2.Structure of Atom

Some Important Points and Terms of the Chapter

1. The word 'atom' has been derived from the Greek word '**a-tomio**' which means 'uncutable'or 'non-divisible'.
2. J. J. Thomson, in 1898, proposed that an atom possesses a spherical shape (radius approximately 10–10 m) in which the positive charge is uniformly distributed. The electrons are embedded into it in such a manner as to give the most stable electrostatic arrangement (Fig. 2.4, NCERT Page 5). Many different names are given to this model, for example, **plum pudding, raisin pudding or watermelon.**
3. **Rutherford's Nuclear Model of Atom**:
   a) Most of the space in the atom is empty as most of the a–particles passed through the foil undeflected.
   b) A few positively charged a– particles were deflected. The deflection must be due to enormous repulsive force showing that the positive charge of the atom is not spread throughout the atom as Thomson had presumed. The positive charge has to be concentrated in a very small volume that repelled and deflected the positively charged a–particles.
   c) Calculations by Rutherford showed that the volume occupied by the nucleus is negligibly small as compared to the total volume of the atom. The radius of the atom is about 10^{-10} m, while that of nucleus is 10^{-15} m.
   d) On the basis of above observations and conclusions, Rutherford proposed the nuclear model of atom (after the discovery of protons). According to this model :(i) The positive charge and most of the mass of the atom was densely concentrated in extremely small region. This very small portion of the atom was called nucleus by Rutherford. (ii) The nucleus is surrounded by electrons that move around the nucleus with a very high speed in circular paths called orbits. Thus, Rutherford's model of atom resembles the solar system in which the nucleus plays the role of sun and the electrons that of revolving planets. (iii) Electrons and the nucleus are held together by electrostatic forces of attraction.

4. The number of protons present in the nucleus is equal to atomic number (Z).the nucleus is equal to atomic number (Z).i.e. **Atomic number (Z)** = number of protons in the nucleus of an atom = number of electronsin a nuetral atom
5. Protons and neutrons present in the nucleus are collectively known as nucleons. The total number of nucleons is termed as **mass number (A)** of the atom.mass number (A) = number of protons (Z) + number of neutrons (n)
6. **Isobars** are the atoms with same mass number but different atomic number for example, $_6C^{14}$and $_7N^{14}$. On the other hand, atoms with identical atomic number but different atomic mass number are known as **Isotopes**. e.g. $_6C^{14}$ $_6C^{13}$ $_6C^{12}$ & $_{17}Cl^{35}$,$_{17}Cl^{37}$
7. **Drawbacks of Rutherford Model** According to the electromagnetic theory of Maxwell, charged particles when accelerated should emit electromagnetic radiation (This feature does not exist for planets since they are uncharged). Therefore, an electron in an orbit will emit radiation, the energy carried by radiation comes from electronic motion. The orbit will thus continue to shrink. Calculations show that it should take an electron only 10^{-8} s to spiral into the nucleus. But this does not happen. Thus, the Rutherford model cannot explain the stability of an atom.
8. The frequency (ν), wavelength (λ) and velocity of light (c) are related by the equation (2.5). **c = ν λ** The other commonly used quantity specially in spectroscopy, is the wavenumber ($\bar{\nu}$). It is defined as the number of wavelengths per unit length. Its units are reciprocal of wavelength unit, i.e., m^{-1}.
9. H. Hertz performed a very interesting experiment in which electrons (or electric current) were ejected when certain metals (for example potassium, rubidium, caesium etc.) were exposed to a beam of light . The phenomenon is called Photoelectric effect. **For photoelectric effect :** $h\nu = h\nu^0 + 1/2\, m\upsilon^2$
10. **Planck's quantum theory.** *(i)* The energy is radiated or absorbed by a body not continuously but discontinuously in form of small packets.

(ii) Each packet is called quantum. In case of light, the quantum is called 'photon'. The energy of quantum is directly proportional to the frequency (v) of the radiation. $E \propto v$ $E = hv$, Where *'h'* is Planck's constant. Its value is 6.625×10^{-34} Joule second.

11. The spectrum of radiation emitted by a substance that has absorbed energy is called an **emission spectrum**. Atoms, molecules or ions that have absorbed radiation are said to be "excited". To produce an emission spectrum, energy is supplied to a sample by heating it or irradiating it and the wavelength (or frequency) of the radiation emitted, as the sample gives up the absorbed energy, is recorded.

12. An **absorption spectrum** is like the photographic negative of an emission spectrum. A continuum of radiation is passed through a sample which absorbs radiation of certain wavelengths.

13. **Line Spectrum of Hydrogen**: When an electric discharge is passed through gaseous hydrogen, the H_2 molecules dissociate and the energetically excited hydrogen atoms produced emit electromagnetic radiation of discrete frequencies. The hydrogen spectrum consists of several series of lines named after their discoverers.

