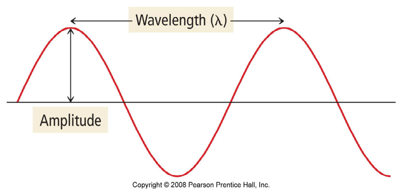
**Chapter 7: The Quantum-Mechanical Model of the Atom**

Light = electromagnetic radiation

Wave-particle duality: light has wave-like AND particle-like properties

dist.



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## The wave nature of light

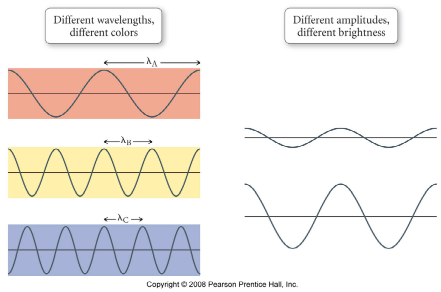
**amplitude**: dist from origin to peak

lambda

**wavelength** (λ): dist from one peak to next

nu ν

**frequency** (ν): # of full waves that pass a point per unit of time



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$\frac{1}{s}$ 10 Hz = 10 cycles per sec

frequency unit: hertz = Hz = $s^{-1}$ (cycles per second)

$c = \lambda\nu$ where $c = 3.00 \times 10^8$ m/s (speed of light in vacuum)

$(m/s) = (m)(s^{-1})$

What is ν (in Hz) of 473 nm blue light?

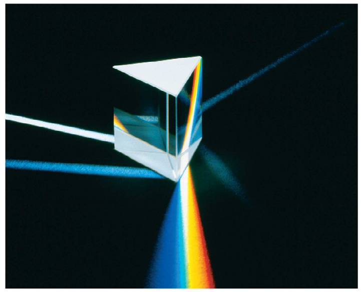
$473\ nm \times \frac{10^{-9}\ m}{1\ nm} = 4.73 \times 10^{-7}\ m$

$\nu = \frac{c}{\lambda} = \frac{3.00 \times 10^8\ m/s}{4.73 \times 10^{-7}\ m} = 6.34 \times 10^{14}\ s^{-1}\ (Hz)$



The electromagnetic spectrum

White visible light can be separated into its component colors through a prism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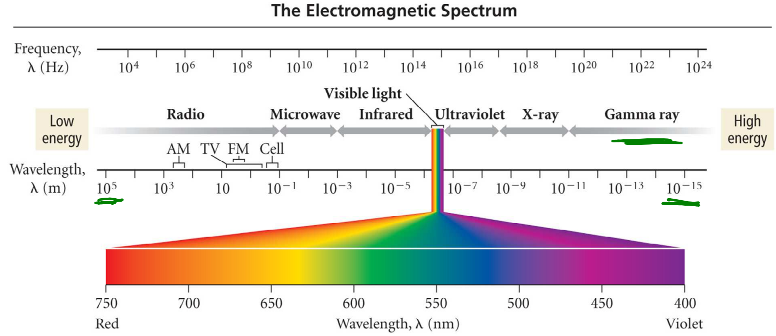

