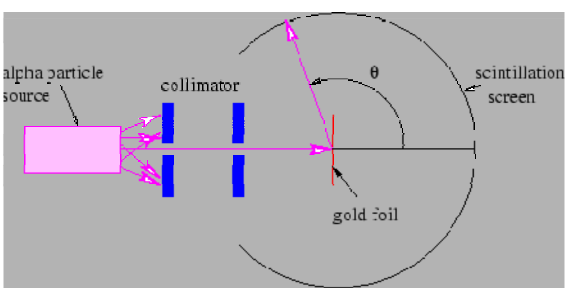
The Atom

The Geiger-Marsden Experiment

- Hans Geiger and Ernest Marsden working with Ernest Rutherford (c. 1911), measure the angular distribution of alpha particles (He nucleus) scattered from a thin gold foil
- The typical scattering should have been very small (around 0.01^{o})
- Most particles suffered only small deflections, but a small fraction scattered through large angles, some greater than 90^{o}



Schematic of Geiger-Marsden experiment

- Rutherford calculated theoretically the number of alpha particles expected at a particular scattering angle based on Coulomb's force law
- His results agreed with the experimental data if the positive atomic charge was confined to a region of linear size of approximately 10^{-15} m
- This and subsequent experiments confirmed the existence of a small massive positive nucleus inside the atom

Calculation

- Consider an alpha particle of charge q shot head on towards a stationary nucleus of charge Q
- Initially the total energy is equal to the kinetic energy of the alpha particle, E_k
- We take the separation distance to be large so that there is no potential energy

- At the point of closest approach, a distance d from the center of the nucleus, the alpha particle stops and is about to turn back
- At this point, the total energy is the electrostatic potential energy given by

$$E = k\frac{qQ}{d}$$

- Then, by conservation of energy

$$E_k = k\frac{qQ}{d}$$

$$d = k\frac{qQ}{E_k}$$

- Assuming a kinetic energy for the alpha particle (q=2e) of 2 MeV directed at a gold nucleus (Q=79e) the result gives $d=1\text{x}10^{-13}$m
- This is outside of the range of the nuclear force, so the alpha particle is simply repelled by the electrical force

- As the energy of the incoming particle increases, the distance of closest approach decreases
- The smallest it can get is the same order as the radius of the nucleus
- Experiments show that the nuclear radius depends on the mass number (the number of protons and neutrons in the nucleus), *A*

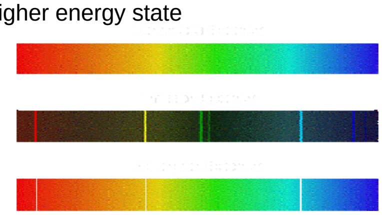
$$\boxed{R = R_0 A^{\frac{1}{3}}}$$

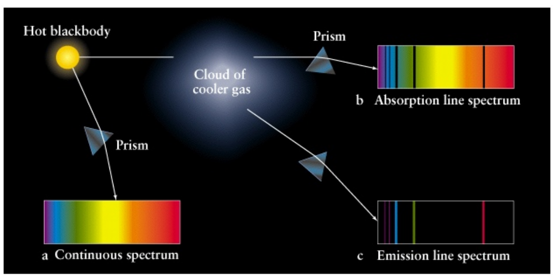
Fermi radius R_0 = $1.2\text{x}10^{-15}$ m

Atomic Spectra

- It had been observed early in the 19^{th} century that the spectrum from excited gases was not continuous but discrete
- Each gas produced its own unique line spectrum

- Emission Spectra
  - The set of wavelength of light emitted by atoms of an element
- Absorption Spectra
  - The wavelengths that are absorbed by the element used by the electrons to jump to a higher energy state





- Johann Balmer (1885) discovered, by trial and error, that the wavelengths of the emission spectrum of Hydrogen were given by:

$$\frac{1}{\lambda} = R\left(\frac{1}{4} - \frac{1}{n^2}\right) \qquad n = 3,4,5,\ldots$$

Bohr Model

- Bohr had studies in Rutherford's lab for several months in 1912 and was convinced that Rutherford's planetary model of the atom had validity
- He felt that it must be related to the newly developing quantum theory
- Bohr argued, that the electrons cannot lose energy continuously. But must do so in quantum "jumps"

- Bohr postulated that electrons move in circular orbits, but only certain orbits are allowed, without radiating energy
  - This violates classical ideas since accelerating charges are supposed to emit EM waves
- He called these specific orbits **stationary states**
- Light is emitted when an electron jumps from a higher stationary state to one of lower energy

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- When the electron jumps a single photon of light is emitted whose energy is given by

$$hf = \Delta E$$

- Bohr set out to determine what energies these orbits would have in Hydrogen
- He found that his theory would agree with Balmer's formula if he assumed that the electron's angular momentum is quantized and equal to an integer n times $h/2\pi$

$L = I\omega$

For a single particle of mass m moving in a circle or radius r with speed v

$I = mr^2$ and $\omega = \frac{v}{r}$

$$L = mr^2\left(\frac{v}{r}\right) = mvr$$

- Bohr's quantum condition is

$$mvr = \frac{nh}{2\pi}$$

Where n is the **principal quantum number** of the orbit

- There was no theoretical foundation for the equation
- Bohr had searched for some "quantum condition" but had not found anything that worked with his experiments
- Bohr's reason for the equation was simply that "it worked"
  - Note that while the equation works for hydrogen it does not work for other atoms

- Calculating the potential and kinetic energies of the electrons in a Hydrogen atom using Bohr's model gives us

$$E = -\frac{13.6}{n^2} eV$$

  - The quantum number n that labels the orbital radii also labels the energy levels
  - The energy is measured in electron volts (as is customary in particle and quantum physics)
  - The lowest energy state (n=1) is referred to as the ground state
  - The binding energy or ionization energy is 13.6 eV (moving from E_1=-13.6 eV to E=0)

- Bohr’s model was successful because
  - It explained the emission line spectra of hydrogen
  - It explained the absorption line spectra
  - It ensured the stability of atoms
    - It did this by decree: the ground state is the lowest state for an electron, it cannot go lower and emit more energy
  - It accurately predicted the ionization energy of hydrogen
- However, the Bohr model was not as successful with other atoms