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,

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

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.

∂ = stands for an infinitesimal change.

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

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

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. $\Psi = \Psi(x, y, z \ldots\ldots times)$

(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.

Quantum numbers and Shapes of orbitals

Quantum numbers

(1) 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).

(2) **Principle quantum number (n)**

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

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

$$\boxed{r=\frac{n^2}{Z}\times 0.529\ Å}$$

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

$$\boxed{E=-\frac{Z^2}{n^2}\times 313.3\ Kcal\ per\ mole}$$

(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) The value of energy increases with the increasing value of n.

(vii) It represents the major energy shell or orbit to which the electron belongs.

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

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

(3) 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 represents 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 value of 'l' is always equal to the value of 'n'.

(x) The energy of any electron is depend on the value of n & l because total energy = $(n + l)$. The electron enters in that sub orbit whose $(n + l)$ value or the value of energy is less.

(4) 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 proved 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.

(5) Spin quantum numbers (s)

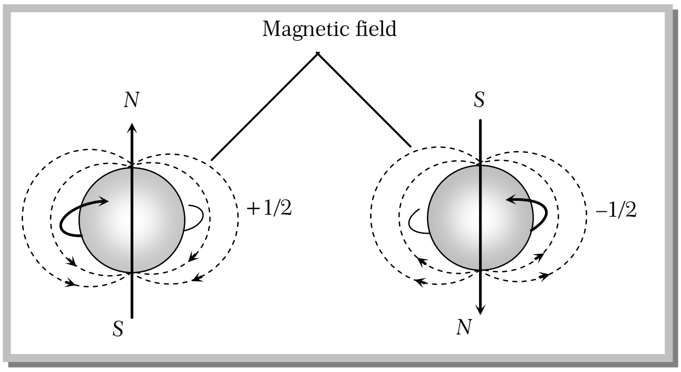
(i) It was proposed by *Goldshmidt & Ulen Back* and denoted by the symbol of 's'.

(ii) The value of 's' is $+1/2$ and $-1/2$, which is 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 Schrödinger equation as solved for H-atom.

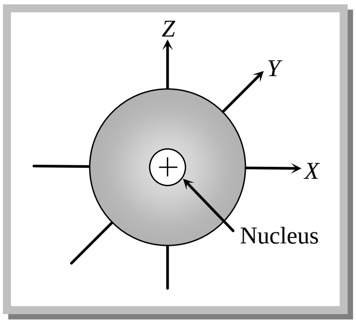
Distribution of electrons among the quantum levels

n	*l*	*m*	*s*	Designation of orbitals	Electrons present	Total no. of electrons
1 (K shell)	0	0	+1/2, −1/2	1s	2	2
2 (L shell)	0	0	+1/2, −1/2	2*s*	2	
		+1	+1/2, −1/2			8
	1	0	+1/2, −1/2	2*p*	6	
		−1	+1/2, −1/2			
3 (M shell)	0	0	+1/2,−1/2	3*s*	2	
		+1	+1/2,−1/2			
	1	0	+1/2,−1/2	3*p*	6	
		−1	+1/2,−1/2			
		+2	+1/2,−1/2			18
		+1	+1/2,−1/2	3*d*		
	2	0	+1/2,−1/2			
		−1	+1/2,−1/2		10	
		−2	+1/2,−1/2			
4(N shell)	0	0	+1/2,−1/2	4*s*	2	
		+1	+1/2,−1/2			
	1	0	+1/2,−1/2	4*p*	6	
		−1	+1/2,−1/2			
		+2	+1/2,−1/2			
		+1	+1/2,−1/2			
	2	0	+1/2,−1/2	4*d*	10	32
		−1	+1/2,−1/2			
		−2	+1/2,−1/2			
		+3				
		+2				
		+1		4*f*	14	
	3	+0				
		−1				

		-2 -3	$\left.\begin{matrix} +1/2,-1/2 \\ +1/2,-1/2 \\ +1/2,-1/2 \\ +1/2,-1/2 \\ +1/2,-1/2 \\ +1/2,-1/2 \\ +1/2,-1/2 \end{matrix}\right]$			

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 $1s < 2s < 3s < 4s$.

