# Atomic and molecular structure

## Introduction to Structure of Atom

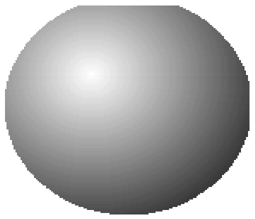
An atom (meaning “indivisible ” in Greek ) is present at the most basic level in everything we see around us. Living organism and every non-living thing is composed of atoms. around you such as tables, chairs, water, etc is made up of matter. But the building blocks of matter can be broken down into very small invisible particles called atoms. Therefore, living or non-living, everything is composed of atoms.

Greek philosophers such as:

- **Democritus** - explained the nature of matter. He also proposed that all substances are made up of matter. He stated that atoms are constantly moving, invisible, minuscule particles that are different in shape, size, temperature and cannot be destroyed.

- **John Dalton** - The first scientific theory of atomic structure was proposed by John Dalton in the 1800s.

1. Atom was indivisible.
2. All elements are composed of atoms.
3. The same atoms for one element are exactly alike.
4. Atoms are neither created nor destroyed in a chemical reaction.
5. In a chemical reaction, atoms are separated, combined, or rearranged



Democritus & Dalton

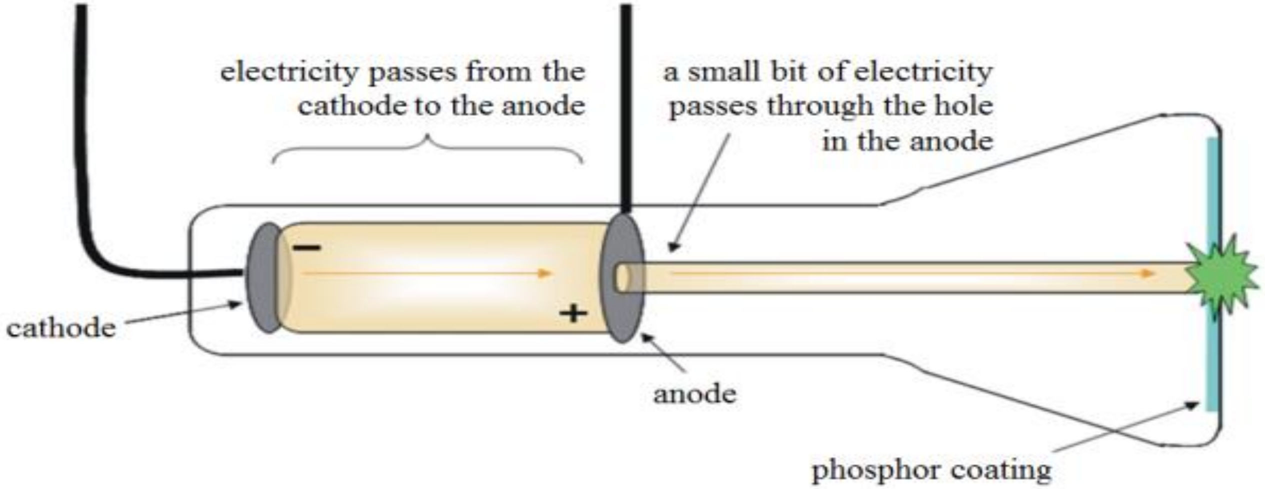
These two are the first set of people to put forward the concept of the atom.

By the end of 18th and the early 20th centuries, many scientists such as J.J Thompson, Gold stein, Rutherford, Bohr among others developed and proposed several concepts on “atom.

## The electrons

The first indication that Atoms had internal structure was the discovery of the Electron by the cathode ray experiment.

