Unit Overview
Quantum Theory & Periodic Properties

Quantum theory seeks to describe the behavior and properties of small objects (electrons) just like classical physics does this for large objects (people, rockets, planets).

We will see that the quantum description of the electronic distribution of atoms helps us understand the arrangements of the elements on the periodic table and, ultimately, the physical and chemical properties they possess.

Lesson Overview
Introduction to Quantum Theory

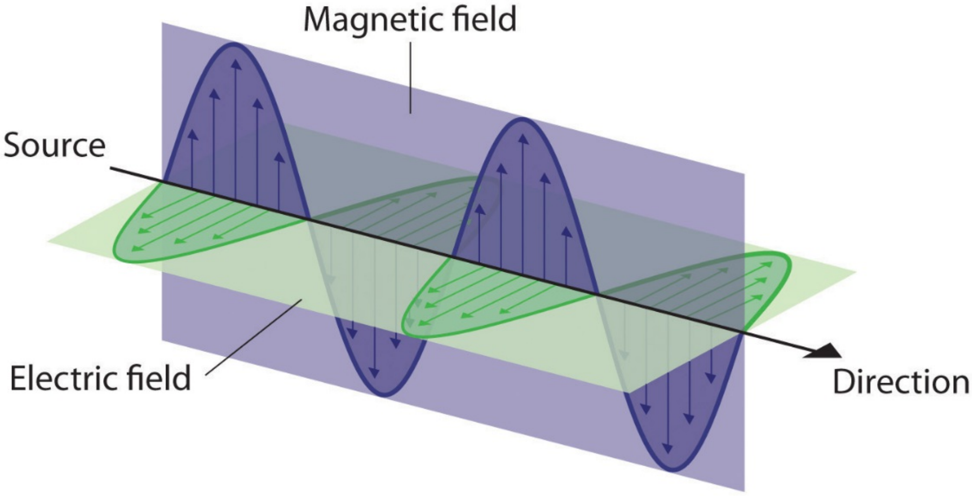
Objective: The student will be able to (1) define the wave nature of light and (2) explain how this definition lead to the development of a model for electrons in an atom.

Lesson Outline:

I. The Wave Nature of Light
II. Quantized Energy and Photons
III. Line Spectra and the Bohr Model

The Wave Nature of Light

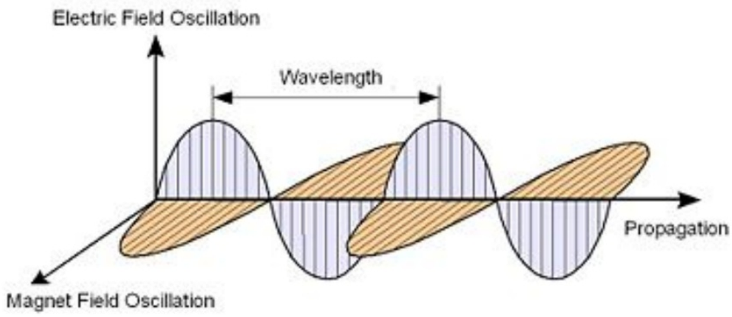
The portion of light we can see is called visible light and makes up only a small portion of electromagnetic radiation.



The Wave Model

There are many types of electromagnetic radiation: radiowaves, infrared, X-rays, etc. but they all share the same fundamental characteristics:

