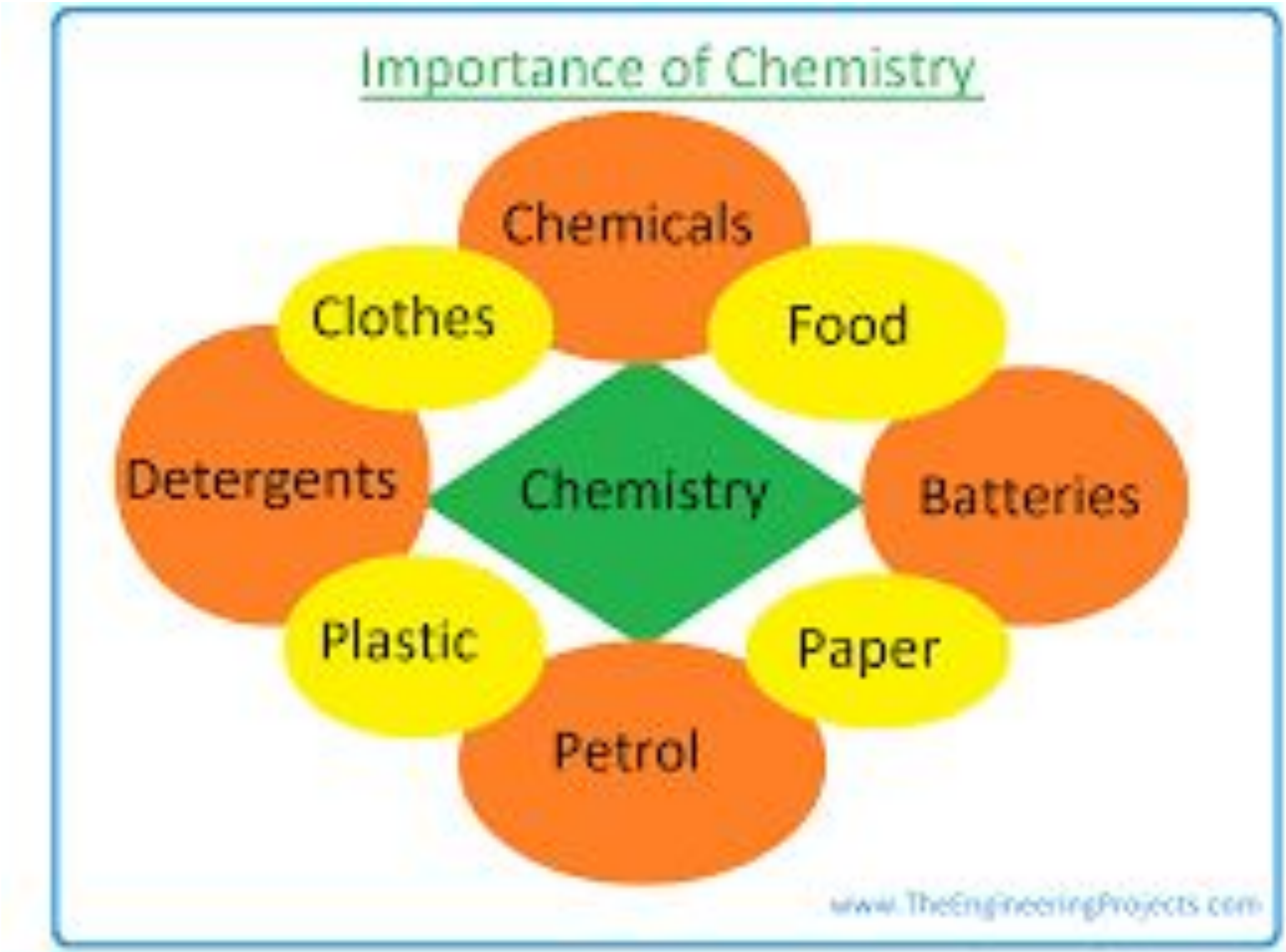
CHAPTER 1

FUNDAMENTAL CONCEPTS

I. Introduction

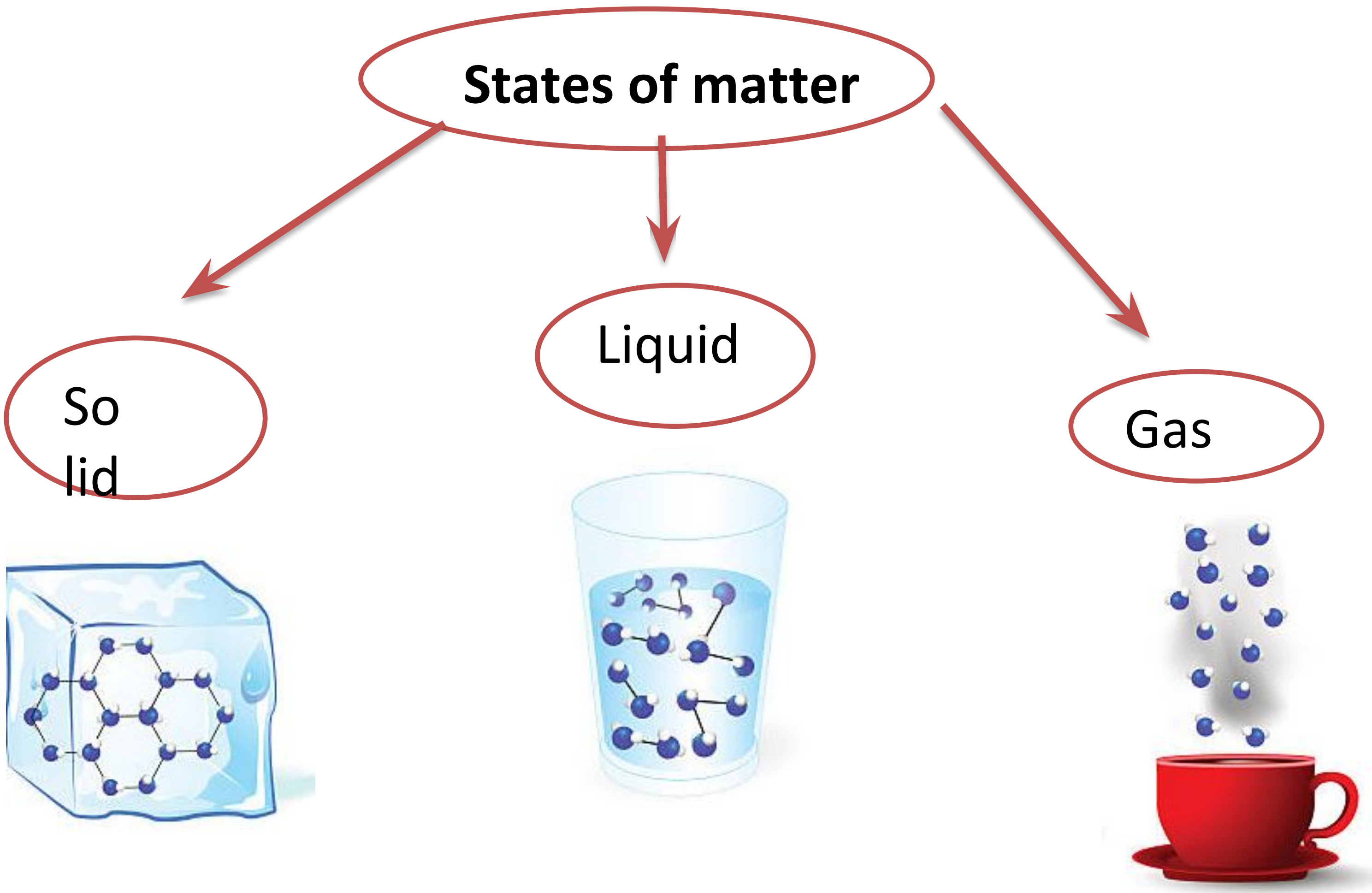
Chemistry is the science of matter and its transformations.



