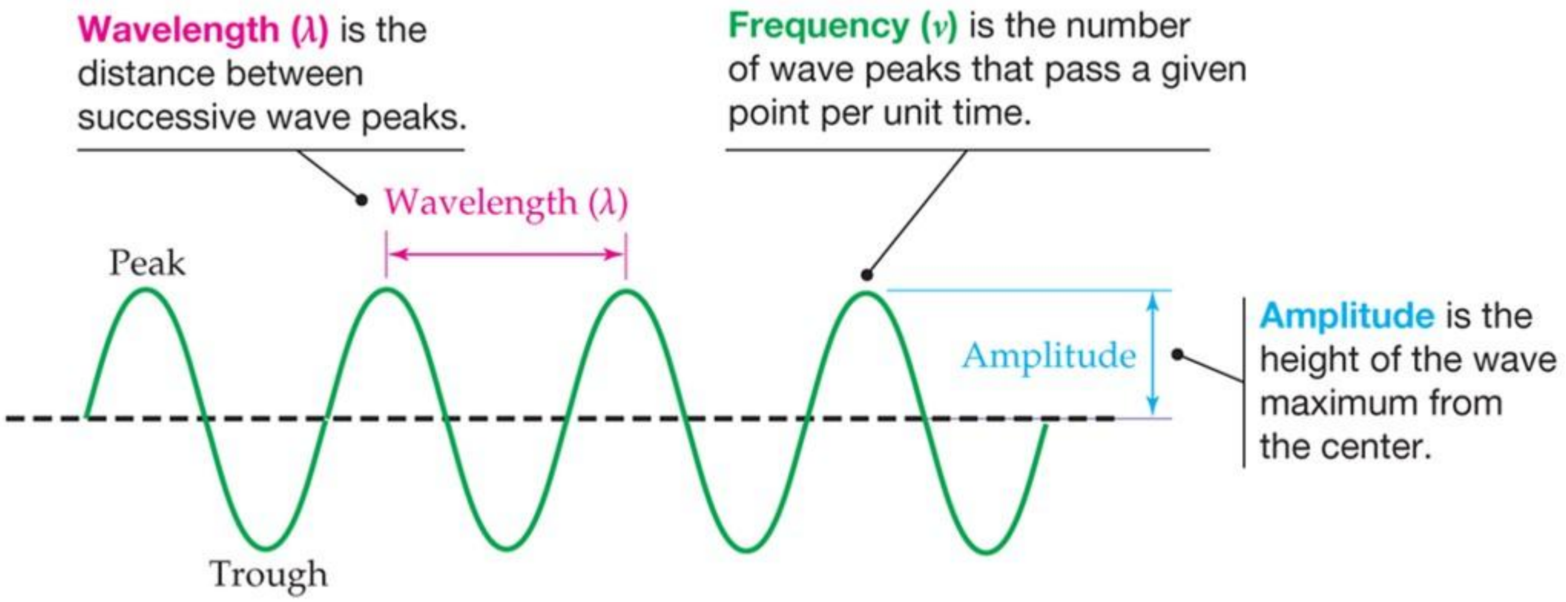
General Chemistry

202-SN1-RE

with Olivia Bibollet-Bahena

Office: 5th floor

Wave Properties of Radiant Energy



Formula

$$c = \nu \lambda$$

ν : frequency (s^{-1})
λ : wavelength (m)
c : speed of light 2.998×10^8 m/s

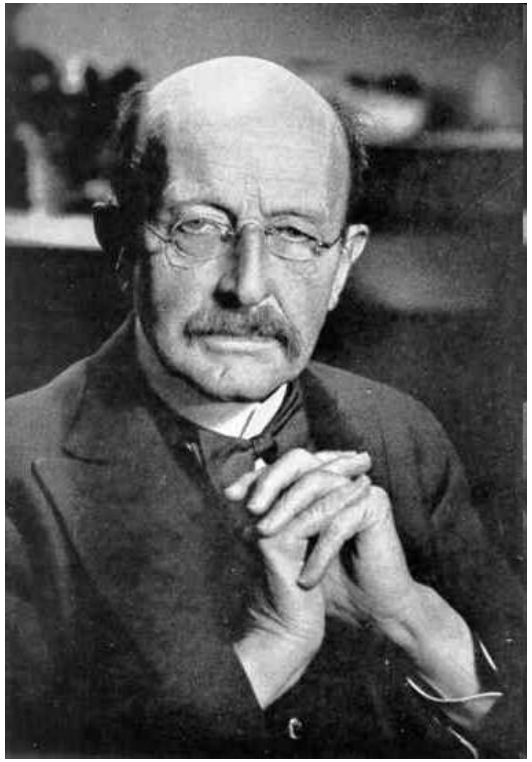
End of the 19th Century

- It seemed like all major problems in physics had been solved.
- That matter and energy were distinct
  - Matter was thought to consist of particles that had a measurable mass and easily specified position.
  - Energy in the form of light (electromagnetic radiation) was described as a wave that had no mass and its position could not be precisely defined.
  - There was no overlap in characteristics!

Is any of this true today?

Nature of Matter

The Quantization of Energy



Max Planck

German physicist
(1858 – 1947)

Planck's radiation profile studies of solid bodies heated to incandescence could not be understood in the terms of the physics of his day (which stated that matter could absorb or emit ***any*** quantity of energy).

First evidence that the early theories of matter and energy were NOT fully correct

The Photoelectric Effect

Einstein’s Next Contribution: The Photoelectric Effect (Nobel Prize)

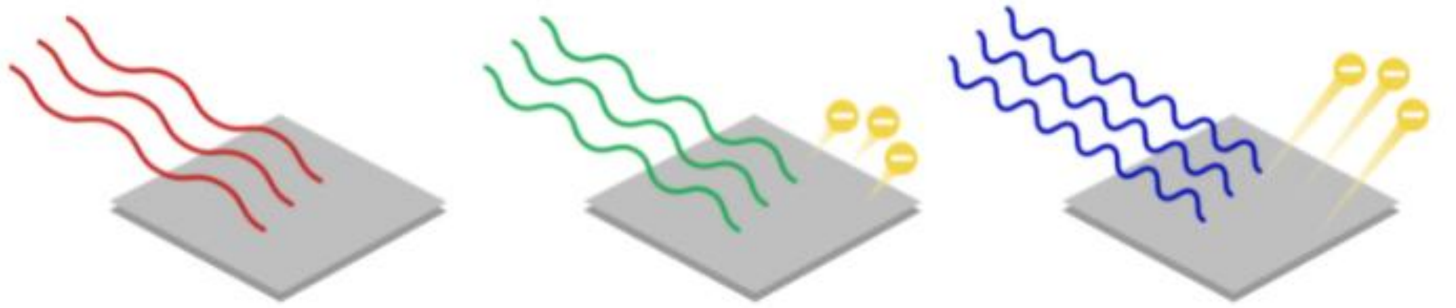
- **Photoelectric Effect**: the phenomenon in which electrons are released from the surface of a metal when light strikes it.

Particle Like Properties of Light

In classical physics, the kinetic energy of the electrons should be related to the intensity. But this is not the case.

Frequency threshold: minimal frequency at which electrons are ejected.

Reality (photoelectric effect):



https://www.chem.fsu.edu/chemlab/chm1020c/lecture%203/04.php

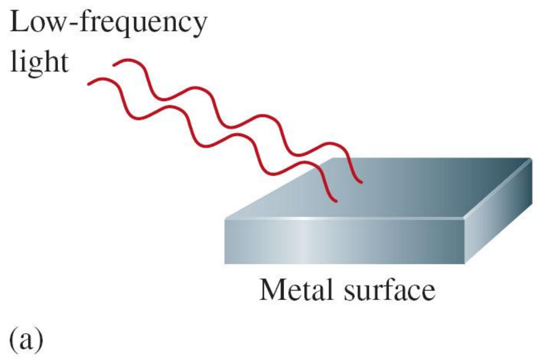
Einstein's Next Contribution: The Photoelectric Effect (Nobel Prize)

- **Photoelectric Effect**: the phenomenon in which electrons are released from the surface of a metal when light strikes it.

