

Biophysics, Part 2: The structure of atoms

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This is the second part of the series of lecture notes prepared for the students of PHYS119 (Biophysics) in Faculty of Dentistry, Eastern Mediterranean University during the Spring semester 2019-20. In this part we go over the models of atoms from the late 19th century to the early of 20th century. The Rutherford's experiment and Bohr's model are highlighted. The main reference of this lecture note is Ref. [1], however I used other references to emphasis the historic aspect of the subjects.

Among his many experiments, Newton found that sunlight passing through a small opening in a window shutter could be refracted by a glass prism so that it would fall on a screen. The white sunlight thus refracted was spread into a rainbow-colored band—a spectrum. He had discovered dispersion, and his experimental arrangement was the prototype of the modern spectroscope (see Fig. 1).

When, 150 years later, Fraunhofer (Fig. 2) dispersed sunlight using an experimental setup similar to that shown in Fig. 1 to test prisms made of glasses that he had developed, he found that the solar spectrum was crossed by more than 600 narrow, or sharp, dark lines.² Soon after, a number of scientists observed sharp bright lines in the spectra of light emitted by flames, arcs, and sparks. Spectroscopy quickly became an important area of research. It soon became clear that chemical elements and compounds emit three general types of spectra. Continuous spectra, emitted mainly by incandescent solids, show no lines at all, bright or dark, in spectroscopes of the highest possible resolving power. Band spectra consist of very closely packed groups of lines that appear to be continuous in instruments of low resolving

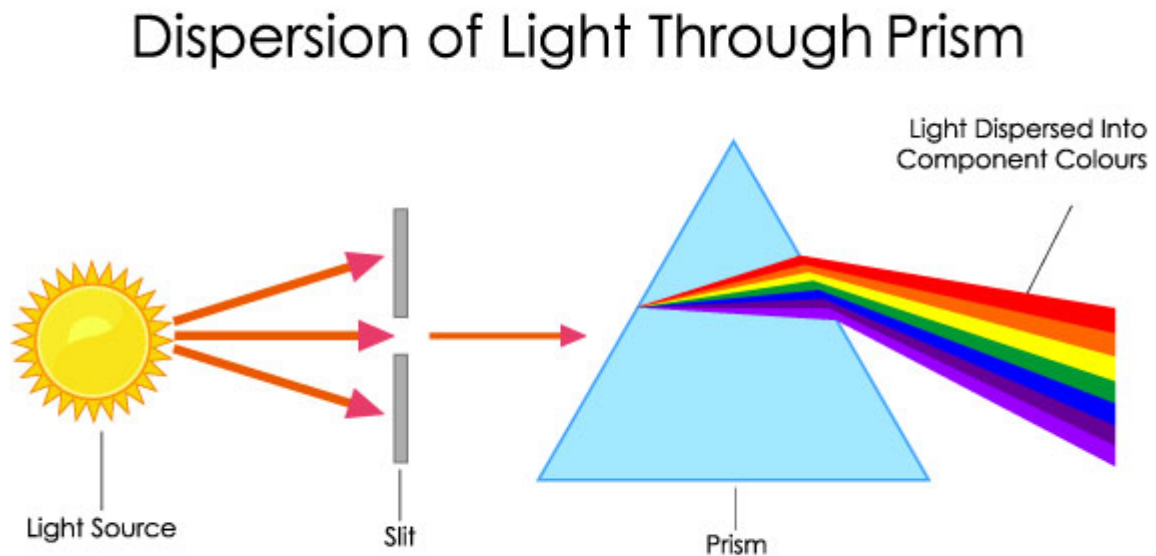


FIG. 1: Dispersion of light from a glass prism.