II. Matter

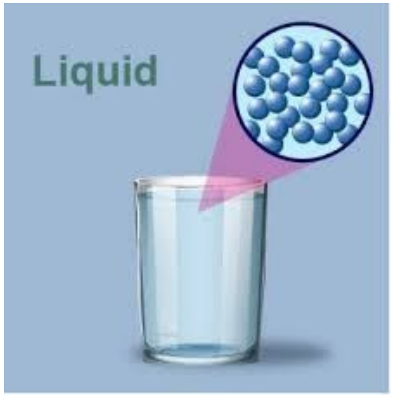
- Matter is anything that has mass and occupies space.

- Matter can exist in different forms, known as the states of matter.

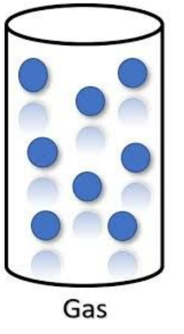


- **Solid** is a rigid form of matter. It has a definite shape and a fixed volume that changes very little with temperature or pressure.

- **Liquid** is a fluid form of matter with a definite volume and a definite surface, but it assumes the shape of its container

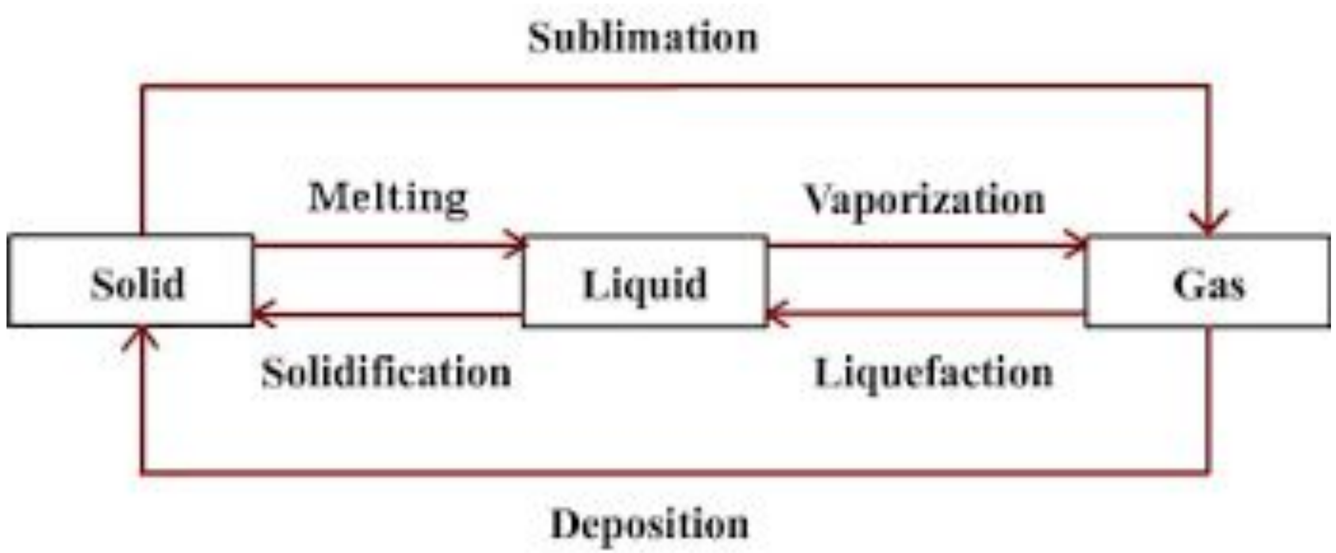


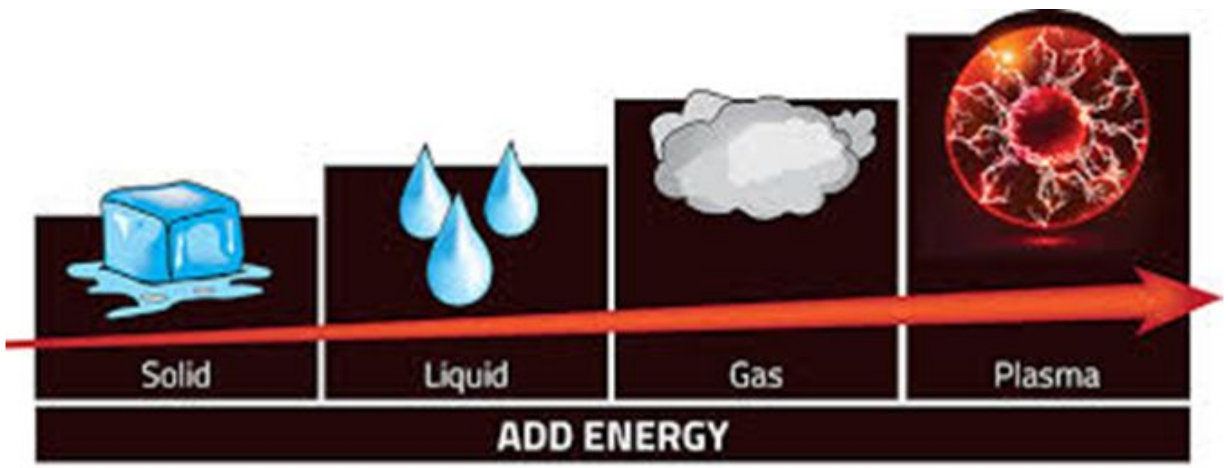
- **Gas** is a fluid form of matter that fills the whole volume of its container, and its pressure depends on temperature.



