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The Discovery of the Nucleus

Physics Group 1

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1 Introduction

Atom, the basic building blocks of all matter and the field of physical chemistry can combine with other atoms to form molecules but cannot be divided into smaller parts by ordinary chemical processes.

Most of the atom is empty space, letting trillions of neutrinos pass through it each second. The rest consists of three basic types of subatomic particles: protons, neutrons, and electrons. The protons and neutrons form the atom's central nucleus. (The ordinary hydrogen atom is an exception; it contains one proton but no neutrons.) As their names suggest, protons have a positive electrical charge, while neutrons are electrically neutral—they carry no charge; overall, then, the nucleus has a positive charge. Circling the nucleus is a cloud of electrons, which are negatively charged. Like opposite ends of a magnet that attract one another, the negative electrons are attracted to a positive force, which binds them to the nucleus. The nucleus is small and dense compared with the electrons, which are the lightest charged particles in nature. The electrons circle the nucleus in orbital paths called shells, each of which holds only a certain number of electrons.

2 Discovery of the Atom

The first truly direct evidence of atoms is credited to Robert Brown, a Scottish botanist. In 1827, he noticed that tiny pollen grains suspended in still water moved about in complex paths. This can be observed with a microscope for any small particles in a fluid. The motion is caused by the random thermal motions of fluid molecules colliding with particles in the fluid, and it is now called Brownian motion. Statistical fluctuations in the numbers of molecules striking the sides of a visible particle cause it to move first this way, then that. Although the molecules cannot be directly observed, their effects on the particle can be. By examining Brownian motion, the size of molecules can be calculated. The smaller and more numerous they are, the smaller the fluctuations in the numbers striking different sides.

It was Albert Einstein who, starting in his epochal year of 1905, published several papers that explained precisely how Brownian motion could be used to measure the size of atoms and molecules. (In 1905 Einstein created special relativity, proposed photons as quanta of EM radiation, and produced a theory of Brownian motion that allowed the size of atoms to be determined. All of this was done in his spare time, since he worked days as a patent examiner. Any one of these very basic works could have been the crowning achievement of an entire career—yet Einstein did even more in later years.) Their sizes were only approximately known to be 10^{-10} m, based on a comparison of latent heat of vaporization and surface tension made in about 1805 by Thomas Young of double-slit fame and the famous astronomer and mathematician Simon Laplace.

Using Einstein's ideas, the French physicist Jean-Baptiste Perrin (1870-1942) carefully observed Brownian motion; not only did he confirm Einstein's theory, he also produced accurate sizes for atoms and molecules. Since molecular weights and densities of materials were well established, knowing atomic and molecular sizes allowed a precise value for Avogadro's number to be obtained. (If we know how big an atom is, we know how many fit into a certain volume.) Perrin also used these ideas to explain atomic and molecular agitation effects in sedimentation, and he received the 1926 Nobel Prize for his achievements. Most scientists were already convinced of the existence of atoms, but the accurate observation and analysis of Brownian motion was conclusive—it was the first truly direct evidence.

The basic idea that matter is made up of tiny indivisible particles is an old idea that appeared in many ancient cultures. The word atom is derived from the ancient Greek word atomos, which means "uncuttable". This ancient idea was based in philosophical reasoning rather than scientific reasoning. Modern atomic theory is not based on these old concepts.

In the early 1800s, the English chemist John Dalton compiled experimental data gathered by him and other scientists and discovered a pattern now known as the "law of multiple proportions". He noticed that in any group of chemical compounds which all contain two particular chemical elements, the amount of Element A per measure of Element B will differ across these compounds by ratios of small whole numbers. This pattern suggested that the elements combine with each other in multiples of a basic unit of weight, with each element having a unit of unique weight. Dalton decided to call these units "atoms".

3 Isotopes

isotopes are members of a family of an element that all have the same number of protons but different numbers of neutrons. You can see the different chemical elements on the periodic table.

Each element is distinguished by the number of protons, neutrons and electrons that it possesses. The atoms of each chemical element have a defining and same number of protons and electrons, but—crucially—not neutrons, whose numbers can vary.

atoms with the same number of protons but different numbers of neutrons are called isotopes. They share almost the same chemical properties, but differ in mass and therefore in physical properties. There are stable isotopes, which do not emit radiation, and there are unstable isotopes, which do emit radiation. The latter are also called radioisotopes.

The fact that all protons and all neutrons are identical means that the only difference between the nucleus of one element and the nucleus of another is the number of each of these particles. protons and neutrons are thus sometimes collectively referred to as nucleons.

