University IBN KHALDOUN of Tiaret

Radiation Biophysics

Mrs CHIBANI F

Second Semester

- Atomic Structure
- Radiation
- Radioactivity (Physics)
- Radioactivity (Radioactive Transformations)
- Course and Exercises
- Radiation/Particle-Matter Interactions
- Dosimetry
- Radioprotection
- Course and Exercises
- Principles of X-Ray Production
- Physics of Lasers and Medical Applications
- Ultrasound Physics: Echography and Doppler
- Physics Basics of MRI

Atomistics:

Atomistics is a discipline that experienced significant development during the 20th century, evolving from a primarily philosophical stage to a full-fledged discipline integrated into all modern sciences.

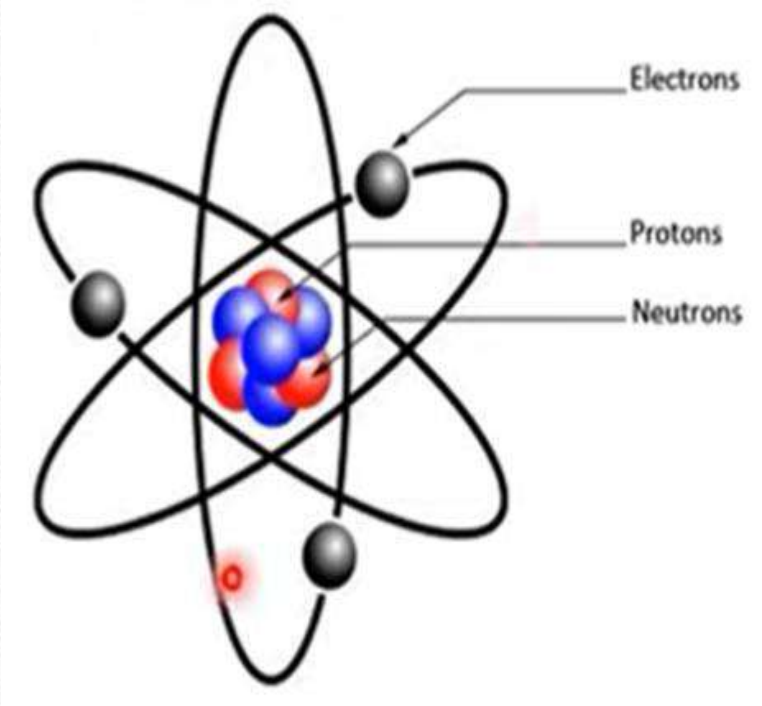
In other words, atomistics is an interdisciplinary branch of science that studies in detail the structure, properties, and behavior of the fundamental particles of matter at the atomic and subatomic scale. It plays a crucial role in many scientific fields, including chemistry, physics, biology, and materials science, and provides the necessary theoretical foundations for understanding the nature of matter and developing new technologies.

Structure of Matter:

Matter exists in three states: solid, liquid, and gas. However, the entire physical universe, including living organisms, is made up of matter. It is characterized by its mass and its energy, which measures its capacity to produce work. Matter is formed from elementary grains, which are atoms.

General Information on the Atom

An atom consists of a nucleus around which one or more electrons orbit.
The nucleus of the atom is composed of nucleons: neutrons and protons.
The atom is electrically neutral (it contains as many protons as electrons).



118 atoms or elements have been discovered (as of 2022), and each one isdesignated by its name and symbol.

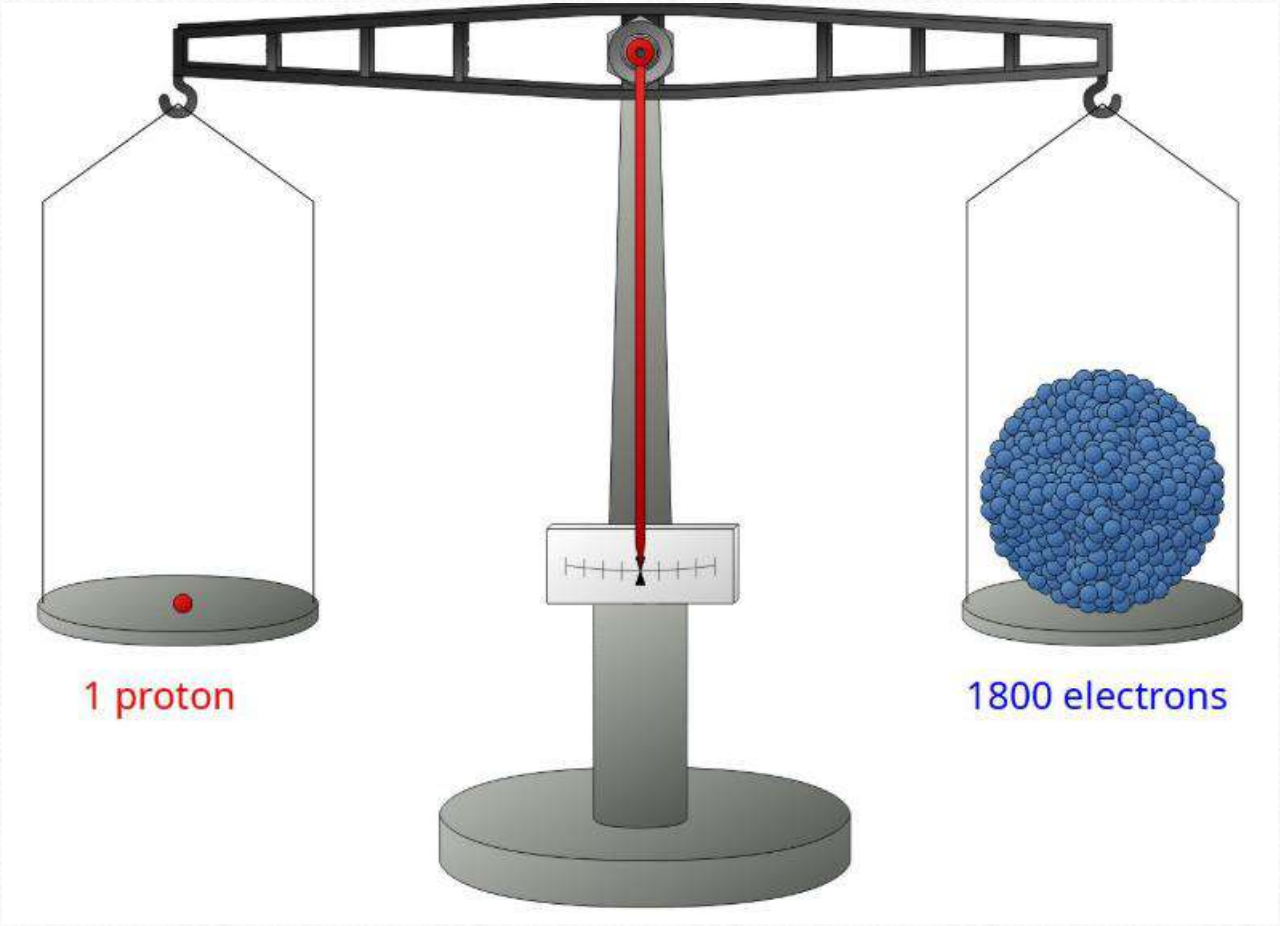
Example: Hydrogen: H; Iron: Fe; Oxygen: O; Carbon: C

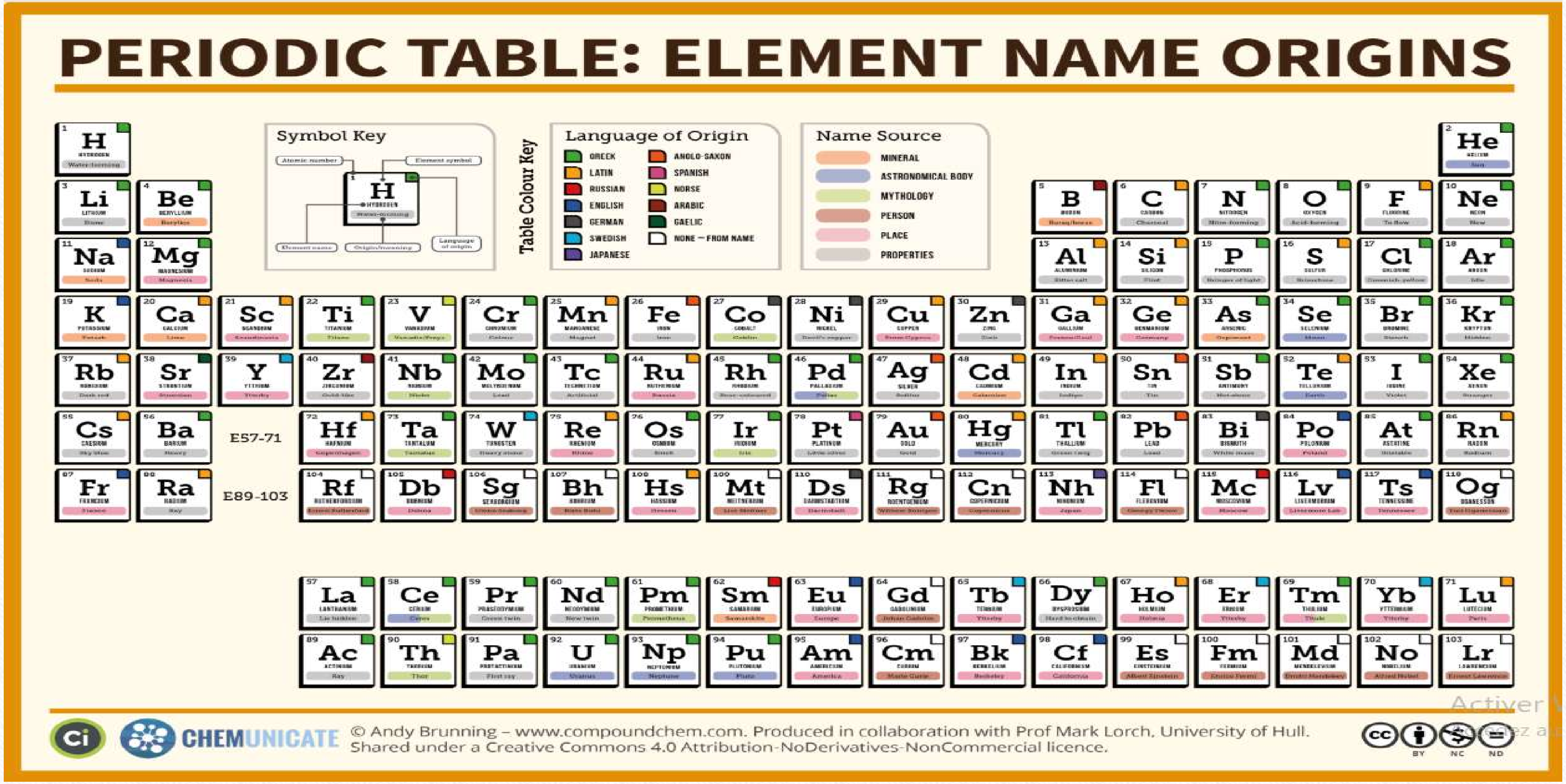
The mass of the atom is primarily contained in the nucleus, and the volume of the atom is mainly due to the size of the electron cloud.

