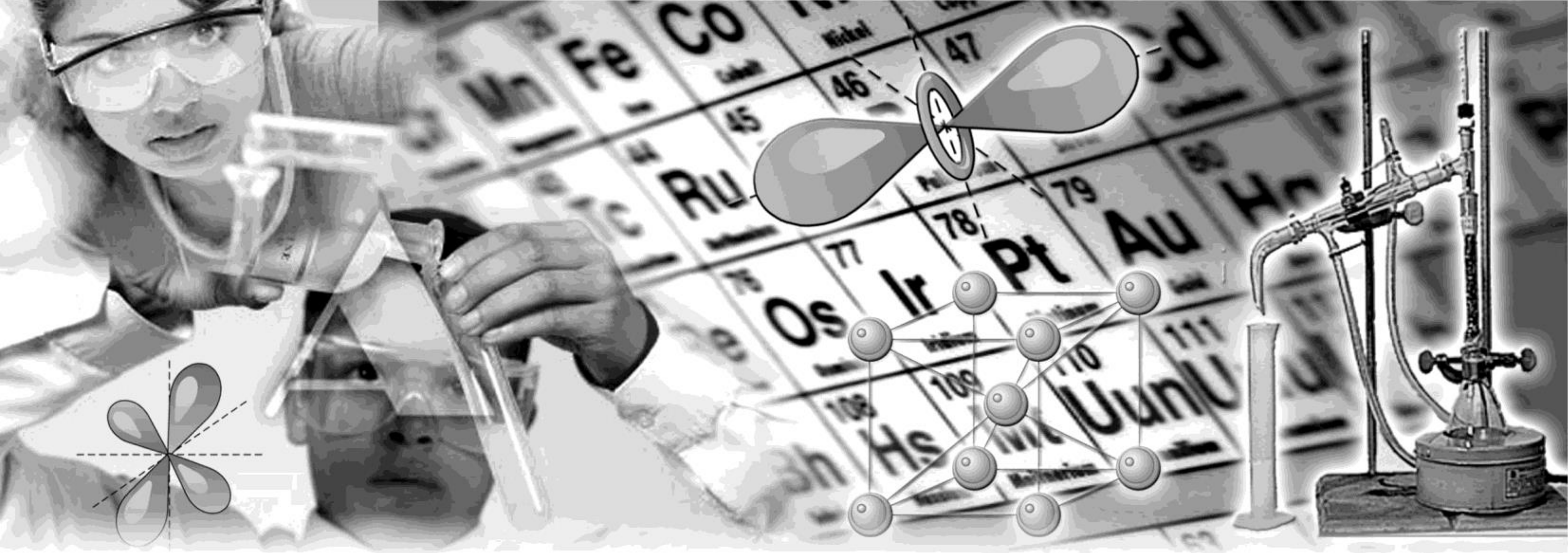


Chapter
2
Structure of atom

John Dalton 1808, believed that matter is made up of extremely minute indivisible particles, called *atom* which can takes part in chemical reactions. These can neither be created nor be destroyed. However, modern researches have conclusively proved that atom is no longer an indivisible particle. Modern structure of atom is based on Rutherford's scattering experiment on atoms and on the concepts of quantization of energy.

Composition of atom

The works of J.J. Thomson and Ernst Rutherford actually laid the foundation of the modern picture of the atom. It is now believed that the atom consists of several *sub-atomic particles* like electron, proton, neutron, positron, neutrino, meson etc. Out of these particles, the electron, proton and the neutron are called *fundamental subatomic particles* and others are *non-fundamental particles.*

Electron ($_{-1}e^{o}$)

(1) It was discovered by ***J.J. Thomson*** (1897) and is *negatively charged particle*. Electron is a component particle of cathode rays.

(2) Cathode rays were discovered by ***William Crooke's*** & ***J.J. Thomson*** (1880) using a cylindrical hard glass tube fitted with two metallic electrodes. The tube has a side tube with a stop cock. This tube was known as ***discharge tube.*** They passed electricity (10,000*V*) through a discharge tube at very low pressure (10^{-2} to $10^{-3}\,mm\ Hg$). Blue rays were emerged from the cathode. These rays were termed as ***Cathode rays***.

(3) **Properties of Cathode rays**

(i) Cathode rays travel in straight line.

(ii) Cathode rays produce mechanical effect, as they can rotate the wheel placed in their path.

(iii) Cathode rays consist of negatively charged particles known as *electron.*

(iv) Cathode rays travel with high speed approaching that of light (ranging between 10^{-9} to 10^{-11} *cm/sec*)

(v) Cathode rays can cause fluorescence.

(vi) Cathode rays heat the object on which they fall due to transfer of kinetic energy to the object.

(vii) When cathode rays fall on solids such as $Cu, X-$rays are produced.

(viii) Cathode rays possess ionizing power *i.e.*, they ionize the gas through which they pass.

(ix) The cathode rays produce scintillation on the photographic plates.

(x) They can penetrate through thin metallic sheets.

(xi) The nature of these rays does not depend upon the nature of gas or the cathode material used in discharge tube.

(xii) The *e/m* (charge to mass ratio) for cathode rays was found to be the same as that for an e^{-} (-1.76×10^{8} coloumb per *gm*). Thus, *the cathode rays are a stream of electrons.*

(xiii) According to *Einstein's theory of relativity,* mass of electron in motion is, $m' = \frac{\text{Rest mass of electron(m)}}{\sqrt{[1-(u/c)^2]}}$

Where u = velocity of electron, c= velocity of light.

When $u=c$ than mass of moving electron =∞.

Proton ($_1H^1$, H^+, P)

(1) Proton was discovered by ***Goldstein*** and is positively charged particle. It is a component particle of anode rays.

(2) **Goldstein (1886)** used perforated cathode in the discharge tube and repeated Thomson's experiment and observed the formation of anode rays. These rays also termed as *positive* or *canal rays*.

(3) **Properties of anode rays**

(i) Anode rays travel in straight line.

(ii) Anode rays are material particles.

(iii) Anode rays are positively charged.

Table : 2.1 Comparison of mass, charge and specific charge of electron, proton and neutron

Name of constant	*Unit*	**Electron(e^-)**	**Proton(p^+)**	**Neutron(n)**
Mass (m)	Amu	0.000546	1.00728	1.00899
	Kg	9.109×10^{-31}	1.673×10^{-27}	1.675×10^{-27}
	Relative	1/1837	1	1
Charge(e)	Coulomb (C)	-1.602×10^{-19}	$+1.602 \times 10^{-19}$	Zero
	Esu	-4.8×10^{-10}	$+4.8 \times 10^{-10}$	Zero
	Relative	– 1	+1	Zero
Specific charge (e/m)	C/g	1.76×10^8	9.58×10^4	Zero
Density	Gram / cc	2.17×10^{-17}	1.114×10^{14}	1.5×10^{-14}

- The atomic mass unit (amu) is 1/12 of the mass of an individual atom of $_6C^{12}$, *i.e.* 1.660×10^{-27} kg.

Table : 2.2 Other non fundamental particles

Particle	**Symbol**	**Nature**	**Charge esu $\times 10^{-10}$**	**Mass (amu)**	**Discovered by**
Positron	$e^+, 1e^0, \beta^+$	+	+ 4.8029	0.0005486	Anderson (1932)
Neutrino	ν	0	0	< 0.00002	Pauli (1933) and Fermi (1934)
Anti-proton	p^-	–	– 4.8029	1.00787	Chamberlain Sugri (1956) and Weighland (1955)
Positive mu meson	μ^+	+	+ 4.8029	0.1152	Yukawa (1935)
Negative mu meson	μ^-	–	– 4.8029	0.1152	Anderson (1937)
Positive pi meson	π^+	+	+ 4.8029	0.1514	Powell (1947)
Negative pi meson	π^-	–	– 4.8029	0.1514	
Neutral pi meson	π^0	0	0	0.1454	

(iv) Anode rays may get deflected by external magnetic field.

(v) Anode rays also affect the photographic plate.

(vi) The e/m ratio of these rays is smaller than that of electrons.

(vii) Unlike cathode rays, their e/m value is dependent upon the nature of the gas taken in the tube. It is maximum when gas present in the tube is hydrogen.

(viii) These rays produce flashes of light on *ZnS* screen.

Neutron ($_0n^1$, N)

(1) Neutron was discovered by ***James Chadwick*** (1932) according to the following nuclear reaction,

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

or $$_5B^{11} + _2He^4 \rightarrow _7N^{14} + _0n^1$$

(2) Neutron is an unstable particle. It decays as follows,

$$ {}_0n^1 \longrightarrow \underset{\text{Proton}}{{}_1H^1} + \underset{\text{electon}}{{}_{-1}e^0} + \underset{\text{antinutrin o}}{{}_0\nu^0} $$
neutron

Atomic number, Mass number and Atomic species

(1) **Atomic number or Nuclear charge**

(i) The number of protons present in the nucleus of the atom is called *atomic number* (Z).

(ii) It was determined by **Moseley** as,

$\sqrt{\nu} = a(Z - b)$ or $aZ - ab$

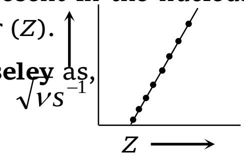


Fig. 2.1

Where, $\nu = X -$ ray's frequency

Z= atomic number of the metal $a \& b$ are constant.

(iii) Atomic number = Number of positive charge on nucleus = Number of protons in nucleus = Number of electrons in nutral atom.

(iv) *Two different elements can never have identical atomic number*.

(2) **Mass number**

Mass number (A) = Number of protons or Atomic number (Z) + Number of neutrons or Number of neutrons = $A - Z$.

(i) Since mass of a proton or a neutron is not a whole number (on atomic weight scale), weight is not necessarily a whole number.

(ii) The atom of an element X having mass number (A) and atomic number (Z) may be represented by a symbol, ${}_ZX^A$.

Table: 2.3 Different types of atomic species

Atomic species	Similarities	Differences	Examples
Isotopes (Soddy)	(i) Atomic No. (Z) (ii) No. of protons (iii) No. of electrons (iv) Electronic configuration (v) Chemical properties (vi) Position in the periodic table	(i) Mass No. (A) (ii) No. of neutrons (iii) Physical properties	(i) ${}^1_1H, {}^2_1H, {}^3_1H$ (ii) ${}^{16}_8O, {}^{17}_8O, {}^{18}_8O$ (iii) ${}^{35}_{17}Cl, {}^{37}_{17}Cl$
Isobars	(i) Mass No. (A) (ii) No. of nucleons	(i) Atomic No. (Z) (ii) No. of protons, electrons and neutrons (iii)Electronic configuration (iv) Chemical properties (v) Position in the perodic table.	(i) ${}^{40}_{18}Ar, {}^{40}_{19}K, {}^{40}_{20}Ca$ (ii) ${}^{130}_{52}Te, {}^{130}_{54}Xe, {}^{130}_{56}Ba$
Isotones	No. of neutrons	(i) Atomic No. (ii) Mass No., protons and electrons. (iii) Electronic configuration (iv) Physical and chemical properties (v) Position in the periodic table.	(i) ${}^{30}_{14}Si, {}^{31}_{15}P, {}^{32}_{16}S$ (ii) ${}^{39}_{19}K, {}^{40}_{20}Ca$ (iii) ${}^3_1H, {}^4_2He$ (iv) ${}^{13}_6C, {}^{14}_7N$
Isodiaphers	Isotopic No. ($N - Z$) or ($A - 2Z$)	(i) At No., mass No., electrons, protons, neutrons. (ii) Physical and chemical properties.	(i) ${}_{92}U^{235}, {}_{90}Th^{231}$ (ii) ${}_{19}K^{39}, {}_9F^{19}$ (iii) ${}_{29}Cu^{65}, {}_{24}Cr^{55}$

Isoelectronic species	(i) No. of electrons (ii) Electronic configuration	At. No., mass No.	(i) $N_2O, CO_2, CNO^-(22e^-)$ (ii) $CO, CN^-, N_2(14e^-)$ (iii) $H^-, He, Li^+, Be^{2+}(2e^-)$ (iv) $P^{3-}, S^{2-}, Cl^-, Ar, K^+ and\ Ca^{2+}(18e^-)$
Isosters	(i) No. of atoms (ii) No. of electrons (iii) Physical and chemical properties.		(i) N_2 and CO (ii) CO_2 and N_2O (iii) HCl and F_2 (iv) CaO and MgS (v) C_6H_6 and $B_3N_3H_6$

Electromagnetic radiations

(1) Light and other forms of radiant energy propagate without any medium in the space in the form of waves are known as *electromagnetic radiations*. These waves can be produced by a charged body moving in a magnetic field or a magnet in a electric field. e.g. $\alpha -$ rays, $\gamma -$ rays, cosmic rays, ordinary light rays etc.

(2) **Characteristics**

(i) All electromagnetic radiations travel with the velocity of light.

(ii) These consist of electric and magnetic fields components that oscillate in directions perpendicular to each other and perpendicular to the direction in which the wave is travelling.

(3) A wave is always characterized by the following five characteristics,

(i) ***Wavelength*** **:** The distance between two nearest crests or nearest troughs is called the wavelength. It is denoted by λ (lambda) and is measured is terms of centimeter(cm), angstrom(Å), micron(μ) or nanometre (nm).

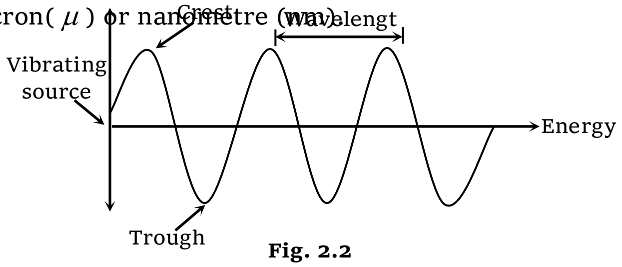


Fig. 2.2

$1Å = 10^{-8}\ cm = 10^{-10}\ m$; $1\mu = 10^{-4}\ cm = 10^{-6}\ m$;

$1nm = 10^{-7}\ cm = 10^{-9}\ m$; $1cm = 10^8\ Å = 10^4\ \mu = 10^7\ nm$

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

$\lambda\nu =$ distance travelled in one second = velocity =c

$$\nu = \frac{c}{\lambda}$$

(iii) ***Velocity*** **:** It is defined as the distance covered in one second by the wave. It is denoted by the letter 'c'. All electromagnetic waves travel with the same velocity, i.e., $3\times 10^{10} cm/\sec$.

$$c = \lambda\nu = 3\times 10^{10}\ cm/\sec$$

(iv) ***Wave number*** **:** This is the reciprocal of wavelength, i.e., the number of wavelengths per centimetre. It is denoted by the symbol $\bar{\nu}$ (nu bar). It is expressed in cm^{-1} or m^{-1}.

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

(v) ***Amplitude*** **:** It is defined as the height of the crest or depth of the trough of a wave. It is denoted by the letter 'A'. It determines the intensity of the radiation.

The arrangement of various types of electromagnetic radiations in the order of their increasing or decreasing wavelengths or frequencies is known as ***electromagnetic spectrum***.

Table: 2.4

Name	Wavelength (Å)	Frequency (Hz)
Radio wave	$3\times 10^{14} - 3\times 10^7$	$1\times 10^5 - 1\times 10^9$
Microwave	$3\times 10^7 - 6\times 10^6$	$1\times 10^9 - 5\times 10^{11}$
Infrared (IR)	$6\times 10^6 - 7600$	$5\times 10^{11} - 3.95\times 10^{16}$
Visible	$7600 - 3800$	$3.95\times 10^{16} - 7.9\times 10^{14}$
Ultraviolet (UV)	$3800 - 150$	$7.9\times 10^{14} - 2\times 10^{16}$
X-Rays	$150 - 0.1$	$2\times 10^{16} - 3\times 10^{19}$
$\gamma -$ Rays	$0.1 - 0.01$	$3\times 10^{19} - 3\times 10^{20}$
Cosmic Rays	0.01- zero	3×10^{20} – infinity

Atomic spectrum - Hydrogen spectrum

Atomic spectrum

Spectrum is the impression produced on a photographic film when the radiation (s) of particular wavelength (s) is (are) analysed through a prism or diffraction grating.

Types of spectrum

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

(i) ***Continuous spectrum* :** When sunlight is passed through a prism, it gets dispersed into continuous bands of different colours. If the light of an incandescent object resolved through prism or spectroscope, it also gives continuous spectrum of colours.

(ii) ***Line spectrum* :** If the radiation's obtained by the excitation of a substance are analysed with help of a spectroscope a series of thin bright lines of specific colours are obtained. There is dark space in between two consecutive lines. This type of spectrum is called line spectrum or atomic spectrum..

(2) **Absorption spectrum :** Spectrum produced by the absorbed radiations is called absorption spectrum.

Hydrogen spectrum

(1) Hydrogen spectrum is an example of line emission spectrum or atomic emission spectrum.

(2) When an electric discharge is passed through hydrogen gas at low pressure, a bluish light is emitted.

(3) This light shows discontinuous line spectrum of several isolated sharp lines through prism.

(4) All these lines of H-spectrum have Lyman, Balmer, Paschen, Barckett, Pfund and Humphrey series. These spectral series were named by the name of scientist discovered them.

(5) To evaluate wavelength of various H-lines Ritz introduced the following expression,

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

Where R is universal constant known as Rydberg's constant its value is 109, 678 cm^{-1}.

Plum pudding model of Thomson

(1) He suggected that atom is a positively charged sphere having electrons embedded uniformly giving an overall picture of plum pudding.

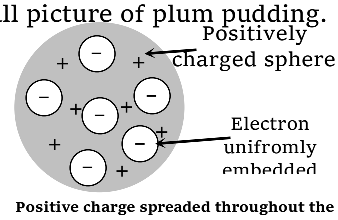


Fig. 2.3

(2) This model failed to explain the line spectrum of an element and the scattering experiment of Rutherford.

Rutherford's nuclear model

(1) Rutherford carried out experiment on the bombardment of thin (10^{-4} mm) Au foil with high speed positively charged $\alpha-$particles emitted from Ra and gave the following observations based on this experiment,

(i) Most of the $\alpha-$ particles passed without any deflection.

(ii) Some of them were deflected away from their path.

(iii) Only a few (one in about 10,000) were returned back to their original direction of propagation.

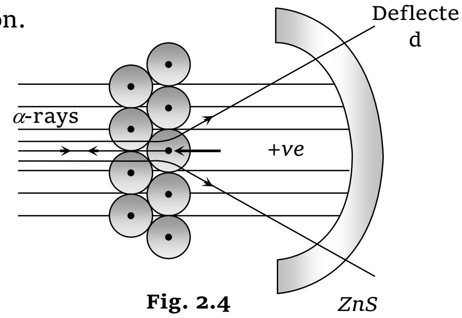


Fig. 2.4

(2) From the above observations he concluded that, an atom consists of

(i) ***Nucleus*** which is small in size but carries the entire mass i.e. contains all the neutrons and protons.

(ii) ***Extra nuclear part*** which contains electrons. This model was similar to the solar system.

(3) **Properties of the nucleus**

(i) Nucleus is a small, heavy, positively charged portion of the atom and located at the centre of the atom.

(ii) All the positive charge of atom (i.e. protons) are present in nucleus.

(iii) Nucleus contains neutrons and protons, and hence these particles collectively are also referred to as ***nucleons.***

(iv) The ***size*** of nucleus is measured in ***Fermi*** (1 Fermi = 10^{-13} cm).

(v) The ***radius*** of nucleus is of the order of $1.5\times10^{-13}\,cm.$ to $6.5\times10^{-13}\,cm.$ i.e. 1.5 to 6.5 Fermi.

Generally the radius of the nucleus (r_n) is given by the following relation,

$$r_n = r_o (= 1.4 \times 10^{-13} cm) \times A^{1/3}$$

This exhibited that nucleus is 10^{-5} times small in size as compared to the total size of atom.

(vi) The ***Volume*** of the nucleus is about $10^{-39}\ cm^3$ and that of atom is $10^{-24} cm^3$, i.e., volume of the nucleus is 10^{-15} times that of an atom.

(vii) The **density** of the nucleus is of the order of $10^{15} g\,cm^{-3}$ or 10^8 tonnes cm^{-3} or $10^{12} kg/cc$. If nucleus is spherical than,

Density

$$= \frac{\text{mass of the nucleus}}{\text{volume of the nucleus}} = \frac{\text{mass number}}{6.023 \times 10^{23} \times \frac{4}{3}\pi r^3}$$

(4) **Drawbacks of Rutherford's model**

(i) It does not obey the ***Maxwell theory of electrodynamics,*** according to it “A small charged particle moving around an oppositely charged centre continuously loses its energy”. If an electron does so, it should also continuously lose its energy and should set up spiral motion ultimately failing into the nucleus.

(ii) It could not explain the line spectra of $H-$ atom and discontinuous spectrum nature.

Planck's quantum theory

When black body is heated, it emits thermal radiation’s of different wavelengths or frequency. To explain these radiations, max planck put forward a theory known as planck’s quantum theory.

(i) The radiant energy which is emitted or absorbed by the black body is not continuous but discontinuous in the form of small discrete packets of energy, each such packet of energy is called a '**quantum**'. In case of light, the quantum of energy is called a '**photon**'.

(ii) The energy of each quantum is directly proportional to the frequency (ν) of the radiation, i.e.

$$E \propto \nu \text{ or } E = h\nu = \frac{hc}{\lambda}$$

Where, $h =$ Planck's constant = 6.62×10^{-27} *erg. sec.* or 6.62×10^{-34} *Joules* sec.

(iii)The total amount of energy emitted or absorbed by a body will be some whole number quanta. Hence $E = nh\nu$, where n is an integer.

Photoelectric effect

(1) When radiations with certain minimum frequency (ν_0) strike the surface of a metal, the electrons are ejected from the surface of the metal. This phenomenon is called *photoelectric effect* and the electrons emitted are called *photo-electrons*. The current constituted by photoelectrons is known as photoelectric current.

(2) The electrons are ejected only if the radiation striking the surface of the metal has at least a minimum frequency (ν_0) called *Threshold frequency*. The minimum potential at which the plate photoelectric current becomes zero is called *stopping potential*.

(3) The velocity or kinetic energy of the electron ejected depend upon the frequency of the incident radiation and is independent of its intensity.

(4) The number of photoelectrons ejected is proportional to the intensity of incident radiation.

(5) **Einstein’s photoelectric effect equation**

According to Einstein,

Maximum kinetic energy of the ejected electron = absorbed energy – threshold energy

$$\frac{1}{2}mv^2_{\max} = h\nu - h\nu_0 = hc\left[\frac{1}{\lambda} - \frac{1}{\lambda_0}\right]$$

Where, ν_0 and λ_0 are threshold frequency and threshold wavelength.

Bohr’s atomic model

Bohr retained the essential features of the Rutherford model of the atom. However, in order to account for the stability of the atom he introduced the concept of the stationary orbits. The Bohr postulates are,

(1) An atom consists of positively charged nucleus responsible for almost the entire mass of the atom (This assumption is retention of Rutherford model).

(2) The electrons revolve around the nucleus in certain permitted circular orbits of definite radii.

(3) The permitted orbits are those for which the angular momentum of an electron is an intergral multiple of $h/2\pi$ where h is the Planck’s constant. If m is the mass and v is the velocity of the electron in a permitted orbit of radius r, then

$$L = mvr = \frac{nh}{2\Pi}\ ;\ n = 1, 2, 3, ...\infty$$

Where L is the orbital angular momentum and n is the number of orbit. The integer n is called the principal quantum number. This equation is known as the Bohr quantization postulate.

(4) When electrons move in permitted discrete orbits they do not radiate or lose energy. Such orbits are called stationary or non-radiating orbits. In this manner, Bohr overcame Rutherford’s difficulty to account for the stability of the atom. Greater the distance of energy level from the nucleus, the more is

the energy associated with it. The different energy levels were numbered as 1,2,3,4 .. and called as $K, L, M, N,$ .... etc.

(5) Ordinarily an electron continues to move in a particular stationary state or orbit. Such a state of atom is called ground state. When energy is given to the electron it jumps to any higher energy level and is said to be in the excited state. When the electron jumps from higher to lower energy state, the energy is radiated.

Advantages of Bohr's theory

(i) Bohr's theory satisfactorily explains the spectra of species having one electron, *viz*. hydrogen atom, He^{+}, Li^{2+} etc.

(ii) ***Calculation of radius of Bohr's orbit*** : According to Bohr, radius of n^{th} orbit in which electron moves is

$$r_n = \left[\frac{h^2}{4\pi^2 me^2 k}\right] \cdot \frac{n^2}{Z}$$

Where, $n =$ Orbit number, $m =$ Mass number $[9.1 \times 10^{-31} kg]$, $e =$ Charge on the electron $[1.6 \times 10^{-19}]$ $Z =$ Atomic number of element, k = Coulombic constant $[9 \times 10^9 Nm^2 c^{-2}]$

After putting the values of *m*,*e*,*k*,*h*, we get.

$$r_n = \frac{n^2}{Z} \times 0.529\, Å$$

(iii) ***Calculation of velocity of electron***

$$V_n = \frac{2\pi e^2 ZK}{nh}, V_n = \left[\frac{Ze^2}{mr}\right]^{1/2};$$

$$V_n = \frac{2.188 \times 10^8 Z}{n} cm. \sec^{-1}$$

(iv) ***Calculation of energy of electron in Bohr's orbit***

Total energy of electron = K.E. + P.E. of electron $= \frac{kZe^2}{2r} - \frac{kZe^2}{r} = -\frac{kZe^2}{2r}$

Substituting of *r*, gives us $E = \frac{-2\pi^2 mZ^2 e^4 k^2}{n^2 h^2}$

Where, n=1, 2, 3.......... ∞

Putting the value of m, e, k, h, π we get

$$E = 21.8 \times 10^{-12} \times \frac{Z^2}{n^2} erg\ per\ atom$$

$$= -21.8 \times 10^{-19} \times \frac{Z^2}{n^2} J\ per\ atom\ (1J = 10^7 erg)$$

$$E = -13.6 \times \frac{Z^2}{n^2} eV\ per\ atom (1\text{eV} = 1.6 \times 10^{-19} J)$$

$$= -13.6 \times \frac{Z^2}{n^2} k.cal/mole \quad (1\ cal = 4.18J)$$

or $\frac{-1312}{n^2} Z^2 kJmol^{-1}$

When an electron jumps from an outer orbit (higher energy) n_2 to an inner orbit (lower energy) n_1, then the energy emitted in form of radiation is given by

$$\Delta E = E_{n_2} - E_{n_1} = \frac{2\pi^2 k^2 me^4 Z^2}{h^2}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

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

As we know that $E = h\bar{\nu}$, $c = \nu\lambda$ and $\bar{\nu} = \frac{1}{\lambda} = \frac{\Delta E}{hc}$,

$$= \frac{2\pi^2 k^2 me^4 Z^2}{ch^3}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

This can be represented as $\frac{1}{\lambda} = \bar{\nu} = RZ^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$

Where, $R = \frac{2\pi^2 k^2 me^4}{ch^3}$; R is known as Rydberg constant. Its value to be used is $109678\ cm^{-1}$.

The negative sign in the above equations shows that the electron and nucleus form a bound system, *i.e.,* the electron is attracted towards the nucleus. Thus, if electron is to be taken away from the nucleus, energy has to be supplied. The energy of the electron in $n = 1$ orbit is called the ground state energy; that in the $n = 2$ orbit is called the first excited state energy, etc. When $n = \infty$ then $E = 0$ which corresponds to ionized atom *i.e.,* the electron and nucleus are infinitely separated $H \rightarrow H^{+} + e^{-}$ (ionization).

(6) **Spectral evidence for quantisation** (Explanation for hydrogen spectrum on the basisof bohr atomic model)

(i) The light absorbed or emitted as a result of an electron changing orbits produces characteristic absorption or emission spectra which can be recorded on the photographic plates as a series of lines, the optical spectrum of hydrogen consists of several series of lines called **Lyman, Balmar, Paschen, Brackett, Pfund** and **Humphrey.** These spectral series were named by the name of scientist who discovered them.

(ii) To evaluate wavelength of various H-lines Ritz introduced the following expression,

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

Where, R is = $\frac{2\pi^2 me^4}{ch^3}$ = Rydberg's constant

It's theoritical value = 109,737 cm^{-1} and It's experimental value = $109{,}677.581\, cm^{-1}$

This remarkable agreement between the theoretical and experimental value was great achievment of the Bohr model.

(iii) Although H-atom consists of only one electron yet it's spectra consist of many spectral lines.

(iv) *Comparative study of important spectral series of Hydrogen is shown in following table.*

(v) If an electron from n^{th} excited state comes to various energy states, the maximum spectral lines obtained will be $= \frac{n(n-1)}{2}$. n= principal quantum number.

As n=6 than total number of spectral lines $= \frac{6(6-1)}{2} = \frac{30}{2} = 15$.

(vi) Thus, at least for the hydrogen atom, the Bohr theory accurately describes the origin of atomic spectral lines.

(7) **Failure of Bohr model**

(i) Bohr theory was very successful in predicting and accounting the energies of line spectra of hydrogen i.e. one electron system. It could not explain the line spectra of atoms containing more than one electron.

(ii) This theory could not explain the presence of multiple spectral lines.

(iii) This theory could not explain the splitting of spectral lines in magnetic field (Zeeman effect) and in electric field (Stark effect). The intensity of these spectral lines was also not explained by the Bohr atomic model.

(iv) This theory was unable to explain of dual nature of matter as explained on the basis of De broglies concept.

(v) This theory could not explain uncertainty principle. (vi) No conclusion was given for the concept of quantisation of energy.

