

Chapter 2: Structure of the atom

- Objectives

By the end of the second chapter, the student will be able to:

- Understand the historical development of atoms and the discovery of the different theories.
- Distinguish between the different constituents of the atom and know their characteristics.
- Define isotopes and understand their characteristics.

- Introduction

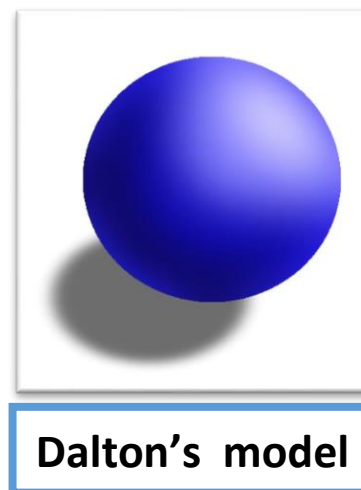
The atom, a particle of extremely small size compared to those we are more familiar with, is the fundamental unit of matter.

The word atom comes from the Greek $\alpha\tau\omicron\mu\omicron\varsigma$ (atomos), meaning “indivisible” — something that cannot be divided further. However, during the late 19th and early 20th centuries, physicists discovered subatomic particles and internal structure within the atom, thereby demonstrating that the so-called “indivisible” atom was, in fact, divisible.

1- Dalton's Atomic Theory 1807

Dalton expressed his theory in a series of postulates.

1. All elements are composed of tiny indivisible particles called atoms.
2. Atoms of the same element are identical. Atoms of any one element are different from those of any other element.
3. Atoms of different elements combine in simple whole-number ratios to form chemical compounds.
4. In chemical reactions, atoms are combined, separated, or rearranged but never changed into atoms of another element.
5. John Dalton in 1807 proposed that matter was composed of small, spherical particles.



Note:

Dalton was wrong about all atoms of the same element being identical. Atoms of the same element can have different numbers of neutrons; these are called isotopes. Frederick Soddy (1877–1956) proposed the idea of isotopes in 1912 and won the Nobel Prize in Chemistry in 1921 for his work.

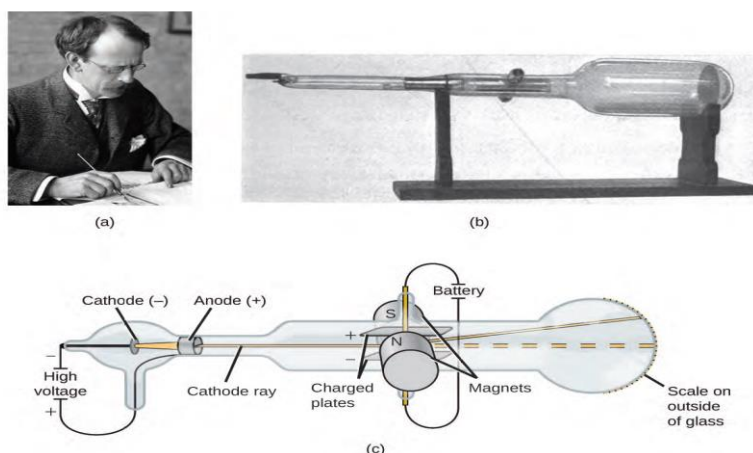
2- Discovery of electron

a- The Joseph John Thomson experiment (J.J. Thomson) 1897:

1897- J.J. Thompson used a cathode ray tube to deduce the presence of a negatively charged particle and he conclude that:

1. Cathode rays have identical properties regardless of the element used to produce them. All elements must contain identically charged electrons.
2. Atoms are neutral, so there must be positive particles in the atom to balance the negative charge of the electrons.
3. Electrons have so little mass that atoms must contain other particles that account for most of the mass.
4. By measuring their deviation, he was able to determine the charge to mass (e/m) ratio of the particles.

