The Contributions of Ernest Rutherford in Nuclear Science - Empirical – An Analysis

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Abstract

This paper seeks to study groundbreaking experiments by Rutherford that would completely change the accepted model of the atom with nucleus. Although some ancient Greeks (such as Democritus) postulated the existence of atoms (units of matter which could not be subdivided), concrete evidence for their existence did not develop until the 19th century. The first direct evidence came from observations of the Brownian motion. Other evidence came from chemistry, such as the experiments of Faraday on electrolysis (1833). Faraday's law states that when a current is passed through a solution or through a molten electrolyte, the mass of a particular element deposited at the cathode or anode is proportional to the electrical charge which has flowed around the circuit (current times time) and to the atomic weight of that element, but inversely proportional to the valence of the element. Although its interpretation was not clear at the time, Faraday's law reflects the fact that electricity passes through an electrolyte in the form of ions whose mass is proportional to the atomic weight and whose charge is equal to the valence, which represents the number of electrons which have been removed from the neutral atom.

The properties of electrons were investigated by J.J.Thomson at the Cavendish Laboratory in 1897 . An electrical discharge in a low-pressure gas was known to produce cathode rays, which could cause an object in their path to emit light (fluoresce), but it was not known if these rays consisted of waves or material particles. The value of e/m was about 2000 times larger than the charge/mass ratio measured (by electrolysis of water) for hydrogen ions, indicating a particle much smaller in mass than the smallest atom. The electronic charge e was first determined by Robert Millikan (in 1909) by observing the motion of individual drops of oil in an vapour, ionized with a radioactive source. A drop acquires a terminal velocity as a result of a balance of the viscous(air-resistance), gravitational and (with an electric field applied) electrostatic forces. Applying a formula for the viscous force on a spherical drop and measuring the velocity with an without applied field, the charge on the drop can be calculated. The electron was discovered by J.J. Thomson in 1897. The existence of protons was also known, as was the fact that atoms were neutral in charge. Since the intact atom had no net charge and the electron and proton had opposite charges, the next step after the discovery of subatomic particles was to figure out how these particles were arranged in the atom. This is a difficult task because of the incredibly small size of the atom. Therefore, scientists set out to design a model of what they believed the atom could look like.

Key words: protons, subatomic particles, electrostatic forces, low-pressure gas, Rutherford, Millikan.

Introduction

Although atoms were originally thought of as being indivisible, evidence began to accumulate (towards the end of the 19th century) that the atom contains component parts: electrons (which can be emitted as cathode rays) and a balancing positive charge (to give overall neutrality and prevent the atom exploding from the electrostatic repulsion of all the negatively charged electrons). The goal of each atomic model was to accurately represent all of the experimental evidence about atoms in the simplest way possible.

In 1898, J.J. Thomson proposed that the electrons are embedded (like *plums in a pudding*) in a sphere of uniformly distributed positive charge. Other physicists came up with other ideas, for example the positive charge might be concentrated in a *central nucleus* with the electrons orbiting around it, analagous to the solar system, except that the attractive forces would be electrostatic rather than gravitational. Following the discovery of the electron, J.J. Thomson developed what became known as the "**plum pudding**" model in 1904. Plum pudding is an English dessert similar to a blueberry muffin. In Thomson's plum pudding model of the atom, the electrons were embedded in a uniform sphere of positive charge like blueberries stuck into a muffin. The positive matter was thought to be jelly-like or a thick soup. The electrons were somewhat mobile. As they got closer to the outer portion of the atom, the positive charge in the region was greater than the neighboring negative charges and the electron would be pulled back more toward the center region of the atom.

Experiments conducted between 1909 and 1924, supervised by Ernest Rutherford but actually carried out by two of his students (Hans Geiger and Ernst Marsden), provided the evidence necessary to choose between these models. The experiments arose out of investigations of **radioactive** materials, discovered by Becquerel in 1896 when he accidentally left wrapped photographic plates in a drawer together with some uranium-salt crystals (he shared a 1903 Nobel prize with Pierre and Marie Curie). According to the accepted atomic model, in which an atom's mass and charge are uniformly distributed throughout the atom, the scientists expected that all of the alpha particles would pass through the gold foil with only a slight deflection or none at all.

