Niels Bohr and the Atomic Structure

M Durga Prasad

Niels Bohr developed his model of the atomic structure in 1913 that succeeded in explaining the spectral features of hydrogen atom. In the process, he incorporated some non-classical features such as discrete energy levels for the bound electrons, and quantization of their angular momenta.

M Durga Prasad is a Professor of Chemistry at the University of Hyderabad. His research interests lie in theoretical chemistry.

Introduction

An atom consists of two interacting subsystems. The first subsystem is the atomic nucleus, consisting of Z protons and $A - Z$ neutrons, where Z and A are the atomic number and atomic weight respectively, of the concerned atom. The electrons that are part of the second subsystem occupy (to zeroth order approximation) *discrete* energy levels, called the orbitals, according to the Aufbau principle with the restriction that no more than two electrons occupy a given orbital. Such paired electrons must have their spins in opposite direction (Pauli's exclusion principle). The two subsystems interact through Coulomb attraction that binds the electrons to the nucleus. The nucleons in turn are trapped in the nucleus due to strong interactions. The two forces operate on significantly different scales of length and energy. Consequently, there is a clear-cut demarcation between the nuclear processes such as radioactivity or fission, and atomic processes involving the electrons such as optical spectroscopy or chemical reactivity. This picture of the atom was developed in the period between 1890 and 1928 (see the timeline in *Box* 1). One of the major figures in this development was Niels Bohr, who introduced most of the terminology that is still in use. We look at his contributions in this article.

Bohr was born on 7th October, 1885 in Copenhagen. His father, Christien Bohr, was a professor of physiology in the Copenhagen University. Niels received his education in Copenhagen, obtaining his masters degree in 1909, and the doctoral degree in 1911.

Keywords
Atomic structure, Bohr's model.

Box 1. Timeline of Events in Atomic Physics

Hydrogen atom spectral lines	1890	J Rydberg
Discovery of electron	1890–97	A Shuster, J J Thomson
Zeeman effect	1897	P Zeeman
α-particle scattering: discovery of nucleus	1909–11	G Geiger, E Marsden, E Rutherford
Radioactive displacement laws and isotopes	1910–12	F Soddy
Bohr's models	**1913**	**N Bohr**
Stark effect	1913	J Stark
Nuclear charge/Atomic number	1913	H Mosley
Inelastic scattering of electrons from mercury atoms	1914	J Franck, G Hertz
Interpretation of fine structure	1915	A Sommerfeld
Wave–particle duality	1923	L de Broglie
Exclusion principle and the periodic table	1925	W Pauli
Quantum mechanics	1925–26	W Heisenberg, E Schrödinger, P A M Dirac
Uncertainty relations	1927	W Heisenberg
Independent particle model	1928	D R Hartree
Discovery of neutron	1932	J Chadwick

After finishing his doctorate, Bohr went to Cavendish Laboratory in Cambridge on a one year fellowship to work with J J Thomson. For one reason or the other, the two did not gel with each other. Bohr moved to Manchester to work with Ernst Rutherford in the second half of his one year tenure. He was introduced to radio activity and atomic structure in Manchester.

Manchester

The first problem that attracted the attention of Bohr in Manchester was the formulation of, what we call today as, the radioactive displacement laws. Early in his sojourn, Bohr learned that several new 'radioactive elements' were discovered over the past few years. The primary question that surfaced as a consequence was regarding their location in the periodic table. Moreover, some of these newly discovered elements were chemically identical to the older elements, making it impossible to separate them by chemical methods. For example, Uranium-X produced from Uranium by the loss of an α particle was indistinguishable from Thorium. The only difference between them is their atomic weight. F Soddy suggested that all such species should be placed in the same box of the periodic table. Later on, such atoms came to be known as the istopes.

Rutherford's planetary model of the atom posited that all the positive charge in the atom is concentrated in the nucleus, and the electrons revolving around it provided the counterbalancing negative charge, such that the atom as a whole is electro-neutral. If so, Bohr reasoned, loss of an α particle must be occurring from the nucleus irrespective of the electronic environment around it, and it reduces the charge of the nucleus by two units of (electronic) charge. Since Thorium is two places before Uranium in the periodic table, it stands to reason that the position of the element in the periodic table is determined by the nuclear charge. The α particle has a mass of four atomic units. So Uranium-X has an atomic weight of 234. The atomic weight of Thorium is 232. Since the two nuclei have the same charge, they are chemically indistinguishable, but have different atomic weights. A similar argument can be made for the β decay processes.