- **J.J. Thomson in 1897 used** a device called a cathode ray tube to conduct his experiments. The particles he detected were attracted to the positive end of the circuit, so they had to be negatively charged which are known as ELECTRONS




Cathode Ray Tube Diagram

- He also applied electric and magnetic field to a beam of electrons and used the deviation from a straight line to calculated the charge to mass ratio for an individual electron. The number he came up with was $-1.76 \times 10^8$ C/g. Where C stands for coulomb

- **Thirteen years later (1910),** The American scientist , **Robert A. Millikan** determine the charge of the electron with great precision by observing the rate at of fall of charged oil droplets. Using his knowledge of electrostatics , he found the charge of an electron to be;

    - -1.602 x 10-19 C

From this data, he calculated the mass of an electron (e -)

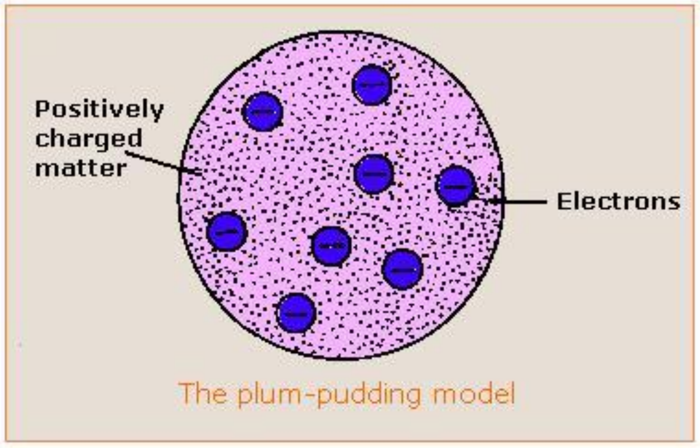
Mass of an electron = charge/ charge to mass

$= - 1.6 \times 10^{-19}$ C/ $-1.76 \times 10^{8}$ c/g

$= 9.10 \times 10^{-28}$ g

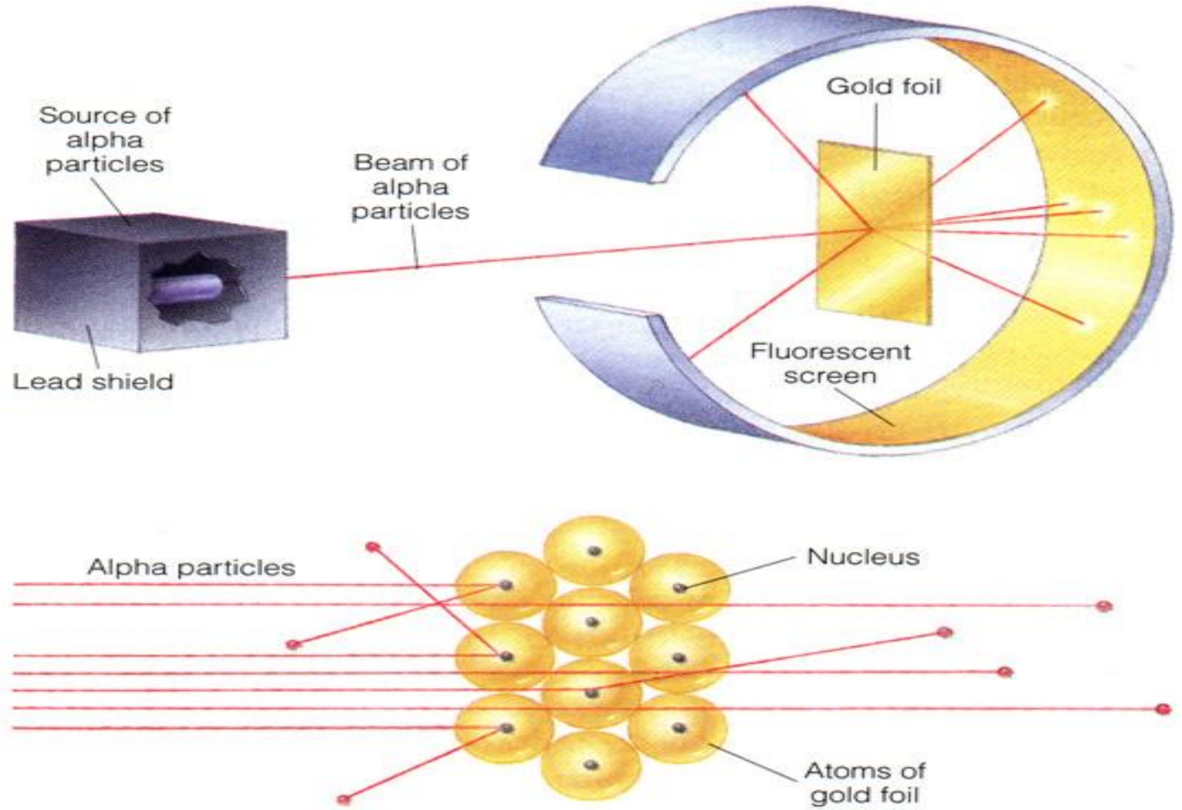
This is an exceedingly small mass.