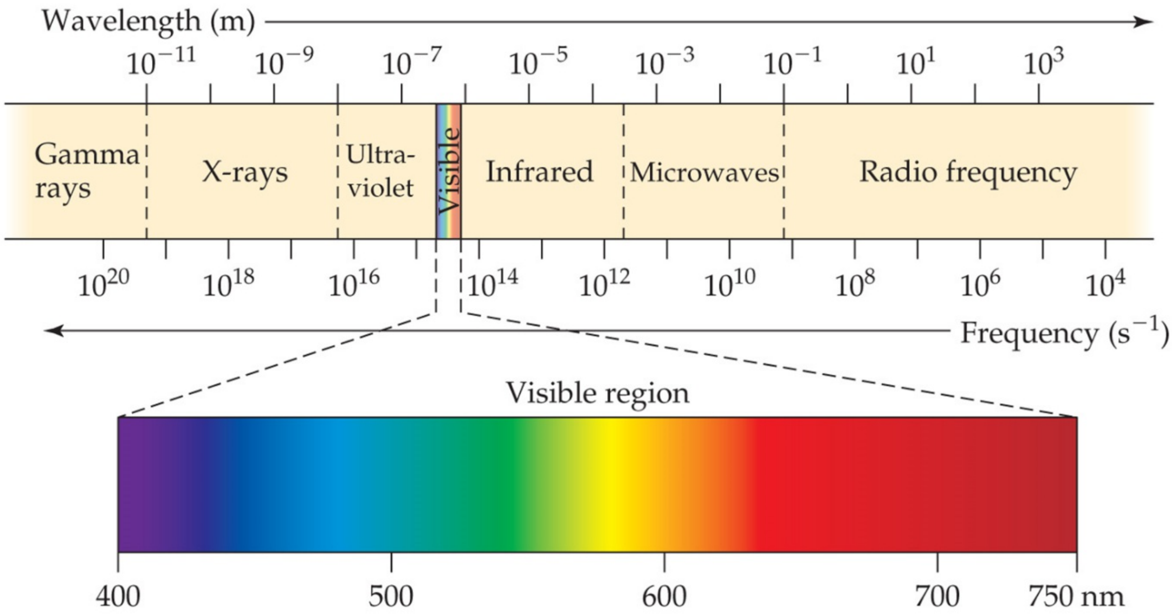
1. **in a vacuum they have the same speed = 3.00×10^8 m/s**
2. **wavelength**
3. **frequency**



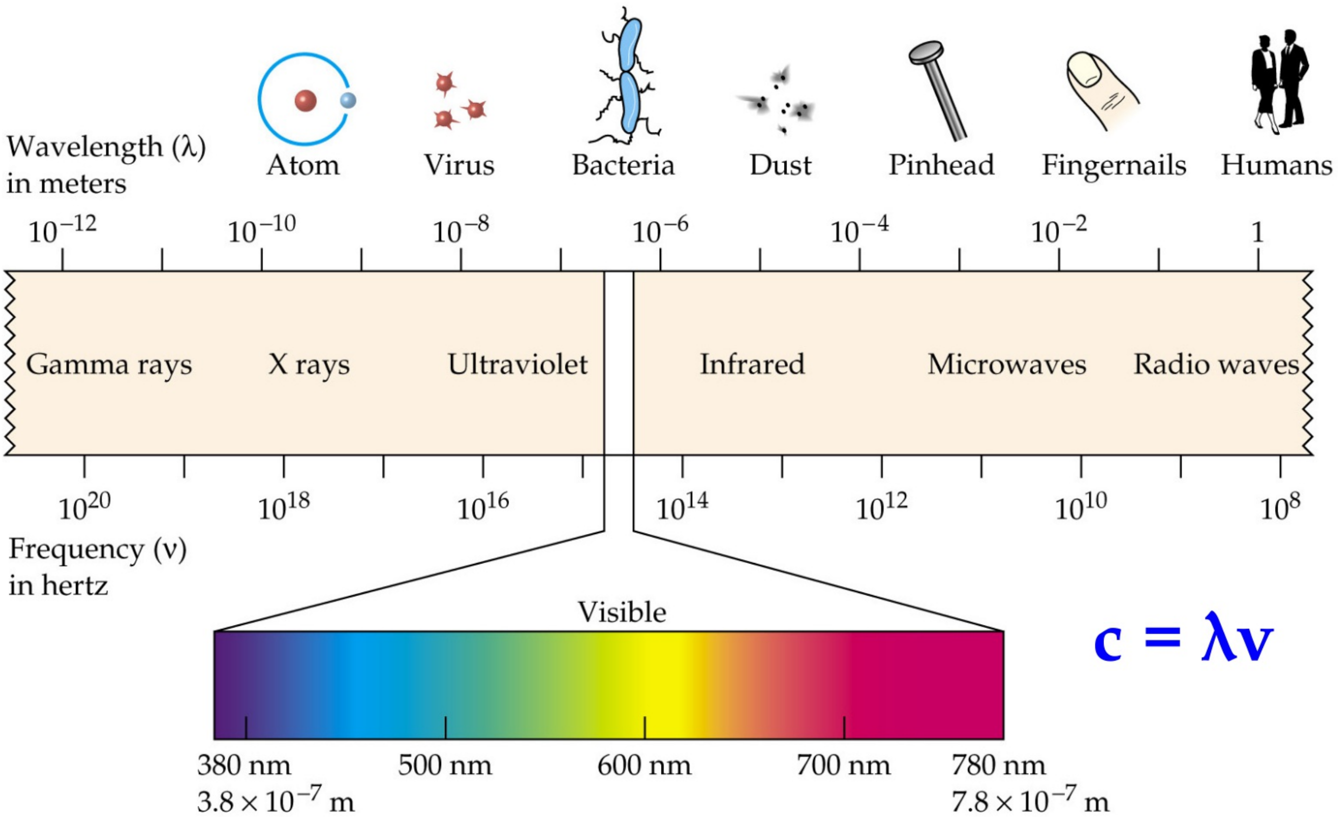
Relationship between wavelength and frequency: indirect

