Hydrogen Emission Spectrum

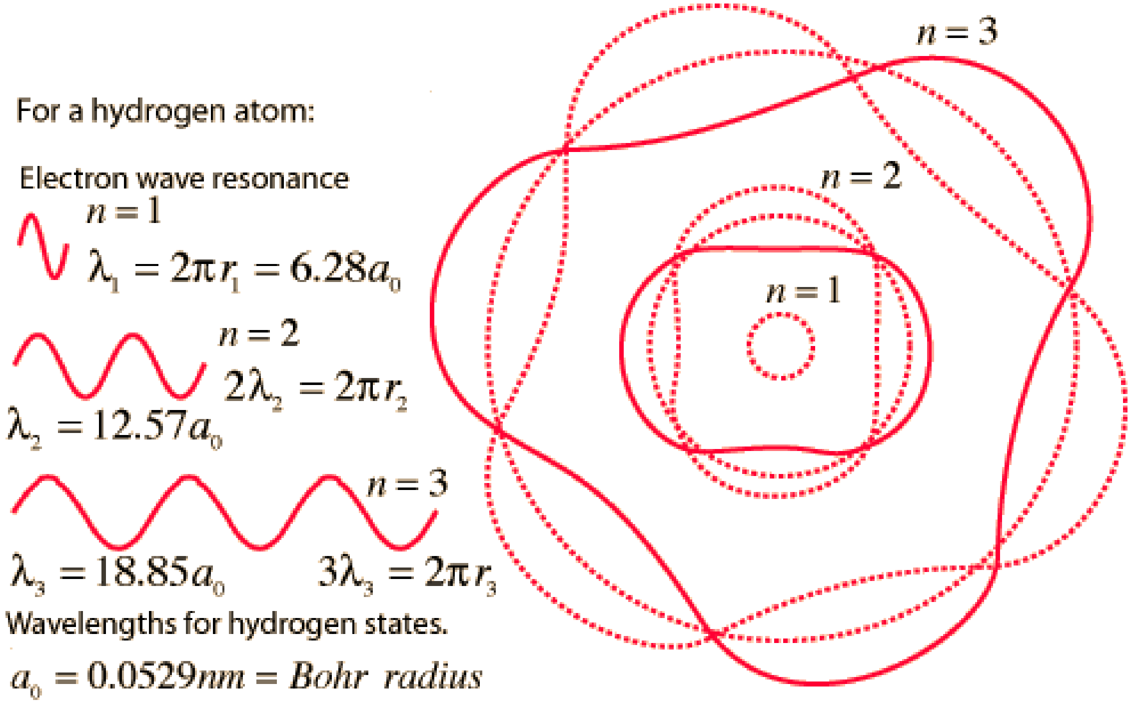
In the Bohr model of the atom the De Broglie Wavelength:

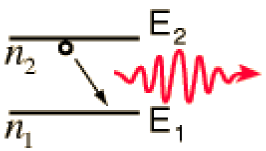
$$\lambda = \frac{h}{mv}$$

Is used to constrain the possible orbits that an electron can occupy by using what is called the "standing wave condition" that is there can only be whole number multiples of the De Broglie wavelength around each circular orbit:

$$n\lambda = 2\pi r$$

Where n is called the "principal quantum number"





A downward transition involves emission of a photon of energy:

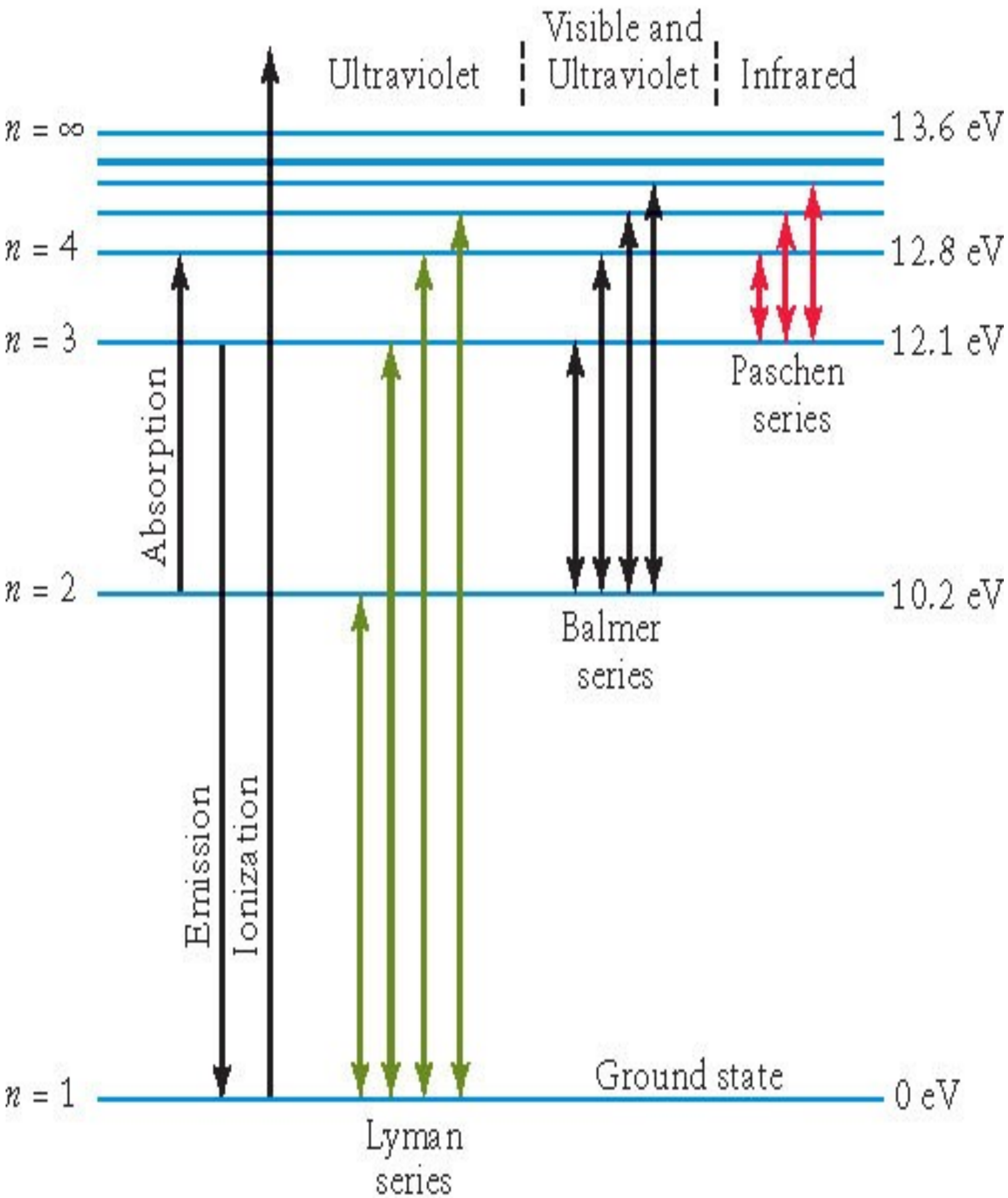
$$E_{photon} = h\upsilon = E_2 - E_1$$

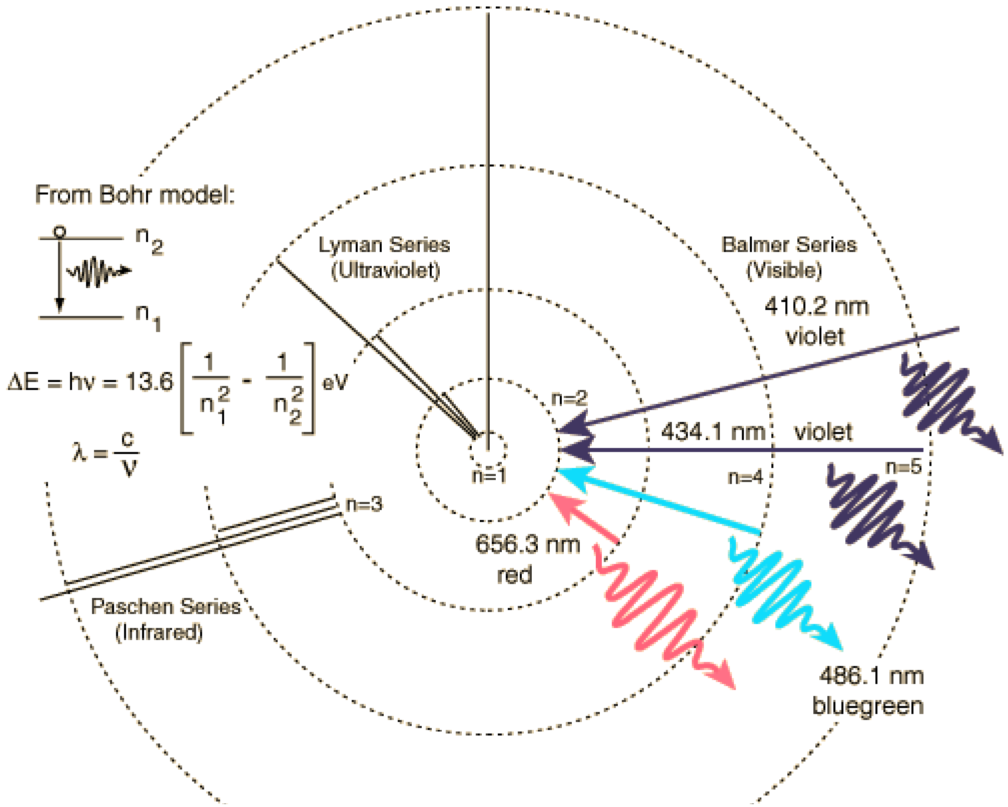
Given the expression for the energies of the hydrogen electron states:

$$h\upsilon = \frac{2\pi^2 me^4}{h^2}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] = -13.6\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]\text{eV}$$

Unfortunately the assumption that the orbits must be circular is too simplistic and the Bohr model is unable to deal with the more complex interactions when more than one electron is involved.

Nevertheless the Bohr model does explain the Hydrogen Spectrum, and more importantly gives us a visualisable model of why atomic energy levels might be quantised and of the importance of the principal quantum number.





Wavelength (nm)	Relative Intensity	Transition	Colour
383.5384	5	9 -> 2	Violet
388.9049	6	8 -> 2	Violet
397.0072	8	7 -> 2	Violet
410.174	15	6 -> 2	Violet
434.047	30	5 -> 2	Violet
486.133	80	4 -> 2	Bluegreen (cyan)
656.272	120	3 -> 2	Red
656.2852	180	3 -> 2	Red

Ref: http://hyperphysics.phy-astr.gsu.edu/hbase/bohr.html#c4 and http://www.ualberta.ca/~pogosyan/teaching/ASTRO_122/lect6/lecture6.html