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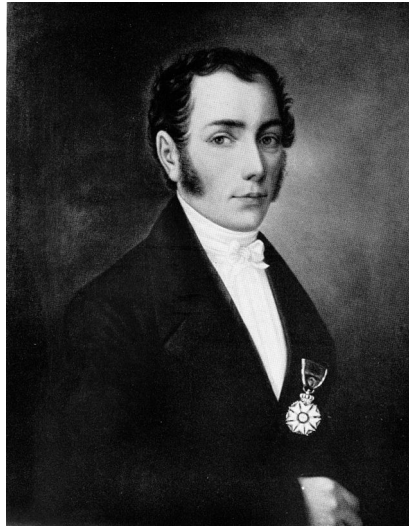


FIG. 2: Joseph Ritter von Fraunhofer, (6 March 1787 – 7 June 1826) was a Bavarian physicist and optical lens manufacturer. He made optical glass and achromatic telescope objective lenses, invented the spectroscope, and developed diffraction grating. In 1814, he discovered and studied the dark absorption lines in the spectrum of the sun now known as Fraunhofer lines. The German research organization Fraunhofer Society is named after him and is Europe's biggest Society for the Advancement of Applied Research.

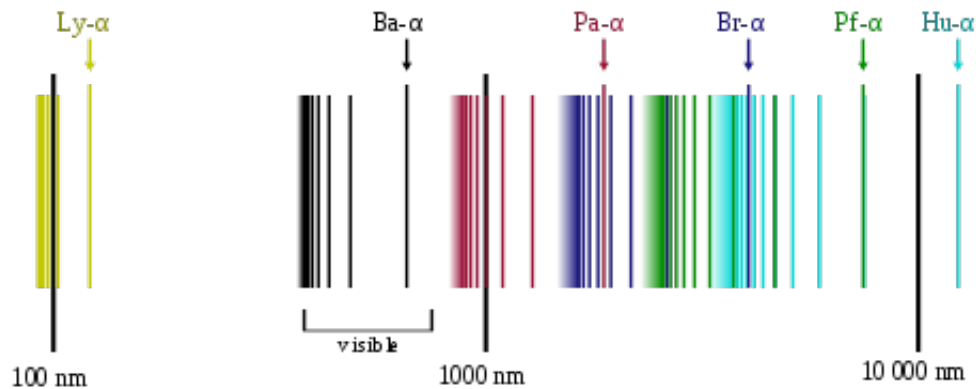


FIG. 3: The spectral series of hydrogen, on a logarithmic scale.

power. These are emitted when small pieces of solid materials are placed in the source flame or electrodes. The line spectra mentioned above arise when the source contains unbound chemical elements. The lines and bands turned out to be characteristic of individual elements and chemical compounds when excited under specific conditions. In Fig. 3 the spectrum line of a Hydrogen atom is shown. Indeed, the spectra could be (and are today) used as a highly sensitive test for the presence of elements and compounds. Line spectra raised an enormous theoretical problem: although classical physics could account for the existence of a continuous spectrum (if not its detailed shape, as we saw with blackbodies), it could in no way explain why sharp lines and bands should exist (see Fig. 4).

Atomic Spectra: The characteristic radiation emitted by atoms of individual elements in a flame or in a gas excited by an electrical discharge was the subject of vigorous study during the late nineteenth and early twentieth centuries. When viewed or photographed through a spectroscope, this radiation appears as a set of discrete lines, each of a particular color or wavelength; the positions and intensities of the lines are characteristic of the element. The wavelengths of these lines could be determined with great precision, and much effort went into finding and interpreting regularities in the spectra. A major breakthrough was made in 1885 by a Swiss schoolteacher, Johann Balmer, who found that the lines in the visible and near ultraviolet spectrum of hydrogen could be represented by the empirical formula