Table: 2.5

S.No.	Spectral series	Lies in the region	Transition $n_2 > n_1$	$\lambda_{max} = \frac{n_1^2 n_2^2}{(n_2^2 - n_1^2)R}$	$\lambda_{min} = \frac{n_1^2}{R}$	$\frac{\lambda_{max}}{\lambda_{min}} = \frac{n_2^2}{n_2^2 - n_1^2}$
(1)	Lymen series	Ultraviolet region	$n_1 = 1$ $n_2 = 2,3,4....\infty$	$n_1 = 1$ and $n_2 = 2$ $\lambda_{max} = \frac{4}{3R}$	$n_1 = 1$ and $n_2 = \infty$ $\lambda_{min} = \frac{1}{R}$	$\frac{4}{3}$
(2)	Balmer series	Visible region	$n_1 = 2$ $n_2 = 3,4,5....\infty$	$n_1 = 2$ and $n_2 = 3$ $\lambda_{max} = \frac{36}{5R}$	$n_1 = 2$ and $n_2 = \infty$ $\lambda_{min} = \frac{4}{R}$	$\frac{9}{5}$
(3)	Paschen series	Infra red region	$n_1 = 3$ $n_2 = 4,5,6....\infty$	$n_1 = 3$ and $n_2 = 4$ $\lambda_{max} = \frac{144}{7R}$	$n_1 = 3$ and $n_2 = \infty$ $\lambda_{min} = \frac{9}{R}$	$\frac{16}{7}$
(4)	Brackett series	Infra red region	$n_1 = 4$ $n_2 = 5,6,7....\infty$	$n_1 = 4$ and $n_2 = 5$ $\lambda_{max} = \frac{16 \times 25}{9R}$	$n_1 = 4$ and $n_2 = \infty$ $\lambda_{min} = \frac{16}{R}$	$\frac{25}{9}$
(5)	Pfund series	Infra red region	$n_1 = 5$ $n_2 = 6,7,8....\infty$	$n_1 = 5$ and $n_2 = 6$ $\lambda_{max} = \frac{25 \times 36}{11R}$	$n_1 = 5$ and $n_2 = \infty$ $\lambda_{min} = \frac{25}{R}$	$\frac{36}{11}$
(6)	Humphrey series	Far infrared region	$n_1 = 6$ $n_2 = 7,8....\infty$	$n_1 = 6$ and $n_2 = 7$ $\lambda_{max} = \frac{36 \times 49}{13R}$	$n_1 = 6$ and $n_2 = \infty$ $\lambda_{min} = \frac{36}{R}$	$\frac{49}{13}$

Bohr–Sommerfeild's model

It is an extension of Bohr's model. The electrons in an atom revolve around the nuclei in elliptical orbit. The circular path is a special case of ellipse. Association of elliptical orbits with circular orbit explains the fine line spectrum of atoms.

Dual nature of electron

(1) In 1924, the French physicist, ***Louis de Broglie*** suggested that if light has both particle and wave like nature, the similar duality must be true for matter. Thus an electron, behaves both as a material particle and as a wave.

(2) This presented a new wave mechanical theory of matter. According to this theory, small particles like electrons when in motion possess wave properties.

(3) According to de-broglie, the wavelength associated with a particle of mass *m*, moving with velocity *v* is given by the relation

$\lambda = \frac{h}{mv}$, where *h* = Planck's constant.

(4) This can be derived as follows according to Planck's equation, $E = h\nu = \frac{h.c}{\lambda}$ $\left(\because \nu = \frac{c}{\lambda}\right)$

energy of photon (on the basis of Einstein's mass energy relationship), $E = mc^2$

Equating both $\frac{hc}{\lambda} = mc^2$ or $\lambda = \frac{h}{mc}$ which is same as de-Broglie relation. $(\because mc = p)$

(5) This was experimentally verified by ***Davisson and Germer*** by observing diffraction effects with an electron beam. Let the electron is accelerated with a potential of V than the Kinetic energy is

$\frac{1}{2}mv^2 = eV$; $m^2v^2 = 2eVm$

$mv = \sqrt{2eVm} = P$; $\lambda = \frac{h}{\sqrt{2eVm}}$

(6) If Bohr's theory is associated with de-Broglie's equation then wave length of an electron can be determined in bohr's orbit and relate it with circumference and multiply with a whole number

$2\pi r = n\lambda$ or $\lambda = \frac{2\pi r}{n}$

From de-Broglie equation, $\lambda = \frac{h}{mv}$.

Thus $\frac{h}{mv} = \frac{2\pi r}{n}$ or $mvr = \frac{nh}{2\pi}$

(7) The de-Broglie equation is applicable to all material objects but it has significance only in case of microscopic particles. Since, we come across macroscopic objects in our everyday life, de-broglie relationship has no significance in everyday life.

Heisenberg's uncertainty principle

This principle states "It is impossible to specify at any given moment both the position and momentum (velocity) of an electron".

Mathematically it is represented as , $\Delta x . \Delta p \geq \frac{h}{4\pi}$

Where $\Delta x =$ uncertainty is position of the particle, $\Delta p =$ uncertainty in the momentum of the particle

Now since $\Delta p = m\,\Delta v$

So equation becomes, $\Delta x . m \Delta v \geq \frac{h}{4\pi}$ or $\Delta x \times \Delta v \geq \frac{h}{4\pi m}$

In terms of uncertainty in energy, ΔE and uncertainty in time Δt, this principle is written as, $\Delta E . \Delta t \geq \frac{h}{4\pi}$

Schrödinger wave equation

(1) Schrodinger wave equation is given by ***Erwin Schrödinger*** in 1926 and based on dual nature of electron.

(2) In it electron is described as a three dimensional wave in the electric field of a positively charged nucleus.

(3) The probability of finding an electron at any point around the nucleus can be determined by the help of Schrodinger wave equation which is,

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

Where x, y and z are the 3 space co-ordinates, m = mass of electron, h = Planck's constant, E = Total energy, V = potential energy of electron, Ψ = amplitude of wave also called as wave function, ∂ = for an infinitesimal change.

(4) The Schrodinger wave equation can also be written as,

$$\nabla^2 \Psi + \frac{8\pi^2 m}{h^2}(E - V)\,\Psi = 0$$

Where ∇ = laplacian operator.

(5) **Physical significance of Ψ and Ψ^2**

(i) The wave function Ψ represents the amplitude of the electron wave. The amplitude Ψ is thus a function of space co-ordinates and time i.e. $\Psi = \Psi(x, y, z...... times)$

(ii) For a single particle, the square of the wave function (Ψ^2) at any point is proportional to the probability of finding the particle at that point.

(iii) If Ψ^2 is maximum than probability of finding e^- is maximum around nucleus and the place where probability of finding e^- is maximum is called *electron density*, electron cloud or an atomic orbital. It is different from the Bohr's orbit.

(iv) The solution of this equation provides a set of number called *quantum numbers* which describe specific or definite energy state of the electron in atom and information about the shapes and orientations of the most probable distribution of electrons around the nucleus.

Radial probability distribution curves : Radial probability is $R = 4\pi r^2 dr \psi^2$. The plats of R distance from nucleus as follows

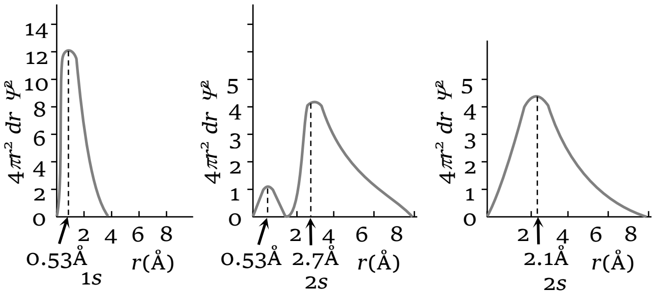


Fig. 2.5

Quantum numbers

Each orbital in an atom is specified by a set of three quantum numbers (*n*, *l*, *m*) and each electron is designated by a set of four quantum numbers (*n*, *l*, *m* and *s*).

(1) **Principle quantum number (*n*)**

(i) It was proposed by *Bohr* and denoted by '*n*'.

(ii) It determines the average distance between electron and nucleus, means it denotes the size of atom.

(iii) It determine the energy of the electron in an orbit where electron is present.

(iv) The maximum number of an electron in an orbit represented by this quantum number as $2n^2$. No energy shell in atoms of known elements possess more than 32 electrons.

(v) It gives the information of orbit *K, L, M, N*------------.

(vi) Angular momentum can also be calculated using principle quantum number

(2) **Azimuthal quantum number (*l*)**

(i) Azimuthal quantum number is also known as angular quantum number. Proposed by *Sommerfield* and denoted by '*l*'.

(ii) It determines the number of sub shells or sublevels to which the electron belongs.

(iii) It tells about the shape of subshells.

(iv) It also expresses the energies of subshells $s < p < d < f$ (increasing energy).

(v) The value of $l = (n-1)$ always. Where '*n*' is the number of principle shell.

(vi) Value of *l*	=	0	1	2	3.....(n-1)
Name of subshell	=	s	p	d	f
Shape of subshell	=	Spherical	Dumbbell	Double dumbbell	Complex

(vii) It represent the orbital angular momentum. Which is equal to $\frac{h}{2\pi}\sqrt{l(l+1)}$

(viii) The maximum number of electrons in subshell $= 2(2l+1)$

$s-$subshell $\rightarrow$ 2 electrons $d-$subshell $\rightarrow$ 10 electrons

$p-$subshell $\rightarrow$ 6 electrons $f-$subshell $\rightarrow$ 14 electrons.

(ix) For a given value of '*n*' the total values of '*l*' is always equal to the value of '*n*'.

(3) **Magnetic quantum number (*m*)**

(i) It was proposed by *Zeeman* and denoted by '*m*'.

(ii) It gives the number of permitted orientation of subshells.

(iii) The value of *m* varies from –*l* to +*l* through zero.

(iv) It tells about the splitting of spectral lines in the magnetic field i.e. this quantum number proves the Zeeman effect.

(v) For a given value of '*n*' the total value of '*m*' is equal to n^2.

(vi) For a given value of 'l' the total value of 'm' is equal to $(2l+1)$.

(vii) *Degenerate orbitals* : Orbitals having the same energy are known as degenerate orbitals. e.g. for p subshell p_x p_y p_z

(viii) The number of degenerate orbitals of s subshell =0.

(4) **Spin quantum numbers (*s*)**

(i) It was proposed by *Goldshmidt & Ulen Back* and denoted by the symbol of 's'.

(ii) The value of $'s'$ is $+1/2$ and $-1/2$, which signifies the spin or rotation or direction of electron on it's axis during movement.

(iii) The spin may be clockwise or anticlockwise.

(iv) It represents the value of spin angular momentum is equal to $\frac{h}{2\pi}\sqrt{s(s+1)}$.

(v) Maximum spin of an atom $=1/2\times$number of unpaired electron.

(vi) This quantum number is not the result of solution of schrodinger equation as solved for H-atom.

Table : 2.6 Distribution of electrons among the quantum levels

n	*l*	*m*	Designation of orbitals	Number of Orbitals in the subshell
1	0	0	1s	1
2	0	0	2s	1
2	1	–1, 0, +1	2p	3
3	0	0	3s	1
3	1	–1, 0, +1	3p	3
3	2	–2, –1, 0, +1, +2	3d	5
4	0	0	4s	1
4	1	–1, 0, +1	4p	3
4	2	–2, –1, 0, +1, +2	4d	5
4	3	–3, –2, –1, 0, +1, +2, +3	4f	7

Shape of orbitals

(1) **Shape of 's' orbital**

(i) For 's' orbital l=0 & m=0 so 's' orbital have only one unidirectional orientation i.e. the probability of finding the electrons is same in all directions.

(ii) The size and energy of 's' orbital with increasing 'n' will be $1s < 2s < 3s < 4s$.

(iii) s-orbitals known as radial node or modal surface. But there is no radial node for 1s orbital since it is starting from the nucleus.

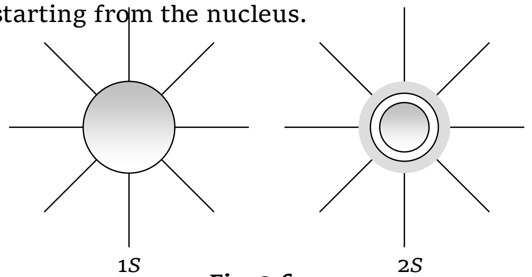


Fig. 2.6

(2) **Shape of '*p*' orbitals**

(i) For 'p' orbital l=1, & m=+1,0,–1 means there are three 'p' orbitals, which is symbolised as p_x, p_y, p_z.

(ii) Shape of 'p' orbital is dumb bell in which the two lobes on opposite side separated by the nodal plane.

(iii) p-orbital has directional properties.

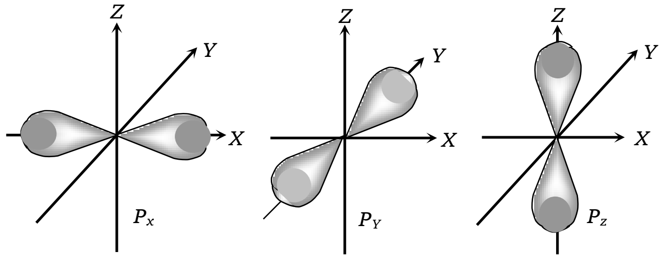


Fig. 2.7

(3) **Shape of '*d*' orbital**

(i) For the 'd' orbital l =2 then the values of 'm' are –2, –1, 0, +1, +2. It shows that the 'd' orbitals has five orbitals as $d_{xy}, d_{yz}, d_{zx}, d_{x^2-y^2}, d_{z^2}$.

(ii) Each '*d*' orbital identical in shape, size and energy.

(iii) The shape of *d* orbital is double dumb bell .

(iv) It has directional properties.

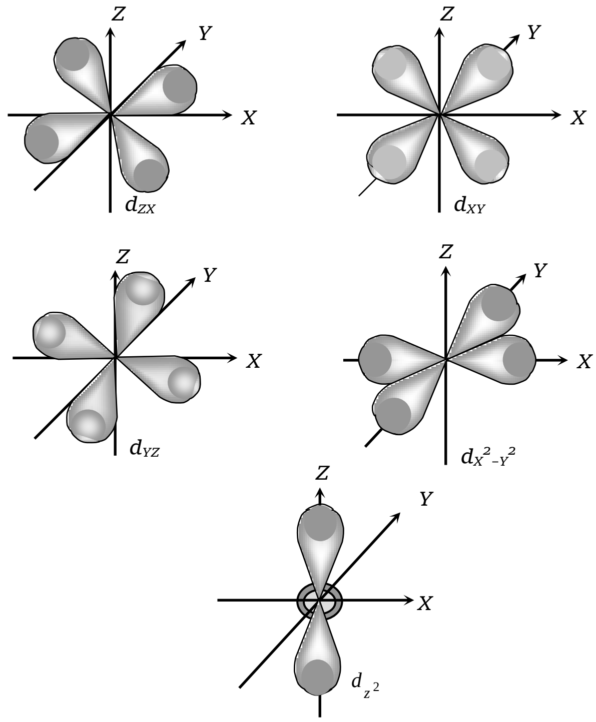


Fig. 2.8

(4) **Shape of '*f*' orbital**

(i) For the '*f*' orbital l=3 then the values of '*m*' are –3, –2, –1,0,+1,+2,+3. It shows that the '*f*' orbitals have seven orientation as $f_{x(x^2-y^2)}, f_{y(x^2-y^2)}, f_{z(x^2-y^2)}, f_{xyz}, f_{z^3}, f_{yz^2}$ and f_{xz^2}.

(ii) The '*f*' orbital is complicated in shape.

Rules for filling of electrons in various orbitals

The atom is built up by filling electrons in various orbitals according to the following rules,

(1) **Aufbau's principle**

This principle states that the electrons are added one by one to the various orbitals in order of their increasing energy starting with the orbital of lowest energy. The increasing order of energy of various orbitals is

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

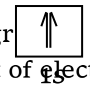
(2) **(n+l) Rule**

In neutral isolated atom, the lower the value of (n + l) for an orbital, lower is its energy. However, if the two different types of orbitals have the same value of (n + l), the orbitals with lower value of n has lower energy.

(3) **Pauli's exclusion principle**

According to this principle "no two electrons in an atom will have same value of all the four quantum numbers".

If one electron in an atom has the quantum numbers $n = 1$, $l = 0$, $m = 0$ and $s = +1/2$, no other electron can have the same four quantum numbers. In other words, we cannot place two electrons with the same value of s in a $1s$ orbital.

The orbital diagr  does not represent a possible arrangement of electrons

Because there are only two possible values of s, an orbital can hold not more than two electrons.

(4) **Hund's Rule of maximum multiplicity**

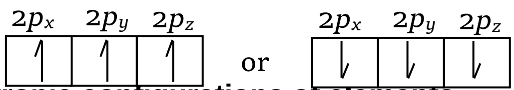
This rule deals with the filling of electrons in the orbitals having equal energy (degenerate orbitals). According to this rule,

"Electron pairing in p, d and f orbitals cannot occur until each orbitals of a given subshell contains one electron each or is singly occupied".

This is due to the fact that electrons being identical in charge, repel each other when present in the same orbital. This repulsion can however be minimised if two electrons move as far apart as possible by occupying different degenerate orbitals. All the unpaired electrons in a degenerate set of orbitals will have same spin.

As we now know the Hund's rule, let us see how the three electrons are arranged in p orbitals.

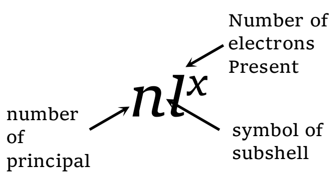
The important point ot be remembered is that all the singly occupied orbitals should have electrons with parallel spins i.e in the same direction either-clockwise or anticlockwise.



Electronic configurations of elements

On the basis of the elecronic configuration principles the electronic configuration of various elements are given in the following table :

The above method of writing the electronic configurations is quite cumbersome. Hence, usually the electronic configuration of the atom of any element is simply represented by the notation.



Some Unexpected Electronic Configuration

Some of the exceptions are important though, because they occur with common elements, notably chromium and copper.

Cu has 29 electrons. Its excepted electronic configuration is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9$ but in reality the configuration is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$ as this configuration is more stable. Similarly *Cr* has the configuration of $1s^2 2s^2 sp^6 3s^2 3p^6 4s^1 3d^5$ instead of $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4$.

Factors responsible for the extra stability of half-filled and completely filled subshells,

(i) *Symmetrical distribution* : It is well known fact that symmetry leads to stability. Thus the electronic configuration in which all the orbitals of the same subshell are either completely filled or are exactly half filled are more stable because of symmetrical distribution of electrons.

(ii) *Exchange energy* : The electrons with parallel spins present in the degenerate orbitals tend to exchange their position. The energy released during this exchange is called exchange energy. The number of exchanges that can take place is maximum when the degenerate orbtials (orbitals of same subshell having equal energy) are exactly half-filled or completely. As a result, the exchange energy is maximum and so it the stability.

Tips & Tricks

- ✍ All lines in the visible region are of Balmer series but reverse is not true i.e., all Balmer lines will not fall in visible region.
- ✍ A part of an atom up to penultimate shell is a kernel or atomic core.
- ✍ If the energy supplied to hydrogen atom is less than 13.6 eV it will awpt or absorb only those quanta which can take it to a certain higher energy level i.e., all those photons having energy less than or more than a particular energy level will not be absorbed by hydrogen atom, but if energy supplied to hydrogen atom is more than 13.6eV then all photons are absorbed and excess energy appear as kinetic energy of emitted photo electron.
- ✍ No of nodes in any orbital $=(n-l-1)$
- ✍ No of nodal planes in an orbitals $=l$
- ✍ The d orbital which does not have four lobes is d_{z^2}
- ✍ The d orbital whose lobes lie along the axis is $d_{x^2-y^2}$
- ✍ Spin angular momentum $=\sqrt{s(s+1)}\,\frac{h}{2\pi}$
- ✍ Total spin $=\pm\frac{n}{2}$; where n is no of unpaired e^-
- ✍ Magnetic moment $=\sqrt{n(n+2)}$ B.M. (Bohr magnetron) of n unpaired e^-
- ✍ Ion with unpaired electron in d or f orbital will be coloured.
- ✍ Exception of E.C. are $Cr(24)$, $Cu(29)$, $Mo(42)$, $Ag(47)$, $W(74)$, $Au(79)$.
- ✍ No. of waves $n=\frac{2\pi r}{\lambda}$ (where $\lambda=\frac{h}{mv}$)
- ✍ No. of revolutions of e^- per second is $=\frac{v}{2\pi r}$.
- ✍ The solution of schrodinger wave equation gives principal, azimuthal and magnetic quantum numbers but not the spin quantum number.
- ✍ In the Rydberg formula, when $n_2=\infty$ the line produced is called the limiting line of that series.
- ✍ Among various forms of visible light, violet colour has shortest wavelength, highest frequency and highest energy.
- ✍ Red coloured light has largest wavelength, least frequency and lowest energy in visible light.
- ✍ Elements give line spectra. The line spectrum is characteristic of the excited atom producing it. No two elements have identical line spectrum.
- ✍ The line spectrum results from the emission of radiations from the atoms of the elements and is therefore called as atomic spectrum.
- ✍ Atoms give line spectra (known as atomic spectrum) and the molecules give band spectra (known as molecular spectrum).
- ✍ The negative potential at which the photoelectric current becomes zero is called cut off potential or stopping potential.
- ✍ When energy or frequency of scattered ray is lesser than the incident ray, it is known as Compton effect.
- ✍ The instrument used to record solar spectrum is called spectrometer or spectrograph developed by Bunsen and Kirchoff in 1859.
- ✍ The intensities of spectral lines decreases with

increase in the value of n. For example, the intensity of first Lyman line $(2 \to 1)$ is greater than second line $(3 \to 1)$ and so on.

✍ In Balmer series of hydrogen spectrum the first line $(3 \to 2)$ is also known as L_α line. The second line $(4 \to 2)$ is L_β line. The line from infinity energy shell is called limiting line.

Ordinary Thinking

Objective Questions

Discovery and Properties of anode, cathode rays neutron and Nuclear structure

1. A neutral atom (Atomic no. > 1) consists of **[CPMT 1982]**

(a) Only protons
(b) Neutrons + protons
(c) Neutrons + electrons
(d) Neutron + proton + electron

2. The nucleus of the atom consists of **[CPMT 1973, 74, 78, 83, 84; MADT Bihar 1980; DPMT 1982, 85; MP PMT 1999]**

(a) Proton and neutron
(b) Proton and electron
(c) Neutron and electron
(d) Proton, neutron and electron

3. The size of nucleus is of the order of **[CPMT 1982; MP PMT 1991]**

(a) $10^{-12} m$ (b) $10^{-8} m$
(c) $10^{-15} m$ (d) $10^{-10} m$

4. Positive ions are formed from the neutral atom by the **[CPMT 1976]**

(a) Increase of nuclear charge
(b) Gain of protons
(c) Loss of electrons
(d) Loss of protons

5. The electron is **[DPMT 1982; MADT Bihar 1980]**

(a) α-ray particle (b) β-ray particle
(c) Hydrogen ion (d) Positron

6. Who discovered neutron **[IIT 1982; BITS 1988;CPMT 1977; NCERT 1974; MP PMT 1992; MP PET 2002]**

(a) James Chadwick (b) William Crooks
(c) J.J. Thomson (d) Rutherford

7. The ratio of charge and mass would be greater for **[BHU 2005]**

(a) Proton (b) Electron
(c) Neutron (d) Alpha

8. Magnitude of *K.E.* in an orbit is equal to **[BCECE 2005]**

(a) Half of the potential energy

(b) Twice of the potential energy

(c) One fourth of the potential energy

(d) None of these

9. The density of neutrons is of the order **[NCERT 1980]**

(a) $10^3 kg/cc$ (b) $10^6 kg/cc$

(c) $10^9 kg/cc$ (d) $10^{11} kg/cc$

10. The discovery of neutron becomes very late because **[CPMT 1987; AIIMS 1998]**

(a) Neutrons are present in nucleus

(b) Neutrons are highly unstable particles

(c) Neutrons are chargeless

(d) Neutrons do not move

11. The fundamental particles present in the nucleus of an atom are

(a) Alpha particles and electrons

(b) Neutrons and protons

(c) Neutrons and electrons

(d) Electrons, neutrons and protons

12. The order of density in nucleus is **[NCERT 1981, CPMT 1981, 2003]**

(a) $10^8 kg/cc$ (b) $10^{-8} kg/cc$

(c) $10^{-9} kg/cc$ (d) $10^{12} kg/cc$

13. Cathode rays are **[JIPMER 1991; NCERT 1976]**

(a) Protons (b) Electrons

(c) Neutrons (d) α -particles

14. Number of neutron in C^{12} is **[BCECE 2005]**

(a) 6 (b) 7

(c) 8 (d) 9

15. Heaviest particle is **[DPMT 1983; MP PET 1999]**

(a) Meson (b) Neutron

(c) Proton (d) Electron

16. Penetration power of proton is **[BHU 1985; CPMT 1982, 88]**

(a) More than electron (b) Less than electron

(c) More than neutron (d) None

17. An elementary particle is **[CPMT 1973]**

(a) An element present in a compound

(b) An atom present in an element

(c) A sub-atomic particle

(d) A fragment of an atom

18. The nucleus of helium contains **[CPMT 1972; DPMT 1982]**

(a) Four protons

(b) Four neutrons

(c) Two neutrons and two protons

(d) Four protons and two electrons

19. Which is correct statement about proton **[CPMT 1979; MP PMT 1985; NCERT 1985; MP PET 1999]**

(a) Proton is nucleus of deuterium

(b) Proton is ionized hydrogen molecule

(c) Proton is ionized hydrogen atom

(d) Proton is α -particle

20. Cathode rays are made up of **[AMU 1983]**

(a) Positively charged particles

(b) Negatively charged particles

(c) Neutral particles

(d) None of these

21. Anode rays were discovered by **[DPMT 1985]**

(a) Goldstein (b) J. Stoney

(c) Rutherford (d) J.J. Thomson

[CPMT 1983, 84]

22. The radius of an atom is of the order of **[AMU 1982; IIT 1985; MP PMT 1995]**

(a) $10^{-10} cm$ (b) $10^{-13} cm$

(c) $10^{-15} cm$ (d) $10^{-8} cm$

23. Neutron possesses **[CPMT 1982]**

(a) Positive charge (b) Negative charge

(c) No charge (d) All are correct

24. Neutron is a fundamental particle carrying **[CPMT 1990]**

(a) A charge of +1 unit and a mass of 1 unit

(b) No charge and a mass of 1 unit

(c) No charge and no mass

(d) A charg of –1 and a mass of 1 unit

25. Cathode rays have **[CPMT 1982]**

(a) Mass only (b) Charge only

(c) No mass and charge (d) Mass and charge both

26. The size of nucleus is measured in **[EAMCET 1988; CPMT 1994]**

(a) amu (b) Angstrom

(c) Fermi (d) cm

27. Which phrase would be incorrect to use **[AMU (Engg.) 1999]**

(a) A molecular of a compound

(b) A molecule of an element

(c) An atom of an element

(d) None of these

28. Which one of the following pairs is not correctly matched **[MP PET 2002]**

(a) Rutherford-Proton

(b) J.J. Thomsom-Electron

(c) J.H. Chadwick-Neutron

(d) Bohr-Isotope

29. Proton was discovered by

(a) Chadwick (b) Thomson

(c) Goldstein (d) Bohr

30. The minimum real charge on any particle which can exist is **[RPMT 2000]**

(a) 1.6×10^{-19} *Coulomb* (b) 1.6×10^{-10} *Coulomb*

(c) 4.8×10^{-10} *Coulomb* (d) Zero

31. The nature of anode rays depends upon **[MP PET 2004]**

(a) Nature of electrode (b) Nature of residual gas

(c) Nature of discharge tube (d) All the above

32. One would expect proton to have very large **[Pb. CET 2004]**

(a) Ionization potential (b) Radius

(c) Charge (d) Hydration energy

33. The mass of a mol of proton and electron is

(a) $6.023\times10^{23}\,g$ (b) $1.008\,g$ and $0.55\,mg$

(c) $9.1\times10^{-28}\,kg$ (d) $2\,gm$

34. The average distance of an electron in an atom from its nucleus is of the order of

(a) $10^{6}\,m$ (b) $10^{-6}\,m$

(c) $10^{-10}\,m$ (d) $10^{-15}\,m$

35. The mass of 1 mole of electrons is **[Pb. CET 2004]**

(a) $9.1\times10^{-28}\,g$ (b) $1.008\,mg$

(c) $0.55\,mg$ (d) $9.1\times10^{-27}\,g$

36. The ratio of specific charge of a proton and an α-particle is **[MP PET 1999]**

(a) 2 : 1 (b) 1 : 2

(c) 1 : 4 (d) 1 : 1

37. Ratio of masses of proton and electron is **[BHU 1998]**

(a) Infinite (b) 1.8×10^{3}

(c) 1.8 (d) None of these

38. Splitting of signals is caused by **[Pb. PMT 2000]**

(a) Proton (b) Neutron

(c) Positron (d) Electron

39. The proton and neutron are collectively called as **[MP PET 2001]**

(a) Deutron (b) Positron

(c) Meson (d) Nucleon

40. Which of the following has the same mass as that of an electron **[AFMC 2002]**

[AFMC 2004]

(a) Photon (b) Neutron

(c) Positron (d) Proton

41. What is the ratio of mass of an electron to the mass of a proton **[UPSEAT 2004]**

(a) 1 : 2 (b) 1 : 1

(c) 1 : 1837 (d) 1 : 3

Atomic number, Mass number, Atomic species

1. The number of electrons in an atom of an element is equal to its **[BHU 1979]**

(a) Atomic weight (b) Atomic number

(c) Equivalent weight (d) Electron affinity

2. The nucleus of the element having atomic number 25 and atomic weight 55 will contain **[CPMT 1986; MP PMT 1987]**

(a) 25 protons and 30 neutrons

(b) 25 neutrons and 30 protons

(c) 55 protons

(d) 55 neutrons

3. If W is atomic weight and N is the atomic number of an element, then **[CPMT 1971, 80, 89]**

[MP PET 1996]

(a) Number of $e^{-1}=W-N$

(b) Number of ${}_0n^1=W-N$

(c) Number of ${}_1H^1=W-N$

(d) Number of ${}_0n^1=N$

4. The total number of neutrons in dipositive zinc ions with mass number 70 is **[IIT 1979; Bihar MEE 1997]**

(a) 34 (b) 40

(c) 36 (d) 38

5. Which of the following are isoelectronic with one another **[NCERT 1983; EAMCET 1989]**

(a) Na^{+} and Ne (b) K^{+} and O

(c) Ne and O (d) Na^{+} and K^{+}

6. The number of electrons in one molecule of CO_2 are **[IIT 1979; MP PMT 1994; RPMT 1999]**

(a) 22 (b) 44

(c) 66 (d) 88

7. Chlorine atom differs from chloride ion in the number of **[NCERT 1972; MP PMT 1995]**

(a) Proton (b) Neutron

(c) Electrons (d) Protons and electrons

8. CO has same electrons as **or** the ion that is isoelectronic with CO is **[CPMT 1984; IIT 1982; EAMCET 1990; CBSE PMT 1997]**

