XI STD-CHEMISTRY LESSON: ATOMIC STRUCTURE-I

1.Define Atom

All matter is composed of very small particles called atoms

2.Define Orbital

The nucleus is surrounded by electrons that move around the nucleus with very high speed in circular paths called **orbits**

3.Define Mass Number

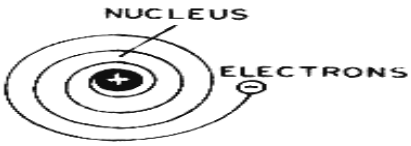
Protons and neutrons present in the nucleus are collectively also known as nucleons. The total number of nucleons is termed as **mass number(A)** of the atom.

4.Define Atomic Number

A number of electron or The number of protons in an atom is called its atomic number**(Z).**

5.State Defects of Rutherford's model

- According to Rutherford's model, an atom consists of a positive nucleus with the electrons moving around it in circular orbits.
- However it had been shown by J. C. Maxwell that whenever an electron is subjected to acceleration, it emits radiation and loses energy.
- As a result of this, its orbit should become smaller and smaller and finally it should drop into the nucleus by following a spiral path.
- This means that atom would collapse and thus Rutherford's model failed to explain stability of atoms.
- Another drawback of the Rutherford's model is that it says nothing about the electronic structure of the atoms i.e.,
- how the electrons are distributed around the nucleus and what are the energies of these electrons.
- Therefore, this model failed to explain the existence of certain definite lines in the hydrogen spectrum.

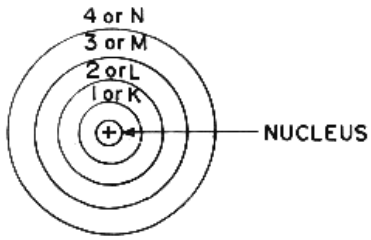


6.Explain Postulates of Bohr's model of an atom

- The electrons revolve round the nucleus only in certain selected circular paths called orbits. These orbits are associated with definite energies and are called energy shells or energy levels or quantum levels. These are numbered as 1, 2, 3, 4 ….. etc. (starting from the nucleus) are designated as K, L, M, N ….etc. (Fig. 3.2).
- As long as an electron remains in a particular orbit, it does not lose or gain energy. This means that energy of an electron in a particular path remains constant. Therefore, these orbits are also called stationary states.
- Only those orbits are permitted in which angular momentum of the electron isa whole number multiple of h/2π,Where 'h' is planck'sconstant. An electron moving in a circular orbit has an angular momentum equal to mvr where m is the mass of the electron and ν, theangular momentum,mvr is a Whole number multiplicity of h/2π

  $$mvr = nh/2\pi \text{ where } n=1,2,3,\ldots..$$

  In other words, angular velocity of electrons in an atom is quantised.
- If an electron jumps from one stationary state to another, it will absorb or emit radiation of a definite frequency giving a spectral line of that frequency which depends upon the initial and final levels. When an electron jumps back to the lower energy level, it radiates same amount of energy in the form of radiation.



7.ExplainLimitation of Bohr's Theory

- Bohr Theory had explained the existence of various lines in H- spectrum, but it predicted that only a series of lines exist. At that time this was exactly what had been observed. However, as better instruments and techniques were developed, it was realized that the spectral line that had been thought to be a single line was actually a collection of several lines very close together (known as fine spectrum). Thus for example, the single H∞-

spectral line of Balmer series consists of many lines very close to each other.

- Thus the appearance of the several lines implies that there are several energy levels, which are close together for each quantum number n. This would require the existence of new quantum numbers.
- Bohr's theory has successfully explained the observed spectra for hydrogen atom and hydrogen like ions (e.g. He^{+}, Li^{2+}, Be^{3+} etc.), it can not explain the spectral series for the atoms having a large number of electrons.
- There was no satisfactory justification for the assumption that the electron can rotate only in those orbits in which the angular momentum of the electron i.e. he could not give any explanation for using the principle of quantisation of angular momentum and it was introduced by him arbitrarily.
- Bohr assumes that an electron in an atom is located at a definite distance from the nucleus and is revolving round it with definite velocity, i.e. it is associated with a fixed value of momentum. This is against the Heisenberg's Uncertainty Principle according to which it is impossible to determine simultaneously with certainty the position and the momentum of a particle
- Does not explain Stark And Zeeman effect

8.State Zeeman Efeect:

If a substance which gives a line emission spectrum, is placed in a magnetic field, the lines of the spectrum get split up into a number of closely spaced lines. This phenomenon is known as Zeeman effect. Bohr's theory has no explanation for this effect.

