## Section 2: ATOMIC STRUCTURE of MATTER

In this section, we present a basic description of the atomic structure of matter together with definitions of terms used in the literature.

### Background:

Matter can be broken down chemically into elements – H, C Fe etc. These consist of atoms, which for many years were regarded as the smallest, indestructible units of matter. However, in the decade around turn of the 19th century, key discoveries were made which completely changed this view.

From studies of electrical discharges in gases, JJ Thompson in 1897 demonstrated the existence of the **electron** as a fundamental unit of negative electricity having a very small mass compared with that of an atom. Evidently, the atom could be split into a positively-charged part [a nucleus] and negatively-charged electrons.

In 1896, Henri Becquerel in France first encountered radioactivity in the form of unknown radiations emanating from uranium-bearing rocks. Later studies showed there are 3 distinct types of radiation: alpha, beta and gamma rays – terms that have been retained to this day. Later work, notably by Ernest Rutherford in England, showed that alpha particles are positively charged helium nuclei, beta particles are electrons and gamma rays were identified as energetic electromagnetic radiation [photons].

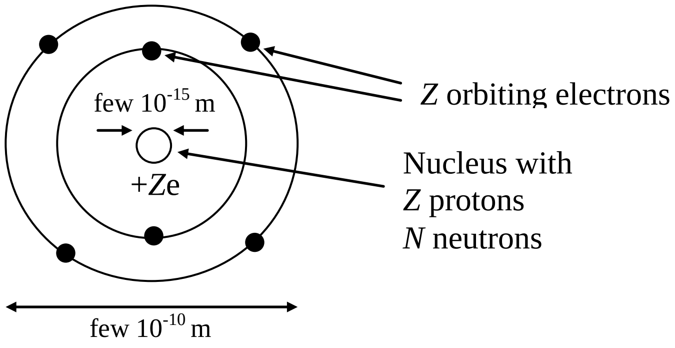
Rutherford and his co-workers, first in Canada (1898-1907) and later in Manchester England (1907-1919) made detailed studies of the way beams of alpha particles pass through matter. It was noted that occasionally, an alpha particle was scattered backwards by even very thin foils of gold. Rutherford realised that this could only be explained if the alpha particle had encountered a tiny, massive charged entity less than 1/1000th of the atom in size. He proposed the **planetary model** of the atom as consisting of a small heavy positively-charged nucleus surrounded by orbiting electrons which occupy the vast bulk of the atom's volume. The simplest [lightest] atom, hydrogen, was pictured as an electron orbiting a **proton** held together by the electrical force acting between them. Other atoms were thought to be combinations of these two fundamental particles. However, further studies and developments in quantum mechanics exposed several fatal shortcomings to this **proton-electron model** of the nucleus.

A breakthrough came some time later, in 1930, when a previously unseen, highly penetrating, uncharged radiation was shown to result from the bombardment of beryllium with alpha particles. Experiments were then done to observe the effect of bombarding different materials with this new radiation. James Chadwick interpreted the results of these experiments and in 1932 correctly deduced the existence of the **neutron**, a new uncharged particle with a mass approximately equal to that of the proton.

The neutron was the critical missing ingredient to understanding atomic structure and lead to the modern picture of the nucleus as consisting of **nucleons** [neutrons and protons] held together by a short-range (nuclear) force whose strength is independent of the type of nucleon. Electrons are not influenced by the nuclear force and form a cloud of negative charge surrounding the nucleus.

**Atomic structure:**

All atoms are composed of a central **nucleus** surrounded by a number of orbiting **electrons** - like planets orbiting the sun



A problem is that orbiting electrons would be accelerating and should radiate energy causing them to spiral into the nucleus. We will see why they don't later.

**Atomic constituents:**

An atomic nucleus is the small heavy central part of an atom consisting of $A$ nucleons: $Z$ protons and $N$ neutrons.

$A$ is referred to as the mass number and $Z$ the atomic number.

Nuclear size is measured in fermis (also called femtometres), where 1 fm = $10^{-15}$ m.