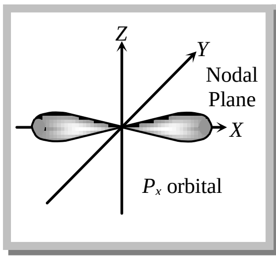
(iii) It does not possess any directional property. *s* orbital has spherical shape.

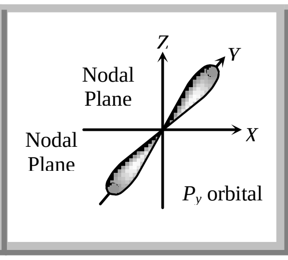
(2) Shape of '*p*' orbitals

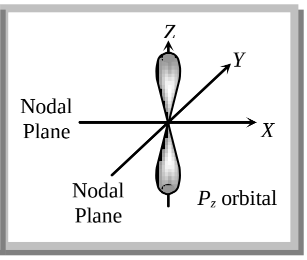
(i) For '*p*' orbital $l=1$, & $m=+1,0,-1$ means there are three '*p*' orbitals, which is symbolised as p_x, p_y, p_z.

(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.







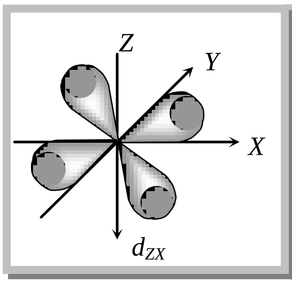
(3) Shape of '*d*' orbital

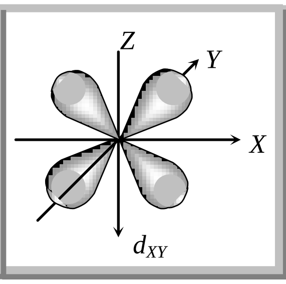
(i) For the 'd' orbital $l=2$ then the values of '*m*' are $-2,-1,0,+1,+2$. It shows that the '*d*' orbitals has five orbitals as $d_{xy}, d_{yz}, d_{zx}, d_{x^2-y^2}, d_{z^2}$.

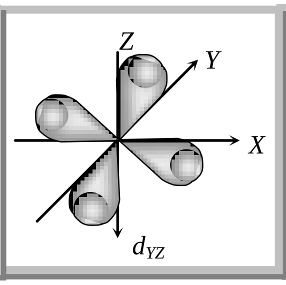
(ii) Each '*d*' orbital identical in shape, size and energy.

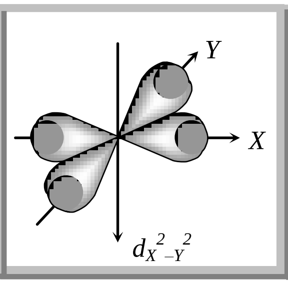
(iii) The shape of *d* orbital is double dumb bell .

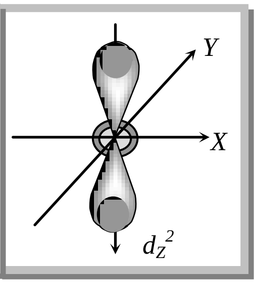
(iv) It has directional properties.











(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^3} and f_{xz^2}.

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

Electronic configuration principles

The distribution of electrons in different orbitals of atom is known as electronic configuration of the atoms.

Filling up of orbitals in the ground state of atom is governed by the following rules:

(1) Aufbau principle

(i) Auf bau is a German word, meaning ‘building up’.

(ii) According to this principle, “In the ground state, the atomic orbitals are filled in order of increasing energies i.e. in the ground state the electrons first occupy the lowest energy orbitals available”.

(iii) In fact the energy of an orbital is determined by the quantum number n and l with the help of $(n+l)$ rule or Bohr Bury rule.

(iv) According to this rule

(a) Lower the value of $n + l$, lower is the energy of the orbital and such an orbital will be filled up first.

