

Creation and development of Bohr's theory (on the 90th anniversary of the Bohr theory of the atom)

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Abstract. The history of the creation and development of Bohr's atomic theory is discussed. Even now, with a consistent quantum theory available, Bohr's theory is not simply the property of history, of methodological interest only. To this day, the ideas of the theory not only provide an excellent introduction to atomic physics, but are also used successfully in treating atomic Rydberg states, exotic atoms, etc.

1. Introduction

The great discoveries of the 19th century — spectral patterns, the electron, X-ray radiation, radioactivity, etc. — led to the understanding that the atom as the smallest indivisible particle of matter does not actually exist in nature; it became clear that the atom must have a structure. Among atomic models that were around in the early 20th century was the drop — or pudding — model proposed by the discoverer of the electron Joseph John Thomson (1856–1940). In this model, the atom is ‘a sphere of uniform positive electrification’ within which, like raisins in a pudding, a certain number of electrons are embedded to neutralize the atom's positive charge. However, the 1909–1912 experiments performed under the guidance of Ernest Rutherford (1871–1937) showed convincingly that the positive charge of an atom is concentrated in a very small region (known as the nucleus) and led Rutherford to the planetary model of the atom, with electrons circling the nucleus similar to planets orbiting the sun [1]. As far as classical physics is concerned, however, such an atom cannot exist in nature: an electron circling the

nucleus must lose its energy due to radiation according to electrodynamics. The energy of the orbiting electron is related to its orbit's radius, so as the energy decreases due to the radiation, also the orbital radius decreases, and since radiation is a continuous process, the electron will eventually collapse into the nucleus. Such a collapse, G A Schott calculated back in 1904, must occur virtually instantaneously, in a time of the order of 10^{-11} s. Besides, the closer the electron comes to the nucleus, the faster it should rotate — continuously increasing the frequency of the light it emits. What is actually observed is entirely different, however: the light emitted by an atom is characterized by a discrete set of wavelengths. To quote a witty remark P Ehrenfest (1880–1933) made on this, “The question is, why does an atom emit a pure tone rather than making a noise similar to a cat's miaow?”

Thus, Rutherford's experiments led to an impasse in that either electrodynamics or the planetary model should be abandoned. As to electrodynamics, most physicists of the time believed it should not be abandoned — given numerous confirming experiments and the practical applications it had already been put to. Besides, the characteristic quantities of the planetary atom — the charges and masses of the electron and nucleus — could not be combined into a quantity of dimensions of length to thus obtain an estimate for the atom's size. For these reasons, up to 1913 physicists tried not to notice this model at all. Rutherford was fully aware of the challenges the planetary model faced. In his words of the time, “The question of the stability of the atom proposed need not be considered at this stage, for this will obviously depend on the minute structure of the atom and on the motion of the constituent charged parts” [1].

It was the great Danish physicist Niels Henrik David Bohr (1885–1962) who found a way out [2].

Niels Bohr was born in Copenhagen into a well-educated family of a physiology professor. In the fall of 1911, with a doctorate in the electronic theory of metals earned at the University of Copenhagen, Bohr went to do postdoctoral research at the famous Cavendish Laboratory in Cambridge,

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where J J Thomson had discovered the electron in 1897. In 1906 J J Thomson had received the Nobel Prize ‘in recognition of the great merits of his investigations on the conduction of electricity by gases’. By the time Bohr came to the laboratory, Thomson was preoccupied with how his atomic model could be perfected and possibly put to use and so was not interested in the slightest in what Bohr was doing. Soon Bohr met Ernest Rutherford, J J Thomson’s most prominent disciple, and asked him for a postdoctoral research position. In the spring of 1912 Bohr moved to Manchester, where Rutherford had a chair and directed a laboratory at Victoria University. That was how fate brought them together, apparently picking Bohr to be the savior of Rutherford’s atomic model. It was for Bohr to become a central figure in the development of the quantum theory of atomic structure and the leading ideologist of the ‘Copenhagen’ interpretation of the quantum mechanics created in 1925–1927. Bohr is also well known for his work in nuclear physics.

