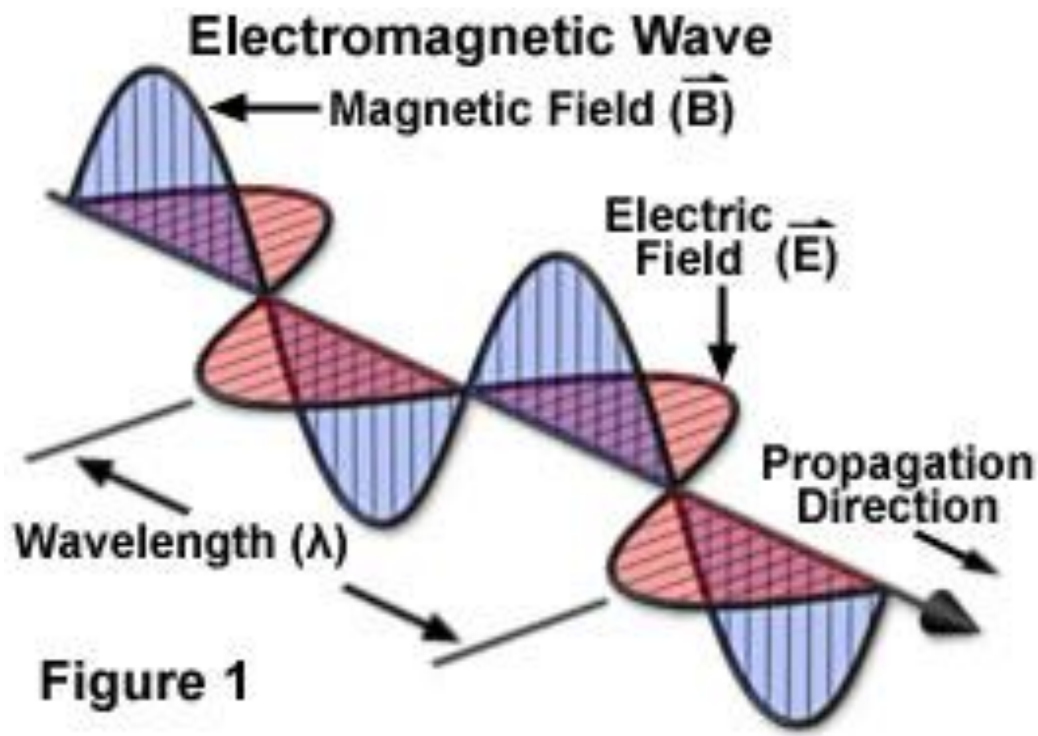
Chemistry 1104 Lab: Line Spectra

Goals:

1. **Introduction to Line Spectra of Atoms.**
2. **Compare Line spectrum of Hydrogen to known values and calculated values.**
3. **Determine the electron transitions observed in Hydrogen.**
4. **Examine the emission spectrum for various light sources and cationic elements.**

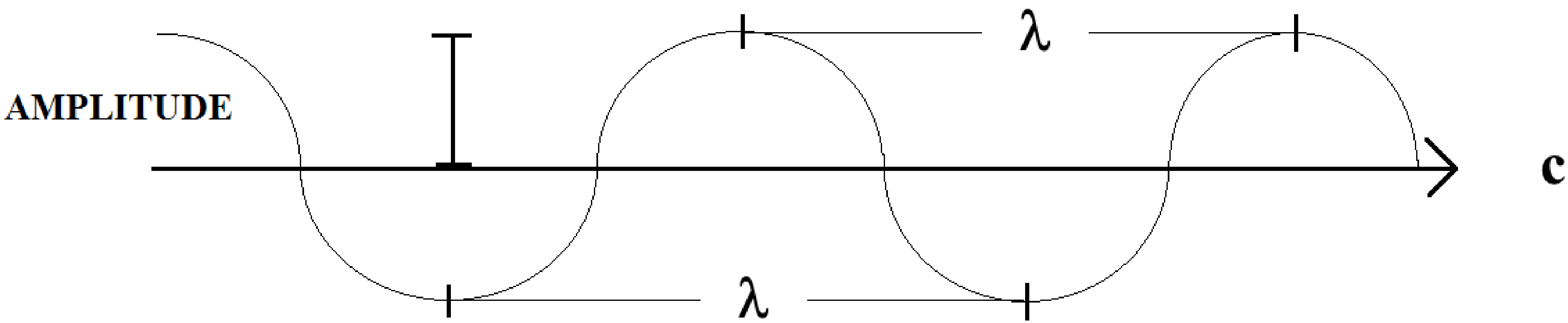
Electromagnetic Radiation:

Electromagnetic(EM) radiation is the transmission of energy in the form of a wave. Consists of an electric and magnetic component.



Wave Parameters of Measurement:

λ: wavelength(in metres or nanometers)
u: speed(in m/s)
c= speed of light in vacuum(3.00×10^8 m/s)
ν: frequency(in Hz or s^{-1})



In vacuum.

$$\lambda = \frac{c}{\nu}$$

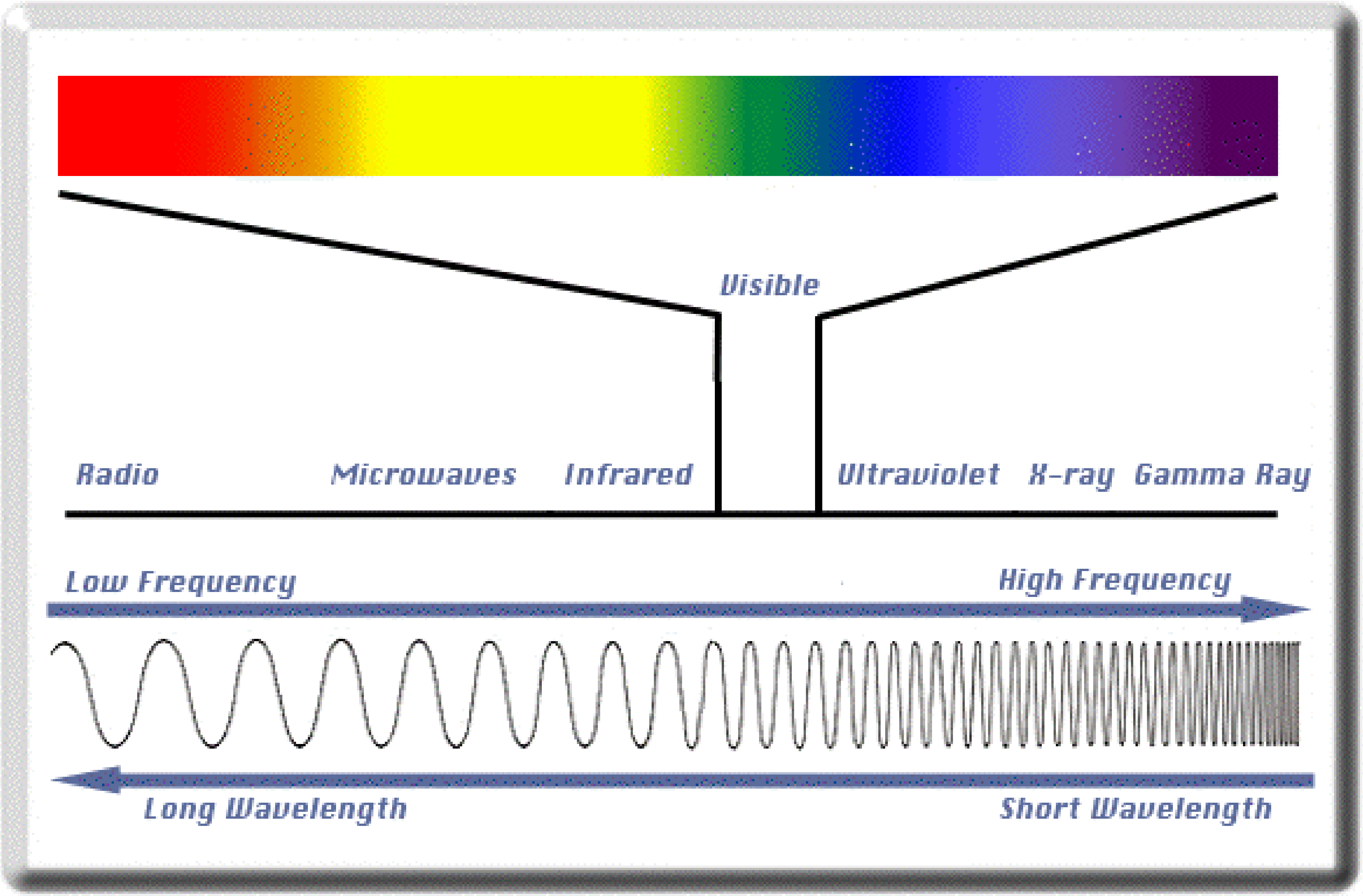
$$c = \lambda \cdot \nu$$

c = 3.00 ×10^{8} m/s

1 Å = 1×10^{-10} m

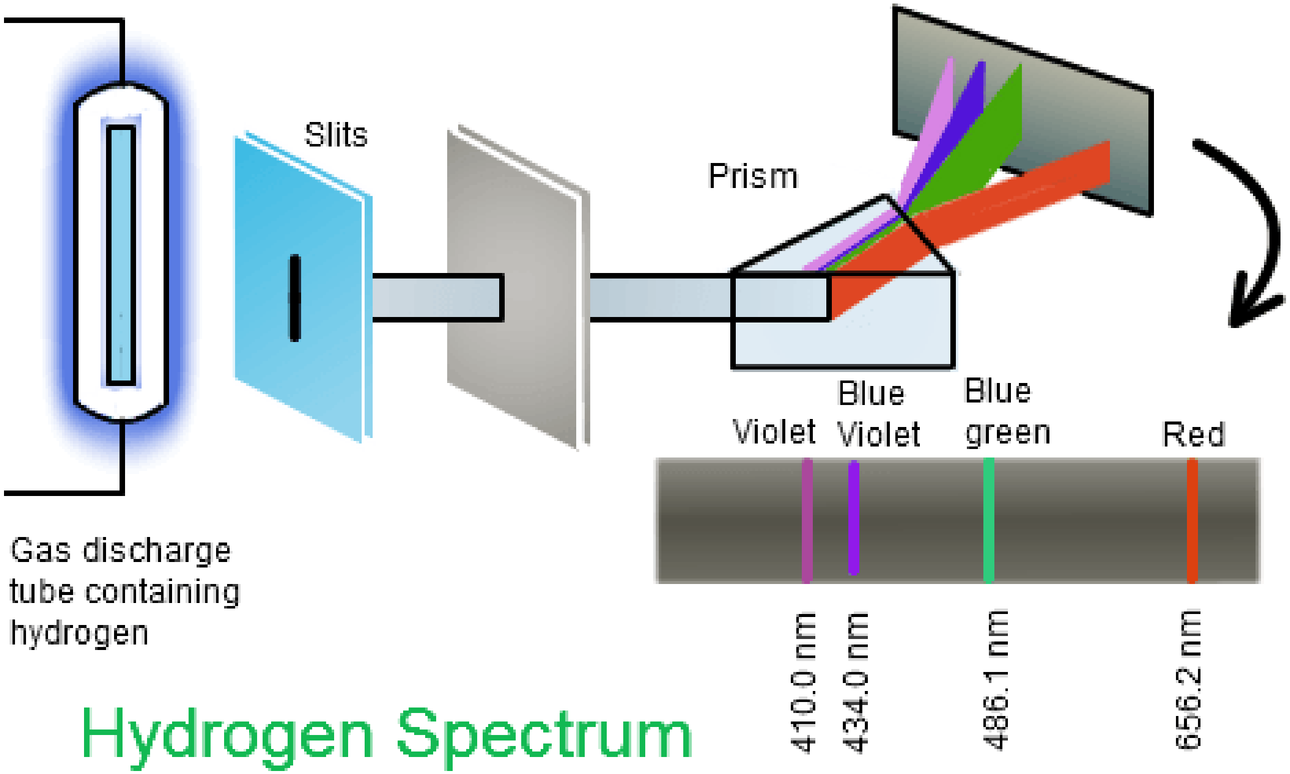
1 nm = 1×10^{-9} m

Electromagnetic Spectrum:



Atomic Spectra:

When elements are exposed to energy(heat, light, electrical energy) certain wavelengths of light are emitted and is known as a line spectrum.