(b) When two orbitals have same value of $(n+l)$ the orbital having lower value of “n” has lower energy and such an orbital will be filled up first .

Thus, order of filling up of orbitals is as follows:

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

(2) Pauli’s exclusion principle

(i) According to this principle, “No two electrons in an atom can have same set of all the four quantum numbers n, l, m and s .

(ii) In an atom any two electrons may have three quantum numbers identical but fourth quantum number must be different.

(iii) Since this principle excludes certain possible combinations of quantum numbers for any two electrons in an atom, it was given the name exclusion principle. Its results are as follows :

(a) The maximum capacity of a main energy shell is equal to $2n^2$ electron.

(b) The maximum capacity of a subshell is equal to $2(2l+1)$ electron.

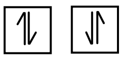
(c) Number of sub-shells in a main energy shell is equal to the value of n.

(d) Number of orbitals in a main energy shell is equal to n^2.

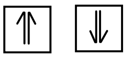
(e) One orbital cannot have more than two electrons.

(iv) According to this principle an orbital can accomodate at the most two electrons with spins opposite to each other. It means that an orbital can have 0, 1, or 2 electron.

(v) If an orbital has two electrons they must be of opposite spin.



Correct



Incorrect

(3) Hund's Rule of maximum multiplicity

(i) This rule provides the basis for filling up of degenerate orbitals of the same sub-shell.

(ii) According to this rule "Electron filling will not take place in orbitals of same energy until all the available orbitals of a given subshell contain one electron each with parallel spin".

(iii) This implies that electron pairing begins with fourth, sixth and eighth electron in p, d and f orbitals of the same subshell respectively.

(iv) The reason behind this rule is related to repulsion between identical charged electron present in the same orbital.

(v) They can minimise the repulsive force between them serves by occupying different orbitals.

(vi) Moreover, according to this principle, the electron entering the different orbitals of subshell have parallel spins. This keep them farther apart and lowers the energy through electron exchange or resonance.

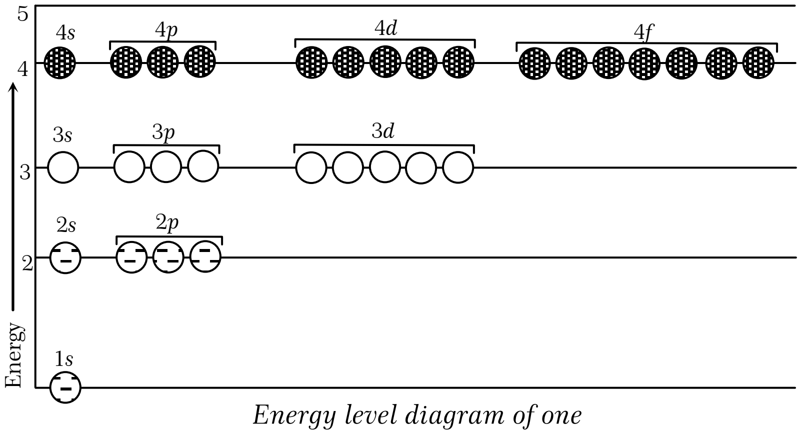
(vii) The term maximum multiplicity means that the total spin of unpaired e^- is maximum in case of correct filling of orbitals as per this rule.

Energy level diagram

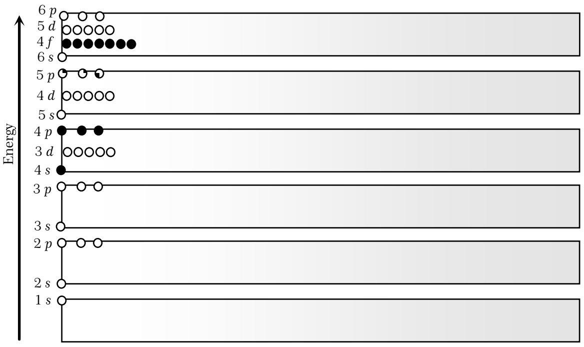
The representation of relative energy levels of various atomic orbital is made in the terms of energy level diagrams.