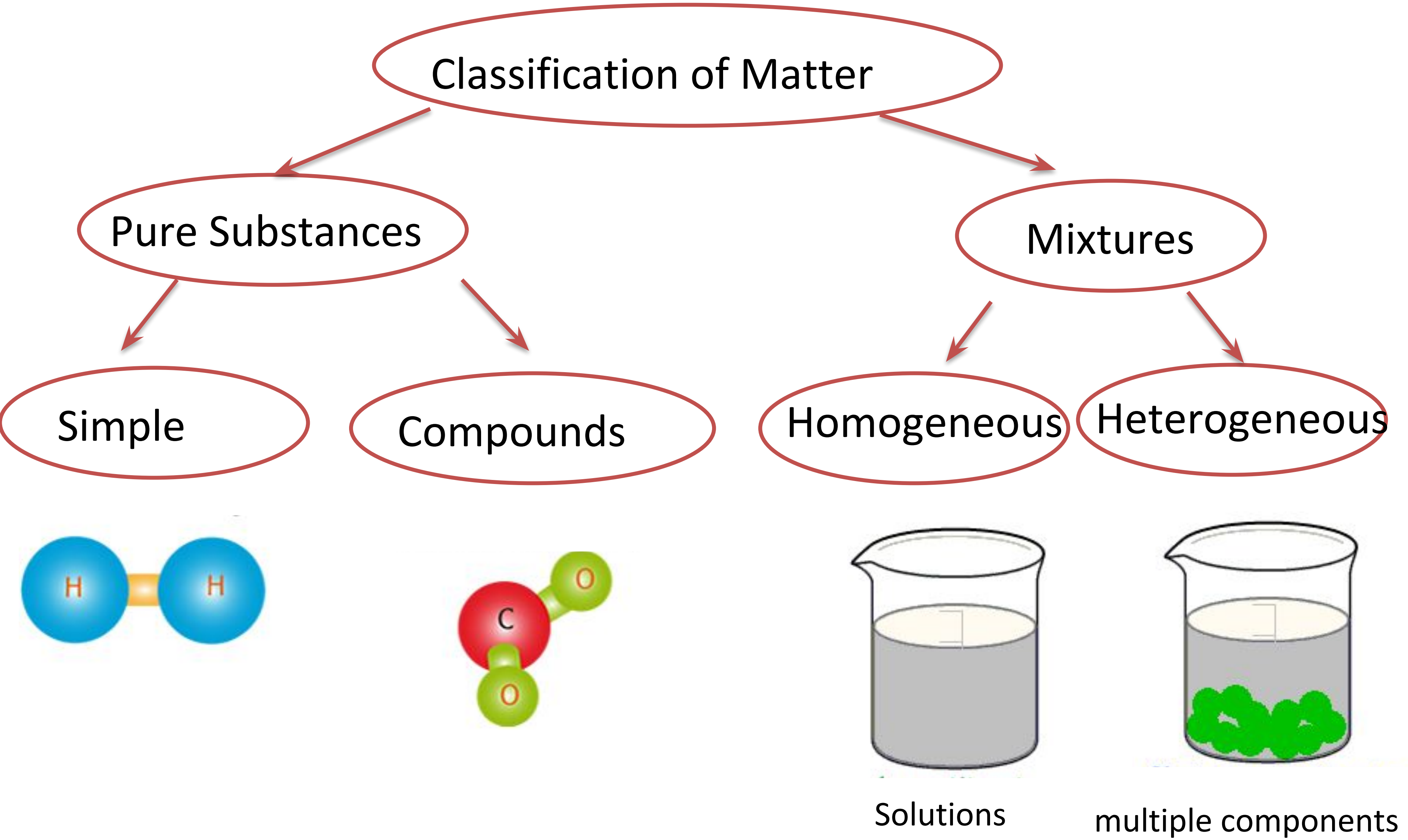
✔ **Plasma** is an ionized state of matter consisting of a mixture of free electrons and ions. It has no definite shape or volume, and it conducts electricity and responds strongly to magnetic and electric fields. Plasma is often referred to as the fourth state of matter