The basic properties of atomic constituents are as follows:

| | charge | mass (u) | spin ($h/2\pi$)* | magnetic moment (J $T^{-1}$)** |
|---|---|---|---|---|
| proton | $e$ | 1.007276 | ½ | $1.411 \times 10^{-26}$ |
| neutron | $e$ | 1.008665 | ½ | $-9.66 \times 10^{-27}$ |
| electron | $e$ | 0.000549 | ½ | $9.28 \times 10^{-24}$ |

* $h = 6.626 \times 10^{-34}$ J s, is Planck's constant [see Section 3]

** The unit T is the tesla - the SI unit of magnetic field.

**Charge:** Protons have a positive charge ($+e = 1.602 \times 10^{-19}$ C (coulomb) equal and opposite to that of the electron. Neutrons are uncharged. Thus a neutral atom ($A$, $Z$) contains $Z$ electrons and can be written symbolically as ${}^{A}_{Z}\mathrm{X}_{N}$.

**Mass:** Nuclear and atomic masses are expressed in atomic mass units (u) based on the definition that the mass of a neutral atom of $^{12}$C is exactly 12.000 u. ($1\ \mathrm{u} = 1.6605 \times 10^{-27}$ kg).

**Spin:** Each of the constituents has a spin ½ in units of $h/2\pi$ and is an example of the class of particles of half-integer spin known as *fermions.* Fermions obey the *Exclusion Principle,* first enunciated by Wolfgang Pauli in 1925, which determines the way electrons can occupy atomic energy states. The same rule applies to nucleons in nuclei as is discussed in a later section.

**Magnetic moment:** Associated with the spin is a magnetic dipole moment. Compared with the magnetic moment of an electron, nuclear moments are very small. However, they play an important role in the theory of nuclear structure. It may be surprising that the uncharged neutron has a magnetic moment. This reflects the fact that it has an underlying quark sub-structure, consisting of charged components.

An important application of nuclear moments, based on their behaviour in electromagnetic fields, is the technique of magnetic resonance imaging [MRI] or nuclear magnetic resonance.

For light nuclei, $N \approx Z$. However, for heavier nuclei, there are significantly ($\approx$50%) more neutrons than protons.

- mass(proton) ≈ mass(neutron) ≈ 2000 × mass(electron).

  Therefore, almost **all** the mass of an atom (~99.98%) is contained within the nucleus.

An atom in its normal state is electrically **neutral** ⇒ number of orbiting electrons = number of protons in the nucleus.

An atom can be **ionized** by

- **losing** one or more electrons and forming a **positive** ion (e.g. $\mathrm{Fe}^{+}$, $\mathrm{Cu}^{+}$)
- **gaining** additional electrons and forming a **negative** ion (e.g. $\mathrm{O}^{-}$, $\mathrm{O}^{--}$).

## Elements, isotopes, isotones and isobars

The name of an **element** uniquely identifies the atomic number $Z$.

| Element | Atomic number *Z* |
|---|---|
| Hydrogen | 1 |
| Helium | 2 |
| Lithium | 3 |
| Beryllium | 4 |
| Boron | 5 |
| Carbon | 6 |
| … | … |
| Uranium | 92 |
| Neptunium | 93 |
| Plutonium | 94 |

**Isotopes** of an element: Atoms whose nuclei have the same *Z* but different *N*.
The electron structure and, therefore, the chemical properties of isotopes are similar.
e.g. Hydrogen ${}^{1}_{1}H_{0}$, ${}^{2}_{1}H_{1}$ and ${}^{3}_{1}H_{2}$, nuclei are the proton p, the deuteron d and the triton t.
Carbon has three naturally occurring isotopes: ${}^{12}_{6}C_{6}$, ${}^{13}_{6}C_{7}$ and ${}^{14}_{6}C_{8}$.

**Isotones:** nuclides with the same *N* and different *Z*.

**Isobars:** nuclides with the same mass number *A*.

It is usual to omit the *N* and *Z* subscripts from the symbolic representation and include only the mass number as superscript:

$$\boxed{\text{e.g. } {}^{12}C}$$

## Forces acting in atoms and nuclei

- The electromagnetic force acts as an attractive force between electrons and the nucleus and as a repulsive force between protons inside the nucleus.
- Within the nucleus, a strong, short-range force holds nucleons together in close proximity to each other.
- There is also a very much weaker nuclear force, which gives rise to radioactivity (see Section 7).
- The gravitational force acts between objects with mass, but is only perceived when masses are large. It is totally negligible in its effect on the structure of atoms and nuclei.