$$\lambda_n = \frac{364.6}{1 - \frac{4}{n^2}} nm, \quad (1)$$

SPECTRUM

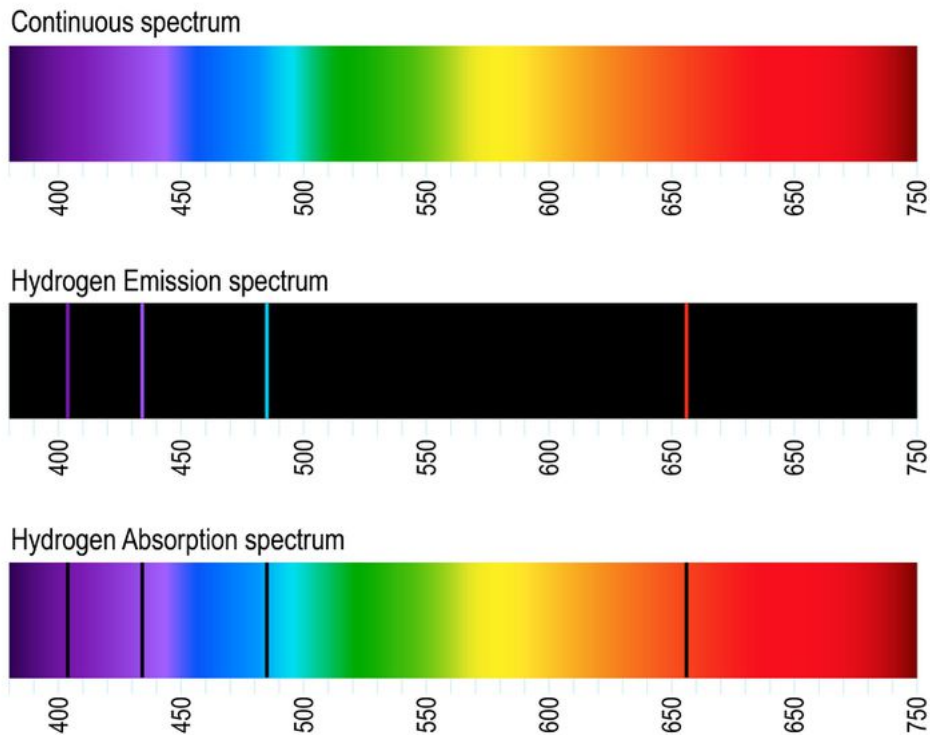


FIG. 4: The hydrogen emission and absorption spectrum in the visible range of wavelength.

where n is a variable integer that takes on the values $n = 3, 4, 5, \dots$ (Fig. 5).

Such an expression, found independently by J. R. Rydberg and W. Ritz and thus called the Rydberg-Ritz formula, gives the reciprocal wavelength

$$\frac{1}{\lambda_{mn}} = R \left(\frac{1}{m^2} - \frac{1}{n^2} \right) \quad (2)$$

for $n > m$. Herein, m and n are integers and R , the Rydberg constant, is the same for all series of spectral lines of the same element and varies only slightly, and in a regular way, from element to element. For hydrogen, the value of R is $R_H = 1.096776 \times 10^7 m^{-1}$. For very heavy elements, R approaches the value of $R_\infty = 1.097373 \times 10^7 m^{-1}$. Such empirical expressions were successful in predicting other series of spectral lines, such as other hydrogen lines outside the visible region.

The hydrogen Balmer series reciprocal wavelengths are those given by Equ. (2), with $m = 2$ and $n = 3, 4, 5, \dots$. Other series of hydrogen spectral lines were found for $m = 1$ (by Theodore Lyman) and $m = 3$ (by Friedrich Paschen).

Exercise: Compute the wavelength and frequency of the series limit for the Lyman, Balmer, and Paschen spectral series of hydrogen.

Exercise : The wavelength of a particular line in the Balmer series is measured to be 379.1 nm . What transition does it correspond to?

Exercise: The series of hydrogen spectral lines with $m = 4$ is called Brackett's series. Compute the wavelengths of the first four lines of Brackett's series.