Series	n_1	n_2	**Spectral region**
Lyman	1	2,3……	Ultraviolet
Balmer	2	3,4………	Visible
Paschen	3	4,5……..	Infrared
Brackett	4	5,6…….	Infrared
Pfund	5	6,7…….	Infrared

The Swedish spectroscopist, Johannes Rydberg, noted that all series of lines in the hydrogen spectrum could be described by the following expression : $\bar{\nu} = 109{,}677 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) \text{ cm}^{-1}$

14. **Bohr's Model For Hydrogen Atom**

  a) The electron in the hydrogen atom can move around the nucleus in a circular path of fixed radius and energy. These paths are called orbits, stationary states or allowed energy states. These orbits are arranged concentrically around the nucleus.

  b) An electron can move only in those orbits for which its angular momentum is integral multiple of h/2∏ that is why only certain fixed orbits are allowed.The angular momentum of an electron in a given stationary state can be expressed as in equation

$$m_e vr = n.\frac{h}{2\pi} \qquad n = 1,2,3...$$

  c) The energy of an electron in the orbit does not change with time. However, the electron will move from a lower stationary state to a higher stationary state when required amount of energy is absorbed by the electron or energy is emitted when electron moves from higher stationary state to lower stationary state . The energy change does not take place in a continuous manner.

  d) The frequency of radiation absorbed or emitted when transition occurs between two stationary states that differ in energy by ΔE, is given by :

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

15. Bohr's theory can also be applied to the ions containing only one electron, similar to that present in hydrogen atom. For example, He^+ Li^{2+}, Be^{3+} and so on. The energies of the stationary states associated with these kinds of ions (also known as hydrogen like species) are given by the

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

and radii by the expression

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

expression

16. **Limitations of Bohr's Model**: It fails to account for the finer details (doublet, that is two closely spaced lines) of the hydrogen atom spectrum observed by using sophisticated spectroscopic techniques. This model is also unable to explain the spectrum of atoms other than hydrogen, for example, helium atom which possesses only two electrons. Further, Bohr's theory was also unable to explain the splitting of spectral lines in the presence of magnetic field (**Zeeman effect**) or an electric field (**Stark effect).**
17. **Dual Behaviour of Matter**: The French physicist, de Broglie in 1924 proposed that matter, like radiation, should also exhibit dual behaviour i.e., both particle and wavelike properties.
18. The **de Broglie relation**. :de Broglie relation state that the wavelength associated with a moving object or an electron is inversely proportional to the momentum of the particle.

    $\lambda = \frac{h}{mv} = \frac{h}{p}$ where p is the momentum of particle = mv.

19. **Heisenberg's Uncertainty Principle**. It is not possible to determine the position and velocity simultaneously for a sub-atomic particle like electron at any given instant to an arbitary degree of precision. Consequently, it is not possible to talk of path of the electron in which it moves. If 'Δx' is uncertainty in position and 'ΔP' is uncertainty in momentum then

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

20. **Orbital.** It is a region or space where there is maximum probability of getting electron.
21. **Quantum numbers.** They are used to get complete information about electron, i.e., location, energy, spin, etc. These quantum numbers also help to designate the electron present in an orbital.
22. **Principal quantum number.** It specifies the location and energy of an electron. It is measure of the effective volume of the electron cloud. It is denoted by '*n*'. Its possible values are 1, 2, 3,4 …..

23. **Angular momentum quantum number.** It is also called 'azimuthal quantum number'. It determines the shape of the orbital. It is denoted by '*l*'. The permitted values of '*l*' are 0, 1, 2, etc., upto n–1. For a given value of *n*, $l = 0$ to $n - 1$ '. e.g., if value of *n* is 4, *l* can have values 0, 1, 2, 3. It determines angular momentum. $mvr = \sqrt{l(l+1)}\frac{h}{2\pi}$

24. **Magnetic quantum number.** It is denoted by '*m*' and its value depends on value of '*l*' since magnetism is due to angular momentum. It determines the magnetic orientation of an orbital, i.e., the direction of orbital relative to magnetic field in which it is placed. Its permitted values are - *l* to + *l* including zero, e.g., when *l* = 1, then *m* = -1, 0, +1. It has total number of values equal to $2l + 1$.

25. **Spin quantum number.** It indicates, the direction in which electron revolves. Spin is magnetic property and is also quantized. It has two permitted values + ½ or – ½. The spin angular momentum of an electron is constant and cannot be changed.

26. **(n+l) rule**: The relative order of energies of various sub-shells in a multi-electron atom can be predicted with the help of (n+l) rule (also called Bohr-Bury rule)According to this rule a sub-shell with lower values of (n+l) has lower energy.In case two sub-shell has equal value of (n+l), the sub-shell with lower value of n has lower energy

27. **Pauli's Exclusion Principle.** No two electrons in an atom can have all the four quantum numbers same. It can also be stated as – An orbital can have maximum two electrons and they must be of opposite spin quantum numbers.

