Chapter 2: The atom's structure

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Objectifs

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

- Understand the historical development of chemistry 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.

I The atom's structure

1. Introduction

The atom, a particle of extremely small dimension in comparison with those we are more familiar with, is the fundamental unit of matter.

The atom has a dimension of about 10^{-8} cm which is well below the resolution of the most powerful optical and scanning electron microscopes. A pin-head, whose dimension represents the resolution (defined as the smallest distance between two points on a specimen that can still be distinguished as two separate entities) power of an human eye (about 1/10 mm or 100 microns), has a size that would contain about one million atoms.

The name atom comes from the Greek ἄτομος (*atomos*, "indivisible"), which means uncuttable, or indivisible, something that cannot be divided further. During the late 19th and early 20th centuries, physicists discovered subatomic components and structure inside the atom, thereby demonstrating that the 'atom' was divisible. The concept of indivisibility can actually be applicable as long as the uniqueness of matter relates with what we macroscopically touch, see, smell and call matter. However, atoms as well has their own internal structure.

2. The constituents of the atom

2.1. The electron

If matter were composed of atoms, what were atoms composed of? Were they the smallest particles, or was there something smaller? In the late 1800s, a number of scientists interested in questions like these investigated the electrical discharges that could be produced in low-pressure gases, the most significant discovery made by English physicist *J. J. Thomson*⁎ using a cathode ray tube. This apparatus consisted of a sealed glass tube from which almost all the air had been removed; the tube contained two metal electrodes. When high voltage was applied across the electrodes, a visible beam called a cathode ray appeared between them. This beam was deflected toward the positive charge and away from the negative charge, and was produced in the same way with identical properties when different metals were used for the electrodes. The results of these measurements indicated that these particles were much lighter than atoms. Based on his observations, here is what Thomson proposed and why: The particles are attracted by positive (+) charges and repelled by negative (−) charges, so they must be negatively charged they are less massive than atoms and indistinguishable, regardless of the source material, so they must be fundamental, subatomic constituents of all atoms. Although controversial at the time, Thomson’s idea was gradually accepted, and his cathode ray particle is what we now call an electron.

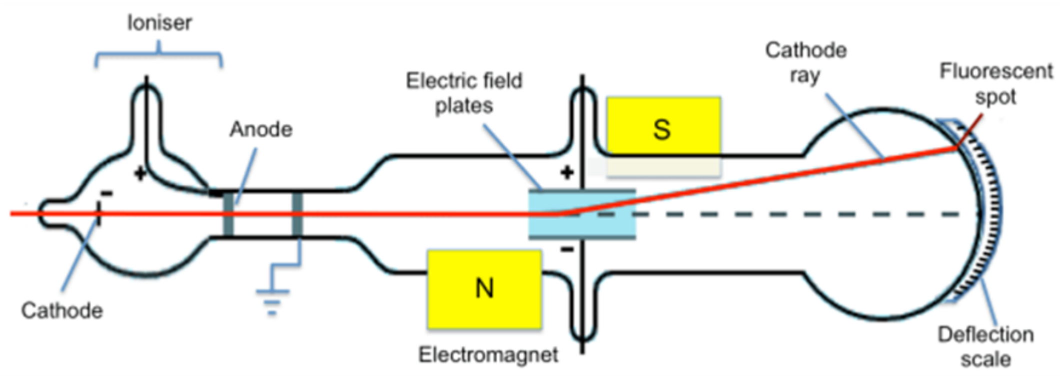
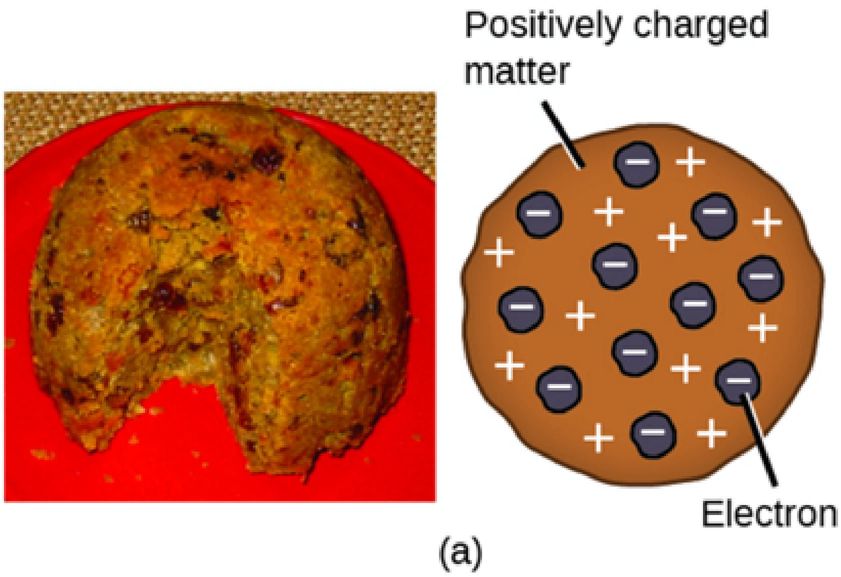


