Physical Electronics

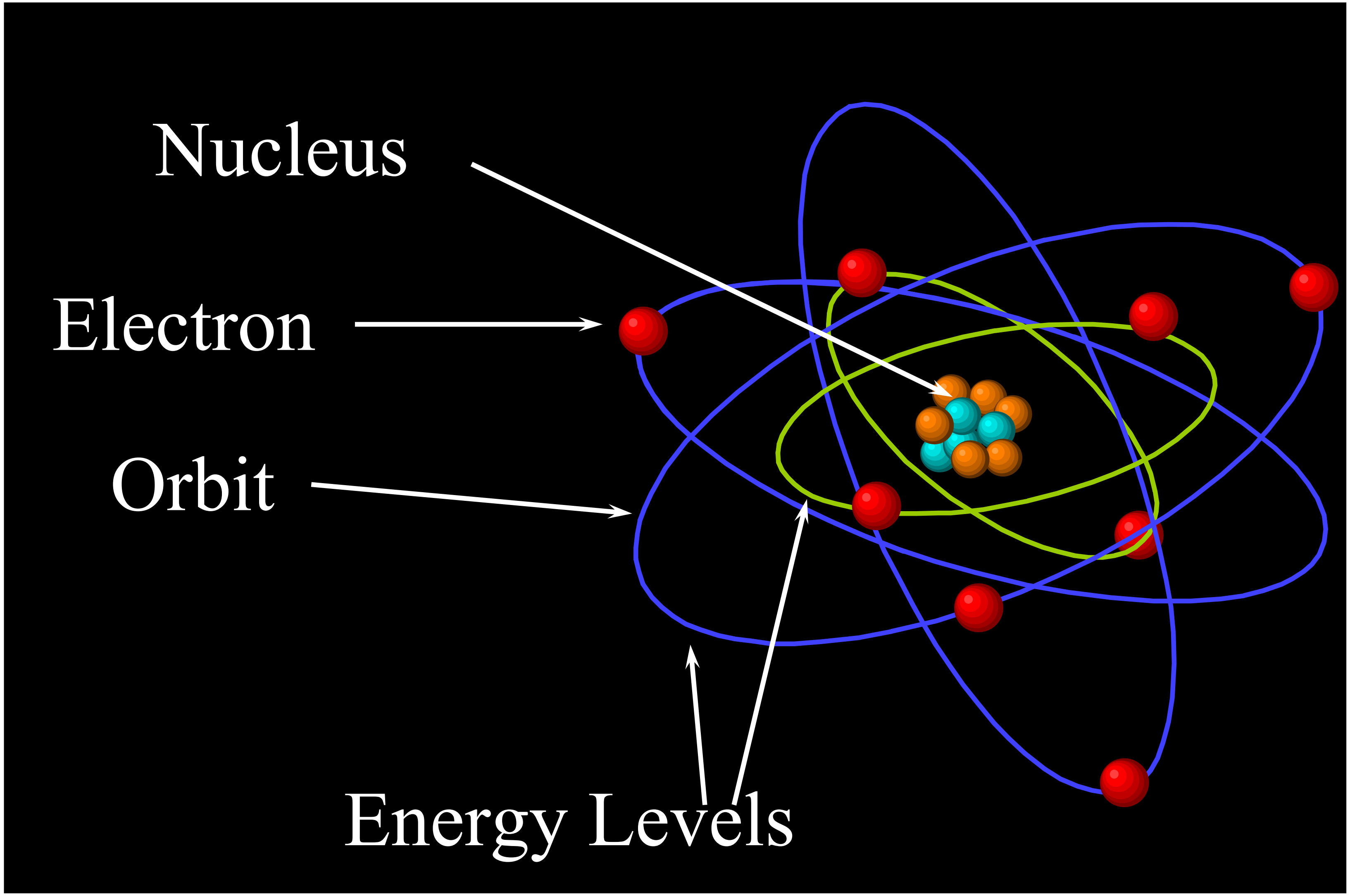
First class

(1)

Bohr's Model

- Why don't the electrons fall into the nucleus?
- Move like planets around the sun.
- In circular orbits at different levels.
- Amounts of energy separate one level from another.

Bohr's Model

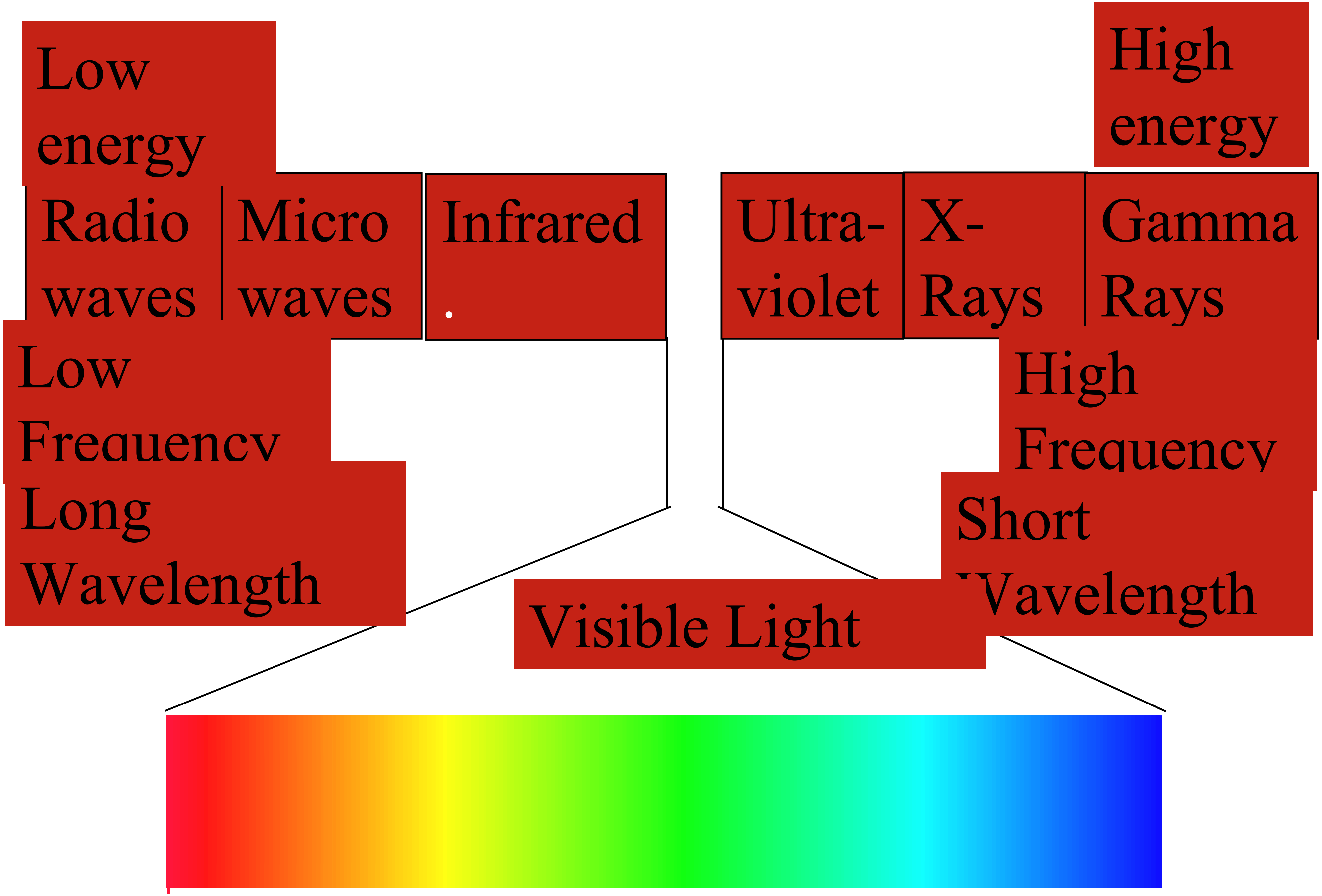


Bohr postulated that:

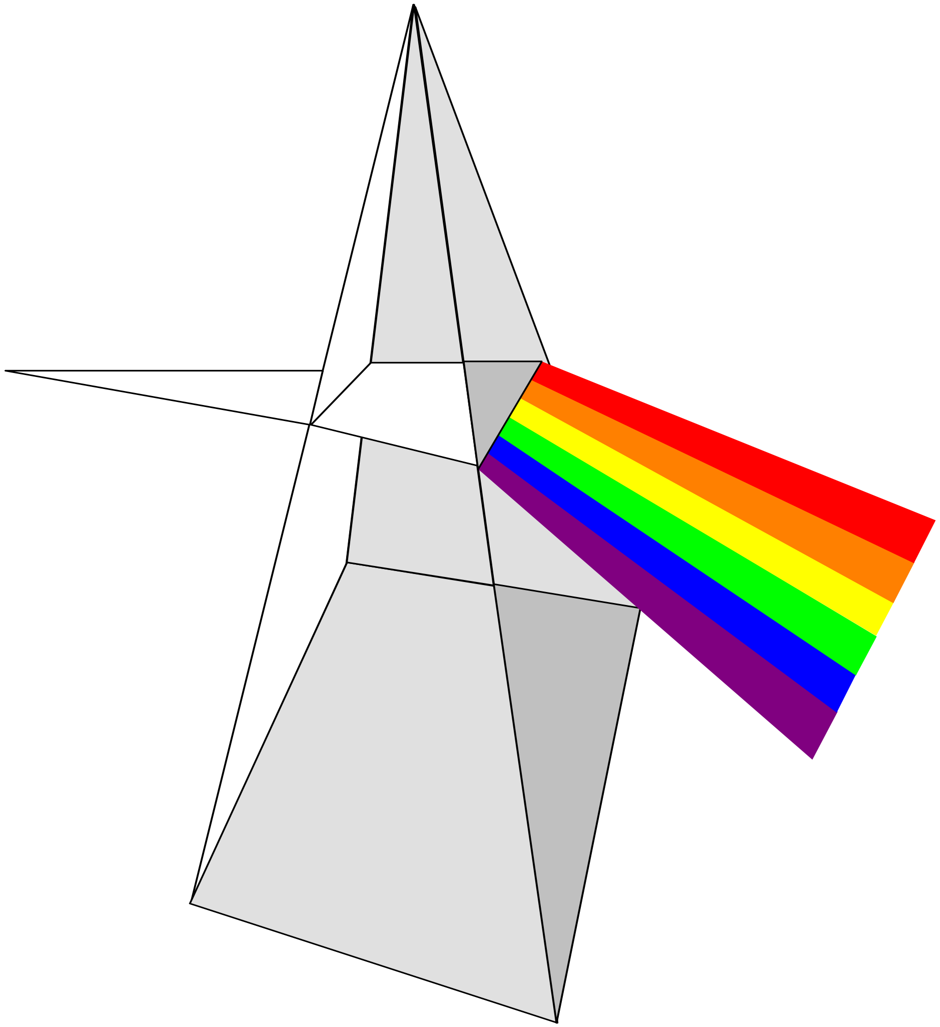
- Fixed energy related to the orbit
- Electrons cannot exist between orbits
- The higher the energy level, the further it is away from the nucleus
- An atom with maximum number of electrons in the outermost orbital energy level is stable (unreactive)

How did he develop his theory?

- He used mathematics to explain the visible spectrum of hydrogen gas

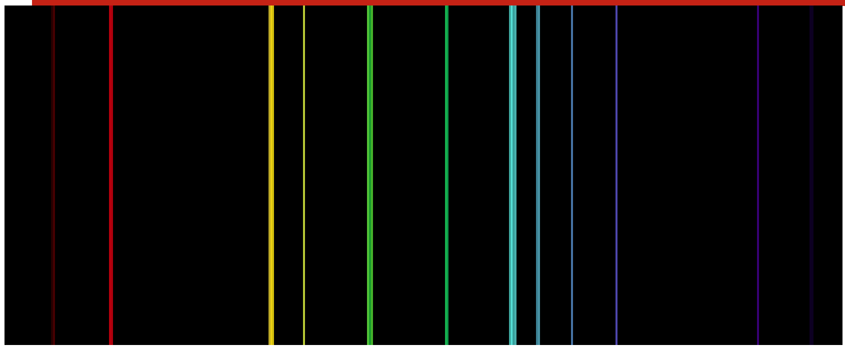

Low energy
Radio waves
Micro waves
Infrared
Ultra-violet
X-Rays
Gamma Rays
High energy
Low Frequency
High Frequency
Long Wavelength
Short Wavelength
Visible Light

The line spectrum

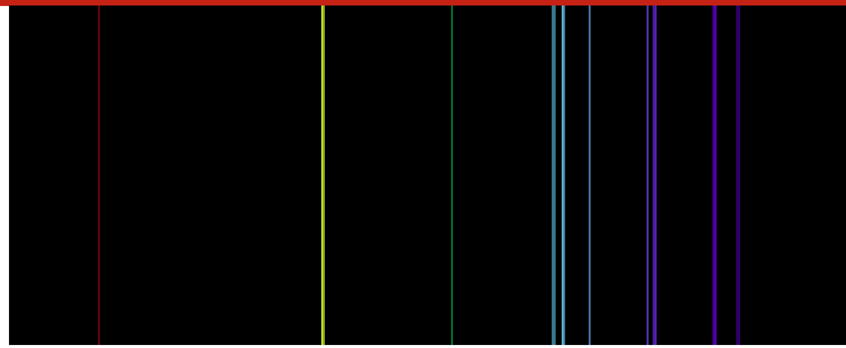


- electricity passed through a gaseous element emits light at a certain wavelength
- Can be seen when passed through a prism
- Every gas has a unique pattern (color)

