

# Periodic Table of the Elements

1 H Hydrogen 1.01																	2 He Helium 4.00
3 Li Lithium 6.94	4 Be Beryllium 9.01											5 B Boron 10.81	6 C Carbon 12.01	7 N Nitrogen 14.01	8 O Oxygen 16.00	9 F Fluorine 19.00	10 Ne Neon 20.18
11 Na Sodium 22.99	12 Mg Magnesium 24.31											13 Al Aluminum 26.98	14 Si Silicon 28.09	15 P Phosphorus 30.97	16 S Sulfur 32.06	17 Cl Chlorine 35.45	18 Ar Argon 39.95
19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.88	23 V Vanadium 50.94	24 Cr Chromium 51.99	25 Mn Manganese 54.94	26 Fe Iron 55.85	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.38	31 Ga Gallium 69.72	32 Ge Germanium 72.63	33 As Arsenic 74.92	34 Se Selenium 78.97	35 Br Bromine 79.90	36 Kr Krypton 83.80
37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybdenum 95.95	43 Tc Technetium 98.91	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.91	46 Pd Palladium 106.42	47 Ag Silver 107.87	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.71	51 Sb Antimony 121.76	52 Te Tellurium 127.6	53 I Iodine 126.90	54 Xe Xenon 131.29
55 Cs Cesium 132.91	56 Ba Barium 137.33	57-71 Lanthanides	72 Hf Hafnium 178.49	73 Ta Tantalum 180.95	74 W Tungsten 183.85	75 Re Rhenium 186.21	76 Os Osmium 190.23	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.97	80 Hg Mercury 200.59	81 Tl Thallium 204.38	82 Pb Lead 207.20	83 Bi Bismuth 208.98	84 Po Polonium [208.98]	85 At Astatine 209.98	86 Rn Radon 222.02
87 Fr Francium 223.02	88 Ra Radium 226.03	89-103 Actinides	104 Rf Rutherfordium [261]	105 Db Dubnium [262]	106 Sg Seaborgium [266]	107 Bh Bohrium [264]	108 Hs Hassium [277]	109 Mt Meitnerium [278]	110 Ds Darmstadtium [281]	111 Rg Roentgenium [282]	112 Cn Copernicium [285]	113 Nh Nihonium [286]	114 Fl Flerovium [289]	115 Mc Moscovium [289]	116 Lv Livermorium [293]	117 Ts Tennessine [294]	118 Og Oganesson [294]

# General Chemistry

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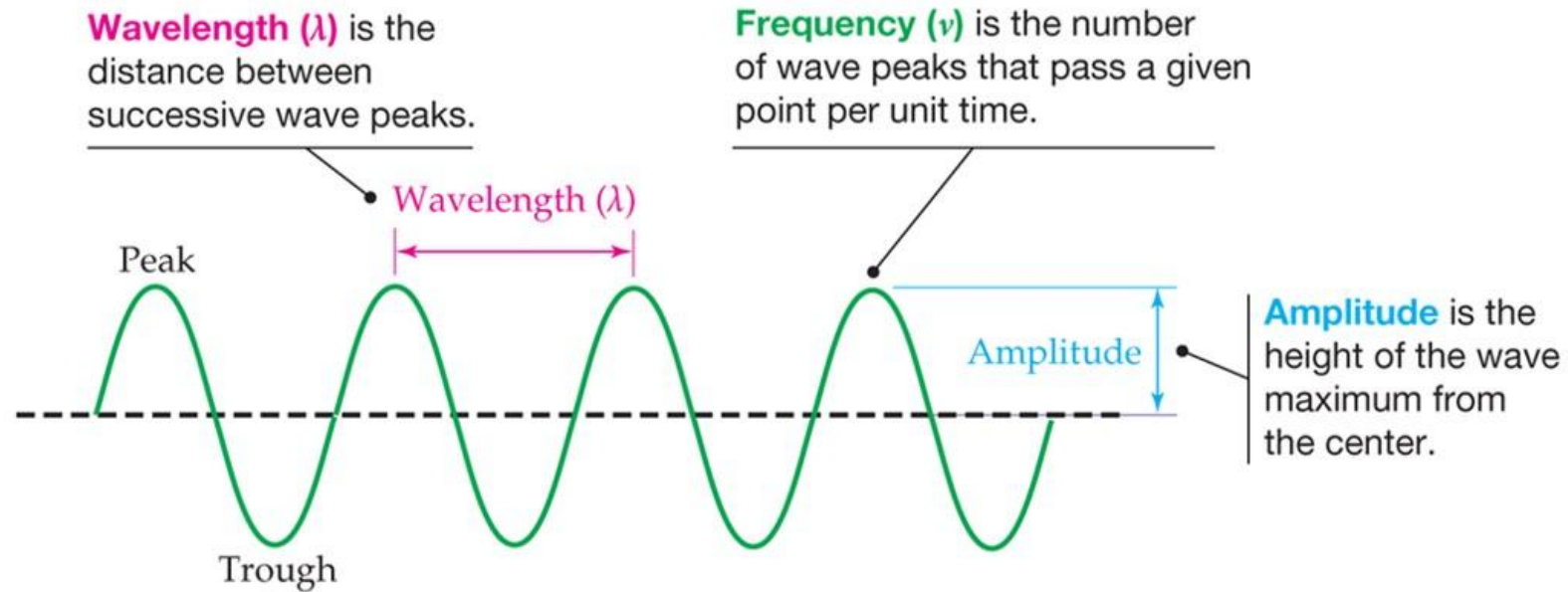
with Olivia Bibollet-Bahena

Office: 5<sup>th</sup> floor

57 La Lanthanum 138.91	58 Ce Cerium 140.12	59 Pr Praseodymium 140.91	60 Nd Neodymium 144.24	61 Pm Promethium 144.91	62 Sm Samarium 150.36	63 Eu Europium 151.96	64 Gd Gadolinium 157.25	65 Tb Terbium 158.93	66 Dy Dysprosium 162.50	67 Ho Holmium 164.93	68 Er Erbium 167.26	69 Tm Thulium 168.93	70 Yb Ytterbium 173.06	71 Lu Lutetium 174.97
89 Ac Actinium 227.03	90 Th Thorium 232.04	91 Pa Protactinium 231.04	92 U Uranium 238.03	93 Np Neptunium 237.05	94 Pu Plutonium 244.06	95 Am Americium 243.06	96 Cm Curium 247.07	97 Bk Berkelium 247.07	98 Cf Californium 251.08	99 Es Einsteinium [254]	100 Fm Fermium 257.10	101 Md Mendelevium 258.10	102 No Nobelium 259.10	103 Lr Lawrencium [262]