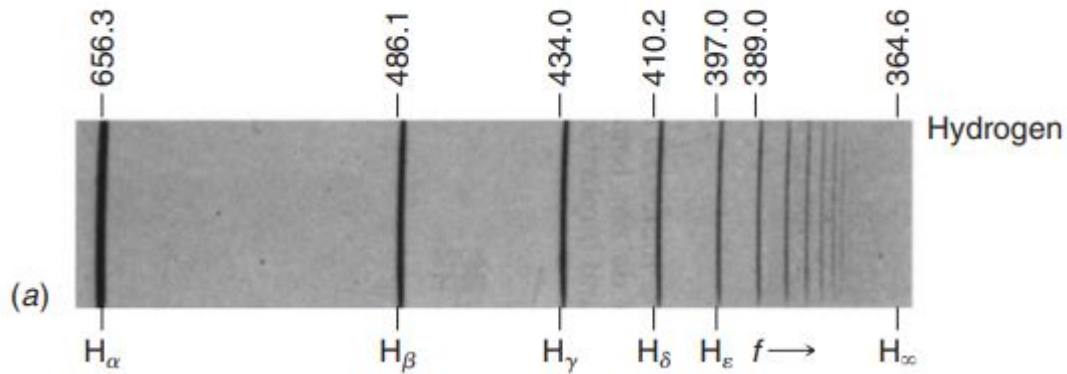


FIG. 5: Emission line spectrum of hydrogen in the visible and near ultraviolet. The lines appear dark because the spectrum was photographed; hence, the bright lines are exposed (dark) areas on the film. The names of the first five lines are shown, as is the point beyond which no lines appear H_{∞} , called the limit of the series.



FIG. 6: Johann Jakob Balmer (1 May 1825 – 12 March 1898) was a Swiss mathematician and mathematical physicist.

Exercise: In a sample that contains hydrogen, among other things, four spectral lines are found in the infrared with wavelengths $7460nm$, $4654nm$, $4103nm$, and $3741nm$. Which one does not belong to a hydrogen spectral series?

Rutherford's Nuclear Model: In the famous Rutherford's experiment, a narrow beam of particles fell on a zinc sulfide screen, which emitted visible light scintillations when struck (Fig. 9).

The distribution of scintillations on the screen was observed when various thin metal foils were placed between it and the source. Most of the α particles were either undeflected or deflected through very small angles of the order of 1° . Quite unexpectedly, however, a few particles were deflected through angles as large as 90° or more. The unexpected scatterings at large angles were described by Rutherford with these words: It was quite the most incredible event that ever happened to me in my life. It was as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you.

Rutherford's atomic model (1911) resembles a miniature solar system. The nucleus is situated in the center of the atom: - its dimensions, as estimated by Rutherford, are in the order of $10^{-15}m$ (for comparison: the radius of the atom is approximately 10,000 times greater), - its charge is $Z \times e$ (Z is the atomic number, e is the electron charge),



FIG. 7: Theodore Lyman IV (November 23, 1874 – October 11, 1954) was a U.S. physicist and spectroscopist, born in Boston. He graduated from Harvard in 1897, from which he also received his Ph.D. in 1900.

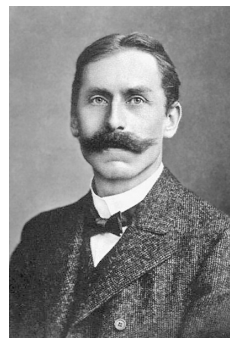


FIG. 8: Louis Carl Heinrich Friedrich Paschen, was a German physicist, known for his work on electrical discharges. He is also known for the Paschen series, a series of hydrogen spectral lines in the infrared region that he first observed in 1908.

- its mass is practically equal to the mass of the atom. Electrons move on circular orbits around the nucleus, held on their orbit by the electrostatic attraction of the nucleus. Although it was a huge step in deciphering the structure of the atom, Rutherford's model had a great deficiency: such an atom was theoretically unstable. According to the laws of electrodynamics an electron moving on a circular orbit must radiate energy, thus slow down and fall into the nucleus on a spiral trajectory. This contradiction was solved by Niels Bohr, danish physicist, who was maybe the greatest genius of the first part of the 20 century.

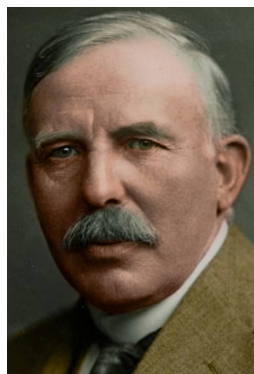


FIG. 9: Ernest Rutherford (1871-1937), New Zealand physicist. Rutherford's early work on radioactivity at McGill University in Canada led to his 1908 Nobel Prize in Chemistry. His most famous work took place in 1909 at the University of Manchester, England. The Geiger-Marsden experiment fired helium ions (alpha particles) at thin gold foil.