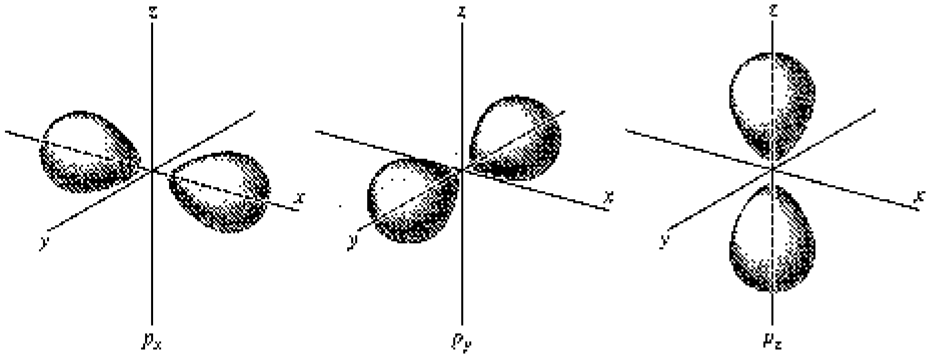
28. **Aufbau principle.** Electrons are filled in the various orbitals in the increasing order of their energies, *i.e.*, orbital having lowest energy will be filled first and the orbital having highest energy will be filled last.**Increasing energy ofatomic orbitals for multi-electron atoms**

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*

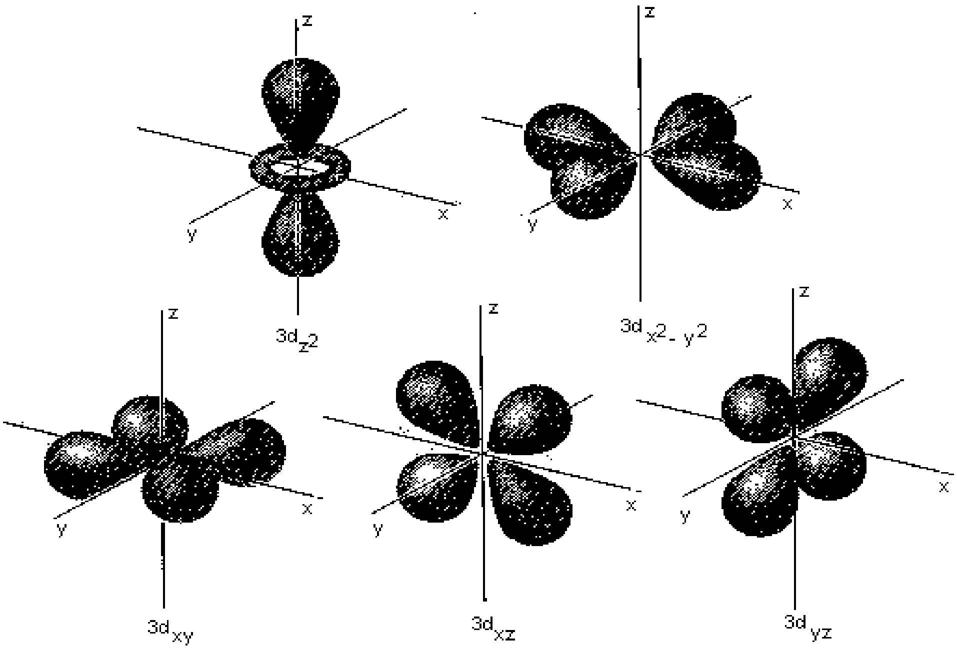
29. **Hund's rule of maximum multiplicity.** No electron pairing takes place in *p, d* and *f* orbitals until each orbital in the given sub-shell contains one electron, *e.g.*, N (7) has electronic configuration $1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$ according to Hund's rule and not $1s^2\ 2s^2\ 2p_x^2\ 2p_y^1$.

30. The valence electronic configurations of Cr and Cu, therefore, are $3d^5\ 4s^1$ and $3d^{10}4s^1$ respectively and not $3d^4\ 4s^2$ and $3d^9 4s^2$. It has been found that there is extra stability(Stability of Completely Filled and Half Filled Subshells) associated with these electronic configurations.

31. **Three orbitals of 2p subshell** ($2p_x$, $2p_y$, and $2p_z$ orbitals).



32. **Five orbitals of 3d subshell** ($3d_{xy}$, $3d_{yz}$, $3d_{zx}$ $3d_{x^2-y^2}$and $3d_{x^2}$ orbitals).



Unit-2

STRUCTURE OF ATOM

1.Questions based on sub-atomic particles (Electron, Protons and Neutrons) atomic models, Thomson Model Of Atom, Rutherford's Nuclear Model Of Atom. Atomic Number,Mass Number, Isobars, Isotopes, Drawbacks Of Rutherford Model.

