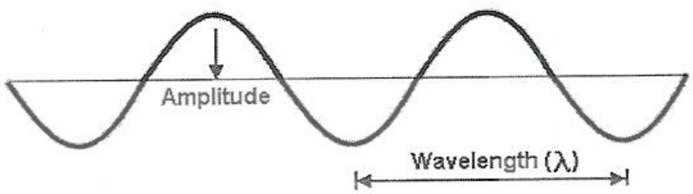
Atomic Structure : Electromagnetic Radiation

26.

Electromagnetic Radiation - Radiation that contains both electric and magnetic components and travels at the speed of light (2.998×10^8 m/s).

- Wave-like properties of electromagnetic radiation
  - Carry energy
  - Speed (s)
  - Frequency (ν) - The number of waves or cycles per unit of time, generally reported in units of cycles per second or hertz (Hz).
  - Wavelength (λ) - The smallest distance between repeating points on a wave.
  - Amplitude - The distance between the highest (or lowest) point on a wave and the center of gravity of the wave.



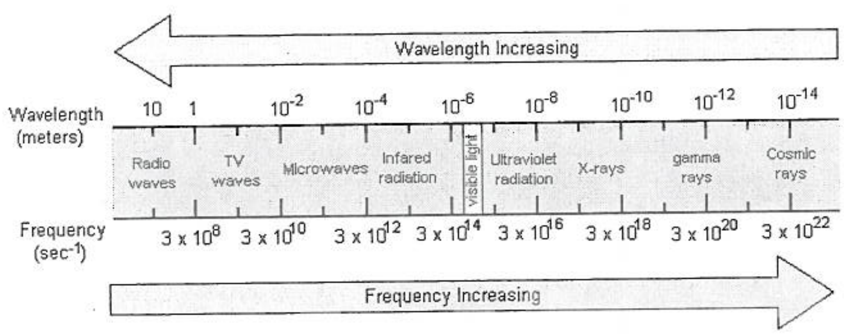
High Frequency

Low Frequency

The product of the frequency and the wavelength is equal to the speed of a wave.

- For electromagnetic radiation, the speed of the waves is equal to the speed of light.

$$\nu\lambda = s = 2.998 \times 10^8 \text{ m/s}$$



- **Visible light lies between the range 700 nm (red) to 400 nm (violet).**

Next: "Atomic Spectra"

Atomic Structure : Atomic Spectra

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1814 - Joseph von Fraunhofer

- Studied the absorption spectrum of the light given off by the sun.
  - **Absorption Spectrum** - The spectrum of dark lines against a light background that results from the absorption of selected frequencies of the electromagnetic radiation by an atom or molecule.

1855-1860 - Robert Bunsen and Gustav Kirchhoff

Robert Bunsen

Gustav Kirchhoff

- Developed a spectroscope that focused the light from a burner flame onto a prism that separated the light into its spectrum. Studied the emission spectrum of several metals.
  - **Emission Spectrum** - The spectrum of bright lines against a dark background obtained when an atom or molecule emits radiation when excited by heat or an electric discharge.

Flame Test

The following metals give emit certain colors of light when their atoms are excited.

Metal	Color
Sodium (Na)	Yellow
Lithium (Li)	Pink/Red
Potassium (K)	Purple
Copper (Cu)	Green
Calcium (Ca)	Pink
Barium (Ba)	Yellow/Orange
Strontium (Sr)	Red/Orange

1885 - Johann Jacob Balmer

- Analyzed the hydrogen spectrum and found that hydrogen emitted four bands of light within the visible spectrum:

Wavelength (nm)	Color
656.2	red
486.1	blue
434.0	blue-violet
410.1	violet

- Balmer found that the data fit to the following equation:

$$\frac{1}{\lambda} = R_H \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

- λ = wavelength (nm)
- R_H = Rydberg's constant = $1.09678 \times 10^{-2}\ nm^{-1}$
- n_1 = the lower energy level
- n_2 = the higher energy level

For example, to calculate the wavelength of light emitted when the electron in a hydrogen atom falls from the fourth energy level to the second energy level:

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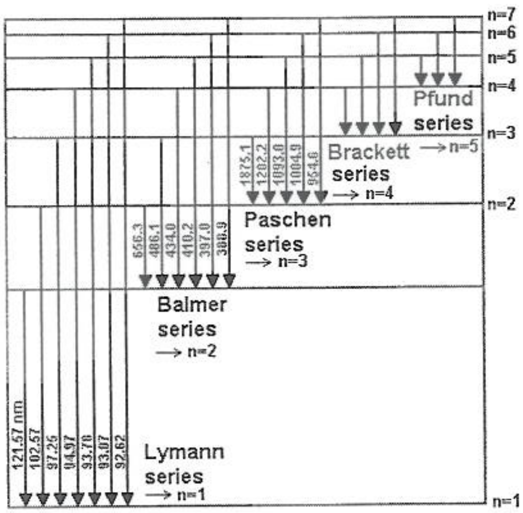
$$\frac{1}{\lambda} = (1.09678 \times 10^{-2}\ \text{nm}^{-1})\left[\frac{1}{2^2} - \frac{1}{4^2}\right]$$
$$= 0.002$$
$$\lambda = 486.3\ \text{nm}$$



Each series is named after its discoverer.

- The **Lyman series** is the wavelengths in the ultra violet (UV) spectrum of the hydrogen atom, resulting from electrons dropping from higher energy levels into the **n = 1** orbit.
- The **Balmer series** is the wavelengths in the visible light spectrum of the hydrogen atom, resulting from electrons falling from higher energy levels into the **n = 2** orbit.
- The **Paschen series** is the wavelengths in the infrared spectrum of the hydrogen atom, resulting from electrons falling from higher energy levels into the **n = 3** orbit.
- The **Brackett series** is the wavelengths in the infrared spectrum of the hydrogen atom, resulting from electrons falling from higher energy levels into the **n = 4** orbit.
- The **Pfund series** is the wavelengths in the infrared spectrum of the hydrogen atom, resulting from electrons falling from higher energy levels into the **n = 5** orbit.

1900 - Max Planck

- Hypothesized that substances were surrounded by oscillating "resonators" which emitted energy that was quantized, or countable, because he assumed that there were only a limited number of energies at which these oscillators could exist.

1905 - Albert Einstein

- Extended Planck's work to include light, hypothesizing that light was also quantized.
- He assumed that:
  1. Light was made up of small, discrete particles of energy called **photons**.
  2. The energy (E) of a photon is proportional to its frequency (v):

where **h is Planck's constant = 6.626 x 10^{-34} J-s.**

$$E = h\nu$$

Next: "Bohr Model"

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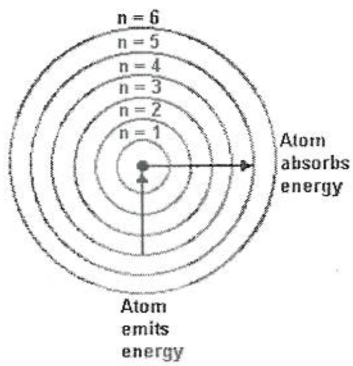
Atomic Structure : Bohr Model

In 1913, **Niels Bohr** proposed a model for the hydrogen atom which explained the spectrum of a hydrogen atom based upon the following assumptions:

Bohr Model for the Hydrogen Atom

1. The electron in a hydrogen atom travels around the nucleus is a circular orbit.
2. The energy of the electron is directly proportional to its distance from the nucleus.
3. There are a limited number of specific allowed energy levels, i.e., the orbitals are quantized.
4. The angular momentum of the electron in any orbital is an integral multiple of Planck's constant divided by 2π.
5. When light is absorbed, the electron jumps from a lower energy level to a higher energy level. When light is emitted, the electron falls from a higher energy level to a lower energy level.
6. The energy of light absorbed or emitted is equal to the difference in the energy levels of the orbits between which the electron jumps of falls.

To simplify these assumptions, remember these key words and concepts:



1. **CIRCULAR ORBIT**
   - in which the electron travels around the nucleus.
2. **Energy $\propto$ Distance**
   - energy of an orbital is directly proportional to the distance from the nucleus.
3. **QUANTIZED**
   - limited number of countable allowed energy levels.
4. **ANGULAR MOMENTUM**
   - fixed for each energy level.
5. **Absorption is Low E to High E**
   Emission is High E to Low E
6. **Energy of absorption/emission = E difference of energy levels.**

Problems with the Bohr Model

1. The model is only valid for the hydrogen atom (with one electron).
2. The first assumption concerning the electrons in fixed circular orbits violates the laws of classical mechanical physics.
3. The ANGULAR MOMENTUM assumption violates the Heisenberg uncertainty principle.

- **Heisenberg Uncertainty Principle** - It is impossible to determine the position (x) and the momentum (mv) of a particle simultaneously with certainty.
- In the Bohr model, it is assumed that there are fixed angular momentums for each quantized orbital.

Werner Heisenberg

Next: "Wave-Particle Duality"