(a) N_2^+ (b) CN^-

(c) O_2^+ (d) O_2^-

9. The mass of an atom is constituted mainly by **[DPMT 1984, 91; AFMC 1990]**

(a) Neutron and neutrino (b) Neutron and electron

(c) Neutron and proton (d) Proton and electron

10. The atomic number of an element represents **[CPMT 1983; CBSE PMT 1990; NCERT 1973; AMU 1984]**

(a) Number of neutrons in the nucleus

(b) Number of protons in the nucleus

(c) Atomic weight of element

(d) Valency of element

11. An atom has 26 electrons and its atomic weight is 56. The number of neutrons in the nucleus of the atom will be **[CPMT 1980]**

(a) 26 (b) 30

(c) 36 (d) 56

12. The most probable radius (in *pm*) for finding the electron in He^+ is **[AIIMS 2005]**

(a) 0.0 (b) 52.9

(c) 26.5 (d) 105.8

13. The number of unpaired electrons in the Fe^{2+} ion is **[MP PET 1989; KCET 2000]**

(a) 0 (b) 4

(c) 6 (d) 3

14. A sodium cation has different number of electrons from

(a) O^{2-} (b) F^-

(c) Li^+ (d) Al^{+3}

15. An atom which has lost one electron would be **[CPMT 1986]**

(a) Negatively charged

(b) Positively charged

(c) Electrically neutral

(d) Carry double positive charge

16. Number of electrons in the outermost orbit of the element of atomic number 15 is **[CPMT 1988, 93]**

(a) 1 (b) 3

(c) 5 (d) 7

17. The atomic weight of an element is double its atomic number. If there are four electrons in $2p$ orbital, the element is **[AMU 1983]**

(a) C (b) N

(c) O (d) Ca

18. An atom has the electronic configuration of $1s^2, 2s^2 2p^6$, $3s^2 3p^6 3d^{10}, 4s^2 4p^5$. Its atomic weight is 80. Its atomic number and the number of neutrons in its nucleus shall be **[MP PMT 1987]**

(a) 35 and 45 (b) 45 and 35

(c) 40 and 40 (d) 30 and 50

19. Which of the following particles has more electrons than neutrons

(a) C (b) F^-

(c) O^{-2} (d) Al^{+3}

20. Compared with an atom of atomic weight 12 and atomic number 6, the atom of atomic weight 13 and atomic number 6 **[NCERT 1971]**

(a) Contains more neutrons (b) Contains more electrons

(c) Contains more protons (d) Is a different element

21. In the nucleus of $_{20}Ca^{40}$ there are **[CPMT 1990; EAMCET 1991]**

(a) 40 protons and 20 electrons

(b) 20 protons and 40 electrons

(c) 20 protons and 20 neutrons

(d) 20 protons and 40 neutrons

22. Na^+ ion is isoelectronic with **[CPMT 1990]**

(a) Li^+ (b) Mg^{+2}

(c) Ca^{+2} (d) Ba^{+2}

23. Ca has atomic no. 20 and atomic weight 40. Which of the following statements is not correct about Ca atom **[MP PET 1993]**

(a) The number of electrons is same as the number of neutrons

(b) The number of nucleons is double of the number of electrons

(c) The number of protons is half of the number of neutrons

(d) The number of nucleons is double of the atomic number

24. Pick out the isoelectronic structures from the following

$\underset{\text{I}}{CH_3^+}$ $\underset{\text{II}}{H_3O^+}$ $\underset{\text{III}}{NH_3}$ $\underset{\text{IV}}{CH_3^-}$ **[IIT 1993]**

(a) I and II (b) I and IV

(c) I and III (d) II, III and IV

25. Number of electrons in $-CONH_2$ is **[AMU 1988]**

(a) 22 (b) 24

(c) 20 (d) 28

26. The atomic number of an element having the valency shell electronic configuration $4s^2 4p^6$ is **[MP PMT 1**

(a) 35 (b) 36

(c) 37 (d) 38

27. The present atomic weight scale is based on

[EAMCET 1988; MP PMT 2002]

(a) C^{12} (b) O^{16}

(c) H^{1} (d) C^{13}

28. Isoelectronic species are **[EAMCET 1989]**

(a) K^{+}, Cl^{-} (b) Na^{+}, Cl^{-}

(c) Na, Ar (d) Na^{+}, Ar

29. If the atomic weight of an element is 23 times that of the lightest element and it has 11 protons, then it contains

[EAMCET 1986; AFMC 1989]

(a) 11 protons, 23 neutrons, 11 electrons

(b) 11 protons, 11 neutrons, 11 electrons

(c) 11 protons, 12 neutrons, 11 electrons

(d) 11 protons, 11 neutrons, 23 electrons

30. Which of the following oxides of nitrogen is isoelectronic with CO_2 **[CBSE PMT 1990]**

(a) NO_2 (b) N_2O

(c) NO (d) N_2O_2

31. The ratio between the neutrons in C and Si with respect to atomic masses 12 and 28 is

(a) 2 : 3 (b) 3 : 2

(c) 3 : 7 (d) 7 : 3

32. The atomic number of an element is always equal to

[MP PMT 1994]

(a) Atomic weight divided by 2

(b) Number of neutrons in the nucleus

(c) Weight of the nucleus

(d) Electrical charge of the nucleus

33. Which of the following is isoelectronic with carbon atom

[MP PMT 1994; UPSEAT 2000]

(a) Na^{+} (b) Al^{3+}

(c) O^{2-} (d) N^{+}

34. CO_2 is isostructural with

[IIT 1986; MP PMT 1986, 94, 95]

(a) $SnCl_2$ (b) SO_2

(c) $HgCl_2$ (d) All the above

35. The hydride ions (H^{-}) are isoelectronic with

[AFMC 1995; Bihar MEE 1997]

(a) Li (b) He^{+}

(c) He (d) Be

36. The number of electrons in the nucleus of C^{12} is

[AFMC 1995]

(a) 6 (b) 12

(c) 0 (d) 3

37. An element has electronic configuration 2, 8, 18, 1. If its atomic weight is 63, then how many neutrons will be present in its nucleus

(a) 30 (b) 32

(c) 34 (d) 33

38. The nucleus of the element ${}_{21}E^{45}$ contains

(a) 45 protons and 21 neutrons

(b) 21 protons and 24 neutrons

(c) 21 protons and 45 neutrons

(d) 24 protons and 21 neutrons

39. Neutrons are found in atoms of all elements except in

[MP PMT 1997]

(a) Chlorine (b) Oxygen

(c) Argon (d) Hydrogen

40. The mass number of an anion, X^{3-}, is 14. If there are ten electrons in the anion, the number of neutrons in the nucleus of atom, X_2 of the element will be

[MP PMT 1999]

(a) 10 (b) 14

(c) 7 (d) 5

41. Which of the following are isoelectronic species $I = CH_3^{+}, II - NH_2, III - NH_4^{+}, IV - NH_3$ **[CPMT 1999]** **[EAMCET 1990]**

(a) I, II, III (b) II, III, IV

(c) I, II, IV (d) I and II

42. The charge on the atom containing 17 protons, 18 neutrons and 18 electrons is **[AIIMS 1996]**

(a) $+1$ (b) -2

(c) -1 (d) Zero

43. Number of unpaired electrons in inert gas is **[CPMT 1996]**

(a) Zero (b) 8

(c) 4 (d) 18

44. In neutral atom, which particles are equivalent

[RPMT 1997]

(a) p^{+}, e^{+} (b) e^{-}, e^{+}

(c) e^{-}, p^{+} (d) p^{+}, n^{o}

45. Nuclei tend to have more neutrons than protons at high mass numbers because **[Roorkee Qualifying 1998]**

(a) Neutrons are neutral particles

(b) Neutrons have more mass than protons

(c) More neutrons minimize the coulomb repulsion

(d) Neutrons decrease the binding energy

46. Which one of the following is not isoelectronic with O^{2-}

[CBSE PMT 1994]

(a) N^{3-} (b) F^{-}

(c) Tl^{+} (d) Na^{+}

47. The number of electrons in $[{}^{40}_{19}K]^{-1}$ is

[CPMT 1997; AFMC 1999]

(a) 19 (b) 20

(c) 18 (d) 40

48. The number of electrons and neutrons of an element is 18 and 20 respectively. Its mass number is

[CPMT 1997; Pb. PMT 1999; MP PMT 1999]

(a) 17 (b) 37
(c) 2 (d) 38

49. Number of protons, neutrons and electrons in the element $^{231}_{89}Y$ is **[AFMC 1997]**

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

50. Be^{2+} is isoelectronic with **[EAMCET 1998]**

(a) Mg^{2+} (b) Na^{+}
(c) Li^{+} (d) H^{+}

51. An isostere is **[UPSEAT 1999]**

(a) NO_2^- and O_3 (b) NO_2^- and PO_4^{3-}
(c) CO_2, N_2O, NO_3^- (d) ClO_4^- and OCN^-

52. Nitrogen atom has an atomic number of 7 and oxygen has an atomic number 8. The total number of electrons in a nitrate ion will be **[Pb. PMT 2000]**

(a) 8 (b) 16
(c) 32 (d) 64

53. If molecular mass and atomic mass of sulphur are 256 and 32 respectively, its atomicity is **[RPET 2000]**

(a) 2 (b) 8
(c) 4 (d) 16

54. The nitride ion in lithium nitride is composed of

[KCET 2000]

(a) 7 protons + 10 electrons
(b) 10 protons + 10 electrons
(c) 7 protons + 7 protons
(d) 10 protons + 7 electrons

55. The atomic number of an element is 17. The number of orbitals containing electron pairs in its valence shell is

[CPMT 2001]

(a) Eight (b) Six
(c) Three (d) Two

56. The atomic number of an element is 35 and mass number is 81. The number of electrons in the outer most shell is

[UPSEAT 2001]

(a) 7 (b) 6
(c) 5 (d) 3

57. Which of the following is not isoelectronic**[MP PET 2002]**

(a) Na^{+} (b) Mg^{2+}
(c) O^{2-} (d) Cl^{-}

58. The charge of an electron is $-1.6\times10^{-19}C$. The value of free charge on Li^{+} ion will be

[AFMC 2002; KCET (Engg.) 2002]

(a) $3.6\times10^{-19}C$ (b) $1\times10^{-19}C$
(c) $1.6\times10^{-19}C$ (d) $2.6\times10^{-19}C$

59. Iso-electronic species is **[RPMT 2002]**

(a) F^-, O^{-2} (b) F^-, O
(c) F^-, O^+ (d) F^-, O^{+2}

60. An element have atomic weight 40 and it's electronic configuration is $1s^2 2s^2 2p^6 3s^2 3p^6$. Then its atomic number and number of neutrons will be **[RPMT 2002]**

(a) 18 and 22 (b) 22 and 18
(c) 26 and 20 (d) 40 and 18

61. The nucleus of tritium contains **[MP PMT 2002]**

(a) 1 proton + 1 neutron (b) 1 proton + 3 neutron
(c) 1 proton + 0 neutron (d) 1 proton + 2 neutron

62. Which one of the following groupings represents a collection of isoelectronic species **[AIEEE 2003]**

(a) Na^+, Ca^{2+}, Mg^{2+} (b) N^{3-}, F^-, Na^+
(c) Be, Al^{3+}, Cl^- (d) Ca^{2+}, Cs^+, Br

63. Which of the following are isoelectronic and isostructural $NO_3^-, CO_3^{2-}, ClO_3^-, SO_3$ **[IIT Screening 2003]**

(a) NO_3^-, CO_3^{2-} (b) SO_3, NO_3^-
(c) ClO_3^-, CO_3^{2-} (d) CO_3^{2-}, SO_3

64. The number of electrons in Cl^- ion is **[MP PMT 2003]**

(a) 19 (b) 20
(c) 18 (d) 35

65. The number of neutron in tritium is **[CPMT 2003]**

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

66. Tritium is the isotope of **[CPMT 2003]**

(a) Hydrogen (b) Oxygen
(c) Carbon (d) Sulpher

67. The atomic number of an element is 35. What is the total number of electrons present in all the *p*-orbitals of the ground state atom of that element

[EAMCET (Engg.) 2003]

(a) 6 (b) 11
(c) 17 (d) 23

68. The nucleus of an element contain 9 protons. Its valency would be

(a) 1 (b) 3
(c) 2 (d) 5

69. The compound in which cation is isoelectronic with anion is

[UPSEAT 2004]

(a) $NaCl$ (b) CsF
(c) NaI (d) K_2S

70. Which among the following species have the same number of electrons in its outermost as well as penultimate shell

[DCE 2004]

(a) Mg^{2+} (b) O^{2-}
(c) F^- (d) Ca^{2+}

71. Six protons are found in the nucleus of

[CPMT 1977, 80, 81; NCERT 1975, 78]

(a) Boron (b) Lithium

(c) Carbon (d) Helium

72. The nitrogen atom has 7 protons and 7 electrons, the nitride ion (N^{3-}) will have

(a) 7 protons and 10 electrons

(b) 4 protons and 7 electrons

(c) 4 protons and 10 electrons

(d) 10 protons and 7 electrons

73. Number of neutrons in heavy hydrogen atom is

[MP PMT 1986]

(a) 0 (b) 1

(c) 2 (d) 3

74. Which of the following is always a whole number

[CPMT 1976, 81, 86]

(a) Atomic weight (b) Atomic radii

(c) Equivalent weight (d) Atomic number

Atomic models and Planck's quantum theory

1. Rutherford's experiment on scattering of particles showed for the first time that the atom has

[IIT 1981; NCERT 1981; CMC Vellore 1991; CPMT 1984; Kurukshetra CEE 1998]

(a) Electrons (b) Protons

(c) Nucleus (d) Neutrons

2. Rutherford's scattering experiment is related to the size of the

[IIT 1983; MADT Bihar 1995; BHU 1995]

(a) Nucleus (b) Atom

(c) Electron (d) Neutron

3. Rutherford's alpha particle scattering experiment eventually led to the conclusion that[IIT 1986; RPMT 2002]

(a) Mass and energy are related

(b) Electrons occupy space around the nucleus

(c) Neutrons are buried deep in the nucleus

(d) The point of impact with matter can be precisely determined

4. Bohr's model can explain [IIT 1985]

(a) The spectrum of hydrogen atom only

(b) Spectrum of atom or ion containing one electron only

(c) The spectrum of hydrogen molecule

(d) The solar spectrum

5. When atoms are bombarded with alpha particles, only a few in million suffer deflection, others pass out undeflected. This is because[MNR 1979; NCERT 1980; AFMC 1995]

(a) The force of repulsion on the moving alpha particle is small

(b) The force of attraction on the alpha particle to the oppositely charged electrons is very small

[NCERT 1977]

(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

6. Positronium consists of an electron and a positron (a particle which has the same mass as an electron, but opposite charge) orbiting round their common centre of mass. Calculate the value of the Rydberg constant for this system.

(a) $R_{\infty}/4$ (b) $R_{\infty}/2$

(c) $2R_{\infty}$ (d) R_{∞}

7. When α-particles are sent through a thin metal foil, most of them go straight through the foil because (one or more are correct)

(a) Alpha particles are much heavier than electrons

(b) Alpha particles are positively charged

(c) Most part of the atom is empty space

(d) Alpha particles move with high velocity

8. When an electron jumps from L to K shell

[CPMT 1983]

(a) Energy is absorbed

(b) Energy is released

(c) Energy is sometimes absorbed and sometimes released

(d) Energy is neither absorbed nor released

9. When beryllium is bombarded with α-particles, extremely penetrating radiations which cannot be deflected by electrical or magnetic field are given out. These are

[CPMT 1983]

(a) A beam of protons (b) α-rays

(c) A beam of neutrons (d) X-rays

10. Which one of the following is not the characteristic of Planck's quantum theory of radiation [AIIMS 1991]

(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 conti- nuously but in the form of small packets called quanta

(d) This magnitude of energy associated with a quantum is proportional to the frequency

11. The spectrum of He is expected to be similar to

[AIIMS 1980, 91; DPMT 1983; MP PMT 2002]

(a) H (b) Li^+

(c) Na (d) He^+

12. Energy of orbit **[DPMT 1984, 91]**

(a) Increases as we move away from nucleus

(b) Decreases as we move away from nucleus

(c) Remains same as we move away from nucleus

(d) None of these

13. Bohr model of an atom could not account for

(a) Emission spectrum

(b) Absorption spectrum

(c) Line spectrum of hydrogen

(d) Fine spectrum

14. Existence of positively charged nucleus was established by

[CBSE PMT 1991]

(a) Positive ray analysis

(b) α -ray scattering experiments

(c) X-ray analysis

(d) Discharge tube experiments

15. Electron occupies the available orbital singly before pairing in any one orbital occurs, it is **[CBSE PMT 1991]**

(a) Pauli's exclusion principle

(b) Hund's Rule

(c) Heisenberg's principle

(b) Prout's hypothesis

16. The wavelength of a spectral line for an electronic transition is inversely related to **[IIT 1988]**

(a) The number of electrons undergoing the transition

(b) The nuclear charge of the atom

(c) The difference in the energy of the energy levels involved in the transition

(d) The velocity of the electron undergoing the transition

17. When an electron drops from a higher energy level to a low energy level, then **[AMU 1985]**

(a) Energy is emitted

(b) Energy is absorbed

(c) Atomic number increases

(d) Atomic number decreases

18. Davisson and Germer's experiment showed that

[MADT Bihar 1983]

(a) β -particles are electrons

(b) Electrons come from nucleus

(c) Electrons show wave nature

(d) None of the above

19. When an electron jumps from lower to higher orbit, its energy **[MADT Bihar 1982]**

(a) Increases (b) Decreases

(c) Remains the same (d) None of these

20. Experimental evidence for the existence of the atomic nucleus comes from

(a) Millikan's oil drop experiment

(b) Atomic emission spectroscopy

(c) The magnetic bending of cathode rays

(d) Alpha scattering by a thin metal foil

21. Which of the following statements does not form part of Bohr's model of the hydrogen atom **[CBSE PMT 1989**

(a) Energy of the electrons in the orbit is quantized

(b) The electron in the orbit nearest 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. When β -particles are sent through a tin metal foil, most of them go straight through the foil as **[EAMCET**

(a) β -particles are much heavier than electrons

(b) β -particles are positively charged

(c) Most part of the atom is empty space

(d) β -particles move with high velocity

23. The energy of second Bohr orbit of the hydrogen atom is –328 $kJ\ mol^{-1}$, hence the energy of fourth Bohr orbit would be

[CBSE PMT 2005]

(a) – 41 $kJ\ mol^{-1}$ (b) –1312 $kJ\ mol^{-1}$

(c) –164 $kJ\ mol^{-1}$ (d) – 82 $kJ\ mol^{-1}$

24. When an electron revolves in a stationary orbit then

[MP PET 1994]

(a) It absorbs energy

(b) It gains kinetic energy

(c) It emits radiation

(d) Its energy remains constant

25. A moving particle may have wave motion, if

(a) Its mass is very high

(b) Its velocity is negligible

(c) Its mass is negligible

(d) Its mass is very high and velocity is negligible

26. The postulate of Bohr theory that electrons jump from one orbit to the other, rather than flow is according to

(a) The quantisation concept

(b) The wave nature of electron

(c) The probability expression for electron

(d) Heisenberg uncertainty principle

27. The frequency of an electromagnetic radiation is $2\times10^{6}\,Hz$. What is its wavelength in metres (Velocity of light $=3\times10^{8}\,ms^{-1}$)

(a) 6.0×10^{14} (b) 1.5×10^{4}

(c) 1.5×10^{2} (d) 0.66×10^{-2}

28. What is the packet of energy called **[AFMC 2005]**

(a) Electron (b) Photon

(c) Positron (d) Proton

29. The energy of an electron in n^{th} orbit of hydrogen atom is **[MP PET 1999]**

(a) $\frac{13.6}{n^4}eV$ (b) $\frac{13.6}{n^3}eV$

(c) $\frac{13.6}{n^2}eV$ (d) $\frac{13.6}{n}eV$

30. If wavelength of photon is $2.2\times10^{-11}\,m$, $h=6.6\times10^{-34}\,J\text{-}sec$, then momentum of photon is **[MP PET 1999]**

(a) $3\times10^{-23}\,kg\,ms^{-1}$ (b) $3.33\times10^{22}\,kg\,ms^{-1}$

(c) $1.452\times10^{-44}\,kg\,ms^{-1}$ (d) $6.89\times10^{43}\,kg\,ms^{-1}$

31. The expression for Bohr's radius of an atom is **[MP PMT 1999]**

(a) $r=\frac{n^2h^2}{4\pi^2me^4z^2}$ (b) $r=\frac{n^2h^2}{4\pi^2me^2z}$

(c) $r=\frac{n^2h^2}{4\pi^2me^2z^2}$ (d) $r=\frac{n^2h^2}{4\pi^2m^2e^2z^2}$

32. The energy of an electron revolving in n^{th} Bohr's orbit of an atom is given by the expression **[MP PMT 1999]**

(a) $E_n=-\frac{2\pi^2m^4e^2z^2}{n^2h^2}$ (b) $E_n=-\frac{2\pi^2me^2z^2}{n^2h^2}$

(c) $E_n=-\frac{2\pi^2me^4z^2}{n^2h^2}$ (d) $E_n=-\frac{2\pi m^2e^2z^4}{n^2h^2}$

33. Who modified Bohr's theory by introducing elliptical orbits for electron path **[CBSE PMT 1999; AFMC 2003]**

(a) Hund (b) Thomson

(c) Rutherford (d) Sommerfield

34. Bohr's radius can have **[DPMT 1996]**

(a) Discrete values (b) $+ve$ values

(c) $-ve$ values (d) Fractional values

35. The first use of quantum theory to explain the structure of atom was made by **[IIT 1997; CPMT 2001; J&K CET 2005]**

(a) Heisenberg (b) Bohr

(c) Planck (d) Einstein

36. An electronic transition from 1s orbital of an atom causes **[JIPMER 1997]**

(a) Absorption of energy

(b) Release of energy

(c) Both release or absorption of energy

(d) Unpredictable

37. In an element going away from nucleus, the energy of particle **[RPMT 1997]**

(a) Decreases (b) Not changing

(c) Increases (d) None of these

38. The α-particle scattering experiment of Rutherford concluded that

(a) The nucleus is made up of protons and neutrons

(b) The number of electrons is exactly equal to number of protons in atom

(c) The positive charge of the atom is concentrated in a very small space

(d) Electrons occupy discrete energy levels

39. Wavelength associated with electron motion **[BHU 1998]**

(a) Increases with increase in speed of electron

(b) Remains same irrespective of speed of electron

(c) Decreases with increase in speed of e^-

(d) Is zero

40. The element used by Rutherford in his famous scattering experiment was

(a) Gold (b) Tin

(c) Silver (d) Lead

41. If electron falls from $n=3$ to $n=2$, then emitted energy is **[AFMC 1997; MP PET 2003]**

(a) $10.2eV$ (b) $12.09eV$

(c) $1.9eV$ (d) $0.65eV$

42. The radius of the nucleus is related to the mass number A by

(a) $R=R_oA^{1/2}$ (b) $R=R_oA$

(c) $R=R_oA^2$ (d) $R=R_oA^{1/3}$

43. The specific charge of proton is $9.6\times10^{6}\,C\,kg^{-1}$ then for an α-particle it will be **[MH CET 1999]**

(a) $38.4\times10^{7}\,C\,kg^{-1}$ (b) $19.2\times10^{7}\,C\,kg^{-1}$

(c) $2.4\times10^{7}\,C\,kg^{-1}$ (d) $4.8\times10^{7}\,C\,kg^{-1}$

44. In hydrogen spectrum the different lines of Lyman series are present is **[UPSEAT 1999]**

(a) UV field (b) IR field

(c) Visible field (d) Far IR field

45. Which one of the following is considered as the main postulate of Bohr's model of atom **[AMU 2000]**

(a) Protons are present in the nucleus

(b) Electrons are revolving around the nucleus

(c) Centrifugal force produced due to the revolving electrons balances the force of attraction between the electron and the protons

(d) Angular momentum of electron is an integral multiple of $\frac{h}{2\pi}$

46. The electronic energy levels of the hydrogen atom in the Bohr's theory are called **[AMU 2000]**
(a) Rydberg levels (b) Orbits
(c) Ground states (d) Orbitals

47. The energy of a photon is calculated by **[Pb. PMT 2000]**
(a) $E = h\nu$ (b) $h = E\nu$
(c) $h = \frac{E}{\nu}$ (d) $E = \frac{h}{\nu}$

48. Visible range of hydrogen spectrum will contain the following series **[RPET 2000]**
(a) Pfund (b) Lyman
(c) Balmer (d) Brackett

49. Radius of the first Bohr's orbit of hydrogen atom is **[RPET 2000]**
(a) 1.06 Å (b) 0.22 Å
(c) 0.28 Å (d) 0.53 Å

50. In Balmer series of hydrogen atom spectrum which electronic transition causes third line **[MP PMT 2000]**
(a) Fifth Bohr orbit to second one
(b) Fifth Bohr orbit to first one
(c) Fourth Bohr orbit to second one
(d) Fourth Bohr orbit to first one

51. Energy of electron of hydrogen atom in second Bohr orbit is **[MP PMT 2000]**
(a) $-5.44 \times 10^{-19}\ J$ (b) $-5.44 \times 10^{-19}\ kJ$
(c) $-5.44 \times 10^{-19}\ cal$ (d) $-5.44 \times 10^{-19}\ eV$

52. If change in energy $(\Delta E) = 3 \times 10^{-8}\ J$, $h = 6.64 \times 10^{-34}\ J\text{-}s$ and $c = 3 \times 10^{8}\ m/s$, then wavelength of the light is **[CBSE PMT 2000]**
(a) 6.36×10^{3} Å (b) 6.36×10^{5} Å
(c) 6.64×10^{-8} Å (d) 6.36×10^{18} Å

53. The radius of first Bohr's orbit for hydrogen is 0.53 Å. The radius of third Bohr's orbit would be **[MP PMT 2001]**
(a) 0.79 Å (b) 1.59 Å
(c) 3.18 Å (d) 4.77 Å

54. Rutherford's α-particle scattering experiment proved that atom has **[MP PMT 2001]**
(a) Electrons (b) Neutron
(c) Nucleus (d) Orbitals

55. Wavelength of spectral line emitted is inversely proportional to **[CPMT 2001]**
(a) Radius (b) Energy
(c) Velocity (d) Quantum number

56. The energy of a radiation of wavelength 8000 Å is E_1 and energy of a radiation of wavelength 16000 Å is E_2. What is the relation between these two **[Kerala C**
(a) $E_1 = 6E_2$ (b) $E_1 = 2E_2$
(c) $E_1 = 4E_2$ (d) $E_1 = 1/2E_2$
(e) $E_1 = E_2$

57. The formation of energy bonds in solids are in accordance with **[DCE 2001]**
(a) Heisenberg's uncertainty principle
(b) Bohr's theory
(c) Ohm's law
(d) Rutherford's atomic model

58. The frequency of yellow light having wavelength 600 *nm* is **[MP PET 2002]**
(a) $5.0 \times 10^{14}\ Hz$ (b) $2.5 \times 10^{7}\ Hz$
(c) $5.0 \times 10^{7}\ Hz$ (d) $2.5 \times 10^{14}\ Hz$

59. The value of the energy for the first excited state of hydrogen atom will be **[MP PET 2002]**
(a) $-13.6\ eV$ (b) $-3.40\ eV$
(c) $-1.51\ eV$ (d) $-0.85\ eV$

60. Bohr model of atom is contradicted by **[MP PMT 2002]**
(a) Pauli's exclusion principle
(b) Planck quantum theory
(c) Heisenberg uncertainty principle
(d) All of these

61. Which of the following is not true in Rutherford's nuclear model of atom **[Orissa JEE 2002]**
(a) Protons and neutrons are present inside nucleus
(b) Volume of nucleus is very small as compared to volume of atom
(c) The number of protons and neutrons are always equal
(d) The number of electrons and protons are always equal

62. The emission spectrum of hydrogen is found to satisfy the expression for the energy change. ΔE (in joules) such that $\Delta E = 2.18 \times 10\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) J$ where n_1 = 1, 2, 3..... and n_2 = 2, 3, 4....... The spectral lines correspond to Paschen series to
(a) $n_1 = 1$ and $n_2 = 2, 3, 4$
(b) $n_1 = 3$ and $n_2 = 4, 5, 6$
(c) $n_1 = 1$ and $n_2 = 3, 4, 5$
(d) $n_1 = 2$ and $n_2 = 3, 3, 5$
(e) $n_1 = 1$ and $n_2 =$ infinity

63. The ratio between kinetic energy and the total energy of the electrons of hydrogen atom according to Bohr's model is **[Pb. PMT 2002]**
(a) 2 : 1 (b) 1 : 1

(c) 1 : – 1 (d) 1 : 2

64. Energy of the electron in Hydrogen atom is given by **[AMU (Engg.) 2002]**

(a) $E_n = -\frac{131.38}{n^2} kJ\,mol^{-1}$ (b) $E_n = -\frac{131.33}{n} kJ\,mol^{-1}$

(c) $E_n = -\frac{1313.3}{n^2} kJ\,mol^{-1}$ (d) $E_n = -\frac{313.13}{n^2} kJ\,mol^{-1}$