One electron system: In this system $1s^2$ level and all orbital of same principal quantum number have same energy, which is independent of (l). In this system l only determines the shape of the orbital.

Multiple electron system: The energy levels of such system not only depend upon the nuclear charge but also upon the another electron present in them.



Energy level diagram of one electron system



Energy level diagram of multiple electron system

Diagram of multi-electron atoms reveals the following points :

(i) As the distance of the shell increases from the nucleus, the energy level increases. For example energy level of 2 > 1.

(ii) The different sub shells have different energy levels which possess definite energy. For a definite shell, the subshell having higher value of l possesses higher energy level. For example in 4th shell.

Energy level order $4f > 4d > 4p > 4s$

$l = 3 \quad l = 2 \quad l = 1 \quad l = 0$

(iii) The relative energy of sub shells of different energy shell can be explained in the terms of the $(n+l)$ rule.

(a) The sub-shell with lower values of $(n + l)$ possesses lower energy.

For $3d$ $n = 3$ $l = 2$ $\therefore n + l = 5$

For $4s$ $n = 4$ $l = 0$ $n + l = 4$

(b) If the value of $(n + l)$ for two orbitals is same, one with lower values of 'n' possess lower energy level.

Extra stability of half filled and completely filled orbitals

Half-filled and completely filled sub-shell have extra stability due to the following reasons :

(i) *Symmetry of orbitals*

(a) It is a well kown fact that symmetry leads to stability.

(b) Thus, if the shift of an electron from one orbital to another orbital differing slightly in energy results in the symmetrical electronic configuration. It becomes more stable.

(c) For example p^3, d^5, f^7 configurations are more stable than their near ones.

(ii) *Exchange energy*

(a) The electron in various subshells can exchange their positions, since electron in the same subshell have equal energies.

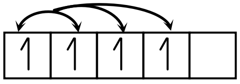
(b) The energy is released during the exchange process with in the same subshell.

(c) In case of half filled and completely filled orbitals, the exchange energy is maximum and is greater than the loss of orbital energy due to the transfer of electron from a higher to a lower sublevel e.g. from 4s to 3d orbitals in case of Cu and Cr .

(d) The greater the number of possible exchanges between the electrons of parallel spins present in the degenerate orbitals, the higher would be the amount of energy released and more will be the stability.

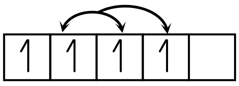
(e) Let us count the number of exchange that are possible in d^4 and d^5 configuraton among electrons with parallel spins.

d^4 (1)



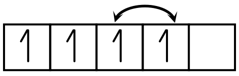
3 exchanges by 1st e^-

(2)



2 exchanges by 2nd e^-

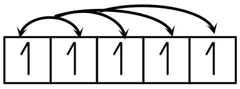
(3)



Only 1 exchange by 3rd e^-

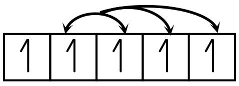
To number of possible exchanges = 3 + 2 + 1 =6

d^5 (1)



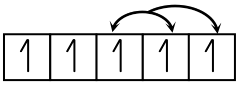
4 exchanges by 1st e^-

(2)



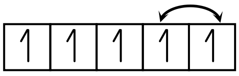
3 exchanges by 2nd e^-

(3)



2 exchange by 3rd e^-

(4)



1 exchange by 4th e^-

To number of possible exchanges = 4 + 3 + 2 +1 = 10

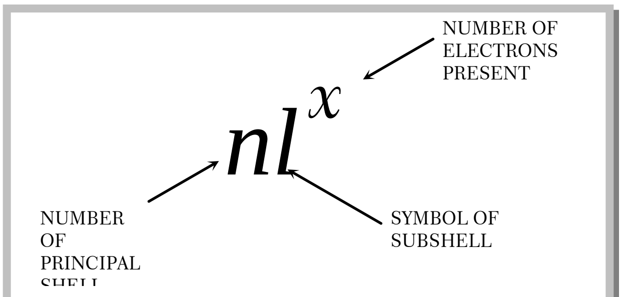
Electronic configurations of Elements

(1) On the basis of the elecronic configuration priciples the electronic configuration of various elements are given in the following table :

