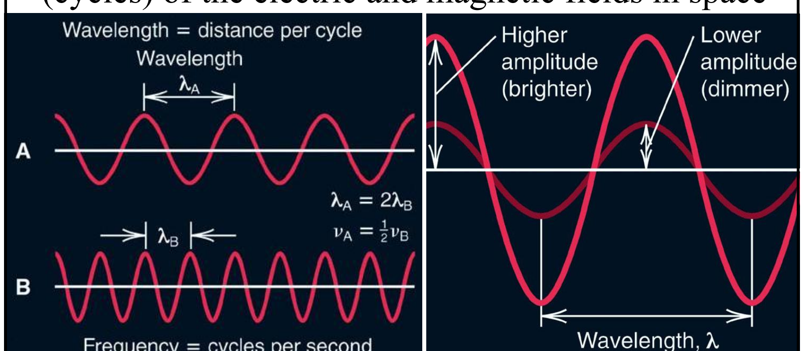
Quantum Theory and Atomic Structure

- **Nuclear atom** – small, heavy, positive nucleus surrounded by a negative electron cloud
- **Electronic structure** – arrangement of the electrons around the nucleus
- **Classical mechanics** – fails in describing the electronic motion
- **Quantum mechanics** – designed to describe the motion of microscopic particles

7.1 The Nature of Light

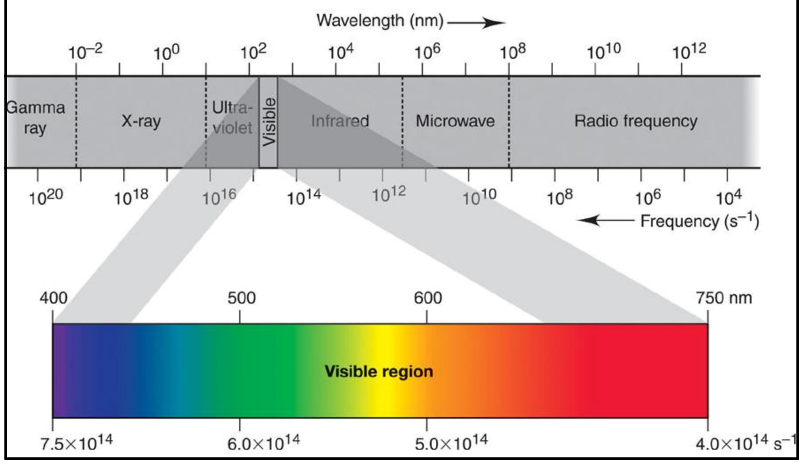
- Light is **electromagnetic radiation** – a stream of energy in the form waves
- **Electromagnetic waves** – periodic oscillations (cycles) of the electric and magnetic fields in space



- **Wavelength (λ)** – distance between two adjacent minima or maxima of the wave
- **Frequency (ν)** – number of oscillations of the electric (or magnetic) field per second
  - units - hertz (Hz) → **1 Hz = 1 s^{-1}**
- **Amplitude** – strength of the oscillation (related to the intensity of the radiation)
- **Speed of light (*c*)** – rate of travel of all types of electromagnetic radiation (3.00×10^8 m/s)

$$\lambda\nu = c \qquad \uparrow\lambda \rightarrow \downarrow\nu$$

- **Electromagnetic spectrum** – classification of light based on the values of λ and ν



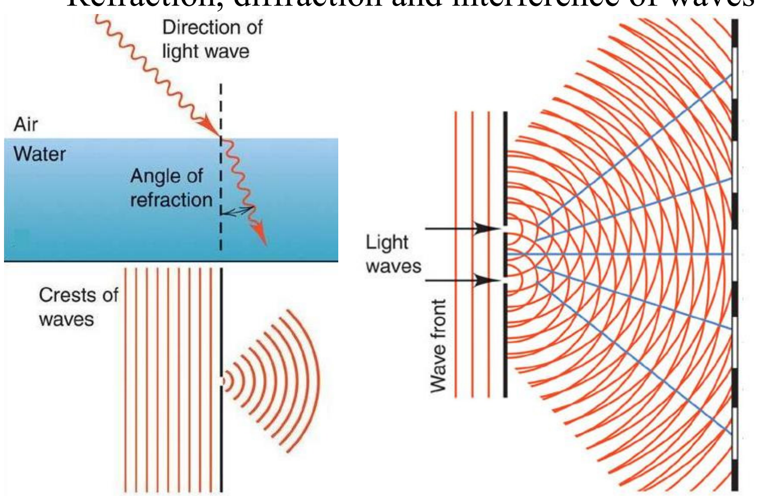
Example:

What is the wavelength of light with frequency **98.9 MHz**.