2. Why were quantum postulates necessary?

Although Bohr immediately became an adherent of the planetary model, he ‘never took [it] literally’ as he confessed years later: it was only an image, not a true picture of the atom for him. Because atoms do exist in nature and because they are ordinarily stable, what turned out to be at issue for the planetary atom was the stability of electrons in their orbits around nuclei — which in turn was related to the question of the stability of atomic sizes. The existence of atoms suggests that there exist in nature such minimum distances from the nucleus closer than which electrons cannot be located. It was when pondering this point that, in a stroke of genius, Bohr realized that a fundamental relation must exist between this minimum distance and the minimum quantum of action. As he remembered afterwards, “In my early time in Manchester, in the spring of 1912, I came to believe that the electron swarm in the Rutherford atom is governed by the quantum of action.” But then the planetary model implied abandoning classical physics. As Bohr said later, “The crucial point about Rutherford’s atomic model was that it most clearly demonstrated that the stability of atoms cannot be accounted for based on classical physics and that the quantum postulate is the only possible way out of the acute dilemma at hand. It is this acute controversy which made me absolutely believe in the correctness of the quantum postulate.”

At the end of 1912, his postdoctoral term over, Bohr presented Rutherford a ‘memo’ which for the first time introduced the idea of stable orbits and in which he suggested that a link must exist between electronic orbits and the periodic system of elements. That was Bohr’s grandiose program, not only to explain the stability of the planetary model but as well to account for “those properties of matter dependent on the system of electrons in the atom.” While Rutherford in his reply advised Bohr not to hasten conclusions, it soon became clear that hasten he certainly should. It turned out that by that time it had been postulated by an astrophysicist J W Nicholson of Cambridge in a number of papers that, for an electron in an atom, the projection of the angular momentum L is quantized: $L = nh/2\pi$, where n is an integer and h , Planck’s constant. Based on this postulate, Nicholson found discrete atomic orbits in each of which, he believed, groups of electrons rotated. According to classical ideas, electrons should emit electromagnetic waves with a frequency equal to the frequency of their rotation around the

nucleus. That was exactly what Nicholson assumed them to do. With this assumption — which was later found to be approximately true for highly excited atomic states — Nicholson was able to explain many features observed in the radiation from the solar corona and from nebulae [3].

It was possibly his chance meeting in early February of 1913 with H Hansen, a spectroscopist and his former university friend, which was the eureka event for Bohr [4]. Bohr explained to Hansen how he believed the structure of a substance might be derived from the planetary model with (as yet) unexplainably stable orbits, and one question Hansen asked in the course of their conversation was how the Balmer–Rydberg–Ritz spectral laws were explained by Bohr’s theory. It turned out that Bohr was fully ignorant of the laws! To quote Bohr’s recollections, “As soon as I saw Balmer’s formula, the whole thing was immediately clear to me.” That was the flash of scientific insight after which, in the short time of less than a month, Bohr prepared the first, most important part of his paper “On the Constitution of Atoms and Molecules,” which he published in *Philosophical Magazine* [Series 6, Vol. 26, pp. 1–15 (1913)] [5].

Before Bohr, the generalized Balmer formula (J Balmer, 1825–1898) and its associated combination principle of Ritz (W Ritz, 1878–1909) had been considered by many as just a funny game of numbers. Bohr was the first to uncover the deep physical meaning of this formula, to see how the quanta of radiation are given birth to by an atom. He realized that, quite contrary to classical views, the frequency of a spectral line is not related to the frequency at which electrons orbit the nucleus, but rather to the energy of the radiation that is emitted when the atom makes a transition from one of its discrete state to another, the generalized Balmer formula playing the role of the energy conservation law in the atom-radiation system.