- Alkali Metal
- Alkaline Earth
- Transition Metal
- Basic Metal
- Metalloid
- Nonmetal
- Halogen
- Noble Gas
- Lanthanide
- Actinide

# Wave Properties of Radiant Energy



# Formula

$$c = \nu \lambda$$

$\nu$  : frequency ( $s^{-1}$ )

$\lambda$  : wavelength (m)

$c$  : speed of light  $2.998 \times 10^8$  m/s

# End of the 19<sup>th</sup> 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)

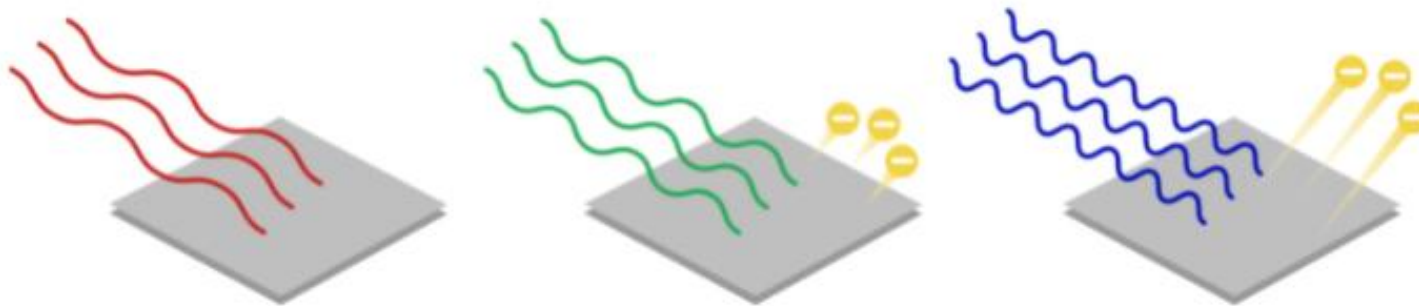
- **Photoelectric Effect**: the phenomenon in which electrons are released from the surface of a metal when light strikes it.

# 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.

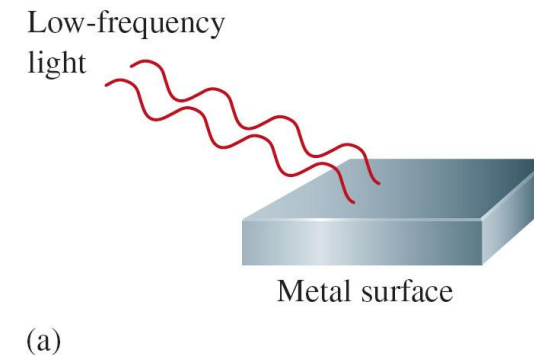
Reality (photoelectric effect):



# Einstein's Next Contribution:

## The Photoelectric Effect (Nobel Prize)

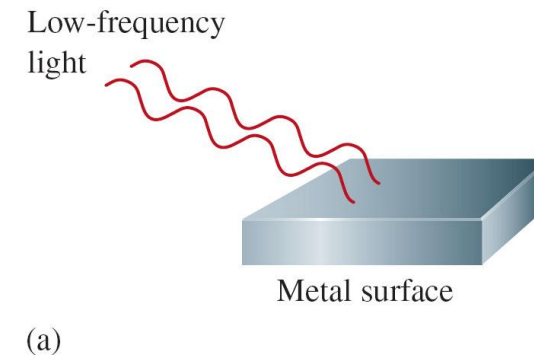
- **Photoelectric Effect:** the phenomenon in which electrons are released from the surface of a metal when light strikes it.
- **Characteristics** of the Photoelectric Effect:
  - 1) No electrons emitted below a threshold frequency ( $\nu_0$ ) when the frequency was varied.
  - 2) No electrons emitted at any intensity for light below  $\nu_0$ .



# 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 ( $\nu_0$ ) when the frequency was varied.
  - 2) No electrons emitted at any intensity for light below  $\nu_0$ .
  - 3) For light above  $\nu_0$ , the number of electrons released increased with increasing intensity.
  - 4) For light above  $\nu_0$ , 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

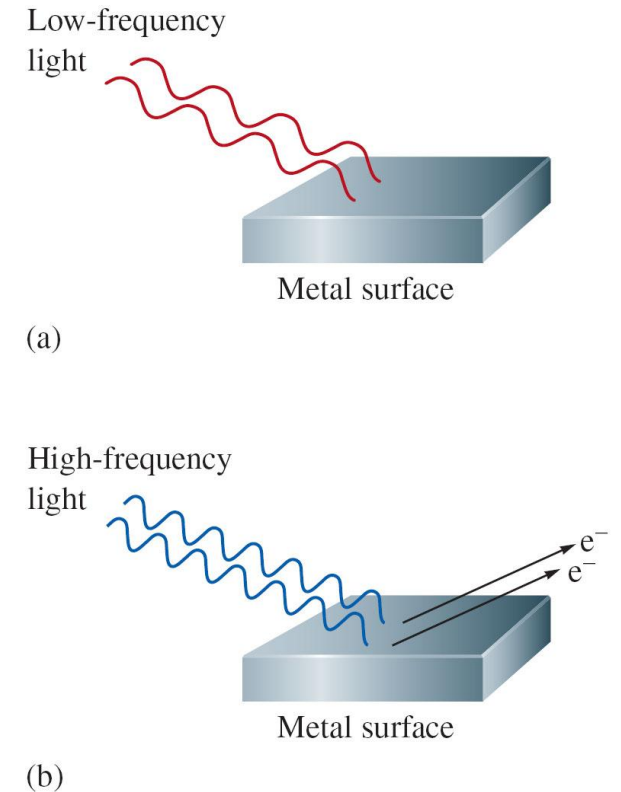
$$E_0 = h\nu_0$$

$$\text{for } \nu > \nu_0$$

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

$$E_{K_{electron}} = \frac{1}{2}mv^2 = h(\nu - \nu_0)$$

Light is quantized and exists as a stream of particles!



