with increase in speed of electron
(b) remains same irrespective of speed of electron
(c) decreases with increase of speed of e^- (electron)
(d) is zero.

68. Which of the following is not correct for electronic distribution in the ground state ?

(a) Co [Ar] $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{}\boxed{}\boxed{}$

(b) Ni [Ar] $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{}\boxed{}$

(c) Cu [Ar] $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{}$

(d) All of the above

69. The electronic configuration of gadolinium (Atomic number 64) is

(a) [Xe] $4f^8 5d^0 6s^2$ (b) [Xe] $4f^3 5d^5 6s^2$
(c) [Xe] $4f^6 5d^2 6s^2$ (d) [Xe] $4f^7 5d^1 6s^2$

70. The following quantum numbers are possible for how many orbital(s) $n = 3, l = 2, m = +2$?

(a) 1 (b) 3 (c) 2 (d) 4

71. The number of spherical nodes in 3p orbitals are

(a) one (b) three
(c) two (d) None of these

72. The order of filling of electrons in the orbitals of an atom will be

(a) $3d, 4s, 4p, 4d, 5s$ (b) $4s, 3d, 4p, 5s, 4d$
(c) $5s, 4p, 3d, 4d, 5s$ (d) $3d, 4p, 4s, 4d, 5s$

73. The orbitals are called degenerate when

(a) they have the same wave functions
(b) they have the same wave functions but different energies
(c) they have different wave functions but same energy
(d) they have the same energy

74. The uncertainties in the velocities of two particles, A and B are 0.05 and 0.02 ms^{-1} respectively. The mass of B is five times to that of the mass of A. What is the ratio of uncertainties $\frac{\Delta x_A}{\Delta x_B}$ in their positions ?

(a) 2 (b) 0.25 (c) 4 (d) 1

75. The orientation of an atomic orbital is governed by

(a) Spin quantum number
(b) Magnetic quantum number
(c) Principal quantum number
(d) Azimuthal quantum number

76. The number of d-electrons retained in Fe^{2+} (At. no. of Fe = 26) ion is

(a) 4 (b) 5 (c) 6 (d) 3

77. The angular speed of the electron in n^{th} orbit of Bohr hydrogen atom is

(a) Directly proportional to n
(b) Inversely proportional of $\sqrt{n}$
(c) Inversely proportional to n^2
(d) Inversely proportional to n^3

78. The energy ratio of a photon of wavelength 3000Å and 6000Å is :

(a) 1 : 1 (b) 2 : 1 (c) 1 : 2 (d) 1 : 4

79. What is the correct orbital designation of an electron with the quantum number, $n = 4, \ell = 3, m = -2, s = 1/2$?

(a) $3s$ (b) $4f$ (c) $5p$ (d) $6s$

80. The electronic configuration of element with atomic number 24 is :

(a) $1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^4, 4s^2$
(b) $1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^{10}$
(c) $1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^6$
(d) $1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^5 4s^1$

81. Which of the following represents correct set of the four quantum numbers for an electron in a $4d$ subshell ?

(a) 4, 2, 1, 0 (b) 4, 2, 1, – 1/2
(c) 4, 3, 2, + 1/2 (d) 4, 3, – 2, – 1/2

82. The de Broglie wavelength associated with a ball of mass 200 g and moving at a speed of 5 m/h, is of the order of ($h = 6.625 \times 10^{-34}$ Js)

(a) 10^{-15} m (b) 10^{-20} m (c) 10^{-25} m (d) 10^{-30} m

83. An electron has principal quantum number 3. The number of its (i) subshells and (ii) orbitals would be respectively

(a) 3 and 5 (b) 3 and 7 (c) 3 and 9 (d) 2 and 5

84. The number of nodal planes 'd' orbital has

(a) 1 (b) 2 (c) 3 (d) 0

85. An element M has an atomic mass 19 and atomic number 9, its ion is represented by

(a) M^+ (b) M^{2+} (c) M^- (d) M^{2-}

86. The measurement of the electron position is associated with an uncertainty in momentum, which is equal to 1×10^{-18} g cm s^{-1}. The uncertainty in electron velocity is, (mass of an electron is 9×10^{-28} g)

(a) 1×10^9 cm s^{-1} (b) 1×10^6 cm s^{-1}
(c) 1×10^5 cm s^{-1} (d) 1×10^{11} cm s^{-1}

87. Maximum number of electrons in a subshell of an atom is determined by the following:

(a) $2l + 1$ (b) $4l - 2$ (c) $2n^2$ (d) $4l + 2$

88. Which of the following is **not** permissible arrangement of electrons in an atom?

(a) $n = 5, l = 3, m = 0, s = + 1/2$
(b) $n = 3, l = 2, m = -3, s = - 1/2$
(c) $n = 3, l = 2, m = -2, s = - 1/2$
(d) $n = 4, l = 0, m = 0, s = - 1/2$

89. The total number of atomic orbitals in fourth energy level of an atom is :

(a) 8 (b) 16 (c) 32 (d) 4

90. Calculate the wavelength (in nanometer) associated with a proton moving at 1.0×10^3 ms^{-1}.
(Mass of proton = 1.67×10^{-27} kg and $h = 6.63 \times 10^{-34}$ Js)

(a) 0.40 nm (b) 2.5 nm (c) 14.0 nm (d) 0.32 nm

91. The frequency of light emitted for the transition $n = 4$ to $n = 2$ of the He^+ is equal to the transition in H atom corresponding to which of the following ?

(a) $n = 2$ to $n = 1$ (b) $n = 3$ to $n = 2$
(c) $n = 4$ to $n = 3$ (d) $n = 3$ to $n = 1$

92. The increasing order of the ionic radii of the given isoelectronic species is :

(a) $Cl^-, Ca^{2+}, K^+, S^{2-}$ (b) $S^{2-}, Cl^-, Ca^{2+}, K^+$
(c) $Ca^{2+}, K^+, Cl^-, S^{2-}$ (d) $K^+, S^{2-}, Ca^{2+}, Cl^-$

93. The velocity of particle A is 0.1 ms^{-1} and that of particle B is 0.05 ms^{-1}. If the mass of particle B is five times that of particle A, then the ratio of de-Broglie wavelengths associated with the particles A and B is
(a) 2 : 5 (b) 3 : 4
(c) 6 : 4 (d) 5 : 2

94. In hydrogen atomic spectrum, a series limit is found at 12186.3 cm^{-1}. Then it belong to
(a) Lyman series (b) Balmer series
(c) Paschen series (d) Brackett series

95. The atom/ion that has the highest number of unpaired electrons is
(a) Mg^{2+} (b) F
(c) N (d) S^{2-}

96. For Balmer series in the spectrum of atomic hydrogen, the wave number of each line is given by $\bar{v} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$ where R_H is a constant and n_1 and n_2 are integers. Which of the following statement(s) is (are) correct?
I. As wavelength decreases, the lines in the series converge.
II. The integer n_1 is equal to 2.
III. The ionization energy of hydrogen can be calculated from the wave number of these lines.
IV. The line of longest wavelength corresponds to $n_2 = 3$.
(a) I, II and III (b) II, III and IV only
(c) I, II and IV (d) II and IV only

97. In which one of the following pairs the two species are both isoelectronic and isotopic?
(Atomic numbers : Ca = 20, Ar = 18, K = 19, Mg = 12, Fe = 26, Na = 11)
(a) $^{40}Ca^{2+}$ and ^{40}Ar (b) $^{39}K^{+}$ and $^{40}K^{+}$
(c) $^{24}Mg^{2+}$ and ^{25}Mg (d) ^{23}Na and $^{24}Na^{+}$

98. Which set of quantum numbers are not possible?

	n	l	m	s
(a)	3	2	0	+1/2
(b)	2	2	1	+1/2
(c)	1	0	0	–1/2
(d)	3	2	–2	+1/2

99. Two fast moving particles X and Y are associated with de Broglie wavelengths 1 nm and 4 nm respectively. If mass of X in nine times the mass of Y, the ratio of kinetic energies of X and Y would be
(a) 3 : 1 (b) 9 : 1
(c) 5 : 12 (d) 16 : 9

100. The wavelength (in cm) of second line in the Lyman series of hydrogen atomic spectrum is (Rydberg constant = R cm^{-1})
(a) $\left(\frac{8R}{9}\right)$ (b) $\left(\frac{9}{8R}\right)$
(c) $\left(\frac{4}{3R}\right)$ (d) $\left(\frac{3R}{4}\right)$

EXERCISE - 2 AAJ KA TOPPER
Applied Questions

1. If the shortest wavelength of the spectral line of H-atom in the Lyman series is X, then the longest wavelength of the line in Balmer series of Li^{2+} is
(a) 9x (b) $\frac{x}{9}$
(c) $\frac{5x}{4}$ (d) $\frac{4x}{5}$

2. Given : The mass of electron is 9.11×10^{-31} kg
Plank constant is 6.626×10^{-34} Js,
the uncertainty involved in the measurement of velocity within a distance of 0.1 Å is
(a) $5.79 \times 10^{7}\ ms^{-1}$ (b) $5.79 \times 10^{8}\ ms^{-1}$
(c) $5.79 \times 10^{5}\ ms^{-1}$ (d) $5.79 \times 10^{6}\ ms^{-1}$

3. The number of nodal planes in a p_x orbital is
(a) one (b) two
(c) three (d) zero

4. The size of isoelectronic species C^{4-}, N^{3-} and Mg^{2+} is affected by
(a) nuclear charge (*Z*)
(b) principle quantum number
(c) interelectronic repulsion
(d) None of these

5. The ratio of magnetic moments of Fe(III) and Co(II) is
(a) 7 : 3 (b) 3 : 7
(c) $\sqrt{7} : \sqrt{3}$ (d) $\sqrt{3} : \sqrt{7}$