1 amu = 1.6605×10^{-24}g = 1.6605×10^{-27} kg

Properties of fundamental particules

Name	Symbol	Absolute charge/C	Relative charge	Mass/kg	Mass/u	Approx. mass/u
Electron	e	-1.6022×10^{-19}	-1	9.10939×10^{-31}	0.00054	0
Proton	p	$+1.6022\times10^{-19}$	+1	1.67262×10^{-27}	1.00727	1
Neutron	n	0	0	1.67493×10^{-27}	1.00867	1


1 proton
1800 electrons


PERIODIC TABLE: ELEMENT NAME ORIGINS
Symbol Key
Atomic number
Element symbol
Element name
Origin/meaning
Language of origin
Table Colour Key
Language of Origin
GREEK
LATIN
RUSSIAN
ENGLISH
GERMAN
SWEDISH
JAPANESE
ANGLO-SAXON
SPANISH
NORSE
ARABIC
GAELIC
NONE – FROM NAME
Name Source
MINERAL
ASTRONOMICAL BODY
MYTHOLOGY
PERSON
PLACE
PROPERTIES
H He
Li Be B C N O F Ne
Na Mg Al Si P S Cl Ar
K Ca Sc Ti V Cr Mn Fe Co Ni Cu Zn Ga Ge As Se Br Kr
Rb Sr Y Zr Nb Mo Tc Ru Rh Pd Ag Cd In Sn Sb Te I Xe
Cs Ba E57-71 Hf Ta W Re Os Ir Pt Au Hg Tl Pb Bi Po At Rn
Fr Ra E89-103 Rf Db Sg Bh Hs Mt Ds Rg Cn Nh Fl Mc Lv Ts Og
La Ce Pr Nd Pm Sm Eu Gd Tb Dy Ho Er Tm Yb Lu
Ac Th Pa U Np Pu Am Cm Bk Cf Es Fm Md No Lr
CHEMUNICATE
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Periodic Table of the Elements

Nonmetals | Metals | Semimetals

ChemTalk
www.chemistrytalk.org

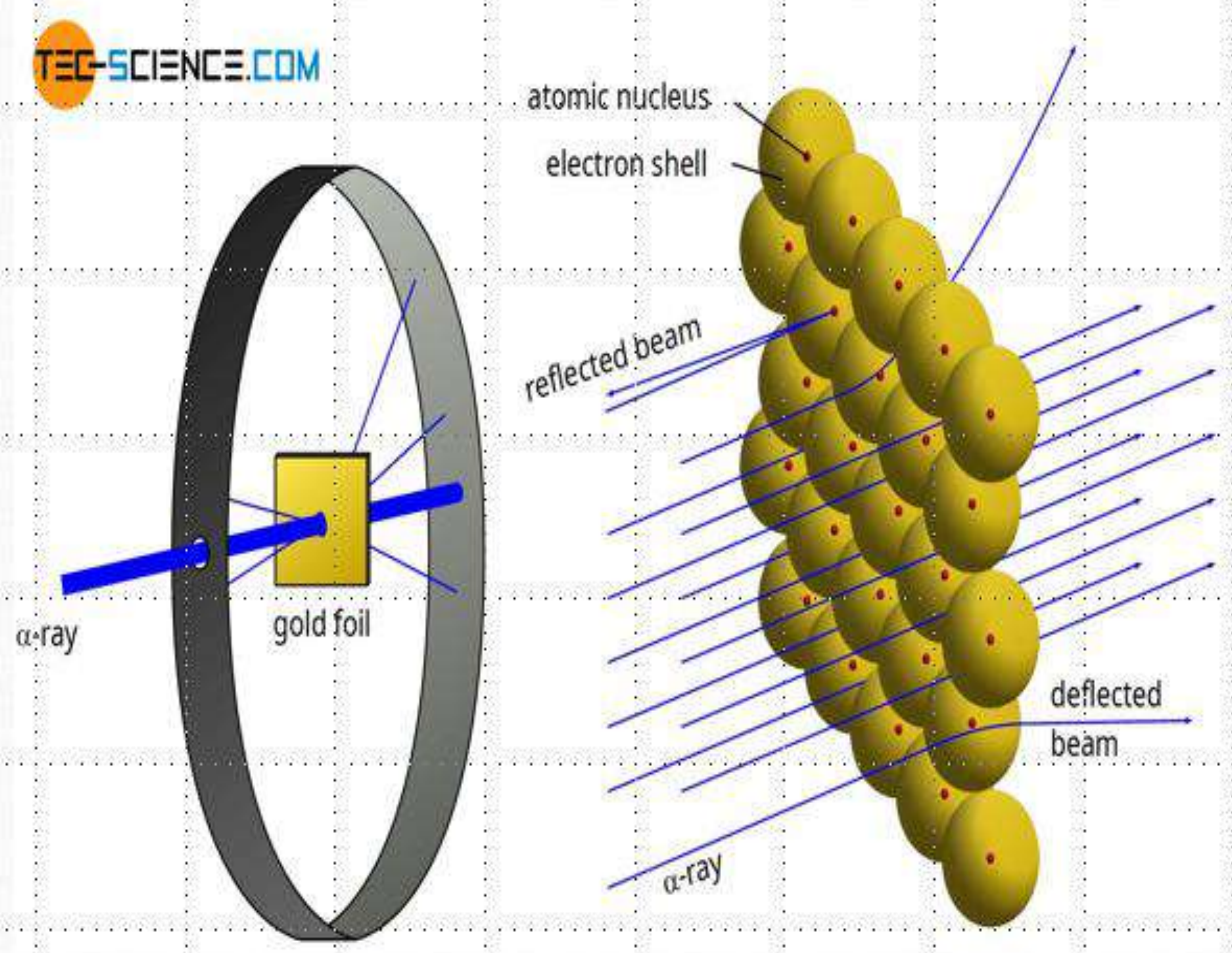
Hydrogen 1																	Helium 2
Lithium 3	Beryllium 4											Boron 5	Carbon 6	Nitrogen 7	Oxygen 8	Fluorine 9	Neon 10
Sodium 11	Magnesium 12											Aluminium 13	Silicon 14	Phosphorus 15	Sulfur 16	Chlorine 17	Argon 18
Potassium 19	Calcium 20	Scandium 21	Titanium 22	Vanadium 23	Chromium 24	Manganese 25	Iron 26	Cobalt 27	Nickel 28	Copper 29	Zinc 30	Gallium 31	Germanium 32	Arsenic 33	Selenium 34	Bromine 35	Krypton 36
Rubidium 37	Strontium 38	Yttrium 39	Zirconium 40	Niobium 41	Molybdenum 42	Technetium 43	Ruthenium 44	Rhodium 45	Palladium 46	Silver 47	Cadmium 48	Indiuom 49	Tin 50	Antimony 51	Tellurium 52	Iodine 53	Xenon 54
Caesium 55	Barium 56	Lanthanides (Below)	Hafnium 72	Tantalum 73	Tungsten 74	Rhenium 75	Osmium 76	Iridium 77	Platinum 78	Gold 79	Mercury 80	Thallium 81	Lead 82	Bismuth 83	Polonium 84	Astatine 85	Radon 86
Francium 87	Radium 88	Actinides (Below)	Rutherfordium 104	Dubnium 105	Seaborgium 106	Bohrium 107	Hassium 108	Meitnerium 109	Darmstadtium 110	Roentgenium 111	Copernicium 112	Nihonium 113	Flerovium 114	Moscovium 115	Livermorium 116	Tennessine 117	Oganesson 118

Lanthanum 57	Cerium 58	Praseodyium 59	Neodymium 60	Promethium 61	Samarium 62	Europium 63	Gadolinium 64	Terbium 65	Dysprosium 66	Holmium 67	Erbium 68	Thulium 69	Ytterbium 70	Lutetium 71
Actinium 89	Thorium 90	Protactinium 91	Uranium 92	Neptunium 93	Plutonium 94	Americium 95	Curium 96	Berkelium 97	Californium 98	Einsteinium 99	Fermium 100	Mendelevium 101	Nobelium 102	Lawrencium 103

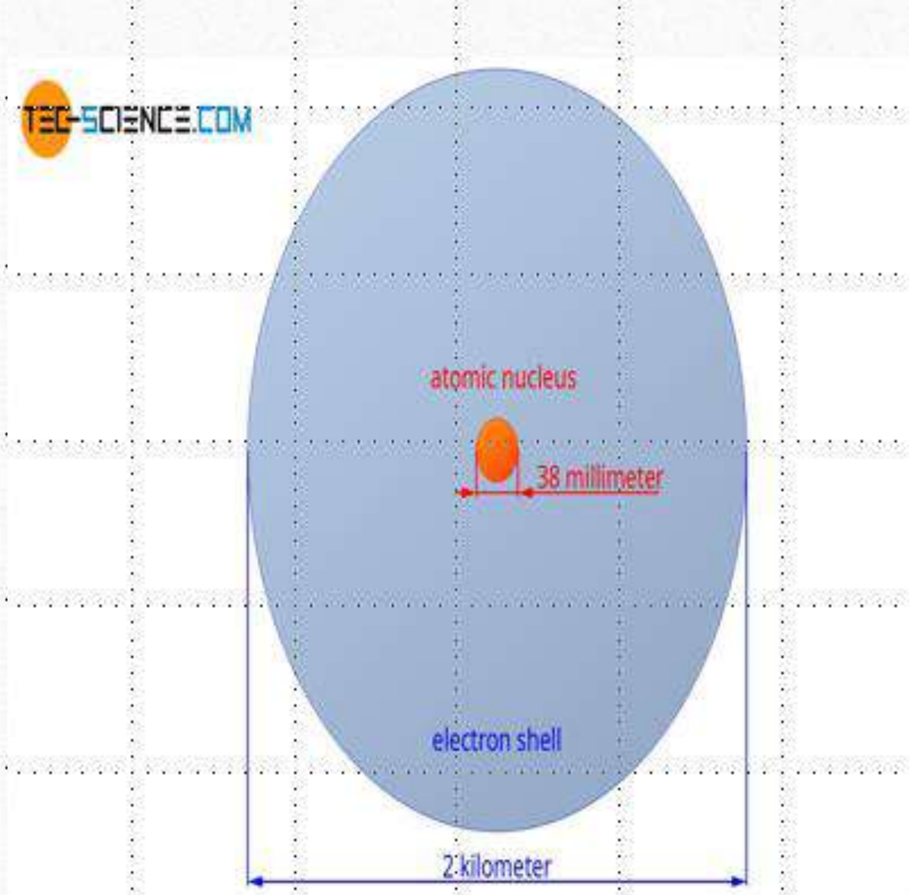
The nucleus:

It was discovered thanks to Rutherford's experiment. He deduced that the entire mass of the atom is concentrated in a positively charged region, which is the central nucleus. Negative electrons revolve around the nucleus like planets around the Sun. The nucleus is spherical, and its volume is calculated using the relation:$V = \frac{4}{3}\pi R^3$

The radius of the atom is 10,000 times larger than the radius of the nucleus:$\boldsymbol{R_{atom}} = \boldsymbol{10^4 R_{nucleus}}$

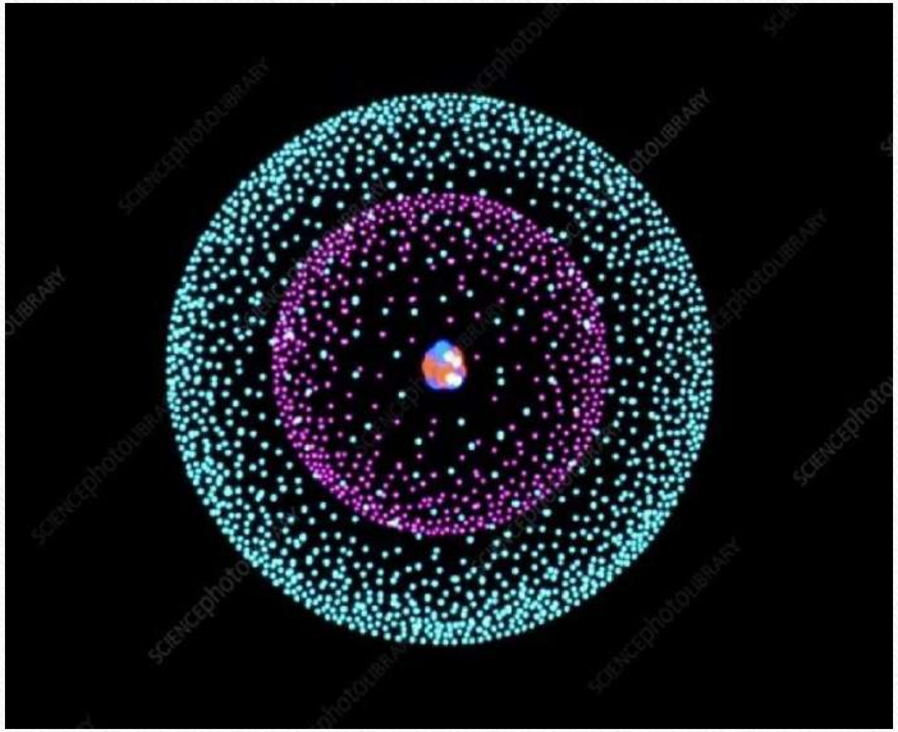


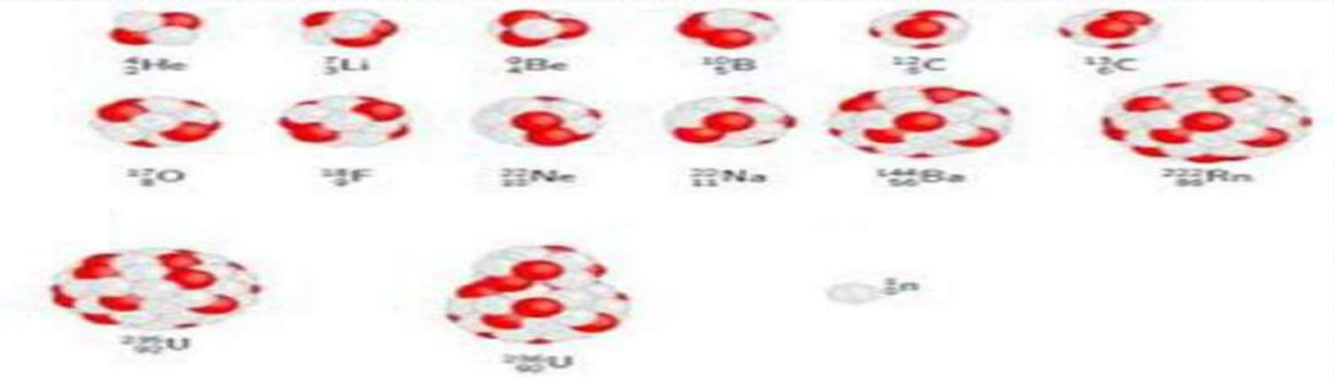
Rutherford's gold foil experiment



Comparison in size between an atomic nucleus and its electron shell

The radius of a nucleus is on the order of 10^{-15}m (femtometer).
It is composed of two types of particles (protons and neutrons), called nucleons, whose number varies from one atom to another.