II-1. Change of state of matter



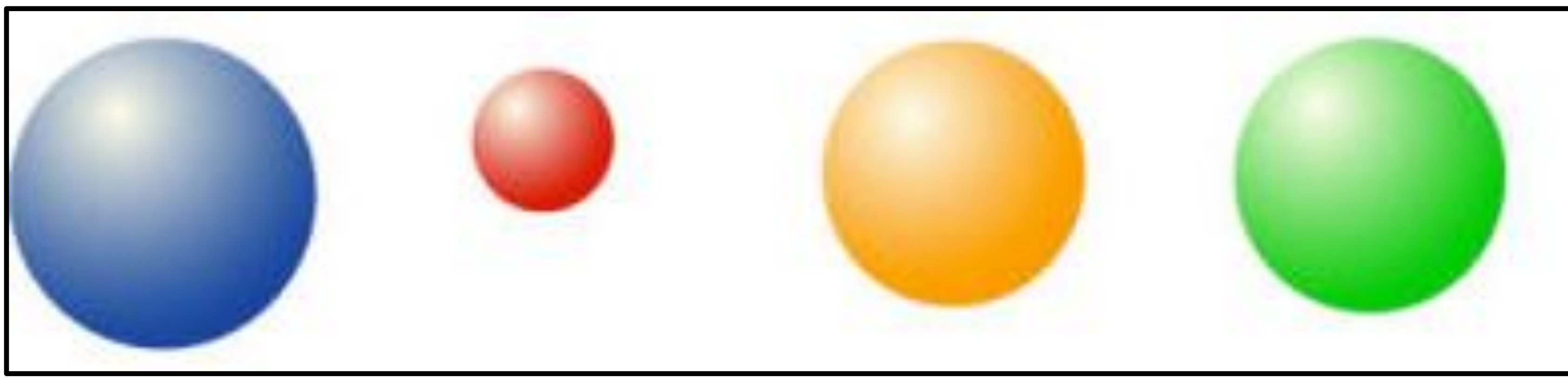


II-1. Classification of Matter



III. Atoms and Molecules

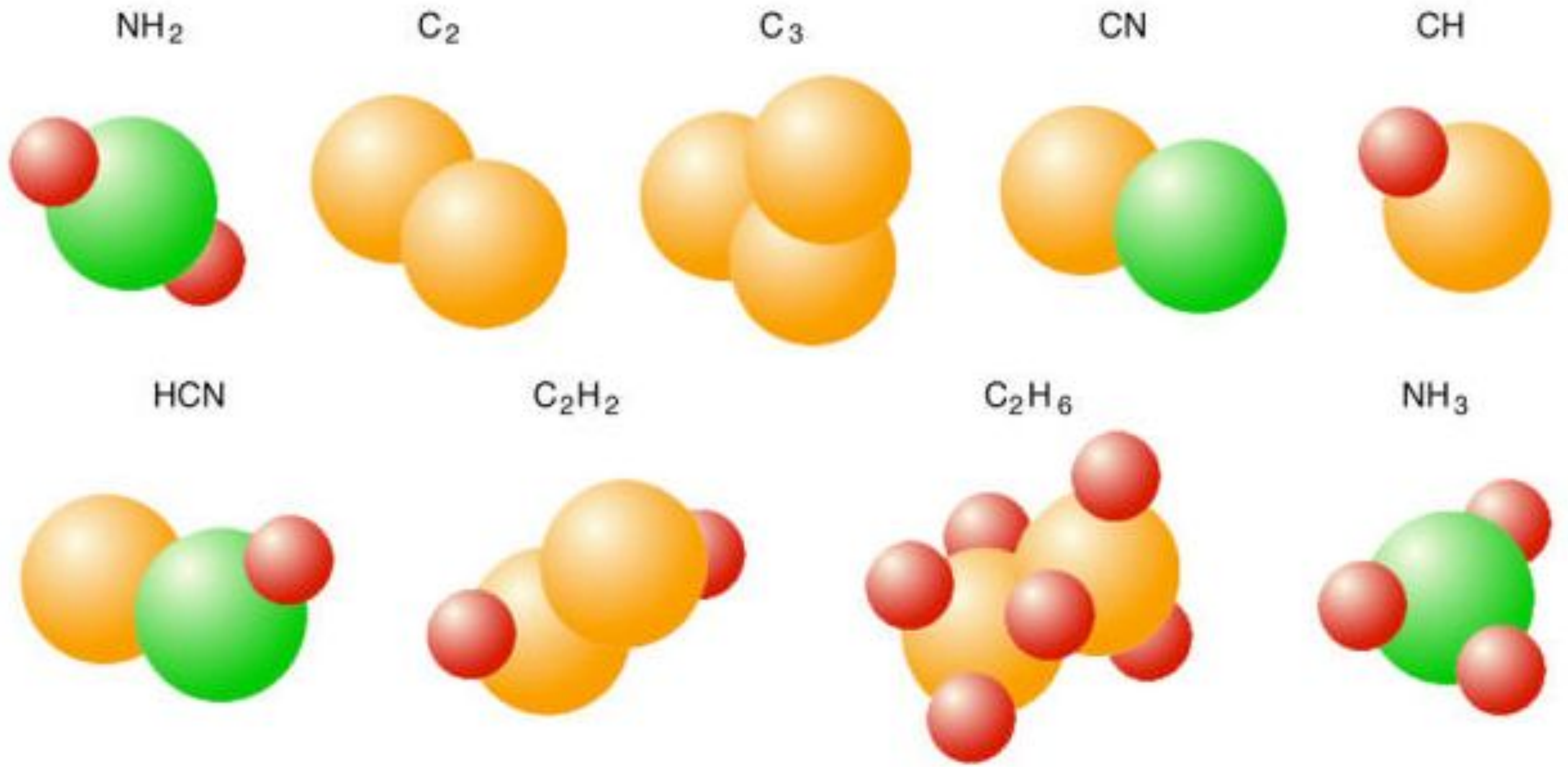
- An atom is the smallest unit of an element that preserves all of its distinctive properties.



Oxygen	Hydrogen	Carbon	Nitrogen
O	H	C	N

III. Atom and Molecule

- A molecule is the smallest unit of matter, composed of at least two atoms—whether alike or different—that retains the properties of the substance.



IV. Measurement Units

Measurements are expressed in the International System of Units (SI). These units are derived from the fundamental base units shown in the table below.

BASE QUANTITY	NAME OF UNIT	SYMBOL
Length	Meter	m
Mass	Kilogram	kg
Time	Second	s
Electrical current	Ampere	A
Temperature	Kelvin	K
Amount of substance	Mole	mol
Luminous intensity	Candela	cd

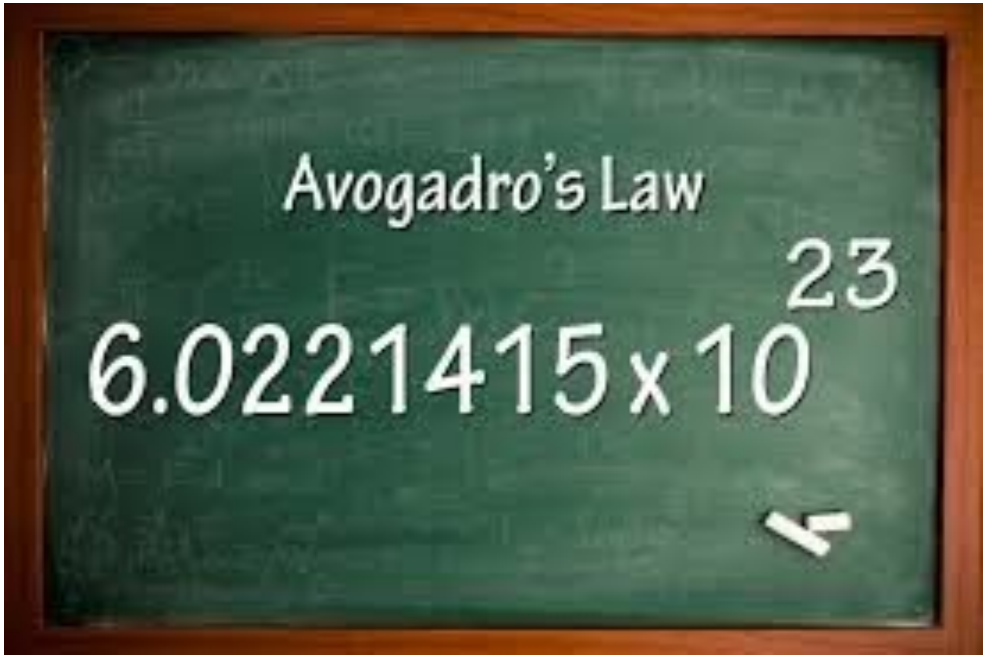
The prefixes applied to SI units are summarized in the table below.

SCIENTIFIC PREFIXES

Prefix	Symbol	Exponent	Prefix	Symbol	Exponent
yotta	Y	10^{24}	deci	d	10^{-1}
zetta	Z	10^{21}	centi	c	10^{-2}
exa	E	10^{18}	milli	m	10^{-3}
peta	P	10^{15}	micro	μ	10^{-6}
tera	T	10^{12}	nano	n	10^{-9}
giga	G	10^{9}	pico	p	10^{-12}
mega	M	10^{6}	femto	f	10^{-15}
kilo	k	10^{3}	atto	a	10^{-18}
hecto	h	10^{2}	zepto	z	10^{-21}
deca	da	10^{1}	yocto	y	10^{-24}

V. Definition of the Mole

- A mole is the amount of substance in a system that contains the same number of elementary entities as there are atoms in 12 grams of carbon-12.
- One mole contains Avogadro's number of entities: $N = 6.02 \times 10^{23}$.