9.State Stark effect:

If a substance which gives a line emission spectrum is placed in an external electric field, its lines get spilt into a number of closely spaced lines. This phenomenon is known as Stark effect. Bohr's theory is not able to explain this observation as well.

10.Define Quantum Numbers

The quantum numbers are nothing but the details that are required to locate an electron in an atom. In an atom a large number of electron orbitals are permissible. An orbital of smaller size means there is more chance of finding the electron near the nucleus. These orbitals are designated by a set of numbers known as quantum numbers.

11.Explain principal quantum number (n)

The electrons inside an atom are arranged in different energy levels called electron shells or orbits. Each shell is characterized by a quantum

number called principal quantum number. This is represented by the letter 'n' and 'n' can have values, 1,2,3,4 etc.

12.Explain subsidiary or azimuthal quantum number (*l*)

According to Sommerfield, the electron in any particular energy level could have circular path or a variety of elliptical paths about the nucleus resulting in slight differences in orbital shapes with slightly differing energies due to the differences in the attraction exerted by the nucleus on the electron. This concept gave rise to the idea of the existence of sub- energy levels in each of the principal energy levels of the atom. This is denoted by the letter '*l*' and have values from 0 to n-1.

if n=1, *l*=0 only one value (one level only) *s* level.

n=2, *l*=0 and 1 (2 values or 2 sub- levels) *s* and *p* level.

13.Magnetic quantum number (m)

This explains the appearance of additional lines in atomic spectra produced when atoms emit light in magnetic field. Each orbitals is designated by a magnetic quantum number m and its values depends on the value of '*l*' . The values are -'*l*' through zero to +'*l*' and thus there are (2*l*+1) values.

Thus when *l*=0, m= 0 (only one value or one orbital)

l=1, m= -1, 0, +1 (3 values or 3 orbitals)

14.Spin quantum number (*s*)

The electron in the atom rotates not only around the nucleus but also around its own axis and two opposite directions of rotation are possible (clock wise and anticlock wise). Therefore the spin quantum number can have only two values +1/2 or –1/2. For each values of m including zero, there will be two values for *s*.

15.State Pauli's exclusion principle

"it is impossible for any two electrons in a given atom to have all the four quantum numbers identical

16.State Hund's rule of maximum multiplicity

No pairing occurs until all orbitals of a given sub- level are half filled. This is known as Hund's rule of maximum multiplicity.

17.State Aufbau Principle

In the ground state of the atoms, the orbitals are filled in order of their increasing energies

(OR)

The lower the value of (n+1) for an orbital, the lower is its energy. If two orbitals have the same (n+1) value, the orbital with lower value of n has the lower energy.

18.Explain why the electronic configuration of Cr andCu are written as $3d^5$, $4s^1$ and $3d^{10}$ $4s^1$ instead of $3d^4$ $4s^2$ and $3d^9$ $4s^2$

Chromium

Expected configuration : $1s^2,2s^2,2p^6,3s^2,3p^6,3d^4,4s^2$ Actual configuration : $1s^2,2s^2,2p^6,3s^2,3p^6,3d^5,4s^1$

Electron exchange

Copper

Expected configuration : $1s^2,2s^2,2p^6,3s^2,3p^6,3d^9,4s^2$ Actual configuration : $1s^2,2s^2,2p^6,3s^2,3p^6,3d^{10},4s^1$

Electron exchange

19.What are the charge and mass of an electron?

Charge of electron is Negative(1.602×10^{-19} C) and mass is 9.11×10^{-31} kg.

20.What is the charge of an electron, proton and a neutron ?

- Electron - Negative
- Proton - Positive
- Neutron - Neutral

21.Explain Shape of Orbital

S-ORBITALS

- $l = 0$,m = 0. This means that the probability of finding the electron in s- orbital is the same in all directions at a particular distance. In other words s-orbitals are spherically symmetrical.
- The electron cloud picture of 1s-orbital is spherical.
- The *s*-orbitals of higher energy levels are also spherically symmetrical.
- They are more diffused and have spherical shells within them where probability of finding the electron is zero.
- These are called nodes. In 2*s*-orbital there is one spherical node.