1. Calculate the number of protons, neutrons and electrons in $_{35}Br^{80}$.
2. The number of electrons, protons and neutrons in a species are equal to 18,16 and 16 respectively. Assign the proper symbol to the species.
3. How many neutrons and protons are there in the following nuclei? $_6C^{13}$,$_8O^{16}$,$_{12}Mg^{24}$,$_{26}Fe^{56}$ $_{38}Sr^{88}$
4. Write the complete symbol for the atom with the given atomic number (Z) and atomic mass (A) (i) Z = 17 , A = 35. (ii) Z = 92 , A = 233. (iii) Z = 4 , A = 9.
5. Which of the following are isoelectronic species i.e., those having the same number of electrons? Na^+, K^+, Mg^{2+}, Ca^{2+}, S^{2-}, Ar.
6. Give the number of electrons in the species H_2^+,H_2 and O_2^+.
7. Calculate the number of electrons which will together weigh one gram.
8. Calculate the mass and charge of one mole of electrons.
9. Compare Electron, Protons and Neutrons.
10. Describe Thomson Model Of Atom.
11. Explain Rutherford's scattering experiment.. What conclusions regarding the structure of atom were drawn by Rutherford on the basis of the observations of experiment? Give the major drawbacks of Rutherford Model.
12. In Rutherford's experiment, generally the thin foil of heavy atoms, like gold, platinum etc. have been used to be bombarded by the a-particles. If the thin foil of light atoms like aluminium etc. is used, what difference would be observed from the above results?
13. Define the terms Atomic Number, Mass Number, Isobars, Isotopes,

ANSWER

1. Calculate the number of protons, neutrons and electrons in $_{35}Br^{80}$.

   Ans: no of protons =no of electrons =35; no of neutrons= 80-35= 45

2. The number of electrons, protons and neutrons in a species are equal to 18,16 and 16 respectively. Assign the proper symbol to the species.

   Ans: $[_{16}S^{32}]^{2-}$

3. How many neutrons and protons are there in the following nuclei?
   a. $_{6}C^{13}$, ans P=6, n=7
   b. $_{8}O^{16}$,ans: P=8, n=8
   c. $_{12}Mg^{24}$, Ans: P=12, n=12
   d. $_{26}Fe^{56}$, Ans: P=26, n=30
   e. $_{38}Sr^{88}$, Ans: P=38, n=50
4. Write the complete symbol for the atom with the given atomic number (Z) and atomic mass (A)
   a. Z = 17 , A = 35.Ans: $_{17}Cl^{35}$
   b. Z = 92 , A = 233.Ans: $_{92}U^{233}$
   c. Z = 4 , A = 9. Ans: $_{4}Be^{9}$

5. Which of the following are isoelectronic species i.e., those having the same number of electrons? Na^{+}, K^{+}, Mg^{2+}, Ca^{2+}, S^{2-}, Ar.

   Ans: i)Na^{+}, Mg^{2+},

   ii)K^{+}, Ca^{2+}, S^{2-}, Ar are isoelectronic.

6. Give the number of electrons in the species H_2^{+},H_2 and O_2^{+}.

   Ans: 1,2,15

7. Calculate the number of electrons which will together weigh one gram.

   Ans: mass of an electron: $9.1X10^{-28}$g;

   No of electrons weighing 1 g= $1/9.1X10^{-28}$ =$1.1X10^{27}$

8. Calculate the mass and charge of one mole of electrons.

Ans: mass of 1 electron: $9.1X10^{-28}$g; mass of 1 mole of electrons: $9.1X10^{-28}$g $X6.022X10^{23}$

$=54.8 X 10^{-5}$g

Charge on an electron= $1.6X 10^{-19}$Col

Charge on one mole of electrons= $1.6 X 10^{-19} X 6.022X10^{23}= 9.635X10^{4}= 96500$ coloumbs = 1faraday

2.Questions based on frequency wavelength wavenumber threshold frequency and work function (W_0), Photoelectric effect, Emission and Absorption Spectra, Line Spectrum of Hydrogen, Bohr's model for hydrogen atom, radii of the stationary states, energy of stationary state, Limitations of Bohr's Model

1. Define the terms frequency wavelength & wave number (Write mathematical forms also).
2. The Vividh Bharati station of All India Radio, Delhi, broadcasts on a frequency of 1,368 kHz (kilo hertz). Calculate the wavelength of the electromagnetic radiation emitted by transmitter. Which part of the electromagnetic spectrum does it belong to?
3. The wavelength range of the visible spectrum extends from violet (400 nm) to red (750 nm). Express these wavelengths in frequencies (Hz). ($1nm = 10^{-9}$ m)
4. Calculate(a) wavenumber and (b) frequency of yellow radiation having wavelength 5800Å.
5. Yellow light emitted from a sodium lamp has a wavelength of 580 nm. Calculate the frequency and wavenumber of the yellow light.
6. Find energy of each of the photons which (i) correspond to light of frequency 3×10^{15}Hz. (ii) have wavelength of 0.50 Å.
7. Calculate the wavelength, frequency and wave number of a light wave whose period is 2.0×10^{-10}s.
8. Electrons are emitted with zero velocity from a metal surface when it is exposed to radiation of wavelength 6800 Å. Calculate threshold frequency. ν_0) and work function (W_0) of the metal.
9. Explain Photoelectric effect, Emission and Absorption Spectra
10. Give the postulates of Bohr's model. also write its Limitations.
11. Write a note on the Spectral Lines for Atomic Hydrogen.
12. What are the frequency and wavelength of a photon emitted during a transition from n = 5 state to the n = 2 state in the hydrogen atom?
13. Calculate the energy associated with the first orbit of He^+ .What is the radius of this orbit?