Diagram of J.J. Thomson's cathode ray tube.

2.2. Thomson's atomic model

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.



Thomson suggested that atoms resembled plum pudding "An English dessert consisting of moist cakewith embedded raisins (plums) ".

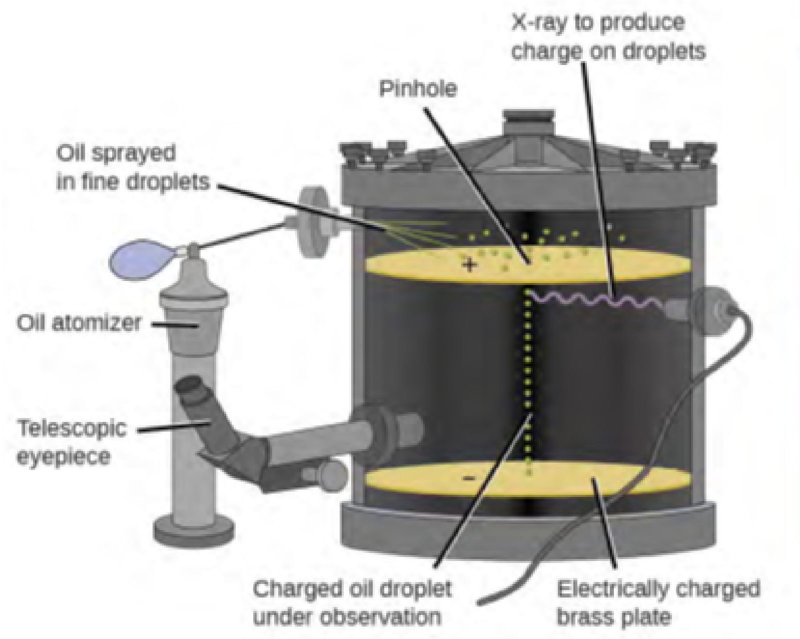
2.3. Millikan experiment: determining the charge of the electron

In 1909, more information about the electron was uncovered by American physicist *Robert A. Millikan** via his “oil drop” experiments. Millikan created microscopic oil droplets, which could be electrically charged by friction as they formed or by using X-rays. These droplets initially fell due to gravity, but their downward progress could be slowed or even reversed by an electric field lower in the apparatus. By adjusting the electric field strength and making careful measurements and appropriate calculations, Millikan was able to determine the charge on individual drops. Looking at the charge data that Millikan gathered, you may have recognized that the charge of an oil droplet is always a multiple of a specific charge, 1.6×10^{-19} C. Millikan concluded that this value must therefore be a fundamental charge—the charge of a single electron—with his measured charges due to an excess of one electron (1 times 1.6×10^{-19} C), two electrons (2 times 1.6×10^{-19} C), three electrons (3 times 1.6×10^{-19} C), and so on, on a given oil droplet[1].

1. https://www.youtube.com/watch?v=nwnjYERS66U

Since the charge of an electron was now known due to Millikan’s research, and the charge-to-mass ratio was already known due to Thomson’s research (1.759 × 10^{11} C/kg), it only required a simple calculation to determine the mass of the electron as well.

$$Mass\,of\,electron = 1{,}602 * 10^{-19}\frac{1\,Kg}{1{,}759{*}10^{11}C} = 9{,}107 * 10^{-31}\,Kg$$



Oil drop	Charge in coulombs (C)
A	4.8×10^{-19} C
B	3.2×10^{-19} C
C	6.4×10^{-19} C
D	1.6×10^{-19} C
E	4.8×10^{-19} C

Millikan’s experiment measured the charge of individual oil drops. The tabulated data are examples of afew possible values