The Electromagnetic Spectrum

They have different wavelengths! This means they have different oscillations in the intensities of the electric and magnetic fields associated with the various types of radiation.



The EM Spectrum:
A visual representation



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

Sample Problem

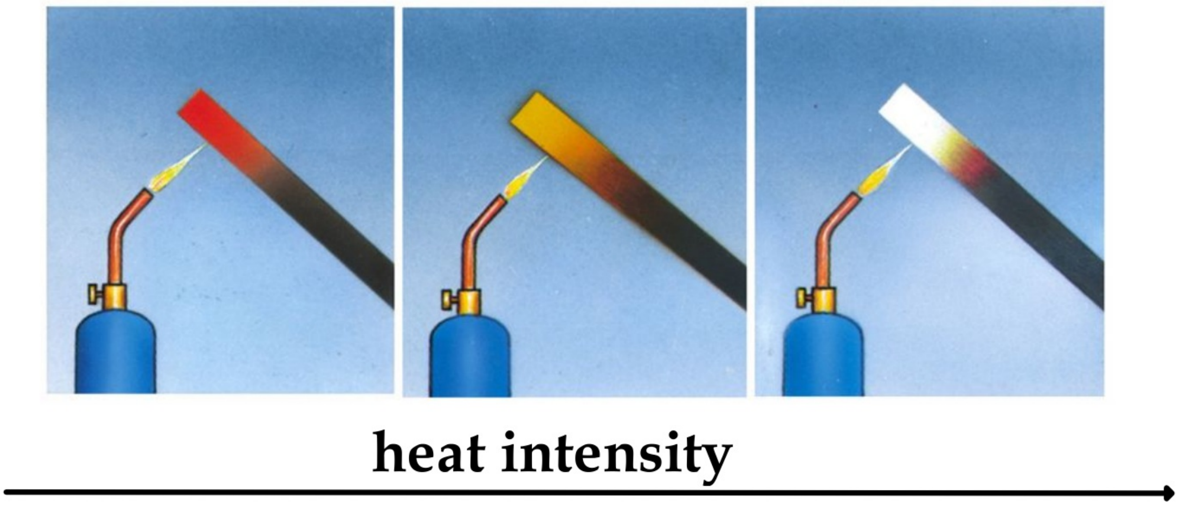
The yellow light given off by a sodium vapor lamp used for public lighting has a wavelength of 589 nm. What is the frequency of this radiation?

Quantized Energy and Photons

The wave model has its limitations, specifically it does not (and cannot) address these issues:

1. **the emission of light from hot objects (black body radiation)**
2. **the emission of electrons from metal surfaces on which light shines (photoelectric effect)**
3. **the emission of light from electronically excited gas atoms (emission spectra)**

Hot Objects & Blackbody Radiation



The laws of the late 1800s physics could not explain this phenomenon.

Max Planck (1900)

Solved the problem of blackbody radiation by assuming that energy can either be released or absorbed by atoms in only discrete "packages" of some minimum size (quantum) as EM Radiation.

Calculated this proportionality constant:

$E = h\nu$

where h is Planck's constant: 6.626×10^{-34} J·s

Quantized Energy

Potential energy of person walking up steps increases in stepwise, quantized manner

Potential energy of person walking up ramp increases in uniform, continuous manner

Imagine viewing the person moving up the stairs from a higher vantage point

Why do energy changes for large systems seem continuous (e.g. heating up a piece of metal)?