- Characteristics of the Photoelectric Effect:

  1) No electrons emitted below a threshold frequency (v_o) when the frequency was varied.

  2) No electrons emitted at any intensity for light below v_o.



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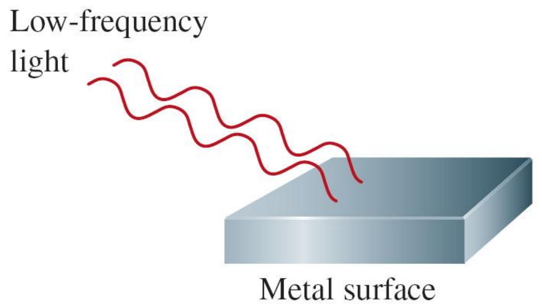


(a)

2) No electrons emitted at any intensity for light below v_o.

3) For light above v_o , the number of electrons released increased with increasing intensity.

4) For light above v_o , the kinetic energy of electrons increased with increasing frequency.

The Photoelectric Effect and the Nature of Light

Conclusions

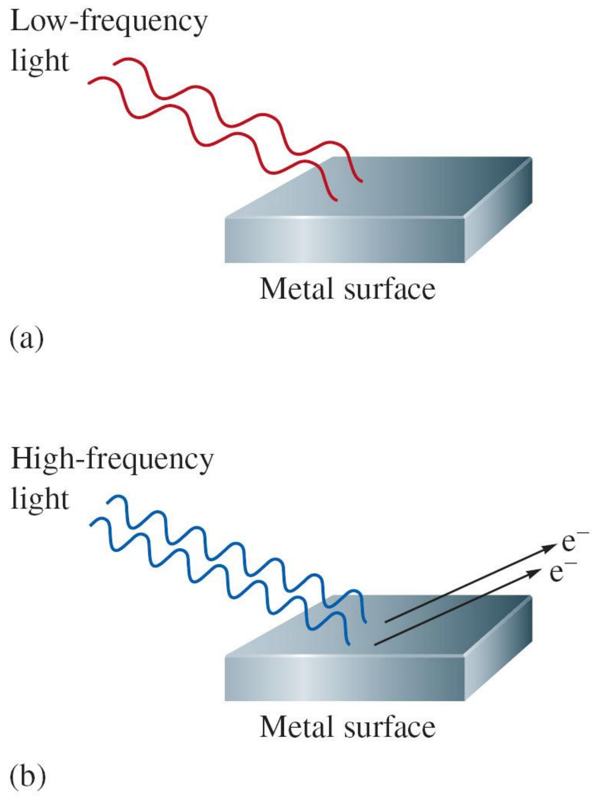
- There is a minimum energy required to remove an electron from the surface of the metal

$$E_0 = hv_0$$

for $v > v_0$

- The excess energy is passed to the electron as kinetic energy:

$$E_{K_{electron}} = \frac{1}{2}mv^2 = h(v - v_0)$$



Light is quantized and exists as a stream of particles!

Particlelike Properties of Radiant Energy: The Photoelectric Effect and Planck's Postulate

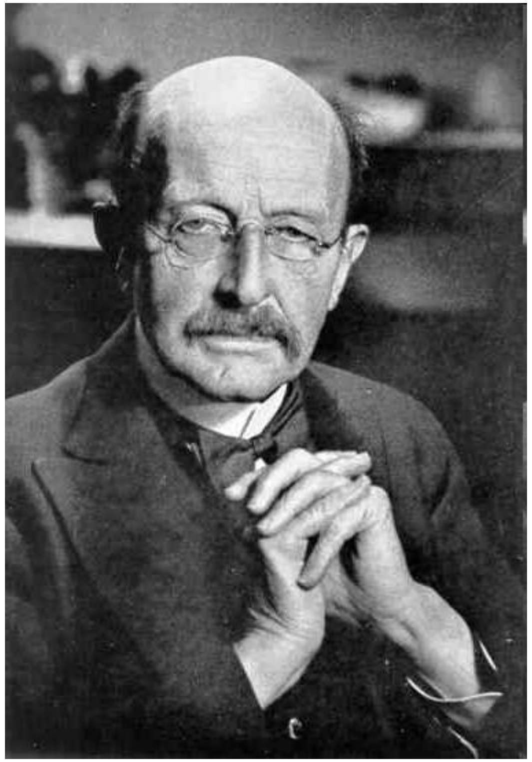
$$E = h\nu$$

$$E \uparrow \quad \nu \uparrow$$

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

Electromagnetic energy (light) is **quantized**.

The Quantization of Energy



Max Planck

Planck concluded that the energy being lost or gained in his observations was only occurring in whole number multiples of hv (where $h = 6.626 \times 10^{-34}$ J·s, today h is known as Plank's constant).

Planck's constant
(6.626×10^{-34} Js)

$$\Delta E = nhv$$

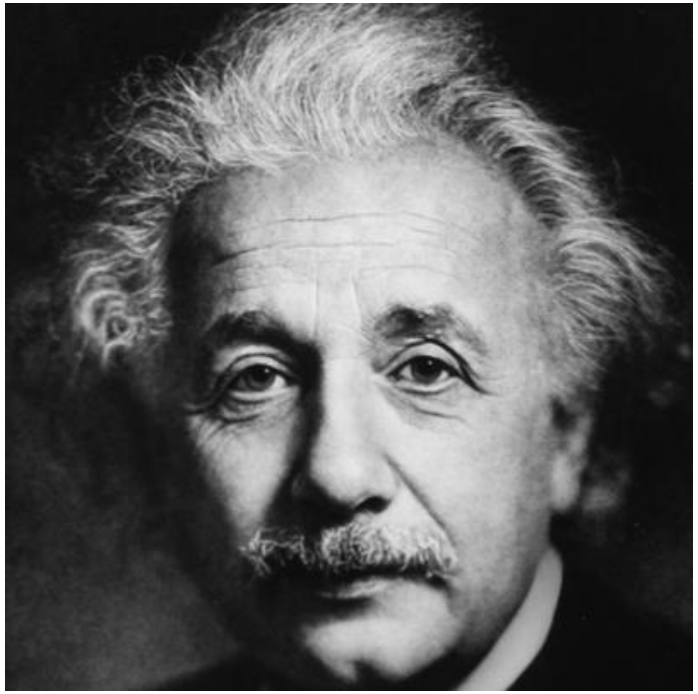
Frequency of EM radiation

Integer values

The Quantization of Energy

- Energy can be gained or lost only in integer multiples of $h\nu$.
- A system can transfer energy only in ‘packets’ called quanta.
- First evidence that energy seems to have particle properties too.

Einstein’s Contribution



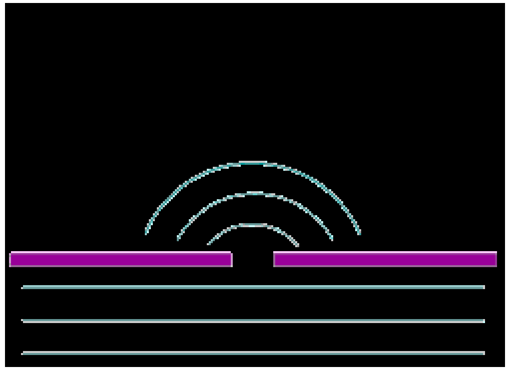
- He suggested that electromagnetic radiation be seen as a stream of “particles” called photons.

