Lecture 20 (2-4), Oct 37, 2022

Spectroscopy and Bohr's Model

- Burning stuff to see what colour it emits
- First there was Thomson's plum pudding model, then it was discovered that atoms are mostly empty space, so came Rutherford's model of electrons orbiting a nucleus
  - However the orbiting electrons should emit radiation, which causes them to lose energy and spiral into the nucleus with a lifetime of 10×10^{-8} s
- Bohr's model had some assumptions:
  1. Electrons are in a circular orbit (Rutherford model)
  2. The electrons are in stationary states/orbits, i.e. they don't radiate energy
  3. Radiation is emitted when electrons change orbits
- Bohr assumed angular momentum is quantized $L = mvr = n\hbar$ where n is a positive integer and $\hbar = \frac{h}{2\pi}$ is the reduced Planck constant
- To solve for the energy of the electron, equate the centripetal force with the Coulomb force:
  - $F_{cent} = m\frac{v^2}{r} = \frac{ke^2}{r}^2 = F_{Coulomb}$
  - $k\frac{e^2}{r^2} = \frac{mv^2}{r} \implies ke^2 = (mvr)v = Lv = n\hbar v$
  - This gives: $v = \frac{ke^2}{n\hbar}, r = \frac{(n\hbar)^2}{kme^2}$