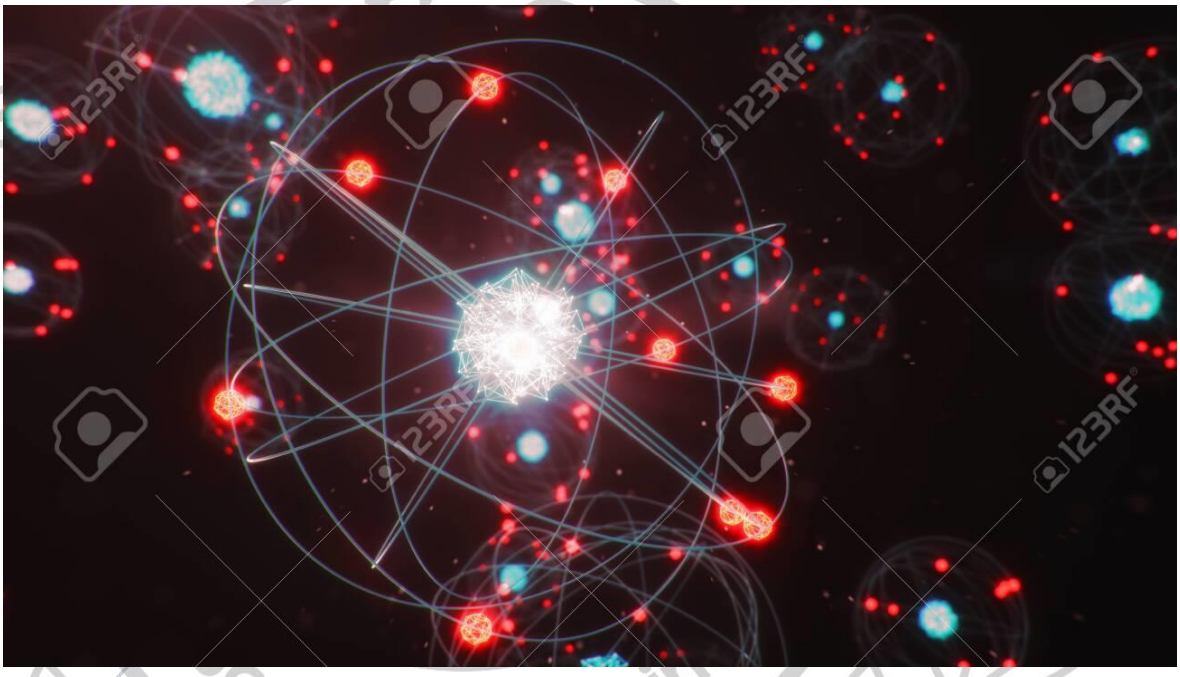
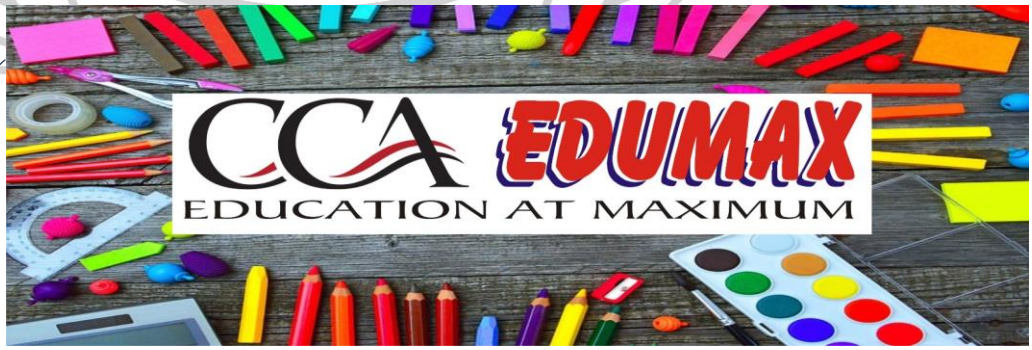
STRUCTURE OF ATOM