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

$$E = h\nu$$

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

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

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  $h\nu$  (where  $h = 6.626 \times 10^{-34}$  J·s, today  $h$  is known as Planck's constant).

$$\Delta E = nh\nu$$

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

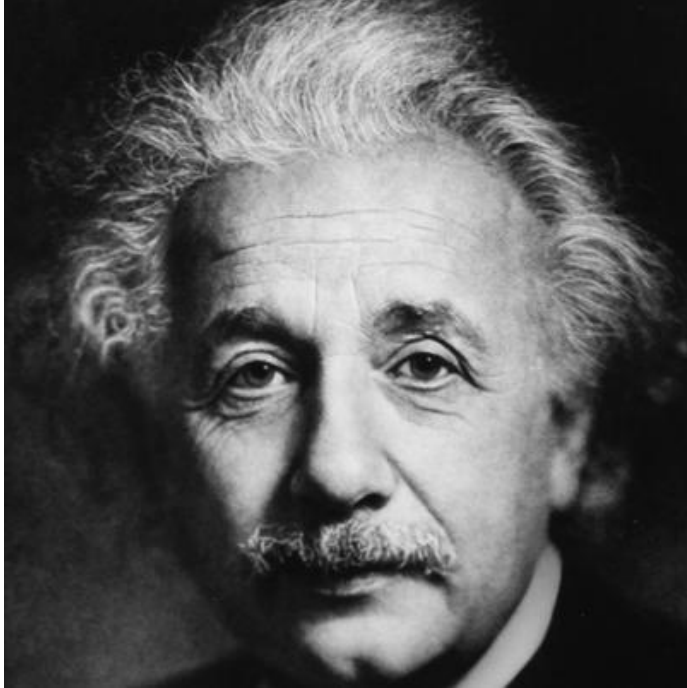
Frequency of  
EM radiation

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**:

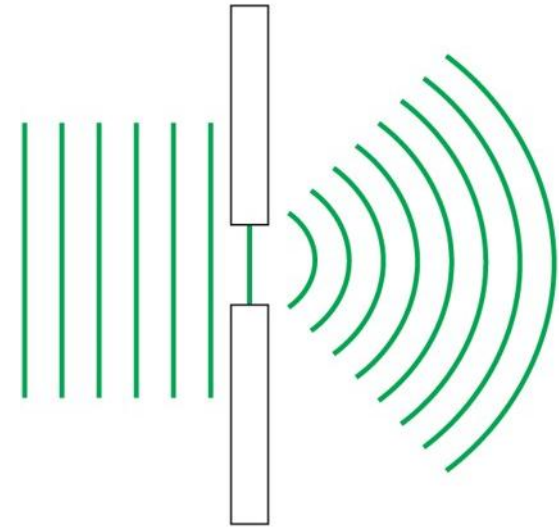
$$E_{\text{photon}} = h\nu = \frac{hc}{\lambda}$$

# Properties of Light

# Diffraction of Light

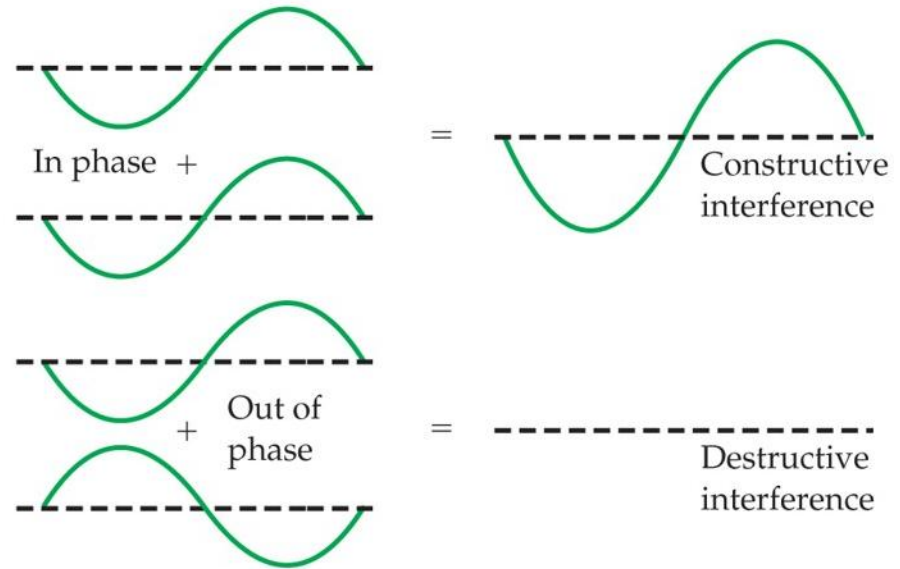
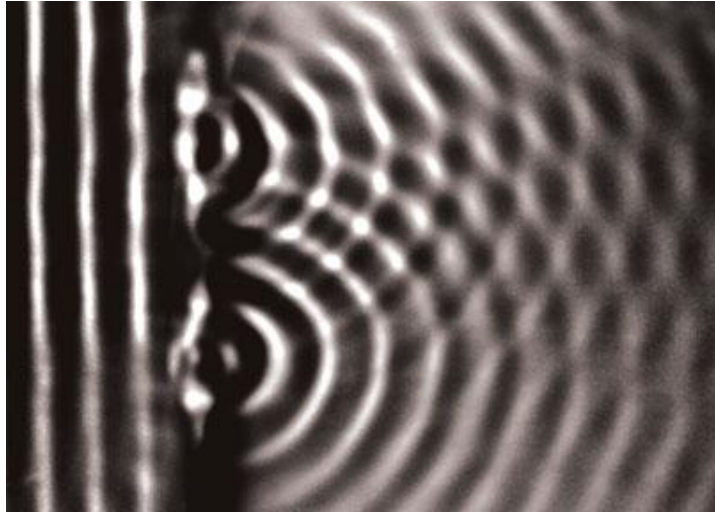


*Passage of a wave through a slit causes a diffraction pattern*



Diffraction is the bending of waves around an object.

