Bohr's Paper

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January 2016

Niels Bohr was a brilliant Danish physicist who came to dominate the world of atomic and nuclear physics during the first half of the twentieth century. Bohr suggested that the planetary model could be saved if one new assumption were made: certain "special states of motion" of the electron, corresponding to different orbital radii, would not result in radiation, and could therefore persist indefinitely without the electron falling into the nucleus. Specifically, Bohr postulated that the angular momentum of the electron, mvr (the mass and angular velocity of the electron and in an orbit of radius r) is restricted to values that are integral multiples of $\frac{h}{2}$. The radius of one of these allowed

Bohr orbits is given by $r = \frac{nh}{2*\pi mv}$ in which h is Planck's constant, m is the mass of the electron, v is the orbital velocity, and n can have only the integer values 1, 2, 3, etc. The most revolutionary aspect of this assumption was its use of the variable integer n; this was the first application of the concept of the quantum number to matter. There is a proportion: larger the value of n, then larger the radius of the electron orbit, and greater the potential energy of the electron.

As the electron moves to orbits of increasing radius, it does so in opposition to the restoring force due to the positive nucleus, and its potential energy is thereby raised. This is entirely analogous to the increase in potential energy that occurs when any mechanical system moves against a restoring force— as, for example, when a rubber band is stretched or a weight is lifted.

Thus what Bohr was saying, in effect, is that the atom can exist only in certain discrete energy states: the energy of the atom is quantized. Bohr noted that this quantization nicely explained the observed emission spectrum of the hydrogen atom. The electron is normally in its smallest allowed orbit, corresponding to n = 1; upon excitation in an electrical discharge or by ultraviolet light, the atom absorbs energy and the electron gets promoted to higher quantum levels. These higher excited states of the atom are unstable, so after a very short time (around 10—9 sec) the electron falls into lower orbits and finally into the innermost one, which corresponds to the atom's ground state. The energy lost on each jump is given off as a photon, and the frequency of this light provides a direct experimental measurement of the difference in the energies of the two states, according to the Planck-Einstein relationship $e = h*$.

Each spectral line represents an energy difference between two possible states of the atom. Each of these states corresponds to the electron in the hydrogen atom being in an "orbit" whose radius increases with the quantum number n. The lowest allowed value of n is 1; because the electron is as close to the nucleus as it can get, the energy of the system has its minimum (most negative) value. This is the "normal" (most stable) state of the hydrogen atom, and is called the ground state.

If a hydrogen atom absorbs radiation whose energy corresponds to the difference between that of n=1 and some higher value of n, the atom is said to be in an excited state. Excited states are unstable and quickly decay to the ground state, but not always in a single step. For example, if the electron is initially promoted to the n=3 state, it can decay either to the ground state or to the $n = 2$ state, which then decays to $n = 1$. Thus this single $n = 1 \Rightarrow 3$ excitation can result in the three emission lines depicted in the diagram above, corresponding to $n = 3 \Rightarrow 1$, $n = 3 \Rightarrow 2$, and $n = 2 \Rightarrow 1$.

If, instead, enough energy is supplied to the atom to completely remove the electron, we end up with a hydrogen ion and an electron. When these two particles recombine ($H + +e- \Rightarrow H$), the electron can initially find itself in a state corresponding to any value of n, leading to the emission of many lines.

The lines of the hydrogen spectrum can be organized into different series according to the value of n at which the emission terminates (or at which absorption originates.) The first few series are named after their discoverers. The most well known (and first-observed) of these is the Balmer series one, which lies mostly in the visible region of the spectrum. The Lyman lines are in the ultraviolet, while the other series lie in the infrared. The lines in each series crowd together as they converge toward the series limit which corresponds to ionization of the atom and is observed as the beginning of the continuum emission. Note that the ionization energy of hydrogen (from its ground state) is 1312 $\frac{kJ}{mol}$.

The line emission spectra we have been discussing are produced when electrons which had previously been excited to values of n greater than 1 fall back to the $n = 1$ ground state, either directly, or by way of intermediate-n states. But if light from a continuous source (a hot body such as a star) passes through an atmosphere of hydrogen (such as the star's outer atmosphere), those wavelengths that correspond to the allowed transitions are absorbed, and appear as dark lines superimposed on the continuous spectrum.

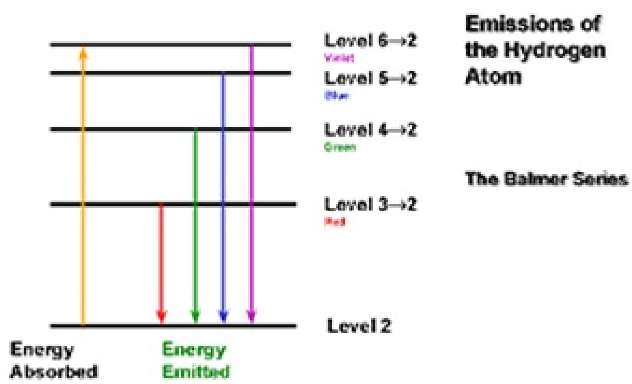

Figure 1: The Universe