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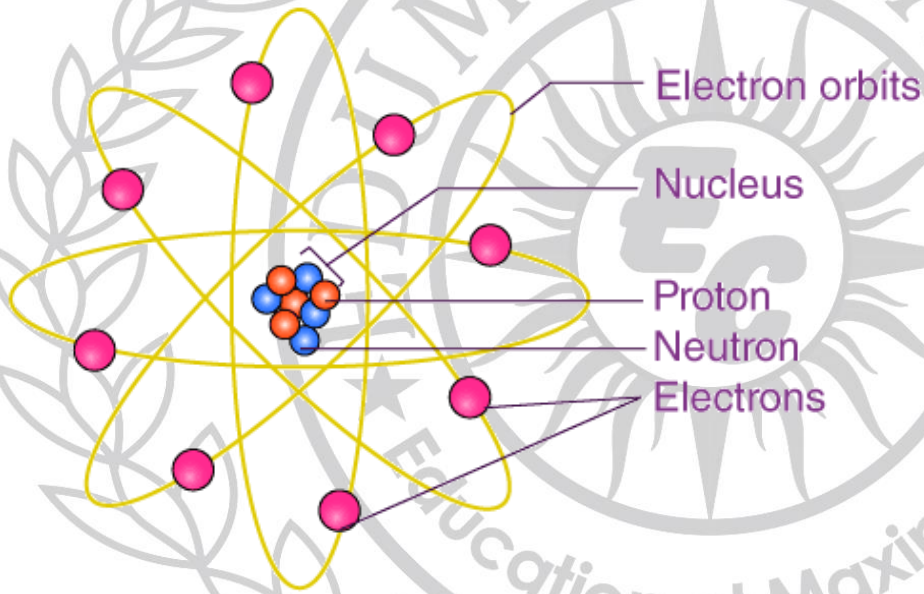
Atomic structure refers to the structure of an atom comprising a **nucleus** (centre) in which the **protons** (positively charged) and **neutrons** (neutral) are present. The negatively charged particles called **electrons** revolve around the **centre of the nucleus**.
The history of atomic structure and quantum mechanics dates back to the times of Democritus, the man who first proposed that matter is composed of atoms. The study about the structure of an atom gives a great insight into the entire class of chemical reactions, bonds and their physical properties. The first scientific theory of atomic structure was proposed by John Dalton in the 1800s.
The advances in atomic structure and quantum mechanics have led to the discovery of other fundamental particles. The **discovery of subatomic particles** has been the base for many other discoveries and inventions.

What is Atomic Structure?
The atomic structure of an element refers to the constitution of its nucleus and the arrangement of the electrons around it. Primarily, the atomic structure of matter is made up of protons, electrons and neutrons.
The **protons and neutrons** make up the nucleus of the atom, which is surrounded by the electrons belonging to the atom. The **atomic number** of an element describes the total number of protons in its nucleus.



The Structure of Atom

Neutral atoms have equal numbers of protons and electrons. However, atoms may gain or lose electrons in order to increase their stability and the resulting charged entity is called an ion.
Atoms of different elements have different atomic structures because they contain different numbers of protons and electrons. This is the reason for the unique characteristics of different elements.

Atomic Models
In the 18th and 19th centuries, many scientists attempted to explain the structure of the atom with the help of atomic models. Each of these models had their own merits and demerits and were pivotal to the development of the **modern atomic model**. The most notable contributions to the field were by the scientists John Dalton, J.J. Thomson, Ernest Rutherford and Niels Bohr. Their ideas on the structure of the atom are discussed in this subsection.

Dalton's Atomic Theory

The English chemist **John Dalton** suggested that all matter is made up of atoms, which were indivisible and indestructible. He also stated that all the atoms of an element were exactly the same, but the atoms of different elements differ in size and mass.

Chemical reactions, according to Dalton's atomic theory, involve a rearrangement of atoms to form products. According to the postulates proposed by Dalton, the atomic structure comprised atoms, the smallest particle responsible for the chemical reactions to occur.

The following are the postulates of his theory:

- Every matter is made up of atoms.
- Atoms are indivisible.
- Specific elements have only one type of atoms in them.
- Each atom has its own constant mass that varies from element to element.
- Atoms undergo rearrangement during a chemical reaction.
- Atoms can neither be created nor be destroyed but can be transformed from one form to another.

Dalton's atomic theory successfully explained the Laws of chemical reactions, namely, the Law of conservation of mass, Law of constant properties, Law of multiple proportions and Law of reciprocal proportions.