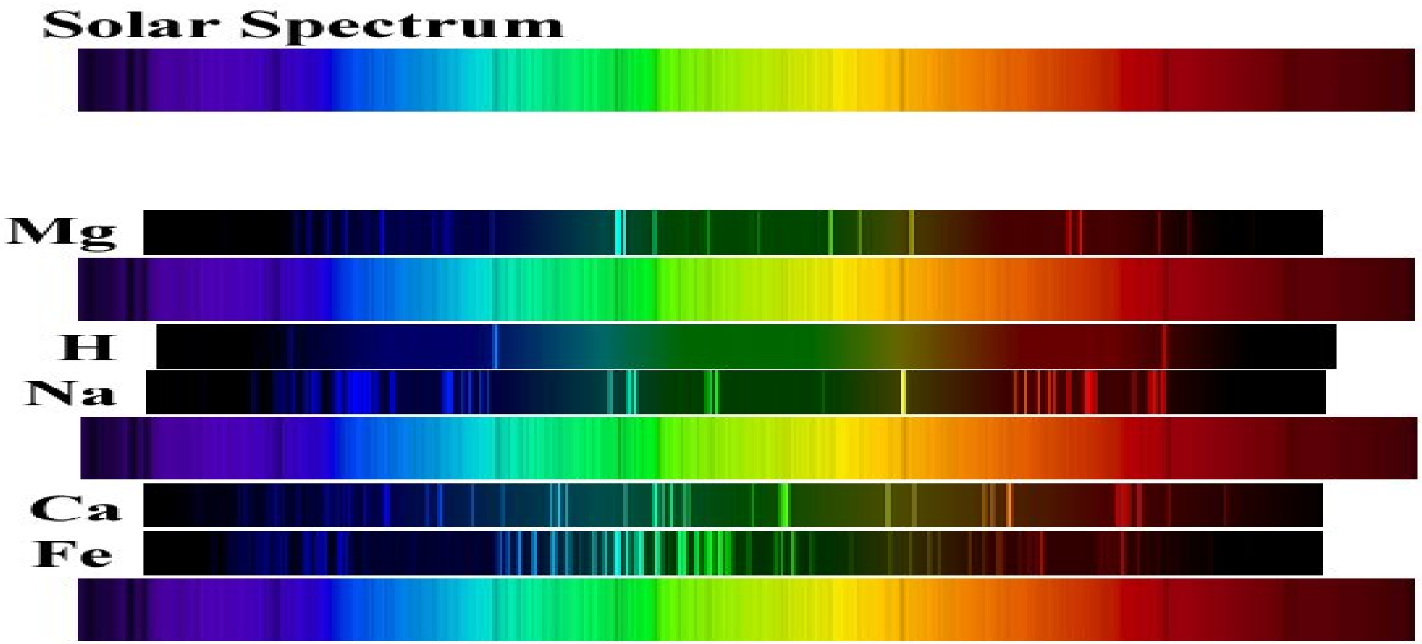
Line spectrum of various elements



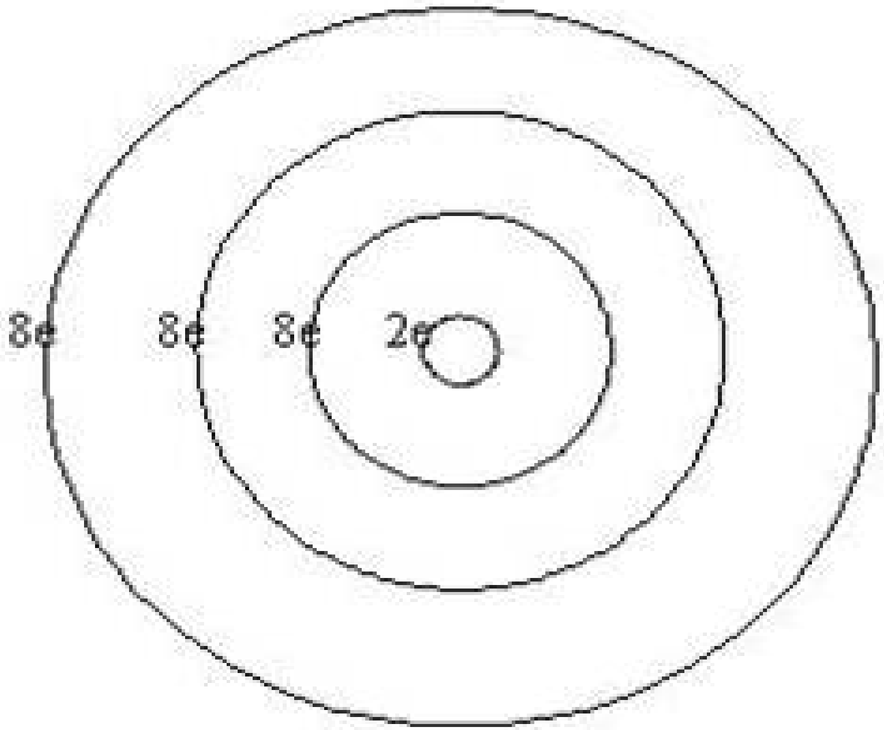
Carbon



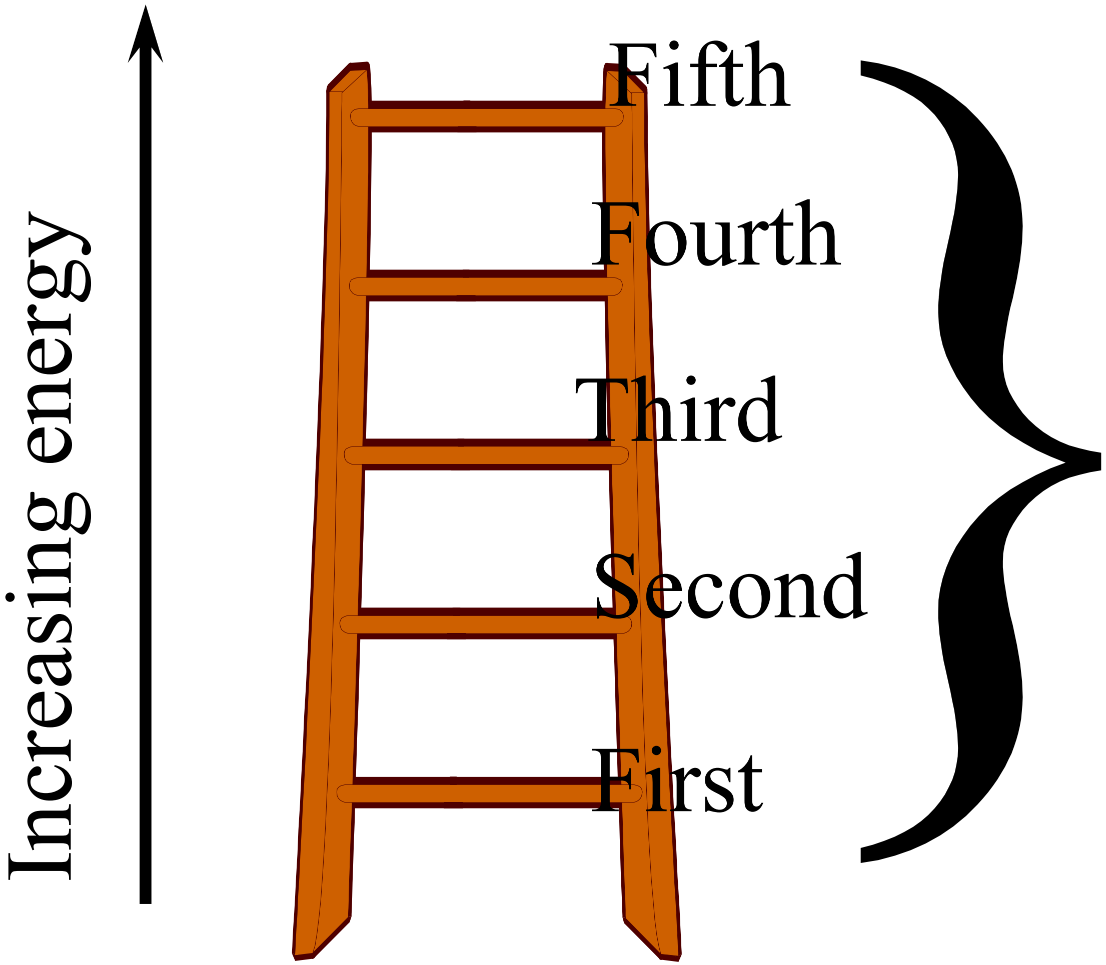
Helium



Drawback



- Bohr's theory did not explain or show the shape or the path traveled by the electrons.
- His theory could only explain hydrogen and not the more complex atoms



- Further away from the nucleus means more energy.
- There is no “in between” energy
- Energy Levels

The Quantum Mechanical Model

- Energy is quantized. It comes in chunks.
- A quanta is the amount of energy needed to move from one energy level to another.
- Since the energy of an atom is never “in between” there must be a quantum leap in energy.
- Schrödinger derived an equation that described the energy and position of the electrons in an atom

- Heated gases emit line spectra (v. heated solids)
- In 1885, J. J. Balmer showed that the wavelengths, λ, in the visible spectrum of hydrogen could be reproduced by a simple formula.