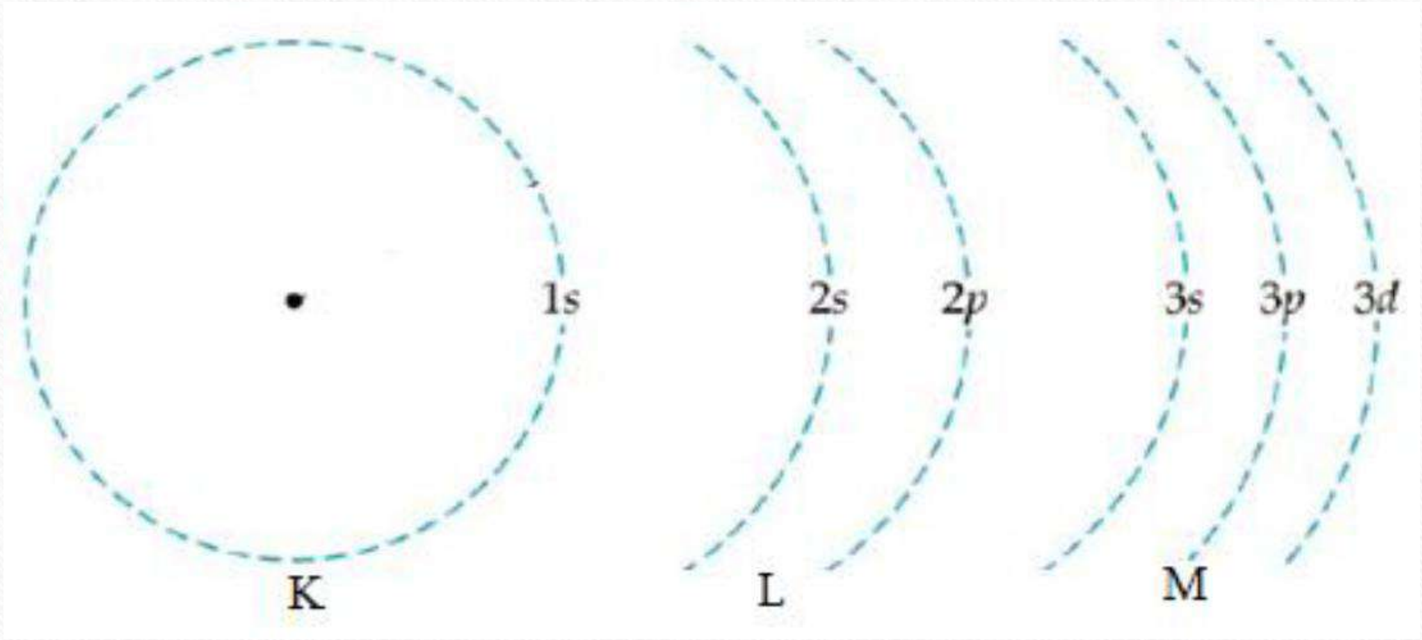
Atoms of the First Six Elements

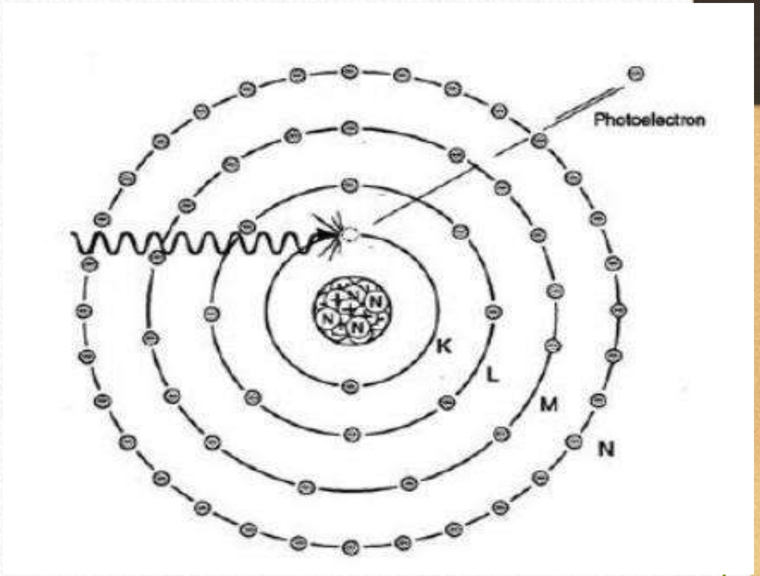
Name	Protons	Neutrons	Electrons	Atomic Number (Z)	Mass Number (A)
Hydrogen	1	0	1	1	1
Helium	2	2	2	2	4
Lithium	3	4	3	3	7
Beryllium	4	5	4	4	9
Boron	5	6	5	5	11
Carbon	6	6	6	6	12

Electron: An electron is an elementary particle with a negative charge that is one of the fundamental components of the atom. It is usually represented as a small particle orbiting around the atomic nucleus. Electrons are responsible for chemical bonding, as well as the electrical and magnetic properties of atoms and the behavior of conductive materials.

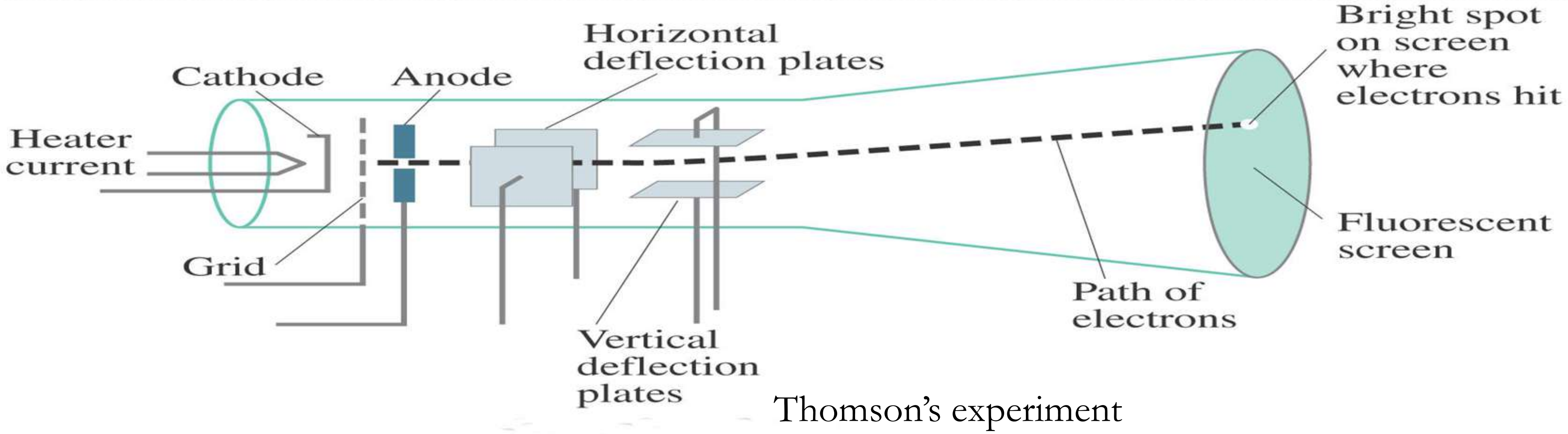
Electron cloud: The electron cloud is a probabilistic representation of the position of electrons around an atomic nucleus. Unlike classical representations of the atom with specific electron orbits, quantum mechanics states that it is impossible to determine precisely both the position and the velocity of an electron at a given moment.Instead, wave functions are used to describe the probability of finding an electron in a certain region of space. The electron cloud is therefore a representation of this probability of electron localization around the atomic nucleus. It is generally depicted as a region of higher electron density where the probability of finding an electron is greater, and lower density where the probability is smaller

Electrons orbiting around the nucleus move at finite distances. These electrons do not escape the atomic domain and remain bound to their trajectories thanks to an energy known as the binding energy. The electrons most strongly bound to the atom are those closest to the nucleus; as the distance from the nucleus increases, this binding energy decreases. Such an electron can be easily removed from the atomic structure (13.6 eV).

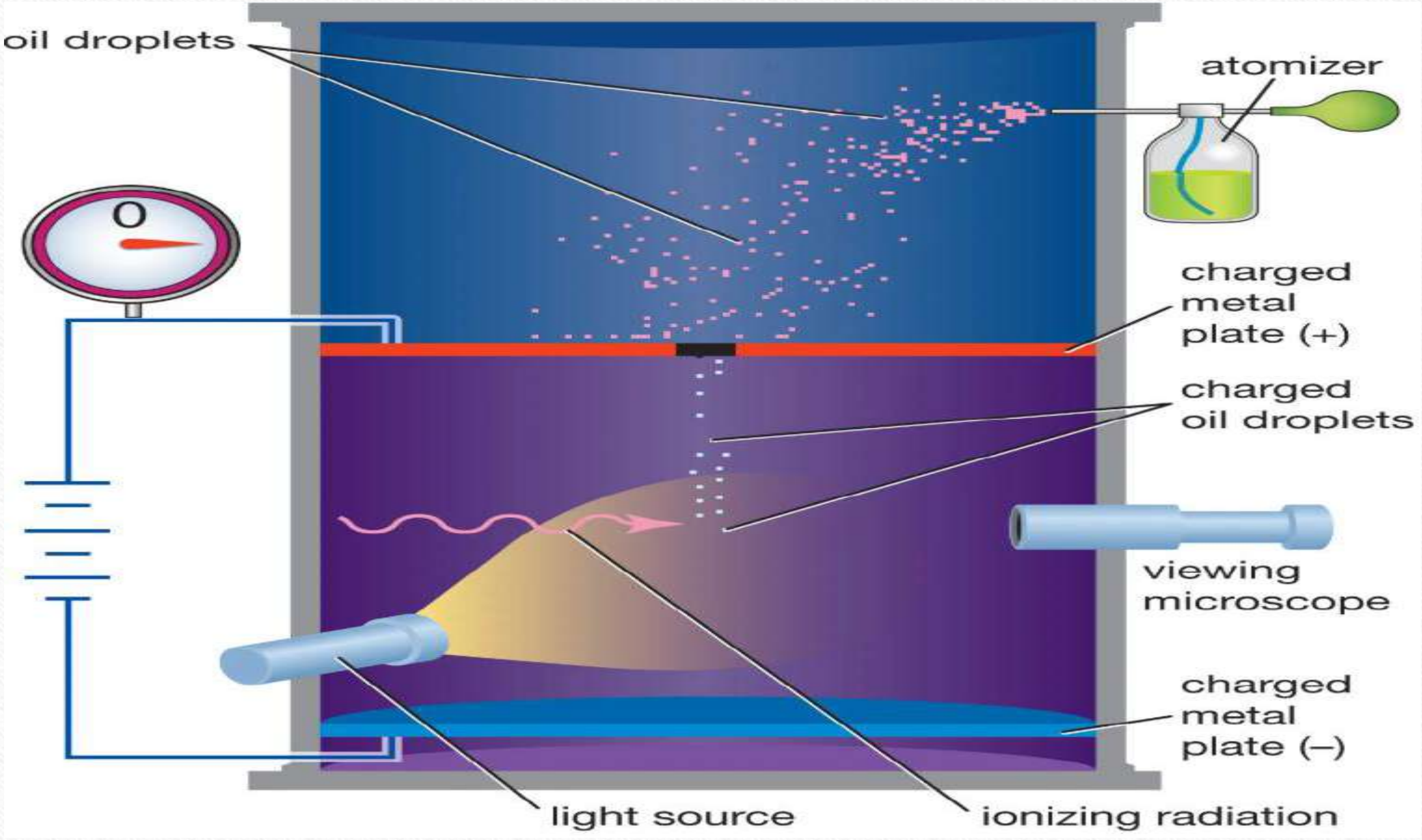




The negative charge of the electron was demonstrated by J. Perrin (a French physicist) in 1895 through the so-called cathode ray experiment. In 1897, the English physicist J. J. Thomson determined the charge-to-mass ratio of this elementary particle. In 1906, Thomson showed that the hydrogen atom contains only one electron. In 1909, the American physicist R. Millikan, using the oil drop experiment, measured the charge of the electron.



