Unit 3 Part 5: Electromagnetic Radiation (EMR)

Electromagnetic radiation: (EMR) is a form of energy that will behave in some ways like a wave, but in some ways it will behave like stream of particles.

- **Photon:** a particle of electromagnetic radiation having zero mass and carrying a specific amount of energy.

The relationship between energy, wavelength and frequency of any EMR wave can be expressed in 2 equations:

$$c = \lambda \nu \qquad \boxed{E} = h\nu$$

energy of 1 photon!

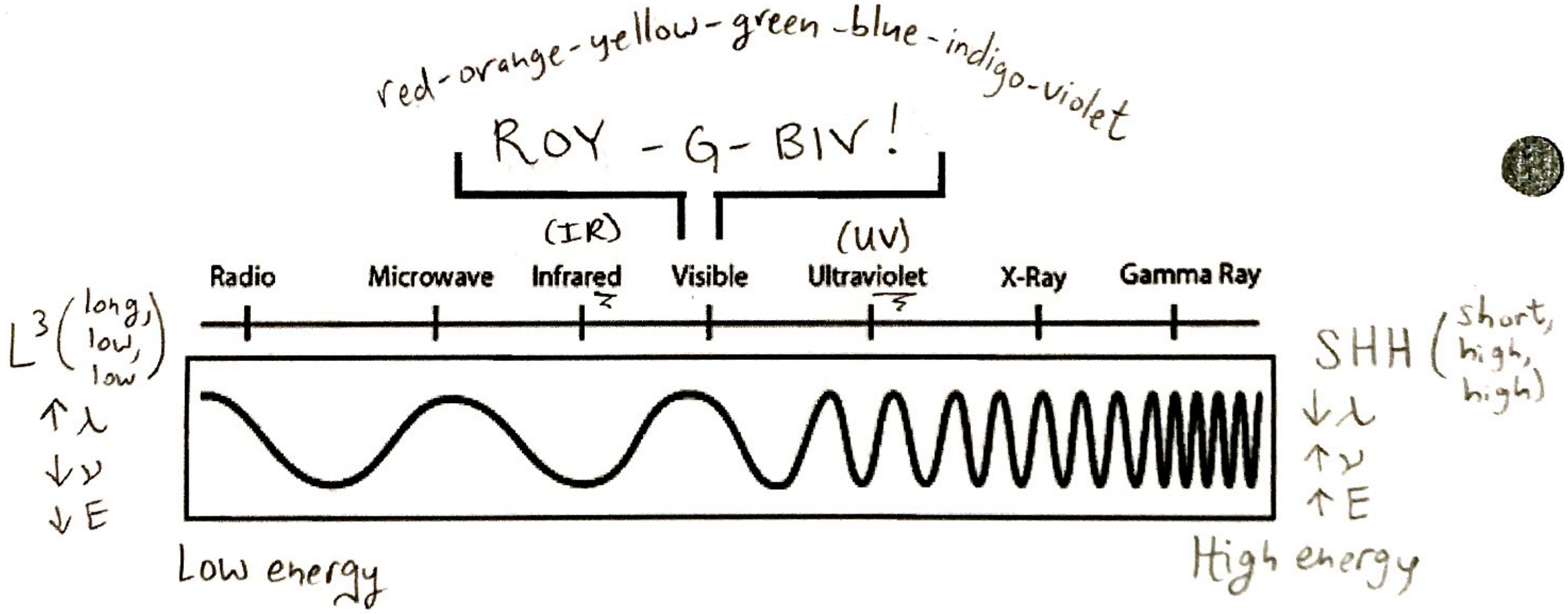
(J) $E = energy$

(Hz = 1/s) $\nu = frequency$

(m) $\lambda = wavelength$

Planck's constant, $h = 6.626 \times 10^{-34}$ J s

Speed of light, $c = 2.998 \times 10^{8}$ m s^{-1}



Guided Practice

1. What is the $\overset{E}{\text{energy}}$ of a photon with a wavelength of $\overset{\lambda}{\boxed{8.27 \times 10^{-7} \text{m?}}}$

a.) $E = h\nu$ and $\nu = \frac{c}{\lambda}$

$$E = \frac{hc}{\lambda} = \frac{(6.626E\text{-}34\ J\cdot s)(2.998\ m/s)}{8.27E\text{-}7m} = \boxed{2.40 \times 10^{-19}\ J}$$

b.) kJ/mol ?

$$\frac{2.40E\text{-}19\ J}{1\ photon} \times \frac{1\ kJ}{1000\ J} \times \frac{6.022E\ 23\ photons}{1\ mol} = \boxed{145\ kJ/mol}$$

2. In an experiment, a molecule of hydrogen, H_2, absorbs a photon of electromagnetic radiation with a wavelength of 300 nm. The bond energy of H_2 gas is 432 kJ/mol.
   a. Calculate the frequency of the photons in Hz (sec^{-1}).

$$\nu = \frac{c}{\lambda} = \frac{2.998\text{E}8\ \text{m/s}}{\underbrace{3.00\text{E-}7\text{m}}_{300\text{ nm}}} = \boxed{9.99 \times 10^{14}\ \text{Hz}} \underset{1/s}{\uparrow}$$

   b. Calculate the number of joules required to break the bond in a single molecule of H_2 gas.

$$\frac{432\text{ kJ}}{1\text{ mol}} \times \frac{1000\text{ J}}{1\text{ kJ}} \times \frac{1\text{ mol}}{6.022\text{E}23\text{ molec.}} = \boxed{7.17 \times 10^{-19}\text{ J (per molec. } H_2)}$$

   c. Does the photon have enough energy to break the bond in a molecule of H_2 gas? Mathematically justify your answer.

$$E = h\nu = (6.626\text{E-}34\ \text{J}\cdot\text{s})(9.99\text{E}14\ \tfrac{1}{s}) = 6.63 \times 10^{-19}\text{ J}$$

$$6.63\text{E-}19\text{ J} < 7.17\text{E-}19\text{ J}$$

$\Rightarrow$ $\boxed{\text{the photon does not have enough energy to break the bond!}}$

Independent Practice

3. A certain green light has a frequency of 6.26×10^{14} Hz.
   a. What is the energy of one photon of this light?

$$E = h\nu = (6.626\text{E-}34\ \text{J}\cdot\text{s})(6.26\text{E}14\ \tfrac{1}{s}) = \boxed{4.15 \times 10^{-19}\text{ J}}$$

   b. What is the energy of this light measured in kilojoules/mole?

$$\frac{4.15\text{E-}19\text{ J}}{1\text{ photon}} \times \frac{1\text{ kJ}}{1000\text{ J}} \times \frac{6.022\text{E}23\text{ photons}}{1\text{ mol}} = \boxed{250.\ \text{kJ/mol}}$$

4. What is the wavelength, in nanometers, of a photon of light that has a frequency of 2.10×10^{14} Hz?

$$\lambda = \frac{c}{\nu} = \frac{2.998\text{E}8\ \text{m/s}}{2.10\text{E}14\ \text{1/s}} = 1.43\text{E-}6\text{ m} \times \frac{1\text{E}9\text{ nm}}{1\text{ m}} = \boxed{1{,}430\text{ nm}}$$