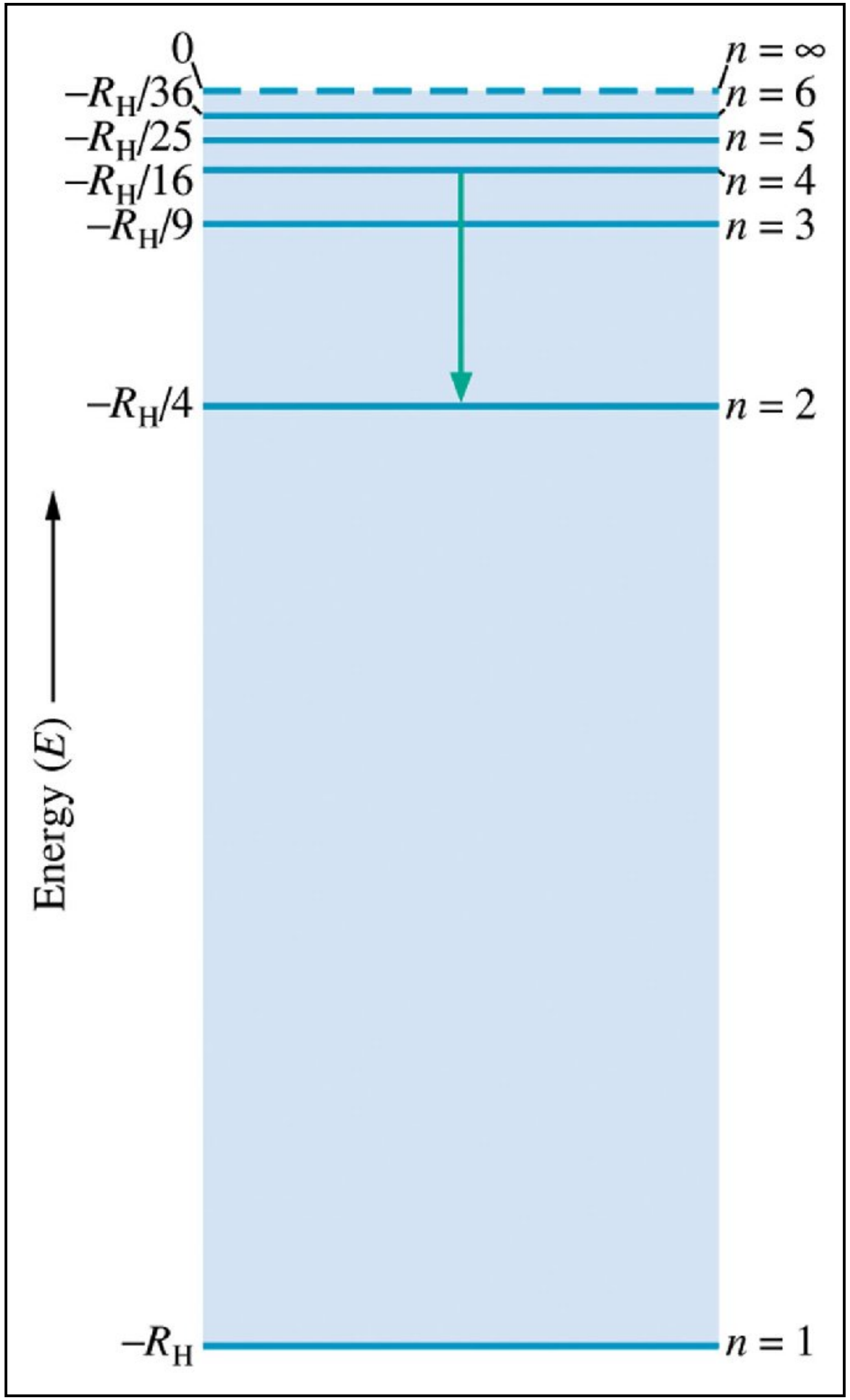
$$\frac{1}{\lambda} = 1.097 \times 10^{7} m^{-1} \left(\frac{1}{2^2} - \frac{1}{n^2}\right)$$

- The known wavelengths of the four visible lines for hydrogen correspond to values of n = 3, n = 4, n = 5, and n = 6.

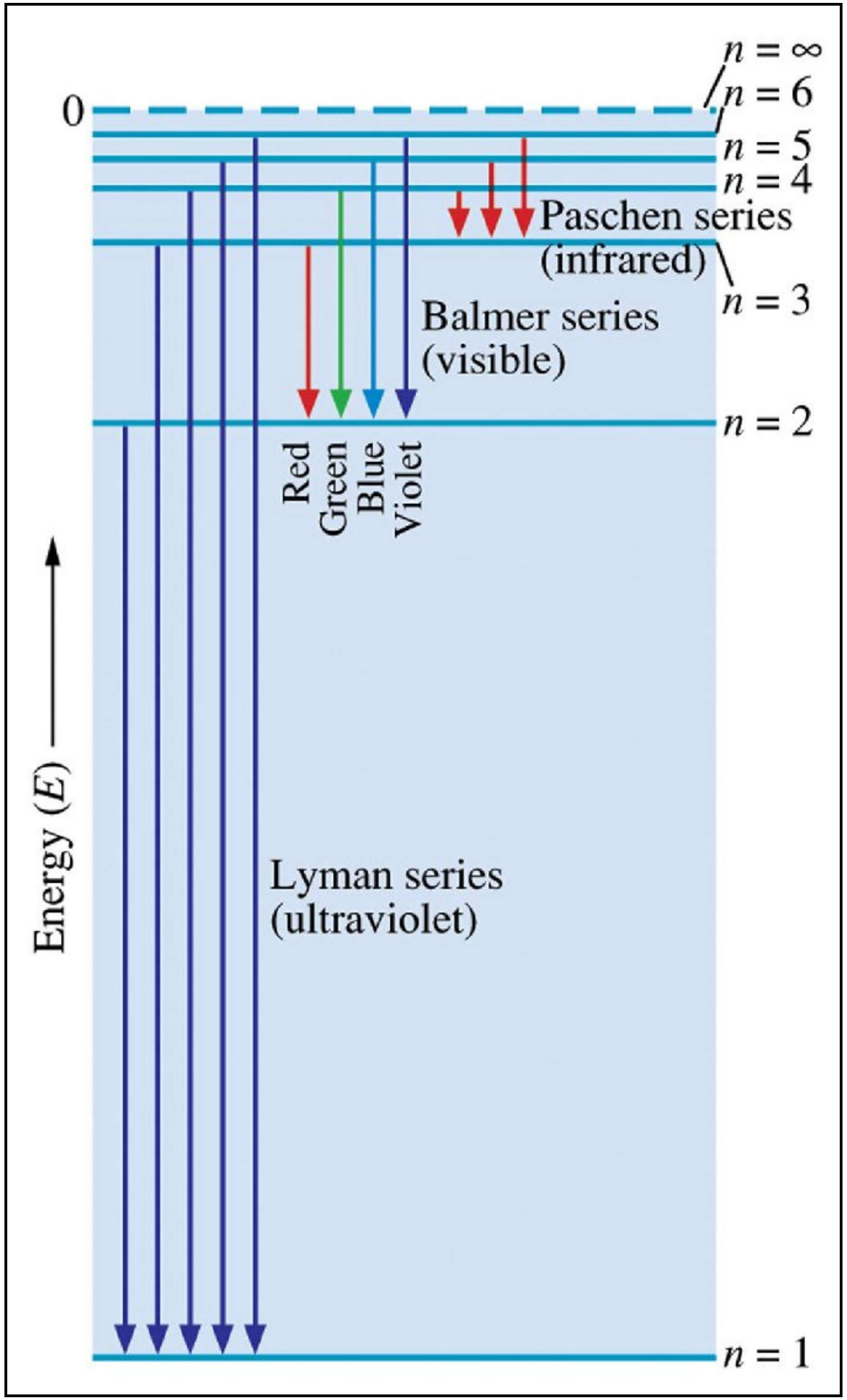
- Prior to the work of **Niels Bohr**, the stability of the atom could not be explained using the then-current theories. How can e^- lose energy and remain in orbit????
- Bohr in 1913 set down postulates to account for (1) the stability of the hydrogen atom and (2) the line spectrum of the atom.