# Diffraction



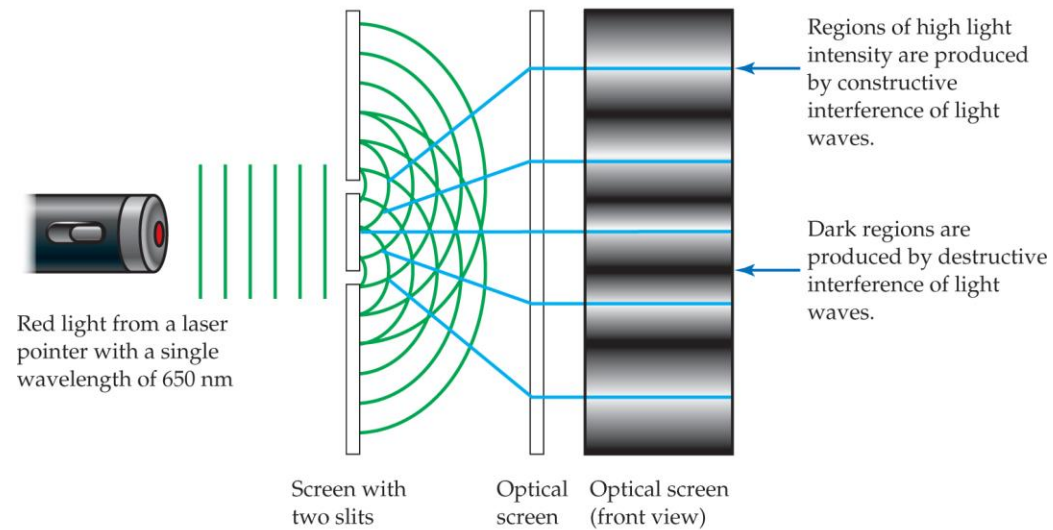
Constructive and destructive interference are wave properties

Diffraction and interference are phenomena exhibited by waves.

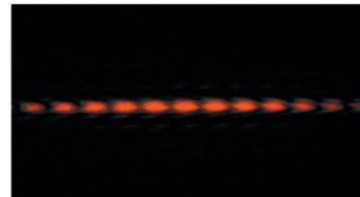
# Wavelike Properties of Light

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

(a)



(b)



The experimentally observed interference pattern from a red laser pointer passed through a double slit assembly with a width of 0.7 mm

# Wave/Particle Duality of Light



Albert Einstein



Light as a wave phenomenon



Light as a stream of photons

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.

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

- **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)

$$\lambda = \frac{h}{mv} \longleftarrow v = \text{velocity}$$

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

# Example

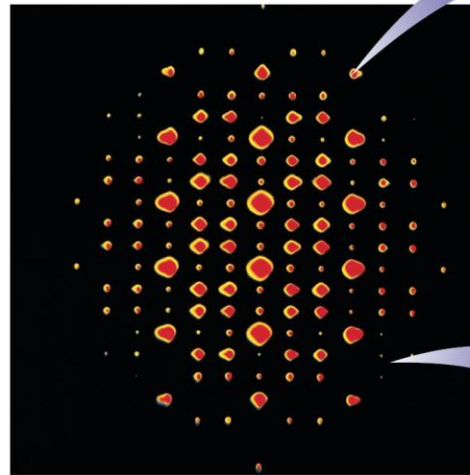
Compare the wavelength for an electron (mass =  $9.11 \times 10^{-31}$  kg) traveling at a speed of  $1.00 \times 10^7$  m/s and that of a baseball (mass = 0.10 kg) traveling at 35 m/s.

In SI units, J = kg x m<sup>2</sup> x s<sup>-2</sup>

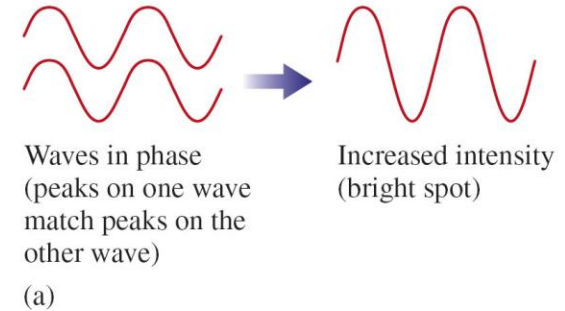
$$\lambda = \frac{h}{mv}$$

# Electron Diffraction Patterns

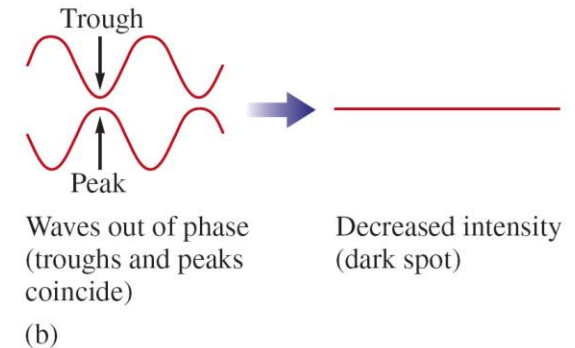
Beam of electrons directed onto a nickel crystal. The scattered radiation produces a diffraction pattern.



## Constructive interference



## Destructive interference



- 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

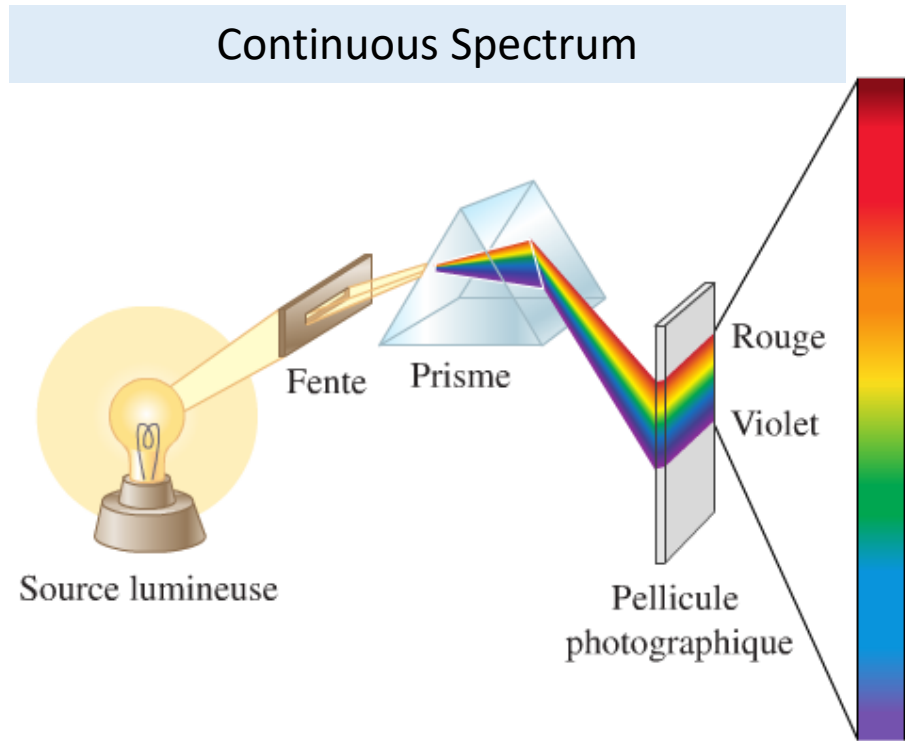
A spectrum is produced when light is separated into its component wavelengths.