$$e/m = 1,759 \times 10^{11} \text{ C/Kg}$$



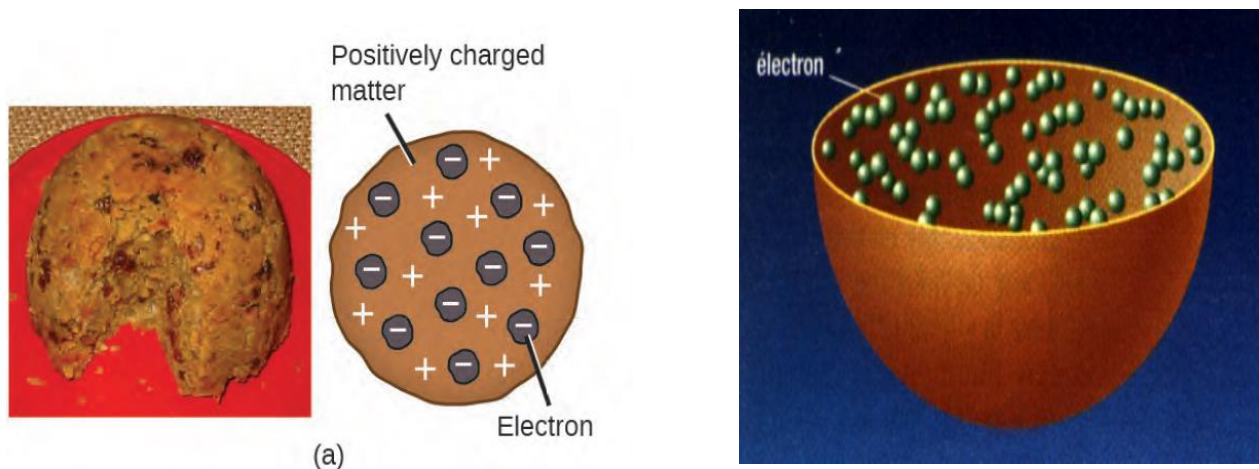
The Joseph John Thomson experiment

Link for demonstrative video

https://www.youtube.com/watch?v=_nLESbIUAHY

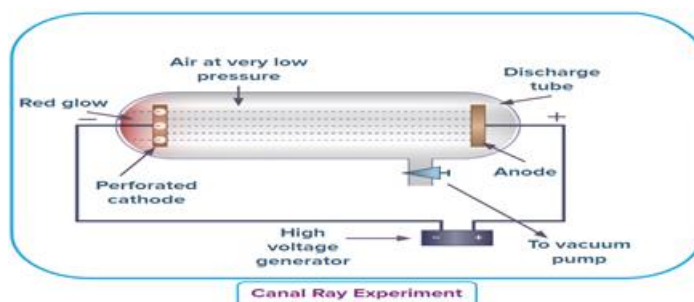
b- The “plum pudding” model of atoms by J-J Thomson 1904

In 1904, Thomson proposed the “plum pudding” model of atoms, which described a positively charged mass with an equal amount of negative charge in the form of electrons embedded in it, since all atoms are electrically neutral.

**3- Discovery Proton****a- Discovery of positive charge in the atom by Eugen Goldstein 1886:**

After his experiment Eugen Goldstein concluded that:

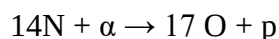
1. The electric field turn positive charged particles.
2. In the magnetic field it deflect towards the North Pole.
3. This experiment proves that hydrogen gaz contain of positive charged particles.

**Link for demonstrative video**

<https://www.youtube.com/watch?v=C3JzHcWl3Tc>

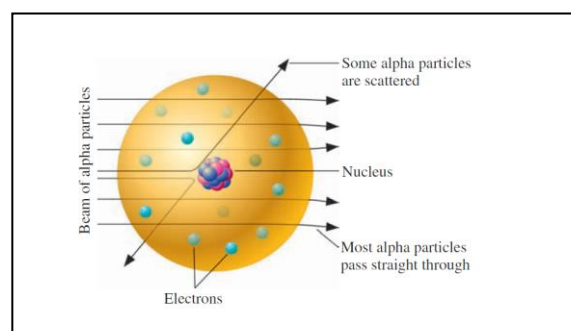
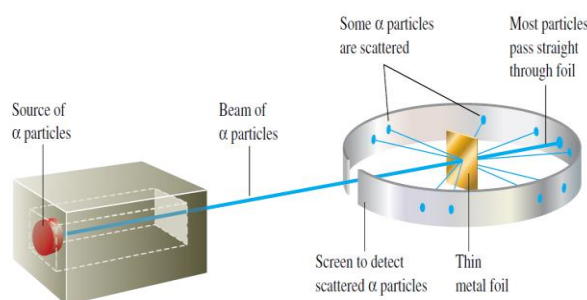
b- Ernest Rutherford experiment 1911

After Eugen Goldstein, Rutherford performed another experiment in the year 1911 to see which particle is responsible for the positive charge of the nucleus. This experiment was the first to report a nuclear reaction, given by the equation:



[Where α is an alpha particle which contains two protons and two neutrons, and 'p' is a proton]

The hydrogen nucleus was later named 'proton' and recognized as one of the building blocks of the atomic nucleus.