1. **Energy level postulate** An electron can have only specific energy levels in an atom.
2. **Transitions between energy levels** An electron in an atom can change energy levels by undergoing a “transition” from one energy level to another.

Energy-level Diagram for the Electron in the Hydrogen Atom



Transitions of the Electron in the Hydrogen Atom



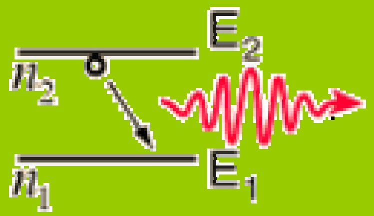
Bohr Theory of the Hydrogen Atom

- Bohr’s Postulates
  - Bohr’s theory explains not only the emission of light, but also the absorption of light.
  - When an electron falls from n = 3 to n = 2 energy level, a photon of red light (wavelength, 685 nm) is emitted.
  - When red light of this same wavelength shines on a hydrogen atom in the n = 2 level, the energy is gained by the electron that undergoes a transition to n = 3.

- Bohr's theory established the concept of atomic energy levels but did not thoroughly explain the "wave-like" behavior of the electron.

- Current ideas about atomic structure depend on the principles of **quantum mechanics**, a theory that applies to subatomic particles such as electrons. Electrons show properties of both waves and particles.

The Bohr model for an electron transition in hydrogen between quantized energy levels with different quantum numbers n yields a photon by emission, with quantum energy



A downward transition involves emission of a photon of energy:

$$E_{photon} = h\nu = E_2 - E_1$$

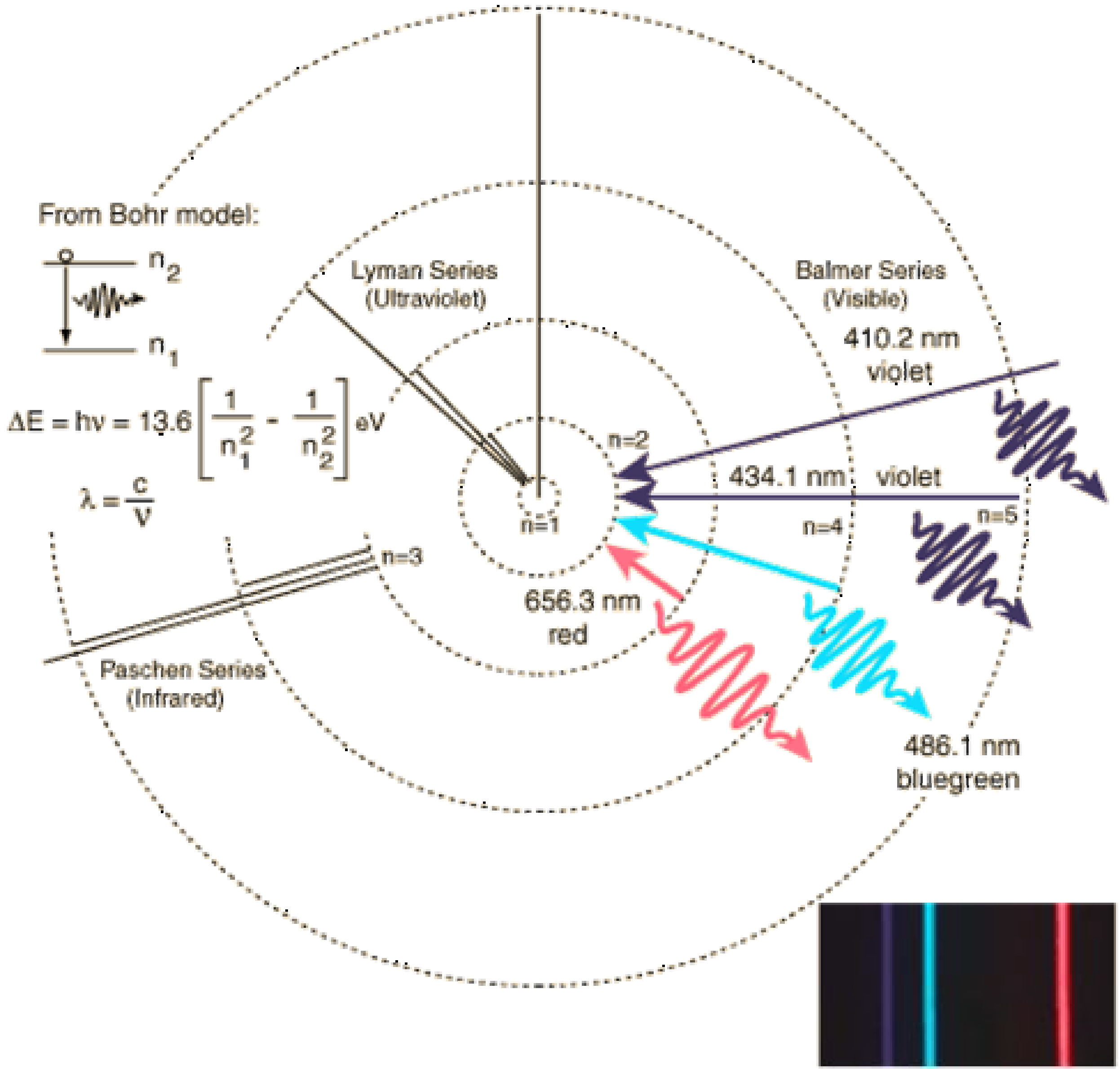
Given the expression for the energies of the hydrogen electron states:

$$h\nu = \frac{me^4}{8\varepsilon_0{}^2 h^2}\left[\frac{1}{n_1{}^2} - \frac{1}{n_2{}^2}\right] = 13.6\left[\frac{1}{n_1{}^2} - \frac{1}{n_2{}^2}\right] \text{ eV}$$