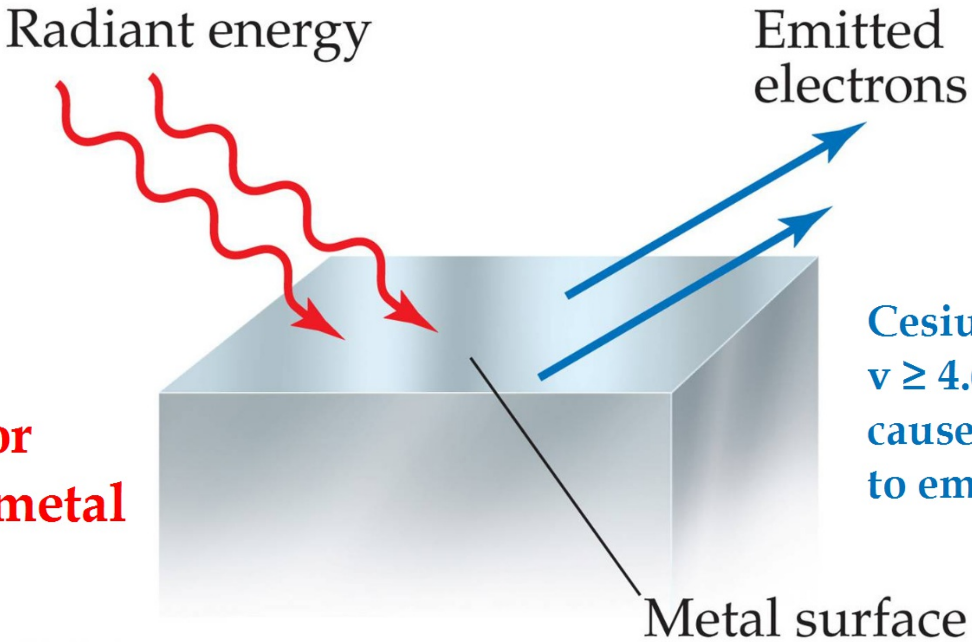
Energy changes for these systems go completely unnoticed!

Sample Problem

Calculate the energy of one photon of yellow light that has a wavelength of 589 nm.

The Photoelectric Effect

Another problem classical physics couldn't explain: light shining on a clean metal surface emits electrons

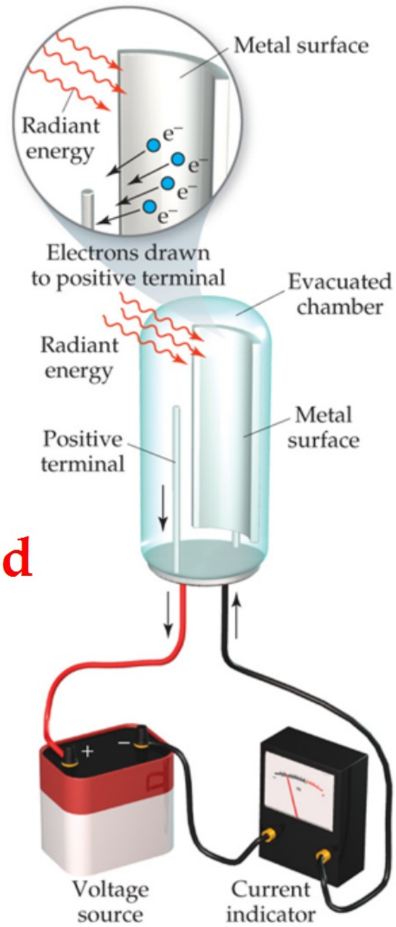


diff. for every metal

Cesium:
$\nu \geq 4.60\text{x}10^{14}$ Hz causes electrons to emit

Einstein (1905)

Used Planck's theory to explain the phenomenon of the photoelectric effect.