- The energy contained in a single photon is equal to a single quantum:

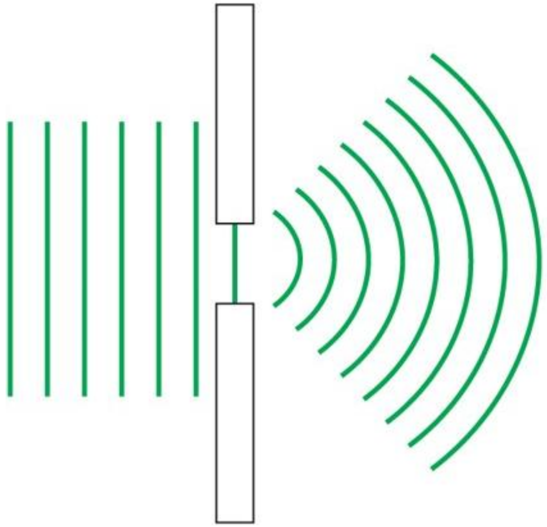
$$E_{\text{photon}} = h\nu = \frac{hc}{\lambda}$$

Properties of Light

Diffraction of Light

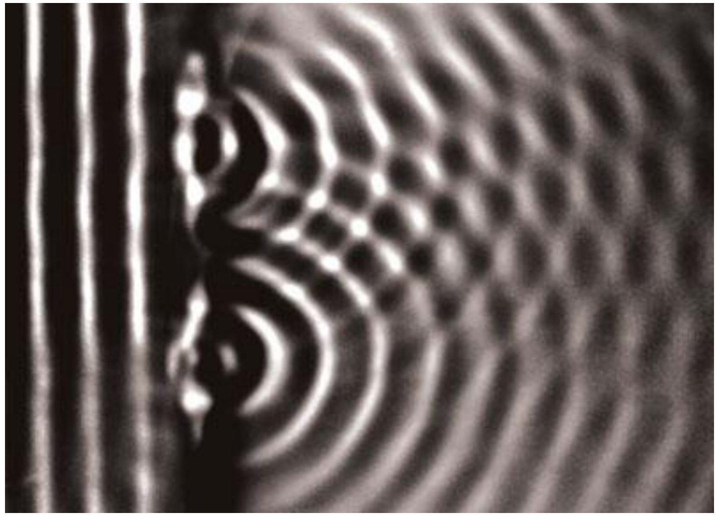


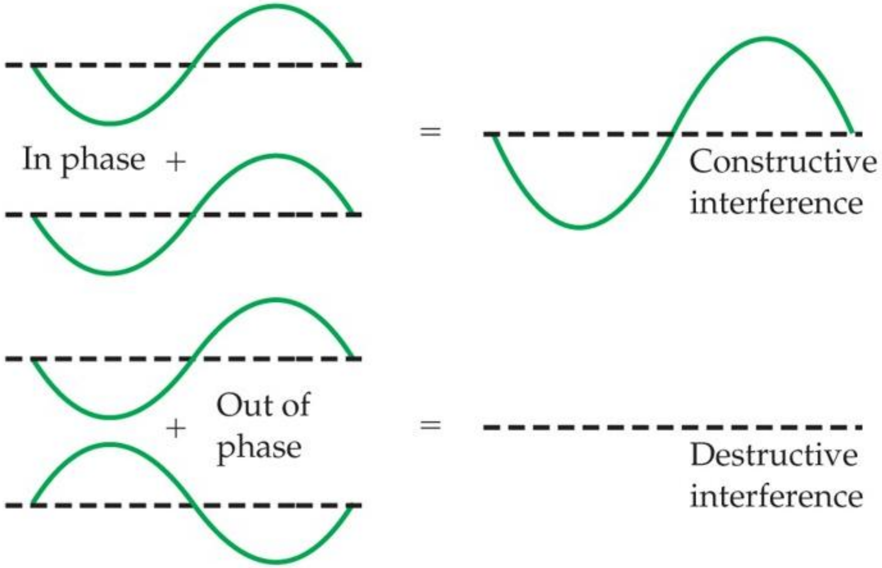
Passage of a wave through a slit causes a diffraction pattern



Diffraction is the bending of waves around an object.

Diffraction



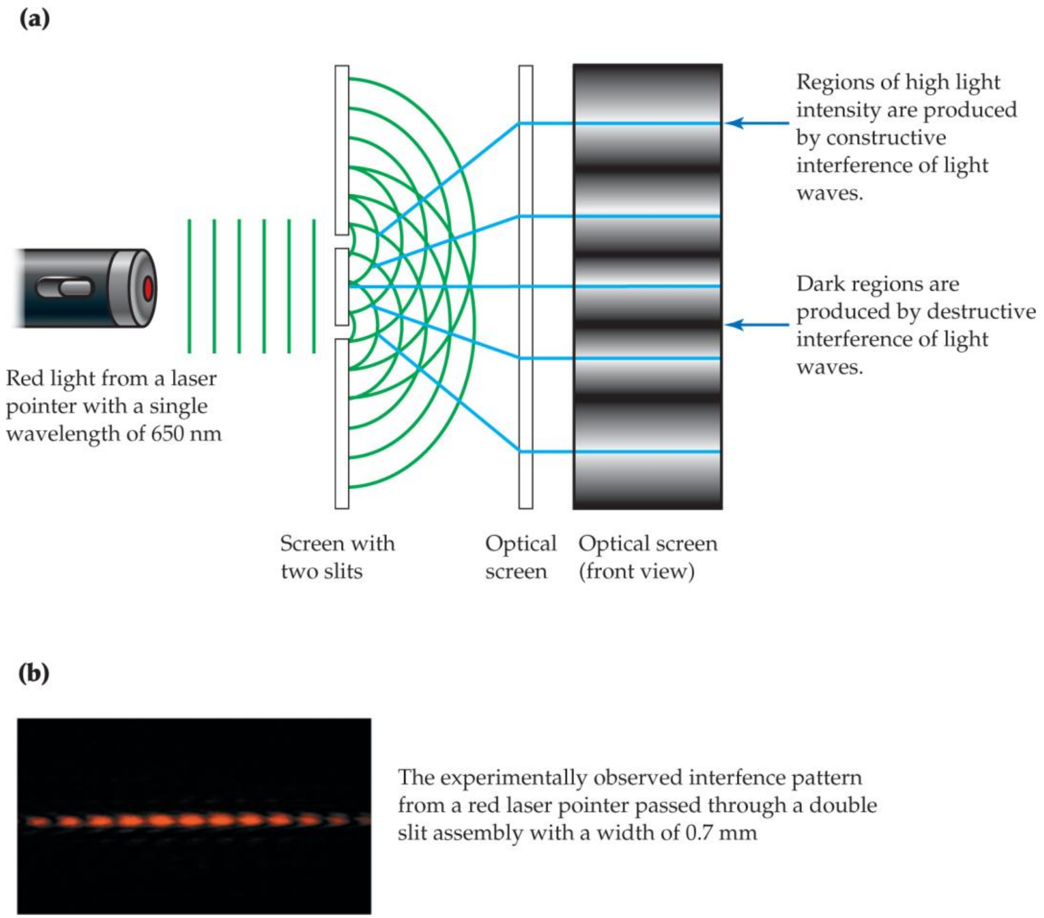


Constructive and destructive interference are wave properties

Diffraction and interference are phenomena exhibited by waves.