14. What is the wavelength of light emitted when the electron in a hydrogen atom undergoes transition from an energy level with n = 4 to an energy level with n = 2?

15 How much energy is required to ionise a H atom if the electron occupies n = 5 orbit? Compare your answer with the ionization enthalpy of H atom (energy required to remove the electron from n =1 orbit).

16. What is the maximum number of emission lines when the excited electron of a H atom in n = 6 drops to the ground state?

17.(i) The energy associated with first orbit in the hydrogen atom is -2.18×10^{-18} J atom^{-1}. What is the energy associated with the fifth orbit?

(ii) Calculate the radius of Bohr's fifth orbit for hydrogen atom.

18. Calculate the wavenumber for the longest wavelength transition in the Balmer series of atomic hydrogen.

19. What is the energy in joules, required to shift the electron of the hydrogen atom from the first Bohr orbit to the fifth Bohr orbit and what is the wavelength of the light emitted when the electron returns to the ground state? The ground state electron energy is $-2.18 \text{ X } 10^{-11}$ ergs.

20. The electron energy in hydrogen atom is given by $E_n = (-2.18 \text{ X } 10^{-18})/n^2$J. Calculate the energy required to remove an electron completely from the n = 2 orbit. What is the longest wavelength of light in cm that can be used to cause this transition?

21. What transition in the hydrogen spectrum would have the same wavelength as the Balmer transition n = 4 to n = 2 of He^+ spectrum?

ANSWER

2.The Vividh Bharati station of All India Radio, Delhi, broadcasts on a frequency of 1,368 kHz (kilo hertz). Calculate the wavelength of the electromagnetic radiation emitted by transmitter. Which part of the electromagnetic spectrum does it belong to?

Ans: $\nu = 1.368 \text{ X} 10^6 = c/\lambda \Rightarrow \lambda = 3 \text{ X } 10^8/1.368 \text{ X } 10^6 = 219$m.

3. The wavelength range of the visible spectrum extends from violet (400 nm) to red (750 nm). Express these wavelengths in frequencies (Hz). (1nm = 10^{-9} m)

Ans: $\nu = c/\lambda = 3\text{X}10^8/400\text{X}10^{-9} = 7.5\text{X}10^{14}$hertz

$\nu = c/\lambda = 3\text{X}10^8/750\text{X}10^{-9} = 4\text{X}10^{14}$hertz

4. Calculate(a) wavenumber and (b) frequency of yellow radiation having wavelength 5800Å.

Ans: $\nu = c/\lambda = 3X10^8/5800X10^{-10} = 5.17X\ 10^{14}$hertz

Wave number $(\bar{\nu}) = 1/\lambda = 1/5.8X10^{-7} = 1.72\ X\ 10^6$

5. Yellow light emitted from a sodium lamp has a wavelength of 580 nm. Calculate the frequency and wavenumber of the yellow light.

Ans: $\nu = c/\lambda = 3X10^8/5800X10^{-10} = 5.17X\ 10^{14}$hertz

Wave number $(\bar{\nu}) = 1/\lambda = 1/5.8X10^{-7} = 1.72\ X\ 10^6$

6. Find energy of each of the photons which

(i) correspond to light of frequency 3×10^{15}Hz.

Ans: $E = h\nu = 6.6X1o^{-34}\ X3X10^{15} = 1.98X\ 10^{-18}$ Joules

(ii) have wavelength of 0.50 Å.

Ans: $E = hc/\lambda = 6.6X10^{-34}X3X10^8/0.5X10^{-10} = 19.8/5X\ 10^{-15}$Joules $=3.96\ X\ 10^{-15}$Joules

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

Ans: $1/T = 1/2X10^{-10} = \nu$;

$\nu = c/\lambda$;

$\lambda = c/\nu = 3X10^8/5X10^9 = 6X10^{-2}$m

Wave number $(\bar{\nu}) = 1/\lambda = 1/6\ X\ 10^{-2} = 167m^{-1}$

8. Electrons are emitted with zero velocity from a metal surface when it is exposed to radiation of wavelength 6800 Å. Calculate threshold frequency. ν_0) and work function (W_0) of the metal.

Ans: : work function = $hc/\lambda = 6.6X10^{-34}X3X10^8/6.8\ X\ 10^{-7} = 2.912\ X10^{-19}$Joules

Threshold frequency = $c/\lambda = 3X10^8/6.8\ X\ 10^{-7} = 4.4X10^{14}$

12. What are the frequency and wavelength of a photon emitted during a transition from n = 5 state to the n = 2 state in the hydrogen atom?