Thompson Knowing that the atom was neutral and the electron was negative, so there must be positive material with a lot more mass. He then proposed that an atom could be thought of as a uniform, positive sphere of matter in which electrons are embedded and scattered like a raisin in a cake. This is referred to as "The plum-puddng model"

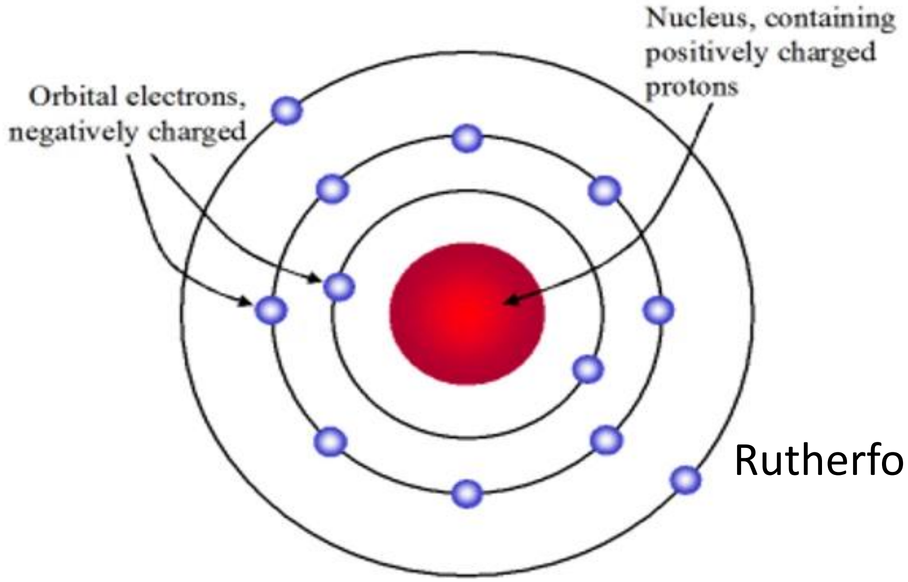


**Proton and the Nucleus**

**In 1909, Ernest Rutherford** discovered the Nucleus of an atom from the observation of his alpha ($\alpha$) scattering experiment using a gold foil. He concluded that the atom's positive charge particles are all concentrated in the nucleus, which is a dense central core within the atom.



He called the positive charged particles **PROTONS.**




Rutherford atomic model

In a separate experiment, it was found that each proton carries the same quantity of charge as an electron and has a mass of -1.6 72262 x $10^{-24}$ g . About 1840 times the mass of the oppositely charged electron.

