Atomic Emission and Quantum Mechanics

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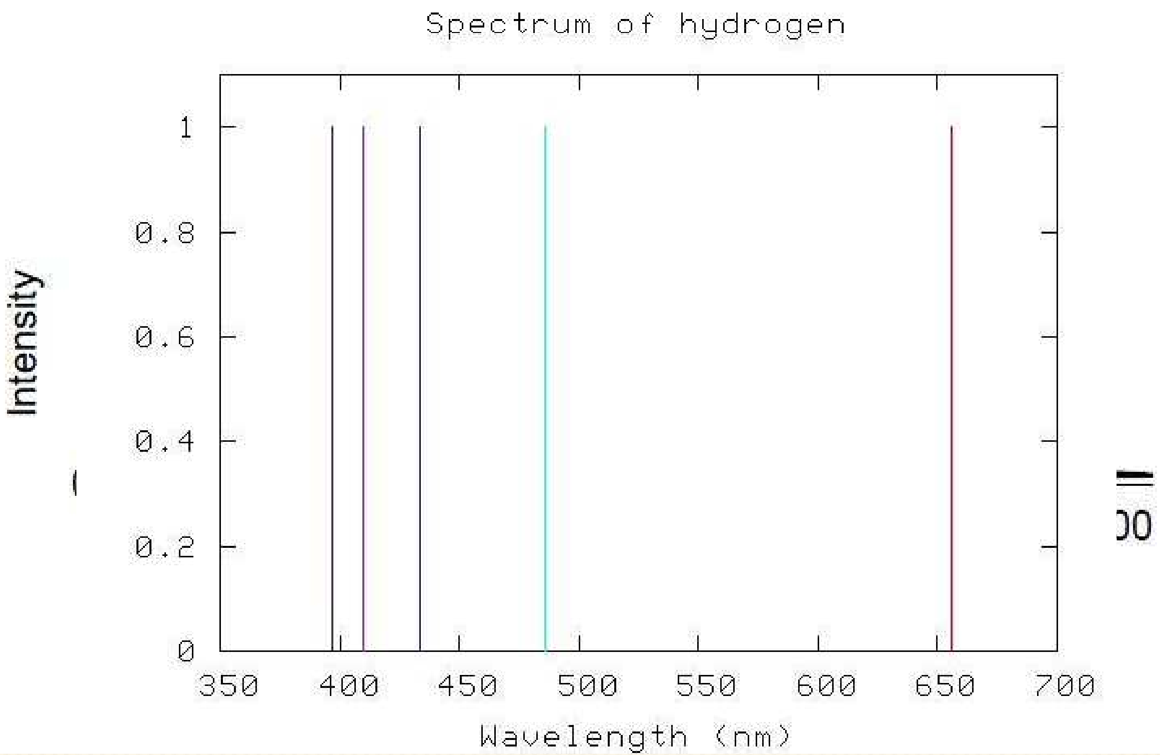
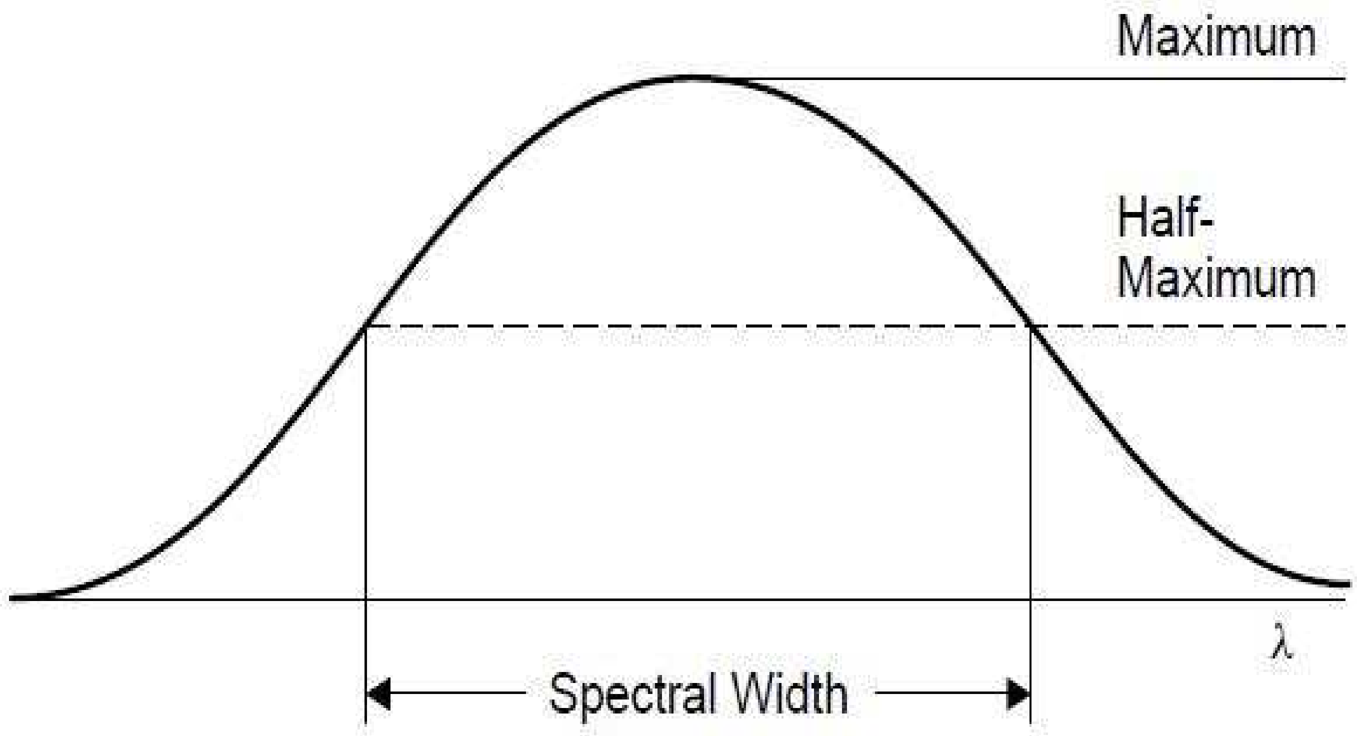
Line Spectrum