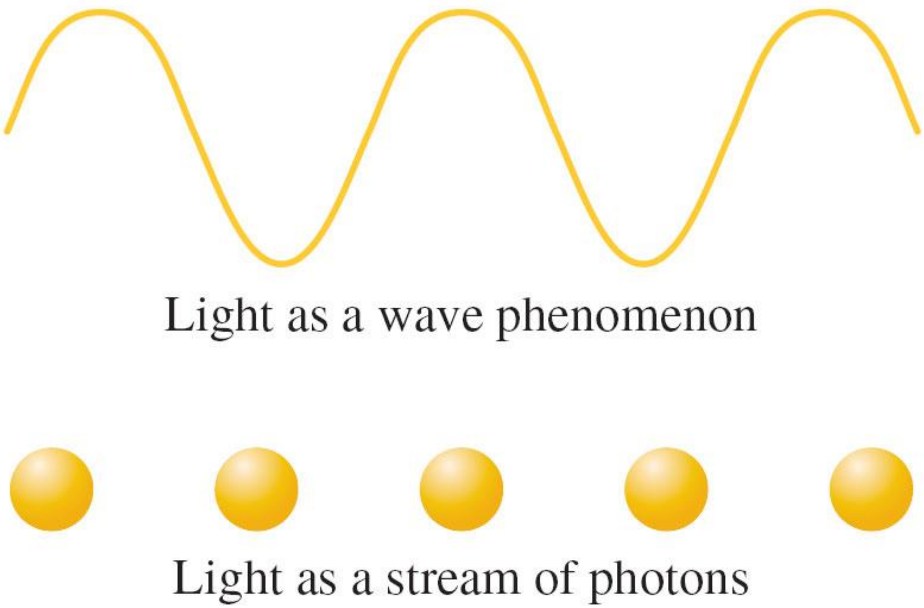
Wavelike Properties of Light

Radiant energy exhibits wave properties in a double-slit experiment.



Wave/Particle Duality of Light

Albert Einstein



Einstein's explanation of the photoelectric effect led to the notion of wave/particle duality of light.

Einstein's Next Contribution

- He went onto develop his most famous theory, the "Theory of relativity":

$$E = mc^2$$

- Therefore, energy has relativistic mass and one can calculate the mass associated with a given quantity of energy.

$$m = \frac{E}{c^2} = \frac{\frac{hc}{\lambda}}{c^2} = \frac{h}{\lambda c}$$

- Dual nature of light:
  - Electromagnetic radiation, which was previously thought to exhibit only wave properties, show particulate properties too.

Matter Waves

- Does matter that is normally assumed to be particulate exhibit wave properties?

Louis de Broglie

(1892-1987)

$$\lambda = \frac{h}{mv}$$

← v = velocity

- The de Broglie equation (rearranged from Einstein's equation) predicts that matter should have an associated wavelength!

Example

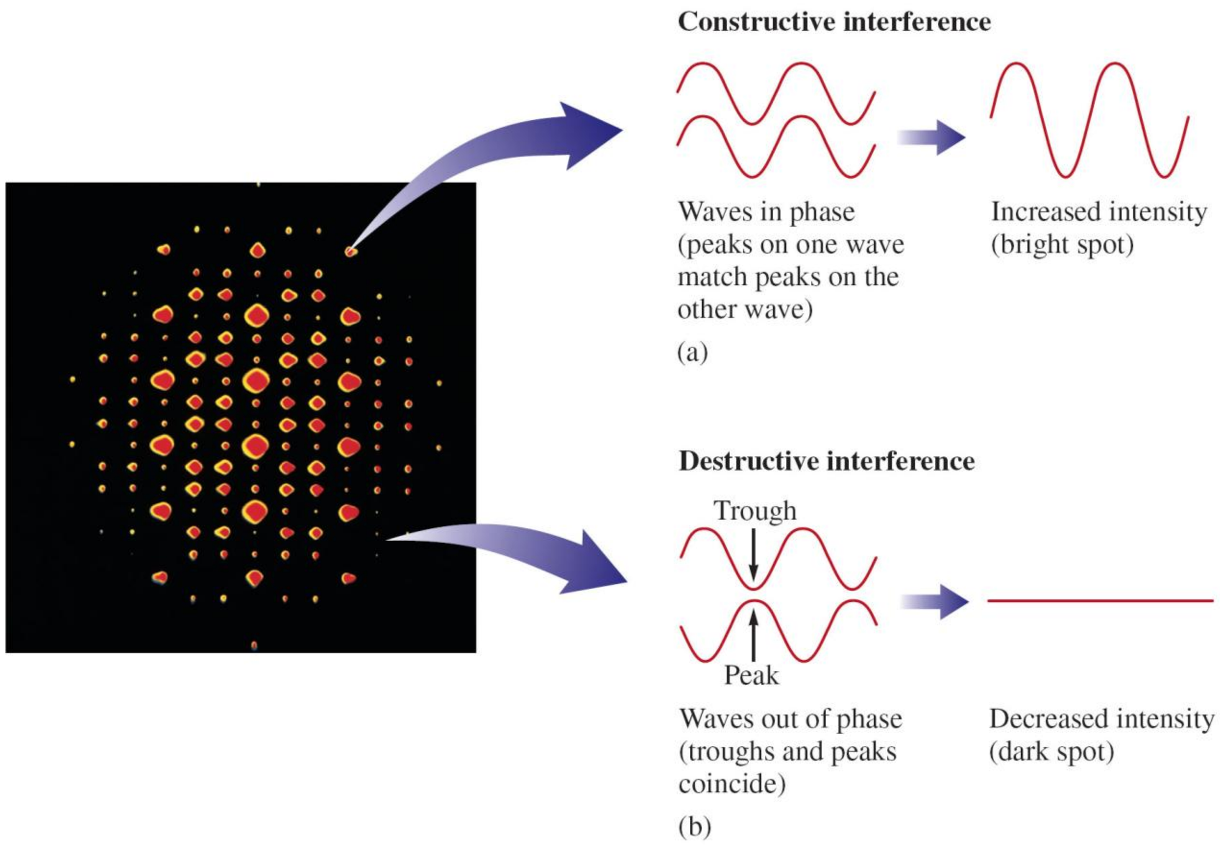
Compare the wavelength for an electron (mass = 9.11×10^{-31} kg) traveling at a speed of 1.00×10^{7} m/s and that of a baseball (mass = 0.10 kg) traveling at 35 m/s.

In SI units, J = kg x m^2 x s^{-2}

$$\lambda = \frac{h}{mv}$$

Electron Diffraction Patterns

Beam of electrons directed onto a nickel crystal. The scattered radiation produces a diffraction pattern.



- Electrons show diffraction patterns! These patterns of diffraction prove that electrons have wave nature as well.

Conclusions

- Electromagnetic radiation originally thought to be a pure waveform, actually possesses particulate properties.

- Electrons originally thought to be particles, actually have a wavelength associated with them.

Conclusions

- Matter and energy are not distinct; energy is a form of matter, and matter shows the same types of properties.

  - All matter exhibits both particulate and wave properties.
  - The extent at which they display these properties is dependent on the size of the piece of matter.

    - Larger matter like a ball has more particle properties than wave properties
    - Smaller matter like photons have more wave properties than particle properties

Atomic Spectrum of Hydrogen

Atomic Spectrum of Hydrogen

A spectrum is produced when light is separated into its component wavelengths.

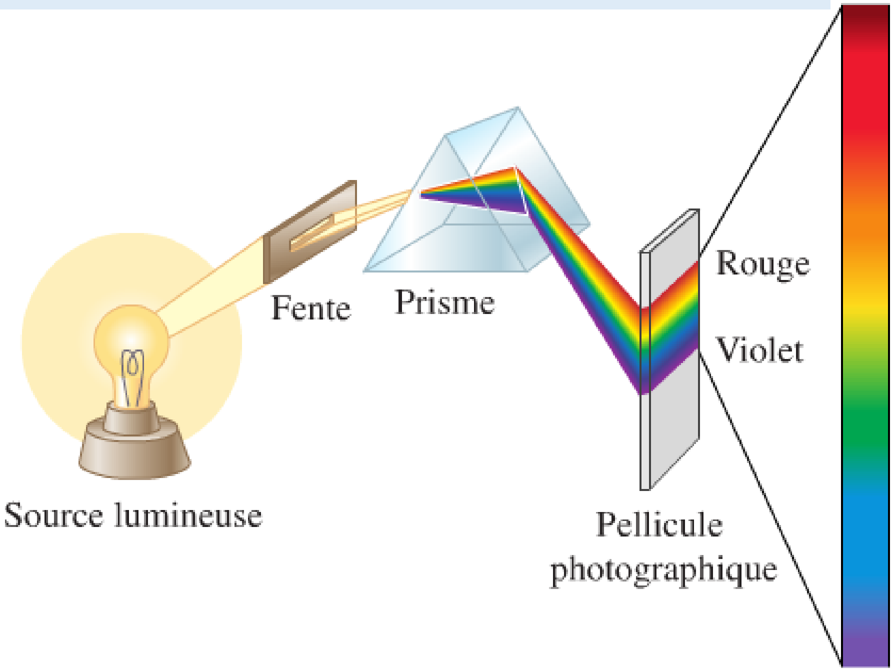
- Continuous spectrum (when white light is passed through a prism) – contains all the wavelengths of visible light.