Thomson's experiment



Millikan oil-drop experiment

An atom is characterized by two non-zero integers, denoted A and Z, and is represented by a nuclide. A nuclide is written as $^{A}_{Z}\boldsymbol{X}$

X: is the name (symbol) of the nuclide,

A: is the mass number, corresponding to the total number of nucleons (protons and neutrons),

Z: is the atomic number, corresponding to the number of protons and also the number of electrons.

$$\boldsymbol{Z}\text{= number of protons = number of electrons.}$$

The number of neutrons N can be found using the relation A=Z+N, therefore $N=A-Z$.

Example: $^{12}_{6}C$ is carbon, which has 6 electrons, 6 protons, and 12−6=6 neutrons.

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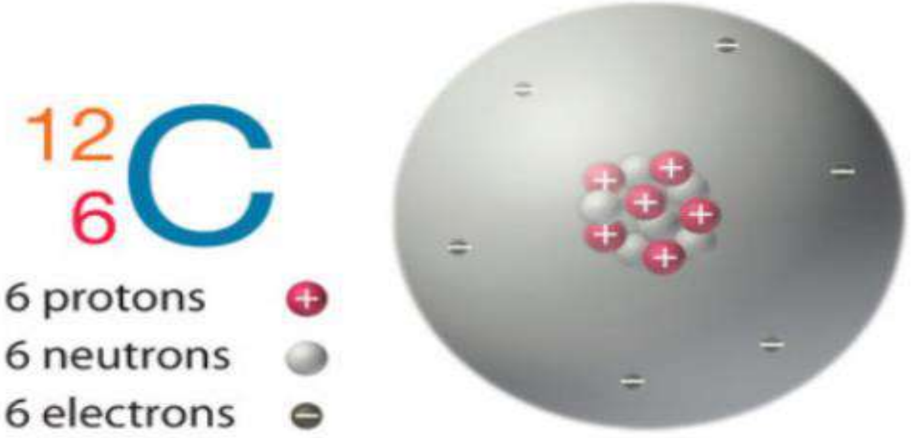
$$\text{mass number } A = (\text{number of protons}) + (\text{number of neutrons})$$

$$\text{mass number} = 6 + 6 = 12$$



The atomic mass unit (amu) is the practical unit used to measure the mass of an atom.

The reference is the carbon-12 atom ${}^{12}_{6}C$, which has a mass of 12 u.

Thus, it is estimated that $\mathbf{1u \sim m_{proton} \sim m_{neutron}}$.

This allows us to define Avogadro's number, which is the number of ${}^{12}_{6}C$ C atoms contained in 12 grams of ${}^{12}_{6}C$: $\mathbf{N_A = 6,022 \, . \, 10^{23} mol^{-1}}$,

.One mole consists of $\mathbf{6,022 \, . \, 10^{23}}$6 entities (molecules, atoms, etc.).

The molar mass *M* of an entity is the mass of one mole of that entity.

For one mole of atoms ${}^{A}_{Z}X$ alors M ~ A g. mol^{-1}.

For example : M (${}^{12}_{6}$C) = 12 g. mol^{-1} ; M (${}^{14}_{7}N$) = 14 g. mol^{-1}

The atomic mass unit:

This unit does not belong to the International System of Units (SI), and its value is determined experimentally. It is defined as **1/12 of the mass of an atom of the nuclide carbon-12**, unbound, at rest, and in its ground state. In other words, a 12C atom has a mass of **12 u**, and if we take **N** (Avogadro's number) atoms of 12C, their total mass is **12 g**.

Consequently,

An **electronvolt** is equal to the energy gained by an electron accelerated through a potential difference of one volt.

Concretely:

$$1 \text{ eV} = 1{,}602\ 177 \cdot 10^{-19} \text{ J}.$$

$$1uma = 931.10^6 eV = 931 MeV$$

Conversion table

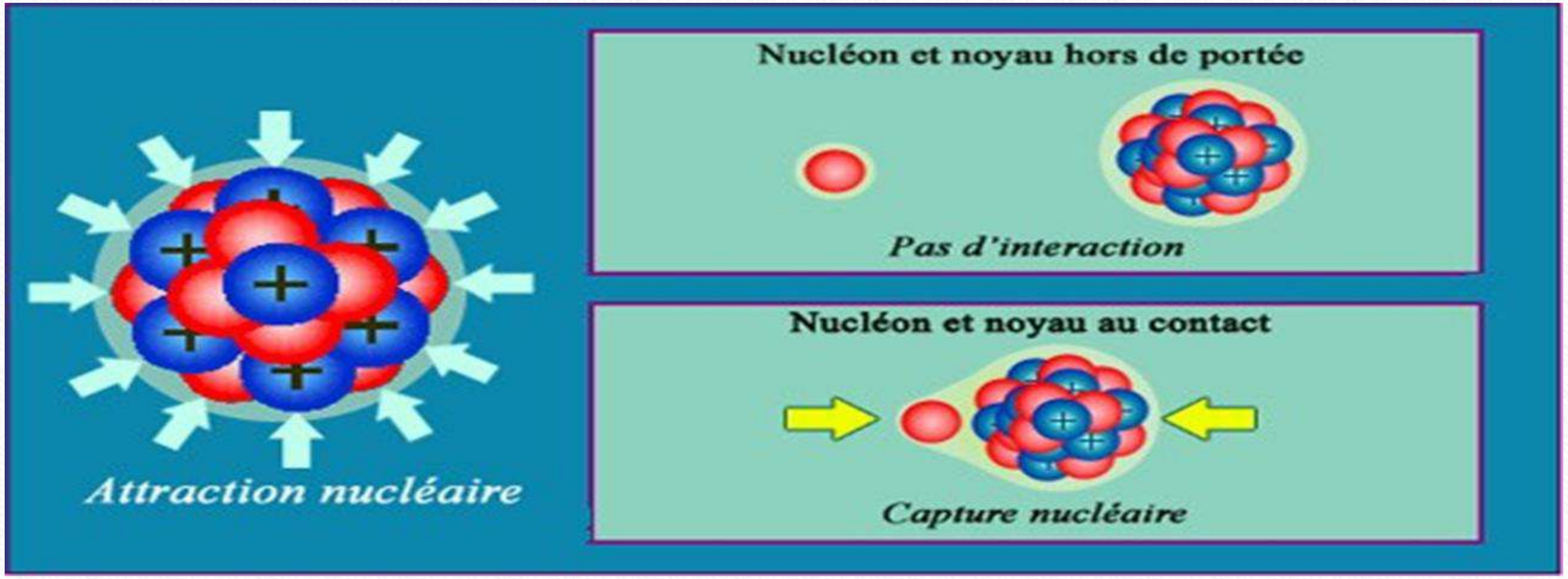
Particule	Kg (10^{-27})	u.m.a	MeV
Proton	1,67	1,007	938
neutron	1,68	1,008	939

Three types of forces compete within the nucleus.

The strong interaction force ensures the cohesion of nuclei because it is attractive. It is also responsible for α-radioactivity. At a distance from a nucleus, a proton or neutron does not feel this force. The attraction becomes significant as soon as the nucleon (here, a neutron) comes into contact with the nucleus. It is then captured. This nuclear glue has been depicted in the figure as enveloping both nucleons and nuclei.

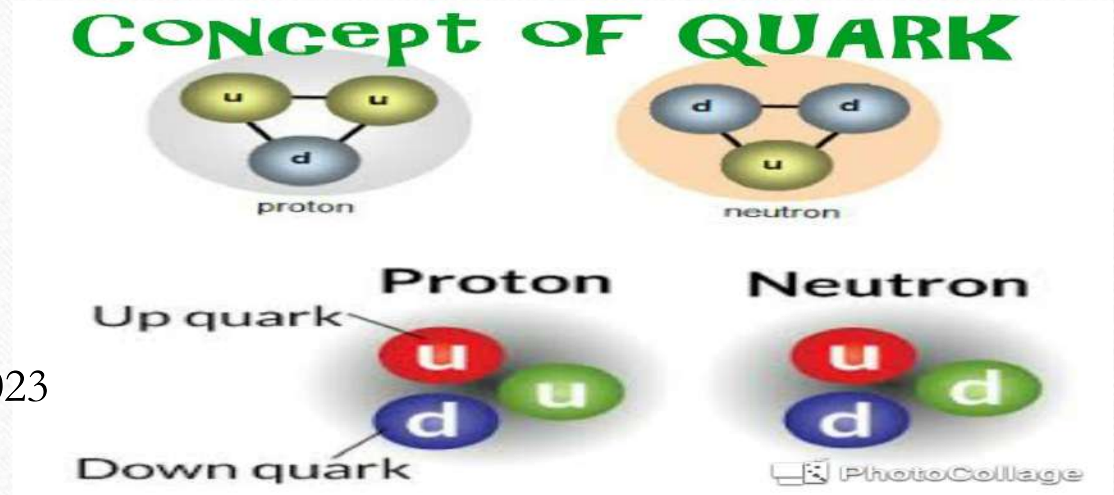
Question: when is a strong force not a strong force? Answer: when it's anywhere outside the atomic nucleus.

The strong force holds together quarks , the fundamental particles that make up the protons and neutrons of the atomic nucleus, and further holds together protons and neutrons to form atomic nuclei. As such it is responsible for the underlying stability of matter. Its huge power is also what is released in the process of nuclear fusion in the sun, or nuclear fission in a nuclear bomb.


Attraction nucléaire
Nucléon et noyau hors de portée
Pas d'interaction
Nucléon et noyau au contact
Capture nucléaire

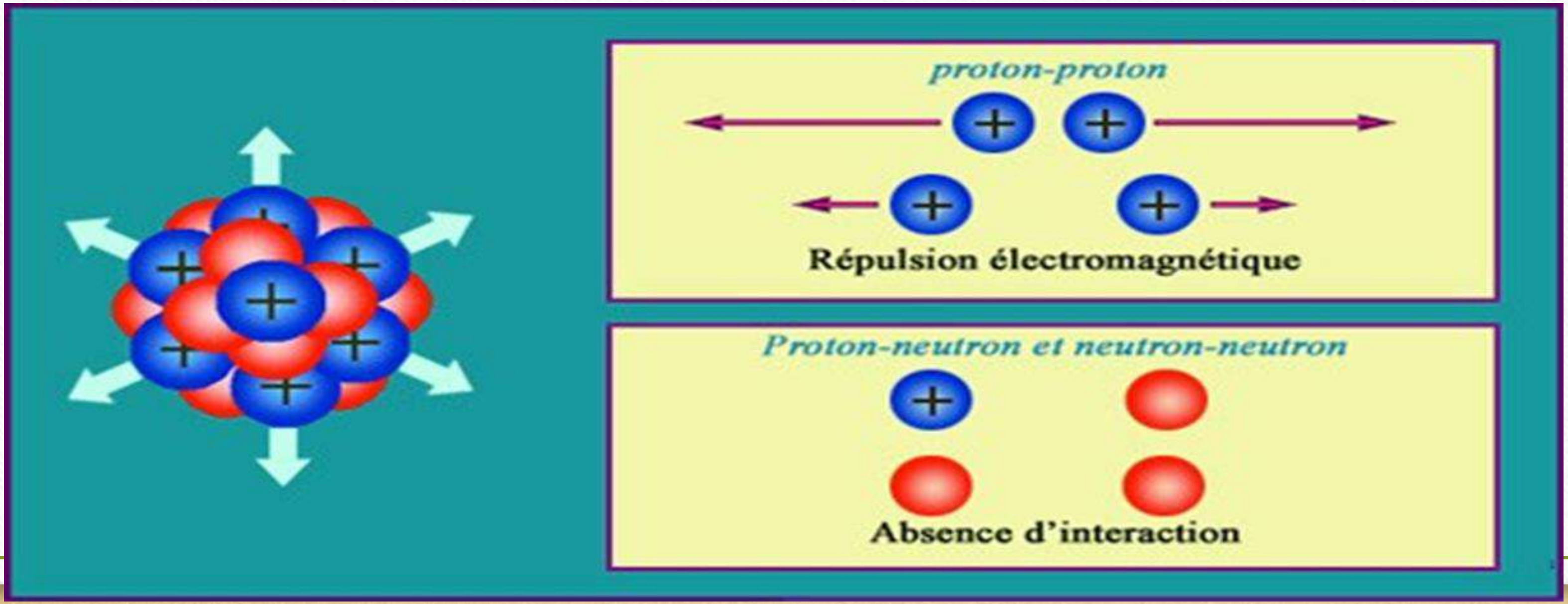
A quark is a fundamental particle and one of the basic building blocks of matter in the Standard Model of particle physics. Quarks combine to form composite particles called **hadrons**, the most familiar of which are protons, neutrons and Pions (π-mesons) , which make up atomic nuclei.