Electronic Configuration (E.C.) of Elements Z=1 to 36

Element	Atomic Number	1*s*	2*s*	2*p*	3*s*	3*p*	3*d*	4*s*	4*p*	4*d*	4*f*
H	1	1									
He	2	2									
Li	3	2	1								
Be	4	2	2								
B	5	2	2	1							
C	6	2	2	2							
N	7	2	2	3							
O	8	2	2	4							
F	9	2	2	5							
Ne	10	2	2	6							
Na	11	2	2	6	1						
Mg	12				2						
Al	13				2	1					
Si	14	10			2	2					
P	15	electrons			2	3					
S	16				2	4					
Cl	17				2	5					
Ar	18	2	2	6	2	6					
K	19	2	2	6	2	6		1			
Ca	20							2			
Sc	21						1	2			
Ti	22						2	2			
V	23						3	2			
Cr	24						5	1			
Mn	25						5	2			
Fe	26						6	2			
Co	27	18					7	2			
Ni	28	electrons					8	2			
Cu	29						10	1			
Zn	30						10	2			
Ga	31						10	2	1		
Ge	32						10	2	2		
As	33						10	2	3		
Se	34						10	2	4		
Br	35						10	2	5		
Kr	36	2	2	6	2	6	10	2	6		

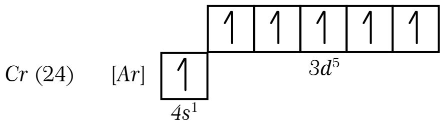
(2) The above method of writing the electronic configurations is quite cumbersome. Hence, usually the electronic configuration of the atom of any element is simply represented by the notation.

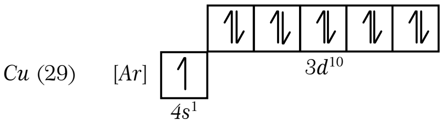


e.g. $1s^2$ means 2 electrons are present in the *s*-subshell of the 1st main shell.

(3) (i) Elements with atomic number 24(*Cr*), 42(*Mo*) and 74(*W*) have $ns^1(n-1)d^5$ configuration and not $ns^2(n-1)d^4$ due to extra stability of these atoms.

(ii) Elements with atomic number 29(*Cu*), 47(*Ag*) and 79(*Au*) have $ns^1(n-1)d^{10}$ configuration instead of $ns^2(n-1)d^9$ due to extra stability of these atoms.





(4) In the formation of ion, electrons of the outer most orbit are lost. Hence, whenever you are required to write electronic configuration of the ion, first write electronic configuration of its atom and take electron from outermost orbit. If we write electronic configuration of $\text{Fe}^{2+}(Z=26,\ 24\ e^-)$, it will not be similar to *Cr* (with $24\ e^-$) but quite different.

$\left.\begin{array}{r} Fe\left[Ar\right]4s^2\,3d^6 \\ Fe^{2+}\left[Ar\right]4s^{\circ}\,3d^6 \end{array}\right\}$ outer most orbit is 4th shell hence, electrons from 4s have been removed to make Fe^{2+}.

(5) Ion/atom will be paramagnetic if there are unpaired electrons. Magnetic moment (spin only) is $\mu=\sqrt{n(n+2)}$ BM (Bohr Magneton). $(1BM=9.27\times10^{-24}\ J/T)$ where *n* is the number of unpaired electrons.

(6) Ion with unpaired electron in *d* or *f* orbital will be coloured. Thus, Cu^+ with electronic configuration $\left[Ar\right]3d^{10}$ is colourless and Cu^{2+} with electronic configuration $\left[Ar\right]3d^9$ (one unpaired electron in 3d) is coloured (blue).