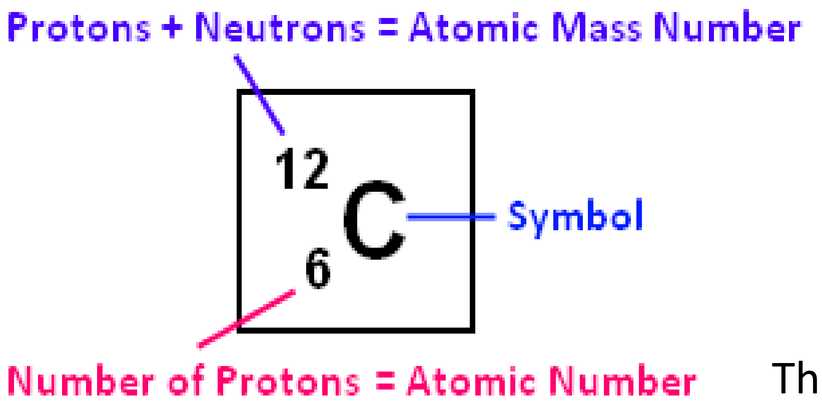
Rutherford's model of atomic structure left one major problems unsolved. It was known that hydrogen (H), the simplest atom contains only one proton and that of helium (He) atom contains two protons. Therefore the ratio of the mass of Helium to that of hydrogen atom should be 2:1. In reality however, the ratio is 4:1. Rutherford and others postulated that there must be another type of sub atomic particle in the atomic nucleus. The proof was provided by another English physics**,** **James Chadwick in 1932 .**

- He showed that the nucleus also had **neutrons** and
- The neutron was basically equal in mass to the proton but had **no electrical** charge.

## Comparison of subatomic Particles

| Particles | Charge | Actual Mass (g) | Relative Mass (amu) | Location |
|---|---|---|---|---|
| Proton | + | $1.67 \times 10^{-24}$ | 1 | Inside nucleus |
| Neutron | 0 | $1.67 \times 10^{-24}$ | 1 | Inside nucleus |
| Electron | - | $9.11 \times 10^{-28}$ | 0 | Outside nucleus |

## Subatomic particles in an atom




The electrons are equal to the number of protons

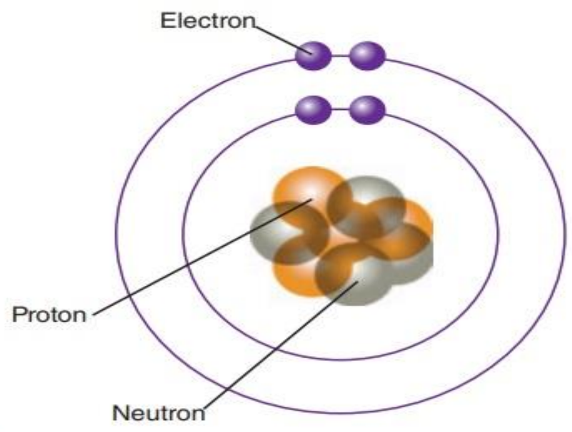
# Structure of Atom

The study about the structure of an atom gives a great insight into the entire class of chemical reactions, bonds and their physical properties.

Atom is the smallest unit of matter that is composed of a *positively charged centre termed as ''nucleus''* and the central nucleus is *surrounded by negatively charged electrons*. Even though an atom is the smallest unit of matter but it retains all the chemical properties of an element .

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.




Structure of an Atom

## Importance of Electrons

1. Stabilizing of the Nucleic charge -

Protons in the nucleus are positively charged, whereas electrons are negatively charged (neutron has no charge but only account for the mass. The charges balanced out, making the atom electrically neutral when there is equal amount of the protons and electron. This concept outline the picture or bonding and electrical interaction between atoms

2. Absorption and release of energy – An atom changes from a ground state to an excited state by taking on energy from its surroundings in a process called absorption. The **electron** absorbs the energy and jumps to a higher energy level. In the reverse process, emission, the electron returns to the ground state by releasing the extra energy it absorbed.