Demerits of Dalton's Atomic Theory

- The theory was unable to explain the existence of isotopes.
- Nothing about the structure of atom was appropriately explained.
- Later, the scientists discovered particles inside the atom that proved, the atoms are divisible.

The discovery of particles inside atoms led to a better understanding of chemical species, these particles inside the atoms are called subatomic particles. The discovery of various subatomic particles is as follows:

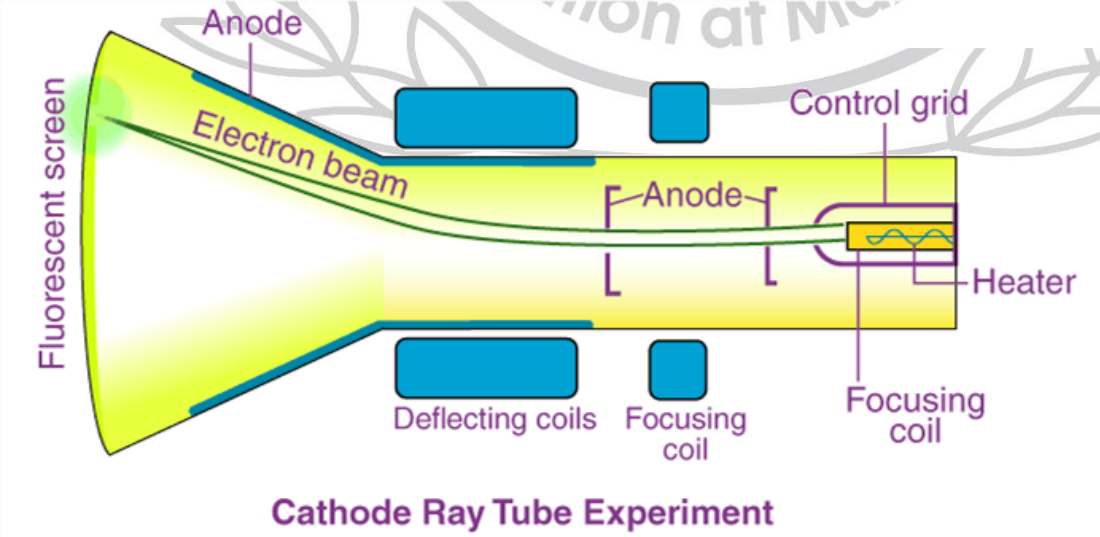
Thomson Atomic Model

The English chemist Sir Joseph John Thomson put forth his model describing the atomic structure in the early 1900s.

He was later awarded the Nobel prize for the **discovery of "electrons"**. His work is based on an experiment called cathode ray experiment. The construction of working of the experiment is as follows:

Cathode Ray Experiment

It has a tube made of glass which has two openings, one for the vacuum pump and the other for the inlet through which a gas is pumped in.

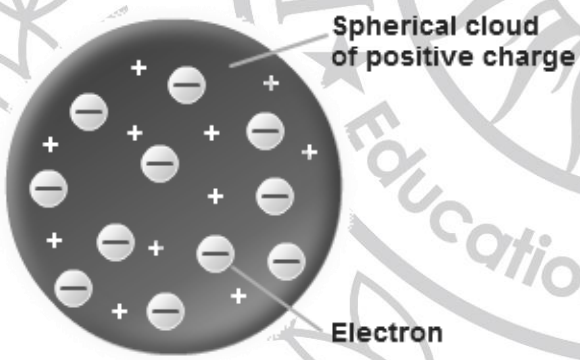


The role of the vacuum pump is to maintain "partial vacuum" inside the glass chamber. A high voltage power supply is connected using electrodes i.e. cathode and Anode is fitted inside the glass tube.
Observations:

- When a high voltage power supply is switched on, there were rays emerging from the cathode towards the anode. This was confirmed by the 'Fluorescent spots' on the ZnS screen used. These rays were called "Cathode Rays".
- When an external electric field is applied, the cathode rays get deflected towards the positive electrode, but in the absence of electric field, they travel in a straight line.
- When rotor Blades are placed in the path of the cathode rays, they seem to rotate. This proves that the cathode rays are made up of particles of a certain mass, so that they have some energy.
- With all this evidence, Thompson concluded that cathode rays are made of negatively charged particles called "electrons".
- On applying the electric and magnetic field upon the cathode rays (electrons), Thomson found the charge to mass ratio (e/m) of electrons. (e/m) for electron: 17588×10^{11} e/bg.