4- Discovery of nucleus**a- Ernest Rutherford Experiment Gold Foil Experiment 1909**

Rutherford found that most (about 99%) of the alpha particles he directed at the gold foil passed straight through it. From these experiments, he concluded that the atom contains a small, dense, positively charged nucleus, while the rest of the atom is mostly empty space containing a few negatively charged electrons.

Because alpha particles are positively charged, any deflections occurred when they came close to another positive charge—the nucleus.

After analyzing a series of such experiments in detail, Rutherford drew two main conclusions:

1. Most of the atom's mass and all of its positive charge are concentrated in a tiny central nucleus.
2. The majority of an atom's volume is empty space through which electrons move.

Link for demonstrative video

https://www.youtube.com/watch?v=IQ1h_gdVlHg

5- Discovery of the charge and the mass of the electron

Robert A. Millikan's experiment 1909

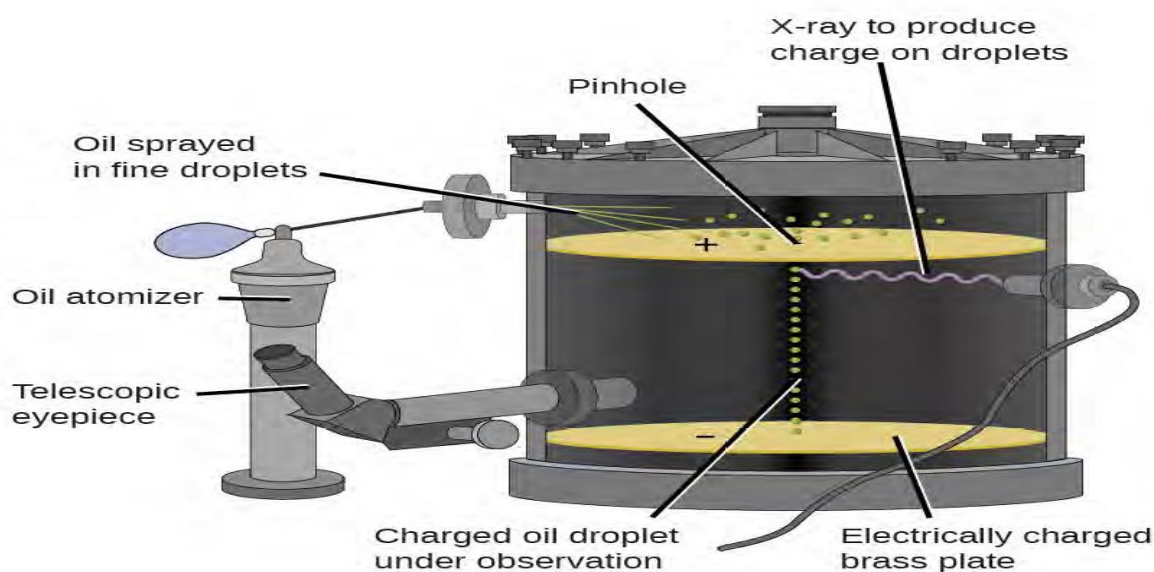
In 1909 Robert Millikan (1868–1953), performed very clever experiments involving charged oil drops. These experiments allowed him to confirmed that the charges was $1.602176634 \times 10^{-19}$ C. He proposed that this was the magnitude of the negative charge of a single electron.

As Thomson determined the charge to mass (e/m) ratio of the particles in his experiment then Milikan was able to determine the mass of an electron :

$$\text{As: } e/m = 1,759 \times 10^{11} \text{ C/Kg} \quad \text{than electron mass} = 9.1093837 \times 10^{-31} \text{ kg}$$

Link for demonstrative video

<https://www.youtube.com/watch?v=2C1o7zXCjCA>



6- Discovery of neutron

Chadwick Experiment 1932

After more than 20 years, in 1932, James Chadwick (1891–1974) discovered the *neutron*, an uncharged dense particle that also resides in the nucleus.

