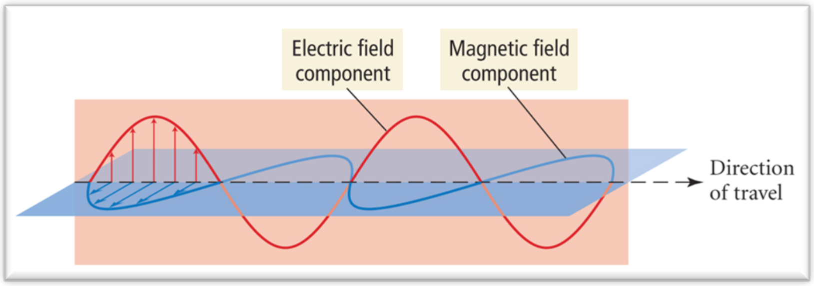
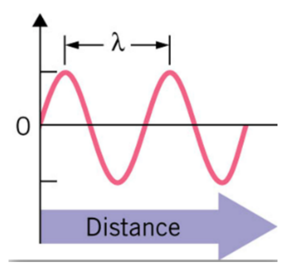
Module-7: Electromagnetic Energy and the Bohr Model of Atom

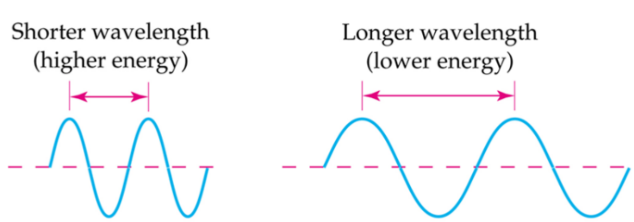
- The electromagnetic radiation composed of perpendicular oscillating waves, one for the electric field and one for the magnetic field.



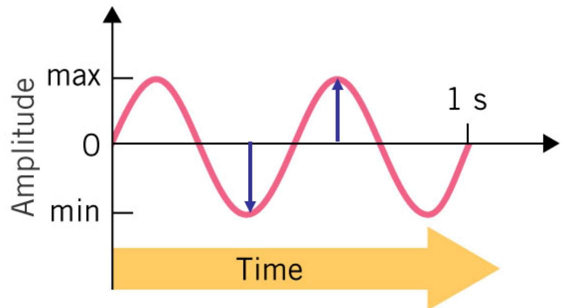
- Electromagnetic waves move through space at the same constant speed of 3.00×10^8 m/s. This velocity is represented by 'c'.
- The peak-to-peak distance is called the **wavelength.** The wavelength is represented by the symbol λ.



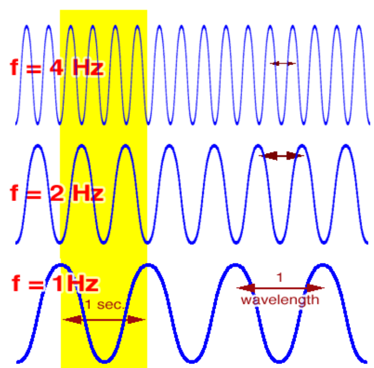
- Wavelength is usually expressed in units of meters, centimeters or **nanometers** (1 nm = 10^{-9} **m)**
- Energy and wavelengths are inversely proportional to one another. Blue light (high energy) has shorter wavelengths and red light has longer wavelengths (low energy).



Amplitude is the vertical distance from the midline of a wave to the peak or trough. Amplitude affects the intensity or the brightness of the radiation.



Frequency (ν) is the number of waves that pass any reference point per unit of time or waves/time. Frequency is measured in **hertz (Hz).** 1 Hz = 1cycle s^{-1}



The frequency can be calculated from the velocity and the and the wavelength of the radiation.