Bohr formulated as follows the basic assumptions of his theory [5]:

“(1) [That] the dynamical equilibrium of the systems in the stationary states can be discussed by help of the ordinary mechanics, while the passing of the systems between different stationary states cannot be treated on that basis.

(2) [That] the latter process is followed by the emission of a homogeneous radiation, for which the relation between the frequency and the amount of energy emitted is the one given by Planck’s theory.”

Bohr then goes on to explain his assumptions: “The first assumption seems to present itself; for it is known that the ordinary mechanics cannot have an absolute validity, but will only hold in calculations of certain mean values of the motion of the electrons. On the other hand, in the calculations of the dynamical equilibrium in a stationary state in which there is no relative displacement of the particles, we need not distinguish between the actual motions and their mean values.” Stationary states, Bohr points out, show the peculiar property of stability, which manifests itself in that every long-duration change in the motion of a closed system is a transition of this system between different stationary states.

Bohr stressed that in accepting Planck’s theory one openly recognizes the insufficiency of ordinary classical electrodynamics and resolutely breaks with the closely linked chain of its propositions. Bohr was fully aware that “the second assumption is in obvious contrast to the ordinary ideas of electrodynamics but appears to be necessary in order to account for experimental facts.” Max von Laue (1879–1960) indignantly opposed this at the time: “But this is



Niels Bohr

rubbish! Maxwell's equations are valid under all circumstances, an electron in an orbit must radiate." [3]

In concluding his work Bohr summarized the assumptions he made. These are

"1. That energy radiation is not emitted (or absorbed) in the continuous way assumed in the ordinary electrodynamics, but only during the passing of the system between different 'stationary' states.

2. That the dynamical equilibrium of the systems in the stationary states is governed by the ordinary laws of mechanics, while these laws do not hold for the passing of the systems between the different stationary states.

3. That the radiation emitted during the transition of the system between two stationary states is homogeneous, and that the relation between the frequency ν and the total amount of energy emitted E is given by $E = h\nu$, where h is Planck's constant.

4. That the different stationary states of a simple system consisting of an electron rotating round a positive nucleus are determined by the condition that the ratio between the total energy, emitted during the formation of configuration, and the frequency of revolution of the electron is an entire multiple of $h/2$. Assuming that the orbit of the electron is circular, this assumption is equivalent with the assumption that the angular momentum of the electron round the nucleus is equal to an entire multiple of $h/2\pi$.

5. That the 'permanent' state of any atomic system — i.e., the state in which the energy emitted is maximum — is determined by the condition that the angular momentum of every electron round the center of its orbit is $h/2\pi$."

Bohr's assumptions for the theory of atoms were later formulated in the form of the following 'quantum postulates':

1. There exist *stationary* (time-independent) *states* in which atoms neither emit nor absorb energy. These states are characterized by a discrete set of energy values E_1, E_2, E_3, \dots

2. The emission and absorption of light occur during a jump transition of the atom from one of its discrete states to another, the energy of the emitted (absorbed) atoms being determined by the equation $h\nu = E_n - E_m$.

The first postulate actually amounts to recognizing the existence of atoms in nature. But then it must be accepted, according to Bohr, that atoms can only exist in states with discrete values of energy and that electrons in them circle the nucleus at specific discrete distances. When at the minimum distance from the nucleus, an electron has nowhere else to jump to and so should stay in this state for an infinitely long time. Hence the stability of the atom is explained.

The second postulate stemmed from Planck's quantum hypothesis and was a generalization of the experimental results that were available on the spectral lines of emitting atoms. "Today," Einstein [6] wrote in 1916 about the postulate of quantum transitions, "this is arguably one of the most reliably laid foundation blocks of our science."

Bohr's postulates are drastically at odds with the notions of classical physics, but it was the abandoning of these notions and the idea of quanta he introduced into the world of the atom that led Bohr to the first quantum theory of the atom.