The neutron was discovered by alpha-particle scattering experiments. When beryllium metal is irradiated with alpha rays, a strongly penetrating radiation is obtained from the metal.



Chadwick showed that this penetrating radiation consists of neutral particles, or neutrons. The neutron is a nuclear particle having a mass almost identical to that of the proton but no electric charge. When beryllium nuclei are struck by alpha particles, neutrons are knocked out.

Link for demonstrative video

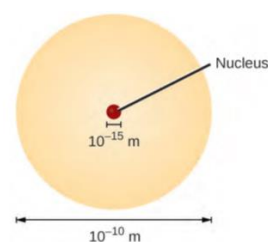
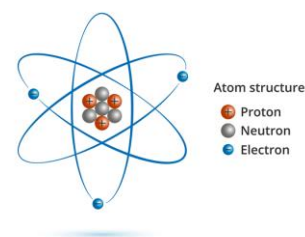
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7- Conclusion about all the experiments

After all these experiments, scientists have now established an atomic model and have been able to determine the characteristics of this model.

The conclusions that scientists were able to establish are:

1. An atom contains a very small nucleus composed of positively charged protons and uncharged neutrons, surrounded by a much larger volume of space containing negatively charged electrons.
2. The nucleus contains the majority of an atom's mass because protons and neutrons are much heavier than electrons, whereas electrons occupy almost all of an atom's volume.
3. The diameter of an atom is on the order of 10^{-10} m, whereas the diameter of the nucleus is roughly 10^{-15} m about 100,000 times smaller.
4. For a perspective about their relative sizes, consider this: If the nucleus were the size of a blueberry, the atom would be about the size of a football stadium



Properties of Subatomic Particles:

Name	Location	Charge (C)	Unit Charge	Mass (amu)	Mass (g)
Electron	Outside the nucleus	-1.602×10^{-19}	-1	0.00055	0.00091×10^{-24}
Proton	In the nucleus	$+1.602 \times 10^{-19}$	+1	1.00727	1.67262×10^{-24}
Neutron	In the nucleus	0	0	1.00866	1.67493×10^{-24}

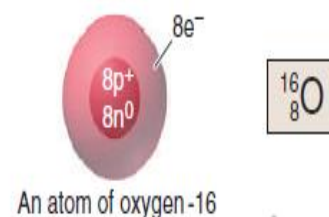
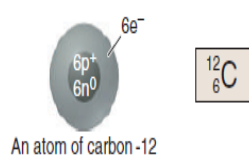
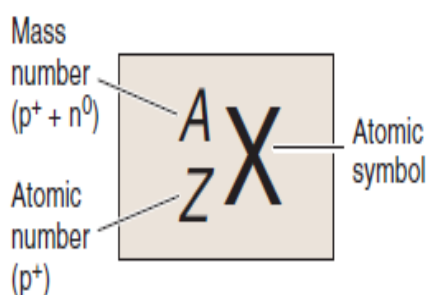
Note: (1 amu = 1.6605×10^{-24} g. (The Dalton (Da) and the unified atomic mass unit (u))

8- Atomic Number, Mass Number, and Atomic Symbol

The mass number (A) is the total number of protons and neutrons in the nucleus of an atom.

The atomic number (Z) of an element equals the number of protons in the nucleus of each of its atoms.

Number of neutrons = mass number - atomic number, or $N = A - Z$

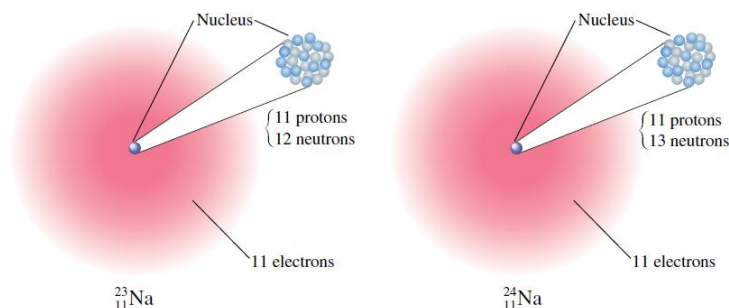


The atomic symbol

Isotopes

All atoms of an element have the same atomic number but not the same mass number.