- Line spectrum – each line corresponds to a discrete wavelength:
  - Hydrogen emission spectrum is a line spectrum.

Continuous Spectrum and Line Spectrum

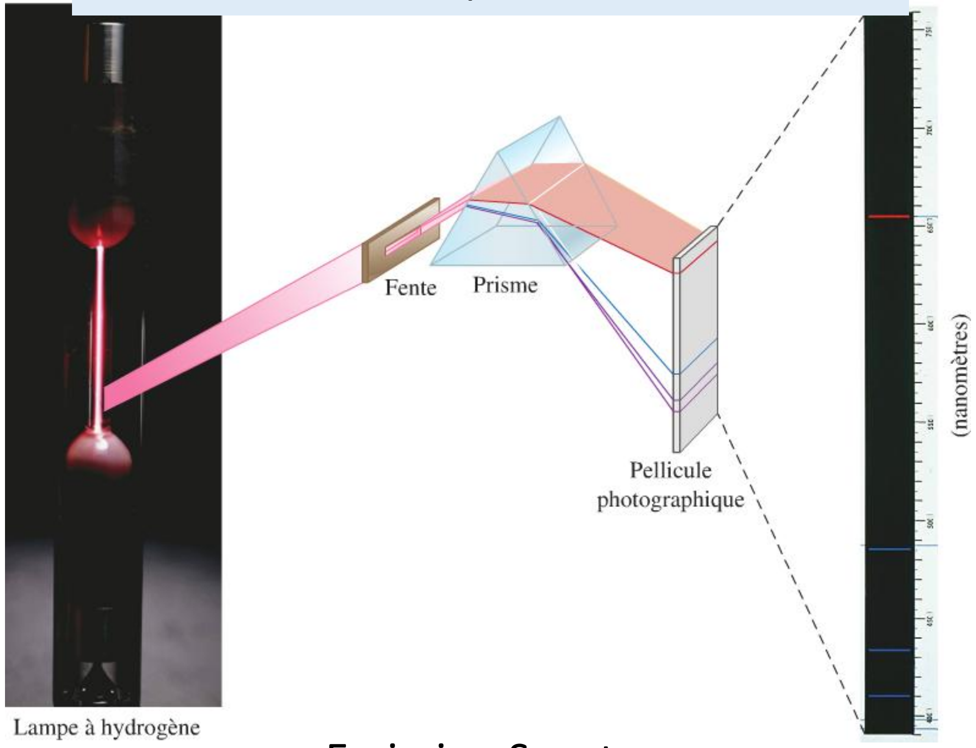
Hydrogen emission spectrum

Continuous Spectrum



When light passes through a prism and separates into a <u>range</u> of wavelengths – from red (780 nm) to violet (380 nm).

Line Spectrum



Emission Spectrum

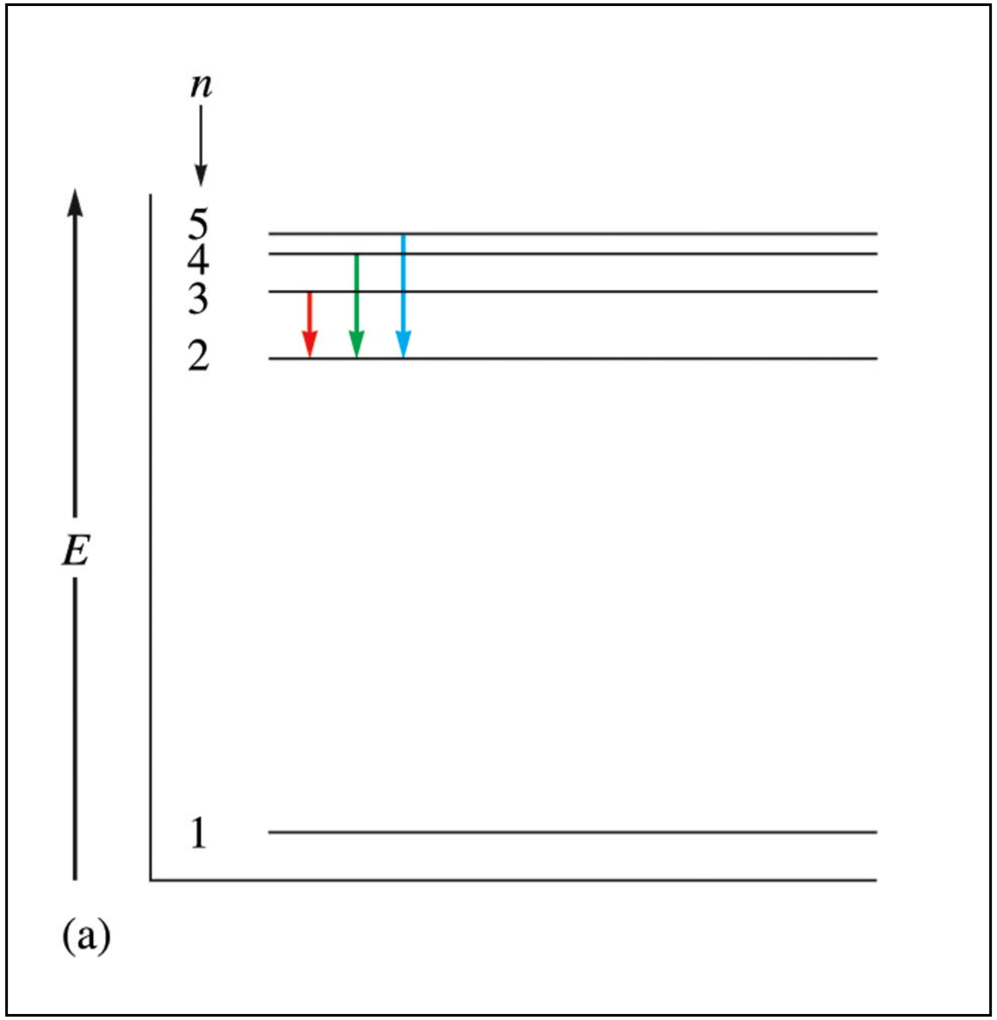
Series of discrete <u>lines</u> as a result of light emitted by excited (or heated) atoms.