- 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.

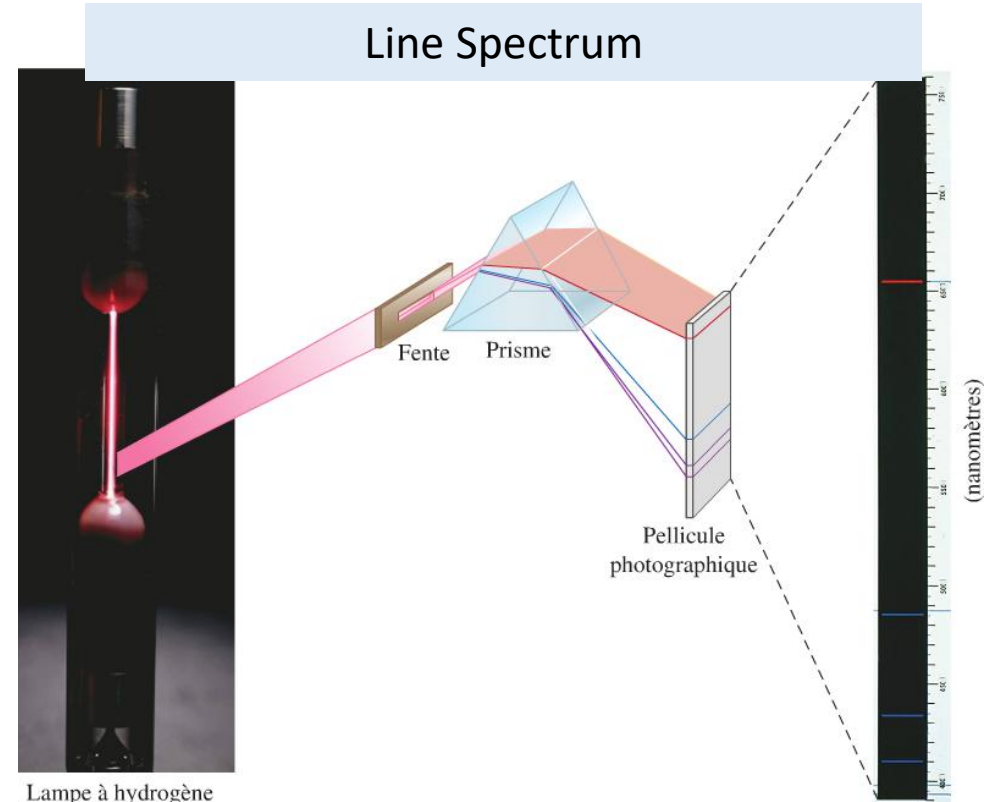
# Continuous Spectrum and Line Spectrum



Hydrogen emission spectrum



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



Emission Spectrum

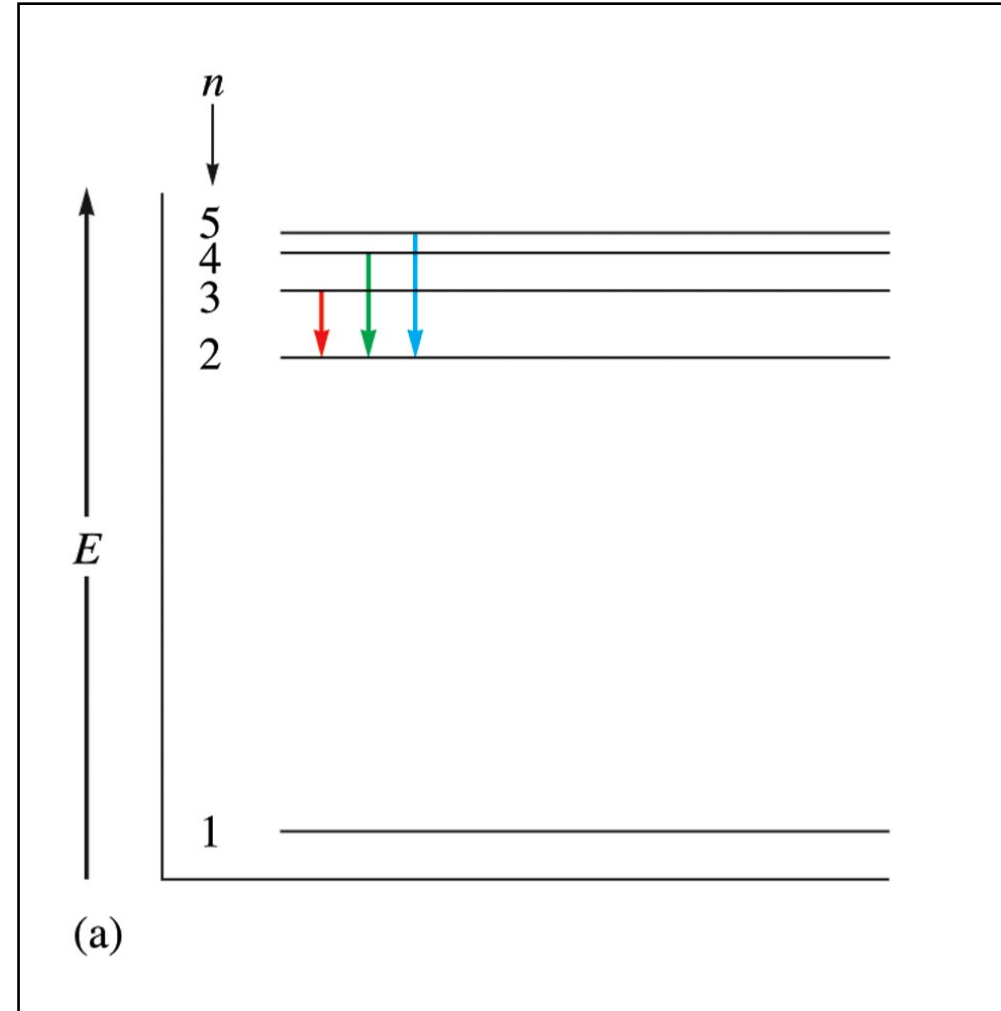
Series of discrete lines as a result of light emitted by excited (or heated) atoms.

