ATOMIC PHYSICS

Atomic Spectra

- If you take an evacuated glass tube filled with a pure atomic gas and apply a high potential difference (V) between electrodes, a current is produced within the gas.
  - The tube will begin to emit light.
  - The frequency of the emissions will depend on the properties of the gas.
  - Ex: Neon lights.
  - The color of the neon lights will vary based off the gases present in the tube.

- **Emission Spectrum** → A diagram that shows the wavelengths of radiant energy that a substance emits.

- **Absorption Spectrum** → A diagram that shows the wavelengths of radiant energy that a substance absorbs.

- Using these spectra is how we have identified the elements present in stars, including the Sun.

- What determines the frequencies that are emitted?
  - In the Bohr model of the atom, electrons jump immediately from one orbit to another.
  - They are **never** in between orbits.
  - Orbits represent energy. Outer orbits are positions of high energy, inner orbits are lower energy.
  - As an electron moves from an outer orbital to an inner orbital, it radiates energy.
  - The frequency of the radiation depends on the energy change of the atom.
  - The energy of the emitted photon is equal to the energy change of the atom.
  - So Planck’s equation can be rewritten as:

\[
E = E_f - E_i = hf
\]

Bohr Model (1913)

- **Ground state** → the lowest energy state.
  - Aka the Bohr radius.
- When light is shone on an atom, photons whose energy (hf) matches the energy difference between two levels can be absorbed.
  - Only photons with that energy will be absorbed.
  - I.e., the frequency of the light will determine if energy is absorbed by the atom or not.
- When the atom absorbs that energy, it causes an e⁻ to jump to a higher energy state.
  - I.e., it jumps to a higher orbit.
  - In an excited state now.
- **Spontaneous Emission** → in that excited state, there is a probability that the e⁻ will jump back down to a lower energy level and emit a photon.
  - The frequencies of these emitted photons are responsible for the bright lines we see in the emission spectrum.
  - There is a direct relationship between the “size” of the e⁻ jump and the energy of the photon.
  - An e⁻ in the 4th energy level could jump to the 3rd, the 2nd, or the ground level.
  - A greater jump means that more energy is emitted.

![Diagram of Energy Levels](image)

Example Problem 1

An electron in a hydrogen atom drops from energy level 4 (-1.4 x 10⁻¹⁹ J) to energy level 2 (-5.5 x 10⁻¹⁹ J). What is the frequency of the emitted photon and which line in the emission spectrum corresponds to this event?

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**The Electron Cloud**

- 1926 → Erwin Schrödinger proposed an equation that when solved would give you the probability of finding an electron in a given area of space.
  - Probability because of the Uncertainty Principle.
  - The probability of finding electrons is highest at the valence levels, but there are probabilities elsewhere, too.
  - An **electron cloud** of probabilities surrounds the nucleus.