***p*-orbital:**

- $l = 1$ m=+1, 0, -1. This means that there are three possible orientations of electron cloud in a *p*-sub-shell.
- The three orbitals of a *p*-sub-shell are designated as p_x, p_y and p_z respectively along x-axis, y-axis and z-axis respectively. Each *p*-orbital has two lobes, which are separated by a point of zero probability called node.
- *p*-orbital is thus dumb bell shaped.

***d*-orbitals:**

- $l = 2$, m = 0, ±1, ±2 indicating that d- orbitals have five orientations, i.e.,
- there are five *d*-orbitals which are

as d_{xy}, d_{yz}, d_{zx}, d_z and $d_{x^2-y^2}$. All these five orbitals, in the absence of magnetic field, are equivalent in energy and are, therefore, said to be five-fold degenerate

- The three orbitals namely d_{xy}, d_{yz} and d_{zx} have their lobes lying symmetrically between the coordinate axes indicated in the subscript to d, e.g. the lobes of d_{xy} orbital are lying between the x-and y-axes. This
- set of three orbitals is known as t_{2g} set. $d_{x^2-y^2}$ and d_z orbitals have their
- lobes along the axes (i.e. along the axial directions), e.g., the lobes of d orbital lie along the x and y-axes, while those of d_z

the z-axis. This set is known as e_g set.

22.What is stability of atom

The extraordinary stability of half-filled and completely filled electron configuration can be explained in terms of symmetry and exchange energy. The half-filled and completely filled electron configurations have symmetrical distribution of electrons and this symmetry leads to stability

23.Mention the Uses of Pauli's exclusion principle

- The greatest use of the principle is that it is helpful in determining the maximum number of electrons that a main energy level can have.
- Let us illustrate this point by considering K and L shells.

K-shell: For this shell n = 1. For n = 1, $l = 0$ and m = 0.

Hence *s* can have a value either +1/2 or –1/2. The different values

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(i) n = 1, l = 0, m = 0 s = +1/2 (1st electron)

(ii) n = 1, l = 0, m = 0, s = -1/2 (2nd electron)

- K shell there is only one sub- shell corresponding to l = 0 value (s-sub-shell) contains only two electrons with opposite spins.

- L-shell: For this shell n = 2. For n = 2 the different values of l, m and s give the following eight combinations of four quantum numbers.

  n= 2, l = 0, m = 0, s = +1/2

- Eight combinations given above show that L shell is divided into two sub-shells corresponding to l = 0 (s sub-shell) and l = 1 (p sub-shell) and this shell cannot contain more than 8 electrons, i.e., its maximum capacity for keeping the electrons is eight.

24.What are the information are given four quantum numbers

- n identifies the shell, determines the size of the orbital and also to a large extent the energy of the orbit.
- There are n subshells in the nth shell. l identifies the subshell and determines the shape of the orbital. There are ($2l$+1) orbitals of each type in a subshell i.e., one s orbital (l=0), three p orbitals (l=1), and five d orbitals (l=2) per subshell. To some extent l also determines the energy of the orbital in a multi-electron atom.
- m_l designates the orientation of the orbital. For a given value of l, m_l has ($2l$+1) values, the same as the number of orbitals per subshell. It means that the number of orbitals is equal to the number of ways in which they are oriented.

  m_s refers to orientation of the spin of the electron.

25.Explain Rutherford's nuclear model of atom.

- An atom consists of a tiny positively charged nucleus at its centre.
- The positive charge of the nucleus is due to protons. The mass of the nucleus, on the other hand, is due to protons and some neutral particles each having mass nearly equal to the mass of proton.
- This neutral particle, called neutron, was discovered later on by Chadwick in 1932. Protons and neutrons present in the nucleus are collectively also known as nucleons. The total number of nucleons is termed as **mass number(A)** of the atom.
- The nucleus is surrounded by electrons that move around the nucleus with very high speed in circular paths called **orbits**.
- Thus, Rutherford's model of atom resembles the solar system in which the sun plays the role of the nucleus and the planets that of revolving electrons.

- The number of electrons in an atom is equal to the number of protons in it. Thus, the total positive charge of the nucleus exactly balances the total negative charge in the atom making it electrically neutral. The number of protons in an atom is called its atomic number**(Z).**
- Electrons and the nucleus are held together by electrostatic forces of attraction.

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