(7) Position of the element in periodic table on the basis of electronic configuration can be determined as,

(i) If last electron enters into *s*-subshell, *p*-subshell, penultimate *d*-subshell and anti penultimate *f*-subshell then the element belongs to *s, p, d* and *f* – block respectively.

(ii) Principle quantum number (*n*) of outermost shell gives the number of period of the element.

(iii) If the last shell contains 1 or 2 electrons (i.e. for *s*-block elements having the configuration ns^{1-2}), the group number is 1 in the first case and 2 in the second case.

(iv) If the last shell contains 3 or more than 3 electrons (i.e. for *p*-block elements having the configuration $ns^2\,np^{1-6}$), the group number is the total number of electrons in the last shell plus 10.

(v) If the electrons are present in the (*n* –1)*d* orbital in addition to those in the *ns* orbital (i.e. for *d*-block elements having the configuration (*n* –1) $d^{1-10}ns^{1-2}$), the group number is equal to the total number of electrons present in the (*n* –1)*d* orbital and *ns* orbital.

EXERCISE

Q.1 X-rays emitted from a copper target and a molybdenum target are found to contain a line of wavelength 22.85 nm attributed to the K_α line of an impurity element. The K_α lines of copper (Z = 29) and molybdenum (Z = 42) have wavelength 15.42 nm and 7.12 nm respectively. Using Moseley's law, $\gamma^{1/2} = a (Z - b)$ calculate the atomic number of the impurity element.

Q.2 Calculate the energy emitted when electrons of 1.0 g atom of hydrogen undergo transition giving the spectral lines of lowest energy in the visible region of its atomic spectra.

Q.3 1.8 g hydrogen atoms are excited to radiations. The study of spectra indicates that 27% of the atoms are in 3rd energy level and 15% of atoms in 2nd energy level and the rest in ground state. If I.P. of H is 21.7×10^{-12} erg. Calculate –

(i) No. of atoms present in III & II energy level.

(ii) Total energy evolved when all the atoms return to ground state.

Q.4 One mole He^+ ions are excited. Spectral analysis showed existence of 50% ions in 3rd orbit, 25% in 2nd and rest in ground state. Calculate total energy evolved when all the ions return to the ground state.

Q.5 The energy of an excited H-atom is –3.4 eV. Calculate angular momentum of e^-.

Q.6 The vapours of Hg absorb some electrons accelerated by a potential diff. of 4.5 volt as a result of which light is emitted. If the full energy of single incident e^- is supposed to be converted into light emitted by single Hg atom, find the wave no. of the light.

Q.7 The hydrogen atom in the ground state is excited by means of monochromatic radiation of wavelength x A^0. The resulting spectrum consists of 15 different lines . Calculate the value of x.

Q.8 The eyes of certain member of the reptile family pass a single visual signal to the brain when the visual receptors are struck by photons of wavelength 850 nm . If a total energy of 3.15×10^{-14} J is required to trip the signal, what is the minimum number of photons that must strike the receptor.

Q.9 If the average life time of an excited state of H atom is of order 10^{-8} sec, estimate how many orbits an e^- makes when it is in the state n = 2 and before it suffers a transition to n =1 state.

Q.10 Calculate the frequency of e^- in the first Bohr orbit in a H-atom.

Q.11 A single electron orbits around a stationary nucleus o f charge +Ze where Z is a constant from the nucleus and e is the magnitude of the electric charge. The hydrogen like species required 47.2 eV to excite the electron from the second Bohr orbit to the third Bohr orbit. Find

(i) the value of Z and give the hydrogen like species formed.

(ii) the kinetic energy and potential energy of the electron in the first Bohr orbit.

Q.12 A stationary He^+ ion emitted a photon corresponding to a first line of the Lyman series. The photon liberated a photon electron from a stationary H atom in ground state. What is the velocity of photoelectron.

Q.13 To what series does the spectral lines of atomic hydrogen belong if its wave number is equal to the difference between the wave numbers of the following two lines of the Balmer series 486.1 and 410.2 nm. What is the wavelength of this.