Ans: $E = -1312(1/5^2 - 1/2^2) = hc/\lambda$ in kJoules/mole= $1312\ X21X10^3/100X6.022X10^{23}$

$\lambda = hc/E = 6.6X10^{-34}X3X10^8X100\ X6.022X10^{23}/1312X21X10^3 = 19.8X6.6X10^{-1}/27552$

$=4.743\ X10^{-7}m = 4743X\ 10^{-10}$

Wavelength of the emitted photon is 4743A

$\nu = c/\lambda$

$\nu = 3X10^8/4.743X10^{-7} = 6.325X10^{14}$

13. Calculate the energy associated with the first orbit of He^+ .What is the radius of this orbit?

$E=-2.18 X10^{-18}Z^2/n^2J = -2.18X10^{-18}X2^2/1^2= -8.72X10^{-18}J.$

14. What is the wavelength of light emitted when the electron in a hydrogen atom undergoes transition from an energy level with n = 4 to an energy level with n =2?

$E=-2.18X10^{-18}(1/2^2-1/4^2)$

$=-2.18X10^{-18}X3/16= -0.40875X10^{-18}$

$E= -0.40875X10^{-18}$ → energy is released.

$E= hc/\lambda= 6.6X10^{-34}X3X10^8/ \lambda$

$\Lambda=6.6X10^{-34}X3X10^{8}/0.40875X10^{-18}= 48.44X10^{-8}m=4844A$

15 How much energy is required to ionize a H atom if the electron occupies n = 5 orbit? Compare your answer with the ionization enthalpy of H atom (energy required to remove the electron from n =1 orbit).

$E_5= -2.18X10^{-18}/5^2= -8.72X10^{-20}J$

$E_1=-2.18X10^{-18}$. Ionization energy of 1 H atom is 25 times the energy required to remove an electron from n=5.

16. What is the maximum number of emission lines when the excited electron of a H atom in n = 6 drops to the ground state?

Ans: 5+4+3+2+1=15 lines

17.(i) The energy associated with first orbit in the hydrogen atom is -2.18×10^{-18}J $atom^{-1}$. What is the energy associated with the fifth orbit?

(ii) Calculate the radius of Bohr's fifth orbit for hydrogen atom.

$R=52.9(n^2)/z= 52.9X25/1=1322.5pm$

18. Calculate the wave number for the longest wavelength transition in the Balmer series of atomic hydrogen.

$\bar{\nu} = 109,677 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) cm^{-1}$ longest wavelength is minimum energy ie3→4 transition

$=109677(1/3^2-1/2^2)$

$=109677X5/36= 15233cm^{-1}$

Λ= 6.565 X10^{-5}cm =6.565X10^{-7}m= 6565A

19. What is the energy in joules, required to shift the electron of the hydrogen atom from the first Bohr orbit to the fifth Bohr orbit and what is the wavelength of the light emitted when the electron returns to the ground state? The ground state electron energy is –2.18 X 10^{-11}ergs.

Ans: E=2.18X10^{-11}(1/5^2-1/1^2)

E=2.18X10^{-11}(24/25) = 2.09X10^{-11}ergs.=2.09X10^{-18}J

Λ=hc/E= 6.6X10^{-34}X3X10^8/2.09X10^{-18}=9.474X10^{-8}m =947.4A. This falls in the uv range

20. The electron energy in hydrogen atom is given by E_n = (–2.18 X 10^{-18})/n^2J. Calculate the energy required to remove an electron completely from the n = 2 orbit. What is the longest wavelength of light in cm that can be used to cause this transition?

Ans:E_2= 2.18X10^{-18}/4=5.36X10^{-19}J

Λ=hc/E= 6.6X10^{-34}X3X10^8/5.36X10^{-19}= 3.694X10^{-7}m= 3.694X10^{-5}cm

Ie light of minimum wavelength3.694X10^{-5}cm is required to remove the electron from n=2

21. What transition in the hydrogen spectrum would have the same wavelength as the Balmer transition n = 4 to n = 2 of He^+ spectrum?

E=-2.18X10^{-18}X2^2X(1/2^2-1/4^2)=2.18X10^{-18}X4X3/16= 1.635X10^{-18}J

Λ=hc/E= 6.6X10^{-34}X3X10^8/1.635X10^{-18}=12.11X10^{-8}m= 1211A. This belongs to the uv range and hence the llyman series.

E=2.18X10^{-18}X(1-1/n^2)=2.18X10^{-18}X12/16

1-1/n^2= 12/16=3/4

1/n^2=1/4

Hence n=2. The transition 2→1 in the Hydrogen spectrum corresponds to the given energy transition.

3..Questions based on Dual behaviour of matter (de Broglie's relation), Heisenberg uncertainty principle.