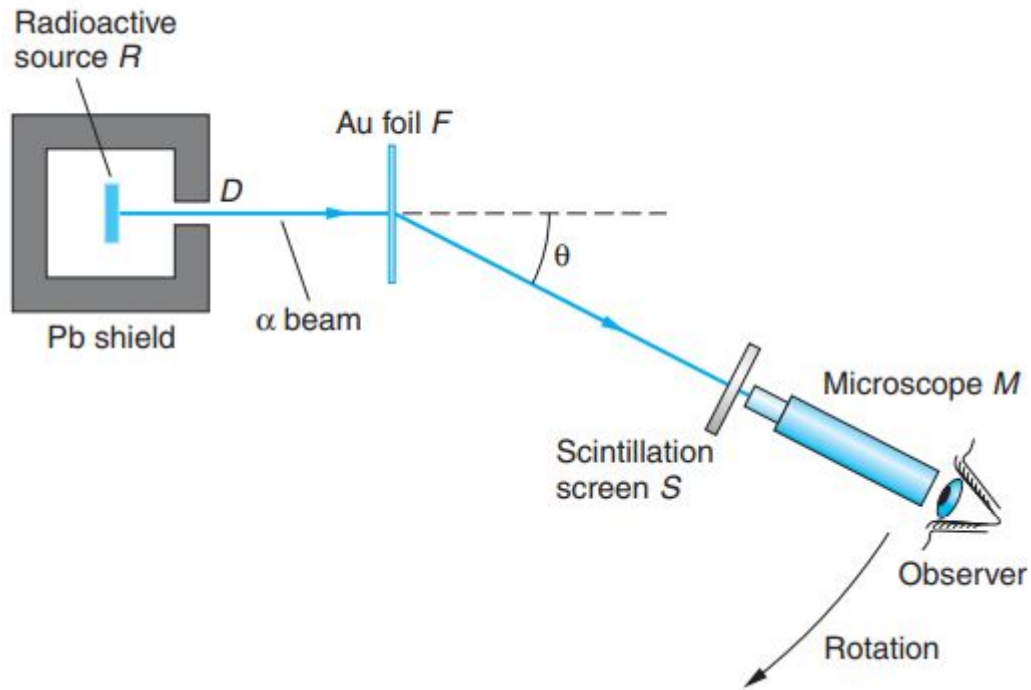


FIG. 10: The beam of particles is defined by the small hole D in the shield surrounding the radioactive source R of ^{214}Bi (called RaC in Rutherford's day). The α beam strikes an ultra-thin gold foil F (about 2000 atoms thick), and the α particles are individually scattered through various angles. Those scattering at the angle θ shown strike a small screen S coated with a scintillator, i.e., a material that emits tiny flashes of light (scintillations) when struck by an α particle. The scintillations were viewed by the observer through a small microscope M . The scintillation screen–microscope combination could be rotated about the center of the foil. The region traversed by the α beam is evacuated. The experiment consisted of counting the number of scintillations as a function of θ .

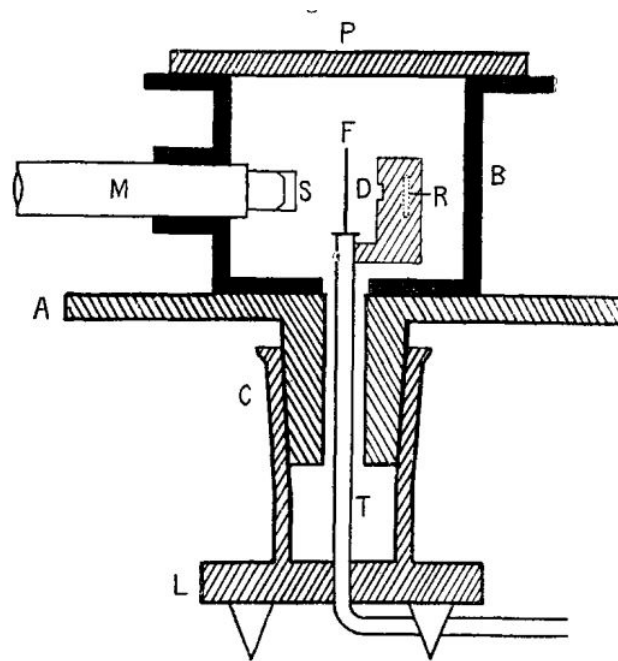


FIG. 11: A diagram of the actual apparatus as it appeared in Geiger and Marsden's paper describing the results. from H. Geiger and E. Marsden, *Philosophical Review*, 25, 507 (1913).