Q.14 A particle of charge equal to that of an electron and mass 208 times the mass of the electron moves in a circular orbit around a nucleus of charge +3e. Assuming that the Bohr model of the atom is applicable to this system, (a) derive an expression for the radius of the nth bohr orbit, (b) find the value of n for which the radius of the orbit is approximately the same as that of the first Bohr orbit for th ehydrogen atom, and (c) find the wavelength of the radiation emitted when the revolving particle jumps from the third orbit to the first.

Q.15 The ionisation energy of the hydrogen atom is given to be 13.6 eV. A photon falls on a hydrogen atom which is initially in the ground state and excites it to the (n = 4)state.

(a) show this transition in the energy-level diagram &

(b) calculate the wavelength of the photon.

Q.16 Find the number of photons of radiation of frequency 5×10^{13} s^{-1} that must be absorbed in order to melt one gm ice when the latent heat of fusion of ice is 330 J/g.

Q.17 The dye acriflavine, when dissolved in water, has its maximum light absorption at 4530 Å and its maximum fluorescence emission at 5080 Å. The number of fluorescence quanta is, on the average, 53% of the number of quanta absorbed. Using the wavelengths of maximum absorption and emission, what % of absorbed energy is emitted as fluorescence?

Q.18 Hydrogen atom in its ground state is excited by means of monochromatic radiation of wavelength 975Å. How many different lines are possible in the resulting spectrum? Calculate the longest wavelength amongst them.

Q.19 An alpha particle after passing through a potential difference of 2×10^6 volt falls on a silver foil. The atomic number of silver is 47. Calculate (i) the K.E. of the alpha-particle at the time of falling on the foil. (ii) K.E. of the α – particle at a distance of 5×10^{-14}m from the nucleus, (iii) the shortest distance from the nucleus of silver to which the α–particle reaches.

Q.20 Suppose the potential energy between electron and proton at a distance r is given by . Use Bohr's theory to obtain energy of such a hypothetical atom.

Q.21 An energy of 68 eV is required to excite a hydrogen like atom from its second Bohr orbit to the third. The nuclear charge is Ze. Find the value of Z, the kinetic energy of the electron in the first Bohr orbit and the wavelength of the radiation required to eject the electrons from the first Bohr orbit to infinity.

Q.23 What is de Broglie wavelength associated with an e^- accelerated through P.D. = 100 KV.

Q.24 Calculate the de-Broglie wavelength associated with motion of earth (mass 6×10^{24} Kg) orbiting around the sun at a speed of 3×10^6 m/s.

Q.25 A base ball of mass 200 g is moving with velocity 30×10^2 cm/s. If we can locate the base ball with an error equal in magnitude to the λ of the light used (5000 Å), how will the uncertainty in momentum be compared with the total momentum of base ball.

Q.26 An electron has a speed of 40 m/s, accurate up to 99.99%. What is the uncertainty in locating its position?

Answer Key

Q.1	24	Q.2	182.5 KJ				
Q.3	292.68×10^{21} atoms, 162.60×10^{21} atoms, 832.50 KJ					Q.4	331.13×10^4 J
Q.5	h/π	Q.6	3.63×10^6 m^{-1}	Q.7	938 Å	Q.8	1.35×10^5
Q.9	8×10^6	Q.10	6530×10^{12}Hz	Q.11	5 ; 340 ev , – 680 eV		
Q.12	3.09×10^8 cm/sec	Q.13	Brackett ; 2.63×10^{-4} cm				
Q.14	r_n = n = 25 ; 55.2 pm	Q.15	973.5 Å				
Q.16	10^{22}	Q.17	47.26%	Q.18	six , 18800 Å		
Q.19	6.4×10^{-13} J, 2.1×10^{-13}J, 3.4×10^{-14}m			Q.20	E =		
Q.21	6 ; 489.6 eV , 25.28 Å			Q.22	+1/2 , +1/2 , +1/2 , +3/2 and 2,2,2,4		
Q.23	3.88 pm	Q.24	3.68×10^{-65} m				
Q.25	1.75×10^{-29}		Q.26 0.0144 m				