The identification of the nuclear charge with the position of the atom in the periodic table (what we call as the atomic number today) was quite bold. The only known heavy particle at that time was the proton. If the charge on the nucleus, and hence the number of protons is identified with the atomic number, then the rest of the atomic mass would require an explanation. This mass-related question was fully answered only after the discovery of neutron by J Chadwick in 1932. Nevertheless, convinced of its validity, Bohr took his theory to Rutherford. Unfortunately, Rutherford did not take the model seriously, discouraging Bohr from pursuing his ideas further. It was rediscovered by Soddy a year later, and fetched him the Nobel Prize in 1921.

Two aspects of Bohr's analysis deserve our attention. First, Bohr could see the patterns in a large body of uncorrelated data, and thus could get to the heart of the problem. Second, he appreciated

Two aspects of Bohr's analysis deserve our attention. First, Bohr could see the patterns in a large body of uncorrelated data, and thus could get to the heart of the problem. Second, he appreciated the distinction between atomic processes and the nuclear processes.

the distinction between atomic processes (involving electrons that filled the extra-nuclear region of the atom) and the nuclear processes (that occur within the nucleus).

The second problem that attracted Bohr's attention was the development of a proper mathematical description of the energy loss of α particles as they passed through matter. It was studied earlier by Thomson using his plum pudding model, and more recently by C G Darwin (a grandson of Charles Darwin) on the basis of the planetary model. The calculation needs the electronic charge density in the atom since α particles lose energy mostly from their collisions with the atomic electrons. Darwin assumed that the electrons were free, instead of being bound to the positively charged nucleus. Bohr took up the problem to correct this deficiency. While working on this, presumably, Bohr's attention was drawn to the inadequacies of the then existing atomic models, and he set out to correct them, leading to his famous model.

Copenhagen

Bohr returned to Copenhagen in the middle of 1912, got married, found a job, and, by early September, he was back to his attempts to find a solution to the problem of atomic structure. The planetary model of Rutherford had two serious problems. The model assumed that the electrons were prevented from collapsing into the nucleus by the centrifugal force much like the planets are prevented from crashing into the sun. Such orbiting electrons are continuously accelerated towards the nucleus due to the centripetal (Coulomb attraction to the nucleus) force. Otherwise, they would fly off at a tangent to the orbit according to Newton's first law. Now, according to classical electrodynamics, accelerating charged particles must emit electromagnetic radiation and loose their energy. As this happens, the electronic orbit would continuously shrink, and the electron would eventually collapse into the nucleus. This raises doubts about the stability of matter. The second problem with the planetary model is that, the orbiting electrons can have arbitrary amounts of energy. The radius of the orbit (that determines the potential energy) and the speed with which they travel along the orbit (that determines the kinetic energy) have to be consistent with each other for the orbit to be stable. Given two variables and one constraint (that centrifugal and centripetal forces should sum up to zero), one variable is free to take any value. Consequently, the energy of the electron can be anything. Since the energy has a continuous range, the range of frequencies that can be observed in the emission or absorption spectrum *should* also be continuous. Experimentally, however, one finds discrete spectral lines.

Instead of condemning the planetary model on the basis of these limitations, Bohr chose to accept that the well-known tenets of classical physics were inadequate to describe the atomic structure.

Instead of condemning the planetary model on the basis of these limitations, Bohr chose to accept that the well-known tenets of classical physics were inadequate to describe the atomic structure. Probably he was inspired by the successes of Planck and Einstein, who, in the previous decade or so, gave up classical physics to explain the experimental features of black-body radiation and photo-electric effect quite successfully. Contrary to the classical wisdom that the energy of radiation field (proportional to the square of the amplitude) can be any positive number, they had demonstrated that by assuming that the energy is transferred in discrete quanta, experimental results could be reproduced correctly. Bohr demonstrated the strength of his instincts by following a similar route in which he gave up some of the features of classical physics.

His first target was the stability of the electronic orbits. Instead of finding the conditions for the stability, Bohr *postulated* that there exist a discrete set of orbits, 'stationary states' as he called them. The stationary states were defined by the requirement that the electrons moving along them did not radiate, and consequently, did not collapse into the nucleus.

Postulating that electrons in stationary states do not emit radiation is one thing, but one needs a mathematical prescription to identify them. Such a prescription should also explain the discrete nature of the spectral lines. Around this time, Bohr became

Box 2. Bohr's Postulates

1. An electron in an atom moves in a circular orbit under the influence of the nuclear Coulomb attraction.
2. The only allowed orbits, called stationary states, are those for which the angular momentum is

$$L = nh/2\pi.$$

3. Electrons in the stationary states do not emit radiation, notwithstanding the fact that they are continuously accelerating.
4. Electromagnetic radiation is emitted or absorbed when an electron jumps discontinuously from one stationary state to another such that

$$h\nu = E_f - E_i.$$

aware of the Rydberg formula for spectral lines of hydrogen atom,

$$\nu = R_H \left(\frac{1}{m^2} - \frac{1}{n^2} \right). \qquad (1)$$

He also came across the work of J W Nicholson who was also developing a model of atomic structure. An important aspect of Nicholson's work was the observation that the Planck's constant, h, has the units of angular momentum. In view of this Nicholson had argued that angular momentum should be quantized as an integral multiple of $(h/2\pi)$.