65. Ratio of radii of second and first Bohr orbits of *H* atom **[BHU 2003]**

(a) 2 (b) 4

(c) 3 (d) 5

66. The frequency corresponding to transition $n = 2$ to $n = 1$ in hydrogen atom is **[MP PET 2003]**

(a) $15.66 \times 10^{10}\ Hz$ (b) $24.66 \times 10^{14}\ Hz$

(c) $30.57 \times 10^{14}\ Hz$ (d) $40.57 \times 10^{24}\ Hz$

67. The mass of a photon with a wavelength equal to $1.54 \times 10^{-8}\ cm$ is **[Pb. PMT 2004]**

(a) $0.8268 \times 10^{-34}\ kg$ (b) $1.2876 \times 10^{-33}\ kg$

(c) $1.4285 \times 10^{-32}\ kg$ (d) $1.8884 \times 10^{-32}\ kg$

68. Splitting of spectral lines under the influence of magnetic field is called **[MP PET 2004]**

(a) Zeeman effect (b) Stark effect

(c) Photoelectric effect (d) None of these

69. The radius of electron in the first excited state of hydrogen atom is **[MP PMT 2004]**

(a) a_0 (b) $4a_0$

(c) $2a_0$ (d) $8a_0$

70. The ratio of area covered by second orbital to the first orbital is **[AFMC 2004]**

(a) 1 : 2 (b) 1 : 16

(c) 8 : 1 (d) 16 : 1

71. Time taken for an electron to complete one revolution in the Bohr orbit of hydrogen atom is **[Kerala PMT 2004]**

(a) $\frac{4\pi^2 mr^2}{nh}$ (b) $\frac{nh}{4\pi^2 mr}$

(c) $\frac{nh}{4\pi^2 mr^2}$ (d) $\frac{h}{2\pi mr}$

72. The radius of which of the following orbit is same as that of the first Bohr's orbit of hydrogen atom **[IIT Screening 2004]**

(a) $He^+(n=2)$ (b) $Li^{2+}(n=2)$

(c) $Li^{2+}(n=3)$ (d) $Be^{3+}(n=2)$

73. 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 $H = 2.18 \times 10^{-18} J$ atom^{-1} and $h = 6.625 \times 10^{-34}\ Js$)

(a) $3.08 \times 10^{15}\ s^{-1}$ (b) $2.00 \times 10^{15}\ s^{-1}$

(c) $1.54 \times 10^{15}\ s^{-1}$ (d) $1.03 \times 10^{15}\ s^{-1}$

74. The wavelength of the radiation emitted, when in a hydrogen atom electron falls from infinity to stationary state 1, would be (Rydberg constant $= 1.097 \times 10^7\ m^{-1}$) **[AIEEE 2004]**

(a) 406 *nm* (b) 192 *nm*

(c) 91 *nm* (d) $9.1 \times 10^{-8}\ nm$

75. In Bohr's model, atomic radius of the first orbit is γ, the radius of the 3rd orbit, is **[MP PET 1997; Pb. CET 200**

(a) $\gamma/3$ (b) γ

(c) 3γ (d) 9γ

76. According to Bohr's principle, the relation between principle quantum number (n) and radius of orbit is **[BHU 2004]**

(a) $r \propto n$ (b) $r \propto n^2$

(c) $r \propto \frac{1}{n}$ (d) $r \propto \frac{1}{n^2}$

77. The ionisation potential of a hydrogen atom is – 13.6 *eV*. What will be the energy of the atom corresponding to $n = 2$ **[Pb. CET 2000]**

(a) –3.4 *eV* (b) –6.8 *eV*

(c) –1.7 *eV* (d) – 2.7 *eV*

78. The energy of electron in hydrogen atom in its grounds state is –13.6 eV. The energy of the level corresponding to the quantum number equal to 5 is **[Pb. CET 2002]**

(a) –0.54 *eV* (b) – 0.85 *eV*

(c) – 0.64 *eV* (d) – 0.40 *eV*

79. The positive charge of an atom is **[AFMC 2002]**

(a) Spread all over the atom

(b) Distributed around the nucleus

(c) Concentrated at the nucleus

(d) All of these

80. A metal surface is exposed to solar radiations **[DPMT 200**

(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 photos of incident light

81. Which of the following transitions have minimum wavelength **[DPMT 2005]**

(a) $n_4 \rightarrow n_1$ (b) $n_2 \rightarrow n_1$

(c) $n_4 \rightarrow n_2$ (d) $n_3 \rightarrow n_1$

Dual nature of electron

1. De broglie equation describes the relationship of wavelength associated with the motion of an electron and its **[CBSE PMT 2004]** **[MP PMT 1986]**

(a) Mass (b) Energy

(c) Momentum (d) Charge

2. The wave nature of an electron was first given by

[CMC Vellore 1991; Pb. PMT 1998; CPMT 2004]

(a) De-Broglie (b) Heisenberg
(c) Mosley (d) Sommerfield

3. Among the following for which one mathematical expression $\lambda = \frac{h}{p}$ stands

(a) De Broglie equation (b) Einstein equation
(c) Uncertainty equation (d) Bohr equation

4. Which one of the following explains light both as a stream of particles and as wave motion

[AIIMS 1983; IIT 1992; UPSEAT 2003]

(a) Diffraction (b) $\lambda = h/p$
(c) Interference (d) Photoelectric effect

5. In which one of the following pairs of experimental observations and phenomenon does the experimental observation correctly account for phenomenon **[AIIMS 1983]**

Experimental observation	**Phenomenon**
(a) X-ray spectra	Charge on the nucleus
(b) α-particle scattering	Quantized electron orbit
(c) Emission spectra	The quantization of energy
(d) The photoelectric effect	The nuclear atom

6. Which of the following expressions gives the de-Broglie relationship **[MP PMT 1996, 2004; MP PET/PMT 1998]**

(a) $h = \frac{\lambda}{mv}$ (b) $\lambda = \frac{h}{mv}$
(c) $\lambda = \frac{m}{hv}$ (d) $\lambda = \frac{v}{mh}$

7. de-Broglie equation is

[MP PMT 1999; CET Pune 1998]

(a) $n\lambda = 2d\sin\theta$ (b) $E = hv$
(c) $E = mc^2$ (d) $\lambda = \frac{h}{mv}$

8. The de-Broglie wavelength of a particle with mass $1gm$ and velocity $100\, m/\sec$ is **[CBSE PMT 1999; EAMCET 1997; AFMC 1999; AIIMS 2000]**

(a) $6.63 \times 10^{-33}\, m$ (b) $6.63 \times 10^{-34}\, m$
(c) $6.63 \times 10^{-35}\, m$ (d) $6.65 \times 10^{-35}\, m$

9. Minimum de-Broglie wavelength is associated with **[RPMT 1999]**

(a) Electron (b) Proton
(c) CO_2 molecule (d) SO_2 molecule

10. The de-Broglie wavelength associated with a material particle is **[JIPMER 2000]**

(a) Directly proportional to its energy
(b) Directly proportional to momentum
(c) Inversely proportional to its energy
(d) Inversely proportional to momentum

11. An electron has kinetic energy $2.8 \times 10^{-23}\, J$. de-Broglie wavelength will be nearly
$(m_e = 9.1 \times 10^{-31}\, kg)$ **[MP PET 2000]**

(a) $9.28 \times 10^{-4}\, m$ (b) $9.28 \times 10^{-7}\, m$
(c) $9.28 \times 10^{-8}\, m$ (d) $9.28 \times 10^{-10}\, m$

12. What will be de-Broglie wavelength of an electron moving with a velocity of $1.2 \times 10^5\, ms^{-1}$ **[MP PET 2000]**

(a) 6.068×10^{-9} (b) 3.133×10^{-37}
(c) 6.626×10^{-9} (d) 6.018×10^{-7}

13. The de-Broglie wavelength associated with a particle of mass $10^{-6}\, kg$ moving with a velocity of $10\, ms^{-1}$, is

[AIIMS 2001]

(a) $6.63 \times 10^{-22}\, m$ (b) $6.63 \times 10^{-29}\, m$
(c) $6.63 \times 10^{-31}\, m$ (d) $6.63 \times 10^{-34}\, m$

14. What is the de-Broglie wavelength associated with the hydrogen electron in its third orbit **[AMU (Engg.)**

(a) $9.96 \times 10^{-10}\, cm$ (b) $9.96 \times 10^{-8}\, cm$
(c) $9.96 \times 10^{4}\, cm$ (d) $9.96 \times 10^{8}\, cm$

15. If the velocity of hydrogen molecule is $5 \times 10^4\, cm\, sec^{-1}$, then its de-Broglie wavelength is **[MP PMT**

(a) 2 Å (b) 4 Å
(c) 8 Å (d) 100 Å

16. A 200*g* golf ball is moving with a speed of 5 *m* per hour. The associated wave length is $(h = 6.625 \times 10^{-34}\, J\text{-}sec)$

[MP PET 2003]

(a) $10^{-10}\, m$ (b) $10^{-20}\, m$
(c) $10^{-30}\, m$ (d) $10^{-40}\, m$

17. A cricket ball of $0.5\, kg$ is moving with a velocity of $100\, m/\sec$. The wavelength associated with its motion is

[DCE 2004]

(a) $1/100\, cm$ (b) $6.6 \times 10^{-34}\, m$
(c) $1.32 \times 10^{-35}\, m$ (d) $6.6 \times 10^{-28}\, m$

18. Dual nature of particles was proposed by **[DCE 2004]**

(a) Heisenberg (b) Lowry
(c) de-Broglie (d) Schrodinger

19. Calculate de-Broglie wavelength of an electron travelling at 1% of the speed of light **[DPMT 2004]**

(a) 2.73×10^{-24} (b) 2.42×10^{-10}
(c) 242.2×10^{10} (d) None of these

20. Which is the correct relationship between wavelength and momentum of particles **[Pb. PMT 2000]**

(a) $\lambda = \frac{h}{P}$ (b) $\pi = \frac{h}{P}$
(c) $P = \frac{h}{\lambda}$ (d) $h = \frac{P}{\lambda}$

21. The de-Broglie equation applies **[MP PMT 2004]**

(a) To electrons only
(b) To neutrons only
(c) To protons only
(d) All the material object in motion

Uncertainty principle and Schrodinger wave equation

1. The uncertainty principle was enunciated by **[NCERT 1975; Bihar MEE 1997]**

(a) Einstein (b) Heisenberg

(c) Rutherford (d) Pauli

2. According to heisenberg uncertainty principle **[AMU 1990; BCECE 2005]**

(a) $E = mc^2$ (b) $\Delta x \times \Delta p \geq \frac{h}{4\pi}$

(c) $\lambda = \frac{h}{p}$ (d) $\Delta x \times \Delta p = \frac{h}{6\pi}$

3. "The position and velocity of a small particle like electron cannot be simultaneously determined." This statement is **[NCERT 1979; BHU 1981, 87]**

(a) Heisenberg uncertainty principle

(b) Principle of de Broglie's wave nature of electron

(c) Pauli's exclusion principle

(d) Aufbau's principle

4. In Heisenberg's uncertainty equation $\Delta x \times \Delta p \geq \frac{h}{4\pi}$; Δp stands for

(a) Uncertainty in energy

(b) Uncertainty in velocity

(c) Uncertainty in momentum

(d) Uncertainty in mass

5. Which one is not the correct relation in the following

(a) $h = \frac{E}{v}$ (b) $E = mc^2$

(c) $\Delta x \times \Delta p = \frac{h}{4\pi}$ (d) $\lambda = \frac{h}{mv}$

6. The maximum probability of finding an electron in the d_{xy} orbital is **[MP PET 1996]**

(a) Along the *x*-axis

(b) Along the *y*-axis

(c) At an angle of 45^o from the x and *y*-axes

(d) At an angle of 90^o from the x and *y*-axes

7. Simultaneous determination of exact position and momentum of an electron is **[BHU 1979]**

(a) Possible

(b) Impossible

(c) Sometimes possible sometimes impossible

(d) None of the above

8. If uncertainty in the position of an electron is zero, the uncertainty in its momentum would be **[CPMT 1988]**

(a) Zero (b) $< \frac{h}{2\lambda}$

(c) $> \frac{h}{2\lambda}$ (d) Infinite

9. The possibility of finding an electron in an orbital was conceived by **[MP PMT 1994]**

(a) Rutherford (b) Bohr

(c) Heisenberg (d) Schrodinger

10. Uncertainty principle gave the concept of

(a) Probability

(b) An orbital

(c) Physical meaning of Ψ the Ψ^2

(d) All the above

11. The uncertainty principle and the concept of wave nature of matter was proposed by ...... and ...... respectively **[MP PET 1997]**

(a) Heisenberg, de Broglie (b)de-Broglie, Heisenberg

(c) Heisenberg, Planck (d) Planck, Heisenberg

12. The uncertainty in momentum of an electron is $1 \times 10^{-5} kg - m/s$. The uncertainty in its position will be ($h = 6.62 \times 10^{-34} kg - m^2/s$) **[AFMC 1998; CBSE PMT 1999; JIPMER 2002]**

(a) $1.05 \times 10^{-28} m$ (b) $1.05 \times 10^{-26} m$

(c) $5.27 \times 10^{-30} m$ (d) $5.25 \times 10^{-28} m$

13. The uncertainty in the position of a moving bullet of mass 10 *gm* is $10^{-5} m$. Calculate the uncertainty in its velocity **[DCE 1999]**

(a) $5.2 \times 10^{-28} m/sec$ (b) $3.0 \times 10^{-28} m/sec$

(c) $5.2 \times 10^{-22} m/sec$ (d) $3 \times 10^{-22} m/sec$

14. The equation $\Delta x . \Delta p \geq \frac{h}{4\pi}$ shows **[MP PET 2000]**

(a) de-Broglie relation

(b) Heisenberg's uncertainty principle

(c) Aufbau principle

(d) Hund's rule

15. Which quantum number is not related with Schrodinger equation **[RPMT 2002]**

(a) Principal (b) Azimuthal

(c) Magnetic (d) Spin

16. Uncertainty in position of a 0.25 *g* particle is 10^{-5}. Uncertainty of velocity is $(h = 6.6 \times 10^{-34}\ Js)$ **[AIEEE 20**

(a) 1.2×10^{34} (b) 2.1×10^{-29}

(c) 1.6×10^{-20} (d) 1.7×10^{-9}

17. The uncertainty in momentum of an electron is $1 \times 10^{-5} kg\, m/s$. The uncertainity in its position will be $(h = 6.63 \times 10^{-34}\ Js)$ **[Pb. CET 2000]**

(a) $5.28 \times 10^{-30} m$ (b) $5.25 \times 10^{-28} m$

(c) $1.05 \times 10^{-26} m$ (d) $2.715 \times 10^{-30} m$

18. According to Heisenberg's uncertainty principle, the product of uncertainties in position and velocities for an electron of mass $9.1 \times 10^{-31} kg$ is

(a) $2.8 \times 10^{-3} m^2 s^{-1}$ (b) $3.8 \times 10^{-5} m^2 s^{-1}$

(c) $5.8 \times 10^{-5} m^2 s^{-1}$ (d) $6.8 \times 10^{-6} m^2 s^{-1}$

19. For an electron if the uncertainty in velocity is Δv, the uncertainty in its position (Δx) is given by **[DPMT 2005]**

(a) $\frac{hm}{4\pi\Delta v}$ (b) $\frac{4\pi}{hm\Delta v}$

(c) $\frac{h}{4\pi m\Delta v}$ (d) $\frac{4\pi m}{h . \Delta v}$

20. Orbital is **[DPMT 2005]**

(a) Circular path around the nucleus in which the electron revolves

(b) Space around the nucleus where the probability of finding the electron is maximum

(c) Amplitude of electrons wave

(d) None of these

Quantum number, Electronic configuration and Shape of orbitals

1. Be's 4th electron will have four quantum numbers **[MNR 1985]**

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

2. The quantum number which specifies the location of an electron as well as energy is **[DPMT 1983]**

(a) Principal quantum number

(b) Azimuthal quantum number

(c) Spin quantum number

(d) Magnetic quantum number

3. The shape of an orbital is given by the quantum number

[NCERT 1984; MP PMT 1996]

(a) n (b) l

(c) m (d) s

4. In a given atom no two electrons can have the same values for all the four quantum numbers. This is called

[BHU 1979; AMU 1983; EAMCET 1980, 83; MADT Bihar 1980; CPMT 1986, 90, 92; NCERT 1978, 84; RPMT 1997; CBSE PMT 1991; MP PET 1986, 99]

(a) Hund's rule

(b) Aufbau's principle

(c) Uncertainty principle

(d) Pauli's exclusion principle

5. Nitrogen has the electronic configuration $1s^2, 2s^2 2p_x^1 2p_y^1 2p_z^1$ and not $1s^2, 2s^2 2p_x^2 2p_y^1 2p_z^0$ which is determined by

[DPMT 1981, 83, 89; MP PMT/PET 1988; EAMCET 1988]

(a) Aufbau's principle (b) Pauli's exclusion principle

(c) Hund's rule (d) Uncertainty principle

6. Which one of the following configuration represents a noble gas

DPMT 1984]

(a) $1s^2, 2s^2 2p^6, 3s^2$ (b) $1s^2, 2s^2 2p^6, 3s^1$

(c) $1s^2, 2s^2 2p^6$ (d) $1s^2, 2s^2 sp^6, 3s^2 3p^6, 4s^2$

7. The electronic configuration of silver atom in ground state is

[CPMT 1984, 93]

(a) $[Kr]3d^{10} 4s^1$ (b) $[Xe]4f^{14} 5d^{10} 6s^1$

(c) $[Kr]4d^{10} 5s^1$ (d) $[Kr]4d^9 5s^2$

8. Principal, azimuthal and magnetic quantum numbers are respectively related to **[CPMT 1988; AIIMS 199**

(a) Size, shape and orientation

(b) Shape, size and orientation

(c) Size, orientation and shape

(d) None of the above

9. Correct set of four quantum numbers for valence electron of rubidium (Z = 37) is

[IIT 1984; JIPMER 1999; UPSEAT 2003]

(a) $5, 0, 0, +\frac{1}{2}$ (b) $5, 1, 0, +\frac{1}{2}$

(c) $5, 1, 1, +\frac{1}{2}$ (d) $6, 0, 0, +\frac{1}{2}$

10. The correct ground state electronic configuration of chromium atom is **[IIT 1989, 94; MP PMT 1993; EAMCET 1997; ISM Dhanbad 1994; AFMC 1997; Bihar MEE 1996; MP PET 1995, 97; CPMT 1999; Kerala PMT 2003]**

(a) $[Ar]3d^5 4s^1$ (b) $[Ar]3d^4 4s^2$

(c) $[AR]3d^6 4s^0$ (d) $[Ar]4d^5 4s^1$

11. $2p$ orbitals have **[NCERT 1981; MP PMT 1993, 97]**

(a) $n = 1, l = 2$ (b) $n = 1, l = 0$

(c) $n = 2, l = 1$ (d) $n = 2, l = 0$

12. Electronic configuration of H^- is **[CPMT 1985]**

(a) $1s^0$ (b) $1s^1$

(c) $1s^2$ (d) $1s^1 2s^1$

13. The quantum numbers for the outermost electron of an element are given below as $n = 2, l = 0, m = 0, s = +\frac{1}{2}$. The atoms is

(a) Lithium (b) Beryllium

(c) Hydrogen (d) Boron

14. Principal quantum number of an atom represents

[EAMCET 1979; IIT 1983; MNR 1990; UPSEAT 2000, 02]

(a) Size of the orbital
(b) Spin angular momentum
(c) Orbital angular momentum
(d) Space orientation of the orbital

15. An element has the electronic configuration $1s^2, 2s^2 2p^6, 3s^2 3p^2$. Its valency electrons are **[NCERT 1973]**

(a) 6 (b) 2
(c) 3 (d) 4

16. The magnetic quantum number specifies **[MNR 1986; BHU 1982; CPMT 1989, 94; MP PET 1999; AFMC 1999; AMU (Engg.) 1999]**

(a) Size of orbitals (b) Shape of orbitals
(c) Orientation of orbitals (d) Nuclear stability

17. Which of the following sets of quantum numbers represent an impossible arrangement **[IIT 1986; MP PET 1995]**

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

18. If $n = 3$, then the value of 'l' which is incorrect **[CPMT 1994]**

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

19. Which orbital is dumb-bell shaped **[MP PMT 1986; MP PET/PMT 1998]**

(a) s-orbital (b) p-orbital
(c) d-orbital (d) f-orbital

20. The total number of unpaired electrons in d-orbitals of atoms of element of atomic number 29 is **[CPMT 1983]**

(a) 10 (b) 1
(c) 0 (d) 5

21. The shape of $2p$ orbital is **[CPMT 1983; NCERT 1979]**

(a) Spherical (b) Ellipsoidal
(c) Dumb-bell (d) Pyramidal

22. The magnetic quantum number for an electron when the value of principal quantum number is 2 can have **[CPMT 1984]**

(a) 3 values (b) 2 values
(c) 9 values (d) 6 values

23. Which one is the correct outer configuration of chromium **[AIIMS 1980, 91; BHU 1995]**

(a)	↑	↑	↑	↑			↑↓
	↑↓	↑↓	↑				
	↑	↑	↑	↑	↑		↑
	↑↓	↑↓	↑	↑	↑		↑

(b)
(c)
(d)

24. The following has zero valency **[DPMT 1991]**

(a) Sodium (b) Beryllium
(c) Aluminium (d) Krypton

25. The number of electrons in the valence shell of calcium is **[IIT 1975]**

(a) 6 (b) 8
(c) 2 (d) 4

26. The valence electron in the carbon atom are **[MNR 1982]**

(a) 0 (b) 2
(c) 4 (d) 6

27. For the dumb-bell shaped orbital, the value of l is **[CPMT 1987, 2003]**

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

28. Chromium has the electronic configuration $4s^1 3d^5$ rather than $4s^2 3d^4$ because

(a) $4s$ and $3d$ have the same energy
(b) $4s$ has a higher energy than $3d$
(c) $4s^1$ is more stable than $4s^2$
(d) $4s^1 3d^5$ half-filled is more stable than $4s^2 3d^4$

29. The electronic configuration of calcium ion (Ca^{2+}) is **[CMC Vellore 1991]**

(a) $1s^2, 2s^2 2p^6, 3s^2 3p^6, 4s^2$
(b) $1s^2, 2s^2 sp^6, 3s^2 3p^6, 4s^1$
(c) $1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^2$
(d) $1s^2, 2s^2 sp^6, 3s^2 3p^6 3d^5$
(e) $1s^2, 2s^2 2p^6, 3s^2 3p^6, 4s^0$

30. The structure of external most shell of inert gases is **[JIPMER 1991]**

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

31. The two electrons in K sub-shell will differ in **[MNR 1988; UPSEAT 1999, 2000; Kerala PMT 2003]**

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

32. A completely filled d-orbital (d^{10}) **[MNR 1987]**

(a) Spherically symmetrical
(b) Has octahedral symmetry
(c) Has tetrahedral symmetry
(d) Depends on the atom

33. If magnetic quantum number of a given atom represented by –3, then what will be its principal quantum number

[BHU 2005]

(a) 2 (b) 3

(c) 4 (d) 5

34. The total number of orbitals in an energy level designated by principal quantum number n is equal to

[AIIMS 1997; J&K CET 2005]

(a) $2n$ (b) $2n^2$

(c) n (d) n^2

35. The number of orbitals in the fourth principal quantum number will be

(a) 4 (b) 8

(c) 12 (d) 16

36. Which set of quantum numbers are not possible from the following

(a) $n=3, l=2, m=0, s=-\frac{1}{2}$

(b) $n=3, l=2, m=-2, s=-\frac{1}{2}$

(c) $n=3, l=3, m=-3, s=-\frac{1}{2}$

(d) $n=3, l=0, m=0, s=-\frac{1}{2}$

37. The four quantum number for the valence shell electron or last electron of sodium (Z = 11) is **[MP PMT 1999]**

(a) $n=2, l=1, m=-1, s=-\frac{1}{2}$

(b) $n=3, l=0, m=0, s=+\frac{1}{2}$

(c) $n=3, l=2, m=-2, s=-\frac{1}{2}$

(d) $n=3, l=2, m=2, s=+\frac{1}{2}$

38. The explanation for the presence of three unpaired electrons in the nitrogen atom can be given by

[NCERT 1979; RPMT 1999; DCE 1999, 2002; CPMT 2001; MP PMT 2002; Pb. PMT / CET 2002]

(a) Pauli's exclusion principle

(b) Hund's rule

(c) Aufbau's principle

(d) Uncertainty principle

39. The maximum energy is present in any electron at

(a) Nucleus

(b) Ground state

(c) First excited state

(d) Infinite distance from the nucleus

40. The electron density between $1s$ and $2s$ orbital is

(a) High (b) Low

(c) Zero (d) None of these

41. For ns orbital, the magnetic quantum number has value

(a) 2 (b) 4

(c) – 1 (d) 0

42. The maximum number of electrons that can be accommodated in the M^{th} shell is

(a) 2 (b) 8

(c) 18 (d) 32

43. For a given value of quantum number l, the number of allowed values of m is given by

(a) $l+2$ (b) $2l+2$

(c) $2l+1$ (d) $l+1$

44. The number of radial nodes of $3s$ and $2p$ orbitals are respectively. **[IIT-JEE 2005]**

(a) 2, 0 (b) 0, 2

(c) 1, 2 (d) 2, 1

45. Which of the sub-shell is circular

(a) $4s$ (b) $4f$

(c) $4p$ (d) $4d$

46. Which electronic configuration for oxygen is correct according to Hund's rule of multiplicity

(a) $1s^2, 2s^2 2p_x^2 2p_y^1 2p_z^1$ (b) $1s^2, 2s^2 2p_x^2 2p_y^2 2p_z^0$

(c) $1s^2, 2s^2 2p_x^3 2p_y^1 2p_z^0$ (d) None of these

47. If value of azimuthal quantum number l is 2, then total possible values of magnetic quantum number will be

(a) 7 (b) 5

(c) 3 (d) 2

48. The type of orbitals present in Fe is

(a) s (b) s and p

(c) s, p and d (d) s, p, d and f

49. The shape of d_{xy} orbital will be

(a) Circular (b) Dumb-bell

(c) Double dumb-bell (d) Trigonal

50. In any atom which sub-shell will have the highest energy in the following

(a) $3p$ (b) $3d$

(c) $4s$ (d) $3s$

51. Which electronic configuration is not observing the $(n+l)$ rule

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

(b) $1s^2, 2s^2 sp^6, 3s^2 3p^6 3d^7, 4s^2$

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

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

52. The four quantum numbers of the outermost orbital of K (atomic no. =19) are **[MP PET 1993, 94]**

(a) $n=2, l=0, m=0, s=+\frac{1}{2}$

(b) $n=4, l=0, m=0, s=+\frac{1}{2}$

(c) $n=3, l=1, m=1, s=+\frac{1}{2}$

(d) $n=4, l=2, m=-1, s=+\frac{1}{2}$

53. The angular momentum of an electron depends on **[BHU 1978; NCERT 1981]**

(a) Principal quantum number
(b) Azimuthal quantum number
(c) Magnetic quantum number
(d) All of these

54. The electronic configuration of copper $({}_{29}Cu)$ is **[DPMT 1983; BHU 1980; AFMC 1981; CBSE PMT 1991; MP PMT 1995]**

(a) $1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^9, 4s^2$
(b) $1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^{10}, 4s^1$
(c) $1s^2.2s^2 2p^6, 3s^2 3p^6, 4s^2 4p^6$
(d) $1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^{10}$

55. The number of orbitals in $2p$ sub-shell is **[NCERT 1973; MP PMT 1996]**

(a) 6 (b) 2
(c) 3 (d) 4

56. The number of orbitals in d sub-shell is **[MNR 1981]**

(a) 1 (b) 3
(c) 5 (d) 7

57. A sub-shell $l=2$ can take how many electrons **[NCERT 1973, 78]**

(a) 3 (b) 10
(c) 5 (d) 6

58. Pauli's exclusion principle states that **[MNR 1983; AMU 1984]**

(a) Two electrons in the same atom can have the same energy
(b) Two electrons in the same atom cannot have the same spin
(c) The electrons tend to occupy different orbitals as far as possible
(d) Electrons tend to occupy lower energy orbitals preferentially
(e) None of the above

59. For d electrons, the azimuthal quantum number is **[MNR 1983; CPMT 1984]**

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

60. For p-orbital, the magnetic quantum number has value

(a) 2 (b) 4, – 4
(c) – 1, 0, +1 (d) 0

61. For $n=3$ energy level, the number of possible orbitals (all kinds) are **[BHU 1981; CPMT 1985; MP PMT 1995]**

(a) 1 (b) 3
(c) 4 (d) 9

62. Which of the following ions is not having the configuration of neon

(a) F^- (b) Mg^{+2}
(c) Na^+ (d) Cl^-

63. Elements upto atomic number 103 have been synthesized and studied. If a newly discovered element is found to have an atomic number 106, its electronic configuration will be **[AIIMS 1980]**

(a) $[Rn]5f^{14}, 6d^4, 7s^2$ (b) $[Rn]5f^{14}, 6d^1, 7s^2 7p^3$
(c) $[Rn]5f^{14}, 6d^6, 7s^0$ (d) $[Rn]5f^{14}, 6d^5, 7s^1$

64. Ions which have the same electronic configuration are those of

(a) Lithium and sodium (b) Sodium and potassium
(c) Potassium and calcium (d) Oxygen and chlorine

65. When the azimuthal quantum number has a value of $l=0$, the shape of the orbital is **[MP PET 1995]**

(a) Rectangular (b) Spherical
(c) Dumbbell (d) Unsymmetrical

66. The magnetic quantum number for valency electrons of sodium is **[CPMT 1988; MH CET 1999]**

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

67. The electronic configuration of an element with atomic number 7 *i.e.* nitrogen atom is **[CPMT 1982, 84, 87]**

(a) $1s^2, 2s^1, 2p_x^3$ (b) $1s^2, 2s^2 2p_x^2 2p_y^1$
(c) $1s^2, 2s^2 2p_x^1 2p_y^1 2p_z^1$ (d) $1s^2, 2s^2 2p_x^1 2p_y^2$

68. In a multi-electron atom, which of the following orbitals described by the three quantum members will have the same energy in the absence of magnetic and electric fields **[AIEEE 2005]**

(1) $n=1, l=0, m=0$ (2) $n=2, l=0, m=0$
(3) $n=2, l=1, m=1$ (4) $n=3, l=2, m=0$
(5) $n=3, l=2, m=0$

(a) (1) and (2) (b) (2) and (3)
(c) (3) and (4) (d) (4) and (5)