VI. Definition of the Molar Mass of an Element

- The molar mass of an element is the mass, in grams, of one mole of that element.

- In numerical value, it is equal to the atomic mass of the element, expressed in atomic mass units (u).

VI. Definition of the Atomic Mass Unit (a.m.u)

The atomic mass unit is denoted by the symbol u.

By convention, 1 u = 1/12 of the mass of a carbon atom.

Mass of a carbon atom, denoted m_{c12}

$$m_{C12} = \frac{M_{C12}}{N_a}$$

M_{C12}: Molar mass of carbon−12

N_a: The Avogadro number

- It is 1/12 of the mass of a carbon-12 atom.

$$1\ \mathrm{u} = \frac{M_{12C}}{N_A} \times \frac{1}{12} = \frac{12{,}0000 \times 10^{-3}}{6{,}02214 \times 10^{23}} \times \frac{1}{12} = 1{,}66054 \times 10^{-27}\ \mathrm{kg}$$

$1\mathrm{u} = 1{,}66\ 10^{-27}$ kg $= 1{,}66\ 10^{-24}$g

Example

Electron = 0.00055 amu

Proton = 1.00728 amu

Neutron = 1.00867 amu

Oxygen = 15.9994 amu

Neon = 20.179 amu

Sodium = 22.9897 amu

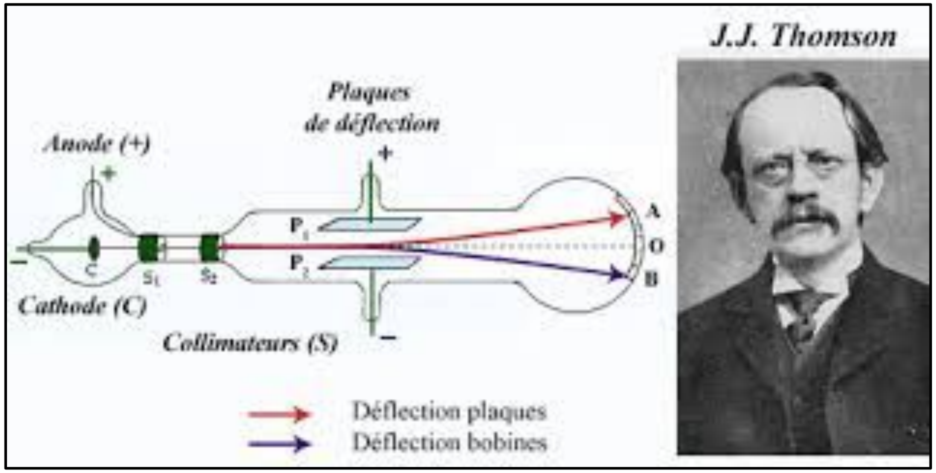
It is defined as 1/12 of the mass of an unbound neutral atom of carbon-12

Atoms

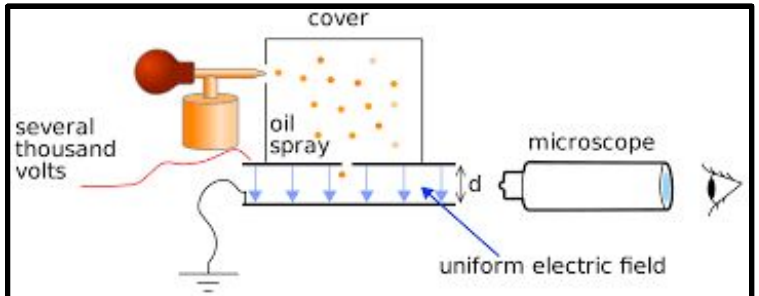
I. Atomic Structure

The experiments that led to the discovery of the nature of the components of the atom are:

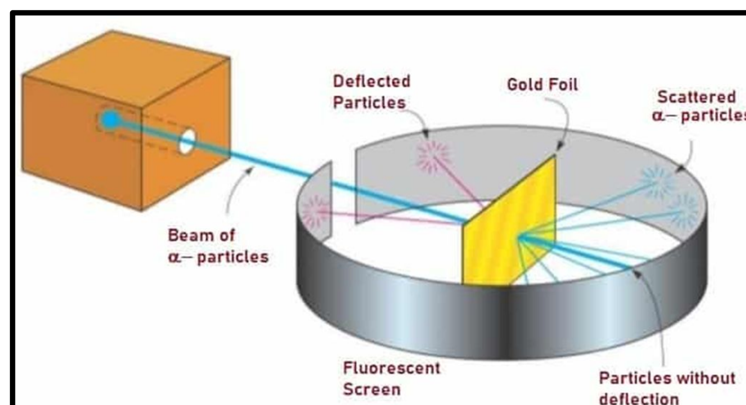
- Eugene Goldstein used discharge tubes to partially demonstrate the existence of positively charged particles.
- J. J. Thomson conducted an experiment using a cathode ray tube to measure the charge-to-mass ratio (e/m) of the electron.



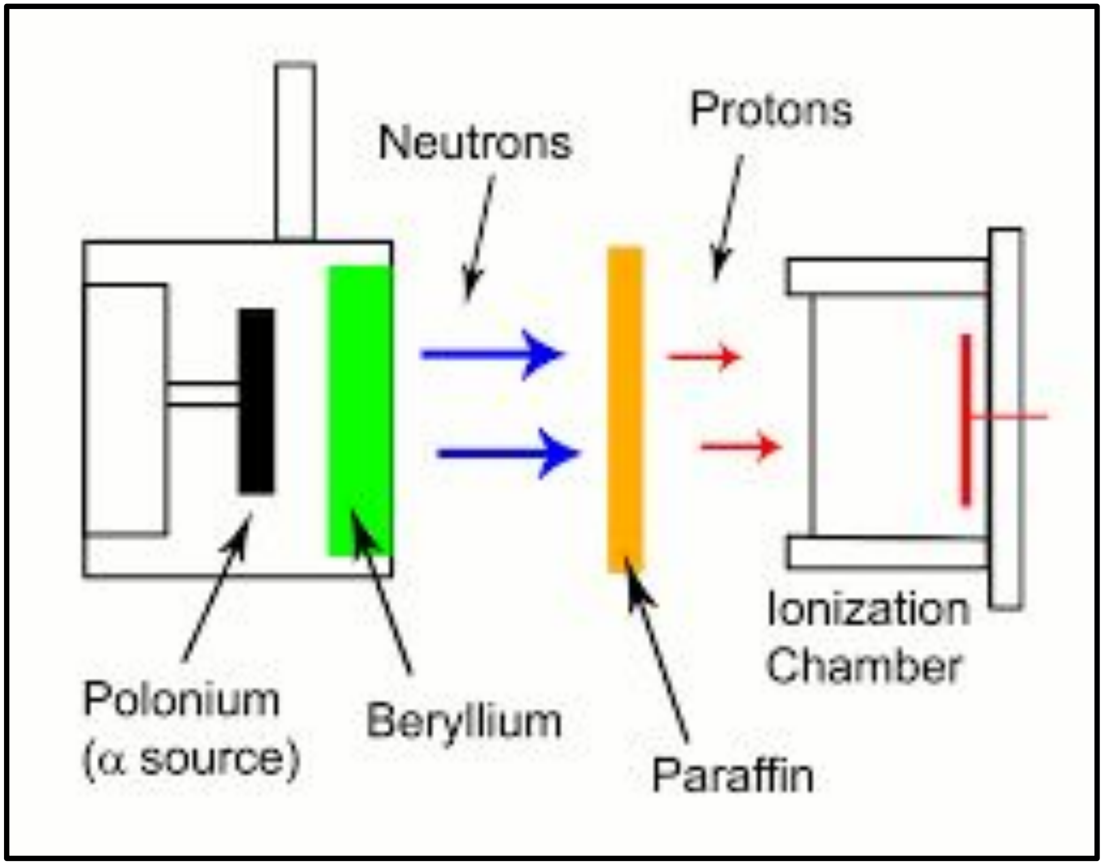
 Robert A. Millikan conducted the oil drop experiment to determine the electric charge of the electron