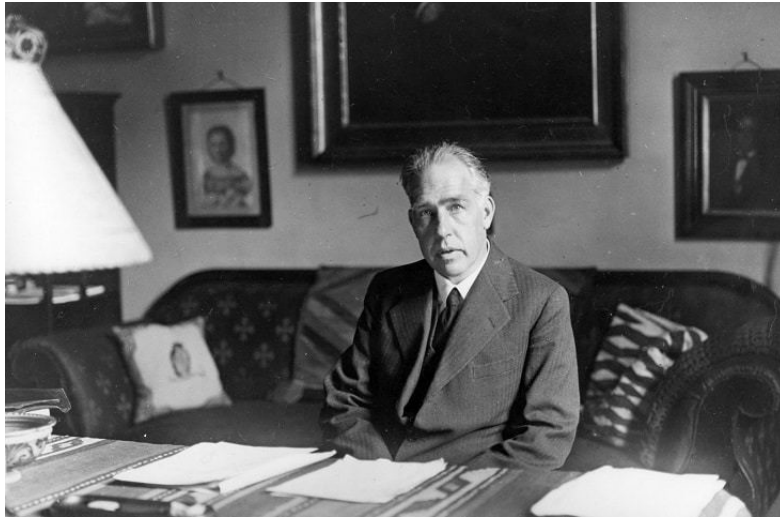


FIG. 12: Niels Bohr (October 7, 1885 – November 18, 1962) was a physicist. His hometown is Copenhagen, Denmark. Niels made his doctoral thesis on the electronic theory of metals in 1911. Niels achieves a Nobel Prize in Physics thanks to his research on atomic structure and radiation.

The Bohr model of Hydrogen atom: In 1913, the Danish physicist Niels H. D. Bohr proposed a model of the hydrogen atom that combined the work of Planck, Einstein, and Rutherford and was remarkably successful in predicting the observed spectrum of hydrogen. The Rutherford model assigned charge and mass to the nucleus but was silent regarding the distribution of the charge and mass of the electrons. Bohr, who had been working in Rutherford's laboratory during the experiments of Geiger and Marsden, made the assumption that the electron in the hydrogen atom moved in an orbit about the positive nucleus, bound by the electrostatic attraction of the nucleus. Classical mechanics allows circular or elliptical orbits in this system, just as in the case of the planets orbiting the Sun. For simplicity, Bohr chose to consider circular orbits.

Bohr's postulates: i) Electrons in an atom can only have defined orbits. The formula defining the radius of the allowed orbits is:

$$L = mvr = n\hbar \quad (3)$$

where L is the angular momentum of the electron ($L = \text{mass} \times \text{velocity} \times \text{radius}$), n is the principal quantum number, and \hbar is the Planck constant. According to the formula, the value of L can only be the multiple of $\hbar = \frac{h}{2\pi}$. The formula can be understood if we rewrite it using the de Broglie formula for the electron wave,

$$2\pi r = n \frac{h}{mv} \Rightarrow 2\pi r = n\lambda \quad (4)$$

where λ is the wavelength of the electron. According to the formula the wavelength of the electron must fit on the circumference of the orbit integral times (see Figure 2), e.g. the wave must join itself in the same phase after going once around the orbit. If this does not occur, the wave would gradually destroy itself by interference. As Bohr postulated, an electron on these defined orbits does not radiate.

ii) When the electron jumps from one allowed orbit to another, the energy difference of the two states is emitted as a photon with the energy of

$$hf = E_{n_i} - E_{n_f} \quad (5)$$

where E_{n_i} and E_{n_f} are the energies of the electron in the initial and final orbits respectively.

