

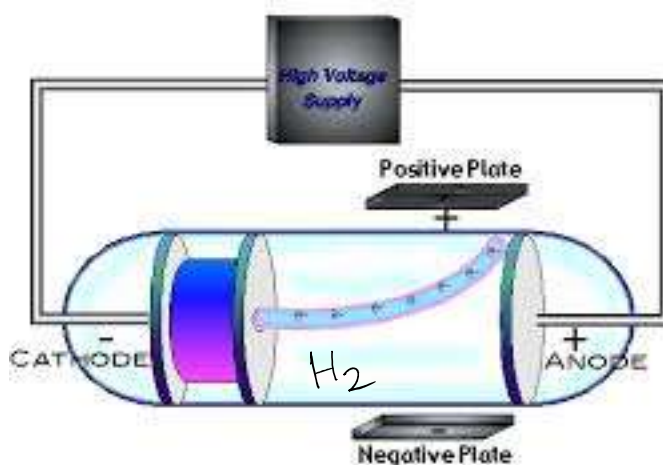
Historical Background:

To track the development of our current understanding of the atomic structure, we need to go back to the ancient Greek. At the time there was, among other debates, a debate about the nature of matter. The question was; if one took a piece of matter, say a steel bar, and cut it into half and take one of the halves and cut it into two and so on, could one continue doing that indefinitely? There were two schools of thinking regarding this problem. One was led by the ancient Greek philosopher Democritus who thought that one will reach a point where the piece is indivisible which then he called ATOMUS (means indivisible in Greek). On the other camp, Aristotle thought that there is no reason to believe that we could not go with the cutting process indefinitely because one will always get half what has been had before. For about 2000 years Aristotle's view that there is no fundamental building block of the matter called atom held the sway.

Fast forward to the 1800th, things began to change largely due to the work of famous chemists. It was already known that there is something called elements such as iron, copper, silver, ... etc. and it was also known for long time that one could produce iron only if the ore contains iron in the first place. So, there seems to be something special of these elements and one could not produce one from the other. It was John Dalton who in 1803 began to piece the atomic model together:

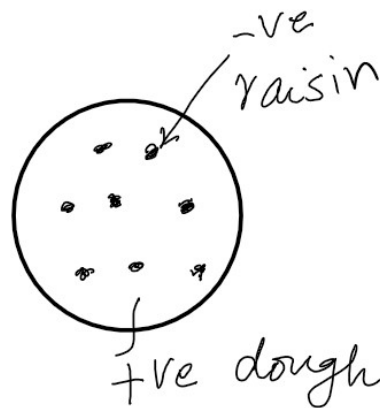
- Matter composed of atoms (He did not know what the atom is).
- All atoms of a given element are identical.
- Atoms of different elements have different weights.
- Atoms combine in simple ratios to form compounds.

So, the idea was born that at the heart of every element there is an indivisible thing called atom. It is rather like a tiny ball, but no body knows what the atom is. J. J. Thomson experiment in 1897 shows that, that was not entirely right. When J. J. Thomson switched on the electricity in his test, the current passes through the tube in the figure below, he noticed that something in the glowing hydrogen started to veer off towards the positive plate



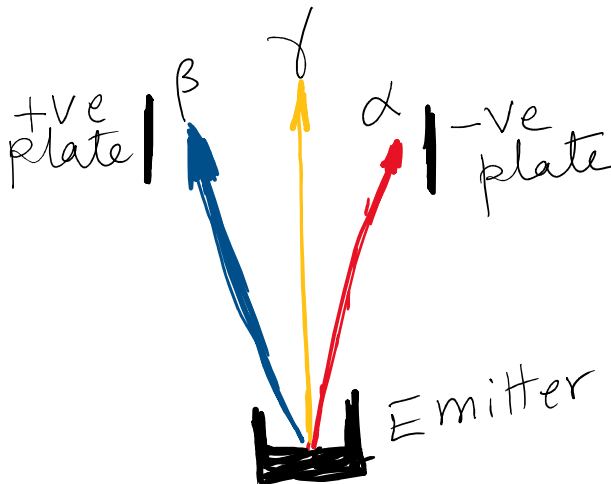
J.J. Thomson's experiment

At the time it was well known that like charges repel and unlike charges attract. So, if we got something attracted to the positively charged plate, it must be of negative charge. But he also knew that the hydrogen inside the tube is neutral. So, what has he made to split a negative piece of the hydrogen atom? and if there is a negative part, then surely there is a positive piece for the two pieces to join to form neutral hydrogen. when he looked closely, he noticed that, while the negatively charged bit veers off significantly toward the positively charged plate, the positively charged piece barely moved toward the negatively charged plate. He concluded that the positively charged bit must be heavier than the negatively charged bit. After J. J. Thomson's experiment, it was concluded that the atom which has been thought of as indivisible was in fact made up of two parts; a negatively charged bit and a much heavier positively charged bit. The best model they could come up with is the plum pudding model (raisin bun model)