98.9 MHz = 98.9×10^{6} Hz = 98.9×10^{6} s^{-1}

$$\lambda = \frac{c}{\nu} = \frac{3.00\times10^{8}\ \text{m/s}}{98.9\times10^{6}\ \text{s}^{-1}} = 3.03\ \text{m}$$

- Differences between waves and particles
  - Refraction, diffraction and interference of waves



The particle nature of light

- **Blackbody radiation** – light emitted from solid objects heated to incandescence
  - The energy profile of the emitted light could not be explained by the classical mechanics which assumes that the energy of an object can be continuously changed
  - **Plank** (1900) explained the energy profiles by assuming that the energy of an object can be changed only in discrete amounts (**quanta**) → **quantization of energy**

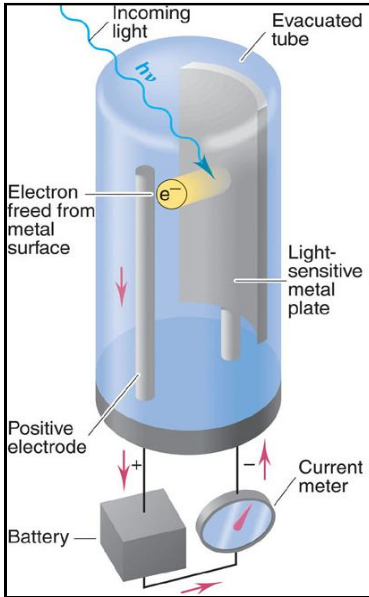
$$\Delta E = n(h\nu)$$

h – Planck's constant $\quad$ **$h = 6.626\times10^{-34}$ J·s**

ν – frequency of the emitted light

n – quantum number (positive integer – 1, 2, 3, …)

***hν* – the energy of one quantum**



- **Photoelectric effect** – ejection of e^- from metals by irradiation with light
  - Ejection of e^- begins only above a certain **threshold frequency** (below this frequency, no ejection occurs no mater how intense the light is)
  - Ejection of e^- begins with no time delay
  - Can't be explained by treating light as waves

- Explanation (**Einstein**, 1905) – the ejection of e^- is caused by particles (**photons**) with energy proportional to the frequency of the radiation
  - ⇒ Only photons with enough energy and therefore high enough frequency can eject electrons
  - ⇒ Ejection results from an encounter of an e^- with a single photon (not several photons), so no time delay is observed
- Energy of the photon (E_{ph}):

$$E_{ph} = h\nu \qquad \nu = c/\lambda \qquad E_{ph} = hc/\lambda$$

⇒ The photon is the **electromagnetic quantum** – the smallest amount of energy atoms can emit or absorb

- **Dual nature of light** – light has both wave and particle like properties
  - wave (refraction, interference, diffraction)
  - particle (photoelectric effect)

Example:

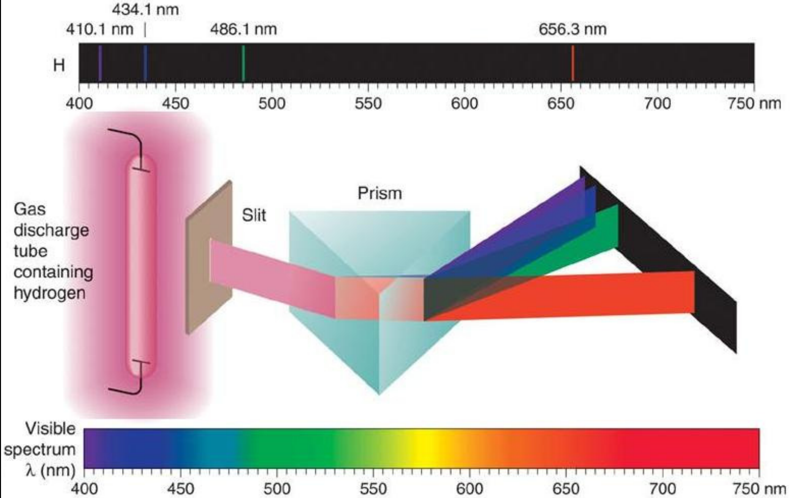
Calculate the energy of a photon of light with wavelength **514 nm**.

$$E_{ph} = h\frac{c}{\lambda} = 6.626\times10^{-34}\,\mathrm{J\cdot s}\,\frac{3.00\times10^{8}\,\mathrm{m/s}}{514\times10^{-9}\,\mathrm{m}} = 3.87\times10^{-19}\,\mathrm{J}$$

7.2 Atomic Spectra

- **Spectroscopy** – studies the interaction of light with matter (emission, absorption, scattering, …)
- **Spectrometer** – instrument that separates the different colors of light and records their intensities
- **Spectrum** – intensity of light as a function of its color (wavelength or frequency)
- **Atomic emission spectrum** – the spectrum emitted by the atoms of an element when they are excited by heating to high temperatures (very characteristic for each element; used for identification of elements)

- **Spectral lines** – images of the spectrometer entrance slit produced by the different colors in the spectrum

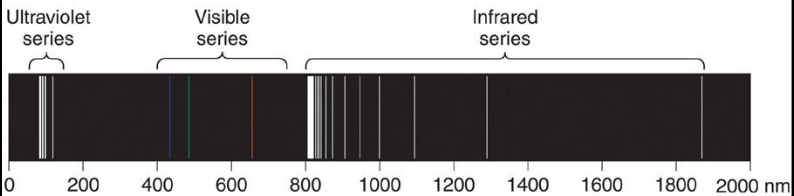


- Atomic emission spectra are **line spectra** – consist of discrete frequencies (lines)
  - Can't be explained by classical physics
- The **Rydberg equation** – fits the observed lines in the hydrogen atomic emission spectrum

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

n_1, n_2 – positive integers (1, 2, 3, ...) and $n_1 < n_2$

R – the **Rydberg** constant (1.096776×10^7 m^{-1})



- Lyman series (UV) – n_1 **= 1** and n_2 **= 2, 3, 4, ...**
- Balmer series (VIS) – n_1 **= 2** and n_2 **= 3, 4, 5, …**
- Paschen series (IR) – n_1 **= 3** and n_2 **= 4, 5, 6, ...**