Reference: https://chemistry.tutorvista.com/inorganic-chemistry/spectral-lines.html

Planck's Equation:

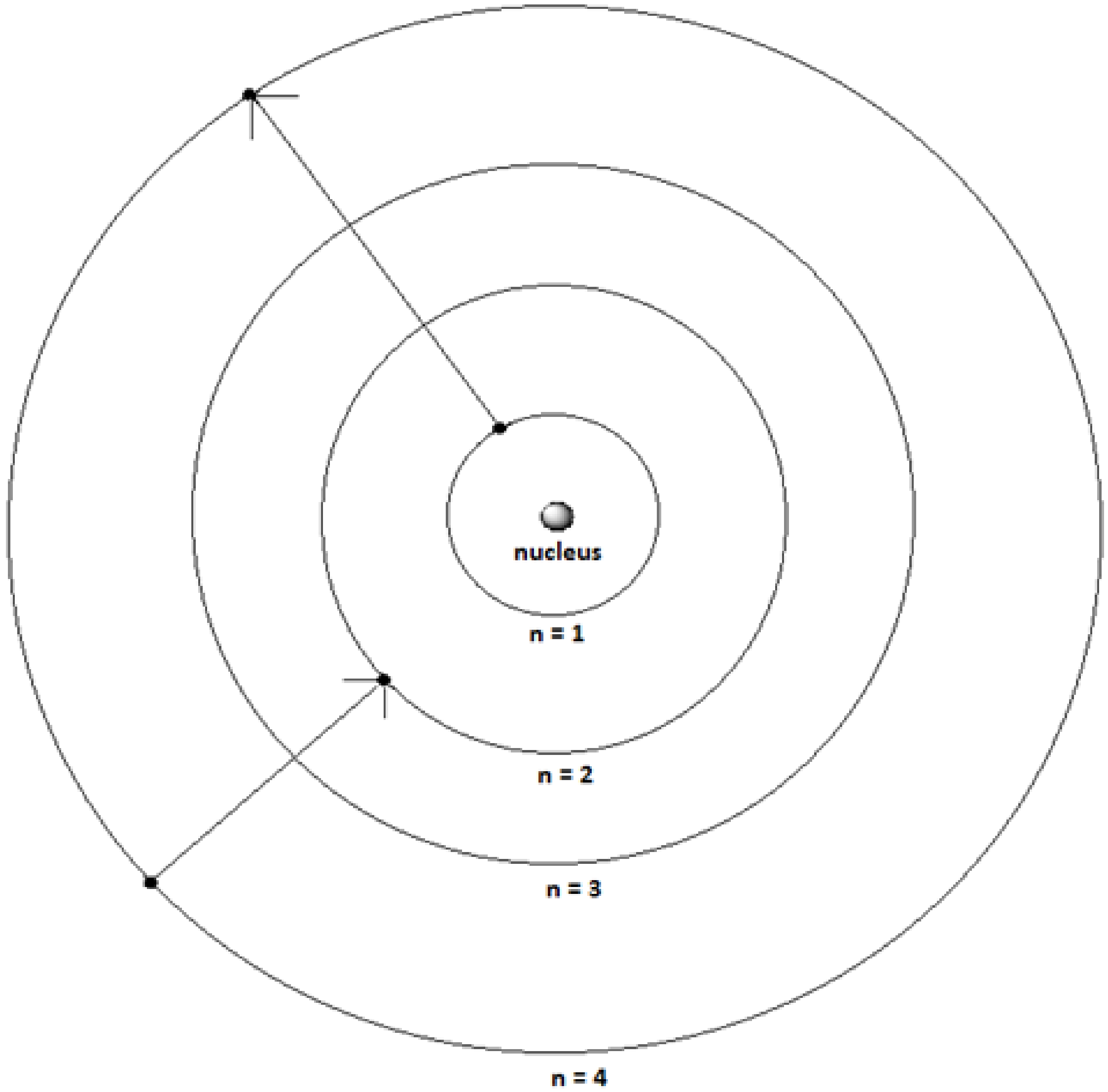
$$E = h\nu$$

E: energy(in joules)