1. Explain Dual behaviour of matter.
2. State de Broglie's relation. Give its mathematical expression.
3. What will be the wavelength of a ball of mass 0.1 kg moving with a velocity of 10 m s^{-1}?
4. The mass of an electron is 9.1 X 10 kg. If its K.E. is 3.0 X10^{-25} J, calculate its wavelength.
5. Calculate the mass of a photon with wavelength 3.6 Å.

6. The velocity associated with a proton moving in a potential difference of 1000 V is $4.37 \times 10^5 ms^{-1}$. If the hockey ball of mass 0.1 kg is moving with this velocity,

Calculate the wavelength associated with this velocity.

7. If the velocity of the electron in Bohr's first orbit is $2.19 \times 10^6 ms^{-1}$, calculate the de Broglie wavelength associated with it.

8. Similar to electron diffraction, neutron diffraction microscope is also used for the determination of the structure of molecules. If the wavelength used here is 800 pm, calculate the characteristic velocity associated with the neutron.

9. Dual behaviour of matter proposed by de Broglie led to the discovery of electron microscope often used for the highly magnified images of biological molecules and other type of material. If the velocity of the electron in this microscope is $1.6 \times 10^6 ms^{-1}$, calculate de Broglie wavelength associated with this electron.

10. Calculate the wavelength of an electron moving with a velocity of $2.05 X 10^7 ms^{-1}$

11. The mass of an electron is $9.1 X 10^{-31}$kg. If its K.E. is $3.0 X 10^{-25}$ J, calculate its wavelength.

12. Why de Broglie's relation is not associated with ordinary objects.

13. State Heisenberg's Uncertainty Principle. Give its mathematical expression.

14. A microscope using suitable photons is employed to locate an electron in an

atom within a distance of 0.1 Å. What is the uncertainty involved in the measurement of its velocity?

15. If the position of the electron is measured within an accuracy of + 0.002 nm, calculate the uncertainty in the momentum of the electron. Suppose the momentum of the electron is h/4∏ × 0.05 nm, is there any problem in defining this value.

16. Show that the circumference of the Bohr orbit for the hydrogen atom is an integral multiple of the de Broglie wavelength associated with the electron revolving around the orbit.

ANSWER

3. What will be the wavelength of a ball of mass 0.1 kg moving with a velocity of $10\ m\ s^{-1}$?

   Ans: $\lambda = h/mv = 6.6X10^{-34}/0.1X10 = 6.6 X 10^{-34}m$

4. The mass of an electron is 9.1 X 10 kg. If its K.E. is $3.0 X10^{-25}$ J, calculate its wavelength.

   Ans: $KE = ½ mv^2 = 3X10^{-25}$; $v^2 = 6X10^{-25} = 60X10^{-26} \Rightarrow v = 7.746X 10^{-13}m/s$

   $\lambda = h/mv = 6.6X10^{-34}/9.1X10^{-31}X7.746X 10^{-13} = 9.363X10^{-12}m$

5. Calculate the mass of a photon with wavelength 3.6 Å.

   Ans: $v=3X10^8 m/s$

λ =h/mv= 6.6X10^{-34}/m X3X10^{8}

m=h/v λ = 6.6X10^{-34}/3X10^{8}X3.6X10^{-10}= 6.11X10^{-33}kg

6. The velocity associated with a proton moving in a potential difference of 1000 V is 4.37 × 10^{5}ms^{-1}. If the hockey ball of mass 0.1 kg is moving with this velocity,

   Calculate the wavelength associated with this velocity.

   λ =h/mv= 6.6X10^{-34}/0.1X4.37X10^{5}= 1.51X10^{-38}m

7. If the velocity of the electron in Bohr's first orbit is 2.19 × 10^{6}ms^{-1}, calculate the de Broglie wavelength associated with it.

   λ =h/mv= 6.6X10^{-34}/9.1X10^{-31}X2.19X10^{6}= 3.3X10^{-10}m

8. Similar to electron diffraction, neutron diffraction microscope is also used for the determination of the structure of molecules. If the wavelength used here is 800 pm, calculate the characteristic velocity associated with the neutron.

   λ =h/mv= 6.6X10^{-34}/6.67X10-27kgXv = 8X10^{-10}

   v=6.6X10^{-34}/6.67X10-27kgX8X10^{-10}=125m/s

9. Dual behavior of matter proposed by de Broglie led to the discovery of electron microscope often used for the highly magnified images of biological molecules and other type of material. If the velocity of the electron in this microscope is 1.6 × 10^{6}ms^{-1}, calculate de Broglie wavelength associated with this electron.

   λ =h/mv= 6.6X10^{-34}/9.1X10^{-31}X1.6X10^{6}= 4.5X10^{-10}m

10. Calculate the wavelength of an electron moving with a velocity of 2.05 X 10^{7}ms^{-1}

    λ =h/mv= 6.6X10^{-34}/9.1X10^{-31}X 2.05X10^{7}=3.538 X10^{-10}m

14. A microscope using suitable photons is employed to locate an electron in an atom within a distance of 0.1 Å. What is the uncertainty involved in the measurement of its velocity?