Pions (π-mesons) are mesons composed of a quark and an antiquark. They play an important role in mediating the strong nuclear force between nucleons (protons and neutrons) in the nucleus.



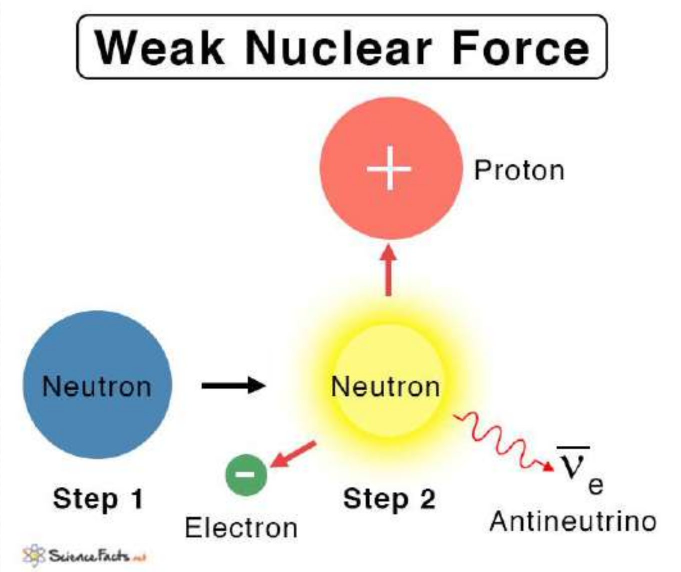
Quarks: a quick introduction. YouTube·Dave Darling. nov.2023

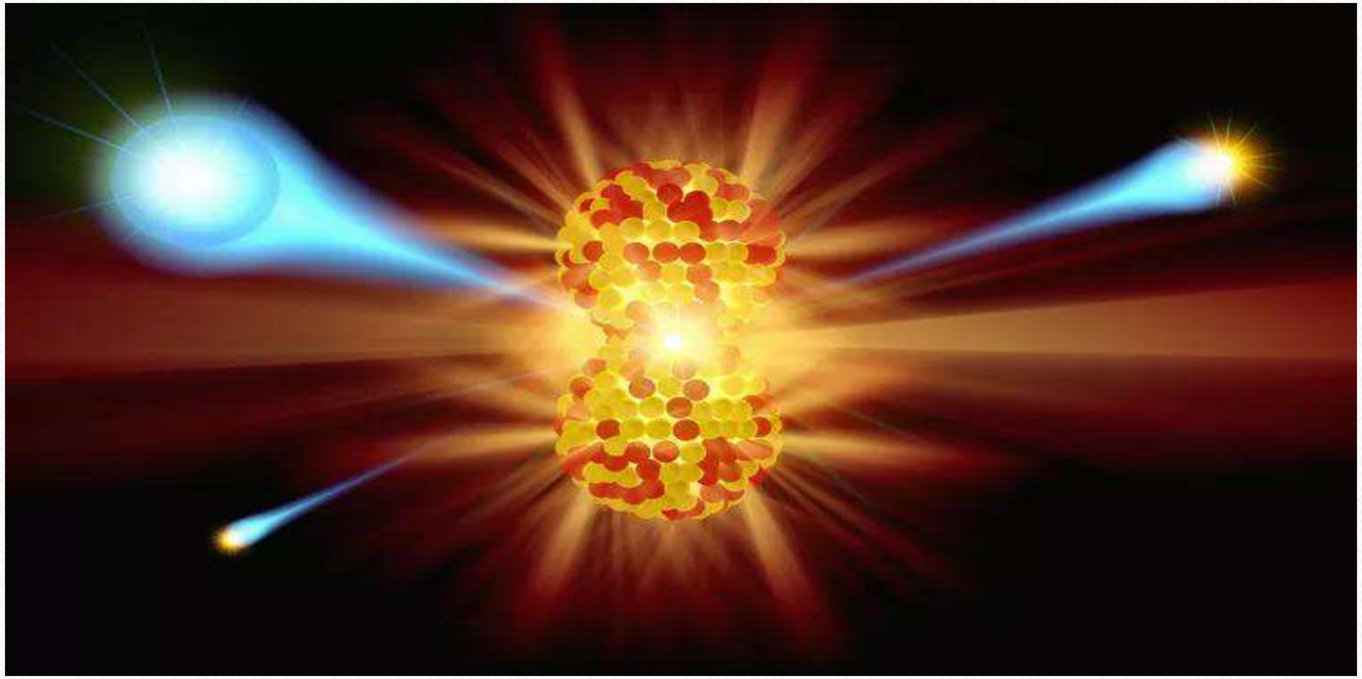
The electromagnetic interaction force, or **the electromagnetic repulsion of protons**, which is less intense.
Two electric charges of the same sign repel each other. This repulsion varies inversely with distance according to Coulomb's law. Protons (in blue) are subject to this repulsion within the nucleus, unlike neutrons (in red), which have no electric charge and therefore do not experience Coulomb repulsion. Without the nuclear forces counteracting the repulsion between protons, the nucleus would explode. These forces must be extremely strong to confine protons within such a small volume as the nucleus.



The weak interaction force is neither attractive nor repulsive; it acts within the nucleons themselves. It transforms one type of nucleon (proton or neutron) into the other and vice versa, causing β-radioactivity. The stability or instability of a nucleus is the result of the competition between these three interactions.

Without the weak force, the Sun would stop shining because it would not be able to fuse hydrogen into deuterium.





Nuclear stability:

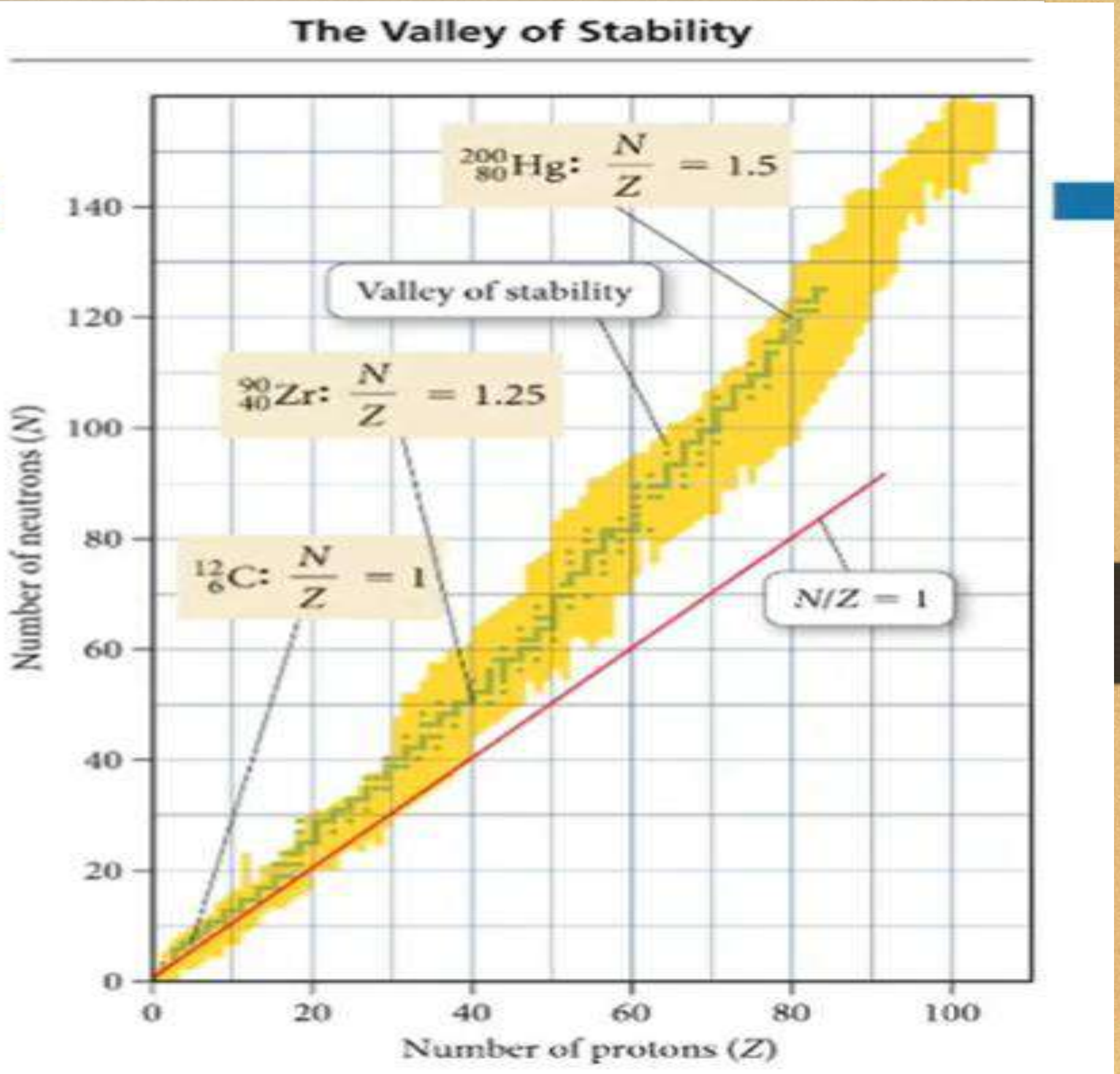
Light atomic nuclei follow the bisector rule (N = Z).The cloud of atoms shifts toward neutron predominance (N > Z) but remains stable.

N/Z too high

- Nuclides above the valley of stability have too many neutrons.
- They tend to convert neutrons to protons via beta decay.

N/Z too low

- Nuclides below the valley of stability have too many protons.
- They tend to convert protons to neutrons via positron emission or electron capture
- Alpha decay also raises N/Z but effect is smaller.



Mass defect and binding energy of a nucleus.The mass of a nucleus is slightly less than the sum of the masses of the protons and neutrons that compose it. The difference between these two masses, called the mass defect and denoted by Δm, is calculated using the following relation:

$$\Delta m = \sum m_{nucleons} - m_{nucleus} = [Z.m_p + (A - Z).m_n] - m_{nucleus}$$

where Z and A are the atomic number and mass number of the nucleus, and $m_{nucleus}$ is the mass of the nucleus.

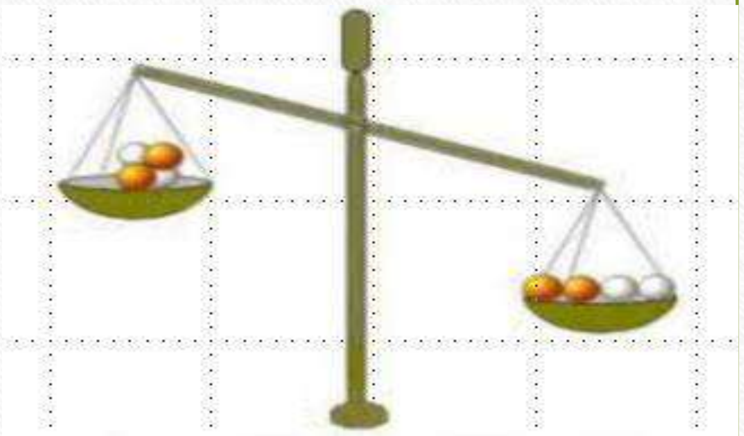
The mass defect corresponds to the binding energy W of all nucleons in the nucleus:$W = \Delta m.c^2$ (c:speed of light)

Example: For the helium nucleus ${}_2^4He$: M=4,003uma

Wheras $2m_p + 2m_n = 4,03uma$

We deduce that : $\Delta m = 0,027uma = 25,13MeV$

The mass of a helium nucleus is smaller than the sum of the masses of its individual protons and neutrons.