This is often expressed in terms of the inverse wavelength or "wave number" as follows:

$$\frac{1}{\lambda} = R_H\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] \quad \text{where} \quad R_H = \frac{me^4}{8\varepsilon_0{}^2 ch^3}$$

$$R_H = 1.097\ 10^{-7}\ \text{m}^{-1}$$


From Bohr model:
n2
n1
ΔE = hν = 13.6 [1/n1² - 1/n2²] eV
λ = c/ν
Lyman Series
(Ultraviolet)
Balmer Series
(Visible)
410.2 nm
violet
434.1 nm
violet
n=2
n=1
n=3
n=4
n=5
656.3 nm
red
486.1 nm
bluegreen
Paschen Series
(Infrared)

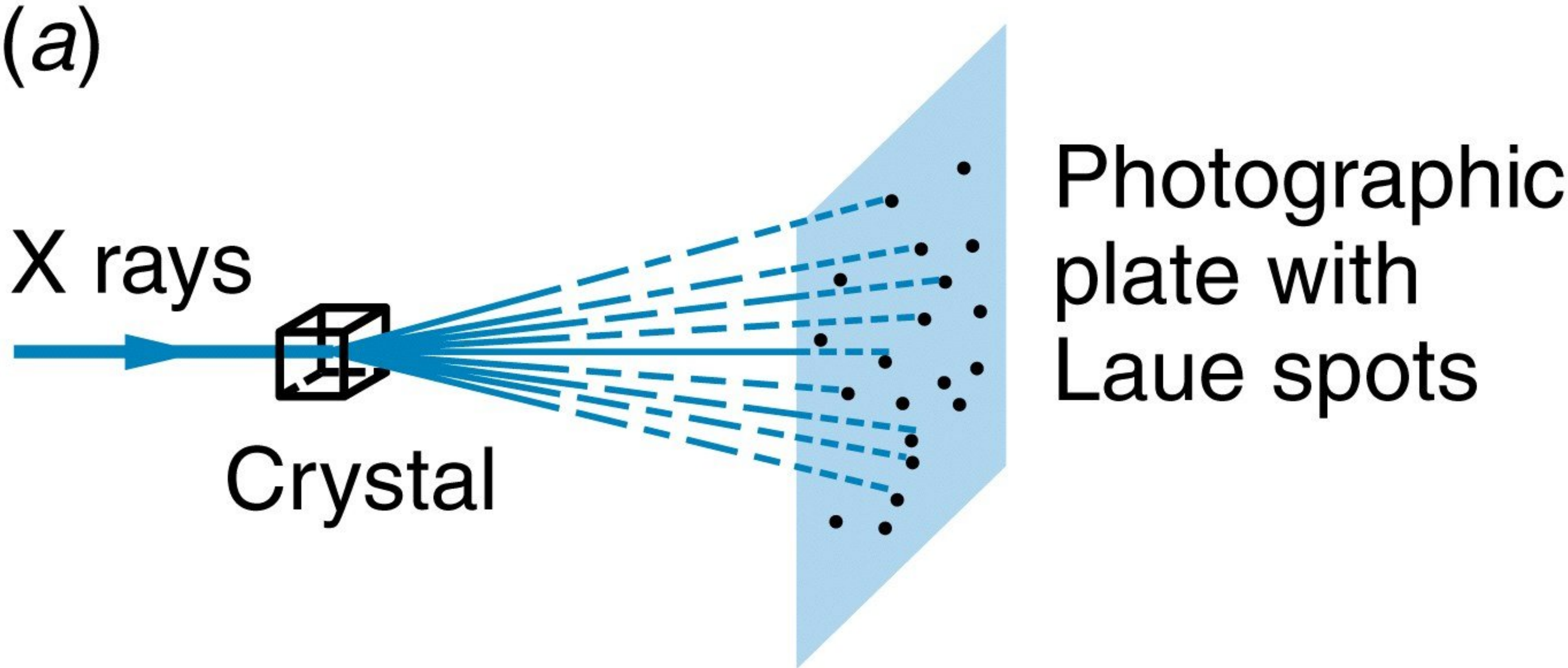
Failures of the Bohr Model

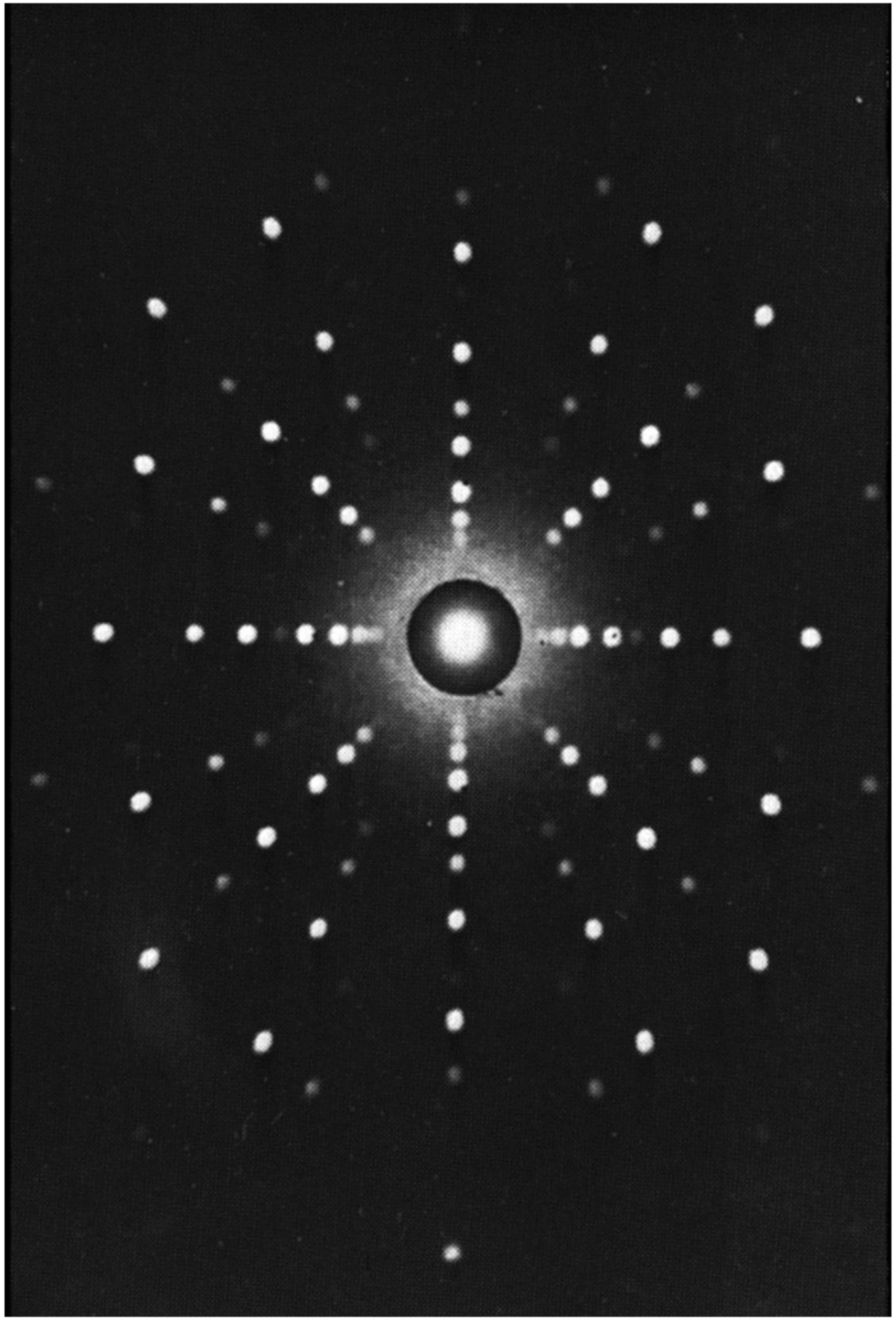
While the Bohr model was a major step toward understanding the quantum theory of the atom, it is not in fact a correct description of the nature of electron orbits. Some of the shortcomings of the model are:

1. It fails to provide any understanding of why certain spectral lines are brighter than others. There is no mechanism for the calculation of transition probabilities.
2. The Bohr model treats the electron as if it were a miniature planet, with definite radius and momentum. This is in direct violation of the uncertainty principle which dictates that position and momentum cannot be simultaneously determined.