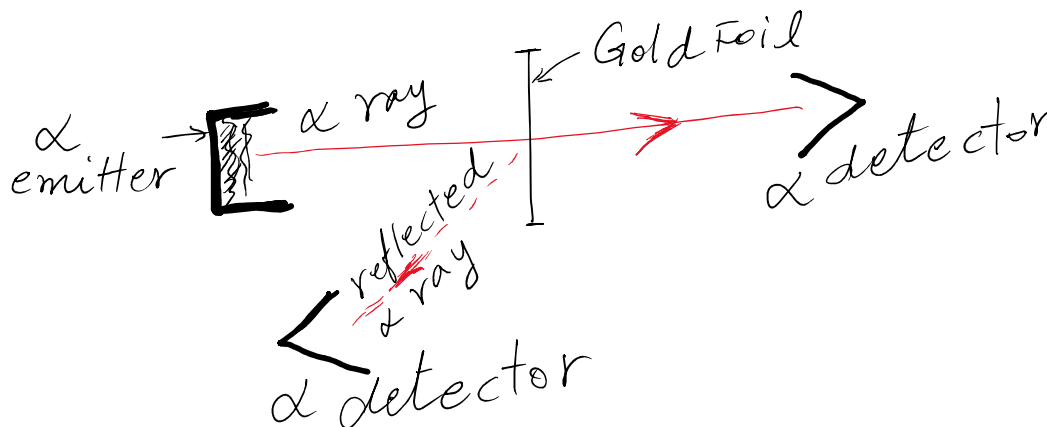


Plum pudding Model of the atom

The story moves now to the discovery of radioactivity by Becquerel. Radioactive materials were seen at the time as materials emitting three types of rays given the names α , β , and γ . If we put them in electric field, α will be attracted to the negatively charged plate which means it is positively charged, β will be attracted to the positively charged plate which means it is negatively charged, and γ will not be attracted at all which means it is not charged.



In 1911 the physicist E. Rutherford performed his famous experiment. He took some radium bromide (known to emit α particles) and put it in a container such that α rays could be emitted in a particular direction. He put a detector that enabled him to detect how many α particles have been emitted. Then he put a very thin gold foil in the path of α particles and measure the number of these particles. To his surprise, there was no difference with or without the gold foil in the path of the emitted α particles. Two of his students, Geiger and Marsden played around the experiment. They moved the detector so that it was facing backward of the path of the α particles. They found a very small number of α particles rebounding toward the backward detector only when the gold



foil was there. No reflection was detected when the gold foil was removed. Rutherford realized that the old plum pudding model is no longer viable. It was concluded that the model of the atom needs to be revised such that the atom has a very small nucleus in the middle and the negative electrons orbit around it like planets orbiting around the sun. In fact, Rutherford was able to work out what the size of the nucleus was compared to that of the atom by knowing the proportion of the rebound α particles and he concluded that the diameter of the nucleus is $1/10^4$ that of the atom. Now we have a model of an atom which consists of a dense, positively charged nucleus at the centre and the negatively charged electrons orbiting around it at a relatively far distance. This means of course that almost all the mass of an atom is concentrated in the nucleus, which means that the density of the nucleus is very big. Indeed, it turns out that a spoon full of nuclear matter would weigh 500 Mega tonnes. It soon be discovered that the positively charged nucleus is made up of protons and a neutral atom has as many protons as electrons and it is the number of protons that determines what the element is (1-H, 6-C, 8-O, 26-Fe, 92-U).

Chadwick, in 1932, discovered that there was another particle that could occupy the nucleus; that was the neutron who does not have any charge. The question now is ; are the proton, the neutron, and the electron the basic unit, that Democratise talking about? such question is beyond the scope of our course.

Rutherford model of the atom gave rise to two important questions; if the negatively charged electrons is orbiting around a positively charged nucleus, They are accelerating due to changing direction (centripetal acceleration $v_i^2/r = \omega r^2$) and emit radiation.