Based on his theory, Bohr calculated the parameters of the electronic orbits in the hydrogen atom. His theory predicted the following:

i) In the hydrogen atom the radius of the first orbit is $a_0 = r_1 = 5.3 \times 10^{-11}m$, called the Bohr's radius. The next orbits are situated farther from the nucleus, such that $r_n = n^2 a_0$.

ii) The energy of the electron on the first orbit is $E_1 = -13.6eV$. The energy is negative because the electron in the atom is in a bound state, and energy must be supplied to bring the electron to the free, unbound state (a free electron has zero energy). The energies of the next orbits are given by

$$E_n = \frac{E_1}{n^2} \quad (6)$$

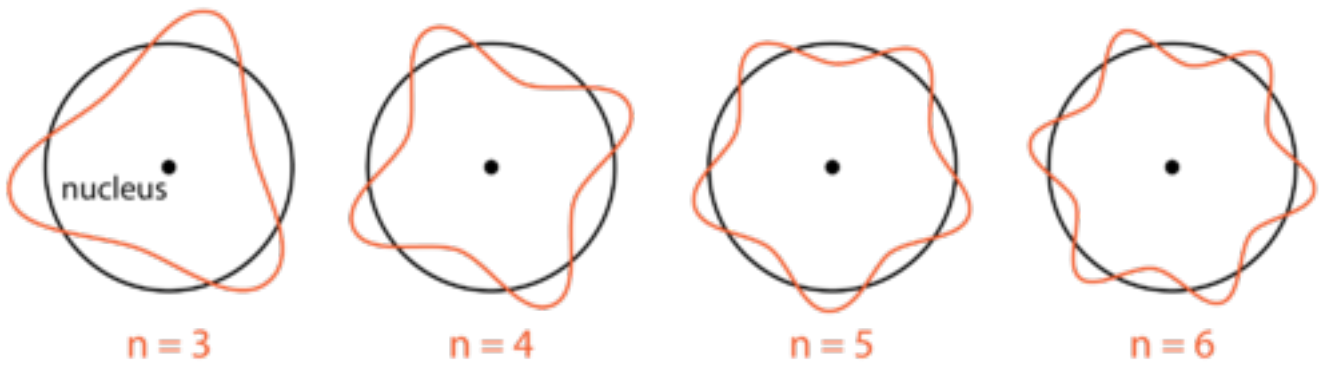


FIG. 13: A moving electron in orbit around the nucleus of an atom behaves more like a spread out wave surrounding the nucleus than a tiny particle existing in one place at a time. Only a whole number of wavelengths can fit around the nucleus. If it's a whole number of wavelengths, the electron waves will form stable standing waves around the nucleus. If it's not a whole number, then the electron wave will wrap round the nucleus, overlap itself and interfere with itself, which is always destructive.