From this ratio, the charge of the electron was found by Mullikin through oil drop experiment. [Charge of e^- = 1.6×10^{-16} C and Mass of e^- = 9.1093×10^{-31} kg].

Conclusions:
Based on conclusions from his cathode ray experiment, Thomson described the atomic structure as a positively charged sphere into which negatively charged electrons were embedded.
It is commonly referred to as the **"plum pudding model"** because it can be visualized as a plum pudding dish where the pudding describes the positively charged atom and the plum pieces describe the electrons.
Thomson's atomic structure described atoms as electrically neutral, i.e. the positive and the negative charges were of equal magnitude.



Thomson's Model of an Atom

Limitations of Thomson's Atomic Structure: Thomson's atomic model does not clearly explain the stability of an atom. Also, further discoveries of other subatomic particles, couldn't be placed inside his atomic model.

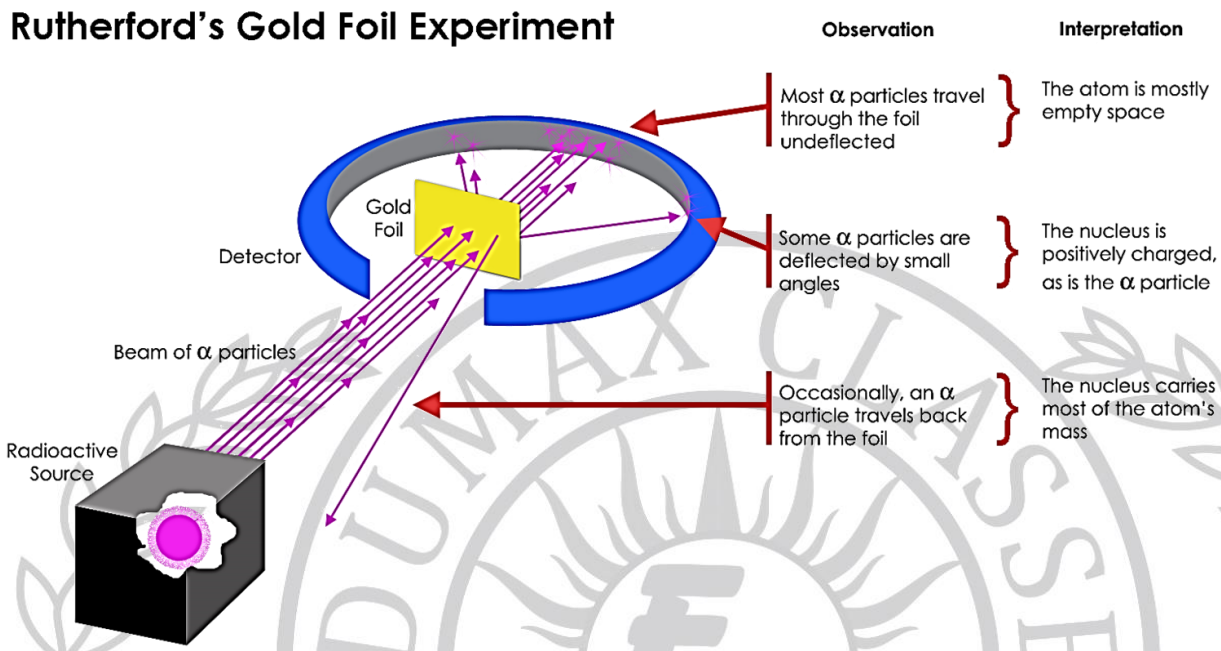
Discovery of Proton:
By Ernest Goldstein in 1886.
He observed in the same gas discharge tube, with different situations that the anode emitted positive particles which he named as Canal Rays. His experiment led to the discovery of proton.

Discovery of Neutron:
By J. Chadwick in1932.
Neutron is present in the nucleus of all atoms.

Rutherford Atomic Theory

Rutherford, a student of J. J. Thomson modified the atomic structure with the discovery of another **subatomic particle called "Nucleus"**. His atomic model is based on the Alpha ray scattering experiment.

Alpha Ray Scattering Experiment



Construction:

- A very thin gold foil of 1000 atoms thick is taken.
- Alpha rays (doubly charged Helium He^{2+}) were made to bombard the gold foil.
- Zn S screen is placed behind the gold foil.