# Electronic Transitions

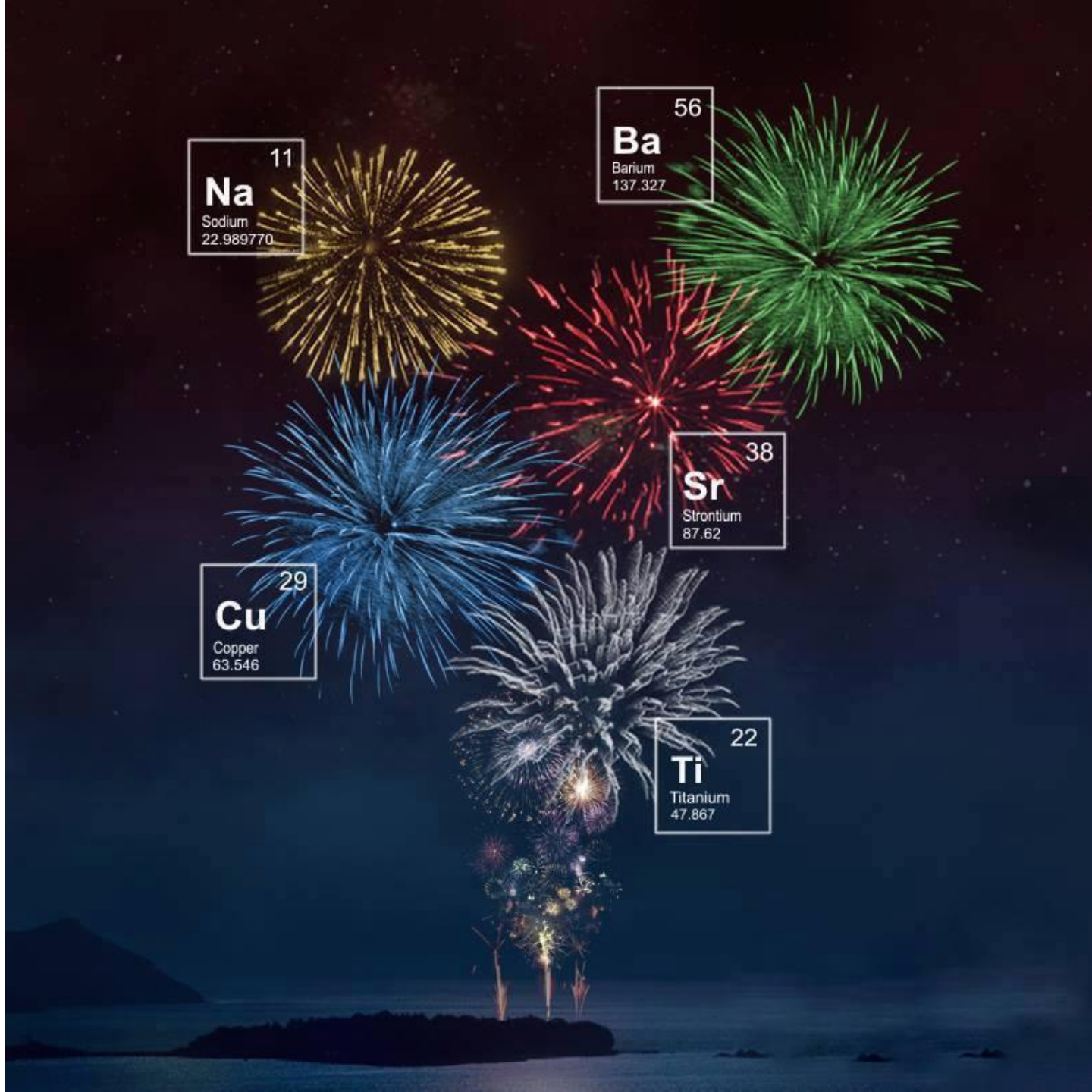
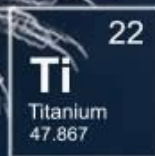
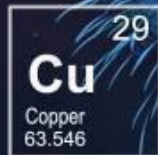
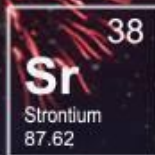
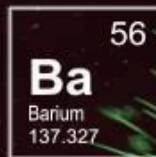
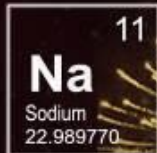
a) An Energy-Level Diagram for Electronic Transitions

( $n$  – principle quantum number)

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







# 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:

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

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

# Exercise

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

Solving for  $\lambda$  in a transition from  $n = 5$  to  $n = 2$ :