69. Which of the following represents the electronic configuration of an element with atomic number 17 **[AMU 1982]**

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

70. The shape of s-orbital is **[NCERT 1978I]**

(a) Pyramidal (b) Spherical

(c) Tetrahedral (d) Dumb-bell shaped

71. When $3d$ orbital is complete, the new electron will enter the **[EAMCET 1980; MP PMT 1995]**

(a) $4p$ -orbital (b) $4f$ -orbital

(c) $4s$ -orbital (d) $4d$ -orbital

72. In a potassium atom, electronic energy levels are in the following order **[EAMCET 1979; DPMT 1991]**

(a) $4s > 3d$ (b) $4s > 4p$

(c) $4s < 3d$ (d) $4s < 3p$

73. Fe (atomic number = 26) atom has the electronic arrangement **[NCERT 1974; MNR 1980]**

(a) 2, 8, 8, 8 (b) 2, 8, 16

(c) 2, 8,14, 2 (d) 2, 8, 12, 4

74. Cu^{2+} will have the following electronic configuration **[MP PMT 1985]**

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

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

(c) $1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^9$

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

75. Which one is the electronic configuration of Fe^{+2} **[MADT Bihar 1982; AIIMS 1989]**

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

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

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

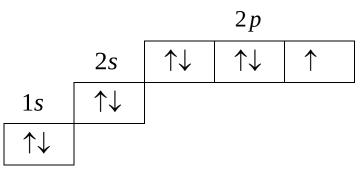
(d) None of these

76. How many electrons can be fit into the orbitals that comprise the 3^{rd} quantum shell $n = 3$ **[MP PMT 1986, 87; Orissa JEE 1997]**

(a) 2 (b) 8

(c) 18 (d) 32

77. Which element is represented by the following electronic configuration **[MP PMT 1987]**



(a) Nitrogen (b) Oxygen

(c) Fluorine (d) Neon

78. If the value of azimuthal quantum number is 3, the possible values of magnetic quantum number would be **[MP PMT 1987; RPMT 1999; AFMC 2002; KCET 2002]**

(a) 0, 1, 2, 3 (b) 0, – 1, – 2, – 3

(c) 0, ± 1, ± 2, $\pm$ 3 (d) ± 1, ± 2, ± 3

79. Krypton $({}_{36}Kr)$ has the electronic configuration $({}_{18}Ar)$ $4s^2, 3d^{10}, 4p^6$. The 37^{th} electron will go into which one of the following sub-levels **[CBSE PMT 1989; CPMT 1989; EAMCET 1991]**

(a) $4f$ (b) $4d$

(c) $3p$ (d) $5s$

80. If an electron has spin quantum number of $+\frac{1}{2}$ and a magnetic quantum number of -1, it cannot be presented in an **[CBSE PMT 1989; UPSEAT 2001]**

(a) d -orbital (b) f -orbital

(c) p -orbital (d) s -orbital

81. The azimuthal quantum number is related to **[BHU 1987, 95]**

(a) Size (b) Shape

(c) Orientation (d) Spin

82. The total number of electrons that can be accommodated in all the orbitals having principal quantum number 2 and azimuthal quantum number 1 is **[CPMT 1971, 89, 91]**

(a) 2 (b) 4

(c) 6 (d) 8

83. Electronic configuration of C is **[CPMT 1975]**

(a) $1s^2, 2s^2 2p^2$ (b) $1s^2, 2s^2 2p^3$

(c) $1s^2, 2s^2$ (d) $1s^2, 2s^2 2p^6$

84. There is no difference between a $2p$ and a $3p$ orbital regarding **[BHU 1981]**

(a) Shape (b) Size

(c) Energy (d) Value of n

85. The electronic configuration of chromium is **[MP PMT 1993; MP PET 1995; BHU 2001; BCECE 2005]**

(a) $[Ne]3s^2 3p^6 3d^4, 4s^2$ (b) $[Ne]3s^2 3p^6 3d^5, 4s^1$

(c) $[Ne]3s^2 3p^6, 4s^2 4p^4$ (d) $[Ne]3s^2 3p^6 3d^1, 4s^2 4p^3$

86. The shape of p -orbital is **[MP PMT 1993]**

(a) Elliptical (b) Spherical

(c) Dumb-bell (d) Complex geometrical

87. The electronic configuration (outermost) of Mn^{+2} ion (atomic number of $Mn = 25$) in its ground state is **[MP PET 1993]**

(a) $3d^5, 4s^0$ (b) $3d^4, 4s^1$

(c) $3d^3, 4s^2$ (d) $3d^2, 4s^2 4p^2$

88. The principal quantum number represents **[CPMT 1991]**

(a) Shape of an orbital

(b) Distance of electron from nucleus

(c) Number of electrons in an orbit

(d) Number of orbitals in an orbit

89. When the azimuthal quantum number has a value of $l = 1$, the shape of the orbital is **[MP PET 1993]**

(a) Unsymmetrical (b) Spherically symmetrical

(c) Dumb-bell (d) Complicated

90. How many electrons can be accommodated in a sub-shell for which $n = 3, l = 1$ **[CBSE PMT 1990]**

(a) 8 (b) 6

(c) 18 (d) 32

91. For azimuthal quantum number $l = 3$, the maximum number of electrons will be

EAMCET 1991; RPMT 2002; CBSE PMT 2002]

(a) 2 (b) 6

(c) 0 (d) 14

92. An ion has 18 electrons in the outermost shell, it is

[CBSE PMT 1990]

(a) Cu^+ (b) Th^{4+}

(c) Cs^+ (d) K^+

93. 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) $3d, 4p, 4s, 4d, 5s$

94. The quantum number which may be designated by s, p, d and f instead of number is **BHU 1980]**

(a) n (b) l

(c) m_l (d) m_s

95. Which of the following represents the correct sets of the four quantum numbers of a $4d$ electron

[MNR 1992; UPSEAT 2001; J&K CET 2005]

(a) $4, 3, 2, \frac{1}{2}$ (b) 4, 2, 1, 0

(c) $4, 3, -2, +\frac{1}{2}$ (d) $4, 2, 1, -\frac{1}{2}$

96. Which of the following statements is not correct for an electron that has the quantum numbers $n = 4$ and $m = 2$

[MNR 1993]

(a) The electron may have the quantum number $s = +\frac{1}{2}$

(b) The electron may have the quantum number $l = 2$

(c) The electron may have the quantum number $l = 3$

(d) The electron may have the quantum number $l = 0, 1, 2, 3$

97. The set of quantum numbers not applicable for an electron in an atom is **[MNR 1994]**

(a) $n = 1, l = 1, m_l = 1, m_s = +1/2$

(b) $n = 1, l = 0, m_l = 0, m_s = +1/2$

(c) $n = 1, l = 0, m_l = 0, m_s = -1/2$

(d) $n = 2, l = 0, m_l = 0, m_s = +1/2$

98. Correct configuration of Fe^{+3} [26] is

[CPMT 1994; BHU 1995; KCET 1992]

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

(b) $1s^2, 2s^2 sp^6, 3s^2 3p^6 3d^3, 4s^2$

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

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

99. Azimuthal quantum number for last electron of Na atom is **[CBSE PMT 1991;**

[BHU 1995]

(a) 1 (b) 2

(c) 3 (d) 0

100. A $3p$ orbital has **[IIT 1995]**

(a) Two spherical nodes

(b) Two non-spherical nodes

(c) One spherical and one non-spherical nodes

(d) One spherical and two non-spherical nodes

101. All electrons on the $4p$ sub-shell must be characterized by the quantum number(s) **[MP PET 1996]** **[CBSE PMT 1991]**

(a) $n = 4, m = 0, s = \pm\frac{1}{2}$ (b) $l = 1$

(c) $l = 0, s = \pm\frac{1}{2}$ (d) $s = \pm\frac{1}{2}$

102. The electronic configuration of the element of atomic number 27 is

(a) $1s^2, 2s^2 2p^6, 3s^2 3p^6, 4s\,(\uparrow\downarrow)\,4p\,(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)\,5s\,(\uparrow)$

(b) $1s^2, 2s^2 2p^6, 3s^2 3p^6 3d\,(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow), 4s\,(\uparrow\downarrow)\,4p\,(\uparrow)$

(c) $1s^2, 2s^2 2p^6, 3s^2 3p^6, 3d\,(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow), 4s\,(\uparrow)$

(d) $1s^2, 2s^2 2p^6, 3s^2 3p^6, 3d\,(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow)(\uparrow)(\uparrow)\,4s\,(\uparrow\downarrow)$

103. When the value of the principal quantum number n is 3, the permitted values of the azimuthal quantum numbers l and the magnetic quantum numbers m, are

	l	m
	0	0
(a)	1	$+1, 0, -1$
	2	$+2, +1, 0, -1, -2$
	1	1
(b)	2	$+2, 1, -1$
	3	$+3, +2, 1, -2, -3$
	0	0
(c)	1	1, 2, 3
	2	$+3, +2, 1, -2, -3$
	1	0, 1
(d)	2	0, 1, 2
	3	0, 1, 2, 3

104. The number of possible spatial orientations of an electron in an atom is given by its

(a) Spin quantum number

(b) Spin angular momentum

(c) Magnetic quantum number

(d) Orbital angular momentum

105. Which of the following sets of orbitals may degenerate

(a) $2s, 2p_x, 2p_y$ (b) $3s, 3p_x, 3d_{xy}$

(c) $1s, 2s, 3s$ (d) $2p_x, 2p_y, 2p_z$

106. The set of quantum numbers $n = 3, l = 0, m = 0, s = -1/2$ belongs to the element

(a) Mg (b) Na

(c) Ne (d) F

107. An electron has principal quantum number 3. The number of its (i) sub-shells and (ii) orbitals would be respectively **[MP PET 1997]**

(a) 3 and 5 (b) 3 and 7

(c) 3 and 9 (d) 2 and 5

108. What is the electronic configuration of $Cu^{2+}(Z = 29)$ of least position **[MP PET/PMT 1998; MP PET 2001]**

(a) $[Ar]4s^1 3d^8$ (b) $[Ar]4s^2 3d^{10} 4p^1$

(c) $[Ar]4s^1 3d^{10}$ (d) $[Ar]3d^9$

109. The correct electronic configuration of $Ti(Z = 22)$ atom is **[MP PMT 1999]**

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

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

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

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

110. Which of the following configuration is correct for iron **[CBSE PMT 1999]**

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

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

(c) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7$

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

111. Which of the following set of quantum numbers belong to highest energy **[CPMT 1999]**

(a) $n = 4, l = 0, m = 0, s = +\frac{1}{2}$

(b) $n = 3, l = 0, m = 0, s = +\frac{1}{2}$

(c) $n = 3, l = 1, m = 1, s = +\frac{1}{2}$

(d) $n = 3, l = 2, m = 1, s = +\frac{1}{2}$

112. Which quantum number will determine the shape of the subshell **[CPMT 1999; Pb. PMT 1998]**

(a) Principal quantum number

(b) Azimuthal quantum number

(c) Magnetic quantum number

(d) Spin quantum number

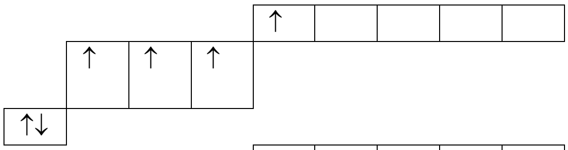
113. For the $n = 2$ energy level, how many orbitals of all kinds are possible **[Bihar CEE 1995]**

(a) 2 (b) 3

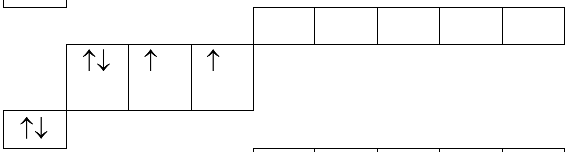
(c) 4 (d) 5

114. Which one is in the ground state **[DPMT 1996]**

(a)



(b)



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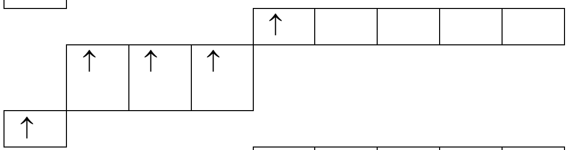
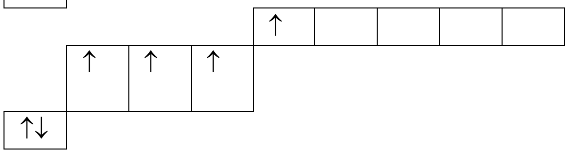


(d)



115. When the principal quantum number $(n = 3)$, the possible values of azimuthal quantum number (l) is **[Bihar MEE 1996; KCET 2000]**

(a) 0, 1, 2, 3 (b) 0, 1, 2

(c) – 2, – 1, 0, 1, 2 (d) 1, 2, 3

(e) 0, 1

116. Which statement is not correct for $n = 5$, $m = 3$ **[CPMT 1996]**

(a) $l = 4$ (b) $l = 0, 1, 3; s = +\frac{1}{2}$

(c) $l = 3$ (d) All are correct

117. $1s^2 2s^2 2p^6 3s^1$ shows configuration of **[CPMT 1996]**

(a) Al^{3+} in ground state (b) Ne in excited state

(c) Mg^+ in excited state (d) None of these

118. Five valence electrons of p^{15} are labelled as

AB		X	Y	Z
3s			3p	

If the spin quantum of B and Z is $+\frac{1}{2}$, the group of electrons with three of the quantum number same are **[JIPMER 1997]**

(a) AB, XYZ, BY (b) AB

(c) XYZ, AZ (d) AB, XYZ

119. Electronic configuration of Sc^{21} is **[BHU 1997]**

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

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

(c) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^0 3d^3$

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

120. If $n+l=6$, then total possible number of subshells would be **[RPMT 1997]**

(a) 3 (b) 4

(c) 2 (d) 5

121. An electron having the quantum numbers $n=4, l=3, m=0$, $s=-\frac{1}{2}$ would be in the orbital **[Orissa JEE 1997]**

(a) $3s$ (b) $3p$

(c) $4d$ (d) $4f$

122. Which of the following sets of quantum numbers is not allowed **[Orissa JEE 1997]**

(a) $n=1, l=0, m=0, s=+\frac{1}{2}$

(b) $n=1, l=1, m=0, s=-\frac{1}{2}$

(c) $n=2, l=1, m=1, s=+\frac{1}{2}$

(d) $n=2, l=1, m=0, s=-\frac{1}{2}$

123. For which of the following sets of four quantum numbers, an electron will have the highest energy **[CBSE PMT 1994]**

	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

124. The electronic configuration of gadolinium (atomic no. 64) is **[CBSE PMT 1997]**

(a) $[Xe]4s^8 5d^9 6s^2$ (b) $[Xe]4s^7 5d^1 6s^2$

(c) $[Xe]4s^3 5d^5 6s^2$ (d) $[Xe]4f^6 5d^2 6s^2$

125. An e^- has magnetic quantum number as -3, what is its principal quantum number **[BHU 1998]**

(a) 1 (b) 2

(c) 3 (d) 4

126. The number of quantum numbers required to describe an electron in an atom completely is **[CET Pune 1998]**

(a) 1 (b) 2

(c) 3 (d) 4

127. The electronic configuration $1s^2 2s^2 2p_x^1 2p_y^1 2p_z^1$

[AFMC 1997; Pb. PMT 1999; CBSE PMT 2001; AIIMS 2001]

(a) Oxygen (b) Nitrogen

(c) Hydrogen (d) Fluorine

128. Which one of the following set of quantum numbers is not possible for $4p$ electron **[EAMCET 1998]**

(a) $n=4, l=1, m=-1, s=+\frac{1}{2}$

(b) $n=4, l=1, m=0, s=+\frac{1}{2}$

(c) $n=4, l=1, m=2, s=+\frac{1}{2}$

(d) $n=4, l=1, m=-1, s=+\frac{1}{2}$

129. Which of the following orbital is not possible **[RPMT 1999]**

(a) $3f$ (b) $4f$

(c) $5f$ (d) $6f$

130. Which set of quantum numbers for an electron of an atom is not possible **[RPMT; DCE 1999]**

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

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

(c) $n=1, l=0, m=0, s=-1/2$

(d) $n=2, l=1, m=-1, s=+1/2$

131. Electronic configuration of ferric ion is **[RPET 2000]**

(a) $[Ar]3d^5$ (b) $[Ar]3d^7$

(c) $[Ar]3d^3$ (d) $[Ar]3d^8$

132. What is the maximum number of electrons which can be accommodated in an atom in which the highest principal quantum number value is 4 **[MP PMT 200**

(a) 10 (b) 18

(c) 32 (d) 54

133. Which of the following electronic configurations is not possible **[CPMT 2000]**

(a) $1s^2\ 2s^2$ (b) $1s^2 2s^2 2p^6$

(c) $3d^{10} 4s^2 4p^2$ (d) $1s^2 2s^2\ 2p^2 3s^1$

134. The electronic configuration of an element is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$. This represents its **[IIT Screening 2000]**

(a) Excited state (b) Ground state

(c) Cationic form (d) Anionic form

135. Which of the following set of quantum numbers is possible **[AIIMS 2001]**

(a) $n=3;\ l=2;\ m=2$ and $s=+\frac{1}{2}$

(b) $n=3;\ l=4;\ m=0$ and $s=-\frac{1}{2}$

(c) $n=4;\ l=0;\ m=2$ and $s=+\frac{1}{2}$

(d) $n=4;\ l=4;\ m=3$ and $s=+\frac{1}{2}$

136. Which of the following set of quantum number is not valid **[AIIMS 2001]**

(a) $n=1,\ l=2$ (b) $3=2,\ m=1$

(c) $m=3,\ l=0$ (d) $3=4,\ l=2$

137. Which one pair of atoms or ions will have same configuration **[JIPMER 2001]**

(a) F^+ and Ne (b) Li^+ and He^-

(c) Cl^- and Ar (d) Na and K

138. Which of the following sets of quantum number is not possible **[MP PET 2001]**

(a) $n=3; l=+2; m=0; s=+\frac{1}{2}$

(b) $n=3; l=0; m=0; s=-\frac{1}{2}$

(c) $n=3; l=0; m=-1; s=+\frac{1}{2}$

(d) $n=3; l=1; m=0; s=-\frac{1}{2}$

139. Which of the following set of quantum numbers is correct for the 19^{th} electron of chromium **[DCE 2001]**

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

140. When the value of azimuthal quantum number is 3, magnetic quantum number can have values **[DPMT 2001]**

(a) + 1, 0, – 1

(b) + 2, + 1, 0, – 1, – 2

(c) – 3, – 2, – 1, – 0, + 1, + 2, + 3

(d) + 1, – 1

141. The quantum numbers $n=2, l=1$ represent **[AFMC 2002]**

(a) 1*s* orbital (b) 2*s* orbital

(c) 2*p* orbital (d) 3*d* orbital

142. The magnetic quantum number of valence electron of sodium (Na) is **[RPMT 2002]**

(a) 3 (b) 2

(c) 1 (d) 0

143. Azimuthal quantum number defines **[AIIMS 2002]**

(a) e/m ratio of electron

(b) Spin of electron

(c) Angular momentum of electron

(d) Magnetic momentum of electron

144. Quantum numbers of an atom can be defined on the basis of **[AIIMS 2002]**

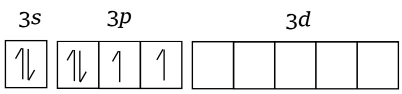
(a) Hund's rule

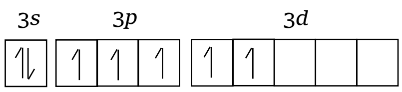
(b) Aufbau's principle

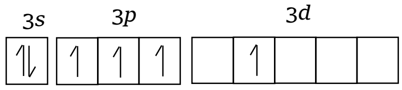
(c) Pauli's exclusion principle

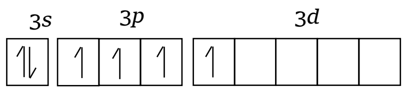
(d) Heisenberg's uncertainty principle

145. Which of the following has maximum energy **[AIIMS 2002]**

(a) 

(b) 

(c) 

(d) 

146. The total magnetic quantum numbers for *d*-orbital is given by

(a) 2 (b) 0, ± 1, ± 2

(c) 0, 1, 2 (d) 5

147. The outer electronic structure $3s^2 3p^5$ is possessed by **[Pb. PMT 2002; Pb. CET 2001]**

(a) Cl (b) O

(c) Ar (d) Br

148. Which of the following set of quantum number is not possible **[Pb. PMT 2002]**

	n	l	m_1	m_2
(a)	3	2	1	+ 1/2
(b)	3	2	1	– 1/2
(c)	3	2	1	0
(d)	5	2	– 1	+ 1/2

149. The configuration $1s^2, 2s^2 2p^5, 3s^1$ shows **[Pb. PMT 2002]**

(a) Excited state of O_2^-

(b) Excited state of neon

(c) Excited state of fluorine

(d) Ground state of fluorine atom

150. The quantum number '*m*' of a free gaseous atom is associated with **[AIIMS 2003]**

(a) The effective volume of the orbital

(b) The shape of the orbital

(c) The spatial orientation of the orbital

(d) The energy of the orbital in the absence of a magnetic field

151. Correct statement is **[BHU 2003]**

(a) $K=4s^1$, $Cr=3d^4 4s^2$, $Cu=3d^{10} 4s^2$

(b) $K=4s^2$, $Cr=3d^4 4s^2$, $Cu=3d^{10} 4s^2$

(c) $K=4s^2$, $Cr=3d^5 4s^1$, $Cu=3d^{10} 4s^2$

(d) $K=4s^1$, $Cr=3d^5 4s^1$, $Cu=3d^{10} 4s^1$

152. Number of orbitats in *h* sub-shell is **[BHU 2003]**

(a) 11 (b) 15

(c) 17 (d) 19

153. Electronic configuration $1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^5, 4s^1$ represents **[CPMT 2003]**

(a) Ground state (b) Excited state

(c) Anionic state (d) All of these

154. Which of the following sets is possible for quantum numbers **[RPET 2003]**

(a) $n=4, l=3, m=-2, s=0$

(b) $n=4, l=4, m=+2, s=-\frac{1}{2}$

(c) $n=4, l=4, m=-2, s=+\frac{1}{2}$

(d) $n=4, l=3, m=-2, s=+\frac{1}{2}$

155. For principle quantum number $n=4$ the total number of orbitals having $l=3$ **[AIIMS 2004]**

(a) 3 (b) 7

(c) 5 (d) 9

156. The number of $2p$ electrons having spin quantum number $s=-1/2$ are **[KCET 2004]**

(a) 6 (b) 0

(c) 2 (d) 3

157. Which of the following sets of quantum numbers is correct for an electron in $4f$ orbital **[AIEEE 2004]**

(a) $n=4, l=3, m=+1, s=+\frac{1}{2}$

(b) $n=4, l=4, m=-4, s=-\frac{1}{2}$

(c) $n=4, l=3, m=+4, s=+\frac{1}{2}$

(d) $n=3, l=2, m=-2, s=+\frac{1}{2}$

158. Consider the ground state of $(Z=24)$. The numbers of electrons with the azimuthal quantum numbers, $l=1$ and 2 are, respectively **[AIEEE 2004]**

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

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

159. The four quantum numbers of the valence electron of potassium are **[DPMT 2004]**

(a) 4, 1, 0 and $\frac{1}{2}$ (b) 4, 0, 1 and $\frac{1}{2}$

(c) 4, 0, 0 and $+\frac{1}{2}$ (d) 4, 1, 1 and $\frac{1}{2}$

160. Which of the following electronic configuration is not possible according to Hund's rule **[Kerala PMT 2004]**

(a) $1s^2 2s^2$ (b) $1s^2 2s^1$

(c) $1s^2 2s^2 2p_x^1 2p_y^1 2p_x^1$ (d) $1s^2 2s^2 2p_x^2$

(e) $1s^2 2s^2 2p_x^2 2p_y^1 2p_z^1$

161. The ground state term symbol for an electronic state is governed by **[UPSEAT 2004]**

(a) Heisenberg's principle

(b) Hund's rule

(c) Aufbau principle

(d) Pauli exclusion principle

162. The electronic configuration of element with atomic number 24 is **[Pb. CET 2004]**

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

163. The maximum number of electrons in p-orbital with $n=5, m=1$ is **[Pb. CET 2003]**

(a) 6 (b) 2

(c) 14 (d) 10

164. Number of two electron can have the same values of ...... quantum numbers **[UPSEAT 2004]**

(a) One (b) Two

(c) Three (d) Four

165. The number of orbitals present in the shell with $n=4$ is **[UPSEAT 2004]**

(a) 16 (b) 8

(c) 18 (d) 32

166. Which of the following electronic configuration is not possible **[MHCET 2003]**

(a) $1s^2 2s^2$ (b) $1s^2, 2s^2 2p^6$

(c) $[Ar]3d^{10}, 4s^2 4p^2$ (d) $1s^2, 2s^2 2p^2, 3s^1$

167. p_x orbital can accommodate **[MNR 1990; IIT 1983; MADT Bihar 1995; BCECE 2005]**

(a) 4 electrons

(b) 6 electrons

(c) 2 electrons with parallel spins

(d) 2 electrons with opposite spins

168. The maximum number of electrons that can be accommodated in 'f' sub shell is **[CPMT 1983, 84; MP PET/PMT 1988; BITS 1988]**

(a) 2 (b) 8

(c) 32 (d) 14

169. The number of electrons which can be accommodated in an orbital is **[DPMT 1981; AFMC 1988]**

(a) One (b) Two

(c) Three (d) Four

170. The number of electrons in the atom which has 20 protons in the nucleus **[CPMT 1981, 93; CBSE PMT 1989]**

(a) 20 (b) 10

(c) 30 (d) 40

171. The maximum number of electrons accommodated in $5f$ orbitals are **[MP PET 1996]**

(a) 5 (b) 10

(c) 14 (d) 18

172. The maximum number of electrons in an atom with $l=2$ and $n=3$ is **[MP PET/PMT 1998]**

(a) 2 (b) 6
(c) 12 (d) 10

173. The configuration $1s^2 2s^2 2p^5 3s^1$ shows **[AIIMS 1997]**
(a) Ground state of fluorine atom
(b) Excited state of fluorine atom
(c) Excited state of neon atom
(d) Excited state of ion O_2^-

174. For sodium atom the number of electrons with $m=0$ will be
(a) 2 (b) 7
(c) 9 (d) 8

175. The number of electrons that can be accommodated in dz^2 orbital is
(a) 10 (b) 1
(c) 4 (d) 2

176. Number of unpaired electrons in $1s^2 2s^2 2p^3$ is
[CPMT 1982; MP PMT 1987; BHU 1987; CBSE PMT 1990; CET Pune 1998; AIIMS 2000]
(a) 2 (b) 0
(c) 3 (d) 1

177. Total number of unpaired electrons in an atom of atomic number 29 is **[CPMT 1984, 93]**
(a) 1 (b) 3
(c) 4 (d) 2

178. The number of unpaired electrons in $1s^2, 2s^2 2p^4$ is
[NCERT 1984; CPMT 1991; MP PMT 1996, 2002]
(a) 4 (b) 2
(c) 0 (d) 1

179. The maximum number of electrons that can be accommodated in a $3d$ subshell is
(a) 2 (b) 10
(c) 6 (d) 14

180. The maximum number of electrons which each sub-shell can occupy is **[Pb. CET 1989]**
(a) $2n^2$ (b) $2n$
(c) $2(2l+1)$ (d) $(2l+1)$

181. Number of unpaired electrons in the ground state of beryllium atom is
(a) 2 (b) 1
(c) 0 (d) All the above

182. How many unpaired electrons are present in Ni^{2+} cation (atomic number = 28)
MP PMT 1995; Kerala PMT 2003]
(a) 0 (b) 2
(c) 4 (d) 6

183. The number of unpaired electrons in an O_2 molecule is
[MNR 1983]
(a) 0 (b) 1
(c) 2 (d) 3

184. The number of unpaired electrons in a chromic ion Cr^{3+} (atomic number = 24) is **[MNR 1986; CPMT 1992]**
(a) 6 (b) 4
(c) 3 (d) 1

185. $3d^{10} 4s^0$ electronic configuration exhibits by
(a) Zn^{++} (b) Cu^{++}
(c) Cd^{++} (d) Hg^{++}

186. **[RPMT 1999]** Which of the following metal ions will have maximum number of unpaired electrons **[CPMT 1996]**
(a) Fe^{+2} (b) CO^{+2}
(c) Ni^{+2} (d) Mn^{+2}

187. **[Kurukshetra CEE 2002]** Which of the metal ion will have highest number of unpaired electrons
(a) Cu^+ (b) Fe^{2+}
(c) Fe^{3+} (d) Co^{2+}

188. The maximum number of unpaired electron can be present in d orbitals are
(a) 1 (b) 3
(c) 5 (d) 7

189. The molecule having one unpaired electron is
(a) NO (b) CO
(c) CN^- (d) O_2

190. A filled or half-filled set of p or d-orbitals is spherically symmetric. Point out the species which has spherical symmetry
(a) Na (b) C
(c) Cl^- (d) Fe

191. The atom of the element having atomic number 14 should have **[AMU 1984]**
(a) One unpaired electron (b) Two unpaired electrons
(c) Three unpaired electrons (d) Four unpaired electrons

192. An atom has 2 electrons in K shell, 8 electrons in L shell and 6 electrons in M shell. The number of s-electrons present in that element is **[CPMT 1989]**
(a) 6 (b) 5
(c) 7 (d) 10

193. The number of unpaired electrons in carbon atom in excited state is **[MNR 1987]**
(a) One (b) Two
(c) Three (d) Four

194. Maximum number of electrons present in 'N' shell is
[IIT 1981; MNR 1984;
[EAMCET 1984]
(a) 18 (b) 32
(c) 2 (d) 8

195. The number of d electrons in Fe^{+2} (atomic number of $Fe=26$) is not equal to that of the **[MNR 1993]**
(a) p-electrons in Ne (At. No.= 10)
(b) s-electrons in Mg (At. No.= 12)

(c) d-electrons in Fe

(d) p-electrons in Cl^- (At. No. of Cl = 17)

196. A transition metal X has a configuration $[Ar]3d^4$ in its $+3$ oxidation state. Its atomic number is **[EAMCET 1990]**

(a) 25 (b) 26

(c) 22 (d) 19

197. The total number of electrons present in all the p-orbitals of bromine are **[MP PET 1994]**

(a) Five (b) Eighteen

(c) Seventeen (d) Thirty five

198. Which of the following has the maximum number of unpaired electrons **[IIT 1996]**

(a) Mg^{2+} (b) Ti^{3+}

(c) V^{3+} (d) Fe^{2+}

199. Which of the following has more unpaired d-electrons

[CBSE PMT 1999]

(a) Zn^+ (b) Fe^{2+}

(c) N^{3+} (d) Cu^+

200. Maximum electrons in a d-orbital are **[CPMT 1999]**

(a) 2 (b) 10

(c) 6 (d) 14

201. The number of unpaired electrons in $Fe^{3+}(Z=26)$ are

[KCET 2000]

(a) 5 (b) 6

(c) 3 (d) 4

202. How many unpaired electrons are present in cobalt $[Co]$ metal **[RPMT 2002]**

(a) 2 (b) 3

(c) 4 (d) 7

203. The number of unpaired electrons in nitrogen is

[Pb. CET 2002]

(a) 1 (b) 3

(c) 2 (d) None of these

204. Which of the following has the least energy

(a) $2p$ (b) $3p$

(c) $2s$ (d) $4d$

205. Pauli's exclusion principle states that **[CPMT 1983, 84]**

(a) Nucleus of an atom contains no negative charge

(b) Electrons move in circular orbits around the nucleus

(c) Electrons occupy orbitals of lowest energy

(d) All the four quantum numbers of two electrons in an atom cannot be equal

206. For the energy levels in an atom, which one of the following statements is correct **[AIIMS 1983]**

(a) There are seven principal electron energy levels

(b) The second principal energy level can have four sub-energy levels and contains a maximum of eight electrons

(c) The M energy level can have maximum of 32 electrons

(d) The $4s$ sub-energy level is at a higher energy than the $3d$ sub-energy level

207. The statements **[AIIMS 1982]**

(i) In filling a group of orbitals of equal energy, it is energetically preferable to assign electrons to empty orbitals rather than pair them into a particular orbital.