Rutherford needed to come up with an entirely new model of the atom in order to explain his results. Because the vast majority of the alpha particles had passed through the gold, he reasoned that most of the atom was empty space. In contrast, the particles that were highly deflected must have experienced a tremendously powerful force within the atom. He concluded that all of the positive charge and the majority of the mass of the atom must be concentrated in a very small space in the atom's interior, which he called the nucleus. The **nucleus** is the tiny, dense, central core of the atom and is composed of protons and neutrons.

Rutherford's atomic model became known as the **nuclear model**. In the nuclear atom, the protons and neutrons, which comprise nearly all of the mass of the atom, are located in the nucleus at the center of the atom. The electrons are distributed around the nucleus and occupy most of the volume of the atom. It is worth emphasizing just how small the nucleus is compared to the rest of the atom. If we could blow up an atom to be the size of a large professional football stadium, the nucleus would be about the size of a marble.

Rutherford's model proved to be an important step towards a full understanding of the atom. However, it did not completely address the nature of the electrons and the way in which they occupied the vast space around the nucleus. For this and other insights, Rutherford was awarded the Nobel Prize in Chemistry in 1908. Unfortunately, Rutherford would have preferred to receive the Nobel Prize in Physics because he considered physics superior to chemistry. In his opinion, "All science is either physics or stamp collecting."

Objective:

This paper intends to explore the discoveries of Ernest Rutherford using alpha and beta rays, set forth the laws of radioactive decay, and identified alpha particles as helium nuclei.Most importantly, he postulated the nuclear structure of the atom.

Rutherford's Model of an Atom

We know a structure of an atom consists of electrons, protons, and neutrons. This was accurately presented after several scientists came up with different models. The classic model of an atom was given by Ernest Rutherford called the Rutherford atomic model or Rutherford model of the atom. However, it is not considered the accurate representation of an atom anymore.

Thomson passed a beam of cathode rays through a uniform electric field E (between two parallel plates) and a magnetic field B (produced by an electromagnet) acting over the same region of space.The two fields were perpendicular to the beam and to each other, such that the magnetic and electrical forces were both perpendicular to the beam but opposite in direction.

By adjusting the strength of one of the fields, so as to produce zero net deflection as observed on a fluorescent screen, he achieved the condition:

$$e E + B e V_x = 0$$

so was able to measure the speed Vx of the particles in the beam (travelling in the x-direction). With the magnetic field turned off the electric field produced a force (and therefore an acceleration a) on each particle over a distance L, deflecting the beam through an angle q which could be measured. The magnitude Vy of the velocity acquired in a direction parallel to the field (i.e. perpendicular to the beam) can be calculated from

Vy = a t = (eE/m) (L/Vx) , so that: tan (theta) = Vy / v = (L E / Vx^2) (e/m)

By measuring (theta) , L and E (=V/d where V is the voltage between the plates, separation d) the charge-to-mass ratio e/m of the particles could be found. The value of e/m was found to be independent of the electrodes and the gas in the discharge tube, suggesting that the cathode rays consisted of particles which were a constituent of all matter.

Rutherford Scattering

The apparatus was quite simple: a radioactive source (such as radon gas) emitted **alpha particles**, which have a charge of +2e and a mass about four times that of the hydrogen atom (they are actually helium-atom nuclei). A parallel beam of these particles (collimated by a lead tube) was directed towards a thin foil of a metal such as gold, mounted in a vacuum chamber. Some of the alpha particles passed through the foil without measurable deviation but some were scattered through appreciable angles (j in Fig.3.9). The **angular distribution** of the scattered a -particles was measured by mounting a fluorescent screen (a slab of glass coated with fine zinc suphide particles) at the entrance of a low-power telescope, which could be rotated about the scattering point; see Fig. 3.9. With a *dark-adapted* eye, single a -particles hitting the screen could be detected as a small flash of light seen in the telescope. Most of the alpha particles were deflected through small angles (of the order of 1 degree) but a small fraction were scattered through larger angles, including some **backscattered** through angles exceeding 90 degrees = p /2 radians; see Fig. 3.11.

Rutherford realized that the existence of large-angle scattering ruled out the Thomson (plum-pudding) model; the relatively heavy a -particles would not be turned around by much lighter electrons or by the combined mass of a gold atom *if this mass were distributed over the whole atomic volume*. On the other hand, if the positive charge and most of the mass of a gold atom were concentrated in a central nucleus, its electrostatic repulsion would repel incoming a -particles and deflect some of them through angles as large as 180 degrees (see Fig. 3.10) *without* absorbing much of the a -particle's energy (*i.e.* the "collision" will be **elastic**). This scattering from the *electrostatic field of the nucleus* is now known as *Rutherford scattering*. Note that if the nucleus is small compared to the whole atom (it typically occupies less than 1 part in 10^15 of the volume!), the *probability* of high-angle a -scattering will be very small, in agreement with the measured angular distribution (note the *logarithmic* vertical scale in Fig. 3.11).