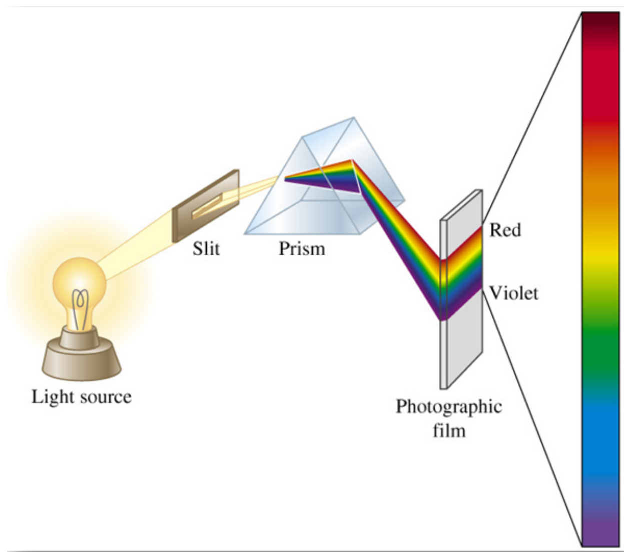
$$\boxed{\nu = \frac{c}{\lambda}}$$

In 1900 the German scientist Max Planck proposed that the electromagnetic radiation could be viewed as a stream of tiny energy packets or **quanta,** we now call **photons.**

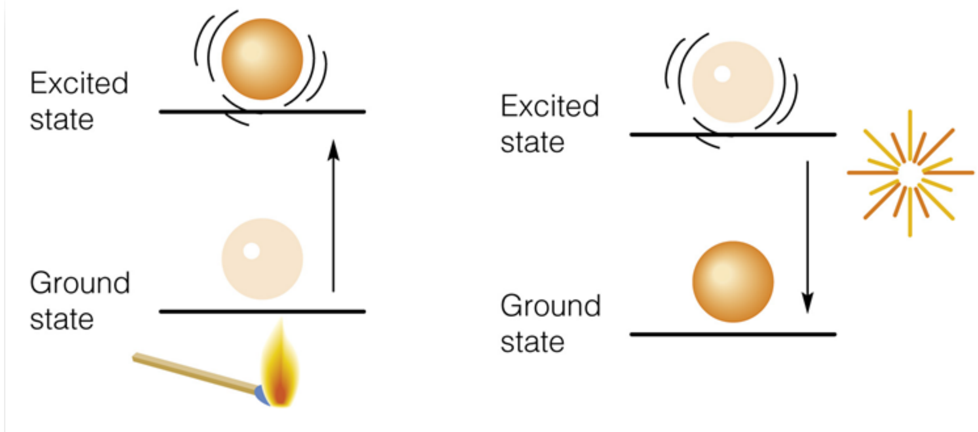
- Planck proposed, and Einstein confirmed, that *the energy of a photon is proportional to its frequency,* **E = h** ν. Where is h is the Plank's constant and it equals = 6.626×10^{-34} J s.
- This means that **both** electrons and electromagnetic radiation can be represented as either **waves** or **particles.**

THE LINE SPECTRA

- The **visible spectrum** of white light is called a **continuous spectrum**, because it contains continuous distribution of all colors.

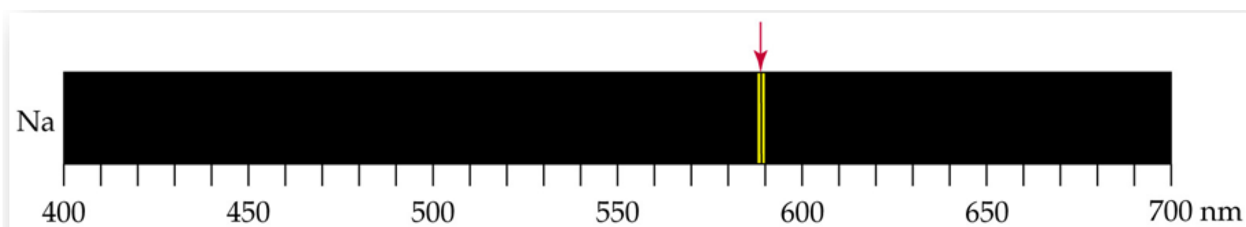


- Atoms when exposed to electromagnetic radiation, they emit the absorbed energy in the form of light as electrons return to a lower energy state.