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

Δx.mΔv=h/4π

Δv=h/4πmΔx =6.6X10^{-34}/4X3.14X9.1X10^{-31}X10^{-11}=5.77X10^{-6}m/s

15. If the position of the electron is measured within an accuracy of + 0.002 nm, calculate the uncertainty in the momentum of the electron. Suppose the momentum of the electron is h/4∏ × 0.05 nm, is there any problem in defining this value.

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

$\Delta P = h/4\pi X \Delta x$

$\Delta P = 6.6X10^{-34}/4X3.14X0.002X10^{-9}$

$\Delta P = 2.69X\ 10^{-22} kgm/s$

4.QUESTIONS BASED ON QUANTUM NUMBERS, AUFBAU RULE, PAULI RULE, HUNDS RULE, ELECTRONIC CONFIGURATION OF ATOMS & IONS.

1. What information is provided by the four quantum numbers?

2. Using s, p, d, f notations, describe the orbital with the following quantum numbers

(a) n = 2, l = 1, (b) n = 4, l = 0, (c) n = 5, l = 3, (d) n = 3, l = 2 (e) n=1, l=0 (f) n = 3 l=1 (g) n = 4; l =2 (h) n= 4; l=3.

3. What is the total number of orbitals associated with the principal quantum number n = 3?

4. What is the lowest value of n that allows g orbitals to exist?

5. An electron is in one of the 3d orbitals. Give the possible values of n, l and m_l for this electron.

6. (i) An atomic orbital has n = 3. What are the possible values of l and m_l?

(ii) List the quantum numbers (m_l and l) of electrons for 3d orbital.

(iii) Which of the following orbitals are possible? 1p, 2s, 2p, 2d, 4f ,6d and 3f.

7. Explain, giving reasons, which of the following sets of quantum numbers are not possible.(a) n = 0, l = 0, m_l = 0, m_s = + ½ (b) n = 1, l = 0, m_l = 0, m s = – ½ (c) n = 1, l = 1, m_l = 0, m_s = + ½ (d) n = 2, l = 1, m_l = 0, m_s = – ½ (e) n = 3, l = 3, m_l = –3, m_s = + ½ (f) n = 3, l = 1, m_l = 0, m_s = + ½

8. (i)How many electrons in an atom may have the following quantum numbers?

(a) n = 4, m s = – ½ (b) n = 3, l = 0

(ii) How many sub-shells are associated with n = 4?

(iii) How many electrons will be present in the sub-shells having m s value of –1/2 for n = 4?

9. State (n+l) rule Aufbau rule & Pauli rule.

10. Give the electronic configuration of first 30 elements.

11. Explain the exceptional configuration of copper and chromium.

12. Give the electronic configurations of the following ions: Cu^{2+} Cr^{3+} Fe^{2+} S^{2-} Fe^{2+} O^{2-} Na^{+}

13. Explain Hunds rule of maximum multiplicity with an example.

14. Indicate the number of unpaired electrons in : (a) P, (b) Si, (c) Cr, (d) Fe and (e) Kr.

15. The quantum numbers of six electrons are given below. Arrange them in order ofincreasing energies. If any of these combination(s) has/have the same energy lists:

(1). $n = 4, l = 2, m_l = -2, m_s = -1/2$ (2). $n = 3, l = 2, m_l = 1, m_s = +1/2$

(3.) $n = 4, l = 1, m_l = 0, m_s = +1/2$ (4.) $n = 3, l = 2, m_l = -2, m_s = -1/2$

(5). $n = 3, l = 1, m_l = -1, m_s = +1/2$ (6). $n = 4, l = 1, m_l = 0, m_s = +1/2$

16. (i) Write the electronic configurations of the following ions: (a) H^-(b) Na^+(c) O^{2-}(d) F^-

(ii) What are the atomic numbers of elements whose outermost electrons are represented by (a) $3s^1$ (b) $2p^3$ and (c) $3p^5$?

(iii) Which atoms are indicated by the following configurations? (a) [He] $2s^1$(b) [Ne] $3s^2 3p^3$(c)[Ar] $4s^2 3d^1$

17. An atom of an element contains 29 electrons and 35 neutrons. Deduce (i) the number of protons and (ii) the electronic configuration of the element.

18. Among the following pairs of orbitals which orbital will experience the larger effective nuclear charge? (i) 2s and 3s, (ii) 4d and 4f, (iii) 3d and 3p.

19. The unpaired electrons in Al and Si are present in 3p orbital. Which electrons will experience more effective nuclear charge from the nucleus?

20 The bromine atom possesses 35 electrons. It contains 6 electrons in 2p orbital, 6 electrons in 3p orbital and 5 electrons in 4p orbital. Which of these electron experiences the lowest effective nuclear charge?

21. Draw the shapes of s,p,d & f orbitals.