Based on his nuclear model of the atom, Rutherford was able to calculate an expression for the angular distribution of the a -particle scattering. He needed to know the magnitide F of the force on an a -particle when it is a distance r from the centre of a nucleus; this is given by Coulomb's law:

F = k (+2e)(Ze)/r^2

where k is the Coulomb constant and (Ze) is the nuclear charge, Z being the atomic number of the atoms in the foil. Applying Newton's second law (and conservation of momentum and energy) to the two-particle interaction led to the following expression for the number n of alpha particles detected at a scattering angle (phi) :

n = C (N Z^2 / K^2) / [sin (phi)/2]^4

where K is the kinetic energy of the alpha -particles and C is a parameter which depends on the strength of the alpha -particle source and the geometry of the particle detector; N is the number of nuclei per unit area of the foil, equal to (rho) t / (A u) where t and (rho) are the foil thickness and density, while A represents atomic weight and u is the atomic-mass unit. Careful experiments by Geiger and Marsden, published in 1913 (Philosophic Magazine 25, p.605), confirmed the t, K and phi dependence implied by this formula.

At that time, atomic numbers for different elements were not known; however, Rutherford was able to *fit* the results obtained from different foils to his formula by trying different values of Z ; see Fig. 3.11. Thereby, he was able to *measure* Z , while the relatively good degree of fit provided additional confirmation of his theory.

Furthermore, Rutherford realized that the remarkable success of his formula provided information about the *size* of the nucleus. He argued that if an alpha actually *reached* the nucleus, the latter would be "deformed" - in other words the force law would depart from the Coulomb's law expression and the angular dependence of scattering would change. A particle which gets closest to a nucleus will be one which directly approaches its centre and is deflected through 180 degrees (as in Fig. 3.10). At the moment of its closest approach, this particle will be momentarily stationary, having exchanged all of its kinetic energy K into electrostatic energy of repulsion, so that:

K = k (Ze) (2e) / r

Smaller values of r will be possible by choosing foils of lower atomic number and alpha particles of higher kinetic energy. The experimenters therefore worked with aluminum foils and other radioactive sources which provided more penetrating radiation, looking for evidence for departures from Rutherford's formula in the measured angular distributions. In 1919, they were able to detect a departure from the formula for 7.7MeV alpha-particles (K = 7.7 x 10^6 / 1.6 x 10^-19 Joule) scattered from a foil of aluminum (Z=13), allowing the radius of the Al nucleus to be estimated as: r = 2 k Z e^2 / K = 4.9 x 10^-15 m. Since the radius of an atom is of the order of 10^-10 m, the atom is seen to be mainly empty space.

Despite the success of the nuclear atomic model, it raised several questions. For example, the atomic number (a measure of the nuclear charge) of an element turned out to be more than a factor two lower than its atomic weight (a measure of the nuclear mass mass). Rutherford speculated that *the nucleus might contain electrons,* which would neutralize some of its positive charge without adding appreciable mass. This seemed plausible at the time, because certain types of radioactive materials emit high-energy electrons (beta- rather than alpha-decay) which might come from the nucleus.

Rutherford Atomic Model

Based on the above observations and conclusions, Rutherford proposed the atomic structure of elements. According to the Rutherford atomic model:

1. The positively charged particles and most of the mass of an atom was concentrated in an extremely small volume. He called this region of the atom as a nucleus.
2. Rutherford model proposed that the negatively charged electrons surround the nucleus of an atom. He also claimed that the electrons surrounding the nucleus revolve around it with very high speed in circular paths. He named these circular paths as orbits.
3. Electrons being negatively charged and nucleus being a densely concentrated mass of positively charged particles are held together by a strong electrostatic force of attraction.

Rutherford thought that the presence of electrons might explain another problem with the nuclear model: why does the nucleus not fly apart because of electrostatic repulsion of the positive particles? Perhaps the electrons act as a kind of glue which holds the nucleus together ? Only later did it become clear that, at the close separations involved in the nucleus, its component particles exhibit strong **nuclear forces** which completely overwhelm the electrostatic repulsion and which represent an entirely different kind of interaction.