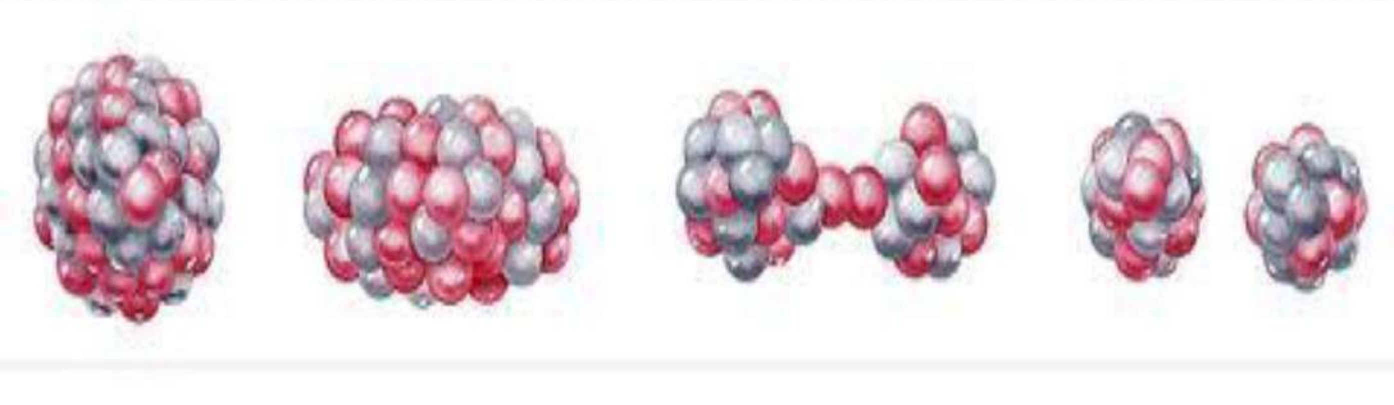


Magic numbers: For light elements, the number of neutrons is approximately equal to the number of protons, while for heavy elements, stability is achieved through a neutron excess.

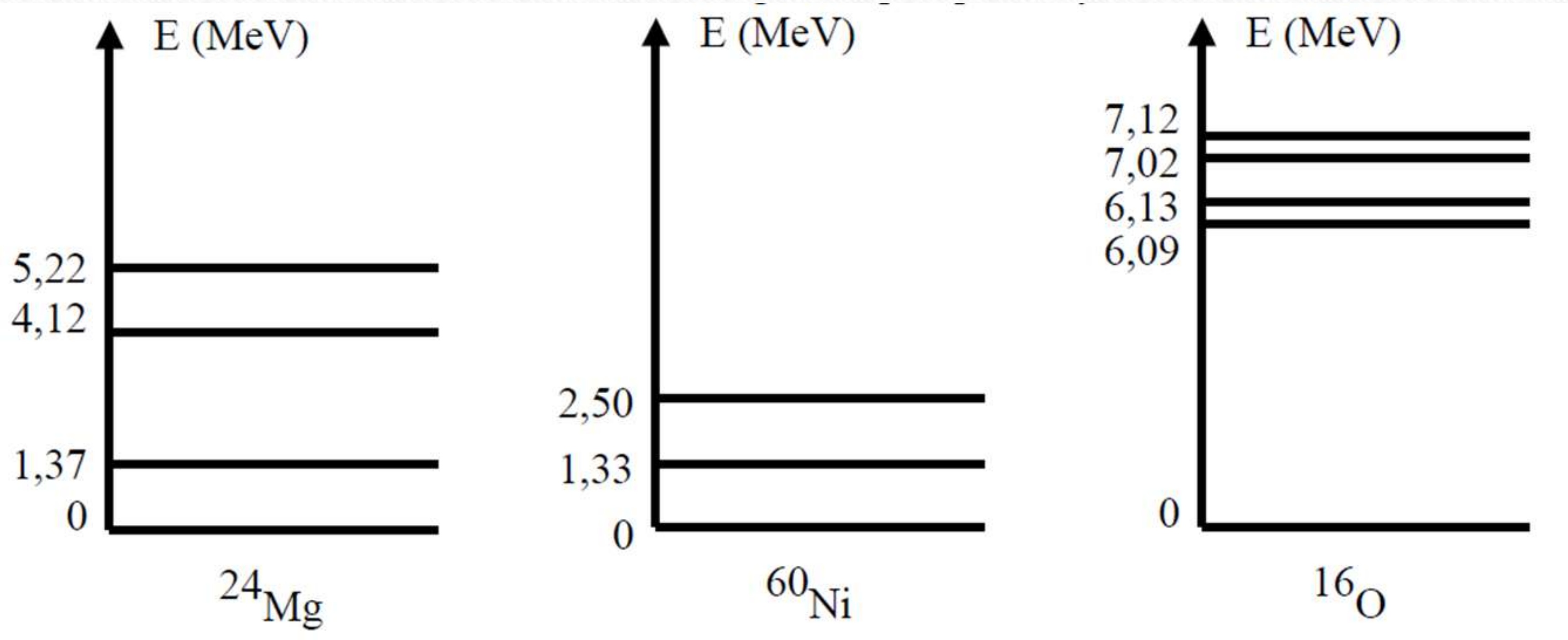
The most stable nuclides are those that have a number of protons or neutrons equal to 2, 8, 20, 28, 50, 82, or 126; these are called magic numbers.

Nuclear models: There are two main models of the nucleus:

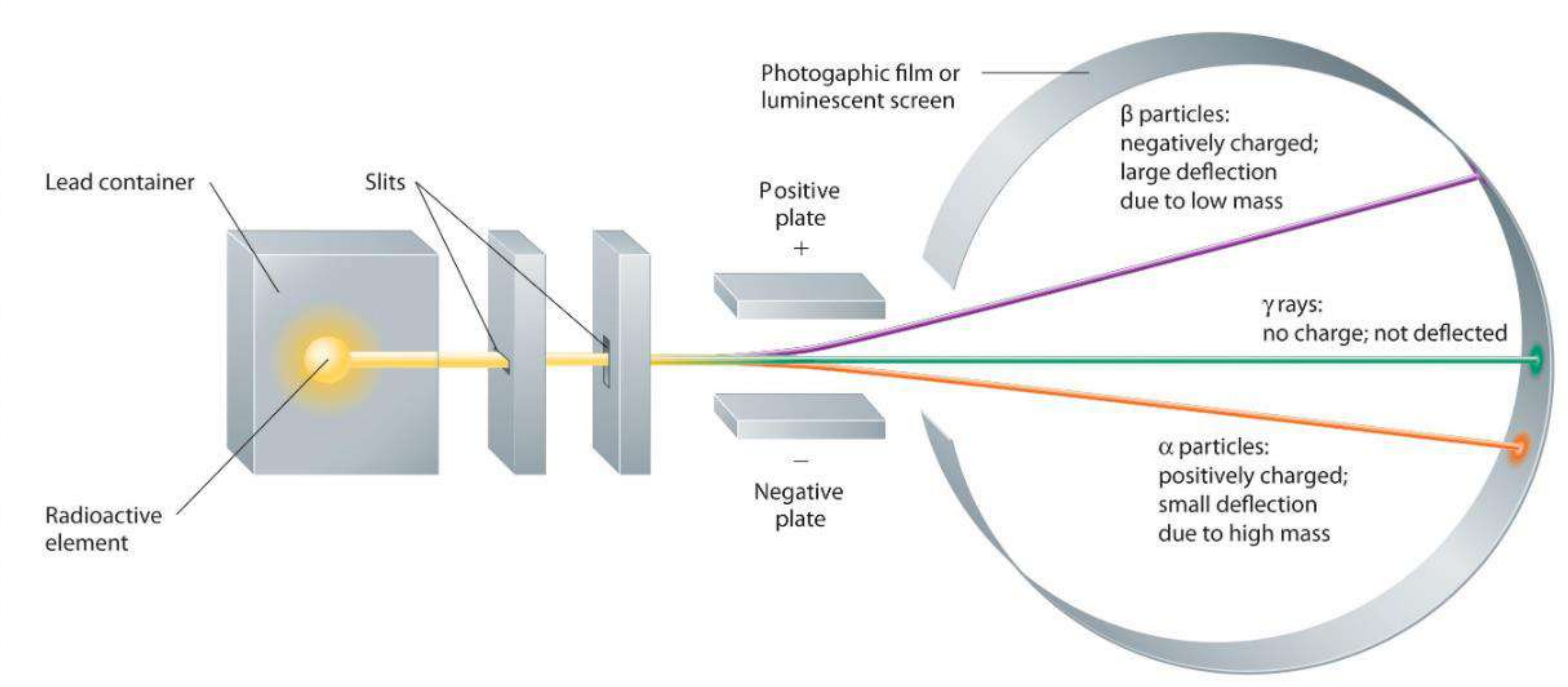
The liquid drop model: In this model, nucleons retain their individual properties and interact with each other via the strong force. This model explains nuclear fission and particle emissions from the nucleus.



The shell model: Nucleons are arranged in energy levels within the nucleus, and this model can explain the emission of gamma rays (γ-rays) from the nucleus.



First energy levels of some nuclei


Photogaphic film or
luminescent screen
β particles:
negatively charged;
large deflection
due to low mass
Lead container
Slits
Positive
plate
+
γ rays:
no charge; not deflected
α particles:
positively charged;
small deflection
due to high mass
−
Negative
plate
Radioactive
element

Isotopes:

Isotopes are atoms that have the same number of protons but a different number of neutrons.

Example: the isotopes of hydrogen ${}^{1}_{1}H$(hydrogen)${}^{2}_{1}H$(deuterium)${}^{3}_{1}H$(tritium)

the isotopes of oxygen: ${}^{15}_{8}O$; ${}^{16}_{8}O$; ${}^{17}_{8}O$

the isotopes of carbon: ${}^{12}_{6}C$; ${}^{13}_{6}C$; ${}^{14}_{6}C$.

Most elements occur naturally as a mixture of isotopes, each with a different natural abundance (%). This leads to the definition of:The average atomic mass:

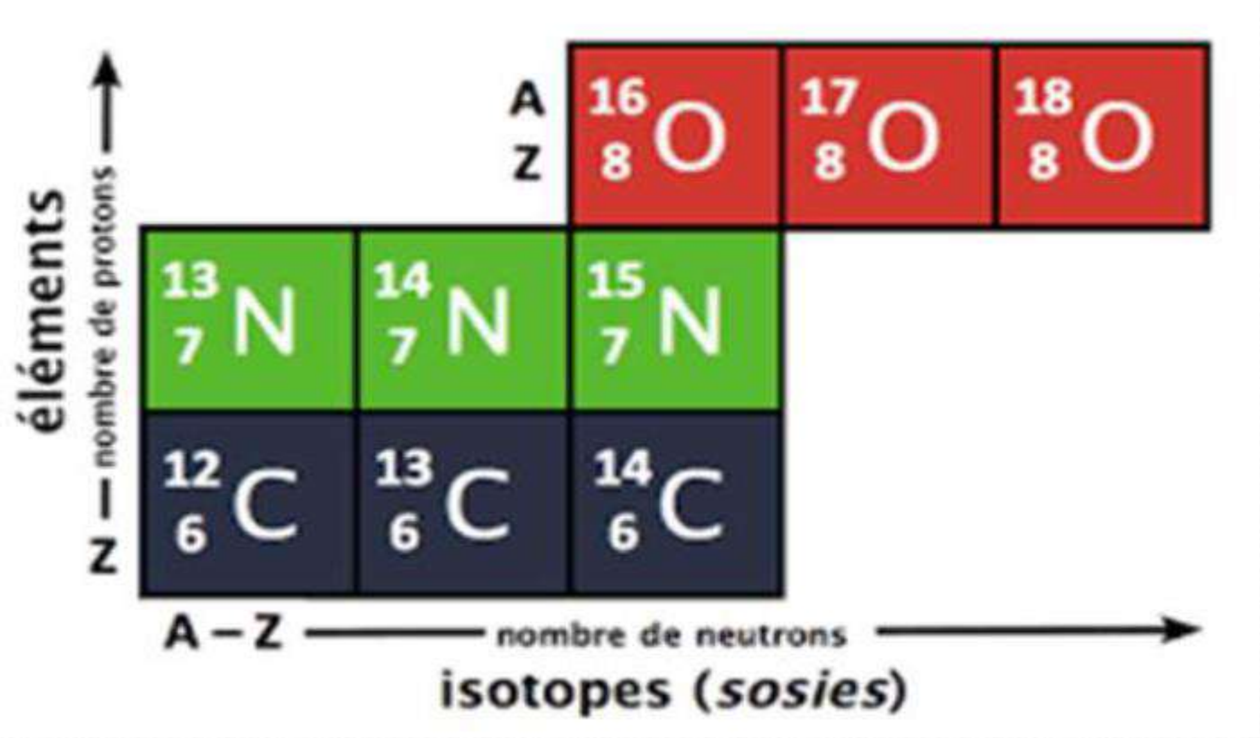
$\bar{M} = \frac{\sum_{i=1}^{n} x_i.m_i}{100}$ where: x_i : percentage abundance of the isotope; m_i : mass of the isotope;

$$\sum_{i=1}^{n} x_i = 100$$

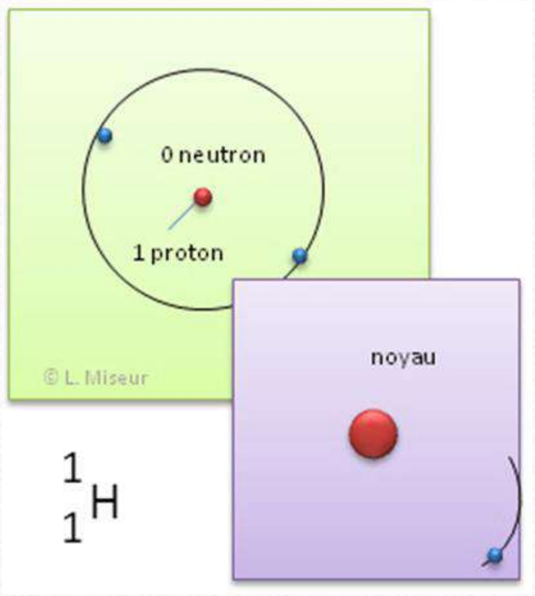
"For example, there are two main naturally occurring isotopes of bromine, ${}^{79}Br$ and ${}^{81}Br$, each with a relative abundance of 50%.