6. If the nitrogen atom has electronic configuration $1s^7$, it would have energy lower than that of the normal ground state configuration $1s^2 2s^2 2p^3$, because the electrons would be closer to the nucleus. Yet $1s^7$ is not observed because it violates.
(a) Heisenberg uncertainty principle
(b) Hund's rule
(c) Pauli exclusion principle
(d) Bohr postulate of stationary orbits

7. The orbital diagram in which the Aufbau principle is violated is :

	2s	**2p**
(a)	[↑↓]	[↑↓][↑][ ]
(b)	[↑]	[↑↓][↑][↑]
(c)	[↑↓]	[↑][↑][↑]
(d)	[↑↓]	[↑↓][↑][↑]

8. If a proton and α -particle are accelerated through the same potential difference, the ratio of de-Broglie wavelengths λ_p and λ_α is

(a) 3 (b) $2\sqrt{2}$
(c) 1 (d) 2

9. In hydrogen atom, energy of first excited state is –3.4 eV. Find out KE of the same orbit of Hydrogen atom

(a) +3.4 eV (b) +6.8 eV (c) –13.6 eV (d) +13.6 eV

10. The ions O^{2-}, F^-, Na^+, Mg^{2+} and Al^{3+} are isoelectronic. Their ionic radii show

(a) a decrease from O^{2-} to F^- and then increase from Na^+ to Al^{3+}
(b) a significant increase from O^{2-} to Al^{3+}
(c) a significant decrease from O^{2-} to Al^{3+}
(d) an increase from O^{2-} to F^- and then decrease from Na^+ to Al^{3+}

11. Kinetic energy of an electron in hydrogen atom increases after transition from an orbit n_1 to another orbit n_2. Then

(a) $n_1 < n_2$
(b) $n_1 > n_2$
(c) it is not possible to predict
(d) none of these

12. In Bohr series of lines of hydrogen spectrum, the third line from the red end corresponds to which one of the following inter-orbit jumps of the electron for Bohr orbits in an atom of hydrogen

(a) $5 \to 2$ (b) $4 \to 1$ (c) $2 \to 5$ (d) $3 \to 2$

13. Consider the ground state of Cr atom (Z= 24). The number of electrons with the azimuthal quantum numbers, $\ell = 1$ and 2 are, respectively

(a) 16 and 4 (b) 12 and 5 (c) 12 and 4 (d) 16 and 5

14. In a multi-electron atom, which of the following orbitals described by the three quantum members will have the same energy in the absence of magnetic and electric fields?

(A) $n=1, l=0, m=0$ (B) $n=2, l=0, m=0$
(C) $n=2, l=1, m=1$ (D) $n=3, l=2, m=1$
(E) $n=3, l=2, m=0$

(a) (D) and (E) (b) (C) and (D)
(c) (B) and (C) (d) (A) and (B)

15. Li and a proton are accelerated by the same potential, their de Broglie wavelengths λ_{Li} and λ_p have the ratio (assume $m_{Li} = 9m_p$)

(a) 1 : 2 (b) 1 : 4
(c) 1 : 1 (d) $1 : 3\sqrt{3}$

16. Which two orbitals are both located between the axes of coordinate system, and not along the axes?

(a) d_{xy}, d_{z^2} (b) d_{yz}, p_x
(c) $d_{x^2-y^2}$, p_z (d) none of these

17. Uncertainty in the position of an electron (mass = 9.1×10^{-31} kg) moving with a velocity 300 ms^{-1}, accurate upto 0.001% will be (h = 6.63×10^{-34} Js)

(a) 1.92×10^{-2} m (b) 3.84×10^{-2} m
(c) 19.2×10^{-2} m (d) 5.76×10^{-2} m

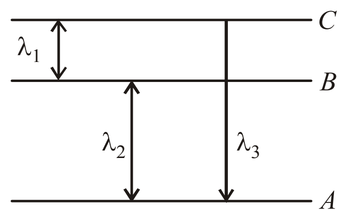
18. The number of radial nodes of 3*s* and 2*p* orbitals are respectively

(a) 2, 0 (b) 0, 2 (c) 1, 2 (d) 2, 1

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

(a) –3.4 eV (b) –4.2 eV
(c) –6.8 eV (d) Both (a) and (c)

20. Energy levels, *A, B, C*, of a certain atom correspond to increasing values of energy i.e., $E_A < E_B < E_C$. If $\lambda_1, \lambda_2, \lambda_3$ are the wave lengths of radiations corresponding to the transition from *C* to *B*, *B* to *A* and *C* to *A* respectively, which of the following statements is correct ?



(a) $\lambda_3 = \lambda_1 + \lambda_2$ (b) $\lambda_3 = \dfrac{\lambda_1\lambda_2}{\lambda_1 + \lambda_2}$
(c) $\lambda_1 + \lambda_2 + \lambda_3 = 0$ (d) $\lambda_3^2 = \lambda_1^2 + \lambda_2^2$

21. Number of protons, neutrons and electrons in the element $^{231}_{89}Y$ is

(a) 89, 231, 89 (b) 89, 89, 242
(c) 89, 142, 89 (d) 89, 71, 89

22. Which one of the following is not the characteristic of Planck's quantum theory of radiation ?

(a) The energy is not absorbed or emitted in whole number or multiple of quantum
(b) Radiation is associated with energy
(c) Radiation energy is not emitted or absorbed continuously but in the form of small packets called quanta
(d) This magnitude of energy associated with a quantum is proportional to the frequency.

23. The magnetic moment of a particular ion is $2\sqrt{6}$ B.M. The ion is

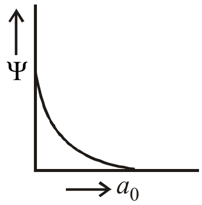
(a) Mn^{2+} (b) Fe^{3+}

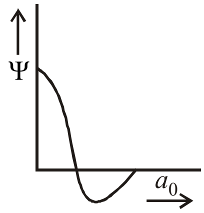
(c) Co^{2+} (d) Co^{3+}

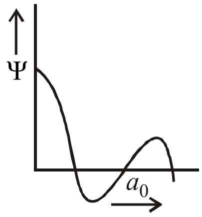
24. If the de-Broglie wavelength of a particle of mass m is 100 times its velocity, then its value in terms of its mass (m) and Planck's constant (h) is

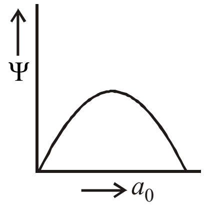
(a) $\frac{1}{10}\sqrt{\frac{m}{h}}$ (b) $10\sqrt{\frac{h}{m}}$ (c) $\frac{1}{10}\sqrt{\frac{h}{m}}$ (d) $10\sqrt{\frac{m}{h}}$

25. Which of the following graph correspond to one node

(a) 

(b) 

(c) 

(d) 

26. The five d-orbitals are designated as $d_{xy}, d_{yz}, d_{xz}, d_{x^2-y^2}$ and d_{z^2}. Choose the correct statement.

(a) The shapes of the first three orbitals are similar but that of the fourth and fifth orbitals are different

(b) The shapes of all five d-orbitals are similar

(c) The shapes of the first four orbitals are similar but that of the fifth orbital is different

(d) Ths shapes of all five d-orbitals are different

27. The magnetic moment of M^{x+} (atomic number of M = 25) is $\sqrt{15}$ B.M. The number of unpaired electrons and the value of x respectively are

(a) 4, 3 (b) 3, 4

(c) 3, 2 (d) 5, 2

28. The magnetic moments of Cu (Z = 29), Ti (Z = 22), and Cr (Z = 24) are in the ratio of

(a) $1:\sqrt{5}:4$ (b) $4:\sqrt{5}:1$

(c) 1 :2:6 (d) $1:\sqrt{3}:\sqrt{21}$

29. If uncertainty in position and momentum are equal, then uncertainty in velocity is :

(a) $\frac{1}{2m}\sqrt{\frac{h}{\pi}}$ (b) $\sqrt{\frac{h}{2\pi}}$ (c) $\frac{1}{m}\sqrt{\frac{h}{\pi}}$ (d) $\sqrt{\frac{h}{\pi}}$

30. The energy absorbed by each molecule (A_2) of a substance is 4.4×10^{-19} J and bond energy per molecule is 4.0×10^{-19} J. The kinetic energy of the molecule per atom will be:

(a) 2.2×10^{-19} J (b) 2.0×10^{-19} J

(c) 4.0×10^{-20} J (d) 2.0×10^{-20} J

31. A 0.66 kg ball is moving with a speed of 100 m/s. The associated wavelength will be ($h = 6.6 \times 10^{-34}$ Js) :

(a) 1.0×10^{-32}m (b) 6.6×10^{-32}m

(c) 6.6×10^{-34}m (d) 1.0×10^{-35}m

32. The energies E_1 and E_2 of two radiations are 25 eV and 50 eV, respectively. The relation between their wavelengths i.e., λ_1 and λ_2 will be :

(a) $\lambda_1 = \lambda_2$ (b) $\lambda_1 = 2\lambda_2$

(c) $\lambda_1 = 4\lambda_2$ (d) $\lambda_1 = \frac{1}{2}\lambda_2$

33. If n = 6, the correct sequence for filling of electrons will be :

(a) $ns \to (n-2)f \to (n-1)\,d \to np$

(b) $ns \to (n-1)\,d \to (n-2)f \to np$

(c) $ns \to (n-2)f \to np \to (n-1)\,d$

(d) $ns \to np \to (n-1)\,d \to (n-2)f$

34. According to the Bohr Theory, which of the following transitions in the hydrogen atom will give rise to the least energetic photon ?