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

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

# 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

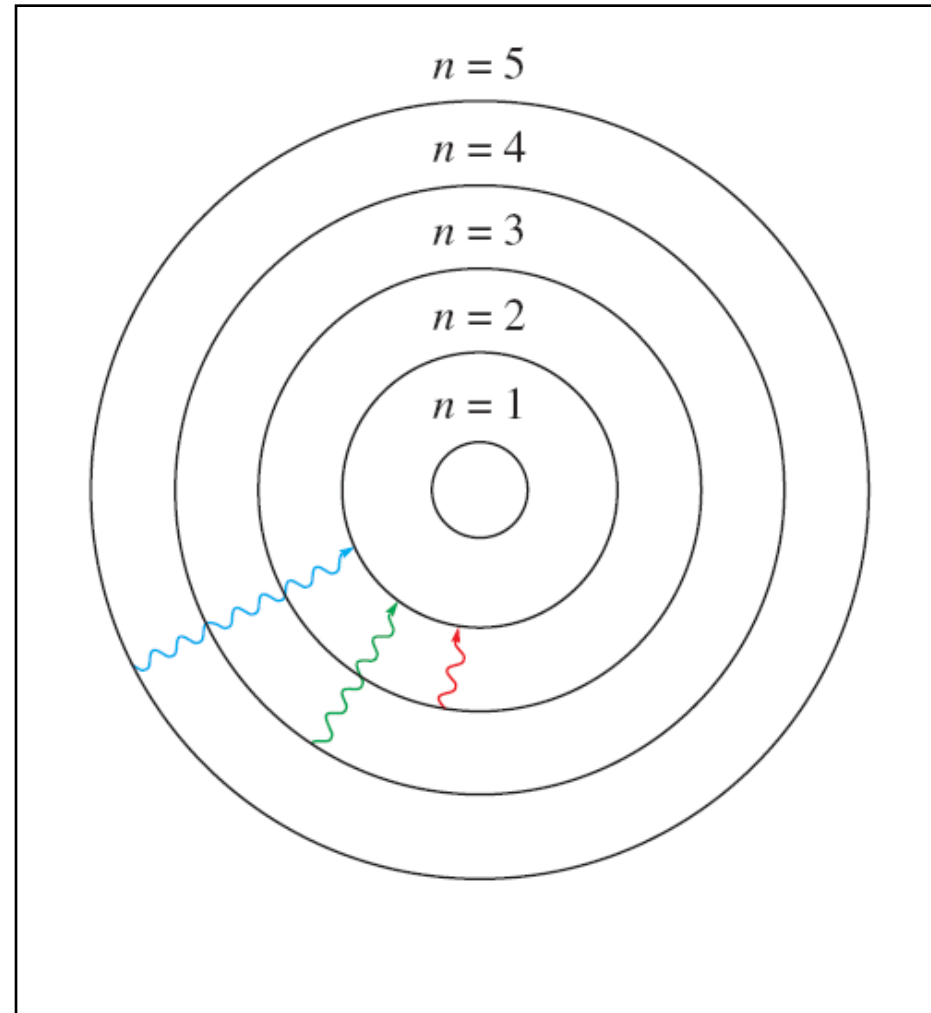
# The Bohr Model

- 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

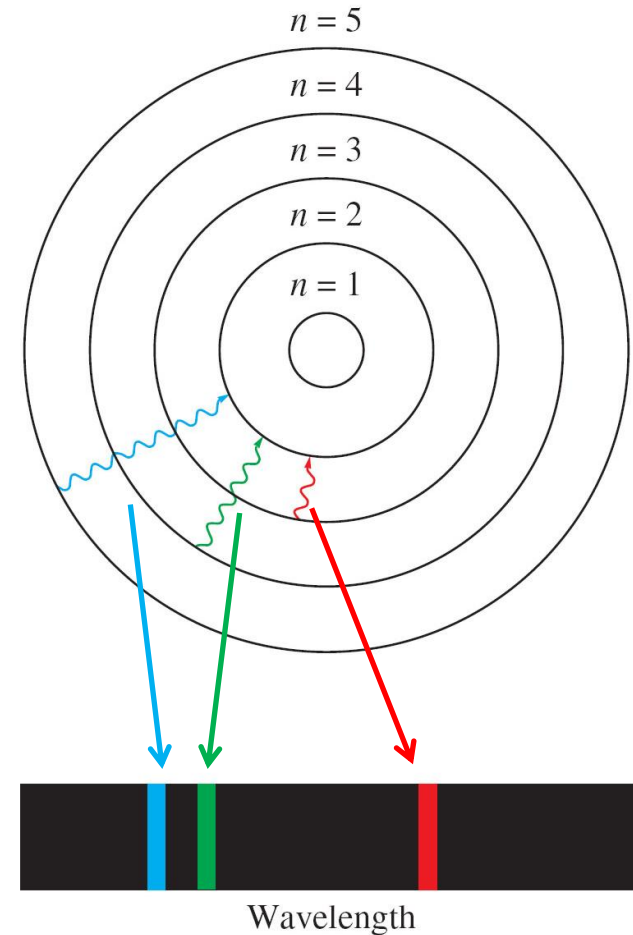
An Orbit-Transition Diagram, which accounts for the experimental spectrum

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

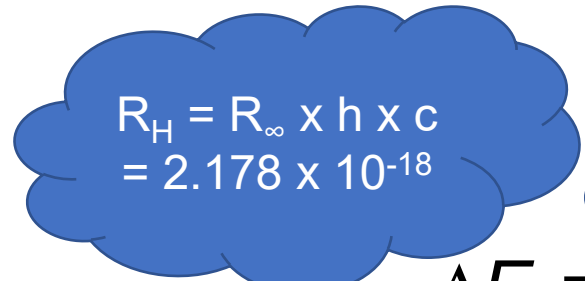
and  $\Delta E_{\text{photon}} = h\nu = \frac{hc}{\lambda} \longrightarrow \frac{1}{\lambda} = \frac{\Delta E_{\text{photon}}}{hc}$

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

Replace  $R_{\infty}$ ,  
h, and c with  
their actual  
values.....

# The Bohr Model

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


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

$\Delta E$  = change in energy of the atom (energy of the **emitted photon**)

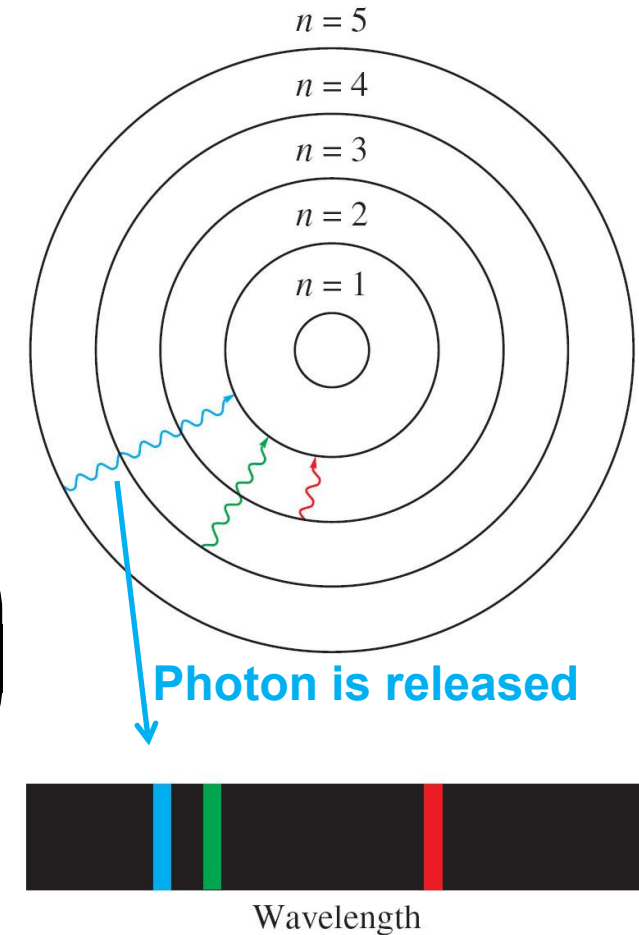
$n_{\text{final}}$  = integer; final distance from the nucleus