(ii) When two electrons are placed in two different orbitals, energy is lower if the spins are parallel.

are valid for

(a) Aufbau principle

(b) Hund's rule

(c) Pauli's exclusion principle

(d) Uncertainty principle

208. According to Aufbau's principle, which of the three $4d, 5p$ and $5s$ will be filled with electrons first **[MADT Bihar 1984]**

(a) $4d$

(b) $5p$

(c) $5s$

(d) $4d$ and $5s$ will be filled simultaneously

209. The energy of an electron of $2p_y$ orbital is **[AMU 1984]**

(a) Greater than that of $2p_x$ orbital

(b) Less than that of $2p_x$ orbital

(c) Equal to that of $2s$ orbital

(d) Same as that of $2p_z$ orbital

210. Which of the following principles/rules limits the maximum number of electrons in an orbital to two **[CBSE PMT 1989]**

(a) Aufbau principle

(b) Pauli's exclusion principle

(c) Hund's rule of maximum multiplicity

(d) Heisenberg's uncertainty principle

211. The electrons would go to lower energy levels first and then to higher energy levels according to which of the following

[BHU 1990; MP PMT 1993]

(a) Aufbau principle

(b) Pauli's exclusion principle

(c) Hund's rule of maximum multiplicity

(d) Heisenberg's uncertainty principle

212. Energy of atomic orbitals in a particular shell is in the order

[AFMC 1990]

(a) $s<p<d<f$ (b) $s>p>d>f$

(c) $p<d<f<s$ (d) $f>d>s>p$

213. Aufbau principle is not satisfied by **[MP PMT 1997]**

(a) Cr and Cl (b) Cu and Ag
(c) Cr and Mg (d) Cu and Na

214. Which of the following explains the sequence of filling the electrons in different shells **[AIIMS 1998; BHU 1999]**
(a) Hund's rule (b) Octet rule
(c) Aufbau principle (d) All of these

215. Aufbau principle is obeyed in which of the following electronic configurations **[AFMC 1999]**
(a) $1s^2 2s^2 2p^6$ (b) $1s^2 3p^3 3s^2$
(c) $1s^2 3s^2 3p^6$ (d) $1s^2 2s^2 3s^2$

216. Following Hund's rule which element contains six unpaired electron **[RPET 2000]**
(a) Fe (b) Co
(c) Ni (d) Cr

217. Electron enters the sub-shell for which $(n+l)$ value is minimum. This is enunciated as **[RPMT 2000]**
(a) Hund's rule
(b) Aufbau principle
(c) Heisenberg uncertainty principle
(d) Pauli's exclusion principle

218. The atomic orbitals are progressively filled in order of increasing energy. This principle is called as **[MP PET 2001]**
(a) Hund's rule (b) Aufbau principle
(c) Exclusion principle (d) de-Broglie rule

219. The correct order of increasing energy of atomic orbitals is **[MP PET 2002]**
(a) $5p < 4f < 6s < 5d$ (b) $5p < 6s < 4f < 5d$
(c) $4f < 5p < 5d < 6s$ (d) $5p < 5d < 4f < 6s$

220. The orbital with maximum energy is **[CPMT 2002]**
(a) $3d$ (b) $5p$
(c) $4s$ (d) $6d$

221. p-orbitals of an atom in presence of magnetic field are **[Pb. PMT 2002]**
(a) Two fold degenerate (b) Non degenerate
(c) Three fold degenerate (d) None of these

222. Orbital angular momentum for a d-electron is **[MP PET 2003]**
(a) $\frac{6h}{2\pi}$ (b) $\frac{\sqrt{6}\,h}{2\pi}$
(c) $\frac{12h}{2\pi}$ (d) $\frac{\sqrt{12}\,h}{2\pi}$

223. Number of nodal centres for $2s$ orbital **[RPET 2003]**
(a) 1 (b) 0
(c) 4 (d) 3

224. The orbital angular momentum of an electron in $2s$-orbital is
(a) $\frac{1}{2}\frac{h}{2\pi}$ (b) $\frac{h}{2\pi}$
(c) $\sqrt{2}\frac{h}{2\pi}$ (d) Zero

225. The maximum number of electrons present in an orbit $l = 3$, is **[Pb. PMT 2004]**
(a) 6 (b) 8
(c) 10 (d) 14

226. Number of unpaired electrons in Mn^{4+} is **[DPMT 2005]**
(a) 3 (b) 5
(c) 6 (d) 4

227. Which of the following sequence is correct as per Aufbau principle **[DPMT 2005]**
(a) $3s < 3d < 4s < 4p$ (b) $1s < 2p < 4s < 3d$
(c) $2s < 5s < 4p < 5d$ (d) $2s < 2p < 3d < 3p$

228. Electronic configuration of deuterium atom is **[J&K CET 2005]**
(a) $1s^1$ (b) $2s^2$
(c) $2s^1$ (d) $1s^2$

Critical Thinking

Objective Questions

1. Which of the following atoms and ions are isoelectronic *i.e.* have the same number of electrons with the neon atom **[NCERT 1978]**
(a) F^- (b) Oxygen atom
(c) Mg (d) N^-

2. Atoms consists of protons, neutrons and electrons. If the mass of neutrons and electrons were made half and two times respectively to their actual masses, then the atomic mass of ${}_6C^{12}$
(a) Will remain approximately the same
(b) Will become approximately two times
(c) Will remain approximately half
(d) Will be reduced by 25%

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

4. The electronic configuration of a dipositive metal M^{2+} is 2, 8, 14 and its atomic weight is 56 a.m.u. The number of neutrons in its nuclei would be **[MNR 1984, 89; Kerala PMT 1999]**
(a) 30 (b) 32
(c) 34 (d) 42

5. **[MP PET 2004]** The ratio of the energy of a photon of $2000\,Å$ wavelength radiation to that of $4000\,Å$ radiation is **[IIT 1986; DCE 2000; JIPMER 2000]**
(a) 1/4 (b) 4

(c) 1/2 (d) 2

6. Discovery of the nucleus of an atom was due to the experiment carried out by **[CPMT 1983; MP PET 1983]**

(a) Bohr (b) Mosley

(c) Rutherford (d) Thomson

7. In a Bohr's model of atom when an electron jumps from $n=1$ to $n=3$, how much energy will be emitted or absorbed **[CBSE PMT 1996]**

(a) $2.15\times10^{-11}\,erg$ (b) $0.1911\times10^{-10}\,erg$

(c) $2.389\times10^{-12}\,erg$ (d) $0.239\times10^{-10}\,erg$

8. The nucleus of an atom can be assumed to be spherical. The radius of the nucleus of mass number A is given by $1.25\times10^{-13}\times A^{1/3}cm$ Radius of atom is one $Å$. If the mass number is 64, then the fraction of the atomic volume that is occupied by the nucleus is **[NCERT 1983]**

(a) 1.0×10^{-3} (b) 5.0×10^{-5}

(c) 2.5×10^{-2} (d) 1.25×10^{-13}

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

[IIT 1998; Orissa JEE 2005]

(a) $-3.4eV$ (b) $-4.2eV$

(c) $-6.8eV$ (d) $+6.8eV$

10. The energy of the electron in the first orbit of He^+ is $-871.6\times10^{-20}\,J$. The energy of the electron in the first orbit of hydrogen would be **[Roorkee Qualifying 1998]**

(a) $-871.6\times10^{-20}\,J$ (b) $-435.8\times10^{-20}\,J$

(c) $-217.9\times10^{-20}\,J$ (d) $-108.9\times10^{-20}\,J$

11. The total number of valence electrons in $4.2\,gm$ of N_3^- ion is (N_A is the Avogadro's number) **[CBSE PMT 1994]**

(a) $1.6N_A$ (b) $3.2N_A$

(c) $2.1N_A$ (d) $4.2N_A$

12. 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 **[CBSE PMT 1998; BHU 1999]**

(a) $0.13\,Å$ (b) $1.06\,Å$

(c) $4.77\,Å$ (d) $2.12\,Å$

13. The frequency of a wave of light is $12\times10^{14}\,s^{-1}$. The wave number associated with this light is **[Pb. PMT 1999]**

(a) $5\times10^{-7}\,m$ (b) $4\times10^{-8}\,cm^{-1}$

(c) $2\times10^{-7}\,m^{-1}$ (d) $4\times10^{4}\,cm^{-1}$

14. The series limit for Balmer series of H-spectra is

[AMU (Engg.) 1999]

(a) 3800 (b) 4200

(c) 3646 (d) 4000

15. The ionization energy of hydrogen atom is $-13.6\,eV$. The energy required to excite the electron in a hydrogen atom from the ground state to the first excited state is (Avogadro's constant = 6.022 × 10^{23}) **[BHU 1999]**

(a) $1.69\times10^{-20}\,J$ (b) $1.69\times10^{-23}\,J$

(c) $1.69\times10^{23}\,J$ (d) $1.69\times10^{25}\,J$

16. The energy required to dislodge electron from excited isolated H-atom, $IE_1=13.6\ eV$ is **[DCE 2000]**

(a) $=13.6\,eV$ (b) $>13.6\,eV$

(c) <13.6 and $>3.4\,eV$ (d) $\le3.4\,eV$

17. The number of nodal planes in a p_x is

[IIT Screening 2000]

(a) One (b) Two

(c) Three (d) Zero

18. The third line in Balmer series corresponds to an electronic transition between which Bohr's orbits in hydrogen

[MP PMT 2001]

(a) $5\rightarrow3$ (b) $5\rightarrow2$

(c) $4\rightarrow3$ (d) $4\rightarrow2$

19. Which of the following has maximum number of unpaired electron (atomic number of Fe 26) **[MP PMT 2001]**

(a) Fe (b) Fe (II)

(c) Fe (III) (d) Fe (IV)

20. The frequency of one of the lines in Paschen series of hydrogen atom is $2.340\times10^{11}\,Hz$. The quantum number n_2 which produces this transition is **[DPMT 2001]**

(a) 6 (b) 5

(c) 4 (d) 3

21. Which of the following electron transition in a hydrogen atom will require the largest amount of energy

[UPSEAT 1999, 2000, 01]

(a) From $n=1$ to $n=2$ (b) From $n=2$ to $n=3$

(c) From $n=\infty$ to $n=1$ (d) From $n=3$ to $n=5$

22. 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) $3\rightarrow2$ (b) $5\rightarrow2$

(c) $4\rightarrow1$ (d) $2\rightarrow5$

23. The value of Planck's constant is $6.63\times10^{-34}\,Js$. The velocity of light is $3.0\times10^{8}\,ms^{-1}$. Which value is closest to the wavelength in nanometres of a quantum of light with frequency of $8\times10^{15}\,s^{-1}$

(a) 3×10^{7} (b) 2×10^{-25}

(c) 5×10^{-18} (d) 4×10^{1}

24. As electron moves away from the nucleus, its potential energy **[UPSEAT 2003]**

(a) Increases (b) Decreases

(c) Remains constant (d) None of these

Assertion & Reason

For AIIMS Aspirants

Read the assertion and reason carefully to mark the correct option out of the options given below :

(a) If both assertion and reason are true and the reason is the correct explanation of the assertion.
(b) If both assertion and reason are true but reason is not the correct explanation of the assertion.
(c) If assertion is true but reason is false.
(d) If the assertion and reason both are false.
(e) If assertion is false but reason is true.

1. Assertion : The position of an electron can be determined exactly with the help of an electron microscope.

Reason : 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. **[NDA 1999]**

2. Assertion : A spectral line will be seen for a $2p_x - 2p_y$ transition.

Reason : Energy is released in the form of wave of light when the electron drops from $2p_x - 2p_y$ orbital. **[AIIMS 1996]**

3. Assertion : The cation energy of an electron is largely determined by its principal quantum number.

Reason : The principal quantum number n is a measure of the most probable distance of finding the electron around the nucleus. **[AIIMS 1996]**

4. Assertion : Nuclide $^{30}Al_{13}$ is less stable than $^{40}Ca_{20}$

Reason : Nuclides having odd number of protons and neutrons are generally unstable **[IIT 1998]**

5. Assertion : The atoms of different elements having same mass number but different atomic number are known as isobars

Reason : The sum of protons and neutrons, in the isobars is always different **[AIIMS 2000]**

6. Assertion : Two electrons in an atom can have the same values of four quantum numbers.

Reason : Two electrons in an atom can be present in the same shell, sub-shell and orbital and have the same spin **[AIIMS 20**

7. Assertion : The value of n for a line in Balmer series of hydrogen spectrum having the highest wave length is 4 and 6.

Reason : For Balmer series of hydrogen spectrum, the value $n_1 = 2$ and $n_2 = 3, 4, 5$. **[AIIMS 1992]**

8. Assertion : Absorption spectrum conists of some bright lines separated by dark spaces.

Reason : Emission spectrum consists of dark lines. **[AIIMS 2002]**

9. Assertion : A resonance hybrid is always more stable than any of its canonical structures.

Reason : This stability is due to delocalization of electrons. **[AIIMS 1999]**

10. Assertion : Cathode rays do not travel in straight lines.

Reason : Cathode rays penetrate through thick sheets **[AIIMS 1996]**

11. Assertion : Electrons revolving around the nucleus do not fall into the nucleus because of centrifugal force.

Reason : Revolving electrons are planetary electrons. **[AIIMS 1994]**

12. Assertion : Threshold frequency is a characteristic for a metal.

Reason : Threshold frequency is a maximum frequency required for the ejection of electron from the metal surface.

13. Assertion : The radius of the first orbit of hydrogen atom is 0.529Å.

Reason : Radius for each circular orbit $(r_n) = 0.529Å\ (n^2/Z)$, where $n = 1,2,3$ and $Z =$ atomic number.

14. Assertion : $3d_{z^2}$ orbital is spherically symmetrical.

Reason : $3d_{z^2}$ orbital is the only d-orbital which is spherical in shape.

15. Assertion : Spin quantum number can have the value +1/2 or −1/2.

Reason : (+) sign here signifies the wave function.

16. Assertion : Total number of orbitals associated with principal quantum number $n = 3$ is 6.

Reason : Number of orbitals in a shell equals to $2n$.

17. Assertion : Energy of the orbitals increases as $1s < 2s = 2p < 3s = 3p < 3d < 4s = 4p = 4d = 4f < ......$

Reason : Energy of the electron depends completely on principal quantum number.

18. Assertion : Splitting of the spectral lines in the presence of magnetic field is known as stark effect.

Reason : Line spectrum is simplest for hydrogen atom.

19. Assertion : Thomson's atomic model is known as 'raisin pudding' model.

Reason : The atom is visualized as a pudding of positive charge with electrons (raisins) embedded in it.

20. Assertion : Atomic orbital in an atom is designated by n, l, m_l and m_s.

Reason : These are helpful in designating electron present in an orbital.

21. Assertion : The transition of electrons $n_3 \rightarrow n_2$ in H atom will emit greater energy than $n_4 \rightarrow n_3$.

Reason : n_3 and n_2 are closer to nucleus tan n_4.

22. Assertion : Cathode rays are a stream of α-particles.

Reason : They are generated under high pressure and high voltage.

23. Assertion : In case of isoelectronic ions the ionic size increases with the increase in atomic number.

Reason : The greater the attraction of nucleus, greater is the ionic radius.

Answers

Discovery and Properties of anode, cathode rays neutron and Nuclear structure

1	d	2	a	3	c	4	c	5	b
6	a	7	b	8	a	9	d	10	c
11	b	12	d	13	b	14	a	15	b
16	b	17	c	18	c	19	c	20	b
21	a	22	d	23	c	24	b	25	d
26	c	27	b	28	d	29	c	30	a
31	b	32	d	33	b	34	c	35	c
36	a	37	b	38	a	39	d	40	c
41	c								

Atomic number, Mass number, Atomic species

1	b	2	a	3	b	4	b	5	a
6	a	7	c	8	b	9	c	10	b
11	b	12	c	13	b	14	c	15	c
16	c	17	c	18	a	19	c	20	a
21	c	22	b	23	c	24	d	25	b
26	b	27	a	28	a	29	c	30	b
31	c	32	d	33	d	34	c	35	c
36	c	37	c	38	b	39	d	40	c
41	b	42	c	43	a	44	c	45	b
46	c	47	d	48	a	49	c	50	c
51	a	52	c	53	b	54	a	55	c
56	a	57	d	58	c	59	a	60	a
61	d	62	b	63	a	64	c	65	b
66	a	67	c	68	a	69	d	70	d
71	c	72	a	73	b	74	d		

Atomic models and Planck's quantum theory

1	c	2	a	3	b	4	b	5	d
6	b	7	c	8	b	9	c	10	a
11	b	12	a	13	d	14	b	15	b
16	c	17	a	18	c	19	a	20	d
21	d	22	c	23	d	24	d	25	c
26	a	27	c	28	b	29	c	30	a
31	b	32	c	33	d	34	b	35	b
36	a	37	c	38	c	39	c	40	a
41	c	42	d	43	d	44	a	45	d
46	b	47	a	48	c	49	d	50	a
51	a	52	c	53	d	54	c	55	b
56	b	57	b	58	a	59	b	60	c
61	c	62	b	63	c	64	c	65	b
66	b	67	c	68	a	69	b	70	d
71	a	72	d	73	a	74	c	75	d
76	b	77	a	78	a	79	c	80	a
81	a								

Dual nature of electron

1	c	2	a	3	a	4	b	5	c
6	b	7	d	8	a	9	d	10	d

11	c	12	c	13	b	14	b	15	b
16	c	17	c	18	c	19	b	20	a
21	d								

Uncertainty principle and Schrodinger wave equation

1	b	2	b	3	a	4	c	5	c
6	c	7	b	8	d	9	d	10	a
11	a	12	c	13	a	14	b	15	d
16	b	17	a	18	c	19	c	20	b

Quantum number, Electronic configuration and Shape of orbitals

1	c	2	a	3	b	4	d	5	c
6	c	7	c	8	a	9	a	10	a
11	c	12	c	13	a	14	a	15	d
16	c	17	c	18	d	19	b	20	c
21	c	22	a	23	c	24	d	25	c
26	c	27	b	28	d	29	e	30	b
31	d	32	a	33	c	34	d	35	d
36	c	37	b	38	b	39	d	40	c
41	d	42	c	43	c	44	a	45	a
46	a	47	b	48	c	49	c	50	b
51	c	52	b	53	b	54	b	55	c
56	c	57	b	58	e	59	c	60	c
61	d	62	d	63	d	64	c	65	b
66	d	67	c	68	d	69	c	70	b
71	a	72	c	73	c	74	c	75	a
76	c	77	c	78	c	79	d	80	d
81	b	82	c	83	a	84	a	85	b
86	c	87	a	88	b	89	c	90	b
91	d	92	a	93	b	94	b	95	d
96	d	97	a	98	a	99	d	100	c
101	b	102	d	103	a	104	c	105	d
106	a	107	c	108	d	109	a	110	d
111	d	112	b	113	c	114	b	115	b
116	a	117	c	118	b	119	a	120	a
121	d	122	b	123	b	124	b	125	d
126	d	127	b	128	c	129	a	130	b
131	a	132	c	133	d	134	b	135	a
136	a	137	c	138	c	139	c	140	c
141	c	142	d	143	c	144	c	145	b
146	d	147	a	148	c	149	b	150	c

151	d	152	a	153	a	154	d	155	b
156	d	157	a	158	b	159	c	160	d
161	c	162	d	163	b	164	c	165	a
166	d	167	d	168	d	169	b	170	a
171	c	172	d	173	c	174	b	175	d
176	c	177	a	178	b	179	b	180	c
181	c	182	b	183	c	184	c	185	a
186	d	187	c	188	c	189	a	190	c
191	b	192	a	193	d	194	b	195	d
196	a	197	c	198	d	199	b	200	b
201	a	202	b	203	b	204	c	205	d
206	b	207	b	208	c	209	d	210	b
211	a	212	a	213	b	214	c	215	a
216	d	217	b	218	b	219	b	220	d
221	b	222	b	223	a	224	d	225	d
226	a	227	b	228	a				

Critical Thinking Questions

1	a	2	d	3	d	4	a	5	d
6	c	7	b	8	d	9	a	10	c
11	a	12	d	13	d	14	c	15	b
16	d	17	a	18	b	19	c	20	b
21	a	22	a	23	d	24	a		

Assertion & Reason

1	d	2	d	3	a	4	a	5	c
6	d	7	e	8	d	9	a	10	e
11	b	12	c	13	a	14	d	15	c
16	d	17	c	18	e	19	a	20	e
21	b	22	d	23	d				

Answers and Solutions

Discovery and Properties of anode, cathode rays neutron and Nuclear structure

1. (d) Neutrons and protons in the nucleus and electrons in the extranuclear region.

2. (a) It consists of proton and neutron and these are also known as nucleones.

3. (c) Radius of nucleus $\simeq 10^{-15} m$.

4. (c) Positive ions are formed from the neutral atom by the loss of electrons.

5. (b) The β-ray particle constitute electrons.

6. (a) James Chadwick discovered neutron $({}_0n^1)$.

7. (b) Charge/mass for

$$n = 0, \alpha = \frac{2}{4}, p = \frac{1}{1} \text{ and } e = \frac{1}{1/1837}$$

9. (d) The density of neutrons is of the order $10^{11} kg/cc.$

10. (c) This is because chargeless particles do not undergo any deflection in electric or magnetic field.

11. (b) Neutron and proton found in nucleus.

13. (b) Cathode rays are made up of negatively charged particles (electrons) which are deflected by both the electric and magnetic fields.

15. (b) Mass of neutron is greater than that of proton, meson and electron.

Mass of neutron = mass of proton + mass of electron

16. (b) Proton is 1837 (approx 1800) times heavier than an electron. Penetration power $\propto \frac{1}{\text{mass}}$

18. (c) Nucleus of helium is ${}_2He^4$ mean 2 neutrons and 2 protons.

19. (c) Proton is the nucleus of $H-$atom ($H-$atom devoid of its electron).

20. (b) Cathode rays are made up of negatively charged particles (electrons, e^-)

26. (c) Size of nucleus is measured in *Fermi* (1 Fermi $= 10^{-15} m$).

27. (b) A molecule of an element is a incorrect statement. The correct statement is "an element of a molecule".

29. (c) Proton is represented by p having charge +1 discovered in 1988 by Goldstein.

31. (b) The nature of anode rays depends upon the nature of residual gas.

32. (d) H^+ (proton) will have very large hydration energy due to its very small ionic size

Hydration energy $\propto \frac{1}{\text{Size}}$

33. (b) Mass of a proton $=1.673\times10^{-24}\,g$

$\therefore$ Mass of one mole of proton

$=9.1\times10^{-24}\times6.02\times10^{23}=10.07\times10^{-1}=1.008\,g$

Mass of a electron $=9.1\times10^{-28}\,g$

$\therefore$ Mass of one mole of electron

$=9.1\times10^{-28}\times6.02\times10^{23}=54.78\times10^{-5}\,g=0.55\,mg$.

35. (c) One mole of electron = 6.023×10^{23} electron

Mass of one electron = 9.1×10^{-28} *gm*

Mass of one mole of electrons

= $6.023\times10^{23}\times9.1\times10^{-28}\,gm$ = $5.48\times10^{-4}\,gm$

= $5.48\times10^{-4}\times1000\,mg$ = $0.548\,gm\approx0.55\,mg$.

36. (a) Charge on proton $=+1$ unit, charge on α particle = + 2 units, 2 : 1.

37. (b) $m_p/m_e\simeq1837\simeq1.8\times10^3$.

38. (a) Splitting of signals is caused by protons attached to adjacent carbon provided these are not equivalent to the absorbing proton.

39. (d) Nucleus consists of proton and neutron both are called as nucleon.

40. (c) Positron $(+1e^0)$ has the same mass as that of an electron $(-1e^0)$.

41. (c) Electron $\frac{1}{1837}$ time lighter than proton so their mass ratio will be 1 : 1837

Atomic number, Mass number, Atomic species

1. (b) The number of electrons in an atom is equal to its atomic number *i.e.* number of protons.

2. (a) No. of protons = Atomic no. = 25 and no. of neutron = 55 – 25 = 30.

3. (b) No. of neutrons = mass number – no. of protons. = *W* – *N*.

4. (b) ${}_{30}Zn^{70}, Zn^{2+}$ has No. of Neutrons = 70 – 30 = 40.

5. (a) Na^+ and *Ne* are isoelectronic which contain 10 electrons.

6. (a) One molecule of CO_2 have 22 electrons.

7. (c) Cl and Cl^- differs in number of electrons. *Cl* has $17e^-$ while Cl^- has $18e^-$.

8. (b) *CO* and CN^- are isoelectronic.

$CO=6+8=14$ and $CN^-=6+7+1=14$.

9. (c) Mass of an atom is due to nucleus (neutron + proton).

10. (b) Atomic number is defined as the number of protons in the nucleus.

11. (b) ${}_{26}X^{56}$ $A=P+N=Z+N=E+N$

$N=A-E=56-26=30$

12. (c) Most probable radius = a_0/Z

where a_0 = 52.9 pm. For helium ion, Z = 2.

$r_{mp}=\frac{52.9}{2}=26.45\,pm.$

13. (b) Four unpaired electron are present in the Fe^{2+} ion $Fe^{2+}_{26}=[Ar]3d^6,4s^0$

14. (c) Na^+ has 10 electron and Li^+ has 2 electron so these are different number of electron from each other.

16. (c) $P_{15}=2,8,5$

17. (c) ${}_8O=1s^22s^22p^4$

18. (a) ${}_{35}Br^{80}=1s^2\,2s^2\,2p^63s^23p^63d^{10}4s^24p^5$

$A=80$, $Z=35$, $N=?$

$N=A-Z=80-35=45$

atomic number (Proton) is 35 and no. of neutron is 45.

19. (c) ${}^{16}_{8}O^{--}$ have more electrons than neutron $p=8,\ e=10,\ n=8$.

20. (a) ${}_6A^{12}$ and ${}_6X^{13}$ both are isotopes but have different no. of neutrons.

${}_6A^{12}$, For *A* have $p=6, e=6$ and $n=6$ and

${}_6X^{13}$, For *B* have $p=6, e=6$ and $n=7$

21. (c) *P* = 20, mass no. (*A*) = 40

$N=A-P=40-20=20$

$P=N=20$.

22. (b) Electrons in $Na^+=11-1=10$

Electrons in $Mg^{2+}=12-2=10$

23. (c) ${}_{20}Ca^{40}$ has 20 proton, 20 neutron.

24. (d) $CH_3^+=6+3-1=8e^-$,

$H_3O^+=3+8-1=10e^-$,

$NH_3=7+3=10e^-$, $CH_3^-=6+3+1=10e^-$

25. (b) $-CONH_2=6+8+7+2+1$ (from other atom to form covalent bond) = 24.

26. (b) Complete $E.C.=[Ar]^{18}\,3d^{10}4s^24p^6$.

Hence no. of e^- = no. of protons $=36=Z$.

28. (a) $K^+=1s^2\,2s^22p^63s^23p^6$

$Cl^-=1s^22s^22p^63s^23p^6$.

29. (c) Mass no. ≈ At. Wt.

Mass no. = no. of protons + no. of neutrons

At. no. = no. of protons.

30. (b) $N_2O = 14 + 8 = 22$

$CO_2 = 6 + 16 = 22$.

31. (c) Neutron in ${}^{12}_{6}C = 6$, , Neutrons in ${}^{28}_{14}Si = 14$

Ratio = 6 : 14 = 3 : 7.

33. (d) $N_7 = 1s^2\,2s^2 2p^3$

$N^+ = 1s^2\,2s^2 2p^2$

$C = 1s^2\,2s^2 2p^2$.

34. (c) $O = C = O$, linear structure 180^o angle

$Cl - Hg - Cl$, linear structure 180^o angle.

35. (c) $H^- = 1s^2$ and $He^+ = 1s^2$.

36. (c) In the nucleus of an atom only proton and neutrons are present.

37. (c) Cu^{63}_{29} Number of neutrons = atomic mass – atomic number = 63 – 29 = 34.

38. (b) 21 Protons and 24 Neutrons are present in nucleus and element is *Sc*.

40. (c) ${}_7X^{14}, n = 14 - 7 = 7$

42. (c) Cl^- have 17 proton, 18 neutron and 18 electron.

43. (a) Number of unpaired electrons in inert gas is zero because they have full filled orbitals.

44. (c) Electrons and Protons are same in neutral atom.