Isotopes of an element are atoms that have different numbers of neutrons N and therefore different mass numbers A .



Two isotopes of sodium. Both have 11 protons and 11 electrons, but they differ in the number of neutrons in their nuclei.

a- Average mass

The mass of an element shown in the periodic table, or listed in a table of atomic masses, is the weighted average masses of all the isotopes present in a naturally occurring sample of that element. It is equal to the sum of each isotope's mass multiplied by its fractional abundance.

$$\text{average mass} = \sum_i (\text{fractional abundance} \times \text{isotopic mass})_i$$

Or

$$\text{average mass} = \frac{\sum_i \text{abundance} \times \text{isotopic mass}}{100}$$

b- Mass defect

Mass defect is the difference between the actual atomic mass and the predicted mass obtained by adding the masses of the protons and neutrons present in the nucleus. The actual atomic mass is less than the predicted mass calculated from the sum of the nucleon masses.

This missing mass corresponds to the binding energy released when the nucleus is formed. During the formation of a nucleus, a portion of the mass is converted into energy, resulting in the mass defect. For this reason, the actual mass of an atomic nucleus is always less than the combined mass of the particles that make it up.

$$\Delta m = (Zm_p + Nm_n) - m_A$$

Δm = mass defect

m_A = mass of the nucleus

m_p = mass of a proton, i.e. 1.00728 amu

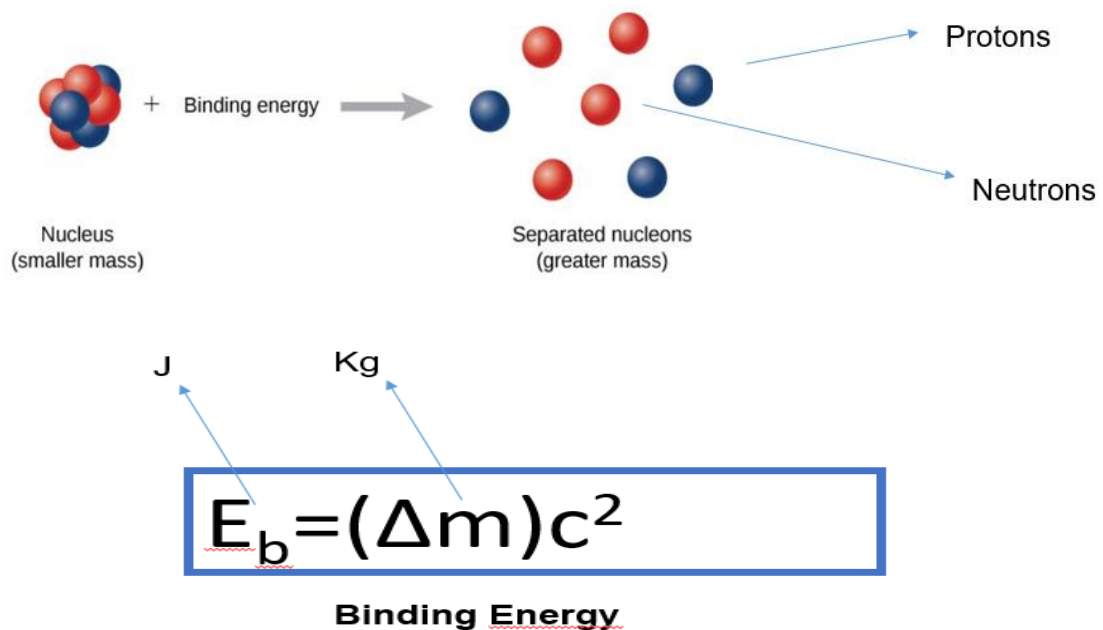
m_n = mass of a neutron, i.e. 1.00867 amu

Z = number of protons

N = number of neutrons

c- Binding Energy

Binding energy is defined as the minimum amount of energy required to remove a particle from a system of particles. In other words, it is the energy needed to separate a system of particles into individual components.



d- Binding Energy per Nucleon and stability

The greater the binding energy per nucleon, the more the stability of the nucleus is.

$\frac{E_b}{A} = \frac{\Delta m \times c^2}{A}$
 Binding Energy Mass defect The mass number