Electronic Transitions

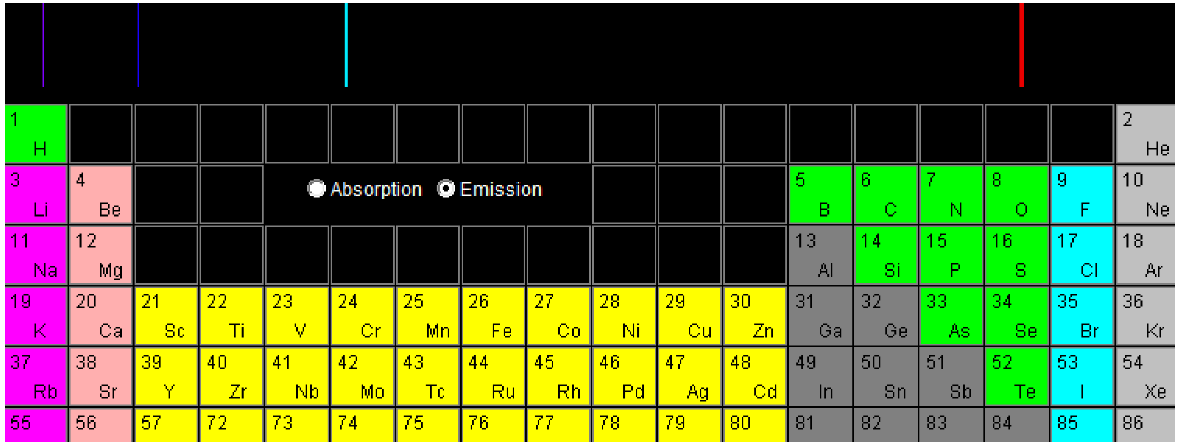
a) An Energy-Level Diagram for Electronic Transitions

(*n* – principle quantum number)

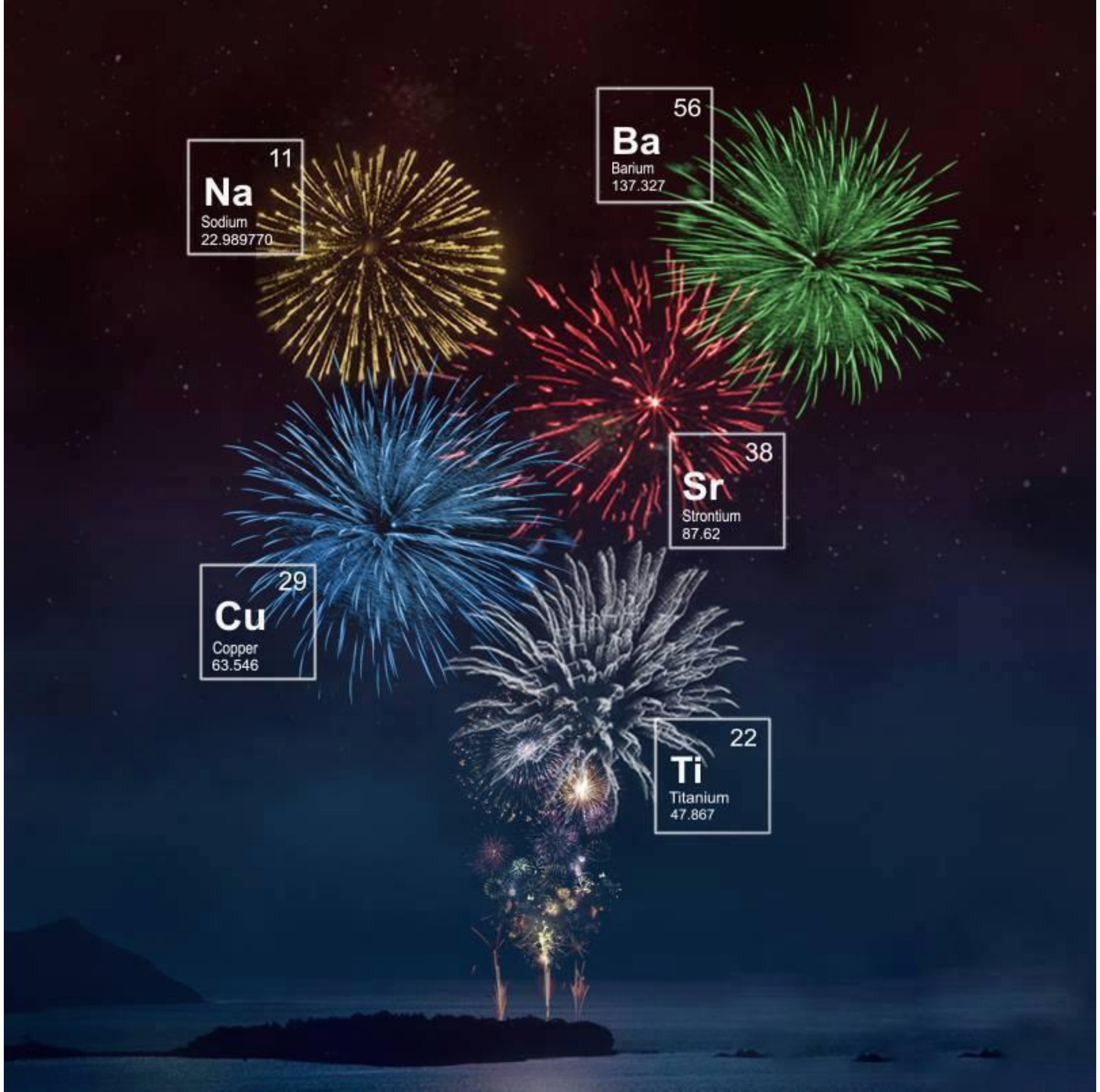
Ground state – lowest possible energy state (for hydrogen *n* = 1)



Line Spectrum of Elements



Every element has its own characteristic emission spectrum (like a fingerprint).


11
Na
Sodium
22.989770
56
Ba
Barium
137.327
38
Sr
Strontium
87.62
29
Cu
Copper
63.546
22
Ti
Titanium
47.867

The Balmer-Rydberg Equation

- Back in the 1880's, Balmer first developed an equation to predict the wavelengths of light seen in the hydrogen line spectrum.

- Eventually more lines were found outside the visible region and the equation was generalized:

$$\frac{1}{\lambda} = (R_\infty)\left(\frac{1}{{n_{inner}}^2} - \frac{1}{{n_{outer}}^2}\right)$$

$$R_\infty = 1.096776 \times 10^7\ m^{-1}$$

Exercise

Calculate the wavelength of the visible light photon corresponding to a transition from *n* = 5 using the Balmer – Rydberg equation.

Solving for λ in a transition from n = 5 to n = 2:

$$\frac{1}{\lambda} = (R_\infty)\left(\frac{1}{{n_{inner}}^2} - \frac{1}{{n_{outer}}^2}\right)$$

$$R_\infty = 1.096776 \times 10^7\ m^{-1}$$

Significance

- Only certain energies are allowed for the electron in the hydrogen atom.
- Energy of the electron in the hydrogen atom is quantized.

Concept Check

A) Why is it significant that the color emitted from the hydrogen emission spectrum is not white?

B) How does the emission spectrum support the idea of quantized energy levels?