48. (d) No. of proton and no. of electron = $18\,[Ar^{36}_{18}]$ and No. of neutron = 20

Mass number = *P* + *N* = $18 + 20 = 38$.

49. (c) In Xe^{231}_{89} number of protons and electrons is 89 and No. of neutrons = *A* – *Z* = 231 – 89 = 142.

51. (a) NO_2^- and O_3 are isostere. The number of atoms in these $(= 3)$ and number of electrons (24) are same.

52. (c) Number of electrons in nitrogen = 7 and number of electron is oxygen = 8 we know that formula of nitrate ion is NO_3^- we also know that number of electron

= (1 × Number of electrons in nitrogen)

+ (3 × number of electrons in oxygen) + 1

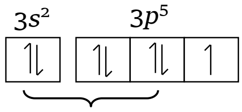
= (1 × 7) + (3 × 8) + 1 = 32.

53. (b) Atomicity = $\frac{\text{Molecular mass}}{\text{Atomic mass}} = \frac{256}{32} = 8 = S_8$.

54. (a) In case of N^{3-}, $p = 7$ and $c = 10$

55. (c) Chlorine $Cl_{17} = [Ne]$

$3s^2$ $3p^5$



Three electron

56. (a) Bromine consists of outer most electronic configuration $[Ar]\,3d^{10}\,4s^2\,4p^5$.

57. (d) $Na^+ = 1s^2\,2s^2\,2p^6$

$Mg^{++} = 1s^2\,2s^2\,2p^6$

$O^{2-} = 1s^2\,2s^2\,2p^6$

$Cl^- = 1s^2\,2s^2\,2p^6 3s^2 3p^6$

60. (a) Ar^{40}_{18} = atomic number 18 and no. of Neutron in case of Ar_{22}

Neutron = Atomic mass – Atomic number

$= 40 - 18 = 22$

61. (d) Nucleus of tritium contain $[H^3_1]$

$p = 1$, $e = 1$, $n = 2$

62. (b) N^{3-}, F^- and Na^+ (These three ions have $e^- = 10$, hence they are isoelectronic)

63. (a) NO_3^- and CO_3^{2-} consist of same electron and show same isostructural.

64. (c) Atomic number of chlorine 17 and in Cl^- ion total no. of electron =18.

65. (b) Tritium (H^3_1) has one proton and two neutron.

67. (c) $X_{35} = 1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4s^2\,4p^5$

Total no. of e^- is all *p*–orbitals $= 6 + 6 + 5 = 17$.

68. (a) Since its nucleus contain 9 proton so its. atomic number is 9 and its electronic configuration is 2, 7. So it require one more electron to complete its octet. Hence its valency is 1.

69. (d) K_2S formed by K^+ and S^{2-} ion. We know that atomic number of *K* is 19 and in K^+ ion its atomic number would be 18 similarly atomic number of *S* is 16 and in form S^{2-} ion its atomic number would be 18 so the K^+ and S^{2-} are isoelectronic with each other in K_2S .

70. (d) ${}_{20}Ca = 2,\ 8,\ 8,\ 2$

Ca^{2+} = 2, 8, 8

Hence, Ca^{2+} has 8 electrons each in outermost and penultimate shell.

71. (c) Atomic no. of *C* = 6 so the number of protons in the nucleus = 6

72. (a) No. of $P = Z = 7$; No. of electrons in $N^{3-} = 7 + 3 = 10$.

73. (b) Heavy hydrogen is ${}^2_1D$ .Number of neutrons = 1

74. (d) Atomic number is always whole number.

Atomic models and Planck's quantum theory

2. (a) The central part consisting whole of the positive charge and most of the mass caused by nucleus, is extremely small in size compared to the size of the atom.

3. (b) Electrons in an atom occupy the extra nuclear region.

4. (b) According to the Bohr model atoms or ions contain one electron.

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

7. (c) α-particles pass through because most part of the atom is empty.

8. (b) An electron jumps from L to K shell energy is released.

9. (c) Neutron is a chargeless particles, so it does not deflected by electric or magnetic field.

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

11. (b) Both He and Li^+ contain 2 electrons each.

18. (c) During the experimental verification of de-Broglie equation, Davisson and Germer confirmed wave nature of electron.

19. (a) Increases due to absorption of energy and it shows absorption spectra.

20. (d) Rutherford α-Scattering experiment.

21. (d) It represents Heisenberg's uncertainty principle.

23. (d) $\frac{E_4}{E_2} = \frac{2^2}{4^2} = \frac{4}{16} = \frac{1}{4}$; $E_4 = \frac{E_2}{4} = \frac{-328}{4} = -82\ kJ/mol$.

27. (c) When $c = \nu \times \lambda$ than $\lambda = \frac{c}{\nu} = \frac{3\times 10^8}{2\times 10^6} = 1.5\times 10^2 m$

28. (b) According to quantum theory of radiation, a hot body emits radiant energy not continuously but discontinuously in the form of small packets of energy called quanta or photons.

30. (a) $p = \frac{h}{\lambda} = \frac{6.6\times 10^{-34}}{2.2\times 10^{-11}} = 3\times 10^{-23} kgms^{-1}$

34. (b) Bohr's radius $= \frac{n^2h^2}{4\pi^2me^2z}$. Which is a positive quantity.

40. (a) Gold used by Rutherford in scatting experiment.

41. (c) $\Delta E = E_3 - E_2 = 13.6\left[\frac{1}{(2)^2} - \frac{1}{(3)^2}\right] = 1.9\, eV$

42. (d) $R = R_0 (= 1.4\times 10^{-13}\ cm) \times A^{1/3}$

43. (d) $\left(\frac{q}{m}\right)_\alpha = \frac{1}{2}\left(\frac{q}{m}\right)_p = \frac{1}{2}\times 9.6\times 10^7 = 4.8\times 10^7\, C\,kg^{-1}$

44. (a) According to Hydrogen spectrum series.

45. (d) The electron can move only in these circular orbits where the angular momentum is a whole number multiple of $\frac{h}{2\pi}$ or it is quantised.

46. (b) Generally electron moving in orbits according to Bohr's principle.

47. (a) According to the planck's law that energy of a photon is directly proportional to its frequency *i.e.* $E = h\nu$

49. (d) Bohr's radius of the hydrogen atom

$r = \frac{n^2 \times 0.529 Å}{z}$; where z = Atomic number,

n = Number of orbitals

51. (a) $E = -\frac{2.172\times 10^{-18}}{n^2} = \frac{-2.172\times 10^{-18}}{2^2}$

$= -5.42\times 10^{-19} J$.

52. (c) $\Delta E = \frac{hc}{\lambda}$ or $\lambda = \frac{hc}{\Delta E}$

$= \frac{6.64\times 10^{-34}\times 3\times 10^8}{3\times 10^{-8}} = 6.64\times 10^{-8} Å$

53. (d) $r_n = r_1 \times n^2$

$r_3 = 0.53\times 3^2 = 0.53\times 9 = 4.77 Å$

54. (c) According to Rutherford an atom consists of nucleus which is small in size but carries the entire mass (P+ N).

55. (b) Wavelength of spectral line emitted is inversely proportional to energy $\lambda \propto \frac{1}{E}$.

56. (b) $E \propto \frac{1}{\lambda}$; $E_1 = \frac{1}{8000}$; $E_2 = \frac{1}{16000}$

$\frac{E_1}{E_2} = \frac{16000}{8000} = 2 \Rightarrow E_1 = 2E_2$

58. (a) $\nu = \frac{c}{\lambda} = \frac{3\times 10^8\ ms^{-1}}{600\times 10^{-9} m} = 5.0\times 10^{14}\ Hz$.

59. (b) $E = \frac{-13.6}{n^2} eV = \frac{-13.6}{2^2} = \frac{-13.6}{4} = -3.40\ eV$

65. (b) Bohr radius $= \frac{r_2}{r_1} = \frac{(2)^2}{(1)^2} = 4$.

66. (b) $\nu = \frac{1}{\lambda} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] = 109678\left[\frac{1}{1} - \frac{1}{4}\right] = 82258.5$

$\lambda = 1.21567\times 10^{-5} cm$ or $\lambda = 12.1567\times 10^{-6} cm$

$= 12.1567\times 10^{-8}\ m$

$\nu = \frac{c}{\lambda} = \frac{3\times 10^8}{12.567\times 10^{-8}} = 24.66\times 10^{14}\ Hz$.

67. (c) We know that $\lambda = \frac{h}{mv}$; $\therefore\ m = \frac{h}{m\lambda}$

The velocity of photon (v) = $3\times 10^8 m\sec^{-1}$

$\lambda = 1.54\times 10^{-8} cm = 1.54\times 10^{-10}$ meter

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

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

68. (a) The spliting of spectral line by the magnetic field is called Zeeman effect.

69. (b) $r \propto n^2$ (excited state $n = 2$)

$r = 4a_0$

70. (d) $r_n \propto n^2 : A_n \propto n^4$

$$\frac{A_2}{A_1} = \frac{n_2^4}{n_1^4} = \frac{2^4}{1^4} = \frac{16}{1} = 16:1$$

71. (a) It will take $\frac{4\pi^2 mr^2}{nh}$

72. (d) $r_H = 0.529 \frac{n^2}{z} Å$

For hydrogen ; $n=1$ and $z=1$ therefore

$r_H = 0.529 Å$

For $Be^{3+} : Z=4$ and $n=2$ Therefore

$$r_{Be^{3+}} = \frac{0.529 \times 2^2}{4} = 0.529 Å.$$

73. (a) $E_{\text{ionisation}} = E_\infty - E_n = \frac{13.6 Z_{eff}^2}{n^2} eV$

$$= \left[\frac{13.6 Z^2}{n_2^2} - \frac{13.6 Z^2}{n_1^2}\right]$$

$$E = h\nu = \frac{13.6 \times 1^2}{(1)^2} - \frac{13.6 \times 1^2}{(4)^2} ; \; h\nu = 13.6 - 0.85$$

$\because \; h = 6.625 \times 10^{-34}$

$$\nu = \frac{13.6 - 0.85}{6.625 \times 10^{-34}} \times 1.6 \times 10^{-19} = 3.08 \times 10^{15} s^{-1}.$$

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

$$\frac{1}{\lambda} = 1.097 \times 10^7 m^{-1}\left[\frac{1}{1^2} - \frac{1}{\infty^2}\right]$$

$\therefore \; \lambda = 91 \times 10^{-9} m$

We know $10^{-9} = 1nm$ So $\lambda = 91nm$

75. (d) $r \propto n^2$

For I[st] orbit $\gamma = 1$

For III[rd] orbit = $\gamma \propto 3^2 = 9$

So it will 9γ.

76. (b) Bohr suggest a formulae to calculate the radius and energy of each orbit and gave the following formulae

$$r_n = \frac{n^2 h^2}{4\pi^2 kme^4 Z}$$

Where except n^2, all other unit are constant so $r_n \propto n^2$.

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

For energy level $(n=2)$

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

78. (a) Energy of ground stage (E_0) = $-13.6 eV$ and energy level = 5

$$E_5 = \frac{-13.6}{n^2} eV = \frac{-13.6}{5^2} = \frac{-13.6}{25} = -0.54 eV.$$

79. (c) Positive charge of an atom is present in nucleus.

81. (a) For $n_4 \rightarrow n_1$, greater transition, greater the energy difference, lesser will be the wavelength.

Dual nature of electron

1. (c) According to de-Broglie equation $\lambda = \frac{h}{mv}$ or $\frac{h}{p}$ or $\frac{h}{mc}$.

4. (b) $\lambda = \frac{h}{p}$ or $\frac{h}{mv}$ or $\frac{h}{mc}$ de-Broglie equation.

5. (c) Emission spectra of different λ accounts for quantisation of energy.

6. (b) According to de-Broglie equation

$$\lambda = \frac{h}{mv}, \; p = mv, \; \lambda = \frac{h}{p}, \; \lambda = \frac{h}{mc}$$

7. (d) According to de-Broglie $\left(\lambda = \frac{h}{mv}\right)$.

8. (a) $\lambda = \frac{h}{mv} = \frac{6.63 \times 10^{-34}}{10^{-3} \times 100} = 6.63 \times 10^{-33} m$

9. (d) $\lambda = \frac{h}{mv}$. For same velocity $\lambda \propto \frac{1}{m}$.

SO_2 molecule has least wavelength because their molecular mass is high.

10. (d) de-Broglie equation is $\lambda = \frac{h}{p}$.

11. (c) Formula for de-Broglie wavelength is

$$\lambda = \frac{h}{p} \text{ or } \lambda = \frac{h}{mv} \Rightarrow eV = \frac{1}{2} mv^2 \text{ or } v = \sqrt{\frac{2eV}{m}}$$

$$\lambda = \frac{h}{\sqrt{2meV}} = \frac{6.62 \times 10^{-34}}{\sqrt{2 \times 9.1 \times 10^{-31} \times 2.8 \times 10^{-23}}}$$

$\lambda = 9.28 \times 10^{-8}$ meter .

12. (c) $\lambda = \frac{h}{p}, \; p = mv$

$$\lambda = \frac{h}{mv} = \frac{6.62 \times 10^{-34}}{9.1 \times 10^{-31} \times 1.2 \times 10^5}$$

$\lambda = 6.626 \times 10^{-9} m$.

13. (b) Mass of the particle $(m) = 10^{-6} kg$ and velocity of the particle $(v) = 10 \, ms^{-1}$

$$\lambda = \frac{h}{mv} = \frac{6.63 \times 10^{-34}}{10^{-6} \times 10} = 6.63 \times 10^{-29} m$$

15. (b) According to de–Broglie

$$\lambda = \frac{h}{mv} = \frac{6.62 \times 10^{-20} erg.\sec}{\frac{2}{6.023 \times 10^{23}} \times 5 \times 10^4 cm/\sec}$$

$= \frac{6.62\times10^{-27}\times6.023\times10^{23}}{2\times5\times10^{4}} cm = 4\times10^{-8} cm = 4 Å$.

16. (c) $\lambda = \frac{h}{mv} = \frac{6.625\times10^{-34}}{0.2\,kg\times\frac{5}{60\times60\,ms^{-1}}} = 10^{-30}\,m$.

17. (c) From de Broglie equation

$\lambda = \frac{h}{mv} = \frac{6.62\times10^{-34}}{0.5\times100} = 1.32\times10^{-35}\,m$.

18. (c) Dual nature of particle was proposed by de-broglie who gave the following equation for the wavelength.

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

19. (b) One percent of the speed of light is

$v = \left(\frac{1}{100}\right)(3.00\times10^{8}\,ms^{-1}) = 3.00\times10^{6} ms^{-1}$

Momentum of the electron $(p) = m\,v$

$= (9.11\times10^{-31}kg)(3.00\times10^{6} ms^{-1})$

$= 2.73\times10^{-24}\,kg\,ms^{-1}$

The de-broglie wavelength of this electron is

$\lambda = \frac{h}{p} = \frac{6.626\times10^{-34}}{2.73\times10^{-24}\,kgms^{-1}}$

$\lambda = 2.424\times10^{-10}\,m$.

20. (a) We know that the correct relationship between wavelength and momentum is $\lambda = \frac{h}{p}$.

Which is given by de-Broglie.

21. (d) De-broglie equation applies to all the material object in motion.

Uncertainty principle and Schrodinger wave equation

1. (b) The uncertainty principle was enunciated by Heisenberg.

2. (b) According to uncertainty principle, the product of uncertainties of the position and momentum, is $\Delta x\times\Delta p \geq h/4\pi$.

5. (c) $\Delta x\times\Delta p = \frac{h}{4\pi}$ is not the correct relation. But correct Heisenberg's uncertainty equation is $\Delta x\times\Delta p \geq \frac{h}{4\pi}$.

7. (b) According to the Heisenberg's uncertainty principle momentum and exact position of an electron can not be determined simultaneously.

8. (d) $\Delta x.\Delta p \geq \frac{h}{4\pi}$, if $\Delta x = 0$ then $\Delta p = \infty$.

12. (c) According to $\Delta x\times\Delta p = \frac{h}{4\pi}$

$\Delta x = \frac{h}{\Delta p\times4\pi} = \frac{6.62\times10^{-34}}{1\times10^{-5}\times4\times3.14} = 5.27\times10^{-30}\,m$.

13. (a) Uncertainty of moving bullet velocity

$\Delta v = \frac{h}{4\pi\times m\times\Delta v} = \frac{6.625\times10^{-34}}{4\times3.14\times.01\times10^{-5}}$

$= 5.2\times10^{-28}\,m/sec$.

14. (b) $\Delta x.\Delta p \geq \frac{h}{4\pi}$ This equation shows Heisenberg's uncertainty principle. According to this principle the product of uncertainty in position and momentum of particle is greater than equal to $\frac{h}{4\pi}$.

15. (d) Spin quantum number does not related with Schrodinger equation because they always show $+1/2$, $-1/2$ value.

16. (b) According to $\Delta x\times m\times\Delta v = \frac{h}{4\pi}$; $\Delta v = \frac{h}{\Delta x\times m\times4\pi}$

$= \frac{6.6\times10^{-34}}{10^{-5}\times0.25\times3.14\times4} = 2.1\times10^{-29}\ m/s$

17. (a) Uncertainity in position $\Delta x = \frac{h}{4\pi\times\Delta p}$

$= \frac{6.63\times10^{-34}}{4\times3.14\times(1\times10^{-5})} = 5.28\times10^{-30}\,m$.

18. (c) Given that mass of electron $= 9.1\times10^{-31}\,kg$

Planck's constant $= 6.63\times10^{-34}\,kg\,m^2 s^{-1}$

By using $\Delta x\times\Delta p = \frac{h}{4\pi}$; $\Delta x\times\Delta v\times m = \frac{h}{4\pi}$

where : Δx = uncertainity in position

Δv = uncertainity in velocity

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

$= \frac{6.63\times10^{-34}}{4\times3.14\times9.1\times10^{-31}} = 5.8\times10^{-5}\,m^2 s^{-1}$.

Quantum number, Electronic configuration and Shape of orbitals

3. (b) The shape of an orbital is given by azimuthal quantum number 'l'.

5. (c) Hund's rule states that pairing of electrons in the orbitals of a subshell (orbitals of equal energy) starts when each of them is singly filled.

6. (c) $1s^2, 2s^2, 2p^6$ represents a noble gas electronic configuration.

7. (c) The electronic configuration of Ag in ground state is $[Kr]4d^{10}5s^1$.

8. (a) n, l and m are related to size, shape and orientation respectively.

9. (a) Electronic configuration of $Rb_{(37)}$ is

$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^1$

So for the valence shell electron $(5s^1)$

$n=5, l=0, m=0, s=+\frac{1}{2}$

10. (a) $3d$ subshell filled with 5 electrons (half-filled) is more stable than that filled with 4 electrons. $1, 4s$ electrons jumps into $3d$ subshell for more sability.

11. (c) In $2p$ – orbital, 2 denotes principal quantum number (n) and p denotes azimuthal quantum number $(l=1)$.

12. (c) Electronic configuration of H^- is $1s^2$. It has 2 electrons in extra nuclear space.

13. (a) The electronic configuration must be $1s^2\ 2s^1$. Hence, the element is lithium $(z=3)$.

14. (a) Principal quantum no. tells about the size of the orbital.

15. (d) An element has the electronic configuration $1s^2, 2s^2\ 2p^6,\ 3s^2\ 3p^2, (Si)$. It's valency electrons are four.

16. (c) The magnetic quantum number specifies orientation of orbitals.

17. (c) If $l=2, m\neq -3.\ =(-e$ to $+e)$.

18. (d) If $n=3$ then $l=0, 1, 2$ but not 3.

20. (c) Atomic number of Cu is $29=(Ar)4s^1 3d^{10}$.

21. (c) The shape of $2p$ orbital is dumb-bell.

22. (a) When the value of $n=2$, then $l=1$ and the value of $m=-1, 0, +1$ *i.e.* 3 values.

23. (c) $Cr_{24}=(Ar)3d^5\ 4s^1$ electronic configuration because half filled orbital are more stable than other orbitals.

24. (d) Kr has zero valency because it contains 8 electrons in outermost shell.

25. (c) 2 electron in the valence shell of calcium $Ca_{20}=(2, 8, 8, 2)$.

27. (b) Value of $l=1$ means the orbital is p (dumb-bell shape).

28. (d) Cr has $[Ar]4s^1 3d^5$ electronic configuration because half filled orbital are more stable than other orbitals.

31. (d) The two electrons will have opposite spins.

33. (c) If $m=-3$, then $l=3$, for this value n must be 4.

34. (d) No. of electrons $=2n^2$ hence no. of orbital $=\frac{2n^2}{2}=n^2$.

35. (d) No. of electrons $=2n^2$ hence no. of orbital $=\frac{2n^2}{2}=n^2$.

36. (c) If $n=3$ then $l=0$ to $n-1$ & $m=-l$ to $+l$

37. (b) $Na_{11}=2, 8, 1=1s^2, 2s^2\ 2p^6,\ 3s^1$

$n=3, l=0, m=0, s=+1/2$

38. (b) Hund's rule states that pairing of electrons in the orbitals of a subshell (orbitals of equal energy) starts when each of them is singly filled.

39. (d) As a result of attraction, some energy is released. So at infinite distance from the nucleus energy of any electron will be maximum. For bringing electrons from ∞ to the orbital of any atom some work has to be done be electrons hence it bill loose its energy for doing that work.

40. (c) This space is called nodal space where there is no possibility of oressene of electrons.

41. (d) For s orbital $l=0\ \ m=0$.

42. (c) For M^{th} shell, $n=3$; so maximum no. of electrons in M^{th} shell $=2n^2=2\times 3^2=18$.

43. (c) $m=-l$ to $+l$ including zero.

44. (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)

45. (a) It consists only s orbital which is circular.

46. (a) Hund's rule states that pairing of electrons in the orbitals of a subshell (orbitals of equal energy) starts when each of them is singly filled.

47. (b) If value of l is 2 then $m=-2, -1, 0, +1, +2$.

$m=-l$ to $+l$ including zero.

(5 values of magnetic quantum number)

48. (c) *s, p, d* orbitals present in Fe

$Fe_{26}=1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6, 4s^2\ 3d^6$

50. (b) According to Aufbau rule.

51. (c) $3d$ subshell filled with 5 electrons (half-filled) is more stable than that filled with 4 electrons. $1, 4s$ electrons jumps into $3d$ subshell for more sability.

52. (b) $K_{19}=1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6, 4s^1$

for $4s^1$ electrons.

$n=4, l=0, m=0$ and $s=+\frac{1}{2}$.

54. (b) $3d$ subshell filled with 5 electrons (half-filled) is more stable than that filled with 4 electrons. $1, 4s$ electrons jumps into $3d$ subshell for more sability.

55. (c) It has 3 orbitals p_x, p_y, p_z.

57. (b) If $l = 2$ then it must be d orbital which can have 10 electrons.

59. (c) for d orbital $l = 2$.

60. (c) $m = -l$ to $+l$ including zero.

61. (d) When $n = 3$ shell, the orbitals are $n^2 = 3^2 = 9$.

No. of electrons $= 2n^2$

Hence no. of orbital $= \frac{2n^2}{2} = n^2$.

62. (d) Configuration of $Ne = 1s^2 2s^2 2p^6$

$F^- = 1s^2 2s^2 2p^6$

$Na^+ = 1s^2 2s^2 2p^6$

$Mg^{++} = 1s^2 2s^2 2p^6$

$Cl^- = 1s^2 2s^2 2p^6 3s^2 3p^6$.

63. (d) $Unh_{106} = [Rn] 5f^{14}, 6d^5, 7s^1$

64. (c) K^+ and Ca^{++} have the same electronic configuration $(1s^2, 2s^2 2p^6, 3s^2 3p^6)$

65. (b) For s-orbital, $l = 0$.

66. (d) $3s^1$ is valency electrons of Na for this $n = 3, l = 0, m = 0, s = \frac{+1}{2}$

67. (c) $_7N = 1s^2, 2s^2 2p_x^1, 2p_y^1, 2p_z^1$. Hund's rule states that pairing of electrons in the orbitals of a subshell (orbitals of equal energy) starts when each of them is singly filled.

68. (d) (4) and (5) belong to d-orbital which are of same energy.

69. (c) Atomic no. 17 is of chlorine.

70. (b) The s-orbital has spherical shape due to its non- directional nature.

71. (a) According to the Aufbau's principle the new electron will enter in those orbital which have least energy. So here $4p$-orbital has least energy then the others.

72. (c) According to Aufbau's principle.

73. (c) $1s^2 2s^2 2p^6, 3s^2 3p^6, 4s^2 3d^6 = 2, 8, 14, 2$.

74. (c) Ground state of $Cu^{29} = 1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$

$Cu^{2+} = 1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^9$.

76. (c) No. of electrons in 3^{rd} shell $= 2n^2 = 2(3)^2 = 18$

77. (c) $F_9 = 1s^2 2s^2 2p^5$

78. (c) When $l = 3$ then $m = -3, -2, -1, 0, +1, +2, +3$. $m = -l$ to $+l$ including zero.

80. (d) $m = -1$ is not possible for s orbital $(l = 0)$.

84. (a) Both $2p$ and $3p$-orbitals have dumb-bell shape.

85. (b) $3d$ subshell filled with 5 electrons (half-filled) is more stable than that filled with 4 electrons. $1, 4s$ electrons jumps into $3d$ subshell for more sability.

86. (c) The shape of $2p$ orbital is dumb-bell.

87. (a) $_{25}Mn = [Ar] 3d^5 4s^2 \xrightleftharpoons[-2]{} Mn^{2+} = [Ar]\, 3d^5\, 4s^0$

89. (c) For p-orbital, $l = 1$ means dumb-bell shape.

91. (d) $l = 3$ means f subshell maximum number of e^- in f subshell = 14.

93. (b) As per Aufbau principle.

94. (b) $l = 0$ is s, $l = 1$ is p and $l = 2$ is d and so on hence spd may be used in state of no..

95. (d) For $4d, n = 4, l = 2, m = -2, -1, 0, +1, +2,\ s = +\frac{1}{2}$.

96. (d) m cannot be greater than $l (= 0, 1)$.

97. (a) For $n = 1, l = 0$.

99. (d) $Na_{11} = 1s^2 2s^2 p^6 3s^2$

$n = 3, l = 0, m = 0$ and $s = +\frac{1}{2}$.

102. (d) According to Aufbau's rule.

105. (d) $2p_x, 2p_y, 2p_z$ sets of orbital is degenerate.

106. (a) Mg_{12} have $1s^2 2s^2 2p^6 3s^2$ electronic configuration

$n = 3,\ l = 0,\ m = 0,\ s = -\frac{1}{2}$.

107. (c) The principle quantum number $n = 3$. Then azimuthal quantum number $l = 3$ and number of orbitals $= n^2 = 3^2 = 9$. 3 and 9

108. (d) $_{29}Cu = [Ar] 3d^{10} 4s^1, Cu^{2+} = [Ar] 3d^9 . 4s^0$.

Ground state of $Cu^{29} = 1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$

$Cu^{2+} = 1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^9$.

110. (d) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$ it shows electronic configuration of Iron.

111. (d) Orbitals are $4s, 3s, 3p$ and $3d$. Out of these $3d$ has highest energy.

113. (c) For the $n = 2$ energy level orbitals of all kinds are possible $2^n, 2^2 = 4$.

114. (b) $n = 2$ than no. of orbitals $= n^2$, $2^2 = 4$

118. (b) For both A & B electrons $s = -1/2$ & $+1/2$ respectively, $n = 3, l = 0, m = 0$

119. (a) According to Aufbau's rule.

120. (a) Possible number of subshells would be (6*s*, 5*p*, 4*d*).

121. (d) For f orbital $l = 3$.

123. (b) 4*d*-orbital have highest energy in given data.

125. (d) If $m = -3, l = 3$ and $n = 4$.

127. (b) $N_7^{14} = 1s^2 2s^2 2p_x^1 2p_y^1 2p_z^1$.

128. (c) *m* can't be greater than *l*.

130. (b) $n = 1$ and $m = 1$ not possible for *s*-orbitals.

131. (a) $Fe_{26} = [Ar]3d^6 4s^2$

$Fe^{3+} = [Ar]3d^5 4s^0$.

132. (c) Maximum number of electron
$= 2n^2 (\text{where } n = 4) = 2 \times 4^2 = 32$.

133. (d) When 2*p* orbital is completely filled then electron enter in the 3*s*. The capacity of 2*p* orbital containing e^- is 6. So $1s^2, 2s^2 2p^2 3s^1$ is a wrong electronic configuration the write is $1s^2 2s^2 2p^3$.

134. (b) This electronic configuration is *Cr* (chromium element) in the ground state

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

137. (c) No. of electron are same (18) in Cl^- and Ar.

138. (c) For *s*-subshell $l = 0$ then should be $m = 0$.

139. (c) 19th electron of chromium is $4s^1$

$n = 4, l = 0, m = 0, s = +\frac{1}{2}$

140. (c) The value of *m* is – *l to* l including zero so for *l* = 3, *m* would be –3, –2, –1, 0, +1, +2, +3.

141. (c) $l = 1$ is for p orbital.

142. (d) Magnetic quantum number of sodium $(3s^1)$ final electron is *m* = 0.

143. (c) Generally azimuthal quantum number defines angular momentum.

146. (d) $m = (2l + 1)$ for *d* orbital $l = 2$ $m = (2 \times 2 + 1) = 5$.

147. (a) The atomic number of chlorine is 17 its configuration is $1s^2 2s^2 2p^6 3s^2 3p^5$

148. (c)

n	l	m_1	m_2
3	2	1	0

This set (c) is not possible because spin quantum number values $= \pm \frac{1}{2}$.

149. (b) The ground state of neon is $1s^2 2s^2 2p^6$ on excitation an electron from $2p$ jumps to $3s$ orbital. The excited neon configuration is $1s^2 2s^2 2p^5 3s^1$.

152. (a)

	s	*p*	*d*	*f*	*g*	*h*
$l =$	0	1	2	3	4	5

Number of orbitals $= 5 \times 2 + 1 = 11$

153. (a) It is the ground state configuration of chromium.

155. (b) $n = 4 \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^6, 4d^{10}, 4f^{14}$

So $l = (n-1) = 4 - 1 = 3$ which is *f* orbit contain 7 orbital.