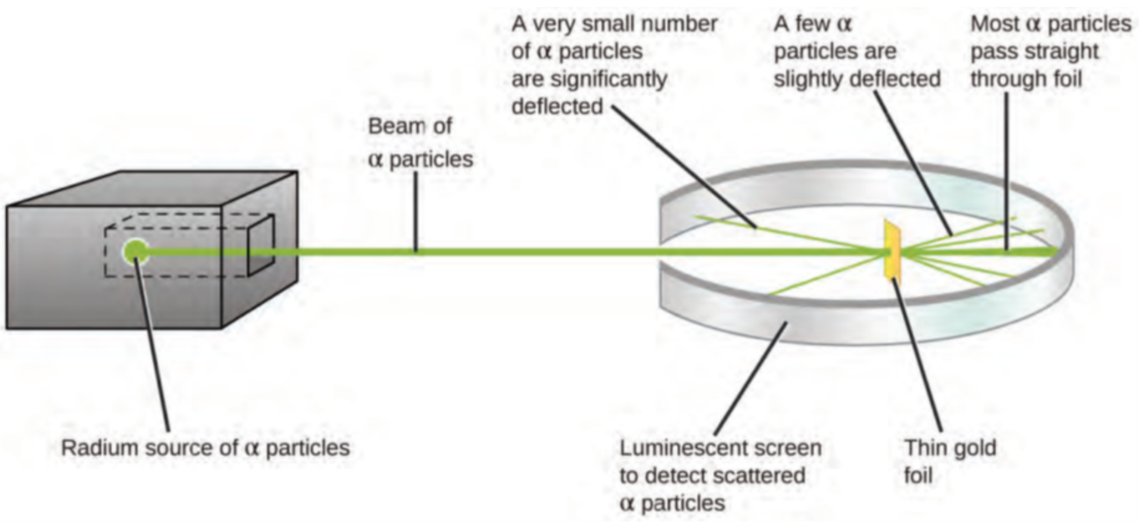
2.4. Rutherford experiment

The next major development in understanding the atom came from *Ernest Rutherford.*∗ He performed a series of experiments using a beam of high-speed, positively charged **alpha particles (α particles)** Rutherford aimed a beam of α particles, the source of which was embedded in a lead block to absorb most of the radiation, at a very thin piece of gold foil and examined the resultant scattering of the α particles using a luminescent screen that glowed briefly where hit by an α particle.

Most particles passed right through the foil without being deflected at all. However, some were diverted slightly, and a very small number were deflected almost straight back toward the source. Rutherford deduce that : Because most of the fast-moving α particles passed through the gold atoms undeflected, they must have traveled through essentially empty space inside the atom. Alpha particles are positively charged, so deflections arose when they encountered another positive charge (like charges repel each other). Since like charges repel one another, the few positively charged α particles that changed paths abruptly must have hit, or closely approached, another body that also had a highly concentrated, positive charge. Since the deflections occurred a small fraction of the time, this charge only occupied a small amount of the space in the gold foil. Analyzing a series of such experiments in detail, Rutherford drew two conclusions:

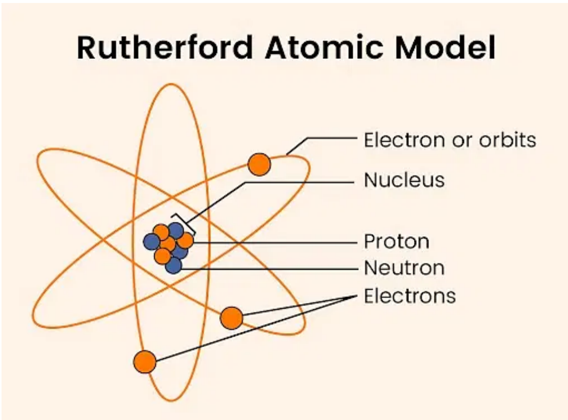
1. The volume occupied by an atom must consist of a large amount of empty space.

2. A small, relatively heavy, positively charged body, the **nucleus**, must be at the center of each atom.



Rutherford experiment

Rutherford devised a new atomic structure: the atom is largely made up of a vacuum, which allows most alpha particles to pass through the sheet. All the positive charge and almost all the mass are concentrated in the centre of the atom, in the nucleus. So the alpha particles that are deflected are those that have hit the nucleus[2].



Rutherford's atomic model

2.5. Discovery of Protons

Ernest Rutherford observed that his scintillation detectors detected hydrogen nuclei when a beam of alpha particles was shot into the air. After investigating further, Rutherford found that these hydrogen nuclei were produced from the nitrogen atoms present in the atmosphere.