(a) $n = 6$ to $n = 1$ (b) $n = 5$ to $n = 4$

(c) $n = 6$ to $n = 5$ (d) $n = 5$ to $n = 3$

35. Which of the following sets of quantum numbers represents the highest energy of an atom?

(a) $n = 3, l = 0, m = 0, s = +1/2$ (b) $n = 3, l = 1, m = 1, s = +1/2$

(c) $n = 3, l = 2, m = 1, s = +1/2$ (d) $n = 4, l = 0, m = 0, s = +1/2$

36. Which one of the following constitutes a group of the isoelectronic species?

(a) C_2^{2-}, O_2^-, CO, NO (b) $NO^+, C_2^{2-}, CN^-, N_2$

(c) $CN^-, N_2, O_2^{2-}, C_2^{2-}$ (d) N_2, O_2^-, NO^+, CO

37. The ionization enthalpy of hydrogen atom is 1.312×10^6 J mol^{-1}. The energy required to excite the electron in the atom from $n = 1$ to $n = 2$ is

(a) 8.51×10^5 J mol^{-1} (b) 6.56×10^5 J mol^{-1}

(c) 7.56×10^5 J mol^{-1} (d) 9.84×10^5 J mol^{-1}

38. In an atom, an electron is moving with a speed of 600 m/s with an accuracy of 0.005%. Certainity with which the position of the electron can be located is ($h = 6.6 \times 10^{-34}$ kg m^{2}s^{-1}, mass of electron, $e_m = 9.1 \times 10^{-31}$ kg)

(a) 5.10×10^{-3} m (b) 1.92×10^{-3} m

(c) 3.84×10^{-3} m (d) 1.52×10^{-4} m

39. The energy required to break one mole of Cl – Cl bonds in Cl_2 is 242 kJ mol^{-1}. The longest wavelength of light capable of breaking a single Cl – Cl bond is
($c = 3 \times 10^8$ ms^{-1} and $N_A = 6.02 \times 10^{23}$ mol^{-1}).

(a) 594 nm (b) 640 nm (c) 700 nm (d) 494 nm

40. Ionisation energy of He^+ is 19.6×10^{-18} J atom^{-1}. The energy of the first stationary state (n = 1) of Li^{2+} is

(a) 4.41×10^{-16} J atom^{-1} (b) -4.41×10^{-17} J atom^{-1}

(c) -2.2×10^{-15} J atom^{-1} (d) 8.82×10^{-17} J atom^{-1}

41. The electrons identified by quantum numbers n and ℓ :

(*A*) $n = 4, \ell = 1$ (*B*) $n = 4, \ell = 0$

(*C*) $n = 3, \ell = 2$ (*D*) $n = 3, \ell = 1$

can be placed in order of increasing energy as :

(a) $(C)<(D)<(B)<(A)$ (b) $(D)<(B)<(C)<(A)$

(c) $(B)<(D)<(A)<(C)$ (d) $(A)<(C)<(B)<(D)$

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

(a) $\frac{h^2}{4\pi^2 m a_0^2}$ (b) $\frac{h^2}{16\pi^2 m a_0^2}$ (c) $\frac{h^2}{32\pi^2 m a_0^2}$ (d) $\frac{h^2}{64\pi^2 m a_0^2}$

43. Given that the abundances of isotopes ^{54}Fe, ^{56}Fe and ^{57} Fe are 5%, 90% and 5%, respectively, the atomic mass of Fe is

(a) 55.85 (b) 55.95 (c) 55.75 (d) 56.05

44. Energy of an electron is given by $E = -2.178 \times 10^{-18} J\left(\frac{Z^2}{n^2}\right)$.

Wavelength of light required to excite an electron in an hydrogen atom from level n = 1 to n = 2 will be :

($h = 6.62 \times 10^{-34}$ Js and $c = 3.0 \times 10^8\ ms^{-1}$)

(a) 1.214×10^{-7} m (b) $2.816 \times 10.^{-7}$ m

(c) 6.500×10^{-7} m (d) 8.500×10^{-7} m

45. A 600 W mercury lamp emits monochromatic radiation of wavelength 331.3 nm. How many photons are emitted from the lamp per second?

$h = 6.62 \times 10^{-34}$ Js velocity of light $= 3 \times 10^8\ ms^{-1}$

(a) 1.0×10^{19} (b) 1.0×10^{23}

(c) 1.0×10^{21} (d) 2.0×10^{23}

46. The de-Broglie wavelength of a particle with mass 1 kg and velocity 100 m/s is

(a) 6.6×10^{-33} m (b) 6.6×10^{-36} m

(c) $3.3 \times 10^{+33}$ m (d) 3.3×10^{-36} m

DIRECTIONS for Qs. 47 to 49 : These are Assertion-Reason type questions. Each of these question contains two statements: Statement-1 (Assertion) and Statement-2 (Reason). Answer these questions from the following four options.

(a) Statement- 1 is True, Statement-2 is True, Statement-2 is a correct explanation for Statement -1

(b) Statement -1 is True, Statement -2 is True ; Statement-2 is NOT a correct explanation for Statement - 1

(c) Statement - 1 is True, Statement- 2 is False

(d) Statement -1 is False, Statement -2 is True

47. **Statement-1 :** The position of an electron can be determined exactly with the help of an electron microscope.

Statement-2 : The product of uncertainty in the measurement of its momentum and the uncertainty in the measurement of the position cannot be less than a finite limit.

48. **Statement-1 :** The radius of the first orbit of hydrogen atom is 0.529Å.

Statement-2 : Radius of each circular orbit (r_n) - 0.529Å (n^2/Z), where n = 1, 2, 3 and Z = atomic number.

49. **Statement-1 :** Nuclide $^{30}Al_{13}$ is less stable than $^{40}Ca_{20}$

Statement-2 : Nuclide having odd number of protons and neutrons are generally unstable.

50. **Statement-1 :** Angular momentum of an electron in any orbit is given by angular momentum $= \frac{n.h}{2\pi}$, where n is the principal quantum number.

Statement-2 : The principal quantum number, n, can have any integral value.

EXERCISE - 3

Exemplar & Past Years NEET/AIPMT Questions

Exemplar Questions

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

(a) Most of the space in the atom is empty.

(b) The radius of the atom is about 10^{-10} m while that of nucleus is 10^{-15} m.

(c) Electrons move in a circular path of fixed energy called orbits.

(d) Electrons and the nucleus are held together by electrostatic forces of attraction.

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

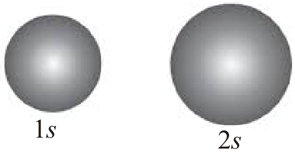
(a) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^8 4s^2$

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

3. The probability density plots of 1*s* and 2*s* 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) 1*s* and 2*s* 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 2*s* orbital decreases uniformly as distance from the nucleus increases.

4. Which of the following statement is not correct about the characteristics of cathode rays?

(a) They start from the cathode and move towards the anode.

(b) They travel in straight line in the absence of an external electrical or magnetic field.

(c) Characteristics of cathode rays do not depend upon the material of electrodes in cathode ray tube.

(d) Characteristics of cathode rays depend upon the nature of gas present in the cathode ray tube.

5. Which of the following statements about the electron is incorrect?
   (a) It is a negatively charged particle.
   (b) The mass of electron is equal to the mass of neutron.
   (c) It is a basic constituent of all the atoms.
   (d) It is a constituent of cathode rays.
6. Which of the following properties of atom could be explained correctly by Thomson model of atom?
   (a) Overall neutrality of atom
   (b) Spectra of hydrogen atom
   (c) Position of electrons, protons and neutrons in atom
   (d) Stability of atom
7. Two atoms are said to be isobars if
   (a) they have same atomic number but different mass number
   (b) they have same number of electrons but different number of neutrons
   (c) they have same number of neutrons but different number of electrons
   (d) sum of the number of protons and neutrons is same but the number of protons is different
8. The number of radial nodes for $3p$ orbital is ....
   (a) 3 (b) 4
   (c) 2 (d) 1
9. Number of angular nodes for $4d$ orbital is ....
   (a) 4 (b) 3
   (c) 2 (d) 1
10. Which of the following is responsible to rule out the existence of definite paths or trajectories of electrons?
    (a) Pauli's exclusion principle
    (b) Heisenberg's uncertainty principle
    (c) Hund's rule of maximum multiplicity
    (d) Aufbau principle
11. Total number of orbitals associated with third shell will be ....
    (a) 2 (b) 4
    (c) 9 (d) 3
12. Orbital angular momentum depends on ....
    (a) l (b) n and l
    (c) n and m (d) m and s
13. Chlorine exists in two isotopic forms Cl-37 and Cl-35, but its atomic mass is 35.5. This indicates the ratio of Cl-37 and Cl-35 is approximately
    (a) 1 : 2 (b) 1 : 1
    (c) 1 : 3 (d) 3 : 1
14. 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+}
15. For the electrons of oxygen atom, which of the following statements is correct?
    (a) Z_{eff} for an electron in a $2s$ orbital is the same as Z_{eff} for an electron in a $2p$ orbital.
    (b) An electron in the $2s$ orbital has the same energy as an electron in the $2p$ orbital.
    (c) Z_{eff} for an electron in $1s$ orbital is the same as Z_{eff} for an electron in a $2s$ orbital.
    (d) The two electrons present in the $2s$ orbital have spin quantum numbers m_s but of opposite sign.
16. If travelling at same speeds, which of the following matter waves have the shortest wavelength?
    (a) Electron (b) Alpha particle (He^{2+})
    (c) Neutron (d) Proton