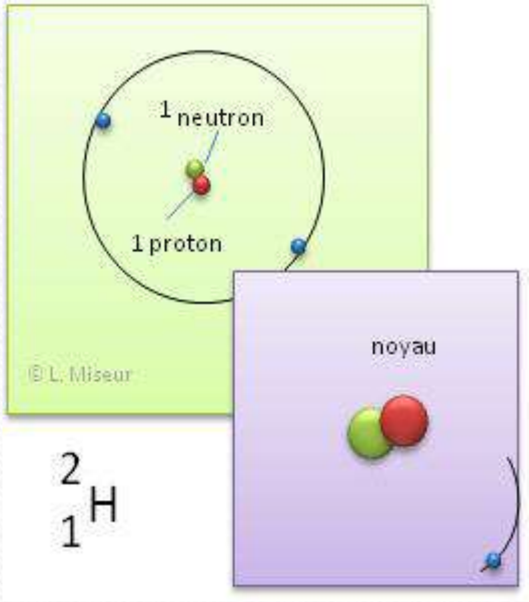
The molar mass is therefore: M=0.5×79+0.5×81=80 $g.mol^{-1}$.

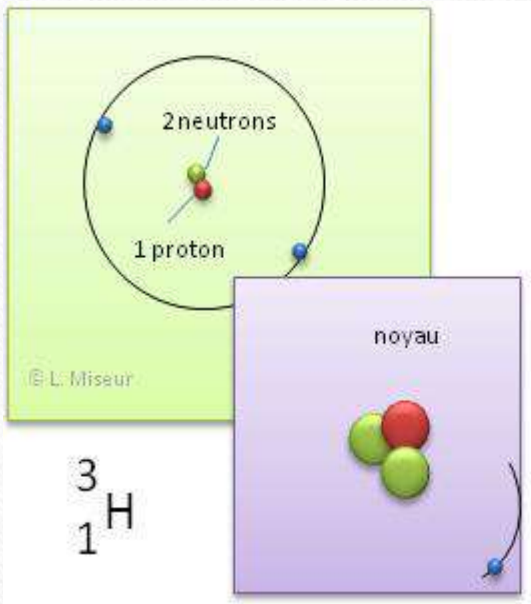
"Unstable isotopes undergo radioactive decay at varying rates, leading to the formation of other nuclei and the release of energy. This phenomenon is referred to as natural or artificial radioactivity."



Hydrogen. There exist other atoms that possess the same chemical properties as the element hydrogen. They have the same number of protons and electrons, but differ in their number of neutrons. These atoms have the same chemical properties, since such properties are determined by the number of electrons, which remains identical. However, they do not have the same physical properties, as their masses are different."







Isobars:

These are nuclides that have the same mass number A but different atomic numbers Z.

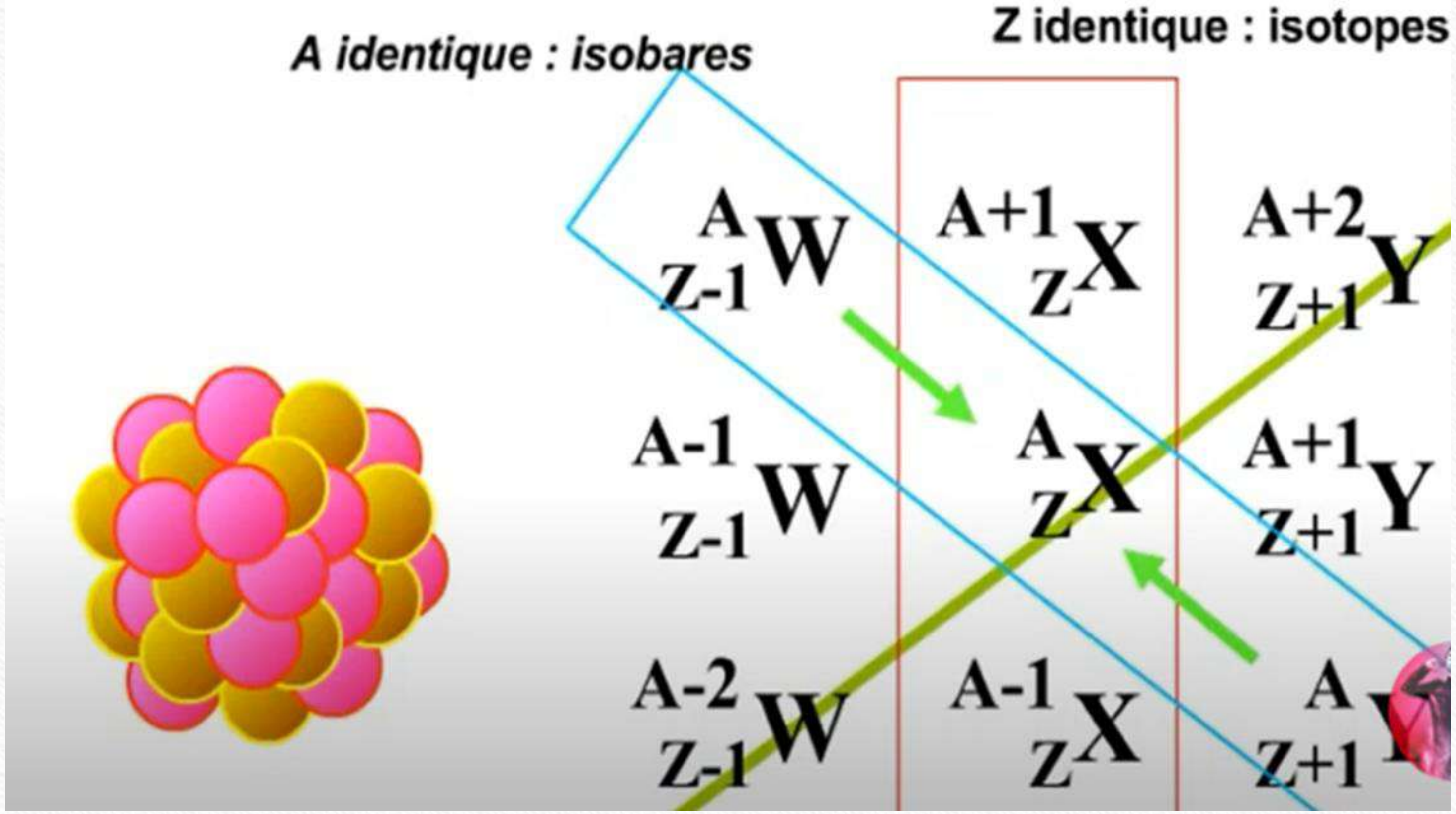
Example: ${}^{15}_{8}O\ and\ {}^{15}_{7}N$

Isotones: These are nuclides that have the same number of neutrons.

Example: ${}^{16}_{8}O$, ${}^{15}_{7}N$ and ${}^{17}_{9}F$.

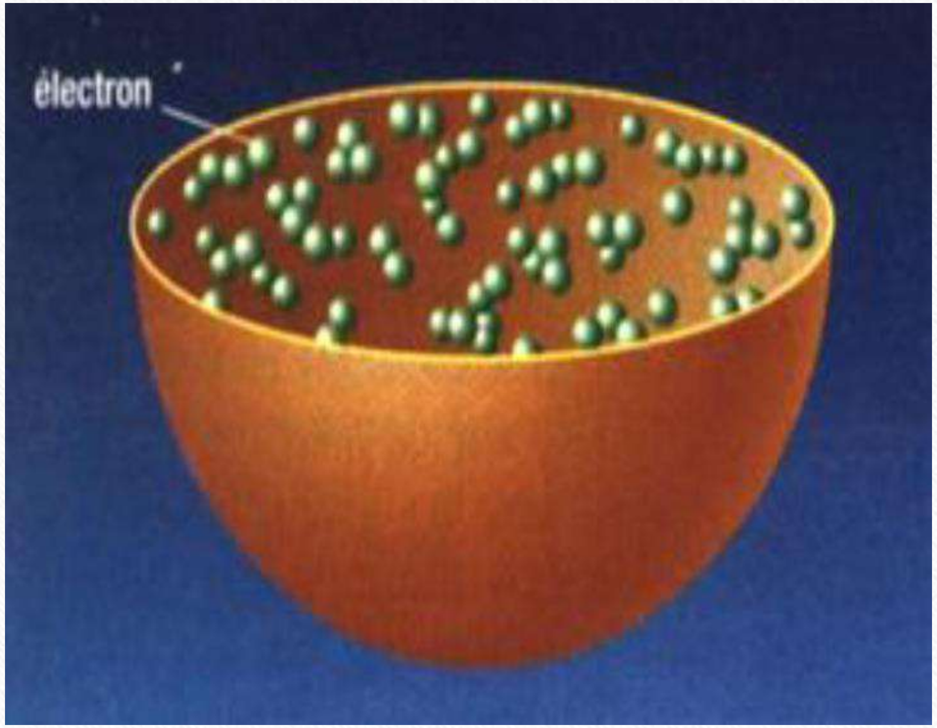
Isomers: These are atoms that have the same number of protons (same Z), the same number of neutrons (same N), and the same mass number (same A), but differ in their energy states.

Example: Technetium-99m, denoted ${}^{99m}Tc$, is a nuclear isomer of the technetium isotope ${}^{99}Tc$, with a mass number equal to 99. It is widely used in nuclear medicine for many diagnostic procedures.