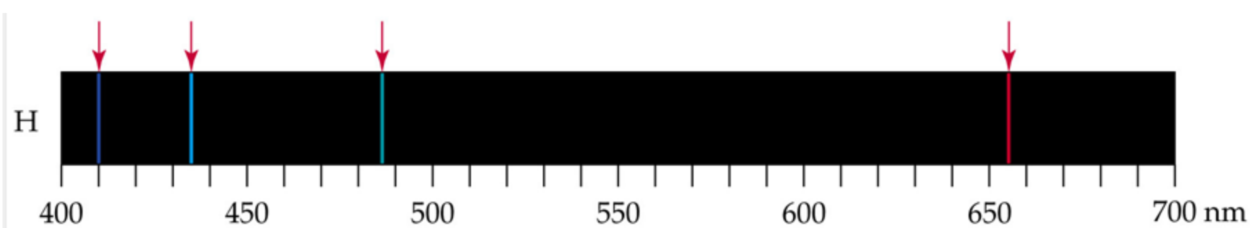


(a) Absorption of energy by an electron within an atom (b) Emission of energy by an electron within an atom

- When the white light is replaced by the light emitted by energetically excited elements as the source, we get a different kind of spectra namely the line spectra. The visible line spectrum of energetically excited sodium atoms consists of a closely spaced pair of yellow lines

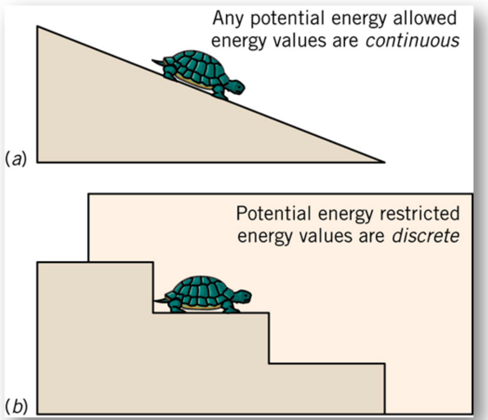


- The visible line spectrum of excited hydrogen atoms consists of four lines, from indigo at 410 nm to red at 656 nm. Note, that this is not a continuous spectrum.



- Atomic line spectra of sodium and hydrogen tell us that *electronic energy levels are not continuous.*

- The line spectra are possible because electrons are restricted to certain **energy levels**. The energy of the electron is said to be **quantized.**

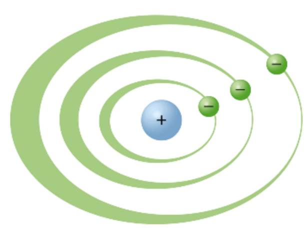


(a) Continuous energy level

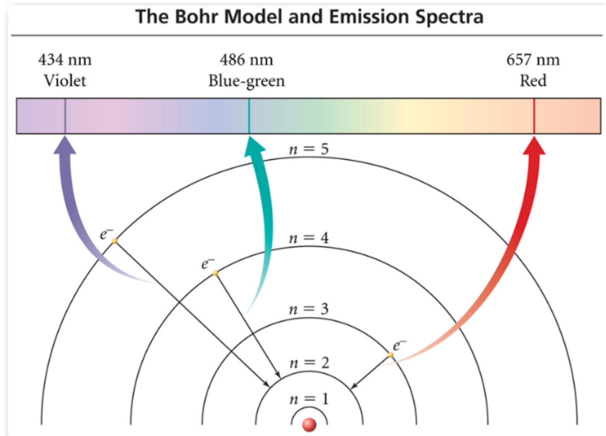
(b) The energy level are quantized.