3. Experimental proof of Bohr's postulates

The existence of discrete atomic states was directly proved in experiments by James Frank (1882–1964) and Gustav Hertz (1887–1975) in 1914. Still unaware of Bohr's theory, the task Frank and Hertz originally set themselves was to measure the ionization potentials of some atoms — of mercury, in particular. Frank confessed years later: "Because the physicists of the time were dominantly openly distrustful of attempts to build a model of the atom with the then available knowledge level, few took the trouble of carefully reading a paper on the atom. A special point to note is that Gustav Hertz and I first failed to realize the huge significance of Bohr's work." The discussion of Frank and Hertz's results (Fig. 1) led to the understanding that what they actually

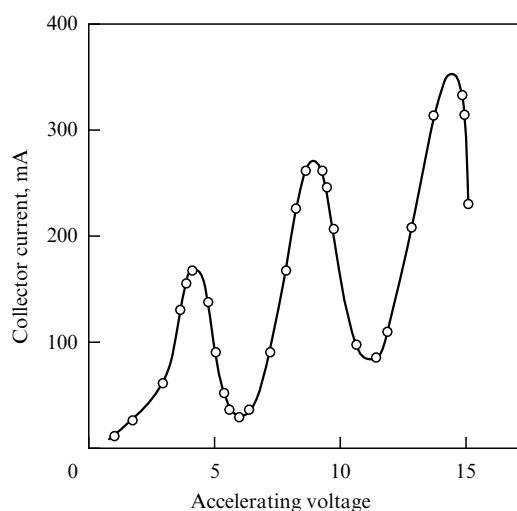


Figure 1. Results of the Frank – Hertz experiment.

measured in their experiments was not the ionization potential — but the excitation potential — of mercury atoms.

Frank and Hertz's experiments were received with satisfaction by Bohr. "The results of these experiments have shown," Bohr wrote, "that in collisions with an electron an atom cannot acquire an arbitrary amount of energy — but only such an amount of it which exactly corresponds to the energy necessary to move the atom from the normal state to one of the remaining stationary states; the existence of these latter is known to us from data on spectra, because the energies of these states are closely related to the values of the spectral terms."

4. Bohr's atom model

The first attempt at a quantum theory of the hydrogen atom was undertaken in 1910 by a young Austrian physicist A Haas (1884–1941), who calculated the Rydberg constant based on quantum ideas and Thomson's atomic model. His Rydberg constant, however, differed many fold from the experimental one. At the time, Haas's ideas became an object of derision, though — as a 'naive attempt' to combine such 'incompatible things' as spectroscopy and the quantum theory of radiation.

For Bohr, the starting point was the planetary model of hydrogen and hydrogen-like atoms [5]. Accordingly, Bohr considered an electron moving in a closed orbit around the nucleus at a speed much lower than that of light. "The frequency of revolution ν and the major-axis of the orbit $2a$ will depend on the amount of energy W which must be transferred to the system in order to remove the electron to an infinitely great distance apart from the nucleus." These quantities are given by the formulas (in modern notation) [7]:

$$\nu = \frac{\sqrt{2}}{\pi} \frac{W^{3/2}}{Ze^2\sqrt{m}}, \quad 2a = \frac{Ze^2}{W}. \quad (1)$$

Bohr then assumed that the orbit in question was a circular one and employed Planck's idea that the amount of energy emitted in each emission event is $h\nu$. Bohr's further reasoning surely shows him to be a man of genius and demonstrates his amazing physical intuition. He assumes that the electron is originally far from the nucleus, that it has no noticeable velocity relative to the nucleus, and that after the encounter with the nucleus it finds itself in a stationary orbit around the nucleus. "Let us now assume that ... the electron [emits] homogeneous radiation of a frequency ν , equal to **half the frequency of revolution of the electron in its final orbit.**" It is this assumption — or in fact quantum postulate — which leads to a correct result! Thus, setting

$$W = \frac{nh\nu}{2}, \quad (2)$$

and using formula (1) Bohr obtains the expressions

$$W = \frac{2\pi^2 m Z^2 e^4}{n^2 h^2}, \quad \nu = \frac{4\pi^2 m Z^2 e^4}{n^3 h^3}, \quad 2a = Z \frac{n^2 h^2}{2\pi^2 m e^2}. \quad (3)$$