- Probably the most serious problem with the planetary model is that an orbiting electron has a centripetal **acceleration** and (according to Maxwell's theory of electromagnetism) ought to lose energy by emitting **electromagnetic radiation** at a frequency **equal to that of the orbital motion** (the reciprocal of the orbital period). This radiated energy would be at the expense of the electrostatic *potential* energy of the electron, which would become more negative - implying that the electron approaches closer to the nucleus and experiences an increased electrostatic force. This increased force implies an increased centripetal acceleration and a higher angular velocity of the orbiting electron; the frequency of the emitted radiation would increase and the electron would spiral into the nucleus, as indicated in Fig. 3.20. Calculations showed that this process should happen in a small fraction of a second; in other words, the atom should not be stable ! The problem was not solved by Rutherford; it took the genius of Niels Bohr to propose a solution. • To start, Rutherford gave no explanation to support the stability of the atom. According to him, the electrons revolved at high speed around the nucleus in fixed paths. This was in contradiction with Maxwell's findings that said accelerated charged particles always released electromagnetic radiations when in motion. Therefore, the revolving electrons had to emit radiations when in their orbit.

- This would have caused the orbits to gradually shrink as a result of the outflow of electromagnetic energy when the electron is in motion. Finally, the electron would collapse into the nucleus. If calculated, the life of the atom would be 10^{-8}seconds resulting out of shrinking orbits. Since this was not the case, the model failed to explain atomic stability.

- Another major drawback was the non-explanation of electrons in their orbits around the nucleus. This made the theory virtually incomplete.

Additionally, Rutherford's atomic model failed to adequately explain the atomic emission spectra of hydrogen. Why did Hydrogen absorb some specific wavelengths of light and not others? These drawbacks were further built on by Neil's Bohr, James Chadwick and many scientists that came after him. Although these didn't sufficiently explain all observed phenomena, the revolutionary experiments, and observations that they conducted changed man's understanding of the atomic theory and formed the basis of today's modern physics which has seen many applications throughout an array of industries and utilities.

Observations of Rutherford's Alpha Scattering Experiment

The observations made by Rutherford led him to conclude that:

1. A major fraction of the α-particles bombarded towards the gold sheet passed through it without any deflection, and hence **most of the space in an atom is empty**.
2. Some of the α-particles were deflected by the gold sheet by very small angles, and hence the **positive charge** in an atom **is not uniformly distributed**. **The positive charge** in an atom **is concentrated in a very small volume**.
3. Very few of the α-particles were deflected back, that is only a few α-particles had nearly 180o angle of deflection. So the **volume occupied by the positively charged particles in an atom is very small as compared to the total volume of an atom**.

Limitations of Rutherford Atomic Model

Although the Rutherford atomic model was based on experimental observations it failed to explain certain things.

- Rutherford proposed that the electrons revolve around the nucleus in fixed paths called orbits. According to Maxwell, accelerated charged particles emit electromagnetic radiations and hence an electron revolving around the nucleus should emit electromagnetic radiation. This radiation would carry energy from the motion of the electron which would come at the cost of shrinking of orbits. Ultimately the electrons would collapse in the nucleus. Calculations have shown that as per Rutherford model an electron would collapse in the nucleus in less than 10-8 seconds. So Rutherford model was **not in accordance with Maxwell's theory and could not explain the stability of an atom**.

- One of the drawbacks of the Rutherford model was also that he **did not say anything about the arrangement of electrons in an atom** which made his theory incomplete.
- Although the early atomic models were inaccurate and failed to explain certain experimental results, **they were the base for future developments in the world of quantum mechanics**.

Conclusion

Rutherford tested Thomson's hypothesis by devising his "gold foil" experiment. Rutherford reasoned that if Thomson's model was correct then the mass of the atom was spread out throughout the atom. Then, if he shot high velocity alpha particles (helium nuclei) at an atom then there would be very little to deflect the alpha particles. He decided to test this with a thin film of gold atoms. As expected, most alpha particles went right through the gold foil but to his amazement a few alpha particles rebounded almost directly backwards.

These deflections were not consistent with Thomson's model. Rutherford was forced to discard the Plum Pudding model and reasoned that the only way the alpha particles could be deflected backwards was if most of the mass in an atom was concentrated in a nucleus. He thus developed the planetary model of the atom which put all the protons in the nucleus and the electrons orbited around the nucleus like planets around the sun.

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