 By bombarding nitrogen gas with alpha particles, Ernest Rutherford was able to observe the emission of hydrogen nuclei, which he identified as protons.



- James Chadwick conducted experiments that led to the discovery of the neutron, a neutral particle located in the atomic nucleus

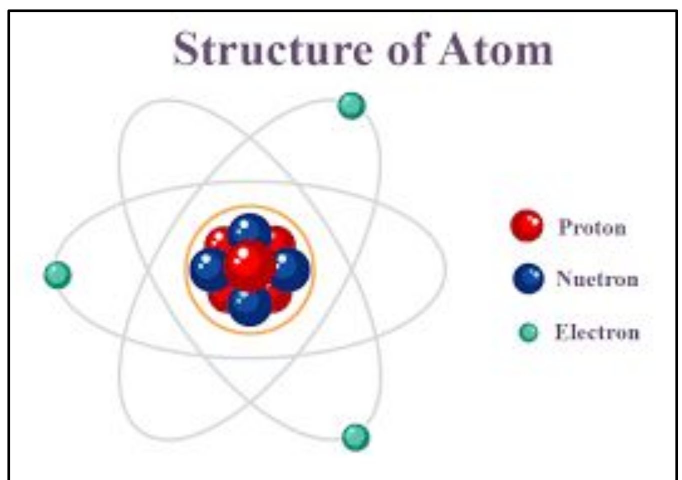


Atoms consist of three main subatomic particles: positively charged protons, neutral neutrons, and negatively charged electrons.

The characteristics of these particles are outlined in the following table:

particle	Mass		Charge	Symbol	Location
	(g)	(u)			
Electron	$9{,}109382\ 10^{-28}$	0,0005485 799	-1	e^-	Outside the nucleus
Proton	$1{,}672622\ 10^{-24}$	1,007 276	+1	p^+	Inside the nucleus
Neutron	$1{,}674927\ 10^{-24}$	1,008 665	0	n^0	Inside the nucleus

The electric charge of the electron is $e^{-} = 1.602\ 176 \times 10^{-19}$ C, which corresponds to a value of (–1). The nucleus of an atom is its central part, containing all the positive charge and nearly all its mass. It is made up of protons and neutrons. The components of the nucleus (protons and neutrons) are called nucleons.



JUST FOR YOUR INFORMATION

Scientists have also discovered other tiny particles like neutrinos, and particles called quarks, which are smaller parts of protons and neutrons — but not of electrons.

Moreover, it has been discovered that electrons are part of a group of fundamental particles known as leptons.

II. Atomic Number and Atomic Mass

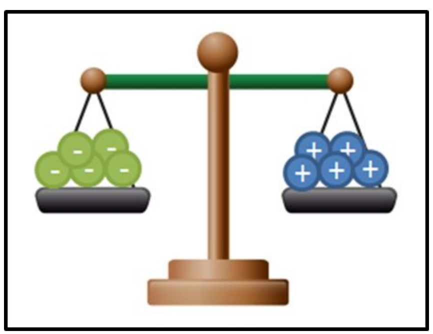
II-1. Atomic Number (Z)

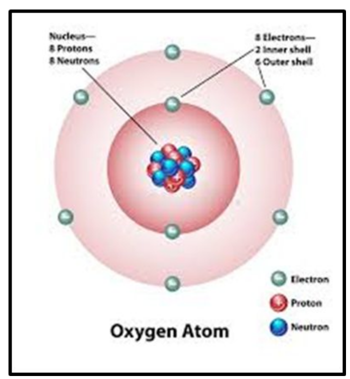
- Every atom of a given element contains the same number of protons in its nucleus

- The number of protons in an element's nucleus, denoted by Z, is called its atomic number, which essentially defines the identity of the element.

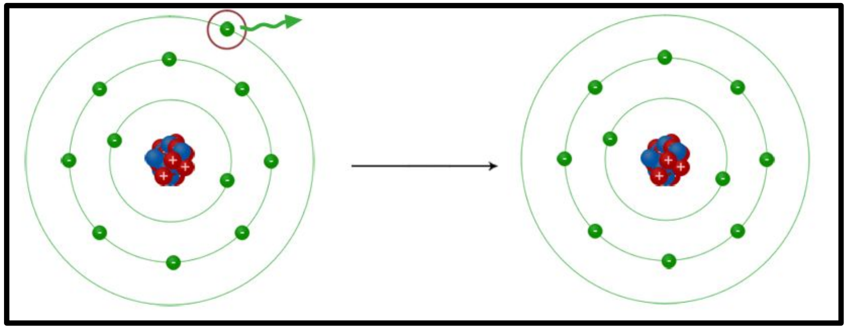
- For a neutral atom, the number of protons equals the number of electrons.

$$\text{Number of } e^- = \text{number of protons}$$





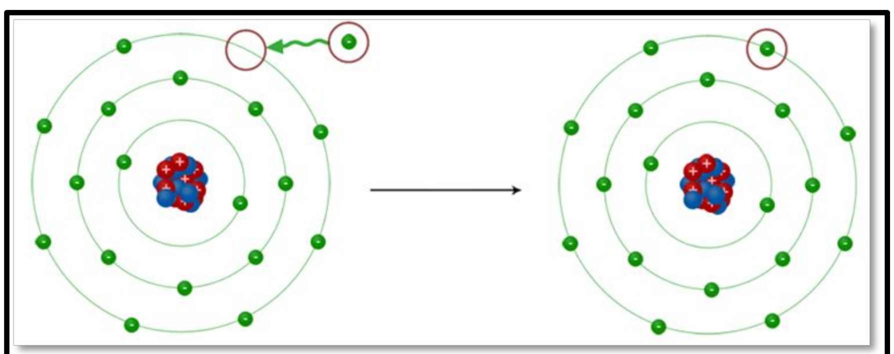
A cation is an atom that has a positive charge due to the loss of one or more electrons.