3. Bonding – (Chemical bonding)

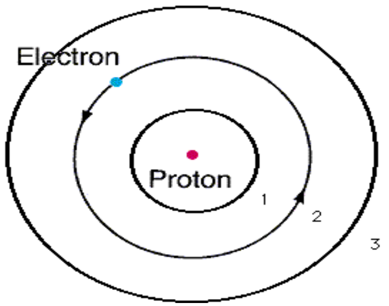
Bonds are formed by electrical tendencies of electrons between atoms. E.g. include Covalent bonding, ionic bonding, metallic bonding

4. The number of electrons in an atom of an element controls the chemical properties of the element empirical observation that led to the grouping together of elements with similar chemical characteristics and to the construction of the periodic table

**Niels Bohr's model**

In 1913, Neil Bohr studied the hydrogen atom and the light it produces when excited (heated, electrified). He revised the atomic theory to include the following points:

- That electrons moved around the nucleus in definite orbits or energy levels.
- Each orbit can hold a certain maximum number of electrons: 2 in the first orbit, 8 in the second, and 8 in the third.
- Electrons can jump from orbit to orbit. They release energy as light when they jump from higher to lower orbits
- Each electron in an orbit has a specific amount of energy. The farther the electron is from the nucleus, the higher its energy.



Inaccuracies in Bohr’s model/Limitations

- It cannot predict the relative intensities of spectral lines.
- It does not explain the Zeeman Effect, when the spectral line is split into several components in the presence of a magnetic field.
- The Bohr Model does not account for the fact that **accelerating electrons do not emit electromagnetic radiation**
- It only works for hydrogen atom (though can be adapted to other one electron ions). If there are 2 or more electrons, the mathematical formula does not match real data.
- It is fundamentally incorrect in that electrons *do not* move in fixed orbits!

Bohr’s model was a very important advance in the development of Atomic theory and still in use apart from some inaccuracy. Because;

1. Its simple approach to atomic theory
2. it was the first model to postulate the quantization of electron orbits in atoms

The idea of discreet energies for electromagnetic radiation i.e. the energy is quantized was developed by Max Plank in the late 19th century. i.e. A photon's energy is equal to its frequency multiplied by the Planck constant. Due to mass–energy equivalence, the Planck constant also relates mass to frequency.

The relationship between energy E and frequency $\upsilon$ is given as ;

$$E = h\,\upsilon$$

The proportionality constant, h is called the plank's constant ($h = 6.626 \times 10^{-34}$ Js or $6.626 \times 10^{-34}\ m^2\ kg / s$ )

## Wave/Particle Duality

Plank's model assumed that light was an electromagnetic wave. In 1905, Albert Einstein proposed that electromagnetic could exhibit particle - like behavior. i.e. light has a dual nature. If energy can behave like a particle and a wave, then matter can also exhibit this duality.

In the 1924, Louis Victor de Broglie showed that all moving particles can be described as wave-like. We call this general phenomenon *wave/particle duality*.

The De Broglie relationship shows that a particle (this include electrons) wit a momentum mv has an associated wave of wavelength ($\lambda$)

$$\lambda = h/mv$$

Where h = plank's constant , m = mass and v = velocity of the particle

## Heisenberg Uncertainty Principle

The **Heisenberg Uncertainty Principle** states that it is impossible to determine simultaneously both the position and the velocity of a particle. The detection of an electron, for example, would be made by way of its interaction with photons of light. Since photons and electrons have nearly the same energy, any attempt to locate an electron with a photon will knock the electron off course, resulting in uncertainty about where the electron is located. We do not have to worry about the uncertainty principle with large everyday objects because of their mass. If you are looking for something with a flashlight, the photons coming from the flashlight are not going to cause the thing you are looking for to move. This is not the case with atomic-sized particles, leading scientists to a new understanding about how to envision the location of the electrons within atoms. we can only talk about the ***PROBABILITY*** of finding an electron in a certain time and place.

Together, the quantization of energy, wave / particle duality and the Uncertainty Principle all led to *Wave Mechanics*