The Bohr model gives us a basic conceptual model of electrons orbits and energies. The precise details of spectra and charge distribution must be left to quantum mechanical calculations, as with the Schrödinger equation.

Bragg Diffraction



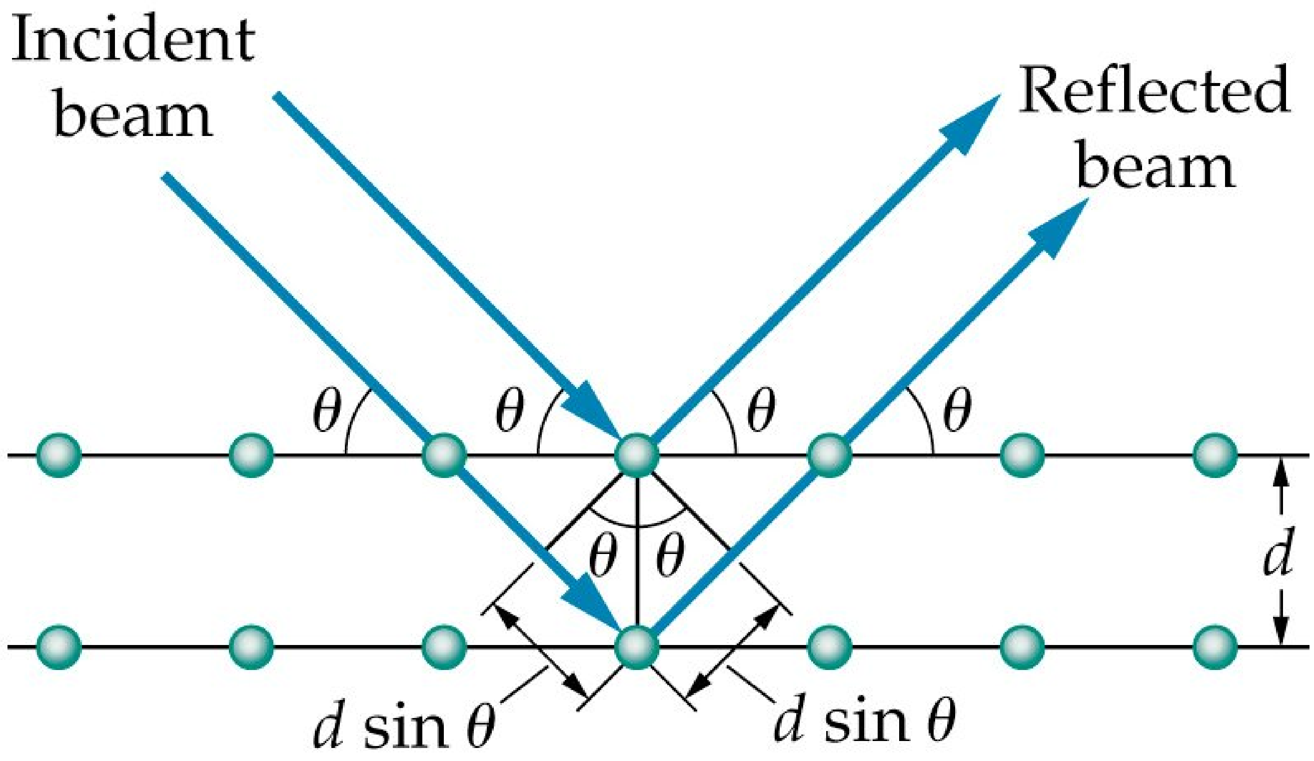


Diffraction

Pathlength difference $= 2d\sin\theta$.
Constructive Interference:
$2d\sin\theta = n\lambda$

$\lambda = h/p$

Photon $E=pc$
Slow Neutron:
$E = p^2/(2M)$



See pictures in text

Wave-Particle Duality: Light

Does light consist of particles or waves? When one focuses upon the different types of phenomena observed with light, a strong case can be built for a wave picture:

Phenomenon	Can be explained in terms of waves.	Can be explained in terms of particles.
Reflection	✓	✓
Refraction	✓	✓
Interference	✓	$\otimes$
Diffraction	✓	$\otimes$
Polarization	✓	$\otimes$
Photoelectric effect	$\otimes$	✓
Compton scattering	$\otimes$	✓

Most commonly observed phenomena with light can be explained by waves. But the photoelectric effect and the Compton scatering suggested a particle nature for light. Then electrons too were found to exhibit dual natures.

- The first clue in the development of quantum theory came with the discovery of the **de Broglie relation.**
- In 1923, Louis de Broglie reasoned that if light exhibits particle aspects, perhaps particles of matter show characteristics of waves.
- He postulated that a particle with mass ***m*** and a velocity ***v*** has an associated wavelength.
- The equation $\lambda = h/mv$ is called the **de Broglie relation.**

de Broglie Waves and the Bohr Model

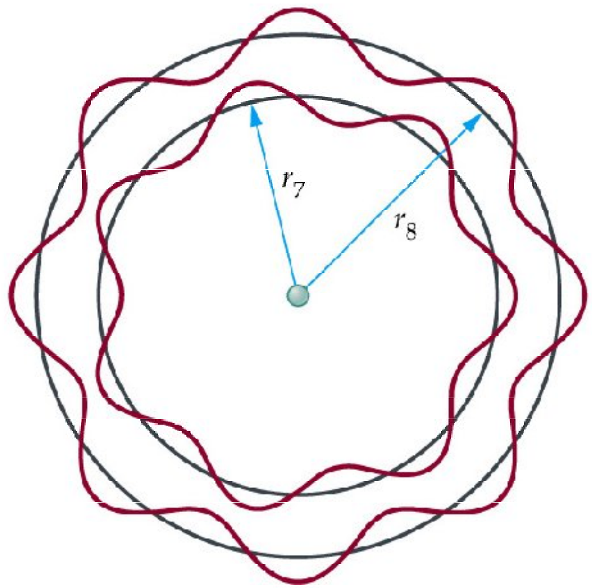
Why is the angular momentum of the electron restricted to certain values?

