AL-Mustaqbal University

College of Sciences

Department of Medical Physics

Atomic physics

MCs. Sarah Jaleel Ahmad

MCs. Baraa Abd Alrda

MCs. Baraa Abd Alrda

First lecture

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1. Introduction

The familiar model of an atom is that of a small nucleus composed of protons and neutrons surrounded by rapidly moving electrons. Typically, the atomic diameter is on the order of $l0^{-10}$ m while that of the nucleus is on the order of 10^{-15} m. Protons and neutrons have about the same mass (1.00728 and 1.00867 amu respectively) and each is about 1800 times as heavy as an electron. A neutron is electrically neutral, but a proton has a positive charge (+1.6 x 10^{-19} coulomb) which is exactly the opposite of the negative charge of an electron. In a neutral atom, the number of electrons around the nucleus equals the number of protons in the nucleus. The number of protons in the nucleus (the "atomic number", Z) characterizes a chemical element. Atoms of a given element all have the same number of protons, yet may have different masses. The atomic mass number of an atom, A, is given by A = Z + N, where N is the number of neutrons in the nucleus. Since an element is characterized solely by Z, it follows that atoms of a given chemical element may have a varying number of neutrons. Subspecies of chemical elements with the same Z but differing N and A are called isotopes. The atomic weight of an element is the weighted average of the atomic masses of the various naturally occurring isotopes of the element, and the atomic weight scale is based on a value of exactly 12, after the carbon isotope that has an atomic mass number of 12.

Table 1 : Fundamental particles of atom and their characteristics

Particle	Symbol	Mass/ kg	Actual Charge / C	Relative charge
Electron	e	$9.109\,389 \times 10^{-31}$	$-1.602\,177 \times 10^{-19}$	-1
Proton	p	$1.672\,623 \times 10^{-27}$	$1.602\,177 \times 10^{-19}$	+1
Neutron	n	$1.674\,928 \times 10^{-27}$	0	0

Since atoms are made up of still smaller particles, they must have an internal structure. In the next section we shall take up some of the earlier ideas about the internal structure of atom.

2. Atomic Models

Once it was established that the atom is not indivisible, the scientists made attempts to understand the structure of the atom. A number of models have been proposed for the internal structure of the atom. The first attempt to describe the structure of atom in terms of a model was made by J.J Thomson.

2.1 Thomson's Model

On the basis of his experiments on discharge tubes, Thomson proposed that atoms can be considered as a large positively charged body with a number of small negatively charged electrons scattered throughout it. This model (Fig. 1) was called as plum pudding model of the atom.

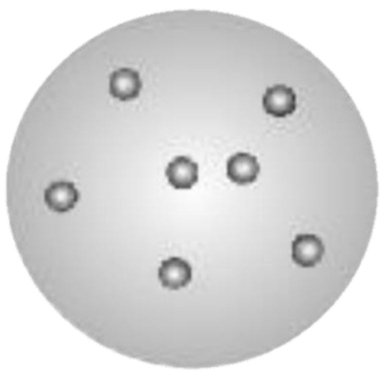

Fig. 1 : A pictorial representation of Thomson's plum-pudding model.

The electrons represent the plums in the pudding made of positive charge. It is sometimes also called as watermelonmodel. In this, the juicy pulp of the watermelon represents the positive charge and the seeds represent the electrons.

J.J.Thomson
(1856-1940)
Won Nobel prize in Physics in 1906

Ernest Rutherford
(1871-1937)
Won Nobel prize in Chemistry in 1908

2.2 Rutherford's Experiment

Ernest Rutherford performed an experiment called 'Gold Foil Experiment' or 'α-ray scattering experiment' to test the structure of an atom as proposed by Thomson. In this experiment a beam of fast moving alpha particles (positively charged helium ions) was passed through a very thin foil of gold. He expected that the alpha particles would just pass straight through the gold foil and could be detected by a photographic plate. But, the actual results of the experiment (Fig. 3.2) were quite surprising. It was observed that most of the □-particles did pass straight through the foil but a number of particles were deflected from their path. Some of these deflected slightly while a few deflected through large angles and about 1 in 10,000 α- particles suffered a rebound.

