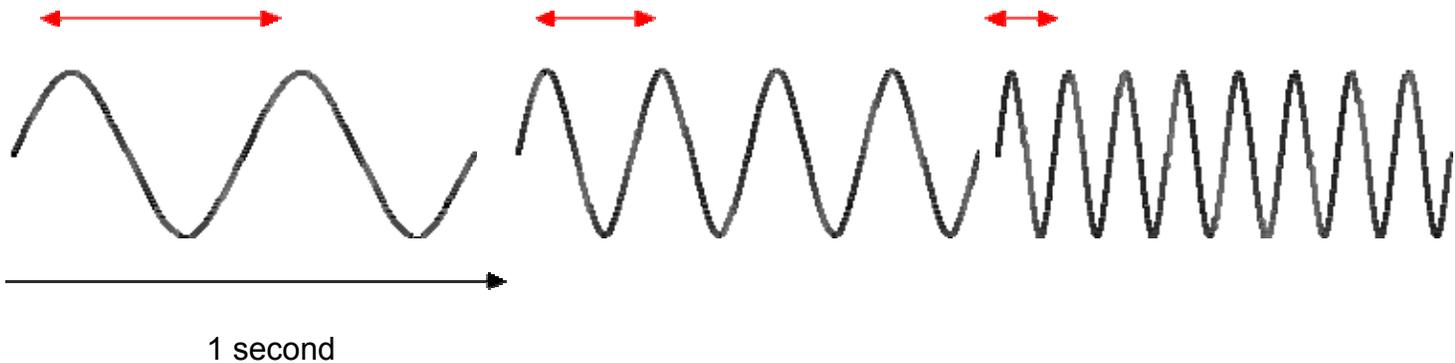


## Worksheet 10 - Electromagnetic Radiation and the Bohr Atom

Light is known to have the **wave-like** properties of **frequency** ( $\nu$ ) and **wavelength** ( $\lambda$ ). These are illustrated below. The x-axis is a measure of time. The distance between the peaks is called the **wavelength** and the number of waves per unit time (1 second in this example) is called the number of **cycles**. The first wave pattern has **2 cycles per second**, the middle example has **4 cycles per second** and the example on the right has **8 cycles per second**. This is the **frequency** of the wave, and has the units of **hertz, Hz** (cycles/s).



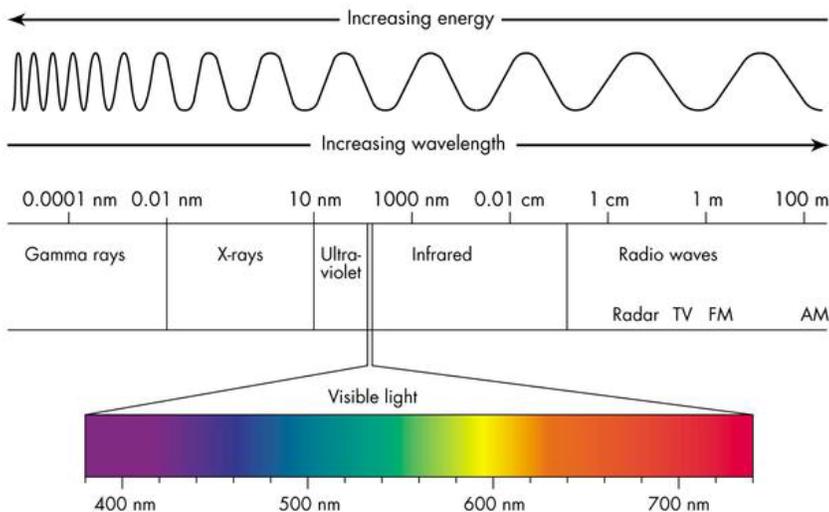
As the frequency increases, the wavelength decreases.

In **Electromagnetic radiation** (light) these are related by the equation:

$$c = \lambda\nu$$

where **c** = the speed of light,  $2.998 \times 10^8$  m/s,  $\lambda$  = wavelength (m) and  $\nu$  = frequency ( $s^{-1}$  or Hz). The electromagnetic spectrum (EMS) is shown below.

Which color of visible light has the shortest wavelength? Which radiation has wavelengths longer than visible light?



1. The wavelength of green light is about 522 nm. What is the frequency of this radiation?
2. What is the wavelength of a photon that has a frequency of  $2.10 \times 10^{14}$  Hz? Answer in nm and determine what type of radiation this is.

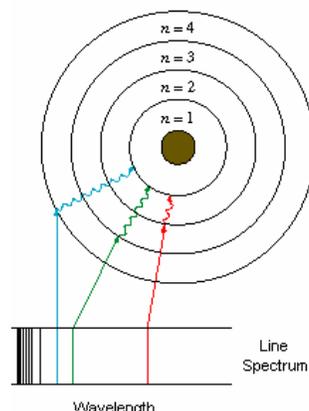
**Planck** recognized that energy is **quantized** and related the energy of radiation (emitted or absorbed) to its frequency.

$$\Delta E = n h \nu$$

where **n** = integer and **h** = Planck's constant =  $6.626 \times 10^{-34}$  J s

3. Which of the following are **directly** related?
  - a) energy and wavelength
  - b) wavelength and frequency
  - c) frequency and energy
4. A classical radio station broadcasts at 93.5 MHz ( $M = 10^6$ ). Find the wavelength of this radiation, in meters, and the energy of one of these photons, in J. What type of radiation is this?
5. What is the energy of a photon with:
  - a) a wavelength of 827 nm? What type of radiation is it?
  - b) a wavelength of 1 nm? What type of radiation is it?

**Bohr** applied this concept to the **line spectra** of elements. When elements are excited they emit radiation at fixed wavelengths. He proposed that only certain **energy levels** are allowed within the structure of an atom. Electrons are allowed to move between these energy levels. The light emitted by the elements is a measure of the energy gap between the two electronic states.



For the hydrogen atom,

$$\Delta E = -R_H Z^2 \left[ \frac{1}{n_f^2} - \frac{1}{n_i^2} \right] \quad R_H = \text{Rydberg constant} = 2.178 \times 10^{-18} \text{ J}$$

$Z = \text{nuclear charge} = 1 \text{ for H, } 2 \text{ for He}$

6. Calculate the  $\Delta E$  for the  $n = 4$  to the  $n = 2$  transition in hydrogen. Where on the EMS would this appear? What does the sign mean?

7. A hydrogen atom in its ground state absorbs light with a wavelength of 102.6 nm. Calculate the energy level of the resulting excited state ( $n = ?$ ).

**Ionization energy** is the energy required to **completely remove** an electron from an atom. This can be thought of as the transition between  $n = 1$  and  $n = \infty$ .

8. Calculate the energy needed to remove the electron from hydrogen in its ground state.

This is the energy to remove an electron from the ground state of hydrogen. What wavelength of light would work? Where is this on the EMS?

9. What is the energy needed to remove the remaining electron from  $\text{He}^+$  in its ground state? Is it easier or harder to remove the electron from  $\text{He}^+$  than from H?