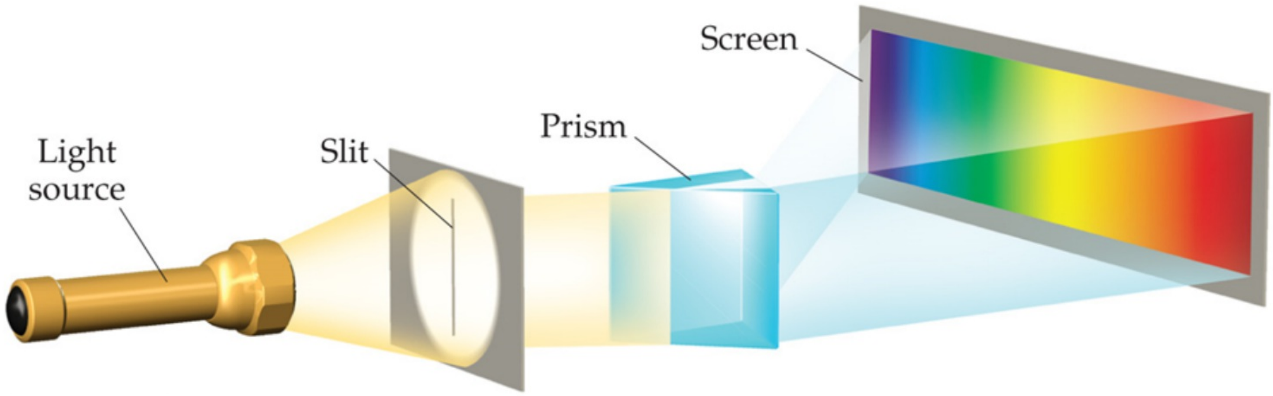


Conclusion: radiant energy is quantized

i.e. radiant energy is only allowed to have certain values proportional to Planck's constant.

Continuous Spectra

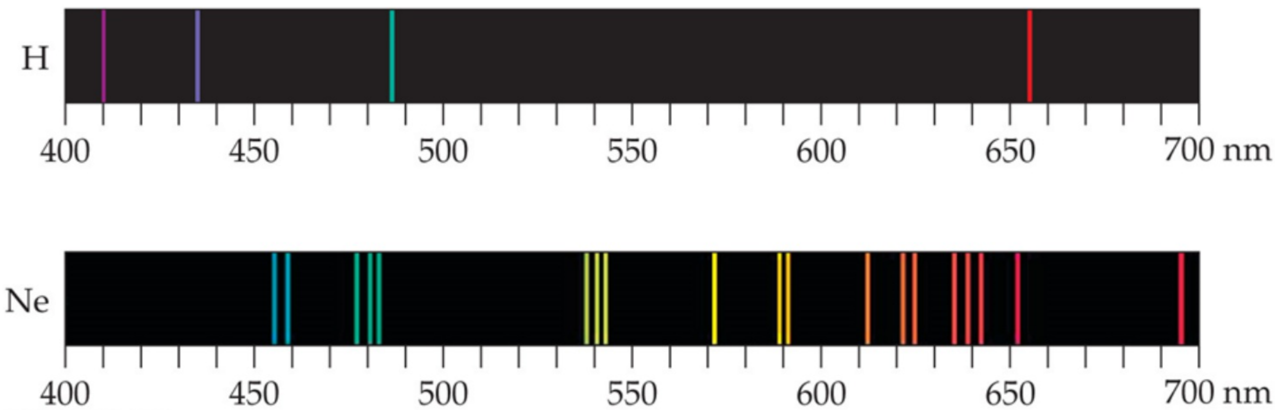
The third, and final, phenomenon scientists could not explain were line spectra.



We know that white light is composed of all colors and produces a continuous spectrum when sorted by a prism.