The Bohr Model

The Bohr Model

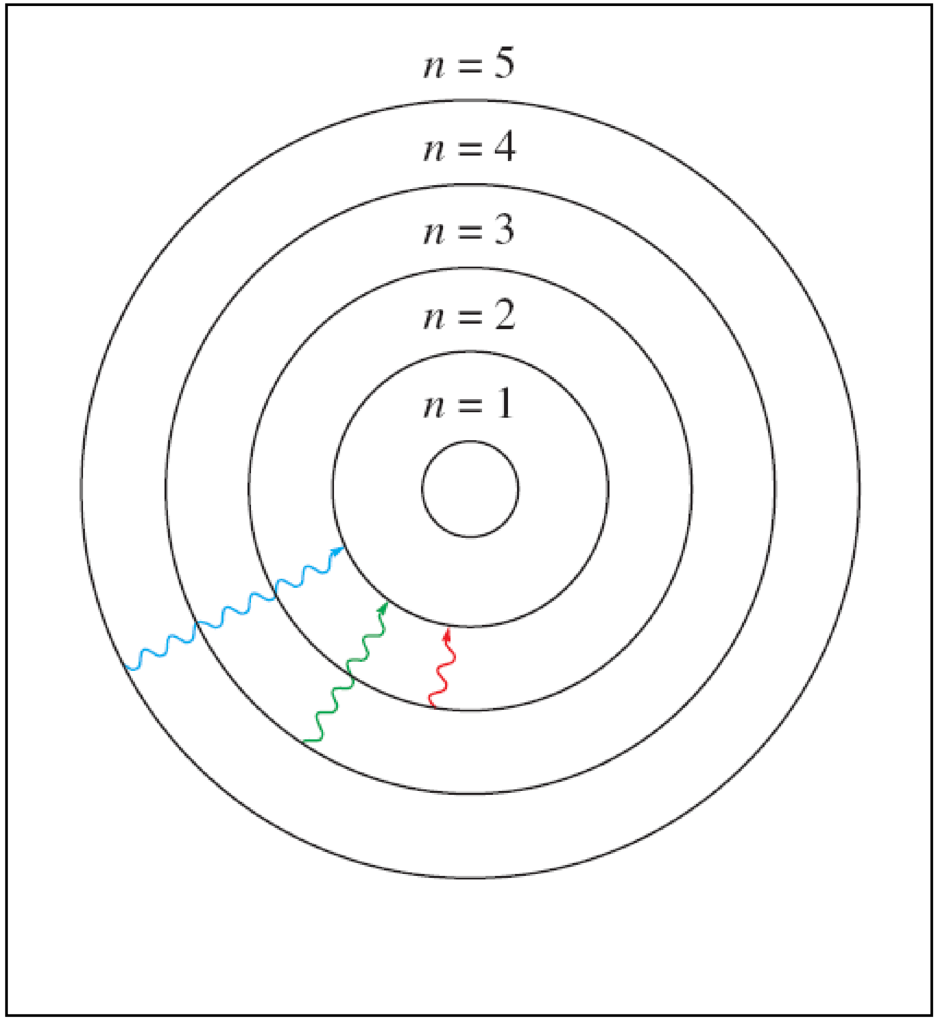
- Electron in a hydrogen atom circulates around the nucleus in orbits of certain allowed radii.

- An electron in orbit is in an “allowed” energy state and does not radiate energy.

- Energy is emitted or absorbed by the electron as it changes from one “allowed” state to another. The energy is absorbed or emitted as a photon of energy $E = h\nu$.

Electronic Transitions in the Bohr Model for the Hydrogen Atom

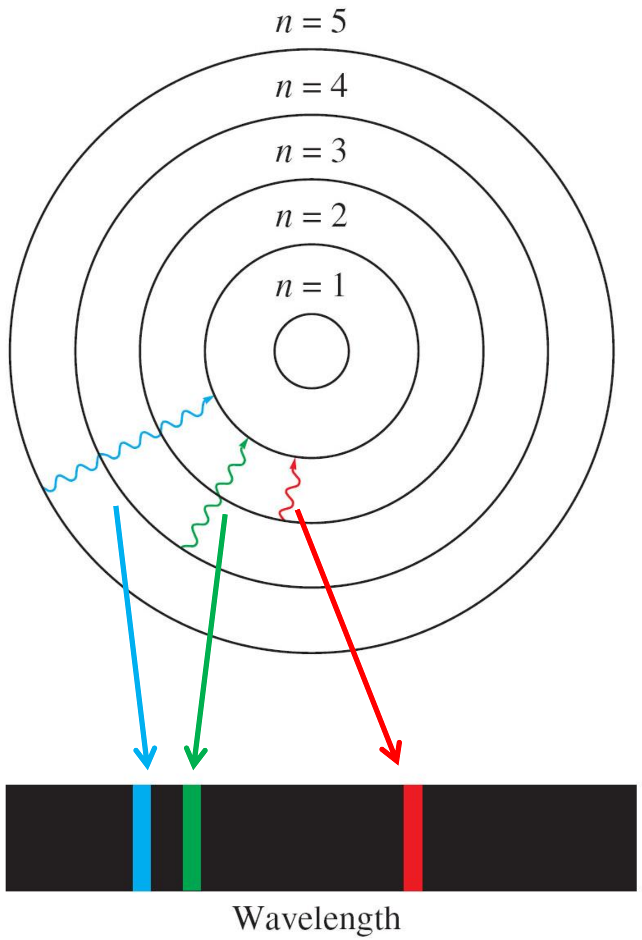
An Orbit-Transition Diagram, which accounts for the experimental spectrum



Ground state – lowest possible energy state (for hydrogen *n* = 1)

Bohr’s Atomic Model

- Bohr’s model gave hydrogen atom energy levels consistent with the hydrogen emission spectrum.
- Lines in the visible region of the line spectrum correspond to transitions from higher levels to the *n* = 2 level.



Line spectrum

The Bohr Model

- From the Balmer–Rydberg equation and the energy equations developed by Planck and Einstein:

If $$\frac{1}{\lambda} = (R_{\infty})\left(\frac{1}{n_{inner}^{2}} - \frac{1}{n_{outer}^{2}}\right)$$

and $$\Delta E_{photon} = h\nu = \frac{hc}{\lambda} \quad \Rightarrow \quad \frac{1}{\lambda} = \frac{\Delta E_{photon}}{hc}$$

then $$\frac{\Delta E_{photon}}{hc} = (R_{\infty})\left(\frac{1}{n_{inner}^{2}} - \frac{1}{n_{outer}^{2}}\right)$$

Replace R_{∞}, h, and c with their actual values……

The Bohr Model

- For a single electron transition from one energy level to another:

$R_H = R_\infty \times h \times c$
$= 2.178 \times 10^{-18}$

$$\Delta E = -2.178 \times 10^{-18}\ \text{J}\left(\frac{1}{n_{\text{final}}^2} - \frac{1}{n_{\text{initial}}^2}\right)$$

ΔE = change in energy of the atom (energy of the emitted photon)

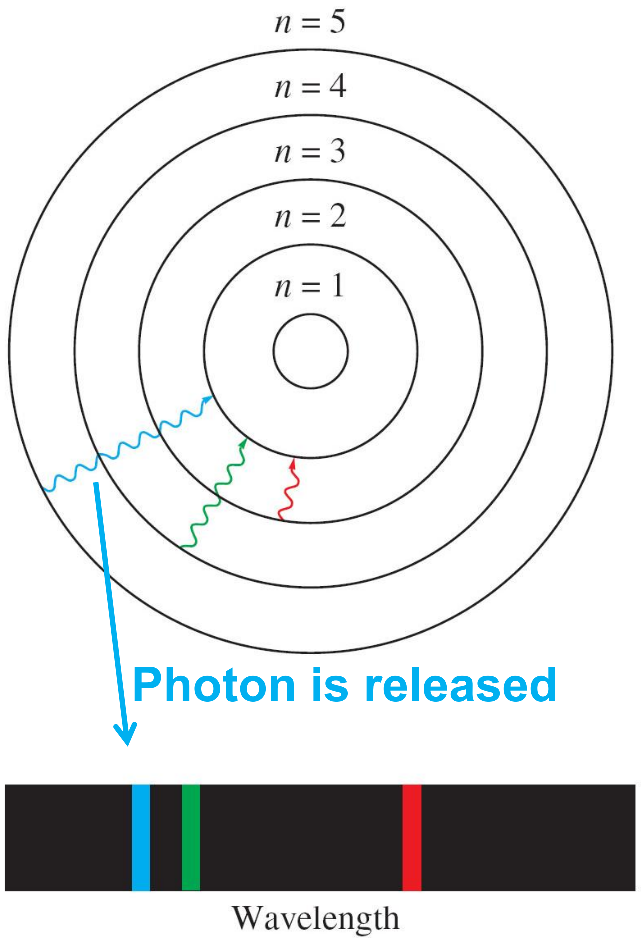
n_{final} = integer; final distance from the nucleus

n_{initial} = integer; initial distance from the nucleus

Bohr's Atomic Model

- When an electron originally in a higher energy orbit (ex. n = 5) falls back to a lower-energy orbit (ex. n = 2), a photon with energy *ΔE is released*.

$$\Delta E = -2.178 \times 10^{-18}\ \text{J}\left(\frac{1}{n_{\text{final}}^2} - \frac{1}{n_{\text{initial}}^2}\right)$$



Line spectrum

Exercise

What energy is released when an excited electron in the hydrogen atom falls from:

a) n = 5 to n = 2
b) n = 4 to n = 2
c) n = 3 to n = 2

Which transition results in the longest wavelength of light?