h: Planck's constant(h = 6.626×10^{-34} J·s)

Bohr's Model of the Hydrogen Atom:

Atom absorbs energy. Electron excited from n = 1 to n = 4.



Electron moves from n = 4 to n = 2. Light emitted. 486 nm(Blue-green).

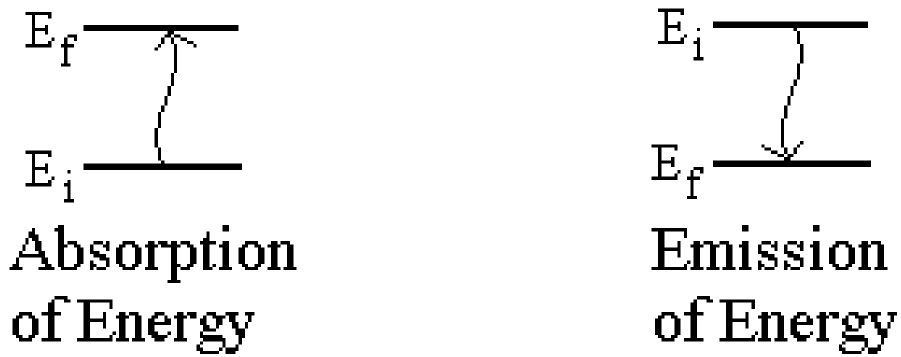
Bohr's Model of the Hydrogen Atom:

The electron travels around the nucleus in well defined fixed orbits.

Electrons in orbits closest to the nucleus are the lowest in energy.

For an electron to be excited from a low orbit to a high orbit it must absorb a specific amount of energy. When the electron goes from a high orbit to a lower orbit a specific amount of energy is given off.

Bohr's Model of the Hydrogen Atom:



$$\Delta E = R_H\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

ΔE: energy absorbed or emitted

R_H: Rydberg Constant(2.179×10^{-18}J)

Ex: Calculate the energy absorbed or given off when an electron travels from the n =5 to the n = 2 level.

$$\Delta E = R_H\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$\mathbf{R_H = 2.179\times10^{-18}J}$$

Ex: Calculate the wavelength of this light?

$$\lambda = \frac{\mathbf{hc}}{\Delta E}$$

Part A:

Observe and record the wavelengths of light given off by a Hydrogen Lamp.

Note Wavelength measured in spectroscope. Scale 4 to 7 corresponds to 400 nm to 700 nm.

Determine line spectrum for Hydrogen and one other element.

Part B:

Certain ionic compounds give off certain colors when excited.

Color Absorbed	Color Observed	Wavelenght Range Absorbed(nm)
Violet	Yellow-Green	400-435
Blue	Yellow	435-480
Green-Blue	Orange	480-490
Blue-Green	Red	490-500
Green	Purple	500-560
Yellow-Green	Violet	560-580
Yellow	Blue	580-595
Orange	Green-Blue	595-605
Red	Blue-Green	605-750

Part B: cont

Record color of flame produced when certain ionic compounds are aspirated into an open flame.

Determine emission line for cations.

CAUTION: Open flame. Danger of Burns.

Part C:

Compare the light spectrum for fluorescent light to that of sunlight. Note the differences.

Part D:

Calculate the energy in Joules(J) and wavelength in nanomoeters (nm) for transition of an electron in Hydrogen.

Complete Table 3 and 4.

Part E:

Compare your experimental observations from Part A for Hydrogen to the calculated results for Part D.

Determine the electron transition causing the observed line spectrum for hydrogen.