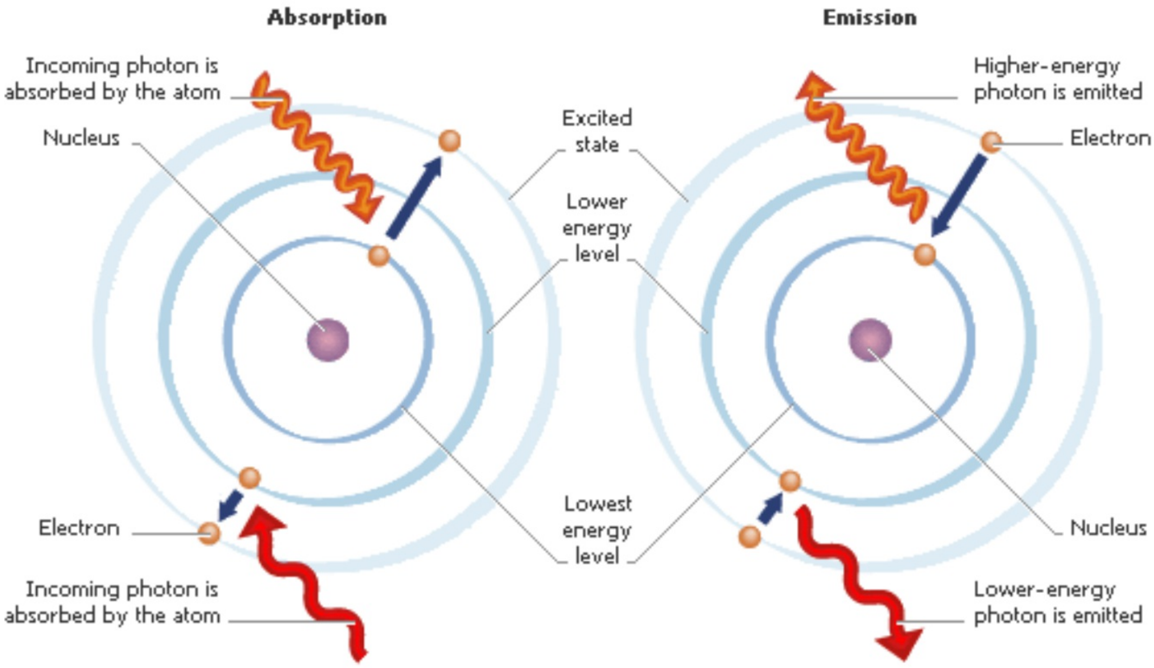
Line Spectra

Not all sources produce a continuous spectrum:



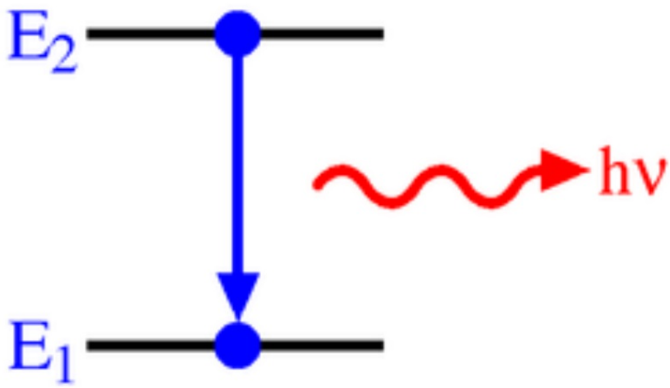
Scientists attempted to find an equation which would explain the values presented in these spectra. They believed it would shed some light on the location of electrons.