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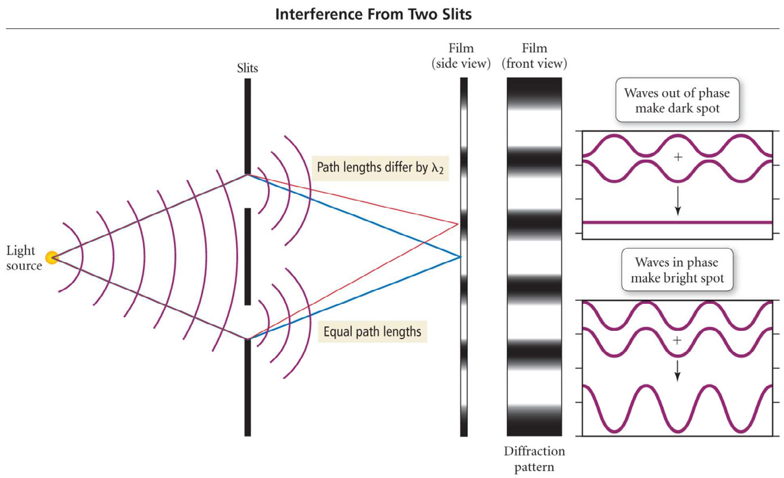
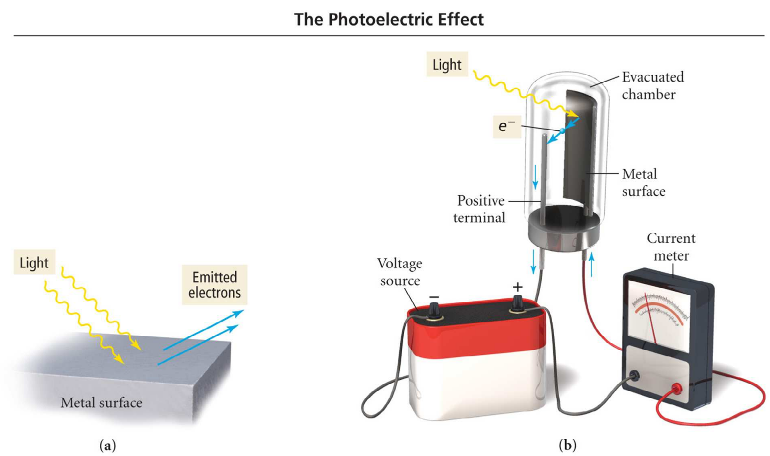
high ν
high energy
short λ

Evidence for the wave and particle natures of light










The particle nature of light

1905: Albert Einstein: **photoelectric effect**

- electrons are ejected from metal only after a certain frequency ($\nu$) of light hits it
- 1 photon of light can eject 1 electron - IF that photon has enough energy

**photon**: a individual packet or "particle" of light

particle equation

$E = h\nu$ where: $(J) = (J \cdot s)(s^{-1})$

- $E$ = energy of one photon
- $h$ = Planck's constant = $6.63 \times 10^{-34}$ J·s

since $c = \lambda\nu$, $\nu = \frac{c}{\lambda}$ and $E = \frac{hc}{\lambda}$ combined particle/wave eqn.

wave equation

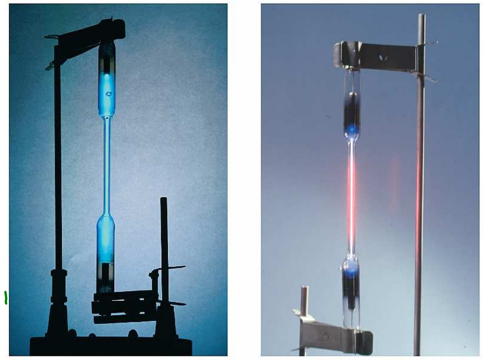
How much energy is in one photon of blue light with a wavelength of 473 nm? $c = 3.00 \times 10^8$ m/s

$$E = \frac{hc}{\lambda} = \frac{(6.63 \times 10^{-34}\ J \cdot s)(3.00 \times 10^8\ m/s)}{473 \times 10^{-9}\ m} =$$

$$= 4.21 \times 10^{-19}\ J$$

Atomic emission spectroscopy

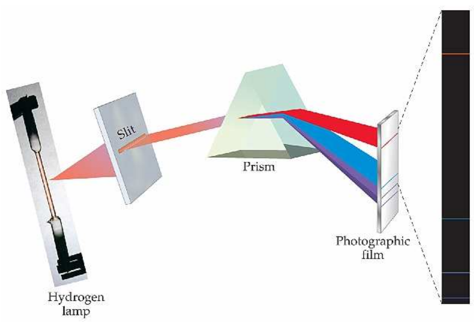
A **gas lamp** is a sealed glass tube that contains a gas sample, and **glows** when a high voltage is applied to it.



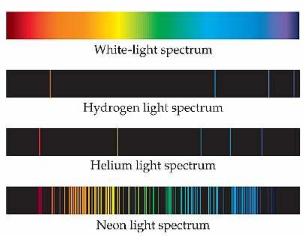


$Hg(g)$ $H_2(g)$

But **only certain wavelengths** of light are given off by a gas lamp.




Compare with the continuous spectrum given off by a white light source like a light bulb.

