Bohr's Theory of Hydrogen atom and hydrogen spectra

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1 Introduction to Bohr's Theory

This is a theory of atomic structure that explains the spectrum of hydrogen atoms. It assumes that the electron orbiting around the nucleus can exist only in certain energy states, a jump from one state to another being accompanied by the emission or absorption of a quantum of radiation.

2 Bohr's Theory and Hydrogen atom spectra

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

3 Bohr's atomic model

In 1913 Bohr proposed his quantized shell model of the atom to explain how electrons can have stable orbits around the nucleus. The motion of the electrons in the Rutherford model was unstable because, according to classical mechanics and electromagnetic theory, any charged particle moving on a curved path emits electromagnetic radiation; thus, the electrons would lose energy and spiral into the nucleus. To remedy the stability problem, Bohr modified the Rutherford model by requiring that the electrons move in orbits of fixed size and energy. The energy of an electron depends on the size of the orbit and is lower for smaller orbits. Radiation can occur only when the electron jumps from one orbit to another. The atom will be completely stable in the state with the smallest orbit, since there is no orbit of lower energy into which the electron can jump.

$$e = hv$$

4 formulas related to Bohr's Theory

In the Bohr model, the wavelength associated with the electron is given by the DeBroglie relationship.

$$\lambda = h \div mv$$

and the standing wave condition that circumference = whole number of wavelengths. In the hydrogenic case, the number n is the principal quantum number.

$$2\pi r = L\lambda$$

These can be combined to get an expression for the angular momentum of the electron in orbit.

$$L = nh \div 2\pi$$

5 Hydrogen spectral series

The emission spectrum of atomic hydrogen is divided into a number of the spectral series.
Lyman series is for n=1 in Hydrogen spectrum and all the wavelengths in the Lyman series are in the ultraviolet band.
Almer lines are for n=2 in Hydrogen atom. Almer lines are historically referred to as "H-alpha", "H-beta", "H-gamma" and so on, where H is the element hydrogen.[8] Four of the Balmer lines are in the technically "visible" part of the spectrum, with wavelengths longer than 400 nm and shorter than 700 nm. Parts of the Balmer series can be seen in the solar spectrum. H-alpha is an important line used in astronomy to detect the presence of hydrogen.
Bohr series or in the other words Paschen lines are for n=3 in Hydrogen spectrum.The Paschen lines all lie in the infrared band.[9] This series overlaps with the next (Brackett) series, i.e. the shortest line in the Brackett series has a wavelength that falls among the Paschen series. All subsequent series overlap.