BOHR MODEL OF ATOM

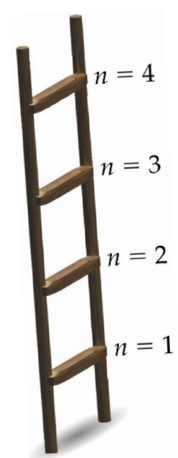
- The first theoretical model that successfully accounted for the Rydberg equation was proposed in 1913 by the Danish physicist **Niels Bohr**.
- Bohr proposed that the electrons moved around the nucleus is fixed paths or orbits much like the planets move around the sun



- Electrons emit radiation when they "jump" from an orbit with higher energy down to an orbit with lower energy. The energy gap between the orbits determines the energy of the photon of light produced.



Bohr orbits are like steps in a ladder. It is possible to be on one step or another, but it is impossible to be between steps. The analogy is limited, in that the distance between the steps is constant, from step to step, but the difference in the energy of the Bohr orbits varies by the orbit's quantum number.



In general, the line spectrum of an element is rather complicated. The line spectrum of hydrogen, with a single electron, is the simplest. The **Rydberg equation** can be used to calculate all the spectral lines of hydrogen.

$$\boxed{\frac{1}{\lambda} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)}$$

Where, λ = wavelength; n_1 and n_2 are represent electronic energy levels and they are positive integers, n_2 must be larger than n_1. The Rydberg constant, R, is an empirical constant with a value of 109,678 cm^{-1}.

Practice Problems

1. A clean metal surface is irradiated with light of three different wavelengths λ_1, λ_2 and λ_3. The kinetic energies of the ejected electrons are as follows: λ_1: 2.9×10^{-20} J; λ_2: 8.1×10^{-19} J; λ_3: 1.2×10^{-20} J. Which light has the shortest wavelength and which has the longest wavelength?

2. The average distance between Mars and Earth is about 1.3×10^8 miles. How many minutes it will take for a signal traveling at the velocity of light to reach the Earth from the spacecraft, '*Mars Rover Curiosity*'.

3. What is the wavelength, in *nm*, of radiation that has energy of 1.0×10^3 kJ/mol? In which region of electromagnetic spectrum this radiation is found?

4. The blue color of sky results from scattering of sunlight by air. The blue light has frequency of about 7.5×10^{14} Hz. (a) calculate the wavelength, in *nm*, associated with this radiation, and (b) calculate the energy, in joules, of a single photon associated with this energy.

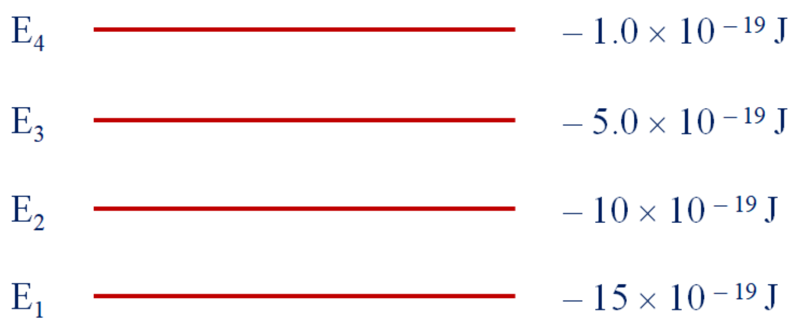
5. One of the popular radio stations in Philadelphia is KYW 1060. This radio station broadcasts at 1060 MHz frequency.

    (a) convert the frequency to wavelength in meters?

    (b) what is the energy in joules for one quantum of this broadcast signal?

    (c) to what region of electromagnetic spectrum would you assign this signal?

6. Consider the following energy levels of a hypothetical atom:



(a) What is the wavelength of the photon needed to excite an electron from E_1 to E_4?

(b) What is the energy in joules a photon must have in order to excite an electron from E_2 to E_3?

(c) When an electron drops from E_3 to E_1, calculate the wavelength of the photon emitted in this process.

7. Calculate the frequency (in Hz) and wavelength (in *nm*) of the emitted photon when an electron drops from $n = 4$ to the $n = 2$ level in a hydrogen atom?