$n_{\text{initial}}$  = 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  $\Delta E$  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

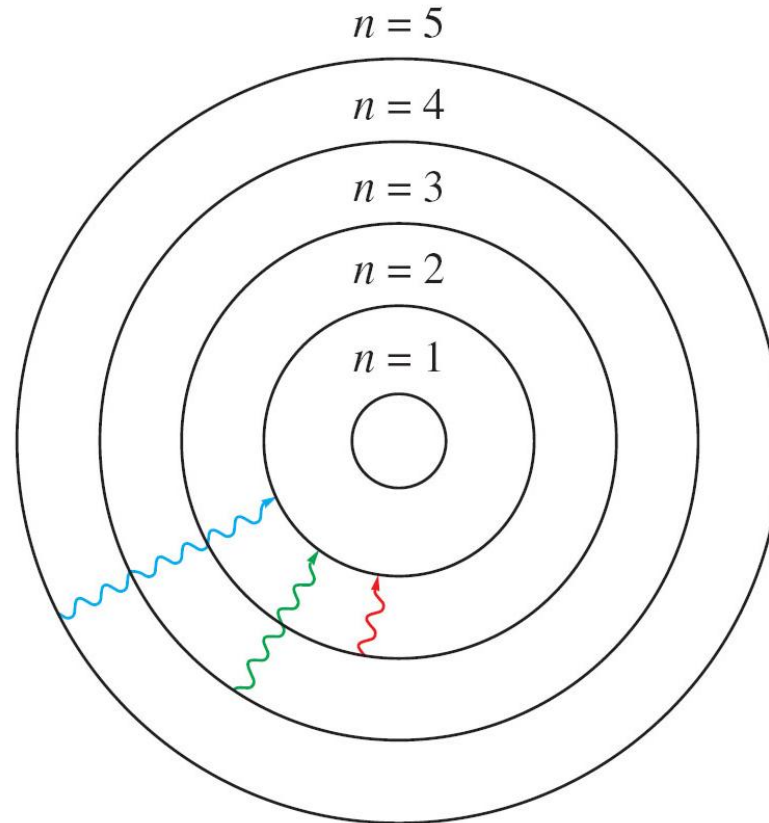
Two important points about the Bohr model:

- 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



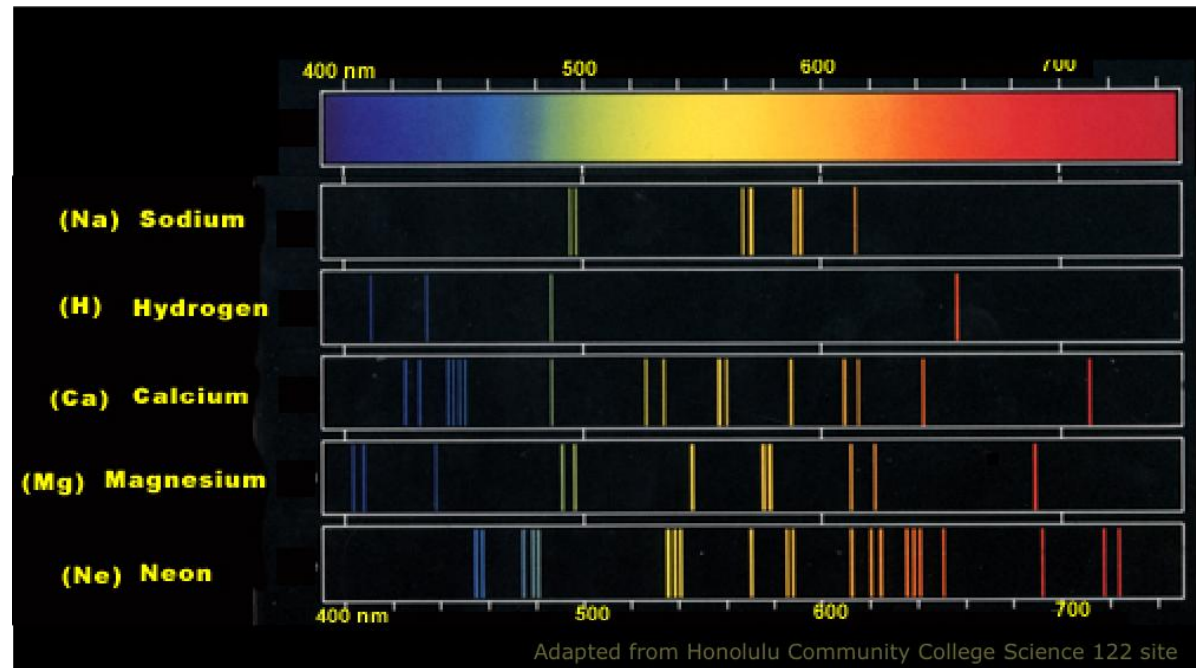
Niels Bohr



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

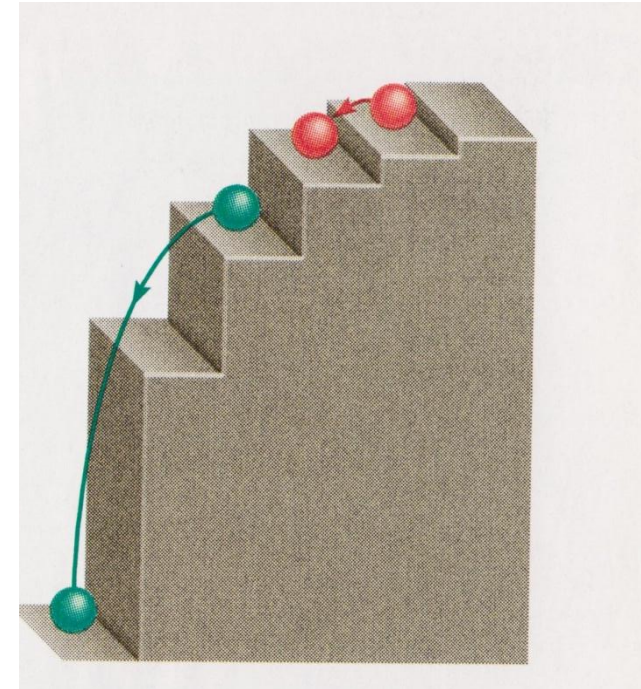
# The Bohr Model

- 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