NEET/AIPMT (2013-2017) Questions

17. According to law of photochemical equivalence the energy absorbed (in ergs/mole) is given as ($h = 6.62 \times 10^{-27}$ ergs, $c = 3 \times 10^{10}$ cm s^{-1}, $N_A = 6.02 \times 10^{23}$ mol^{-1}) ***[NEET Kar. 2013]***
    (a) $\frac{1.196 \times 10^{16}}{\lambda}$ (b) $\frac{1.196 \times 10^{8}}{\lambda}$
    (c) $\frac{2.859 \times 10^{5}}{\lambda}$ (d) $\frac{2.859 \times 10^{16}}{\lambda}$
18. Based on equation $E = -2.178 \times 10^{-18} \text{ J}\left(\frac{Z^2}{n^2}\right)$, certain conclusions are written. Which of them is not correct ? ***[2013]***
    (a) Larger the value of n, the larger is the orbit radius.
    (b) Equation can be used to calculate the change in energy when the electron changes orbit.
    (c) For n = 1, the electron has a more negative energy than it does for n = 6 which mean that the electron is more loosely bound in the smallest allowed orbit.
    (d) The negative sign in equation simply means that the energy or electron bound to the nucleus is lower than it would be if the electrons were at the infinite distance from the nucleus.
19. What is the maximum numbers of electrons that can be associated with the following set of quantum numbers? n = 3, l = 1 and m = –1 ***[2013]***
    (a) 6 (b) 4 (c) 2 (d) 10
20. The value of Planck's constant is 6.63×10^{-34} Js. The speed of light is 3×10^{17} nm s^{-1}. Which value is closest to the wavelength in nanometer of a quantum of light with frequency of 6×10^{15} s^{-1}? ***[2013]***
    (a) 25 (b) 50 (c) 75 (d) 10
21. What is the maximum number of orbitals that can be identified with the following quantum numbers? ***[2014]***
    n = 3, ℓ = 1, m_ℓ = 0
    (a) 1 (b) 2 (c) 3 (d) 4
22. Calculate the energy in joule corresponding to light of wavelength 45 nm :
    (Planck's constant h = 6.63×10^{-34} Js; speed of light c = 3×10^8 ms^{-1}) ***[2014]***
    (a) 6.67×10^{15} (b) 6.67×10^{11}
    (c) 4.42×10^{-15} (d) 4.42×10^{-18}
23. Two electrons occupying the same orbital are distinguished by
    (a) Principal quantum number ***[2016]***
    (b) Magnetic quantum number
    (c) Azimuthal quantum number
    (d) Spin quantum number
24. Which one is the wrong statement ? ***[2017]***
    (a) The uncertainty principle is $\Delta E \times \Delta t \geq h / 4\pi$
    (b) Half filled and fully filled orbitals have greater stability due to greater exchange energy, greater symmetry and more balanced arrangement.
    (c) The energy of $2s$ orbital is less than the energy of $2p$ orbital in case of Hydrogen like atoms
    (d) de-Broglies's wavelength is given by $\lambda = \frac{h}{mv}$, where m = mass of the particle, ν = group velocity of the particle

Hints & Solutions

EXERCISE - 1

1. (a) We know that ions which have the same number of electrons are called isoelectronic. We also know that both CO and CN^- have 14 electrons, therefore these are isoelectronic.
2. (c) The molecule which contains same number of electrons are called isoelectronic. eg. $N_2 = CO = CN^- = O_2^{+2}$ $= 14e$
3. (b) The plot given in option (b) is the correct sketch of wave function (ψ) for a 2s- electron as a function of its distance from the nucleus (r). As ψ has no significate, so the plot of ψ versus r also has no significance.
4. (c) Three 2p orbitals of one atom and three 2p orbitals of another atom give rise to four (π 2p) and two (σ 2p) molecular orbitals.
   Among $2p_x$, $2p_y$ and $2p_z$ orbitals of two atoms $2p_z$ combines with $2p_z$ along z-aixs give rise to two σ 2p molecular orbitals, while $2p_x$ combines with $2p_x$ and $2p_y$ combines with $2p_y$ give rise to two $\pi\ 2p_x$ and two $\pi\ 2p_y$ molecular orbitals.
5. (d) Those species which contain same no. of electrons are called isoelectronic species. C_2H_4 has 16 electrons. The no. of electrons in other given species are
   $O_2^+ = 15e^-; CN^- = 14e^-; N_2^+ = 13e^-;$
   $O_2 = 16e^-$
   So, O_2 is isoelectronic with C_2H_4.
6. (c) Cathode rays are made up of negatively charged particles (electrons) which are deflected by both the electric and magnetic fields.
7. (d) Number of neutrons = Mass number – Atomic number = 70 – 30 = 40.
8. (b) e/m for $He^+ = 1/4$ $\quad$ e/m for $He^{2+} = 2/4$
   e/m for $D^+ = 1/2$ $\quad$ e/m for $H^+ = 1/1$
   $\therefore$ Value of e/m is highest for H^+.
9. (d) Size of H^+ (proton) is very small, so its hydration energy is very large.

   Hydration energy $\propto \dfrac{1}{\text{Size}}$
10. (d) Proton is the nucleus of H-atom (H-atom devoid of its electron)
11. (d)

	F^-	O^{2-}	Al^{3+}	N^{3-}
No. of e^-s	10	10	10	10
Nuclear charge (Number of protons)	9	8	13	7

   All the four given species are isoelectronic and size of isoelectronic species decreases with increase in nuclear charge. Among the four concerned atoms, N has lowest atomic number (nuclear charge), hence N^{3-} ion will be largest in size.
12. (b) For p orbital with n = 6 and m = 0 indicates $6p_z$ orbital. It contains maximum of 2 electrons with spins opposite to each other.
13. (d) NaCl: No. of e^- in Na^+ = At. No. of Na–1 = 11 – 1 = 10
   No. e^- in Cl^- = At. No. of Cl + 1 = 17 + 1 = 18
   CsF : No. of e^- in Cs^+ = 55 – 1 = 54
   No. of e^- in F^- = 9 + 1 = 10
   NaI : No. of e^- in Na^+ = 11 – 1 = 10
   No. of e^- in I^- = 53 + 1 = 54
   K_2S : No. of e^- in K^+ = 19 – 1 = 18
   No. of e^- in S^{2-} = 16 + 2 = 18
14. (c)
15. (d) $\dfrac{e}{m}$ for (i) neutron $= \dfrac{0}{1} = 0$
   (ii) α-particle $= \dfrac{2}{4} = 0.5$
   (iii) proton $= \dfrac{1}{1} = 1$
   (iv) electron $= \dfrac{1}{1/1837} = 1837$
16. (c) No. of neutrons = Mass number – Atomic number $= W - N$.
17. (c) Energy of electron in 2nd orbit of $Li^{+2} = -13.6\dfrac{z^2}{n^2}$
   $$= \frac{-13.6 \times (3)^2}{(2)^2} = -30.6 \text{ eV}$$
   Energy required = 0 – (–30.6) = 30.6 eV
18. (b) $E = mc^2$ (Einstien relation) .... (1)
   $E = h\nu$ .... (2)
   From (1) and (2),
   $mc^2 = h\nu$
   $\Rightarrow mc^2 = \dfrac{hc}{\lambda} \quad \left(\because \nu = \dfrac{c}{\lambda}\right)$
   $\Rightarrow m c \lambda = h$
   $\Rightarrow \lambda = \dfrac{h}{mc}$
   $\Rightarrow \lambda = \dfrac{h}{p} \quad (\because mc = p)$
   $\Rightarrow p = \dfrac{h}{\lambda}$
19. (b)
   - No. of electrons in N^{3-}, O^{2-} and Cr^{3-} are respectively 10, 10 and 27 so these species are not isoelectronic with each other.
   - No. of neutrons in $_{14}Si^{30}$, $_{15}P^{31}$ and $_{16}S^{32}$ are respectively 16, 16 and 16. So, $_{14}S^{30}$, $_{15}P^{31}$ and $_{16}S^{32}$ are isotones.

- Atoms of the same element having same atomic no. and different mass no. are called isotopes. So $_{20}Ca^{40}$ and $_{19}K^{40}$ are not isotopes.
- Isobars are the atoms of different elements having same mass no. and different atomic no. So, $_8O^{16}$, $_8O^{17}$ and $_8O^{18}$ are isobars.

20. (b) Both He and Li^+ contain 2 electrons each therefore their spectrum will be similar.

21. (d) This statement is known as uncertainty principle which was given by Heisenberg it is not a Bohr's postulate.

22. (a) For hydrogen atom $(n) = 1$ (due to ground state)
Radius of hydrogen atom (r) = 0.53 Å.
Atomic number of Li (Z) = 3.

$$\text{Radius of } Li^{2+} \text{ ion} = r_1 \times \frac{n^2}{Z} = 0.53 \times \frac{(1)^2}{3} = 0.17$$

23. (d) Given : Radius of hydrogen atom = 0.530 Å, Number of excited state $(n) = 2$ and atomic number of hydrogen atom $(Z) = 1$. We know that the Bohr radius.

$$(r) = \frac{n^2}{Z} \times \text{Radius of atom} = \frac{(2)^2}{1} \times 0.530$$

$$= 4 \times 0.530 = 2.12 \text{ Å}$$

24. (a) He_2^- contains 5 electrons. So, the molecular orbital electronic configuration is
$\sigma 1s^2, \sigma^* 1s^2, \sigma 2s^1$

25. (c) Principal quantum number (n) gives the information about the size of electron cloud i.e., the approximate distance of electron cloud from the nucleus.

26. (d) The given statement is related to Aufbau principle. According to this principle the electrons first occupy the orbital with lowest energy available to them and then enter into higher energy orbitals only when the lower energy orbitals are filled.

27. (b) We know that $E_n = \frac{-1312}{n^2}$ kJ mol^{-1}

$n = 4$ (Fourth Bohr orbit)

$$\text{Given } E_4 = \frac{-1312}{4^2} = -82 \text{ kJ mol}^{-1}$$

28. (a) For s-electron, $\ell = 0$

$$\therefore \text{ Orbital angular momentum} = \sqrt{0(0+1)}\frac{h}{2\pi} = 0$$

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

$$\frac{1}{\lambda} = 1.097 \times 10^7 \left(\frac{1}{1} - \frac{1}{\infty}\right) = 1.097 \times 10^7 \text{ m}^{-1}$$

$$\lambda = 91.15 \times 10^{-9} \text{ m} \approx 91 \text{nm}$$

30. (c) $\nu = \frac{1}{h} \times \text{IE} \times \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$

$$= \frac{2.18 \times 10^{-18}}{6.625 \times 10^{-34}} \times \left[\frac{1}{1} - \frac{1}{16}\right] = 3.08 \times 10^{15} \text{ s}^{-1}$$

31. (b) Angular momentum of an electron in nth orbit is given by

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

For $n = 5$, we have

$$\text{Angular momentum of electron} = \frac{5h}{2\pi} = \frac{2.5h}{\pi}$$

32. (c) ΔE for two energy levels $= 21.79\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$ J/atom

33. (c) For every ℓ, quantum number 'm' can take values from $-\ell$ to $+\ell$. So for $\ell = 2$, $m = -2, -1, 0, +1, +2$

34. (c)

(i)	$\overset{+}{C}H_3$	8 electrons
(ii)	$H_3\overset{+}{O}$	10 electrons
(iii)	NH_3	10 electrons
(iv)	$\overset{-}{C}H_3$	10 electrons

So, (ii), (iii) and (iv) are isoelectronic species with 10 electrons each.