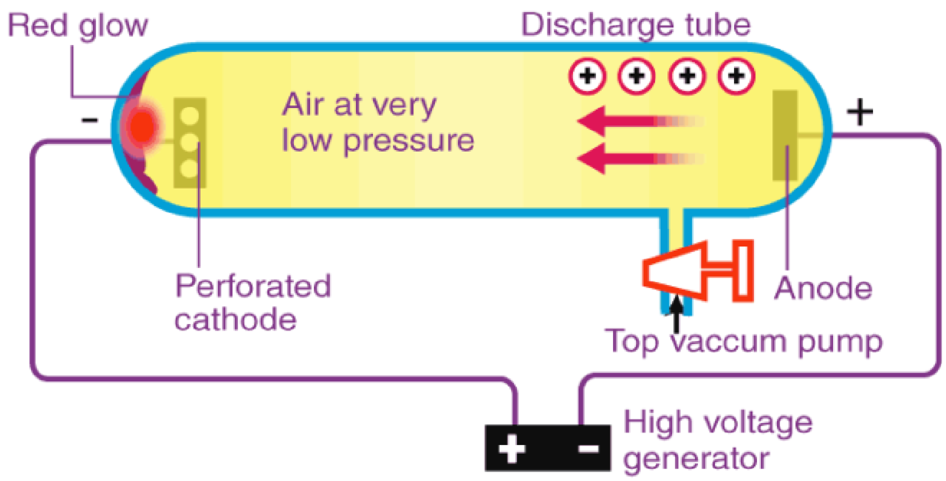
He then proceeded to fire beams of alpha particles into pure nitrogen gas and observed that a greater number of hydrogen nuclei were produced. He concluded that the hydrogen nuclei originated from the nitrogen atom, proving that the hydrogen nucleus was a part of all other atoms.

This experiment was the first to report a nuclear reaction, given by the equation:

$$14N + \alpha \rightarrow 170 + P$$

[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.

2. https://www.youtube.com/watch?v=kHaR2rsFNhg



Discovery of Proton canal ray experiment

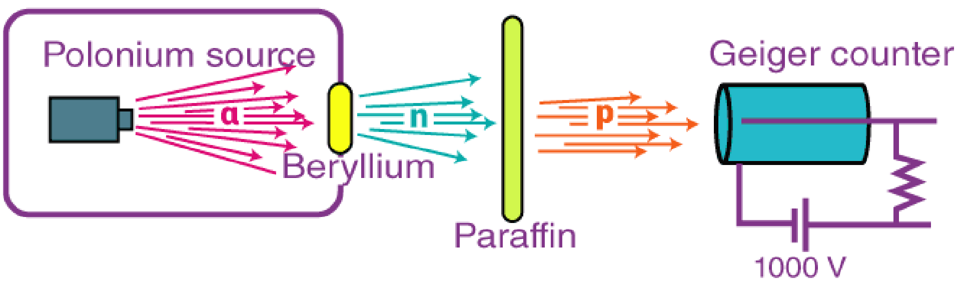
Proton charge= + 1.6 x 10^{-19} Proton mass= 1.673 x 10^{-27} Kg.

2.6. Discovery of Neutrons

James Chadwick fired alpha radiation at beryllium sheet from a polonium source. This led to the production of an uncharged, penetrating radiation. This radiation was made incident on paraffin wax, a hydrocarbon having a relatively high hydrogen content.

The protons ejected from the paraffin wax (when struck by the uncharged radiation) were observed with the help of an ionization chamber. The range of the liberated protons was measured and the interaction between the uncharged radiation and the atoms of several gases was studied by Chadwick.

He concluded that the unusually penetrating radiation consisted of uncharged particles having (approximately) the same mass as a proton (1.675 x 10^{-27} kg). These particles were later termed ‘neutrons’.



Experimentation of Neutron discovery

Mass of a neutron: 1.675 x 10^{-27} kg

Matter	Melting Temperature (°C)	Boiling point (°C)
water : H_2O	0	100
Sodium chloride: NaCl	801	1465
Butane : C_4H_{10}	-138	-1

Tableau 1

3. Stucture of the Atom

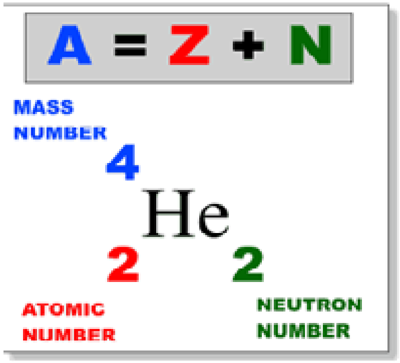
The number of protons in the nucleus of an atom is its atomic number (Z). This is the defining trait of an element: Its value determines the identity of the atom. For example, any atom that contains six protons is the element carbon and has the atomic number 6, regardless of how many neutrons or electrons it may have. A neutral atom must contain the same number of positive and negative charges, so the number of protons equals the number of electrons. Therefore, the atomic number also indicates the number of electrons in an atom. The total number of protons and neutrons in an atom is called its mass number (A). The number of neutrons is therefore the difference between the mass number and the atomic number: A – Z = number of neutrons.