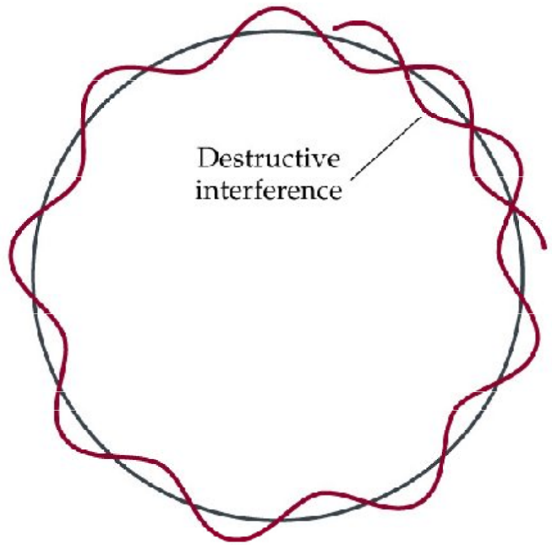
- The electron has a wavelength and forms standing waves in its orbit around the nucleus.
- An integral number of electron wavelengths must fit into the circumference of the circular orbit.

$$n\lambda = 2\pi r$$

$$p = mv = \frac{h}{\lambda} = \frac{h}{(2\pi r / n)} = \frac{nh}{2\pi r} \quad n = 1, 2, 3, \ldots$$

$$L = rmv = \frac{nh}{2\pi} \quad n = 1, 2, 3, \ldots$$





(b)

- If matter has wave properties, why are they not commonly observed?
  - The de Broglie relation shows that a baseball (0.145 kg) moving at about 60 mph (27 m/s) has a wavelength of about 1.7 x 10^{-34} m.

$$\lambda = \frac{6.63\times10^{-34}\ \frac{\mathrm{kg\cdot m^2}}{\mathrm{s}}}{(0.145\ \mathrm{kg})(27\ \mathrm{m/s})} = 1.7\times10^{-34}\ \mathrm{m}$$

  - es

  - Electrons have wavelengths on the order of a few picometers (1 pm = 10^{-12} m).

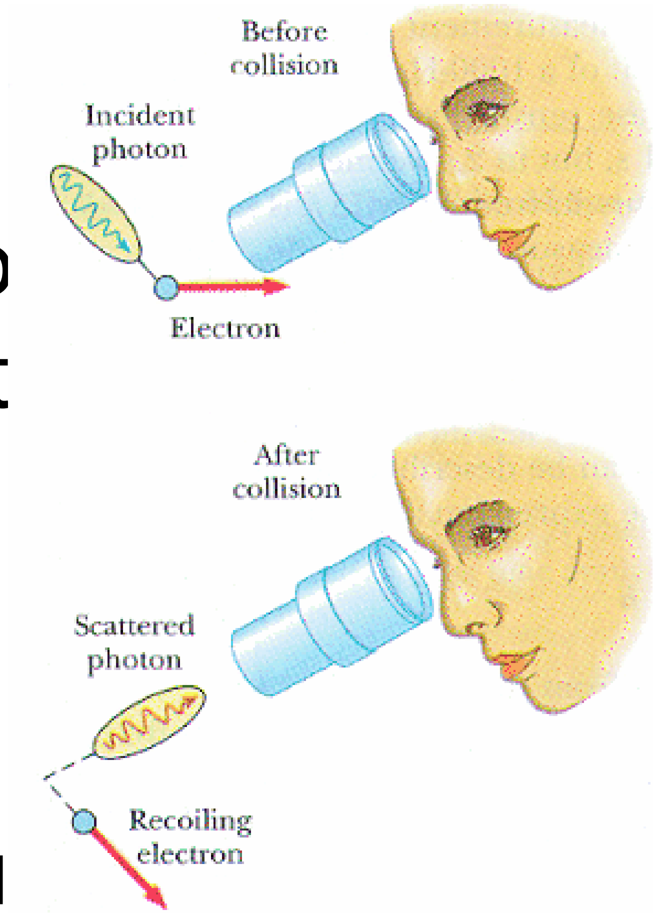
Quiz

Which of the following particles has the **shortest** wavelength? **($\lambda = h/mv$)**

a. an electron traveling at x m/s
b. a proton traveling at x m/s
c. a proton traveling at 2x m/s

Measuring the position and momentum of an electron

- Shine light on electron and detect reflected light using a microscope
- Minimum uncertainty in positio is given by the wavelength of t light
- So to determine the position accurately, it is necessary to u light with a short wavelength



Measuring the position and momentum of an electron (cont'd)

- By Planck's law $E = hc/\lambda$, a photon with a short wavelength has a large energy
- Thus, it would impart a large 'kick' to the electron
- But to determine its momentum accurately, electron must only be given a small kick
- This means using light of long wavelength!

- Use light with short wavelength:
  - accurate measurement of position but not momentum
- Use light with long wavelength:
  - accurate measurement of momentum but not position

Heisenberg's Uncertainty Principle

uncertainty in momentum

$$\Delta x \, \Delta p \geq \frac{h}{4\pi}$$

uncertainty in position

The more accurately you know the position (i.e., the smaller Δx is) , the less accurately you know the momentum (i.e., the larger Δp is); and vice versa

Heisenberg's Uncertainty Principle involving energy and time

$$\Delta E \, \Delta t \geq \frac{h}{4\pi}$$

- The more accurately we know the energy of a body,
  the less accurately we know how long it possessed
  that energy
- The energy can be known with perfect precision ($\Delta E = 0$), only if the measurement is made over an infinite period of time ($\Delta t = \infty$)

- **Quantum mechanics** is the branch of physics that mathematically describes the wave properties of submicroscopic particles.
  - We can no longer think of an electron as having a precise orbit in an atom.
  - To describe such an orbit would require knowing its exact position and velocity.
  - In 1927, Werner Heisenberg showed (from quantum mechanics) that it is <u>impossible</u> to know both simultaneously.

- **Heisenberg's uncertainty principle** is a relation that states that the product of the uncertainty in position (Δx) and the uncertainty in momentum ($m\Delta v_x$) of a particle can be no smaller than **h/4**π.

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

– When ***m*** is large (for example, a baseball) the uncertainties are small, but for electrons, high uncertainties disallow defining an exact orbit.

Quantum Mechanics

- Although we cannot precisely define an electron's orbit, we can obtain the ***probability*** of finding an electron at a given point around the nucleus.
  - Erwin Schrodinger defined this probability in a mathematical expression called a **wave function**, denoted ψ (psi).
  - The probability of finding a particle in a region of space is defined by ψ^2.

Quantum Numbers and Atomic Orbitals

- According to quantum mechanics, each electron is described by four quantum numbers.
  - Principal quantum number (n)
  - Angular momentum quantum number (l)
  - Magnetic quantum number (m_l)
  - Spin quantum number (m_s)
    - The first three define the wave function for a particular electron. The fourth quantum number refers to the magnetic property of electrons.

Quantum Numbers and Atomic Orbitals

- The **principal quantum number(n)** represents the "shell number" in which an electron "resides"
  - The smaller n is, the smaller the orbital
  - The smaller n is, the lower the energy of the electron

Quantum Numbers and Atomic Orbitals

- The **angular momentum quantum number (l)** distinguishes “sub shells” within a given shell that have different shapes.
  - Each main “shell” is subdivided into “sub shells.” Within each shell of quantum number n, there are n sub shells, each with a distinctive shape.
  - l can have any integer value from 0 to (n - 1)
  - The different subshells are denoted by letters.

Letter	**s**	**p**	**d**	**f**	**g**	**…**
l	**0**	**1**	**2**	**3**	**4**	**….**

Quantum Numbers and Atomic Orbitals

- The **magnetic quantum number (m_l)** distinguishes orbitals within a given sub-shell that have different shapes and orientations in space.
  - Each sub shell is subdivided into “orbitals,” each capable of holding a pair of electrons.
  - m_l can have any integer value from -l to +l.
  - Each orbital within a given sub shell has the same energy.

Quantum Numbers and Atomic Orbitals

- The **spin quantum number (m_s)** refers to the two possible spin orientations of the electrons residing within a given orbital.
  - Each orbital can hold only two electrons whose spins must oppose one another.
  - The possible values of m_s are +1/2 and –1/2.