"If in these expressions we give n different values we get a series of values for W , ν , and a corresponding to a series of configurations of the system. According to the above considerations, we are led to assume that these configurations will correspond to states of the system in which there is no radiation of energy; states which consequently will be

stationary as long as the system is not disturbed from outside. We see that the value of W is greatest if n has its smallest value 1. This case will therefore correspond to the most stable state of the system, i.e. will correspond to the binding of the electron for the breaking up of which the greatest amount of energy is required."

Using the values of the physical constants as known at the time, Bohr obtained the estimates

$$2a = 1.1 \times 10^{-8} \text{ cm}, \quad \nu = 6.2 \times 10^{15} \text{ s}^{-1}, \quad \frac{W}{e} = 13 \text{ V}.$$

From this "we see that these values are of the same order of magnitude as the linear dimensions of the atoms, the optical frequencies, and the ionization potentials."

According to formulas (3), the energy emitted during the formation of one of the stationary states of the hydrogen atom is equal to $W = 2\pi^2 m e^4 / (n^2 h^2)$. Then "the amount of energy emitted during the passing of the system from a state corresponding to $n = n_1$ to one for which $n = n_2$ " is equal to

$$W_{n_2} - W_{n_1} = \frac{2\pi^2 m e^4}{h^2} \left(\frac{1}{n_2^2} - \frac{1}{n_1^2} \right),$$

which, when combined with the quantum postulate $W_{n_2} - W_{n_1} = h\nu$, immediately yields the Balmer formula

$$\nu = \frac{2\pi^2 m e^4}{h^3} \left(\frac{1}{n_2^2} - \frac{1}{n_1^2} \right) \equiv R \left(\frac{1}{n_2^2} - \frac{1}{n_1^2} \right).$$

In addition to the above, there were two more ways in which Bohr derived the formula. This was a brilliant victory for the theory: not only did it account for the Balmer series in a natural way but it also yielded the theoretical value of the Rydberg constant:

$$R = \frac{2\pi^2 m_e e^4}{h^3} = 3.1 \times 10^{15} \text{ s}^{-1}.$$

Given that the atomic nucleus is in motion, the theoretical Rydberg constant agrees excellently with what the highest-precision spectroscopic measurements yield.

As the energy (or, equivalently, the number n) increases, the separation between the energy levels becomes small enough to speak of the energy as continuously changing. But continuously changing quantities are a feature of classical mechanics. Thus, for large enough values of the number n — the 'principal quantum number' as Bohr called it — the results of the quantum theory should be identical to those based on classical concepts. This is what makes up the content of the 'correspondence principle', first formulated in Bohr's 1918 paper 'On the quantum theory of line spectra' [5]. Bohr wrote there: "It turned out that although we must give up applying mechanics to describing a transition from one stationary state to another, still it is possible to construct a consistent theory of these states using ordinary mechanics for describing motion in the stationary states themselves. Beyond that the process of emission associated with the transition from one stationary state to another cannot be traced in detail using ordinary electromagnetic ideas. From the point of view of these ideas, the properties of the radiation of the atom are due directly to the motion of the system and to the decomposition of these motions to harmonic components. Nevertheless, it turned out that there is a far-reaching *correspondence* between various types of possible transitions

from one stationary state to another on one hand and various harmonic components — on the other.” Direct calculations using Bohr's theory show that, for example, the frequencies of the spectral lines for transitions between highly excited states (for $n \gg 1$) are equal either to the electron's frequency of rotation around the nucleus or to the harmonics of this frequency. Thus, at large values of the principal quantum number, quantum results go over smoothly to their classical counterparts — thus keeping nature's unity inviolated. On the other hand, Bohr pointed out, “The correspondence principle should be regarded as a purely quantum-theoretical law which can in no way diminish the contrast between these postulates and classical electrodynamics.”