4 Discovery of Protons

In 1886, it was observed that the charge-to-mass ratio of the hydrogen ion was the highest among all gases. This observation hinted that hydrogen ions (or protons) are fundamental particles constituting all atoms.

A proton is a positively charged particle that is present in the nucleus of an atom. The mass of a proton is 1.676×10^{-24} grams (or 1.676×10^{-27} kilograms). The charge of a proton is equal but opposite to that of an electron (it possesses a positive charge of 1.602×10^{-19} coulombs). The total number of protons in the atoms of an element is always equal to the atomic number of the element.

5 Discovery of Neutrons

The discovery of neutrons was the central theme of James Chadwick's research in the year 1932. In this experiment, he made use of alpha particles to discover a neutral atomic particle with a mass close to that of a proton. This particle was eventually termed a neutron.

The experiments that led to the discovery of the neutron began by tracking the radiation emitted from beryllium when it was bombarded with alpha particles (which are doubly charged particles with a mass equal to that of helium nuclei). It was believed that this radiation was made up of high-energy gamma photons.

The results of the experiment showed that when this radiation fell on paraffin wax (or any other hydrogen-containing compound), protons were ejected from the surface of the compound. This prompted Chadwick to propose that the radiation was composed of particles with a mass approximately equal to that of the proton but with no charge. These particles were later termed neutrons.

The discovery of neutrons made it possible to penetrate and split the nuclei of heavy atoms. Therefore, neutron bombardment was made use of in several nuclear reactions, such as nuclear fission and nuclear fusion.

6 Discovery of Electrons

In 1897, British physicist J.J. Thomson was experimenting with a cathode ray tube (a vacuum tube containing two electrodes connected to a high-voltage source). When he turned on the electrical current, he saw a glowing beam travel from the cathode (negatively charged) to the anode (positively charged). He concluded that the beam was made of particles, which he called "corpuscles". Today, we know these particles as electrons.

Thomson's experiments showed that electrons were much smaller than atoms and that they carried a negative charge. His discovery proved that atoms were not indivisible, as John Dalton had thought, but were instead made of smaller particles. This was a groundbreaking discovery that led to the development of the modern atomic model.

7 The Rutherford Gold Foil Experiment

In 1909, physicist Ernest Rutherford and his colleagues conducted the famous gold foil experiment. They aimed alpha particles (helium nuclei) at a thin sheet of gold foil and observed the scattering pattern using a fluorescent screen. Most of the alpha particles passed straight through the foil, but a small fraction was deflected at large angles.

Rutherford concluded that atoms must have a small, dense, positively charged nucleus at their center, where most of the mass is concentrated. This discovery led to the Rutherford model of the atom, where electrons orbit a central nucleus, much like planets orbit the sun.

8 The Nuclear Model of the Atom

The discovery of the nucleus led to the development of the nuclear model of the atom. In this model, the atom consists of a dense central nucleus containing protons and neutrons, surrounded by a cloud of electrons. The protons give the nucleus a positive charge, while the neutrons contribute to the mass but not the charge.

This model explained many of the properties of atoms, including their stability and the behavior of electrons in chemical reactions. It also paved the way for the development of quantum mechanics, which provided a deeper understanding of the behavior of electrons and other subatomic particles.

9 Nuclear Stability

Nuclear stability refers to the stability of a nucleus of an atom. A stable nucleus does not decay spontaneously. Radioactive elements contain unstable nuclei and decay spontaneously emitting various radiations.

Nuclei of atoms contain protons and neutrons. Positively charged protons repel each other due to electrostatic repulsion between them. This electrostatic repulsion is overcome by the strong nuclear force, the attractive force present between nucleons. Neutrons are important for stabilizing the nucleus. If the attractive force between nucleons is less than the electrostatic repulsion then it makes the nucleus unstable and results in decay.

It defines the stability of an isotope of the elements. Nucleons with high binding energy are more stable. Stability of an isotope can be determined by calculating the ratio of neutrons to protons present in a nucleus (N/Z). Elements having atomic number less than 20, mostly have proton and neutron ratio 1:1. The number of neutrons increases as the atomic number increases. Most of the stable nuclei have neutrons to protons ratio more than 1. Only 1H and3He have neutrons to protons ratio less than one but are stable.

The first 80 elements of the periodic table have stable isotopes. All the elements with the atomic number more than 82 are unstable and radioactive, irrespective of the number of neutrons.

Glossary

atom The basic building block of matter.

electron A negatively charged subatomic particle that orbits the nucleus of an atom.

isotope Variants of a particular chemical element that have the same number of protons but different numbers of neutrons.

neutron A neutrally charged subatomic particle found in the nucleus of an atom.

proton A positively charged subatomic particle found in the nucleus of an atom.