35. (a) $\Delta E = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$;

First line in Balmer series results in the transition : $n_2 = 3$ to $n_1 = 2$.

36. (d) Those orbitals which are equal in energies are called degenerate orbitals. All sp^3 orbitals are degenerate as these are equal in energies.

37. (a) Rubidium is an alkali metal. It has one valence electron in 5s orbital. So, principal quantum number, n = 5
azimuthal quantum number $\ell = 0$

Note :

orbital :	s	p	q	f
value of ℓ :	0	1	2	3

magnetic quantum number,
$m = -\ell$ to $\ell = 0$

spin quantum number, $s = +\frac{1}{2}$

So, correct set of quantum number is 5, 0, 0, $+\frac{1}{2}$

38. (c) Zeeman effect refers to splitting of lines of an emission spectrum in magnetic field.

39. (b) If we assume the atom to be hydrogen like, energy of nth energy level

$E_n = \frac{E_1}{n^2}$ where E_1 is energy of first energy level

$$E_2 = -\frac{E_1}{2^2} = -\frac{E_1}{4} = \frac{-21.79 \times 10^{-12}}{4}$$

$$= -5.447 \times 10^{-12} \text{ erg per atom.}$$

40. (b) 41. (d)

42. (d) The nucleus occupies much smaller volume compared to the volume of the atom.

43. (a) Energy of an electron $E = \frac{-E_0}{n^2}$

For energy level $(n = 2)$

$$E = -\frac{13.6}{(2)^2} = \frac{-13.6}{4} = -3.4 \text{ eV}.$$

44. (c) The energy of photon,

$$E = \frac{hc}{\lambda} = 3.03 \times 10^{-19}$$

or $\lambda = \frac{6.626 \times 10^{-34} \times 3 \times 10^8}{3.03 \times 10^{-19}}$

or $\lambda = \frac{19.878}{3.03} \times 10^{-7} = 6.56 \times 10^{-7}\text{m} = 656 \text{ nm}$

45. (c) K.E. of emitted electron

$= h\nu - h\nu_0$ (i.e. smaller than $h\nu$).

46. (a) Given mass of an electron(m) $= 9.1 \times 10^{-28}$ g;

Velocity of electron $(v) = 3 \times 10^4$ cm/s;

Accuracy in velocity $= 0.001\% = \frac{0.001}{100}$;

Actual velocity of the electron

$(\Delta v) = 3 \times 10^4 \times \frac{0.001}{100} = 0.3$ cm/s.

Planck's constant $(h) = 6.626 \times 10^{-27}$ erg-sec.

$\therefore$ Uncertainty in the position of the electron

$$(\Delta x) = \frac{h}{4\pi m \Delta v} = \frac{6.626 \times 10^{-27} \times 7}{4 \times 22 \times (9.1 \times 10^{-28}) \times 0.3}$$

$= 1.93$ cm

47. (b) Atomic number of the given element = 10
Electronic configuration $= 1s^2, 2s^2p^6$
$1s^2 2s^2 2p^6$ is electronic configuration of Ne.
$1s^2 2s^2 2p^5 3s^1$ is excited oxidation state.

48. (a) Positron is electron with positive charge, $_{+1}e^0$

49. (d) By Heisenberg uncertainty Principle $\Delta x \times \Delta p = \frac{h}{4\pi}$ (which is constant)
As Δx for electron and helium atom is same thus momentum of electron and helium will also be same therefore the momentum of helium atom is equal to 5×10^{-26} kg. m.s^{-1}.

50. (a) $\lambda = h/mv$; for the same velocity, λ varies inversely with the mass of the particle.

51. (d) $\lambda = \frac{\text{h}}{\text{mv}} = \frac{6.6 \times 10^{-34}}{60 \times 10^{-3} \times 10} = 10^{-33}$ m

52. (d) (i) Interference and diffraction support the wave nature of electron.

(ii) $E = mc^2$ supports the particle nature of electron.

(iii) $E = h\nu = \frac{hc}{\lambda}$ is de-Broglie equation and it supports both wave nature and particle nature of electron.

53. (c) Radius of nth orbit $r_n = \frac{a_0}{Z} \times n^2$; $\therefore r_3 = 9a_0$

54. (b) $r_n = \frac{a_0 n^2}{Z}, r_1 = \frac{a_0 (1)^2}{1}$ for hydrogen

$r_2 = \frac{a_0 (1)^2}{1}$ for deuterium $\therefore \frac{r_1}{r_2} = \frac{1}{1} = 1:1$

55. (b) For electron in the ground state,

$$mvr = \frac{h}{2\pi} \Rightarrow mv = \frac{h}{2\pi r}$$

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

So, $\frac{h}{\lambda} = \frac{h}{2\pi r} \Rightarrow \lambda = 2\pi r$

$\lambda = 2 \times 3.14 \times 0.53 \text{Å} = 3.328 \text{Å}$

$= 3.328 \times 10^{-10} m$

$= 0.3328 \times 10^{-9} m = 0.3328 \ nm$

56. (c) We know that $\lambda = \frac{h}{mu}$; $\therefore m = \frac{h}{u\lambda}$
The velocity of photon (u) = 3×10^8 m sec^{-1}
$\lambda = 1.54 \times 10^{-8}$ cm $= 1.54 \times 10^{-10}$ meter

$$\therefore m = \frac{6.626 \times 10^{-34} \text{ Js}}{1.54 \times 10^{-10} \text{m} \times 3 \times 10^8 \text{m sec}^{-1}}$$

$= 1.4285 \times 10^{-32}$ kg

57. (a)

58. (a) $\lambda = \frac{\text{h}}{\text{mv}} = \frac{\text{h}}{\sqrt{2\text{mE}}}$

$$\lambda = \frac{6.6 \times 10^{-34}}{\sqrt{2 \times 9.1 \times 10^{-31} \times 4.55 \times 10^{-25}}}$$

$= \frac{6.6 \times 10^{-34}}{9 \times 10^{-28}} = 0.727 \times 10^{-6}$

$= 7.27 \times 10^{-7}$ metres

59. (c) Noble gas must have 8 electrons in its outer most shell. Among the given configuration only option (c) has 8 electron in its outermost orbital.
i.e. the configuration is 2, 8, 18, 18, 8. So option (c) is the right answer.

60. (c) $N(7) = 1s^2 \, 2s^2 \, 2p^3$
$N^{2+} = 1s^2, 2s^2 \, 2p_x^1$
Unpaired electrons = 1.

61. (c) n = 2, l = 1 means 2p–orbital. Electrons that can be accommodated = 6 as p sub-shell has 3 orbital and each orbital contains 2 electrons.

62. (a) $Cu^+ = 29 - 1 = 28 \, e^-$
thus the electronic confingration of Cu^+ is

$$Cu^+(28) = 1s^2 \; 2s^2 \; 2p^6 \; \underbrace{3s^2 \; 3p^6 \; 3d^{10}}_{18e^-}$$

63. (d) $\ell = 3$ means f-subshell. Maximum no. of electrons $= 4\ell + 2 = 4 \times 3 + 2 = 14$

64. (c) If $m = -3$; $\ell = 3$,
[m ranges from $-\ell$ to $+\ell$]
So n = 4 as nature of ℓ ranges from 0 to (n – 1).
So option (c) is the answer.

65. (b) The sub-shell are $3d$, $4d$, $4p$ and $4s$, $4d$ has highest energy as $n + \ell$ value is maximum for this.

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

67. (c) $\lambda = \frac{h}{mv}$; $\therefore \lambda \propto \frac{1}{v}$ hence answer (c).

68. (d) According to Hund's rule electron pairing in P, d and f orbitals cannot occur until each orbital of a given subshell contains one electron each or is singly occupied.

69. (d) We know that atomic number of gadolinium is 64. Therefore the electronic configuration of gadolinium is [Xe] $4f^7\,5d^1\,6s^2$. Because the half filled and fully filled orbitals are more stable.

70. (a) Quantum number $n = 3$, $l = 2$, $m = +2$ represent an orbital with

$$s = \pm\frac{1}{2} \quad \left(3d_{xy} \text{ or } 3d_{x^2-y^2}\right)$$

which is possible only for one electron.

71. (a) No. of radial nodes in $3p$-orbital $= (n - \ell - 1)$
[for p ortbital $\ell = 1$]
$= 3 - 1 - 1 = 1$

72. (b) The sub-shell with lowest value of $(n + \ell)$ is filled up first. When two or more sub-shells have same $(n + \ell)$ value the subshell with lowest value of 'n' is filled up first therefore the correct order is

orbital	$4s$	$3d$	$4p$	$5s$	$4d$
$n+\ell$	4 + 0	3 + 2	4 + 1	5 + 0	4 + 2
value	= 4	=5	= 5	=5	= 6

73. (d) The orbitals which have same energy are called degenerate orbitals eg. p_x, p_y and p_z.