In a broad sense, the correspondence principle can be viewed as assuming that quantum theory contains classical mechanics as a limiting case. This idea was proposed as early as 1906 by Planck, who showed that “classical theory can be characterized simply as a theory in which the quantum of action is infinitely small” [8]. Afterwards, in 1927, it was shown by Ehrenfest that in this limiting case the averages of quantum quantities obey the corresponding equations of classical mechanics. Niels Bohr's was a different premise. He conjectured, namely, that when Planck's constant is kept unchanged and the transition frequency tends to zero, quantum-theoretical results should reduce to those obtained based on classical concepts. Bohr's correspondence principle proved to be a highly flexible and fruitful tool for his theory.

Bohr's studies of the structure of atoms garnered him the 1922 Nobel Prize for physics.

5. The triumph of the Bohr theory

Bohr's theory was quick to get a very high evaluation from the time's most prominent physicists: Einstein, Lorentz, Planck, Rutherford, and Jeans; Sommerfeld, on his part, joined actively in the development of the theory. Bohr, in his first large (over 60 pages) paper already mentioned above, analyzed all the experimental facts available at the time concerning the structure of atoms and molecules — not as much concerned about being mathematically correct as about estimating and qualitatively understanding the phenomena discussed. Here is what W Heisenberg (1901–1976) wrote about that in his paper ‘Quantum theory and its interpretations’ many years later: “Mathematical clarity in itself was of no special value to Bohr. He feared that formal mathematical structure will obscure the physical content of the problem, and he was convinced that a complete physical description should precede a mathematical formulation.”

The conclusive proof of Bohr's theory was the explanation of a spectral series — a series the astronomer E Pickering (1846–1919) had discovered in the spectrum of the star ζ Puppis in 1896. Very reminiscent of its Balmer counterpart, this ‘Pickering series’ was believed to be somehow related to some special state in which hydrogen might be in stars. Similar series were observed by A Fowler in 1912 in his laboratory studies of a hydrogen-helium mixture in vacuum tubes. Thus, a set of contradictory facts emerged, and it was for Bohr to find a way out. It followed from his theory that the Rydberg constant for a hydrogen-like atom depends on the charge of its nucleus, $R_Z = Z^2 R_H$, where R_H is the Rydberg constant for the hydrogen atom. Bohr's conjecture was that the series was not due to hydrogen at all — but was produced by ionized helium He^+ . Before long, F Paschen (1865–1947) and E J Evans confirmed Bohr's predictions experimentally.

Given that the atomic nucleus is in motion, the Rydberg constant for a hydrogen-like atom proves to be dependent on the mass of the nucleus: $R_Z = Z^2 R_H / (1 + m_e / M_Z)$, where M_Z is the mass of the nucleus with atomic number Z . This is how the isotopic shift of spectral lines was explained. This effect enabled H Urey (1893–1981) to discover the heavy isotope of hydrogen — deuterium — in 1932.

Based on his theory, Bohr showed further that, contrary to the belief of the time, the ortho- and parahelium spectra relate not to two different elements but rather to the somewhat different states of one and the same element — helium. In 1915 Bohr for the first time applied the theory of relativity to electrons orbiting around the nucleus. The conclusion he arrived at was that, taking into account the relativistic mass of the electron, only circular orbits can be stationary; elliptic ones must precess.