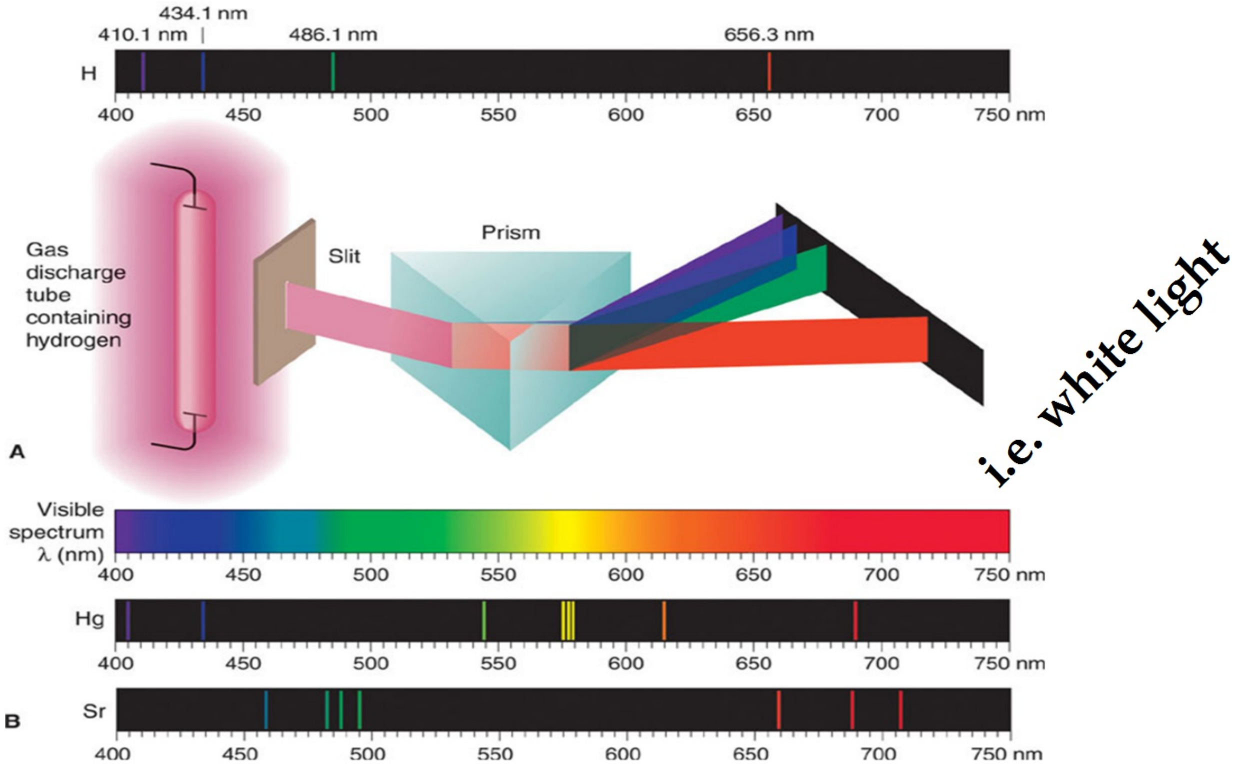
Absorption spectrum: The spectrum produced when atoms absorb specific wavelengths of incoming light and become excited from lower to higher energy levels.



Emission spectrum: The line spectrum produced when excited electrons return to lower energy levels and emit photons characteristic of that element.

Emission vs. Absorption





An atomic spectrum is not continuous because the atom's energy is not continuous, but rather has only certain states.

Johannes Balmer

Derived a very specific equation which described the visible line spectra of hydrogen.

His equation lead to the development of a more general equation which described all types:

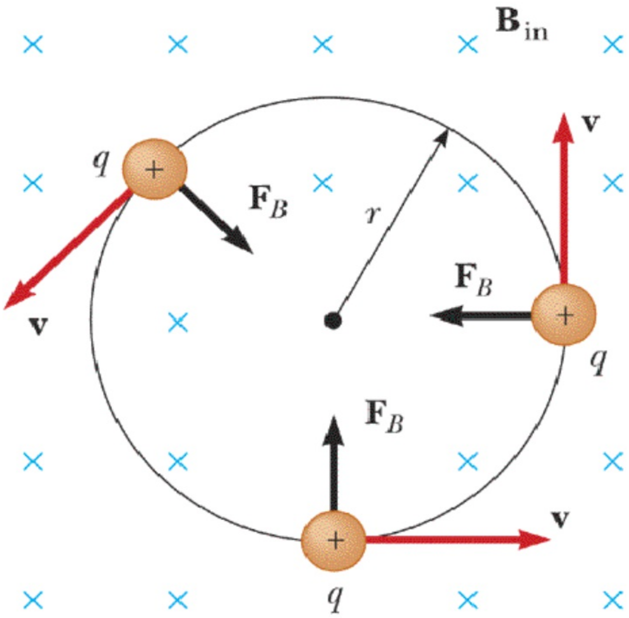
The Rydberg Equation (2 forms)

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

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

Physics Interruption

Classical physics indicates that a charged particle moving in a circle would continuously lose energy.



The Bohr Model of the Atom

Bohr was among many scientists searching for an answer to the configuration of electrons outside of the nucleus.

One major problem!

The Bohr Model of the Atom

Bohr formulated three postulates:

1. **Only orbits of certain radii, corresponding to certain specific energies, are permitted for the electron in a hydrogen atom**

2. **An electron is in a permitted is in an allowed energy state. It does not radiate energy and, therefore, does not spiral into the nucleus.**

3. **Energy is emitted or absorbed by the electron changes from one allowed energy state to another in the form of E=hv**

The Bohr Model of the Atom
& its Limitations

Limitations of the Model:

1. **Cannot explain the line spectra of other atoms.**

2. **Avoided the problem of the spiraling of the nucleus by simply assuming it did not happen.**