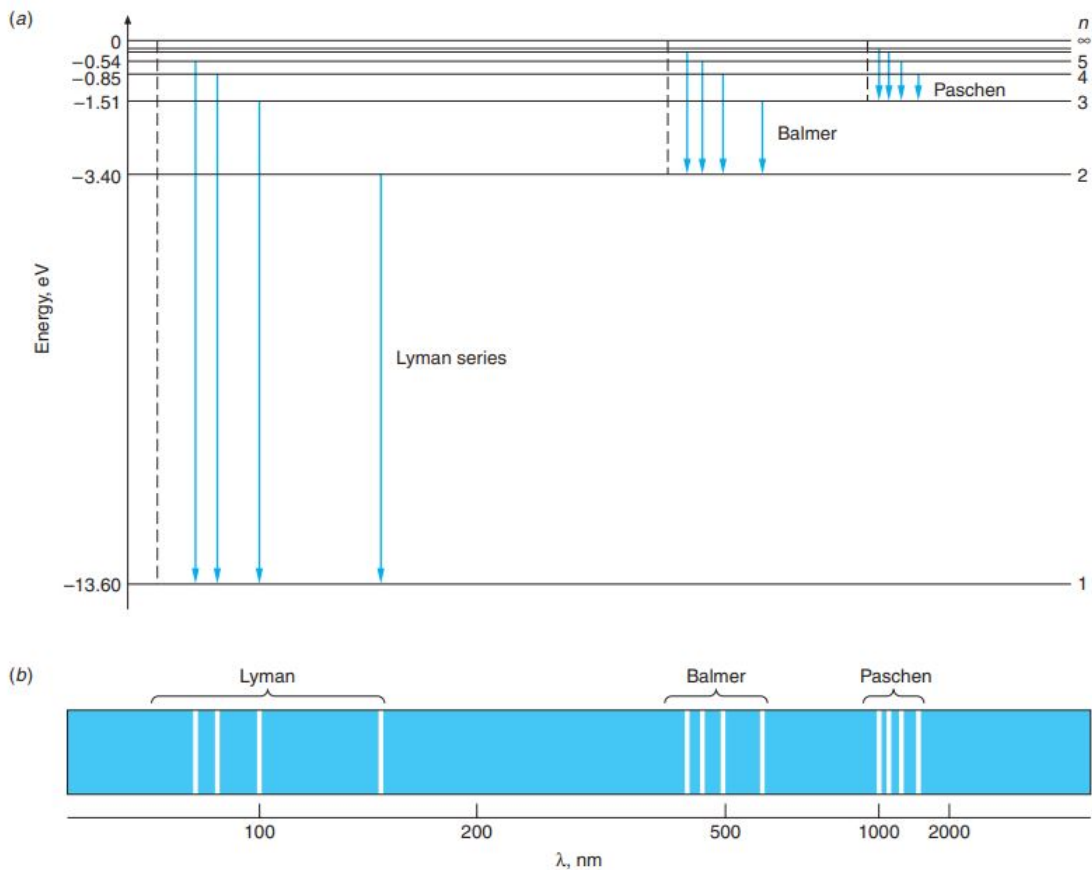


FIG. 14: Energy-level diagram for hydrogen showing the seven lowest stationary states and the four lowest energy transitions each for the Lyman, Balmer, and Paschen series. There are an infinite number of levels. Their energies are given by $E_n = -13.6\text{eV}/n^2$ where n is an integer. The dashed line shown for each series is the series limit, corresponding to the energy that would be radiated by an electron at rest far from the nucleus ($n \rightarrow \infty$) in a transition to the state with $n = n_f$ for that series. The horizontal spacing between the transitions shown for each series is proportional to the wavelength spacing between the lines of the spectrum. (b) The spectral lines corresponding to the transitions shown for the three series. Notice the regularities within each series, particularly the short-wavelength limit and the successively smaller separation between adjacent lines as the limit is approached. The wavelength scale in the diagram is not linear.

where $n = 1, 2, 3, \dots$. One of the greatest achievements of the Bohr-model was that it explained the line spectrum of hydrogen, and predicted exactly the observed wavelengths. It also gave a theoretical explanation for the Balmer-formula which gives the wavelengths correspond to the “jump” of the electron to the second state from higher energy states. According to Bohr’s theory:

$$hf = E_n - E_2 \Rightarrow hf = E_1 \left(\frac{1}{n^2} - \frac{1}{2^2} \right). \quad (7)$$

Hence, we obtain

$$h \frac{c}{\lambda} = E_1 \left(\frac{1}{n^2} - \frac{1}{2^2} \right) \Rightarrow \frac{1}{\lambda} = \frac{-E_1}{hc} \left(\frac{1}{2^2} - \frac{1}{n^2} \right). \quad (8)$$

To be in agreement with Eq. (2), we find

$$R = \frac{-E_1}{hc}. \quad (9)$$

The match was perfect (why?). From his theory Bohr was also able to calculate the value of the Rydberg constant, which was the same as the experimentally determined values. Although Bohr’s postulates were questionable, the predictions of his theory were in a very good agreement with the experimental findings.

Exercise: Compute the wavelength of the H_β spectral line, i.e., the second line of the Balmer series predicted by Bohr’s model. The H_β line is emitted in the transition from $n_i = 4$ to $n_f = 2$.

[1] P. A. Tipler and R. A. Llewellyn, MODERN PHYSICS, (2008) W. H. Freeman and Company, New York.
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