The study of X-ray spectra was another area where Bohr's theory was progressing triumphantly. W Kossel (1888–1956), in 1914, employed Bohr's ideas to build a theory of X-ray spectra. H G J Moseley (1887–1915), Rutherford's talented disciple who died young in World War I, discovered an experimental law (named after him) which, given the measured frequency (wavelength) of the characteristic X-ray radiation from a given element, yields accurate values of the atomic number — and hence of the charge of the nucleus. The Moseley law demonstrated for the first time that rather than the atomic weight, it is the atomic number — the quantity determining the charge of the nucleus — which is the key characteristic of the atom as far as its position in the periodic systems of elements is concerned. Thereby Moseley convincingly confirmed the hypothesis, due to van den Broek (1870–1926), that it is not the atomic weight but rather the charge of the nucleus (equal to the atomic number) which is key in determining the position of an atom in the periodic system. These data were, in Moseley's opinion, a crucial criterion in the problem of the structure of the atom and gave definitive support to Rutherford and Bohr's views.

After his theory's striking success in explaining a number of experimental facts, Bohr next tried to see how, using the chemical and spectroscopic data available at the time, the periodic system of elements might be interpreted in terms of this theory. That was the subject of his talk at a Physical Society meeting in Copenhagen in 1921. What Bohr did was take a nucleus with a charge Ze , taking — successively — Z electrons, and placing them in different orbits. “In doing so one has to ask oneself, how can an atom be formed by successively adding and binding electrons in the force field of the nucleus?” Clearly, the order of electron shell filling cannot be correct without introducing the Pauli principle. However, still unaware of the major quantum discoveries to follow, Bohr gave a definitive “explanation for the characteristic departures from periodicity in the system of elements” and traced “how the family of rare-earth elements forms.” He concluded that the 72nd element in its properties does not belong to the row of the rare-earth elements (as A Dauwillier believed at the time) but rather is an analogue of zirconium. Late in 1922, a new element — number 72 in the periodic table — was indeed discovered in zirconium ores by D Coster and G Hevesy, marking another brilliant victory for Bohr's quantum theory. The new element was called *hafnium* in honor of the ancient name of the Danish capital.

Bohr's studies of the distribution of electrons in the subgroups of atoms in the periodic table were crucial to the discovery of the *exclusion principle*. Although unaware of this

principle at the time, it was for good reason that Bohr asked himself, Why do not all the electrons concentrate in the lowest-energy state, i.e., in the K-shell? W Pauli (1900–1958) wrote about this afterwards: “... the essential progress made by Bohr’s considerations at that time was in explaining, by means of the spherically symmetric atomic model, the formation of the intermediate shells of the atom and the general properties of the rare earths. The question, as to why all electrons for an atom in its ground state were not bound in the innermost shell, has already been emphasized by Bohr as a fundamental problem in his earlier works. ... he treated particularly the closing of this innermost K-shell in the helium atom and its essential connection with the two non-combining spectra of helium, the ortho- and para-helium spectra. However, no convincing explanation for this phenomenon could be given on the basis of classical mechanics. It made a strong impression on me that Bohr at that time and in later discussions was looking for a *general* explanation which should hold for the closing of *every* electron shell ...” [9].

Bohr’s theory, together with the correspondence principle, accounted for a huge number of experimental facts, explained numerous phenomena, and was rightly considered beautiful at the time. The theory had a powerful impact on the whole further development of atomic physics. Here is how Rutherford summarized the theory in 1931: “I consider the original quantum theory of spectra put forward by Bohr as one of the most revolutionary ever created in science, and I know of no other theory that enjoyed greater success.” Einstein thought equally highly of Bohr’s theory: “All my attempts ... to adapt the theoretical foundation of physics to this [new type of] knowledge failed completely. It was as if the ground had been pulled out from under one, with no firm foundation to be seen anywhere upon which one could have built. That this insecure and contradictory foundation was sufficient to enable a man of Bohr’s unique instinct and sensitivity to discover the principal laws of the spectral lines and of the electron shells of atoms, together with their significance for chemistry, appeared to me as a miracle — and appears to me a miracle even today. This is the highest form of musicality in the sphere of thought.”