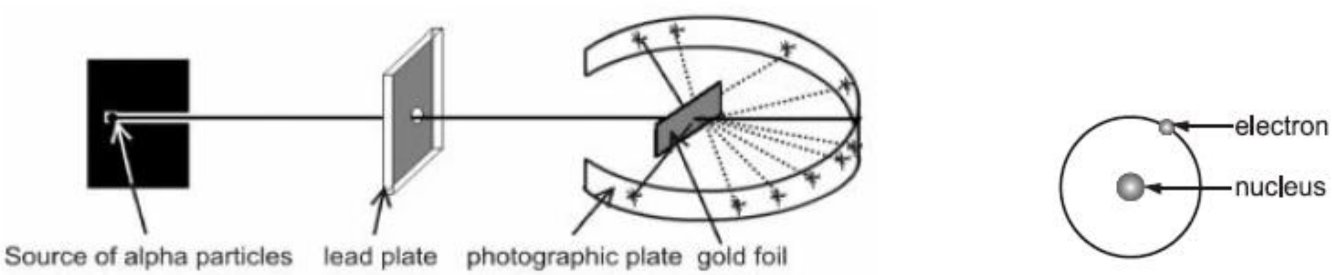

Fig. 2 : Schematic representation of Rutherford's α-ray scattering experiment.

These results led Rutherford to conclude that :

1- the atom contained some dense and positively charged region located at the center of the atom that he called as nucleus.

2- all the positive charge of the atom and most of its mass was contained in the nucleus.

3- the rest of the atom must be empty space which contains the much smaller and negatively charged electrons (Fig. 2).

The model proposed by Rutherford explained the observation in the α-ray scattering experiments as shown below in Fig. 3.

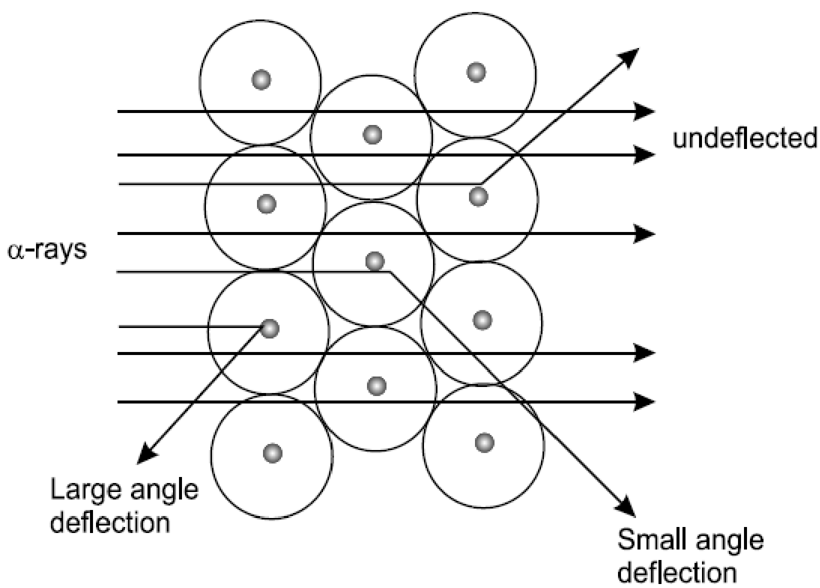

Fig. 3 : Explanation of the results of α-ray scattering experiment.

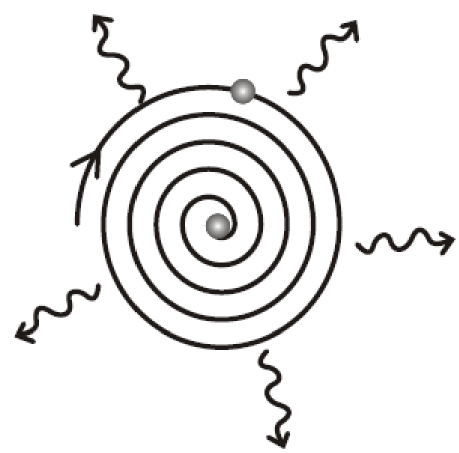

Fig. 4 : Failure of Rutherford's model.