74. (a) Using the relation,

$$\Delta x.\ \Delta v = \frac{h}{4\pi m} \quad \text{[Heisenberg' uncertainty principle]}$$

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

$$\text{Thus, } \Delta x_A = \frac{h}{4\pi \times 0.05 \times m} \quad \ldots\text{(i)}$$

$$\Delta x_B = \frac{h}{4\pi \times 0.02 \times 5m} \quad \ldots\text{(ii)}$$

Dividing (i) by (ii), we get

$$\frac{\Delta x_A}{\Delta x_B} = \frac{0.02 \times 5}{0.05} = \frac{10}{5} \text{ or } 2$$

75. (b) Magnetic quantum no. represents the orientation of atomic orbitals in an atom. For example p_x, p_y & p_z have orientation along X-axis, Y-axis & Z-axis

76. (c) $Fe^{++}(26 - 2 = 24) = 1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^0\,3d^6$ hence no. of d electrons retained is 6. [Two $4s$ electron are removed]

77. (d) Angular speed is $\frac{V}{r}$. $V_n \propto \frac{1}{n}$ and $r_n \propto n^2$.
$\therefore$ Angular speed is inversely proportional to n

78. (b) $E = h\nu = h\,\frac{c}{\lambda}$

$$\text{So, } E \propto \frac{1}{\lambda}$$

$\lambda_1 : \lambda_2 = 3000\,\text{A} : 6000\,\text{A} = 1 : 2$
Hence, $E_1 : E_2 = 2 : 1$

79. (b) $n = 4$ represents 4^{th} orbit
$\ell = 3$ represents f subshell
$m = -2$ represents orientation of f-orbital
$s = 1/2$ represents direction of spin of electron.
$\therefore$ The orbital is $4f$.

80. (d) Electronic configuration of element with atomic no 24 is
$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1, 3d^5$
Exactly half filled d-orbital has extra stability.

81. (b) For $4d$ orbitals, $n = 4$, $l = 2$

$$\begin{bmatrix} \text{For } s \text{ orbital } l = 0 \\ \text{For } p \text{ orbital } l = 1 \\ \text{For } d \text{ orbital } l = 2 \end{bmatrix}$$

$m = -2, -1, 0, +1$ or $+2$

$$s = +\frac{1}{2} \text{ and } -\frac{1}{2}$$

Thus choice b having $n = 4$, $l = 2$, $d = 1$ and $s = \frac{1}{2}$ is correct.

82. (c) $$\lambda = \frac{h}{mv} = \frac{6.63 \times 10^{-34}}{0.200 \times \frac{5}{3600}} = 2.4 \times 10^{-30}\text{ m}$$

83. (c)

84. (b) Number of nodal planes in d orbitals is 2.

85. (c) Atomic number 9 is for F and ion is F^-.

86. (a) $\Delta p = m\Delta v$
Substituting the given values of Δx and m, we get
1×10^{-18} g cm s^{-1} $= 9 \times 10^{-28}$ g $\times \Delta v$

$$\text{or } \Delta v = \frac{1 \times 10^{-18}}{9 \times 10^{-28}}$$

$= 1.1 \times 10^9$ cm s^{-1} $\simeq 1 \times 10^9$ cm s^{-1}
i.e. option (a) is correct.

87. (d) The number of sub shell is $(2l + 1)$. The maximum number of electrons in the sub shell is $2(2l + 1) = (4l + 2)$

88. (b) $m = -l$ to $+l$, through zero thus for $l = 2$, values of m will be $-2, -1, 0, +1, +2$.
Therefore for $l = 2$, m cannot have the value -3.

89. (b) Total no. of atomic orbitals in a shell $= n^2$.
Given $n = 4$; Hence number of atomic orbitals in 4^{th} shell will be 16.

90. (a) $\lambda = \frac{h}{mv} = \frac{6.63\times10^{-34}}{1.67\times10^{-27}\times1\times10^{3}}$

$= 3.97\times10^{-10}$ meter $= 0.397$ nanometer ≈ 0.40 nm

91. (a) For He^{+}

$\bar{v} = \frac{1}{\lambda} = R_H Z^2\left(\frac{1}{2^2}-\frac{1}{4^2}\right)$

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

For H

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

For same frequency,

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

$\therefore \frac{1}{n_1^2}-\frac{1}{n_2^2} = \frac{1}{1^2}-\frac{1}{2^2}$

$\therefore n_1 = 1$ & $n_2 = 2$

92. (c) Among isoelectronic species ionic radii increases as the charge increases.
Order of ionic radii $Ca^{2+} < K^{+} < Cl^{-} < S^{2-}$
The number of electrons remains the same but nuclear charge increases with increase in the atomic number causing decrease in size.

93. (d) Given, $v_A = 0.1\ ms^{-1}$ and $v_B = 0.05\ ms^{-1}$ also, $m_B = 5m_A$

de-Broglie wavelength, $\lambda = \frac{h}{mv}$

$\therefore \frac{\lambda_A}{\lambda_B} = \frac{h/m_Av_A}{h/m_Bv_B} = \frac{m_Bv_B}{m_Av_A}$

$= \frac{5m_A\times0.05}{m_A\times0.1} = 5\times0.5 = 2.5 = 5/2$

$\therefore \lambda_A : \lambda_B = 5:2$

94. (c) Series limit is the last line of the series, i.e. $n_2 = \infty$.

$\therefore \bar{\upsilon} = \frac{1}{\lambda} = R\left[\frac{1}{n_1^2}-\frac{1}{n_2^2}\right] = R\left[\frac{1}{n_1^2}-\frac{1}{\infty^2}\right] = \frac{R}{n_1^2}$

$\because \bar{\upsilon} = 12186.3 = \frac{109677.76}{n_1^2}$

$\Rightarrow n_1^2 = \frac{109677.76}{12186.3} = 9 \Rightarrow n_1 = 3$

$\therefore$ The line belongs to Paschen series.

95. (c) The electronic configuration of the given species is

Mg^{2+}	:	$1s^2\,2s^2\,2p^6$
F	:	$1s^2 2s^2\,2p^5$
N	:	$1s^2\,2s^2\,2p^3$
S^{2-}	:	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6$
Ti^{3+}	:	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^1$

Therefore, N has the highest number of unpaired electrons.

96. (c) (1) Beyond a certain wavelength the line spectrum becomes band spectrum.
(2) For Balmer series $n_1 = 2$
(3) For calculation of longest wavelength use nearest value of n_2. Hence for longest wavelength in Balmer series of hydrogen spectrum,
$n_1 = 2$ & $n_2 = 3$.

97. (b) $^{39}K^{+}$ and $^{40}K^{+}$ contains same number of electrons so they are isoelectronic. They have same atomic numbers but different mass numbers so they are also isotopes.

98. (b) Value of $l = 0$ ....$(n-1)$
l cannot be equal to n.

99. (d) de Broglie wavelength $\lambda = \frac{h}{mv}$

$\frac{\lambda_1}{\lambda_2} = \frac{m_2v_2}{m_1v_1}$; $\frac{1}{4} = \frac{1}{9}\times\frac{v_2}{v_1}$

$\frac{v_2}{v_1} = \frac{9}{4}$

$\frac{v_1}{v_2} = \frac{4}{9}$

$KE = \frac{1}{2}mv^2$

$\frac{KE_1}{KE_2} = \frac{m_1}{m_2}\times\frac{v_1^2}{v_2^2} = \frac{9}{1}\times\left(\frac{4}{9}\right)^2 = \frac{16}{9}$

100. (a) $\bar{\upsilon} = \frac{1}{\lambda} = R\left[\frac{1}{n_1^2}-\frac{1}{n_2^2}\right]$

For second line in lyman series
$n_2 = 3$

$\therefore \frac{1}{\lambda} = R\left[\frac{1}{1^2}-\frac{1}{3^2}\right] = R\left[\frac{1}{1}-\frac{1}{9}\right] = \frac{8R}{9}$

EXERCISE - 2

1. (d) $\frac{1}{\lambda} = R_H Z^2\left(\frac{1}{n_1{}^2}-\frac{1}{n_2{}^2}\right)$

To calculate shorst wavelength take $n_2 = \infty$ and longest wavelength take nearest value of n_2.
For H-atom,

$\frac{1}{\lambda_{shortest}}$ $n_2 = \infty$, $Z = 1$, $n_1 = 1$

$\therefore \frac{1}{x} = R_H$ (Lyman series)

For $\frac{1}{\lambda_{longest}}$ for Li^{2+}, $Z = 3$, $n_1 = 2$, $n_2 = 3$ (Balmer series)

$$\frac{1}{\lambda_{longest}} = \frac{1}{x} \times 3^2\left(\frac{1}{2^2} - \frac{1}{3^2}\right) = \frac{5}{4x}$$

$$\therefore \quad \lambda_{longest} = \frac{4x}{5}$$

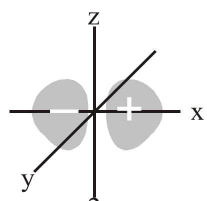
2. (d) We know that $\Delta x . \Delta p \geq \frac{h}{4\pi}$

$$\Delta x . m\Delta v > \frac{h}{4\pi} \qquad \Delta v > \frac{h}{4\pi \Delta x m}$$

$$\Delta v = \frac{6.626 \times 10^{-34}}{4\pi \times 0.1 \times 10^{-10} \times 9.11 \times 10^{-31}}$$

$$= \frac{66}{4\pi \times 9} \times 10^7 = 5.79 \times 10^6 \, m/sec$$

3. (a) p_x orbital being dumbell shaped, have number of nodal planes = 1



$2p_x$

4. (a) Because number of protons (nuclear charge) is different while number of electrons is same in isoelectronic species.

5. (c) Fe (III) = [Ar] $3d^5$ unpaired electrons = 5;

Magnetic moment = $\sqrt{5(5+2)}$;

Ratio = $\sqrt{7} : \sqrt{3}$

Co(II) = [Ar] $3d^7$ unpaired electrons = 3;

Magnetic moment = $\sqrt{3(3+2)}$

Ratio = $\sqrt{7} : \sqrt{3}$

6. (c) As per Pauli exclusion principle "no two electrons in the same atom can have all the four quantum numbers equal or an orbital cannot contain more than two electrons and it can accommodate two electrons only when their directions of spin are opposite".