Observations:

- Most of the rays just went through the gold foil making scintillations (bright spots) in the ZnS screen.
- A few rays got reflected after hitting the gold foil.
- One in 1000 rays got reflected by an angle of 180° (retraced path) after hitting the gold foil.

Conclusions:

- Since most rays passed through, Rutherford concluded that most of the space inside the atom is empty.
- Few rays got reflected because of the repulsion of its positive with some other positive charge inside the atom.
- 1/1000th of rays got strongly deflected because of a very strong positive charge in the center of the atom. He called this strong positive charge as "nucleus".
- He said most of the charge and mass of the atom resides in the Nucleus

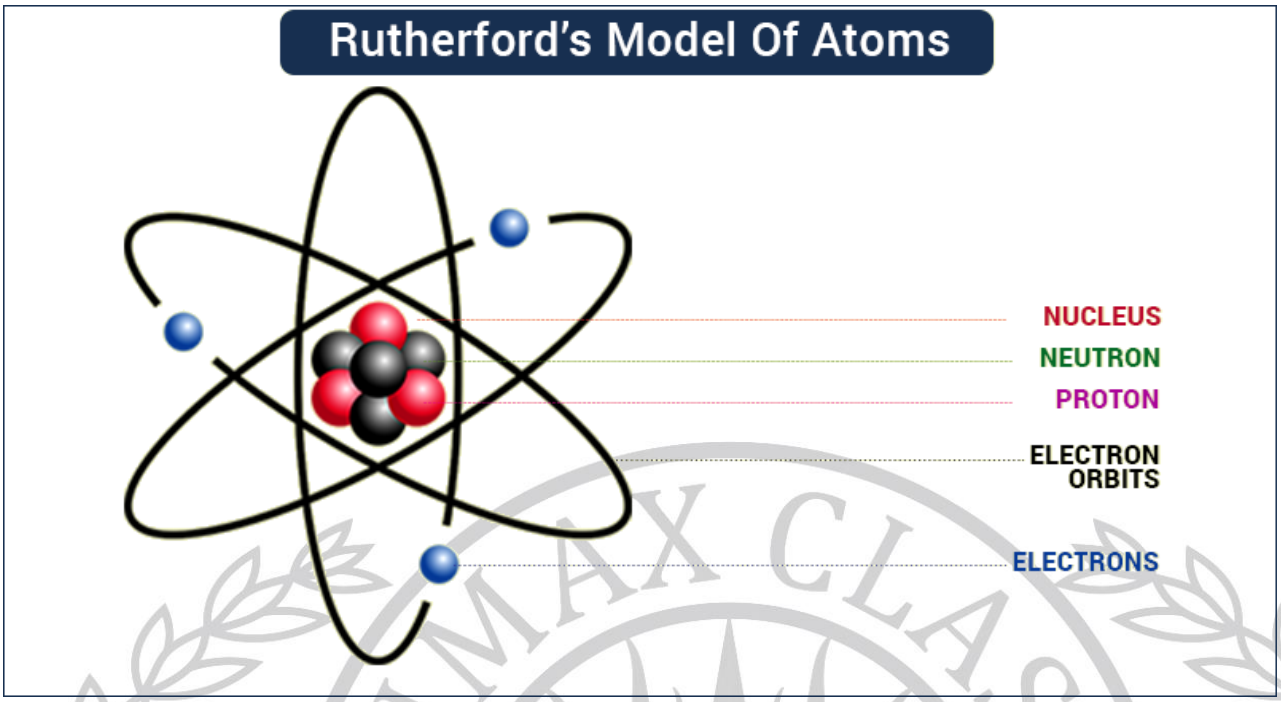
Rutherford's model of an atom

Rutherford concluded the model of the atom from the α-particle scattering experiment as:

(i) There is a positively charged centre in an atom called the nucleus. Nearly all the mass of an atom resides in the nucleus.

(ii) The electrons revolve around the nucleus in well-defined orbits.

(iii) The size of the nucleus is very small as compared to the size of the atom.



Drawbacks of Rutherford's model

- He explained that the electrons in an atom revolve around the nucleus in well-defined orbits. Particles in a circular orbit would experience acceleration.
- Thus, the revolving electron would lose energy and finally fall into the nucleus.
- But this cannot take place as the atom would be unstable and matter would not exist in the form we know.