156. (d) 2*p* have contain maximum 6 electron out of which there are 3 are of + 1/2 spin and 3 are of – 1/2 spin

↑↓	↑↓	↑↓

↑ +1/2 ↓ –1/2

157. (a) For 4*f* orbital electron, $n = 4$

$l = 3$ (Because 0, 1, 2, 3)

s, p, d, f

m = + 3, + 2, +1, 0, – 1, – 2, – 3

s = +1/2

158. (b) $_{24}Cr \rightarrow 1s^2, 2s^2, \underset{l=1}{2p^6}, 3s^2, \underset{l=1}{3p^6}, \underset{l=2}{3d^5}, 4s^1$

(We know that for *p* the value of $l = 1$ and for d, $l = 2$)

For $l = 1$ total number of electron = 12

For $l = 2$ total number of electron = 5.

159. (c) Atomic number of potassium is 19 and hence electronic configuration will be $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1$

Hence for $4s^1$ electron value of Quantum number are

Principal quantum number $n = 4$

Azimuthal quantum number $l = 0$

Magnetic quantum number $m = 0$

Spin quantum number $s = +1/2$

160. (d) According to Hund's rule electron first fill in unpaired form in vacant orbital then fill in paired form to stabilized the molecule by which $1s^2, 2s^2, 2p_x^2$ is not possible. According to Hund's rule. Because $2p_x, p_y, p_z$ have the same energy level so electron first fill in unpaired form not in paired form so it should be $1s^2, 2s^2, 2p_x^1, 2p_y^1$.

161. (c) It is governed by Aufbau principle.

162. (d) The electronic configuration of atomic number

$24 = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^5, 4s^1$

163. (b) The maximum number of electron in any orbital is 2.

164. (c) According to pauli principle 2 electron does not have the same value of all four quantum number. They have maximum same value are 3.

165. (a) Number of orbitals $= n^2 = 4^2 = 16$.

166. (d) We know from the Aufbau principle, that $2p$ orbital will be filled before $3s$ orbital. Therefore, the electronic configuration $1s^2, 2s^2, 2p^2, 3s^1$ is not possible.

167. (d) Each orbital may have two electrons with opposite spin.

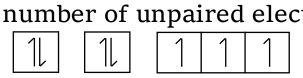
168. (d) Maximum no. of electrons in a subshell $= 2(2l+1)$ for f-subshell, l = 3 so 14 electrons accommodated in f -subshell.

169. (b) Each orbital has atleast two electron.

170. (a) Nucleus of 20 protons atom having 20 electrons.

174. (b) For $m = 0$, electron must be in s-orbital.

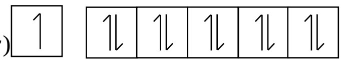
176. (c) In this type of electronic configuration the number of unpaired electrons are 3.

 = 3

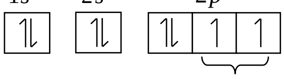
1s 2s 2p

177. (a) Atomic number of Cu is 29 so number of unpaired electrons is 1

$4s^1$ $3d^{10}$

$Cu = (Ar)$ 

$1s^2$ $2s^2$ $2p^4$

178. (b) O_8 = 

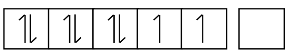
Unpaired electron

181. (c) $Be_4 = 1s^2, 2s^2 =$ (Ground state)

Number of unpaired electrons in the ground state of Beryllium atom is zero.

182. (b) Two unpaired electrons are present in

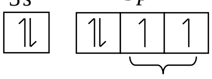
$Ni^{++}(z = 28)$ cation

$Ni_{28}^{++} = [Ar]$ 

$3d^8$ $4s^0$

183. (c) $O_2 = 1s^2\, 2s^2\, 2p^6 3s^2\, 3p^4$

$3s^2$ $3p^4$



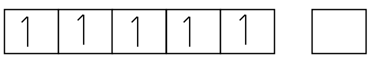
2 Unpaired electrons

184. (c) $Cr_{24} = (Ar)3d^5\, 4s^1$ but $Cr_{24}^{3+} = (Ar)3d^3\, 4s^0$

185. (a) $Zn_{30} = [Ar]3d^{10}\, 4s^2$

$Zn^{++} = [Ar]3d^{10}\, 4s^0$

186. (d) Mn^{+2} ion will have five (maximum) unpaired electrons

$[Ar]$ 

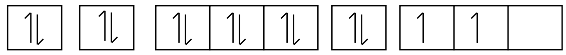
$3d^5$ 4

187. (c) Fe^{3+} ion will have five (maximum) unpaired electrons.

190. (c) Due to full filled d-orbital Cl^- has spherical symmetry.

191. (b) Atomic number 14 leaving 2 unpaired electron

$_{14}Si = 1s^2 2s^2 2p^6 3s^2 3p^2$



1s 2s 2p 3s 3p

192. (a) Shell = K, L, M = $1s^2$ $2s^2 2p^6$ $3s^2 3p^4$

Hence the number of s electron is 6 in that element.

193. (d) $C_6 = 1s^2, 2s^2 2p^2$ (Ground state)

$= 1s^2\, 2s^1\, 2P_x^1\, 2p_y^1\, 2p_z^1$ (Excited state)

In excited state no. of unpaired electron is 4.

194. (b) Max. no. of electrons in N-shell $(n = 4)$

$= 2n^2 = 2 \times 4^2 = 32$.

195. (d) $_{26}Fe = [Ar]3d^6, 4s^2$

$Fe^{2+} = [Ar]3d^6, 4s^0$

Number of d-electrons = 6

$_{17}Cl = [Ne]3s^2, 3p^5$

$Cl^- = [Ne]3s^2, 3p^6$

Number of p-electrons = 6.

196. (a) Electrons in the atom $= 18 + 4 + 3 = 25$ *i.e.* $Z = 25$.

197. (c) The atomic number of bromine is 35 and the electronic configuration of Br is

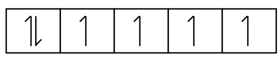
$Br_{35} = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^5$

total electron present in p-orbitals of Br is –

$2p^6 + 3p^6 + 4p^5 = 17$.

198. (d) Fe^{2+} has $1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^6$ configuration with 4 unpaired electron.

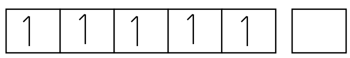
199. (b) $Fe^{2+}\, [Ar]3d^6\, 4s^0$

 = 4

Fe^{2+} consist of maximum 4 unpaired electrons.

201. (a) Fe^{3+} $(z = 26)$

$Fe^{3+} = [Ar]3d^5\, 4s^0$

 = 5

3d 4s

Total no. of unpaired electron=5

202. (b) $Co_{27} = [Ar]\,3d^7 4s^2$

$3d^7$

⇅	⇅	↑	↑	↑

3 unpaired electron are present in cobalt metal.

203. (b) According to Hund's rule, the pairing of electrons will not occur in any orbital of a subshell unit and unless, all the available of it have one electron each.

Electronic configuration of

${}_7N^{14} = 1s^2, 2s^2, 2p_x{}^1 2p_y{}^1 2p_z{}^1$

Hence it has 3 unpaired electron in 2*p*-orbital.

204. (c) $2s$ orbital have minimum energy and generally electron filling increases order of energy according to the Aufbau's principle.

205. (d) According to Pauli's exclusion principle no two electrons in the same atom can have all the set of four quantum numbers identical.

206. (b) The second principal shell contains four orbitals *viz* $2s, 2p_x, 2p_y$ and $2p_z$.

207. (b) Follow Hund's multiplicity rules.

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

210. (b) No two electrons in an atom can have identical set of all the four quantum numbers.

212. (a) In particular shell, the energy of atomic orbital increases with the value of l.

214. (c) Aufbau principle explains the sequence of filling of orbitals in increasing order of energy.

215. (a) According to Aufbau principle electron are filling increasing order of energy. Therefore the electronic configuration $1s^2\,2s^2\,2p^6$ obeys Aufbau principle.

216. (d) Electronic configuration of the Cr_{24} is

$[Ar]4s^1 3d^5$ or

↑	↑	↑	↑	↑		↑	= 6
		3*d*				4*s*	

217. (b) According to the Aufbau principle electron filling minimum to higher energy level.

219. (b) According to Aufbau principle electron are filled in various atomic orbital in the increasing order of energy

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

220. (d) According to Aufbau's rule.

222. (b) We know that for *d*-electron $l = 2$.

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

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

223. (a) Number of nodal centre for $2s$ orbitals $(n-1) = 2-1 = 1$.

224. (d) Since *s*-orbital have $l = 0$

Angular momentum = $\sqrt{l(l+1)} \times \frac{h}{2\pi}$ = $0 \times \frac{h}{2\pi} = 0$.

225. (d) Azimuthal quantum number (l) = 3 shows the presence of f orbit, which contain seven orbitals and each orbital have 2 electrons. Hence $7 \times 2 = 14$ electrons.

227. (b) According to Aufbau principle.

228. (a) Atomic number of deuterium = 1; ${}_1D^2 \rightarrow 1s^1$

Critical Thinking Questions

1. (a) F^- have the same number of electrons with the neon atom.

2. (d) No change by doubling mass of electrons however by reducing mass of neutron to half total atomic mass becomes $6+3$ instead of $6+6$. Thus reduced by 25%.

3. (d) $\frac{e}{m}$ for (i) neutron $= \frac{0}{1} = 0$

(ii) $\alpha -$ particle $= \frac{2}{4} = 0.5$

(iii) Proton $= \frac{1}{1} = 1$

(iv) electron $= \frac{1}{1/1837} = 1837$.

4. (a) Metal is ${}_{56}M^{2+}(2,8,14)$ than $n = A - Z$

$= 56 - 26 = 30$.

5. (d) $E = h\nu = h\frac{c}{\lambda}$ *i.e.* $E \propto \frac{1}{\lambda}$

$\frac{E_1}{E_2} = \frac{\lambda_2}{\lambda_1} = \frac{4000}{2000} = 2.$

6. (c) Rutherford discovered nucleus.

7. (b) According to Bohr's model $\Delta E = E_1 - E_3$

$= 2.179 \times 10^{-11} - \frac{2.179 \times 10^{11}}{9}$

$= \frac{8}{9} \times 2.179 \times 10^{-11} = 1.91 \times 10^{-11} = 0.191 \times 10^{-10}\,erg$

Since electron is going from $n = 1$ to $n = 3$ hence energy is absorbed.

8. (d) Radius of nucleus $= 1.25 \times 10^{-13} \times A^{1/3}\,cm$

$= 1.25 \times 10^{-13} \times 64^{1/3} = 5 \times 10^{-13}\ cm$

Radius of atom = $1Å = 10^{-8} cm$.

$\frac{\text{Volume of nucleus}}{\text{Volume of atom}} = \frac{(4/3)\pi(5\times10^{-13})^3}{(4/3)\pi(10^{-8})^3}$

$= 1.25 \times 10^{-13}$.

9. (a) Values of energy in the excited state $= -\frac{13.6}{n^2} eV = \frac{-13.6}{4} = -3.4\ eV$ in which $n = 2, 3, 4$ *etc.*

10. (c) $E_{1\ He^+} = E_{1\ H} \times z^2$

$-871.6 \times 10^{-20} = E_{1H} \times 4$

$E_{1\ H} = -217.9 \times 10^{-20}\ J$

11. (a) 42*g* of N_3^- ions have $16\ N_A$ valence electrons

4.2*g* of N_3^- ion have $= \frac{16\ N_A}{42} \times 4.2 = 1.6\ N_A$.

12. (d) Ist excited state means $n = 2$

$r = r_0 \times 2^2 = 0.53 \times 4 = 2.12$ Å

13. (d) Frequency $\nu = 12 \times 10^{14} s^{-1}$ and velocity of light $c = 3 \times 10^{10} cm\ s^{-1}$. We know that the wave number $\bar{\nu} = \frac{\nu}{c} = \frac{12 \times 10^{14}}{3 \times 10^{10}} = 4 \times 10^4 cm^{-1}$

14. (c) The last line in any series is called series limit. Series limit for Balmer series is 3646 Å.

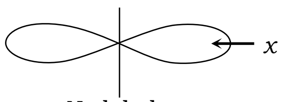
15. (b) $E = \frac{-13.6}{n^2} = \frac{-13.6}{4} = -3.4\ eV$

We know that energy required for excitation $\Delta E = E_2 - E_1 = -3.4 - (-13.6) = 10.2\ eV$

Therefore energy required for excitation of electron per atom $= \frac{10.2}{6.02 \times 10^{23}} = 1.69 \times 10^{-23}\ J$

17. (a) The number of nodal plane are present in a p_x is one or no. of nodal place = l

for p_x orbital l = 1



Nodal plane

18. (b) In Balmer series of hydrogen atomic spectrum which electronic transition causes third line $O \rightarrow L$, $n_2 = 5 \rightarrow n_1 = 2$

20. (b) $\bar{\nu} = \frac{1}{\lambda} = R_H \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$

$= \frac{1}{\lambda} = R_H \left[\frac{1}{3^2} - \frac{1}{n_2^2}\right] = n_2 = 3$ for Paschen series.

21. (a) $E \propto \left[\frac{1}{n_2^2} - \frac{1}{n_1^2}\right]$

23. (d) $\lambda = \frac{c}{\nu} = \frac{3 \times 10^8}{8 \times 10^{15}} = 3.75 \times 10^{-8}$

$= 3.75 \times 10^{-8} \times 10^9 nm = 4 \times 10^1 nm$.

Assertion & Reason

1. (d) The assertion is false but the reason is true exact position and exact momentum of an electron can never be determined as according to Hesenberg's uncertainity principle even with the help of electron microscope because when e^- beam of electron microscope strikes the target e^- of atom, the impact causes the change in velocity of e^- thus attempt to locate the e^- changes ultimately, the momentum & position of e^-.

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

2. (d) Both assertion and reason are false. $2p_x$ and $2p_y$ orbitals are degenerate orbitals, i.e., they are of equal energy and hence no possibility of transition of electron.

3. (a) We know that principal quantum number represent the main energy level or energy shell. Since each energy level is associated with a definite amount of energy, this quantum number determines to a large extent te energy of an electron. It also determines the average distance of an electron around the nucleus. Therefore both Assertion and Reason are true and the Reason is a correct explanation of the Assertion.

4. (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 passes extra stability.

5. (c) The assertion that the isobars are the atoms of different elements having same mass number but different atomic number, is correct but reason is false because atomic mass is sum of number of neutron and protons which should be same for isobars.

6. (d) We know from the Pauli exclusion principle, that two electrons in the same atom can not have same value of all four quantum numbers. This means each electron in an atom has only one set of values for n, l, m and s. Therefore both the Assertion and Reason are false.

7. (e) We know that the line in Balmer series of hydrogen spectrum the highest wavelenght or lowest energy is between $n_1 = 2$ and $n_2 = 3$. And for Balmer series of hydrogen spectrum, the value of $n_1 = 2$ and $n_2 = 3,4,5$. Therefore the Assertion is false but the Reason is true.

8. (d) We know that Absorption spectrum is produced when white light is passed through a substance and transmitted light is analysed by a spectrograph. The dark spaces corresponds to the light radiation absorbed by the substance. And emission spectrum is produced by analysing the radiant energy emitted by an excited substance by a spectrograph. Thus discontinuous spectra consisting of a series of sharp lines and separated by dark bands are obtained. Therefore both the Assertion and Reason are false.

9. (a) We know that a resonance hybrid or the actual molecule is always more stable than any of its canonical structures which is also called hypothetical or imaginary structures. This stability is due to delocalization of electrons and is measured in terms of resonance energy or delocalization energy, it is defined as the difference in internal energy of the resonance hybrid and the most stable canonical structure. Therefore both the Assertion and Reason are true and the Reason is a correct explantion of the Assertion.

10. (e) We know that cathode rays cast shadows of solid objects placed in their path. During experiment performed on these rays, fluorescene (flash of light) is observed in the region, outside the shadow. This shows that cathode rays travel in straight lines. We also known that cathode rays penetrate through a thin sheet of metals but are stopped by thick sheets. Therefore both Assertion and Reason are false.

11. (b) We know that electrons are revolving around the nucleus at high speed in circular paths. The centrifugal force (which arises due to rotation of electrons) acting outwards, balances the electrostatic force of attraction (which arises due to attraction between electrons and nucleus). This prevent the electron from falling into the nucleus. We also know that Rutherford's model of atom is comparable to the "solar system". The nucleeus represent the sun whereas revolving electrons represent the planets revolving around the sun. Thus revolving electron are also called planetary electrons. Therefore both Assertion and Reason are true but Reason is not a correct explanation of Assertion.

12. (c) Assertion is true but Reason is false. Threshold frequency is a minimum frequency required for the emission of electrons from the metal surface.

13. (a) Both assertion and reason are true and reason is the correct explanation of assertion.

Radius, $r = \frac{n^2h^2}{4\pi e^2 mZ} = \frac{n^2}{Z} \times 0.529\,Å . r_n$ also increases indicating a greater separation between the orbit and the nucleus.

14. (d) Both assertion and Reason are false. Only s-orbital is spherically symmertrical. Shape of different d orbitals is as below.

15. (c) Assertion is true but reason is false. Spin angular momentum of the electron, a vector quantity, can have two orientations (represented by + and – sign) relative to a chosen axis. These two orientation are distinguished by the spin quantum number m_s equals to $+\frac{1}{2}$ or $-\frac{1}{2}$. These are called the two spin states of the electron and are normaly represented by the two arrows $\uparrow$ (spin up) and $\downarrow$ (spin down) respectively.

16. (d) Both assertion and reason are false. Total number of orbitals associated with Principal quantum number $n = 3$ is 9. One $3s$ orbital + three $3p$ orbital + five $3d$ orbitals. $\therefore$ Therefore there are a total number of nine orbitals. Number of orbitals in a shell equals to n^2.

17. (c) Assertion is true but reason is false. The order $1s < 2s = 2p < 3s = 3p = 3d < ....$ is true for the energy of an electron in a hydrogen atom and is solely determined by Principal quantum number. For multielectron system energy also depends on azimuthal quantum number. The stability of an electron in a multi electron atom is the net result of the attraction between the electron and the uncleus and the repulsion between the electron and the rest of the electron present. Energies of different subshell (azimuthal quantum number) present within the same principal shell are found to be in order of $s < p < d < f$.

18. (e) Assertion is false but reason is true. Splitting of the spectral lines in the presence of a magnetic field is known as Zeeman effect or in electric field it is known as stark effect. The splitting of spectral lines is due to different orientations which the orbitals can have in the presence of magnetic field.

19. (a) Both assertion and reason are true and reason is the correct explanation of assertion.

20. (e) Assertion is false but reason is true. Atomic orbital is designated by n, l and m_l while state of an electron in an atom is specified by four quantum numbrs n, l, m_l and m_s.

21. (b) Both assertion and reason are true but reason is not the correct explanation of assertion.

The difference between the energies of adjacent energy levels decreases as we move away from the nucleus. Thus in H atom

$E_2 - E_1 > E_3 - E_2 > E_4 - E_3$......

22. (d) Both assertion and reason are false. Cathode rays are stream of electrons. They are generated through gases at low pressure and high voltage.

23. (d) Both assertion and reason are false. In case of isoelectronic, i.e., ions, having the same number of electrons and different nuclear charge, the size decreases with increase in atomic number.

Ion	At. No.	No. of electrons	Ionic radii
Na^+	11	10	0.95Å
Mg^{2+}	12	10	0.65Å
Al^{3+}	13	10	0.50Å

Structure of atom

SET Self Evaluation Test - 2

1. The correct set of quantum numbers for the unpaired electron of chlorine atom is

	n	l	m
(a)	2	1	0
(b)	2	1	1
(c)	3	1	1
(d)	3	0	0

2. The orbital diagram in which the Aufbau's principle is violated is **[IIT 1988; AMU 1999]**

	$2s$	$2p_x$	$2p_y$	$2p_z$
(a)	↑↓	↑↓	↑	
(b)	↑	↑↓	↑	↑
(c)	↑↓	↑	↑	↑
(d)	↑↓	↑↓	↑↓	↑

3. The mass of neutron is nearly

[MNR 1988; UPSEAT 1999, 2000, 02]

(a) $10^{-23}\,kg$ (b) $10^{-24}\,kg$

(c) $10^{-26}\,kg$ (d) $10^{-27}\,kg$

4. Which electronic level would allow the hydrogen atom to absorb a photon but not to emit a photon

[IIT 1984; CPMT 1997]

(a) $3s$ (b) $2p$

(c) $2s$ (d) $1s$

5. Which of the following is not correct for electron distribution in the ground state **[AIIMS 1982]**

		$4s$	$3d$				
(a)	$Co(Ar)$	↑↓	↑↓	↑↓	↑	↑	↑
(b)	$Ni(Ar)$	↑↓	↑↓	↑↓	↑↓	↑	↑
(c)	$Cu(Ar)$	↑↓	↑↓	↑↓	↑↓	↑↓	↑
(d)	$Zn(Ar)$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓

6. If electron, hydrogen, helium and neon nuclei are all moving with the velocity of light, then the wavelengths associated with these particles are in the order **[MP PET 1993]**

(a) Electron > hydrogen > helium > neon

(b) Electron > helium > hydrogen > neon

(c) Electron < hydrogen < helium < neon

(d) Neon < hydrogen < helium < electron

[IIT 1989; MP PET 2004]

7. From the given sets of quantum numbers the one that is inconsistent with the theory is

(a) $n = 3; l = 2; m = -3; s = +1/2$

(b) $n = 4; l = 3; m = 3; s = +1/2$

(c) $n = 2; l = 1; m = 0 l s = -1/2$

(d) $n = 4; l = 3; m = 2; s = +1/2$

8. The uncertainty in the position of an electron (mass = 9.1×10^{-28} g) moving with a velocity of $3.0 \times 10^{4}\,cm\,s^{-1}$ accurate upto 0.001% will be

(Use $\frac{h}{4\pi}$ in the uncertainty expression, where $h = 6.626 \times 10^{-27}\,erg - s$) **[CBSE PMT 1995]**

(a) $1.92\,cm$ (b) $7.68\,cm$

(c) $5.76\,cm$ (d) $3.84\,cm$

9. The orbital angular momentum of an electron in s orbital is

[IIT 1996; AIEEE 2003; MP PET 2004]

(a) $+\frac{1}{2}.\frac{h}{2\pi}$ (b) Zero

(c) $\frac{h}{2\pi}$ (d) $\sqrt{2}.\frac{h}{2\pi}$

10. Values of the four quantum numbers for the last electron in the atom are $n = 4, l = 1, m = +1$ and $s = -1/2$. Atomic number of the atom will be

(a) 22 (b) 32

(c) 33 (d) 36

11. The atomic weight of an element is 39. The number of neutrons in its nucleus is one more than the number of protons. The number of protons, neutrons and electrons respectively in its atom would be **[MP PMT 1997]**

(a) 19, 20, 19 (b) 19, 19, 20

(c) 20, 19, 19 (d) 20, 19, 20

12. The electrons identified by quantum numbers n and l (i) $n = 4, l = 1$ (ii) $n = 4, l = 0$ (iii) $n = 3, l = 2$

(iv) $n=3, l=1$ can be placed in order of increasing energy from the lowest to highest, as

(a) (iv) < (ii) < (iii) <(i)

(b) (ii) < (iv) < (i) < (iii)

(c) (i) < (iii) < (ii) < (iv)

(d) (iii) < (i) < (iv) < (ii)

13. Ground state electronic configuration of nitrogen atom can be represented by

(a)

↑↓	↑↓	↑	↑	↑

(b)

↑↓	↑↓	↑	↓	↑

(c)

↑↓	↑↓	↑	↓	↓

(d)

↑↓	↑↓	↓	↓	↓

14. Which of the following statements (s) is (are) correct

[IIT 1998]

(a) The electronic configuration of Cr is $[Ar]3d^5 4s^1$ (Atomic no. of $Cr = 24$)

(b) The magnetic quantum number may have a negative value

(c) In silver atom, 23 electrons have a spin of one type and 24 of the opposite type (Atomic no. of $Ag = 47$)

(d) The oxidation state of nitrogen in HN_3 is -3

15. The position of both an electron and a helium atom is known within $1.0 nm$ and the momentum of the electron is known within $50 \times 10^{-26} kg\, ms^{-1}$. The minimum uncertainty in the measurement of the momentum of the helium atom is

[CBSE PMT 1998; AIIMS 2001]

(a) $50 kg\, ms^{-1}$ (b) $60\, kg\, ms^{-1}$

(c) $80 \times 10^{-26} kg\, ms^{-1}$ (d) $50 \times 10^{-26} kg\, ms^{-1}$

16. Which of the following pair of orbitals posses two nodal planes **[RPMT 2000]**

(a) $p_{xy}, d_{x^2-y^2}$ (b) d_{xy}, d_{zx}

[IIT 1999]

(c) p_{xy}, d_{zx} (d) $d_{z^2}, d_{x^2-y^2}$

17. The number of atoms in 0.004 *g* of magnesium are

[AFMC 2000]

(a) 4×10^{20} (b) 8×10^{20}

[IIT 1999]

(c) 10^{20} (d) 6.02×10^{20}

18. Which of the following have the same number of unpaired electrons in '*d*' orbitals

(a) *Cr* (b) *Mn*

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

19. The quantum numbers + 1/2 and – 1/2 for the electron spin represent **[IIT Screening 2001]**

(a) Rotation of the electron in clockwise and anticlockwise direction respectively

(b) Rotation of the electron in anticlockwise and clockwise direction respectively

(c) Magnetic moment of the electron pointing up and down respectively

(d) Two quantum mechanical spin states which have no classical analogue

20. The de-Broglie wavelength of a tennis ball of mass 60 *g* moving with a velocity of 10 *metres* per second is approximately **[AIEEE 2003]**

(a) 10^{-33} *metres* (b) 10^{-31} *metres*

(c) 10^{-16} *metres* (d) 10^{-25} *metres*

21. Which of the following are isoelectronic and isostructural $NO_3^-, CO_3^{2-}, ClO_3^-, SO_3$ **[IIT Screening 2003]**

(a) NO_3^-, CO_3^{2-} (b) SO_3, NO_3^-

(c) ClO_3^-, CO_3^{2-} (d) CO_3^{2-}, SO_3

22. The total number of electrons present in all the *s*-orbitals, all the *p*-orbitals and all the *d*-orbitals of cesium ion are respectively

(a) 8, 26, 10 (b) 10, 24, 20

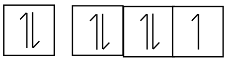
(c) 8, 22, 24 (d) 12, 20, 22

Answers and Solutions

(SET -2)

1. (c) Electronic configuration of Cl is

$3s^2$ $3p$



$[Ne]3s^2 3p^5$ or $[Ne]$ $3p^2{}_x$ $3p^2{}_y$ $3p^1{}_z$

So for the unpaired electron $(3p_z^1)$:

$n = 3, l = 1, m = 1, S = +\frac{1}{2}$

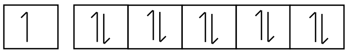
2. (b) According to Aufbau principle the orbitals of lower energy (2*s*) should be fully filled before the filling of orbital of higher energy starts.

3. (d) Mass of neutron $= 1.675 \times 10^{-27}\, kg.$

4. (d) 1*s*-orbital is of lowest energy. Absorption of photon can raise the electron in higher energy state but emission is not possible.

5. (c) The correct electronic configuration

$Cu_{29} = [Ar],\ 4s^1$ $3d^{10}$



6. (a) $\lambda \propto \frac{1}{m}, m_e < m_H < m_{He} < m_{Ne}.$

7. (a) When $l = 2, m \neq -3$.

8. (a) $\Delta p = m \times \Delta v$

$\Delta p = 9.1 \times 10^{-28} \times 3.0 \times 10^4 \times \frac{0.001}{100}$

$\Delta P = 2.73 \times 10^{-24}$

Hence $\Delta x = \frac{h}{\Delta p \times 4\pi} = \frac{6.626 \times 10^{-27}}{2.73 \times 10^{-28} \times 4 \times 3.14}$

$\Delta x = 1.92\, cm.$

9. (b) For 2*s* orbital, l = 0; azimuthal quantum number is not show angular momentum for the $2s$ orbitals.

Angular momentum $= \sqrt{l(l+1)}\frac{h}{2\pi} = 0$.

10. (d) Atomic number is 36 and element is Kr .

11. (a) K_{19}^{39} , $P = 19$, $E = 19$, $N = 20$

12. (a) (*i*) $4p$ (*ii*) $4s$ (*iii*) $3d$ (*iv*) $3p$ order of increasing energy is $3p < 4s < 3d < 4p.$

13. (a,d) According to Hund's principle.

14. (a,b,c) The oxidation state of nitrogen in HN_3 is $-\frac{1}{3}$.

$HN_3:\ 1 + 3x = 0 \Rightarrow 3x = -1$ or $x = \frac{-1}{3}$

15. (d) The product of uncertainties in the position and the momentum of a sub atomic particle $= h/4\pi$. Since Δx is same for electron and helium so Δp must be same for both the particle *i.e.* $50 \times 10^{-26}\, kg\, ms^{-1}$ (given).

16. (b) d_{xy} and d_{zx} has two modal planes.

17. (c) No. of atoms in magnesium

$= \frac{0.004}{24} \times 6.023 \times 10^{23} = 10^{20}$

18. (a,b,c) Cr , Mn and Fe^{3+} have 5 unpaired electron in *d*-orbitals.

$_{24}Cr = 3d^5 4s^1 = 5$

$_{25}Mn = 3d^5 4s^2 = 5$

$_{26}Fe^{3+} = 3d^5 4s^0 = 5$

19. (a,d) Both statement are correct.

20. (a) $\lambda = \frac{h}{mv} = \frac{6.63 \times 10^{-34}}{60 \times 10^{-3} \times 10} = 10^{-33} m$

21. (a) NO_3^- and CO_3^{2-} consist of same electron and show same isostructural.

22. (b) $(Cs_{35}) = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2$

$4p^6, 4d^{10}, 5s^2, 5p^6, 6s^1$

$Cs^+ = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2,$

$4p^6, 4d^{10}, 5s^2, 5p^6$

Total no. of e^- in *s*-orbitals $= 10$

Total no. of e^- in p-ortbitals $= 24$

Total no. of e^- in d-ortbitals $= 20$.

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