7. (b) According to Aufbau principle, the orbital of lower energy (2*s*) should be fully filled before the filling of orbital of higher energy starts.

8. (b) $\lambda_p = \frac{h}{\sqrt{2eVm_p}}$;

$$\lambda_{He^{2+}} = \frac{h}{\sqrt{2 \times 2eVm_{He^{2+}}}} = \frac{h}{\sqrt{2 \times 2eV \times 4m_p}}$$

$$\therefore \qquad \frac{\lambda_p}{\lambda_{He^{2+}}} = 2\sqrt{2}$$

9. (a) Total energy (E_n) = K.E + P.E in first excited state

$$E = \frac{1}{2}mv^2 + \left[-\frac{ze^2}{r}\right] = +\frac{1}{2}\frac{ze^2}{r} - \frac{ze^2}{r}$$

$$-3.4\, eV = -\frac{1}{2}\frac{ze^2}{r}$$

$$\therefore \text{K.E} = \frac{1}{2}\frac{ze^2}{r} = +3.4 \text{ eV}$$

10. (c) Amongst isoelectronic species, ionic radii of anion is more than that of cations. Further size of anion increase with increase in –ve change and size of cation decrease with increase in +ve charge. Hence ionic radii decreases from O^{2-} to Al^{+++}.

11. (b) K.E. of an electron in a Bohr orbit is equal to the magnitude of the total energy but of opposite sign. So it varies inversly to the square of principal quantum number.

12. (a) The lines falling in the visible region comprise Balmer series. Hence the third line from red would be $n_1 = 2$, $n_2 = 5$ i.e, $5 \rightarrow 2$.

13. (b) Electronic configuration of Cr atom $(Z = 24)$

$$= 1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^5, 4s^1$$

when $\ell = 1$, *p*-subshell,
Numbers of electrons = 12
when $\ell = 2$, *d*-subshell,
Numbers of electrons = 5

14. (a) The energy of an orbital is given by $(n + l)$ in (D) and (E). $(n + l)$ value is (3 + 2) = 5 hence they will have same energy, since there n values are also same.

$(h = 6.63 \times 10^{-34}$ Js)

15. (d) $\lambda_p = \frac{h}{\sqrt{2eVm_p}}; \lambda_{Li} = \frac{h}{\sqrt{2 \times 3eVm_{Li}}}$

$$= \frac{h}{\sqrt{2 \times 3eV \times 9m_p}}$$

$$\text{Hence, } \frac{\lambda_{Li^{3+}}}{\lambda_p} = \sqrt{\frac{2eVm_p}{2 \times 3eV \times 9m_p}} = \frac{1}{3\sqrt{3}}$$

16. (d)

17. (a) Given m = 9.1×10^{-31}kg
$h = 6.6 \times 10^{-34}$Js

$$\Delta v = \frac{300 \times .001}{100} = 0.003 ms^{-1}$$

From Heisenberg's uncertainity principle

$$\Delta x = \frac{6.62 \times 10^{-34}}{4 \times 3.14 \times 0.003 \times 9.1 \times 10^{-31}} = 1.92 \times 10^{-2} m$$

18. (a) Number of radial nodes = $(n-l-1)$
For $3s$: $n=3, l=0$ (Number of radial node = 2)
For $2p$: $n=2, l=1$ (Number of radial node = 0)

19. (a) The energy of an electron on Bohr orbits of hydrogen atoms is given by the expression

$$E_n = -\frac{\text{Constant}}{n^2}$$

Where n takes only integral values. For the first Bohr orbit, $n=1$ and it is given that $E_1 = -13.6$ eV

Hence $E_n = -\dfrac{13.6\text{eV}}{n^2}$

Now out of given values of energy, only – 3.4 eV can be obtained by substituting $n=2$ in the above expression.

20. (b) From the given data, we have

$$(E_C - E_B) + (E_B - E_A) = (E_C - E_A)$$

or $\left(\dfrac{hc}{\lambda_1} + \dfrac{hc}{\lambda_2}\right) = \dfrac{hc}{\lambda_3}$ $\left[\text{or } \dfrac{1}{\lambda_1} + \dfrac{1}{\lambda_2} = \dfrac{1}{\lambda_3}\right]$

or $\boxed{\lambda_3 = \dfrac{\lambda_1\lambda_2}{\lambda_1 + \lambda_2}}$ $\left[\because \dfrac{\lambda_1 + \lambda_2}{\lambda_1.\lambda_2} = \dfrac{1}{\lambda_3}\right]$

21. (c) In Y_{89}^{231} number of protons and electrons is 89 and number of neutrons = A – Z = 231 – 89 = 142

22. (a) Energy is always absorbed or emitted in whole number or multiples of quantum.

23. (d) Magnetic moment

$$= 2\sqrt{6} = \sqrt{24} \text{ B.M.} = \sqrt{n(n+2)} \text{ B.M.}$$

Hence, $n = 4$ (unpaired electrons)
Co^{3+} – [Ar] $3d^6$, 4 unpaired electrons.

24. (b) Let wavelength of particle be x

So, velocity $= \dfrac{x}{100}$

$$\lambda = \frac{h}{mv};\quad x = \frac{h \times 100}{m \times x}$$

$$x^2 = 100\frac{h}{m} \quad \text{or} \quad x = 10\sqrt{\frac{h}{m}}$$

25. (b)

26. (c) First four orbitals contain four lobes, while fifth orbital consists of only two lobes. The lobes of d_{xy} orbital lie between x and y axis. Similarly in the case of d_{yz} and d_{zx}. their lobes lie between yz and zx axis respectively. Four lobes of $d_{x^2-y^2}$ orbital are lying along x and y axis while two lobes of d_{z^2} orbital are lying along z-axis.

27. (b) $\sqrt{n(n+2)} = \sqrt{15}$, $n = 3$ (n = number of unpaired electrons)

Therefore, $x = 4$ $\quad \because M^{4+} = [\text{Ar}]3d^3$

28. (a)

29. (a) We know $\Delta p.\Delta x \geq \dfrac{h}{4\pi}$

since $\Delta p = \Delta x$ (given)

$\therefore\ \Delta p.\Delta p = \dfrac{h}{4\pi}$

or $m\Delta v\, m\Delta v. = \dfrac{h}{4\pi}$ $[\therefore \Delta p = m\Delta v]$

or $(\Delta v)^2 = \dfrac{h}{4\pi m^2}$

or $\Delta v = \sqrt{\dfrac{h}{4\pi m^2}} = \dfrac{1}{2m}\sqrt{\dfrac{h}{\pi}}$

Thus option (a) is the correct option.

30. (d) K.E per atom

$$= \frac{\left(4.4\times10^{-19}\right) - \left(4.0\times10^{-19}\right)}{2}$$

$$= \frac{0.4\times10^{-19}}{2} = 2.0\times10^{-20}$$

31. (d) $\lambda = \dfrac{h}{mv} = \dfrac{6.6\times10^{-34}}{0.66\times100} = 1\times10^{-35}$ m

32. (b) Given $E_1 = 25$eV $\quad E_2 = 50$ eV

$E_1 = \dfrac{hc}{\lambda_1}$ $\quad E_2 = \dfrac{hc}{\lambda_2}$ $\qquad \therefore \dfrac{E_1}{E_2} = \dfrac{\lambda_2}{\lambda_1}$

$\therefore \dfrac{\lambda_2}{\lambda_1} = \dfrac{25}{50} = \dfrac{1}{2}$ $\quad \therefore \lambda_1 = 2\lambda_2$

33. (a) $ns \to (n-2)f \to (n-1)d \to np$ $[n=6]$

34. (c) Energy of photon obtained from the transition $n = 6$ to $n = 5$ will have least energy.

$$\Delta E = 13.6Z^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

35. (c)
(a) $n=3,\ \ell=0$ means $3s$-orbital and $n+\ell=3$
(b) $n=3,\ \ell=1$ means $3p$-orbital $n+\ell=4$
(c) $n=3,\ \ell=2$ means $3d$-orbital $n+\ell=5$
(d) $n=4,\ \ell=0$ means $4s$-orbital $n+\ell=4$

Increasing order of energy among these orbitals is
$3s < 3p < 4s < 3d$
$\therefore$ 3d has highest energy.

36. (b) Species having same number of electrons are **isoelectronic**. On calculating the number of electrons in each given species, we get.
CN^- $(6+7+1=14)$; N_2 $(7+7=14)$;
O_2^{2-} $(8+8+2=18)$; C_2^{2-} $(6+6+2=14)$;
O_2^- $(8+8+1=17)$; NO^+ $(7+8-1=14)$
CO $(6+8=14)$; NO $(7+8=15)$
From the above calculation we find that all the species listed in choice (b) have 14 electrons each so it is the correct answer.

37. (d) (ΔE), The energy required to excite an electron in atom of hydrogen from $n = 1$ to n = 2 is ΔE (difference in energy E_2 and E_1)

Values of E_2 and E_1 are,

$$E_2 = \frac{-1.312\times10^6\times(1)^2}{(2)^2} = -3.28\times10^5\ \text{J mol}^{-1}$$

$E_1 = -1.312\times10^6\ \text{J mol}^{-1}$

$\therefore\ \Delta E = E_2 - E_1 = [-3.28\times10^5]-[-1.312\times10^6]\ \text{J mol}^{-1}$

$= (-3.28\times10^5 + 1.312\times10^6)\ \text{J mol}^{-1}$

$= 9.84\times10^5\ \text{J mol}^{-1}$

Thus the correct answer is (d)

38. (b) According to Heisenberg uncertainty principle.