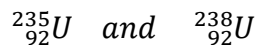
Then electrons are losing energy and they should spiral inward toward the nucleus and the whole atom would annihilate when it hits the nucleus and therefore atom can't exist. Addressing this question gave rise to a branch of physics called "*quantum mechanics*". The second problem with this model is that, other than hydrogen, nucleus have more than one proton in a very tiny space, but we know that like charges repel and all nuclei (apart from hydrogen) should self destruct as they force their protons apart. But that does not happen. Addressing this question gave rise to "*nuclear physics*".

Atomic and Nuclear Nomenclature:

The identity of an atom is specified by the number of protons Z and neutrons N . Z is known as the atomic number. For an electrically neutral atom the number of electrons is equal to the number of protons Z . All atoms of the same element have the same atomic number Thus, e.g. All oxygen atoms have 8 protons in their nucleus. All uranium atoms have 92 protons. Atoms of the same element may have different number of neutrons which are called isotopes. The symbols used to denote an isotope is:



where X is the chemical symbol, $A=N+Z$ is called the mass number. For example, two uranium isotopes which will be discussed extensively later are



Nuclide

A term used to refer to particular atom or nucleus with specific N and Z . Nuclides are either stable, where they do not change unless excited, or radioactive, where they spontaneously change to another nuclide with different Z or N by emitting different kind of radiation.

The electron volt

Energy unit developed to express minute amount of energy liberated or absorbed during chemical reaction of single atom or molecule (changes in the electron bonds in the affected molecule). The electron volt (denoted by eV) is defined as the kinetic energy gained by an electron (mass m_e and charge e) when accelerated through a potential difference of one volt

$$1V = 1 \frac{w}{A} = 1 \frac{j}{s/C} = 1 \frac{j}{C}$$

The Work done on the electron by the electric field of 1 volt is

$$= e\Delta v = (1.602 \times 10^{-19} C) (1 J/C) = 1.602 \times 10^{-19} J$$

Therefore

$$\mathbf{1\text{ev} = 1.602 \times 10^{-19} J}$$

If the electron of rest mass m_e starts from rest, then the kinetic energy (K.E.) of the electron after being affected by a potential difference of 1 volt must equal to 1 ev the work done on the electron

$$K.E. = \frac{1}{2} m_e v_e^2 = 1\text{ev} = 1.602 \times 10^{-19} J$$

Given that the rest mass of the electron $m_e = 9.109 \times 10^{-31} \text{kg}$, the speed of the electron is therefore

$$v_e = \sqrt{\frac{2K.E.}{m_e}} = 5.93 \times 10^5 \text{ m/s}$$

The atomic mass unit

Denoted by amu or u , it is defined to be 1/12 of the mass of the atom of neutral C-12 at ground state. Equivalently, the mass of one mole of C-12 which consists of Avogadro's number of atoms ($6.0221420 \times 10^{23} \text{ mol}^{-1}$) is 0.012 kg, Thus

$$1\text{amu} = \frac{1}{12} \times \frac{0.012}{6.02214 \times 10^{23}} \cong 1.66 \times 10^{-27} \text{ kg}$$

So

$$\mathbf{1\text{amu} = 1.6605387 \times 10^{-27} \text{ kg}}$$

The atomic and molecular weight

The atomic weight $M(^A X)$ of an atom is defined as the ratio of the atom's mass $m(^A X)$ to that of one neutral atom of C-12 in its ground state $m(^{12}C)$; where $m(^{12}C)$ is considered to be exactly 12 *amu*. That is

$$M(^A X) = 12 \times \frac{m(^A X)}{m(^{12}C)}$$

In terms of *amu*

$$M(^A X) = 12 \times \frac{m(^A X)}{12} = m(^A X)$$

most naturally occurring elements are composed of two or more isotopes. The isotopic abundance γ_i of the *i*th isotope in a given element is defined as the relative number of atoms of the different isotopes of one chemical element. It is usually expressed as a percentage of the total of all the long-lived isotopes of a given element, on the average, either in nature or in some reference state. A precise measurement of the isotopic abundance is of great importance in determining the atomic masses of elements. The elemental atomic weight is the weighted average of the atomic weights of all naturally occurring isotopes weighted by the isotopic abundance of each isotope;

$$M = \sum_i \frac{M_i \gamma_i}{100}$$

Avogadro's number

The utility of Avogadro's number is that it facilitates the connection between the microscopic entities in a sample to a macroscopic measure. It is defined as the number of atoms in 12 grams exactly of C-12 ($N_a = 6.022 \times 10^{23}$). The importance of N_a lies in the concept of the mole. A mole is defined to contain as many elementary particles (atoms, molecules, or any entity for that matter) as there are atoms in 12 grams of C-12.