The Bohr Model

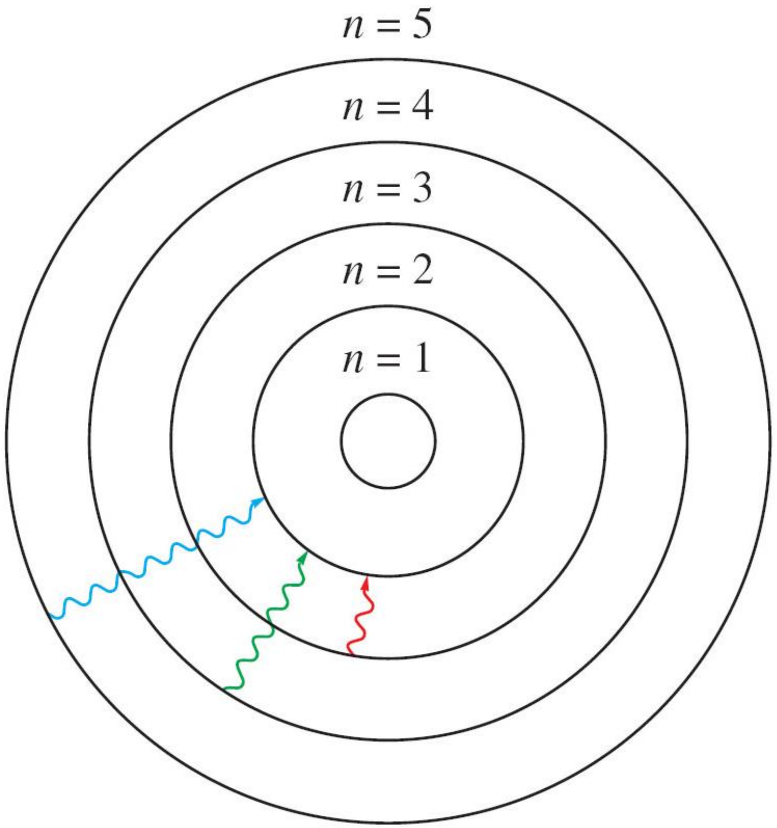
Two important points about the Bohr model:

- The model correctly fits the quantized energy levels of the hydrogen atom and postulates only certain allowed circular orbits for the electron.

- As the electron becomes more tightly bound, its energy becomes more negative relative to the zero-energy reference state (free electron). As the electron is brought closer to the nucleus, energy is released from the system.

The Bohr Model

Niels Bohr

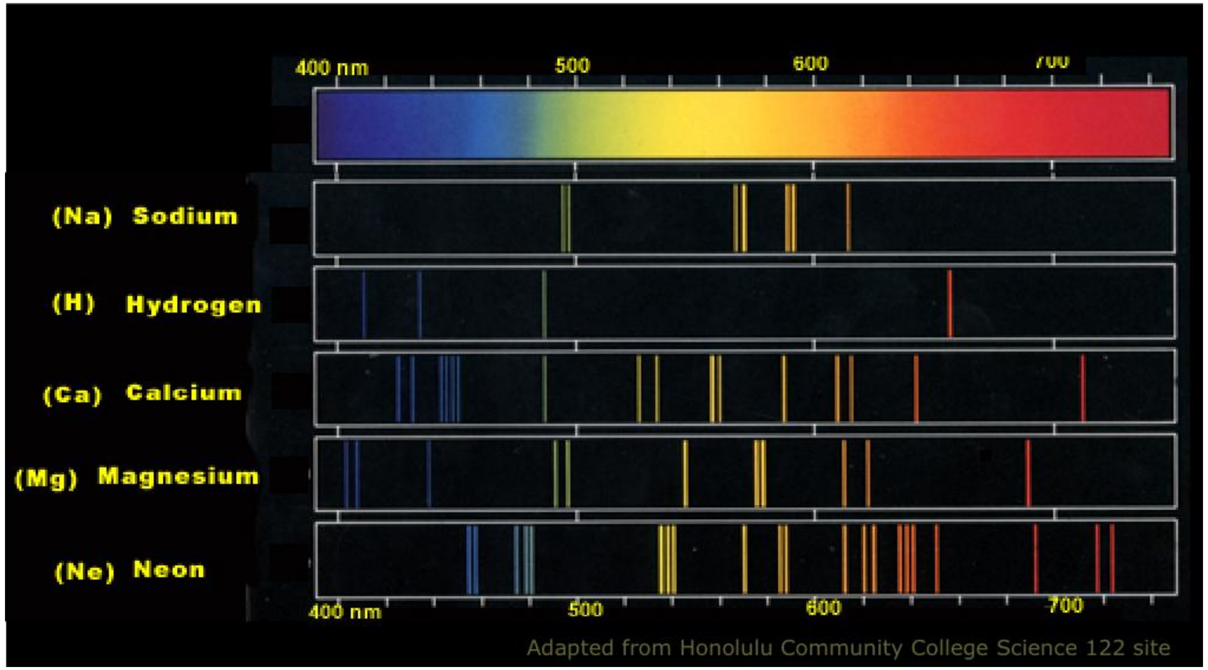


Bohr's model of the hydrogen atom was an early quantum model but could not be applied to other atoms!

The Bohr Model

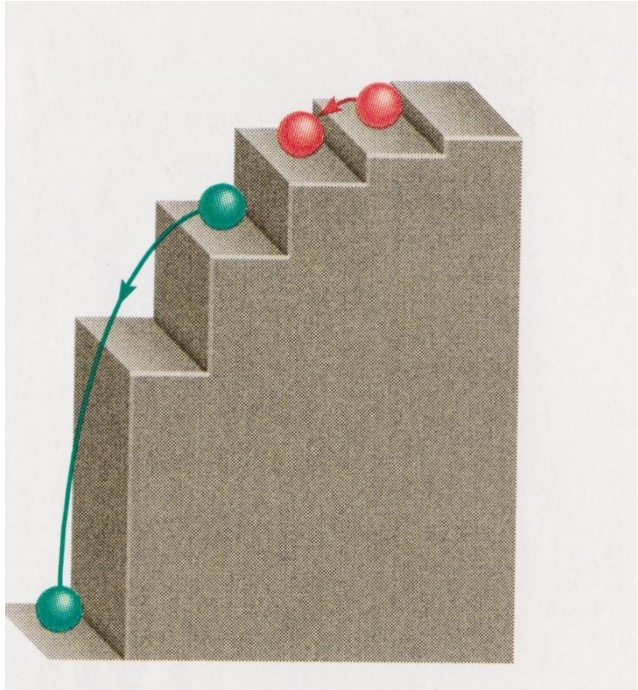
- Bohr's model is incorrect. This model only works for hydrogen.
  - *It does not work for any other element.*

- Electrons *do not* move around the nucleus in circular orbits.



The Bohr Model

- Two of Bohr's postulates however were important steps toward our current understanding:



1. Electrons occupy discrete energy levels, which are described by the principle quantum number.

2. Energy in the form of photons is involved in the transitions between levels.

The quantized movement of the electron from one energy state to another is analogous to the movement of a tennis ball either up or down a set of stairs. The ball can be on any of several steps but never between steps.

Adjusting for "Hydrogen-Like" Atoms

- Can be applied to other elements that have only one electron (ex: $He^+,\ Li^{2+}, Be^{3+}, etc$)

$$\Delta E = -B \cdot Z^2 \left(\frac{1}{{n_f}^2} - \frac{1}{{n_i}^2} \right)$$

- Where B is the constant for that particular element, and Z is the number of protons