Atomic number (Z) = number of protons

Atomic mass (A) = number of protons + number of neutrons

A − Z = number of neutrons

Atoms are electrically neutral if they contain the same number of positively charged protons and negatively charged electrons. When the numbers of these subatomic particles are not equal, the atom is electrically charged and is called an ion. The charge of an atom is defined as follows: Atomic charge = number of protons − number of electrons.



constituent of the Helium atom

The mass of an atom should be calculated simply by summing the masses of its various constituents:

$$m_{atome} = Z * m_p + Z * m_e + N * m_N$$

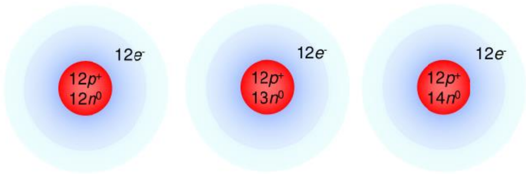
Formule 1

The electron is a much lighter particle; its mass is approximately 2000 times less than that of the proton or neutron (mp/me =1833), so we can neglect it.

4. Isotopes

Isotope are chemical element with the same atomic number and position in the periodic table and nearly identical chemical behaviour but with different atomic masses and physical properties. Every chemical element has one or more isotopes.

For example, magnesium exists as a mixture of three isotopes, each with an atomic number of 12 and with mass numbers of 24, 25, and 26, respectively.



Atomic symbol	$^{24}_{12}Mg$	$^{25}_{12}Mg$	$^{26}_{12}Mg$
Number of protons	**12**	**12**	**12**
Number of electrons	**12**	**12**	**12**
Mass number	**24**	**25**	**26**
Number of neutrons	**12**	**13**	**14**
Isotope Notation	**Mg-24**	**Mg-25**	**Mg-26**

Isotopes of magnesium

4.1. Atomic Mass

Because each proton and each neutron contribute approximately one amu (atomic mass unit) to the mass of an atom, and each electron contributes far less, the atomic mass of a single atom is approximately equal to its mass number (a whole number).

However, the average masses of atoms of most elements are not whole numbers because most elements exist naturally as mixtures of two or more isotopes. The mass of an element shown in a periodic table or listed in a table of atomic masses is a weighted, average mass of all the isotopes present in a naturally occurring sample of that element. This is equal to the sum of each individual isotope's mass multiplied by its fractional abundance.

$$Average\,mass = \Sigma_i(fractional\,abundance \times isotopic\,mass)_i$$

Formule 2

For example, the element boron is composed of two isotopes: About 19.9% of all boron atoms are 10B with a mass of 10.0129 amu, and the remaining 80.1% are 11B with a mass of 11.0093 amu. The average atomic mass for boron is calculated to be:

boron average mass = (0.199 × 10.0129 amu) + (0.801 × 11.0093 amu)

= 1.99 amu + 8.82 amu

= 10.81 amu

It is important to understand that no single boron atom weighs exactly 10.8 amu; 10.8 amu is the average mass of all boron atoms, and individual boron atoms weigh either approximately 10 amu or 11 amu.

Glossaire

J J Thomson

Sir Joseph John Thomson (18 December 1856 – 30 August 1940) was a British physicist and Nobel Laureate in Physics, credited with the discovery of the electron, the first subatomic particle to be found.

Millikan

Robert Andrews Millikan was an American experimental physicist who won the Nobel Prize for Physics in 1923 for the measurement of the elementary electric charge and for his work on the photoelectric effect. Millikan graduated from Oberlin College in 1891 and obtained his doctorate at Columbia University in 1895.

rutherford

Ernest Rutherford, 1st Baron Rutherford of Nelson, (30 August 1871 – 19 October 1937) was a New Zealand physicist who was a pioneering researcher in both atomic and nuclear physics. Rutherford has been described as "the father of nuclear physics