Be More Curious!!!

- The Millikan's Oil Drop Experiment was an experiment performed by Robert A. Millikan and Harvey Fletcher in 1909 to measure the charge of an electron.
- In the experiment, Millikan allowed charged tiny oil droplets to pass through a hole into an electric field.
- By varying the strength of electric field, the charge over an oil droplet was calculated, which always came as an integral value of **'e.'**
- The conclusion of this is that the charge is said to be quantized, i.e. the charge on any particle will always be an integral multiple of e which is $1.6*10^{-19}$

Neil Bohr Model

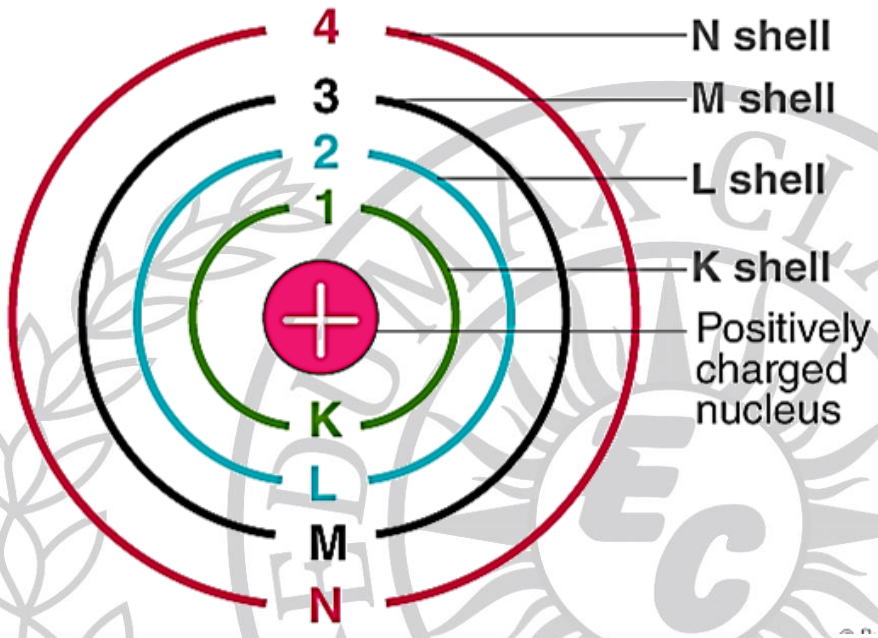
Properties of electrons, protons and neutrons

Particle	*Charge on the particle*	*Mass of the particle*	*Symbol*	*Location in the atom*
1. *Electron*	-1 *unit* $(-1.602 \times 10^{-19}$ *coulomb*$)$	9.11×10^{-31} *kg* $\left(\frac{1}{1840}u\right)$	${}^{0}_{-1}e$	*Outside the nucleus (Extranuclear part)*
2. *Proton*	$+1$ *unit* $(+1.602 \times 10^{-19}$ *coulomb*$)$	1.673×10^{-27} *kg* $(1\ u)$	${}_{+1}p^{1}$	*In the nucleus*
3. *Neutron*	*No charge*	1.675×10^{-27} *kg* $(1\ u)$	${}^{1}_{0}n$	*In the nucleus*

Bohr's Model of an atom

Bohr came up with these postulates to overcome the objections raised against Rutherford's model:

- Electrons revolve around the nucleus in stable orbits without emission of radiant energy. Each orbit has a definite energy and is called an energy shell or energy level.
- An orbit or energy level is designated as K, L, M, N shells. When the electron is in the lowest energy level, it is said to be in the ground state.
- An electron emits or absorbs energy when it jumps from one orbit or energy level to another.
- When it jumps from a higher energy level to lower energy level, it emits energy while it absorbs energy when it jumps from lower energy level to higher energy level.



Orbits

Orbits are energy shells surrounding the nucleus in which electrons revolve.