Mass of the atom

With Avogadro's number, one can estimate the mass of individual atom. Since one mole of a group of identical atoms contains N_a atoms, then the mass of an atom will be the mass of one mole divided by Avogadro's number. (see example at the end of this lecture notes)

Atomic and nuclear radii (size)

The electron clouds around the atom do not provide well-defined edges for the atom. However, a plausible measure of the size of the atom can be estimated by the average distance from the nucleus to the outermost electron. This distance is approximately the same for all atoms, about $2 \times 10^{-10} \text{m}$. The nucleus edge is also not very well defined, though less fuzzy than that of the atoms. To first approximation, measurements have shown that

$$R = R_0 A^{\frac{1}{3}}$$

With $R_0 = 1.25 \times 10^{-15} \text{m}$, and A represents the mass number (total number of protons and neutrons). The volume of nucleus can reasonably be approximated by a sphere. As such, the volume of a nucleus V is

$$V = \frac{4}{3} \pi R^3 = \frac{4}{3} \pi R_0^3 A$$

Atomic number density:

In many calculations we will need the number of atoms in unit volume (e.g. in 1 cm^3 of a substance). Avogadro's number helps us to calculate the atomic number density N . For a single species homogeneous matter with a mass density $\rho \text{ g/cm}^3$, one cubic centimetre will contain ρ/M moles of the substance. The number of atoms in one cubic centimeter (the number density N) is given by:

$$N = \frac{\rho N_a}{M}$$

The atom density of isotope i of isotopic abundance γ_i is obviously given by

$$N_i = \frac{\gamma_i N}{100}$$

It is important to note that N_i represents the number of atoms of isotope i in a unit volume of the substance, including all isotopes.

This definition can be extended to compounds composed of identical molecules. In this case N will be the number of molecules in cm^3

when the composition of a mixture is specified by the mass fraction w_i of each constituents i , then the mass density of constituent i is $\rho_i = w_i \rho$ and

$$N_i = \frac{\rho_i N_a}{M_i} = \frac{w_i \rho N_a}{M_i}$$

For a chemical compound X_nY_m , the molecular weight $M = nM_x + mM_y$ and the mass fraction of the component x is given by

$$w_x = \frac{nM_x}{nM_x + mM_y}$$

Values of some important physical constants

Constant	Symbol	Value
Speed of light (in vacuum)	c	$2.997\,924\,58 \times 10^8 \text{ m s}^{-1}$
Electron charge	e	$1.602\,176\,46 \times 10^{-19} \text{ C}$
Atomic mass unit	u	$1.660\,538\,7 \times 10^{-27} \text{ kg}$ ($931.494\,013 \text{ MeV}/c^2$)
Electron rest mass	m_e	$9.109\,381\,9 \times 10^{-31} \text{ kg}$ ($0.510\,998\,90 \text{ MeV}/c^2$) ($5.485\,799\,11 \times 10^{-4} \text{ u}$)
Proton rest mass	m_p	$1.672\,621\,6 \times 10^{-27} \text{ kg}$ ($938.272\,00 \text{ MeV}/c^2$) ($1.007\,276\,466\,9 \text{ u}$)
Neutron rest mass	m_n	$1.674\,927\,2 \times 10^{-27} \text{ kg}$ ($939.565\,33 \text{ MeV}/c^2$) ($1.008\,664\,915\,8 \text{ u}$)
Planck's constant	h	$6.626\,068\,8 \times 10^{-34} \text{ J s}$ $4.135\,667\,3 \times 10^{-15} \text{ eV s}$
Avogadro's constant	N_a	$6.022\,142\,0 \times 10^{23} \text{ mol}^{-1}$
Boltzmann constant	k	$1.380\,650\,3 \times 10^{-23} \text{ J K}^{-1}$

Examples

1. Find the atomic weight of boron given the following table:

Isotope	Abundance γ	Atomic weight M
^{10}B	19.9	10.012937
^{11}B	80.1	11.009306

2. How many atoms of ^{10}B in 5 grams of boron ? Use M from example 1.
3. A uranium-aluminum alloy contains 10 w/o U and 90 w/o Al and has a density of 3.4 g/cm³. Compute the number density of the aluminum atoms present.
4. The mass density of a Carbon sample is 1.60 g/cm³• Calculate its atomic number density. The atomic weight of carbon is 12.01.
5. Calculate the mass density of nuclear matter. Given that the radius of a proton is 1.25×10^{-15} m