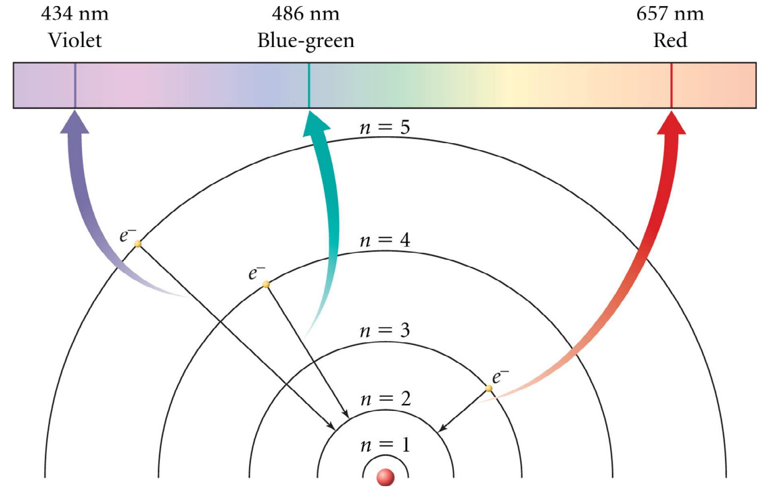







Bohr model and emission spectra

**The Bohr Model and Emission Spectra**

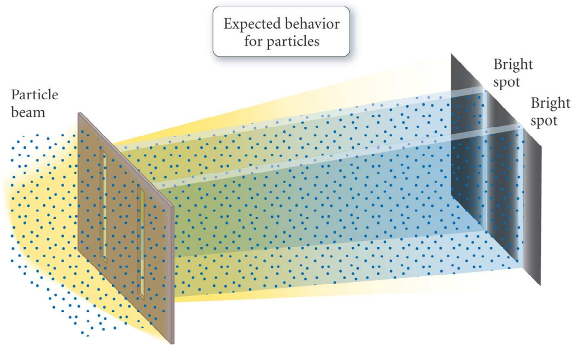





**Bohr's hydrogen atom model:** (Niels Bohr, ~ 1910)

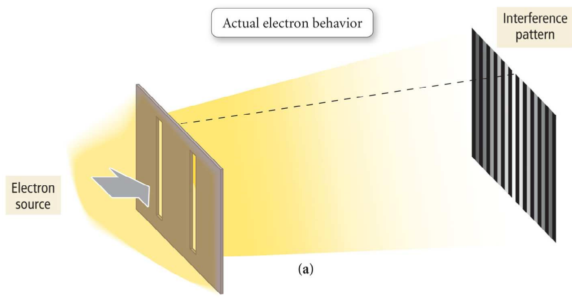
- Electrons in the H atom can occupy only certain energy levels, and the energy of the electron determines which energy level it occupies.
- If an electron is promoted to a higher energy level, it must absorb energy
- If an electron drops to a lower energy level, it gives off energy
- The amount of energy transferred = the energy difference between the levels

The wave-particle duality for electrons



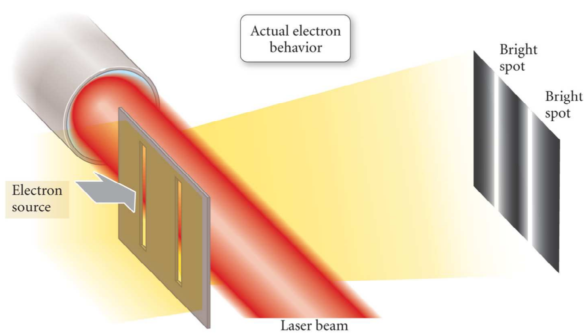

(b)






(a)










Uncertainty and indeterminacy

The wave and particle natures of the electron are **complimentary** properties - the more you know about one, the less you know about the other

**Heisenberg uncertainty principle**:

- Position of an electron: particle nature
- Momentum of an electron: wave nature
- It's impossible to know both precisely at any one time

$$(\Delta x) \cdot (m\Delta v) \geq \frac{h}{4\pi}$$

uncertainty in position ($\Delta x$); uncert in momentum ($m\Delta v$)

But, quantum mechanics allows us to calculate the **probability** of an electron behaving a certain way:

**Wavefunction** ($\psi$): mathematical equation that describes the wavelike properties of an electron

**Quantum numbers**: 4 variables in the wavefunction that, combined, describe a single electron

**Orbital**: a solution to a wavefunction with a certain combination of quantum numbers -
a 3-dimensional volume inside of which an electron is likely to be found