6. Generalization of Bohr’s theory

In 1916 A Sommerfeld (1868–1951) extended Bohr’s theory to systems with several degrees of freedom [10] by formulating quantization rules necessary for such systems. Sommerfeld first showed that the condition for the quantization of Bohr orbits follows from the requirement $\oint p dq = nh$, where p and q are canonically conjugate momentum and coordinate and the integration is carried out over a complete cycle of periodic motion. Sommerfeld assumed that this requirement should be imposed on every degree of freedom of the system under study and that to do this a quantum number n_k defined by $\oint p_k dq_k = n_k h$ should be introduced for this degree of freedom. While Sommerfeld regarded the generalized quantization rules as the foundation of quantum theory, they were “unproved and probably unprovable statements” for him. It was only after the creation of quantum mechanics that it proved possible, using the Wentzel–Kramers–Brillouin (WKB) method, to derive the Bohr–Sommerfeld quantum conditions rigorously as certain approximations.

The development of Bohr’s theory led to the notion of the spatial quantization of orbits [10], meaning that the

plane of an orbit may have certain discrete orientations in space. As Sommerfeld wrote, “Undoubtedly, the spatial quantization of Kepler orbits is one of the most surprising consequences of quantum theory. It almost seems magic in terms of how simple its derivation and result are.” Experimentally, Sommerfeld believed, the Zeeman splitting of spectral lines and the Stern-Gerlach experiments provided evidence for the spatial quantization of orbits. Further analysis has shown, however, that the Stern-Gerlach experiment was in fact the discovery of the electron spin.

7. Failures of Bohr’s theory

Although Bohr’s postulates were directly confirmed experimentally and although his theory was successful in explaining numerous experimental facts, over the period 1919–1925 the shortcomings of the theory and the difficulties it faced became increasingly clear. While the theory accounted for the major properties of atoms, the meaning of the quantization rules remained a mystery: it was for good reason that Bohr called them postulates, i.e., unproved assumptions. It was only the creation of consistent quantum mechanics that provided the answer. The Bohr theory, in fact, proved to be applicable to the hydrogen and hydrogen-like atoms only. Even extending the theory to helium — the next atom in the periodic table — proved a futile attempt. Besides, by far not all physical quantities could be calculated with this theory. Even for the hydrogen atom, while the theory produced the frequencies (wavelengths) of spectral lines, there was no general principle with which to calculate the brightness or intensity of the lines. To do this, additional assumptions were introduced by using the correspondence principle — this ‘magic wand’ as Sommerfeld called it. All this, Bohr noted in frustration, turned his theory into “almost intuitive guess-work for true relations.” The period 1919 to 1925 was called that of ‘systematically guess-working’ based on the correspondence principle. Sommerfeld wrote prophetically in 1924: “Using his correspondence principle, Bohr tries to closely link quantum theory with the classical theory of radiation. He proceeds inductively and based on physical considerations, by progressively putting the period of a certain motion in correspondence to each quantum number. The magic power of the correspondence principle asserted itself completely in the derivation of selection rules for quantum numbers and in how serial and line spectra are treated. The principle became a clue to further new discoveries by Bohr and his school. In spite of this, I cannot consider it finally satisfactory even for the reason that there are classical and quantum viewpoints mixed in it. I would like to see it as a certain very special and very important consequence of the future augmented quantum theory, not its foundation.” And this was said literally on the eve of the creation of quantum mechanics!

In essence, Bohr’s theory was not yet a true theory but, in fact, a set of postulates that brought together classical continuity (revolution in orbits) and its clearly irreconcilable opposite — quantum jumps. By that time, new experimental facts appeared which did not fit in any way with Bohr’s theory, and in 1925–1927 new ideas and concepts that were needed to explain them were suggested and developed by Heisenberg, Dirac, de Broglie, Schrödinger, Born, Pauli, Einstein, and other founding fathers of modern quantum theory [11].

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