A identique : isobares
Z identique : isotopes
$^{A}_{Z-1}W$
$^{A+1}_{Z}X$
$^{A+2}_{Z+1}Y$
$^{A-1}_{Z-1}W$
$^{A}_{Z}X$
$^{A+1}_{Z+1}Y$
$^{A-2}_{Z-1}W$
$^{A-1}_{Z}X$
$^{A}_{Z+1}Y$

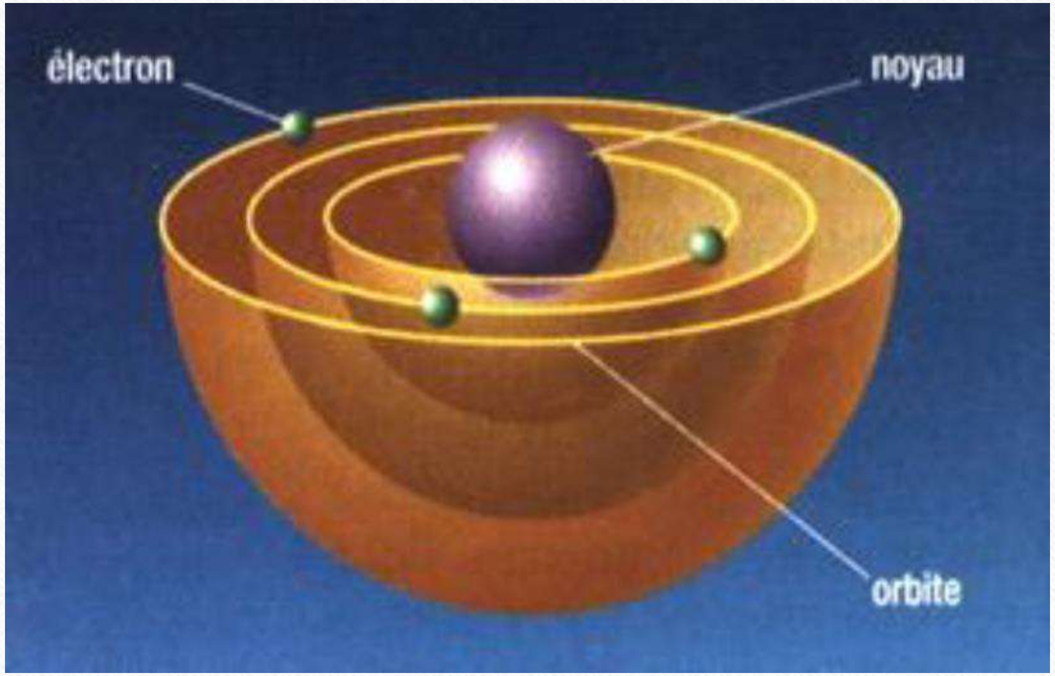
Atomic models:

In 1897, J. J. Thomson discovered the electron and gave it its name. He then proposed a model of the atom in which negative electrons are embedded within a positively charged matter, resembling a plum pudding, so as to ensure the electrical neutrality of the atom.



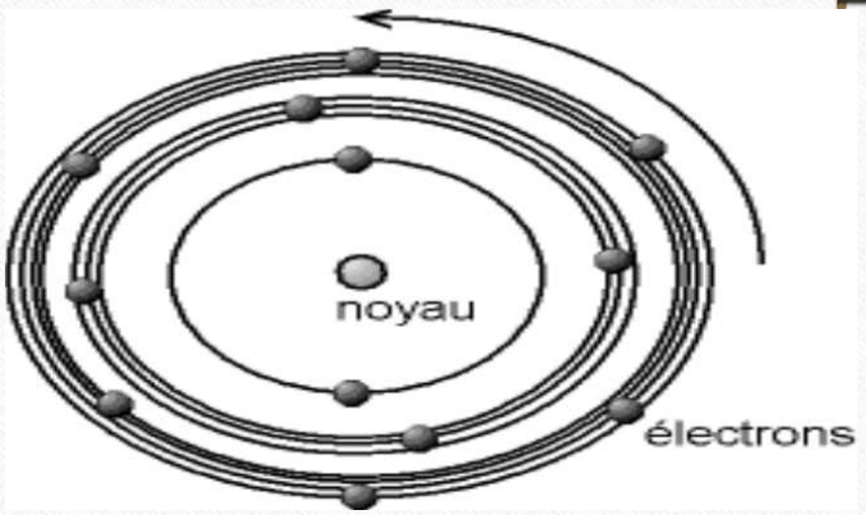
Rutherford's atomic model (1911)

Based on the gold foil experiment, Ernest Rutherford demonstrated that the atom is mostly empty space and contains a small, dense, positively charged nucleus at its center, where most of the mass is concentrated. Electrons move around this nucleus. However, this classical model could not explain the stability of atoms or their discrete emission spectra.



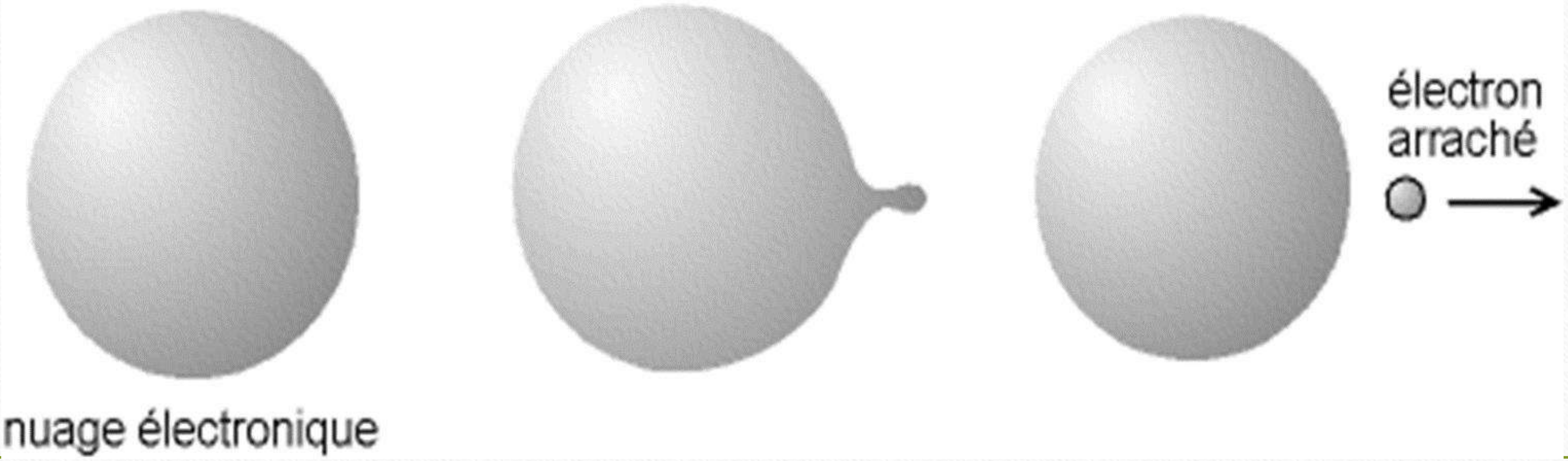
Bohr's atomic model (1913):

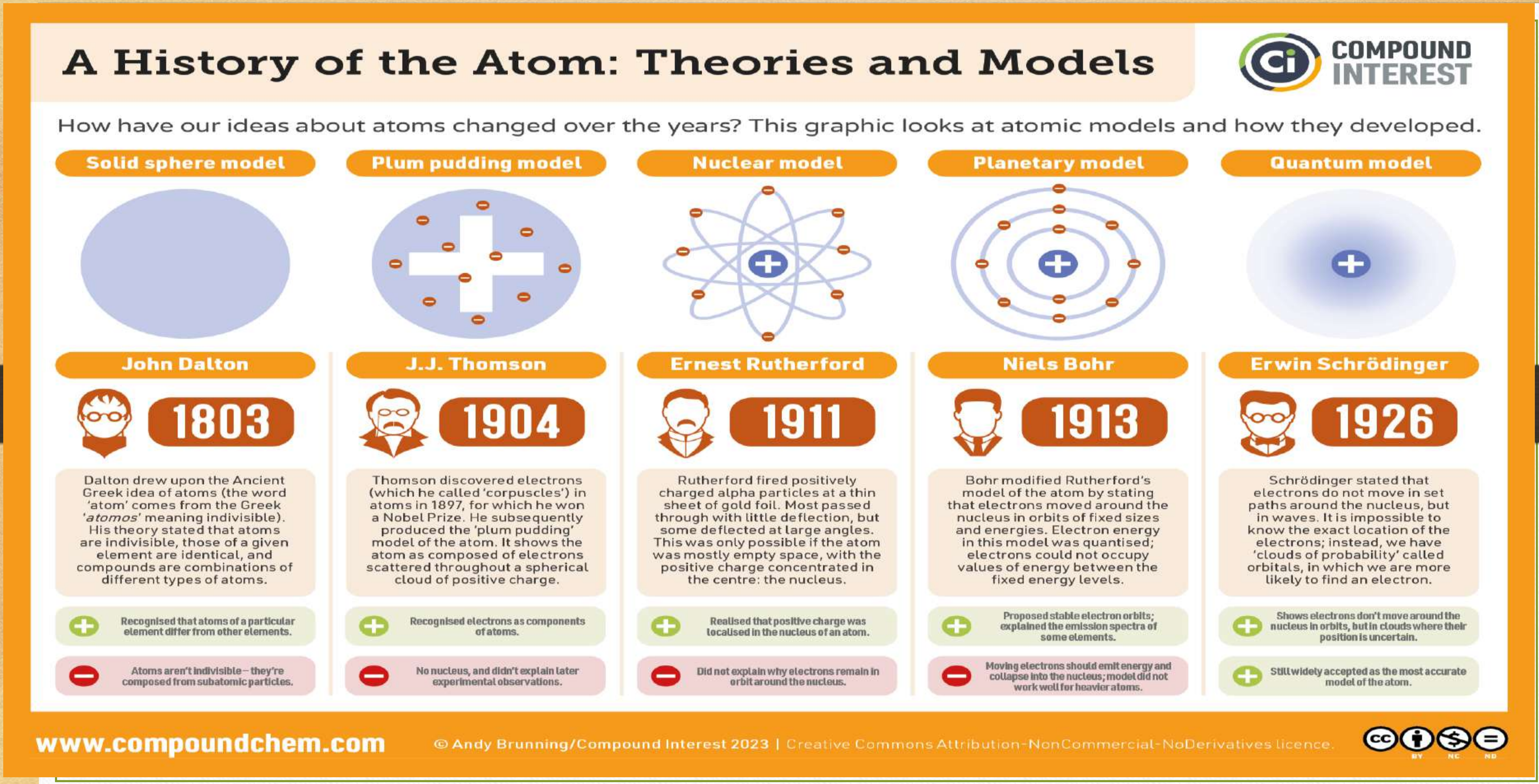
Niels Bohr improved Rutherford's model by introducing the concept of quantized energy levels. Electrons can only occupy specific, well-defined orbits around the nucleus without radiating energy. A transition of an electron between two energy levels is accompanied by the emission or absorption of a photon whose energy corresponds to the difference between these levels. This model successfully explains the emission spectrum of the hydrogen atom but fails for multi-electron atoms. These works were taken up by Einstein, who introduced the quantum model of the atom.



Quantum mechanical model (1926–)

The quantum mechanical model, developed notably by Schrödinger and Heisenberg, describes the electron not as a particle following a precise trajectory, but as a wave function whose squared modulus gives the probability of finding the electron in a given region of space. Electrons occupy atomic orbitals, characterized by quantum numbers, rather than fixed orbits. This model provides an accurate description of atomic structure and underlies modern atomic and molecular physics.




A History of the Atom: Theories and Models
COMPOUND INTEREST
How have our ideas about atoms changed over the years? This graphic looks at atomic models and how they developed.
Solid sphere model
Plum pudding model
Nuclear model
Planetary model
Quantum model
John Dalton
1803
Dalton drew upon the Ancient Greek idea of atoms (the word 'atom' comes from the Greek 'atomos' meaning indivisible). His theory stated that atoms are indivisible, those of a given element are identical, and compounds are combinations of different types of atoms.
Recognised that atoms of a particular element differ from other elements.
Atoms aren't indivisible – they're composed from subatomic particles.
J.J. Thomson
1904
Thomson discovered electrons (which he called 'corpuscles') in atoms in 1897, for which he won a Nobel Prize. He subsequently produced the 'plum pudding' model of the atom. It shows the atom as composed of electrons scattered throughout a spherical cloud of positive charge.
Recognised electrons as components of atoms.
No nucleus, and didn't explain later experimental observations.
Ernest Rutherford
1911
Rutherford fired positively charged alpha particles at a thin sheet of gold foil. Most passed through with little deflection, but some deflected at large angles. This was only possible if the atom was mostly empty space, with the positive charge concentrated in the centre: the nucleus.
Realised that positive charge was localised in the nucleus of an atom.
Did not explain why electrons remain in orbit around the nucleus.
Niels Bohr
1913
Bohr modified Rutherford's model of the atom by stating that electrons moved around the nucleus in orbits of fixed sizes and energies. Electron energy in this model was quantised; electrons could not occupy values of energy between the fixed energy levels.
Proposed stable electron orbits; explained the emission spectra of some elements.
Moving electrons should emit energy and collapse into the nucleus; model did not work well for heavier atoms.
Erwin Schrödinger
1926
Schrödinger stated that electrons do not move in set paths around the nucleus, but in waves. It is impossible to know the exact location of the electrons; instead, we have 'clouds of probability' called orbitals, in which we are more likely to find an electron.
Shows electrons don't move around the nucleus in orbits, but in clouds where their position is uncertain.
Still widely accepted as the most accurate model of the atom.
www.compoundchem.com
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Conclusion:

The knowledge gained from the study of the atom has led to the development of numerous scientific fields, with the medical sciences being among the most significantly impacted, particularly biochemistry, nuclear medicine, and radiology.

Thank you.
I look forward to your comments or questions.