Electrons < Protons

Na **Na^+**

An anion is an atom that has a negative charge due to the gain of one or more electrons.



Electrons > Protons

Cl **Cl^-**

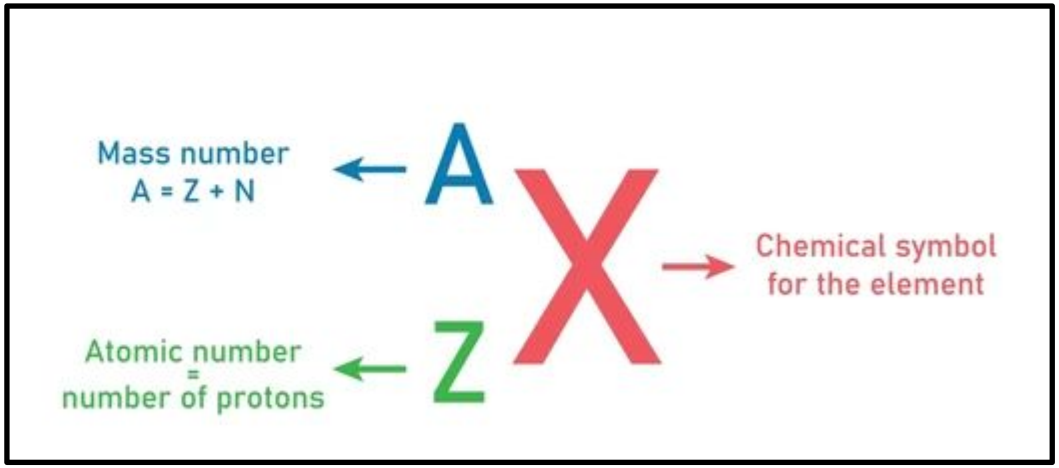
II-2. Mass Number (A)

An atom's mass in atomic units (u) is roughly the sum of its protons and neutrons.

This is known as the mass number, denoted by A.

Mass Number (A) = Number of Protons + Number of Neutrons

The atom can be represented using the following symbolic notation:



III. Isotopes

Isotopes are atoms of the same element that share the same atomic number (Z) but differ in their mass number (A).

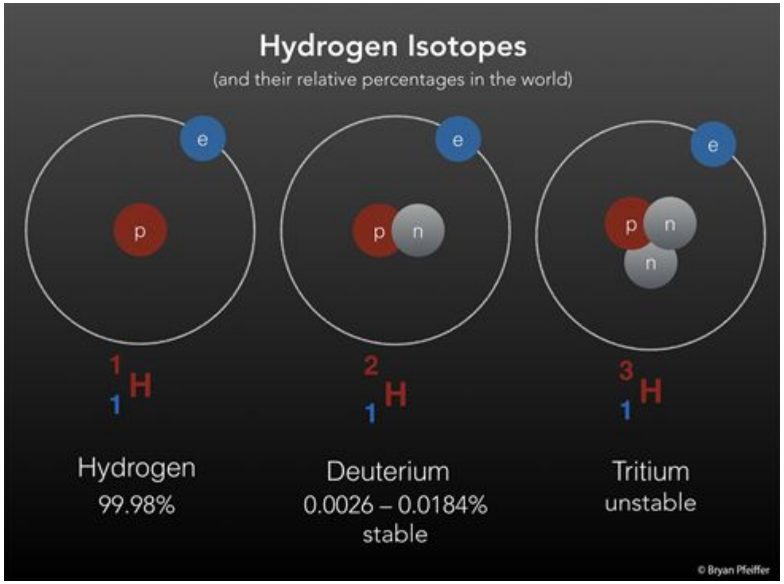
Example:

The isotopes of hydrogen, Z = 1:

^{1}H, protium, more commonly called hydrogen or the proton (p).

^{2}H, deuterium (D).

^{3}H, tritium (T).



IV. Isotopic Abundance

An element is often composed of several isotopes present in different proportions, called abundance.

$$\boldsymbol{Isotopic\ abundance = \frac{number\ of\ atoms\ of\ a\ given\ isotope}{total\ number\ of\ atoms\ of\ all\ isotopes\ of\ that\ element} \times 100}$$

- Isotopic abundance refers to the mole fraction of a given isotope in a mixture, expressed as a percentage.
- The total of all isotopic abundances always adds up to 100.
- The isotopic composition of an element is therefore described by the isotopic abundances of its constituent isotopes

Isotopic abundance = percentage of an element’s isotopes.

V. Atomic mass

The atomic mass of an atom is the weighted average of the masses of an element’s naturally occurring isotopes, expressed in atomic mass units (u). It can be calculated using the following formula::

$$Average\ atomic\ mass = \sum_{i=1}^{n} m_i \times a_i$$

Where:

mi = mass of isotope *i*

ai = relative abundance of isotope *i* (expressed as a decimal)

n = total number of isotopes .

Average atomic mass = Atomic mass shown on the periodic table, which takes into account the isotopes and their relative abundances.

Example :

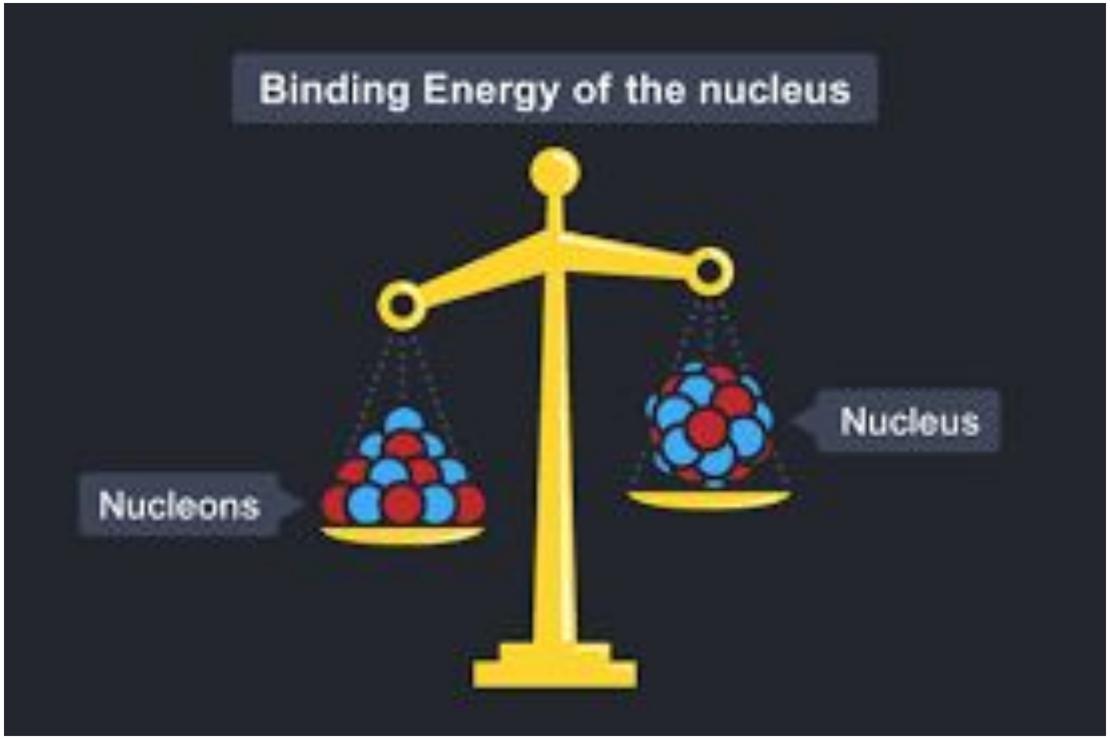
If an element has two isotopes:

- Isotope A: mass = 10 u, abundance = 20% → 0.20
- Isotope B: mass = 11 u, abundance = 80% → 0.80

Then:

Average atomic mass$=(10\times0.20)+(11\times0.80)=2+8.8=10.8$

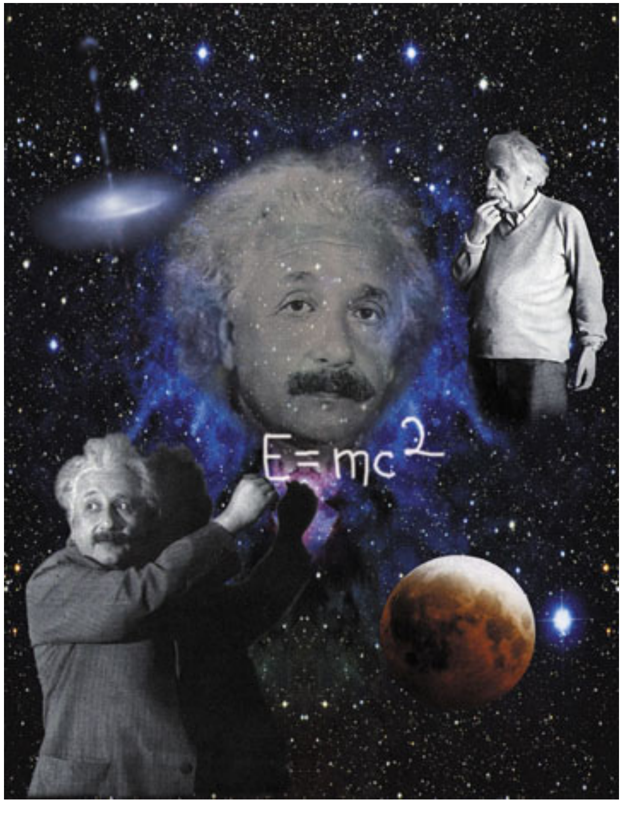
VI. Atomic binding energy



This raises a question: where, then, has the missing mass gone?

Could there be a connection between mass defect and binding energy? Let us put the question to Albert Einstein.

Albert Einstein postulated the principle of mass–energy equivalence in 1905.



VI-1. Einstein's theory of relativity

An object's mass reflects its energy content; therefore, the total energy E and the mass m are connected through Einstein's famous equation:

$$\boxed{E = m\,c^2}$$

The energy E is expressed in joules (J).
The mass m is expressed in kilograms (kg).
The speed of light in vacuum is $c = 2.9979 \times 10^8\ ms^{-1}$
approximately $c \approx 3.0 \times 10^8\ ms^{-1}$

- This equation demonstrates that whenever energy is lost, there is also a corresponding loss of mass.

IV-2. Nuclear binding energy (E_{bn})

The nuclear binding energy (E_{bn}) is the energy released as protons and neutrons bind to form a nucleus.

$$Z\,p^+ + (A - Z)\,n^0 \longrightarrow {}^{A}_{Z}X + \Delta E.$$

The binding energy can be calculated using the following equation: :

$$E_{bn} = |\Delta m| \times C^2$$

$$\Delta m = \sum m(products) - \sum m(reactants)$$

$$\Delta m = m({}^{A}_{Z}X) - (Zm_p + (A - Z)m_n)$$

Binding energies are usually expressed in electron volts (eV) or in mega-electron volts (MeV).

1 electronvolt (eV) is the amount of kinetic energy gained or lost by a single electron when it is accelerated through an electric potential difference of one volt.

$$1eV = 1,602\ 18 \times 10^{-19}\ J$$

$$1\ MeV = 10^{6}\ eV$$

Since nuclide masses are very small, they are usually given in atomic mass units (u).

The energy that corresponds to a mass change of 1 atomic mass unit (u):

$$E_{bn} = |\Delta m| \times C^{2}$$

$$|\Delta m| = 1u = 1,66\ 10^{-27} kg$$

$$c = 3\ 10^{8} ms^{-1}$$

$$E = 1,66\ 1O^{-27}\ \times (3\ 10^{8})^{2} \qquad E = 14,925 \times 10^{-11} J$$

$$1eV \rightarrow 1,6 \times 10^{-19} J$$

$$E \rightarrow 14,925 \times 10^{-11} J \qquad E = 931,5\ 10^{6} eV$$

$$1MeV \rightarrow 10^{6} eV$$

$$E\ (MeV) \rightarrow 931,5\ 10^{6} eV \qquad E = 931,5\ MeV$$

$$\boxed{\Delta m = 1u \longrightarrow E = 931,5\ MeV}$$

Bending energy per nucleon

The binding energy is often expressed as energy per nucleon, that is, the amount of energy needed to separate a nucleon from the nucleus, and it can be calculated using the following equation:

$$E_{bn}(nucleon) = \frac{E_{bn}}{A}$$

A: the mass number of the atom..

The stability of nuclei **increases** as the binding energy per nucleon becomes **higher**

For the helium nucleus

Nuclear mass

$m_n = 6{,}6647 \times 10^{-27}$ kg

Nucleide masses

$= 2\ m_p + 2\ m_n$

$= 2 \times 1{,}6726 \times 10^{-27} + 2 \times 1{,}6750 \times 10^{-27}$

$= 6{,}6952 \times 10^{-27}$ kg

The mass defect is equal to Δm = 0,0305 x 10^{-27}kg

Therefore, the mass defect calculated for the helium nucleus corresponds to its binding energy.

$$E_1 = \Delta m.c^2 = 0{,}0305 \times 10^{-27} \times (2{,}9979 \times 10^8)^2 = 2{,}741 \times 10^{-12}\ \text{J}$$