However, there was a problem with the Rutherford's model. According to the Maxwell's theory of electromagnetic radiation, a charged particle undergoing acceleration would continuously emit radiation and lose energy. Since the electron in the atom is also a charged particle and is under acceleration, it is expected to continuously lose energy. As a consequence, the electron moving around the nucleus would approach the nucleus by a spiral path (Fig. 4) and the atom would collapse. However, since it does not happen we can say that the Rutherford's model failed to explain the stability of the atom. The next attempt to suggest a model for atom was made by Neils Bohr- a student of Rutherford. This model used the concept of quantisation of energy of electrons in the atom. Since this fact was suggested by line spectrum of hydrogen atom it is worthwhile to understand the meaning of a spectrum. For this we begin with the understanding of the nature of an electromagnetic radiation.

2.3 Bohr's Model

In 1913, Niels Bohr (1885-1962) proposed another model of the atom where electrons move around the nucleus in circular paths. Bohr's atomic model is built upon a set of postulates, which are as follows :

Bohr won the Nobel Prize in Physics in 1922 for his work.

1. The electrons move in a definite circular paths around the nucleus. He called these circular paths as orbits and postulated that *as long as the electron is in a given orbit its energy does not change* (or energy remains fixed). These orbits were therefore referred to as stationary orbits or stationary states or non-radiating orbits.

2. The *electron can change its orbit by absorbing or releasing energy*. An electron at a lower (initial) state of energy, E_i can go to a (final) higher state of energy, E_f by absorbing (Fig 7) a single photon of energy given by

$$E = h\upsilon = E_f - E_i \quad \text{... (2)}$$

Similarly, when electron changes its orbit from a higher initial state of energy E_i to a lower final state of energy E_f , a single photon of energy hν is released.

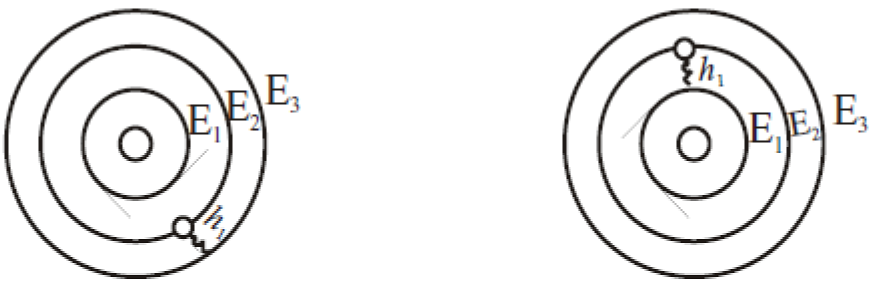

Fig. 7 : Absorption and emission of photon causes the electron to change its energy level.

3. The angular momentum of an electron of mass me moving in a circular orbit of radius r and velocity v is an integral multiple of $h/2\pi$.

$$m_e vr = \frac{n\hbar}{2\pi} \quad \text{... (3)}$$

where n is a positive integer, known as the principal quantum number. Bohr obtained the following expressions for the energy of an electron in stationary states of hydrogen atom by using his postulates :

Energy of the orbit,

$$E_n = -R_H\left(\frac{1}{n^2}\right) \quad \text{... (4)}$$

The negative sign in the energy expression means that the there is an attractive interaction between the nucleus and the electron. This means that certain amount of energy (called ionisation energy) would be required to remove the electron from the influence of the nucleus in the atom. You may note here that the energies of the Bohr orbits are inversely proportional to the square of the quantum number n. As n increases the value of the energy increases (becomes lesser negative or more positive). It means that as we go farther from the nucleus the energy of the orbit goes on increasing.