Armed with these two pieces of the puzzle, Bohr came up with his model in the form of four postulates (*Box* 2). Using these, he derived a formula for the Rydberg constant (R_H) in terms of the fundamental constants. The calculated R_H is quite close to the experimental value.

Notwithstanding the success of Bohr's model in predicting the hydrogen atom spectrum, his mentor Rutherford had two serious reservations about it. According to classical electrodynamics, an oscillating charge emits radiation with the same frequency at which it is oscillating. An electron, transiting from one stationary state to another, has two frequencies associated with the two stationary states. Bohr's model on the other hand, postulates that the electron emits radiation corresponding to their difference. A second objection was that when an electron emits radiation in a downward transition, it seems to be aware of its destination *before* it makes the transition so that it knows the frequency of the radiation to be emitted. However, putting aside his reservations, Rutherford communicated the three papers of Bohr to the *Philosophical Magazine* with the stamp of his approval. They were published between July and November 1913.

Additional experimental support came up over the next year. Bohr interpreted the Pickering–Fowler lines in the solar spectrum as the spectral lines of singly ionised He^+ ions based on his model. This was contrary to the contemporary belief that they were part of the hydrogen atom spectrum. Careful experiments on the spectral lines of He^+ in the laboratory settled the matter in favour of Bohr's analysis. A second experiment that lent support to Bohr's theory was the data of J Franck and G Hertz who studied the inelastic scattering of electrons off mercury atoms. This experiment showed that the energy levels of mercury atom were discrete. With the support from these two additional experiments, acceptance of Bohr's theory was complete.

Bohr's model was extended by A Sommerfeld to stationary states with elliptic orbits to interpret the fine structure that was seen when spectrometers with better resolution became available.

Bohr's interest in atomic physics ended more or less at this point. Later experiments challenged his model far more seriously than the experimental evidence that was available when he developed his model. The model was extended by A Sommerfeld to stationary states with elliptic orbits to interpret the fine structure that was seen when spectrometers with better resolution became available. This led to the identification of a new *orbital* quantum number, k, in addition to the original *principal* quantum number, n. In the process, he also replaced the angular momentum quantization condition of Bohr with a quantization condition on the phase integral, following an earlier work of W Wilson. The Zeeman Effect, splitting of spectral lines in the presence of external magnetic fields, could not be explained by Bohr's theory either. Sommerfeld introduced a third, *magnetic* quantum number, m to indicate the spatial orientation of the plane in which electrons are orbiting with respect to the unique axis – the direction of the magnetic field – to explain the Zeeman effect. The discovery of the atomic number, and the Aufbau principle helped in providing a rationale to the periodic table. De Broglie's hypothesis on the wave–particle duality provided a justification to the quantization of angular momentum. However, the spectral features of many-electron atoms were not amenable to treatment by Bohr's model. Similarly, the spectral intensities of various transitions, where some lines were bright, and others were weak could not be explained by the model. Under these circumstances, the writing on the wall was clear that Bohr's model cannot go on much further. Quantum mechanics which came up a decade later provided the answers to these questions, as well as the correct interpretation of the quantum numbers that Bohr and Sommerfeld defined.

Legacy

Bohr's theory came at a time when classical mechanics was at the crossroads. New experiments such as the black-body radiation theory needed new concepts and mathematical tools, but classical mechanics was unable to provide them. For a while, all that one could do was to develop a phenomenological description of the physical system rather than attempt a full-fledged microscopic theory. Planck and Einstein took this path when they quantized the radiation field without knowing why they were doing so. Bohr took a similar road by postulating the existence of stationary states of an electron in an atom and quantizing them. The most enduring feature of his work is, perhaps, the concept of stationary state – a state from which an electron will not move out unless it is otherwise disturbed. Quantum mechanics defines the stationary states as those states whose probability distribution functions do not change with time. The existence of shells and sub-shells continues into quantum mechanics, though their origin is for different reasons. The electrons stay out of the nucleus, not due to the centrifugal forces, but because of the uncertainty-relati on induced excess kinetic energy. All these came up later. Bohr's theory prepared the ground for these developments to be accepted more easily.

Address for Correspondence
M Durga Prasad
School of Chemistry
University of Hyderabad
Hyderabad 500 046, India.
Emai: mdpsc@uohyd.ac.in

Suggested Reading

[1] Manjit Kumar, *Quantum*, Hachette India, ISBN:978-93-80143-10-1, 2009.