Electron distribution in different orbits

The distribution was suggested by Bohr and Bury;

- The maximum number of electrons present in a shell is given by the formula $2n^2$, where 'n' is the orbit number or energy level index, 1,2,3,….
- The maximum number of electrons in different shells are as follows: the first orbit will have $2*1^2=2$, the second orbit will have 2*2Msup>2=8, the third orbit will have $2*3^2=18$, fourth orbit $2*4^2=32$ and so on.
- The shells are always filled in a step-wise manner from the lower to higher energy levels. Electrons are not filled in the next shell unless previous shells are filled.

Valency

- The electrons present in the outermost shell of an atom are known as the valence electrons.
- The combining capacity of the atoms or their tendency to react and form molecules with atoms of the same or different elements is known as valency of the atom.
- Atoms of elements, having a completely filled outermost shell show little chemical activity.
- Their combining capacity or valency is zero.
- For example, we know that the number of electrons in the outermost shell of hydrogen is 1, and in magnesium, it is 2.
- Therefore, the valency of hydrogen is 1 as it can easily lose 1 electron and become stable.
- On the other hand, that of magnesium is 2 as it can lose 2 electrons easily and also attain stability.

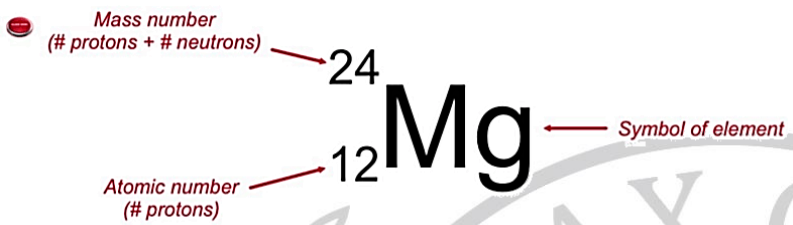
Atomic Number

The number of protons found in the nucleus of an atom is termed as the atomic number. It is denoted by the letter 'Z'.

Mass number and representation of an atom

Protons and neutrons are present in the nucleus, so the mass number is the total of these protons and neutrons.

A simple way to give us all the information we need about the subatomic make-up of an atom is via chemical or isotopic notation:

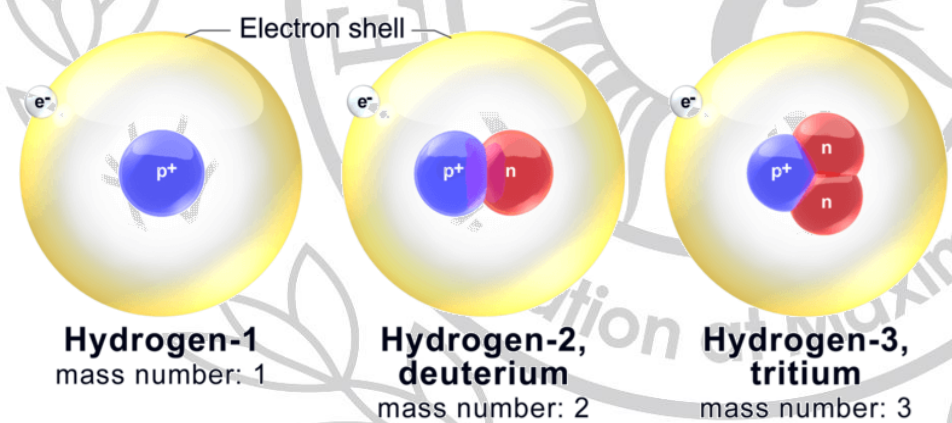


protons (p^+) = 12
electrons (e^-) = 12
neutrons (n^0) = mass # − p^+
= 12

Isotopes and Isobars

Isotopes are defined as the atoms of the same element, having the same atomic number (number of protons) but different mass numbers (number of protons+neutrons).

For example: In the case of Hydrogen we have:



Atoms of different elements with different atomic numbers, which have the same mass number, are known as isobars.

For example, Calcium and Argon: both have the same mass number – 40

$_{20}Ca^{40}$ and $_{18}Ar^{40}$

Calculation of mass number for isotopic elements

When an element has an isotope, the mass number can be calculated by the different proportions it exists in.

For example take 98% Carbon-12u and 2% Carbon-13u

$$(12 \times \frac{98}{100}) + (13 \times \frac{2}{100}) = 12.02u$$

This does not mean that any Carbon atoms exists with the mass number of 12.02u. If you take a certain amount of Carbon, it will contain both isotopes of Carbon, and the average mass is 12.02 u.