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

Here $\Delta v = \frac{600\times0.005}{100} = 0.03$

So, $\Delta x = \frac{6.6\times10^{-34}}{4\times3.14\times9.1\times10^{-31}\times0.03}$

$= 1.92\times10^{-3}$ meter

39. (d) Energy required to break one mole of Cl – Cl bonds in Cl_2

$$= \frac{242\times10^3}{6.023\times10^{23}} = \frac{hc}{\lambda}$$

$$= \frac{6.626\times10^{-34}\times3\times10^8}{\lambda}$$

$$\therefore \lambda = \frac{6.626\times10^{-34}\times3\times10^8\times6.023\times10^{23}}{242\times10^8}$$

$= 0.4947\times10^{-6}\ \text{m} = 494.7\ \text{nm}$

40. (b) I. $E = \frac{Z^2}{n^2}\times13.6\,\text{eV}$ ...(i)

or $\frac{I_1}{I_2} = \frac{Z_1^2}{n_1^2}\times\frac{n_2^2}{Z_2^2}$ ...(ii)

Given $I_1 = -19.6\times10^{-18}, Z_1 = 2,$ $n_1 = 1, Z_2 = 3$ and $n_2 = 1$

Substituting these values in equation (ii).

$$-\frac{19.6\times10^{-18}}{I_2} = \frac{4}{1}\times\frac{1}{9}$$

or $I_2 = -19.6\times10^{-18}\times\frac{9}{4}$

$= -4.41\times10^{-17}$ J/atom

41. (b) (A) $4p$ (B) $4s$ (C) $3d$ (D) $3p$

According to Bohr Bury's $(n+\ell)$ rule, increasing order of energy will be (D) < (B) < (C) < (A).

Note : If the two orbitals have same value of $(n+\ell)$ then the orbital with lower value of n will be filled first.

42. (c) As per Bohr's postulate,

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

So, $v = \frac{nh}{2\pi mr}$

$$\text{KE} = \frac{1}{2}mv^2$$

So, $\text{KE} = \frac{1}{2}m\left(\frac{nh}{2\pi mr}\right)^2$

Since, $r = \frac{a_o\times n^2}{z}$

So, for 2nd Bohr orbit

$$r = \frac{a_o\times2^2}{1} = 4a_o$$

$$\text{KE} = \frac{1}{2}m\left(\frac{2^2h^2}{4\pi^2m^2\times(4a_o)^2}\right)$$

$$\text{KE} = \frac{h^2}{32\pi^2ma_o^2}$$

43. (b) Average atomic mass of Fe

$$= \frac{(54\times5)+(56\times90)+(57\times5)}{100} = 55.95$$

44. (a) $\Delta E = 2.178\times10^{-18}\left(\frac{1}{1^2}-\frac{1}{2^2}\right) = \frac{hc}{\lambda}$

$$\Rightarrow 2.178\times10^{-18}\times\frac{3}{4} = \frac{hc}{\lambda}$$

$$= \frac{6.62\times10^{-34}\times3\times10^8}{\lambda}$$

$$\lambda = \frac{6.62\times10^{-34}}{2.178\times10^{-18}}\frac{3\times10^8\times4}{\times3}$$

$= 1.214\times10^{-7}$m

45. (c) Energy emitted by the bulb = 600 W = 600 Js^{-1} ($1\text{W} = 1\ \text{Js}^{-1}$)

$\lambda = 331.3\times10^{-9}$ m

Energy of one photon $= h\nu = \frac{hc}{\lambda} = \frac{6.62\times10^{-34}\times3\times10^8}{331.3\times10^{-9}}$

$= 0.059\times10^{-17} \approx 0.06\times10^{-17}$ J

No. of photon emitted from the lamp per second

$$= \frac{600}{0.06\times10^{-17}} = 1.0\times10^{21}$$

46. (b) de Broglie wavelength $\lambda = \frac{\text{h}}{\text{mv}} = \frac{6.6\times10^{-34}}{1\times100}$

$= 6.6\times10^{-36}$m

47. (d) The statement-1 is false but the statement-2 is true exact position and exact momentum of an electron can never be determined according to Heisenberg's uncertainty principle. Even not with the help of electron microscope because when electron beam of electron microscope strikes the target electron of atom, the impact causes the change in velocity and position of electron .
Thus the product of uncertainty in position and momentum is

$$\Delta x.\Delta p \geq \frac{h}{4\pi} \approx 0.57 \text{ ergs sec/gram}$$

48. (a) Both statement-1 and statement-2 are true and Statement-2 is the correct explanation of Statement-1.

$$\text{Radius, } r_n = \frac{n^2h^2}{4\pi e^2 mZ} = \frac{n^2}{Z} \times 0.529\text{Å}.r_n$$

For first orbit of H-atom
$n = 1$

$$r_1 = \frac{(1)^2}{1} \times 0.529\,\text{Å} = 0.529\,\text{Å}$$

49. (a) It is observed that a nucleus which is made up of even number of nucleons (No. of n & p) is more stable than nuclie which consist of odd number of nucleons. If number of neutron or proton is equal to some numbers i.e., 2, 8, 20, 50, 82, or 126 (which are called magic numbers), then these posses extra stability.

50. (b) Both Statement-1 and Statement-2 are correct. Statement-2 is **not** the correct explanation of Statement-1.

EXERCISE - 3

Exemplar Questions

1. (c) The concept of electrons move in a circular path of fixed energy called orbits was given by Bohr and not derived from Rutherford's scattering experiment.
2. (b) The correct configuration for copper ($z = 29$) should be $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$. Due to extra stability of fully filled orbital of d-subshell, the last electron enter into d-orbital instead of s-orbital.
3. (d) The probability density of electrons in $2s$ orbital first increases then decreases and after that it increases again as distance increases from nucleus.
4. (d) The characteristics of cathode rays do not depend upon the material of electrodes and the nature of the gas present in the cathode ray tube.
5. (b) The mass of electron is very small as compared to the mass of the neutron.
Mass of electron = 9.1×10^{-31} kg
Mass of neutron = 1.67×10^{-27} kg
6. (a) J. J. Thomson, in 1898, proposed plum pudding model of atom. An important feature of this model is that the mass of the atom is assumed to be uniformly distributed over the atom. This model was able to explain the overall neutrality of the atom.
7. (d) Isobars have the same mass number (i.e., sum of protons and neutrons) but different atomic number (i.e., number of protons) e.g., $_{26}Fe^{58}$ and $_{27}Ni^{58}$ are isobars.
8. (d) For hydrogen atom ($1s$) = number of radial nodes $= n - l - 1$
Number of radial nodes for $3p$ orbital $= 3 - 1 - 1 = 1$
9. (c) Number of angular nodes = l
$l = 2$ for d-orbital
$\therefore$ Number of angular nodes = 2
10. (b) The important implications of Heisenberg uncertainty principle is that it rules out existence of definite paths or trajectories of electrons and other similar particles.
11. (c) Total number of orbitals associated with n^{th} shell $= n^2$
$\therefore$ Total number of orbitals associated with third shell $= (3)^2 = 9$
12. (a) Orbital angular momentum, $mvr = \frac{h}{2\pi}\sqrt{l(l+1)}$
Hence, it depends only on 'l', l can have values ranging from 0 to $(n - 1)$.
13. (c) The fractional atomic mass (35.5) of chlorine is due to the fact that in ordinary chlorine atom, Cl-37 and Cl-35 are present in the ratio of 1 : 3.
$\therefore$ Average atomic mass of Cl

$$= \frac{3 \times 35 + 1 \times 37}{4} = 35.5 \text{ amu}$$

14. (b)
$_{24}Cr = [Ar]3d^5, 4s^1$ $\quad$ $Cr^{3+} = [Ar]3d^3$
$_{26}Fe = [Ar]3d^6, 4s^2$ $\quad$ $Fe^{3+} = [Ar]3d^5$
$_{25}Mn = [Ar]3d^5, 4s^2$ $\quad$ $Mn^{2+} = [Ar]3d^5$
$_{27}Co = [Ar]3d^7, 4s^2$ $\quad$ $Co^{3+} = [Ar]3d^6$
$_{21}Sc = [Ar]3d^1, 4s^2$ $\quad$ $Sc^{3+} = [Ar]$
Thus, Fe^{3+} and Mn^{2+} have the same electronic configuration.
15. (d) For the two electrons of $2s$ orbital, the value of m_s is between $+\frac{1}{2}$ and $-\frac{1}{2}$

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16. (b) From de-Broglie equation wavelength,

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

For same speed of different particles, $\lambda \propto \frac{1}{m}$

As h is constant, greater the mass of matter waves, lesser is wavelength and vice-versa. Among these matter waves, alpha particle (He^{2+}) has higher mass, therefore, shortest wavelength.

NEET/AIPMT (2013-2017) Questions

17. (b) $E = \dfrac{hc}{\lambda} \times N_A$

$$= \frac{6.62 \times 10^{-27} \times 3 \times 10^{10} \times 6.02 \times 10^{23}}{\lambda}$$

$$= \frac{1.19 \times 10^{8}}{\lambda} \text{ ergs mol}^{-1}$$

18. (c) Energy of an electron at infinite distance from the nucleus is zero. As an electron approaches the nucleus, the electron attraction increases and hence the energy of electron decreases and thus becomes negative. Thus as the value of n decreases, *i.e.* lower the orbit is, more negative is the energy of the electron in it.

19. (c) $n = 3 \rightarrow 3^{rd}$ shell
$l = 1 \rightarrow$ p sub shell.
m = – 1 is possible for two electrons present in an orbital.

20. (b) $c = \nu\lambda$

$$\lambda = \frac{c}{\nu} = \frac{3 \times 10^{17}}{6 \times 10^{15}} = 50 \text{ nm}$$

21. (a) Given: $n = 3, l = 1, m = 0$
Hence orbital is $3p$

–1	0	+1

hence the number of orbital identified by $m = 0$ can be one only.

22. (d) $E = \dfrac{hc}{\lambda} = \dfrac{6.63 \times 10^{-34} \times 3 \times 10^{8}}{45 \times 10^{-9}}$
$= 4.42 \times 10^{-18}$ J

23. (d) Two electrons occupying the same orbital should have opposite spins i.e. they differ in spin quantum number.

24. (c) For hydrogen like atoms energy of $2s$-orbital and $2p$-orbital is equal.