- Electromagnetic spectrum consisting of discrete lines.
- Thermal light → Line spectrum is continuous peaking at a particular wavelength.
- Gas → Line spectrum consists of discrete, well-defined components.



Spectral Width



Spectroscope

- An instrument used to measure properties of light over a specific portion of the electromagnetic spectrum.
- The variable measured is usually wavelength.

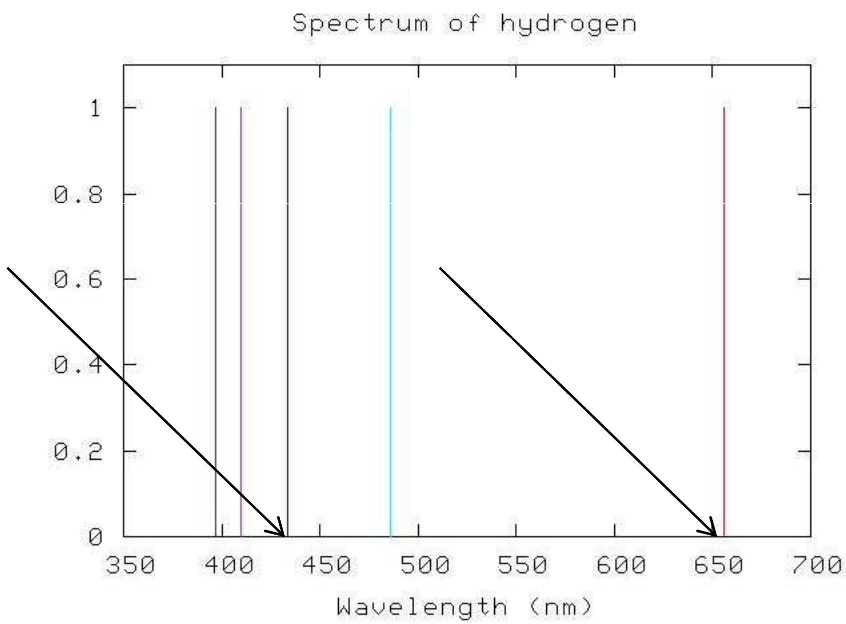
$$m\lambda = d \sin\theta$$

$m \rightarrow$ order of diffraction
$\lambda \rightarrow$ wavelength
$d \rightarrow$ spacing between lines on the grating
$\theta \rightarrow$ angle at which light is diffracted

Einstein and Planck Relation

$$E = h\nu = \frac{hc}{\lambda}$$

$\nu \rightarrow$ Frequency
$\lambda \rightarrow$ Wavelength
$c \rightarrow$ Velocity
$E \rightarrow$ Energy



Example:

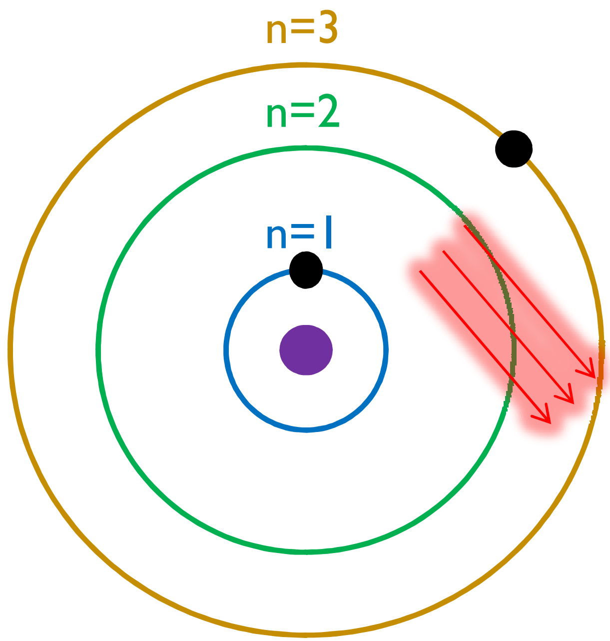
Blue $\rightarrow$ $\lambda = 434.1$ nm
$E = 2.86$ eV

Red $\rightarrow$ $\lambda = 656.3$ nm
$E = 1.9$ eV

Photoelectric Effect

Bohr’s Atomic Model

- Electrons around an atom orbit in a number of possible discrete energy states.
- Atoms do not radiate energy as long as they are fixed in that orbit.
- Atoms may jump from one energy state to another and in doing so will absorb or emit radiation in the form of a photon.



Bohr's Atomic Model (Ctd.)

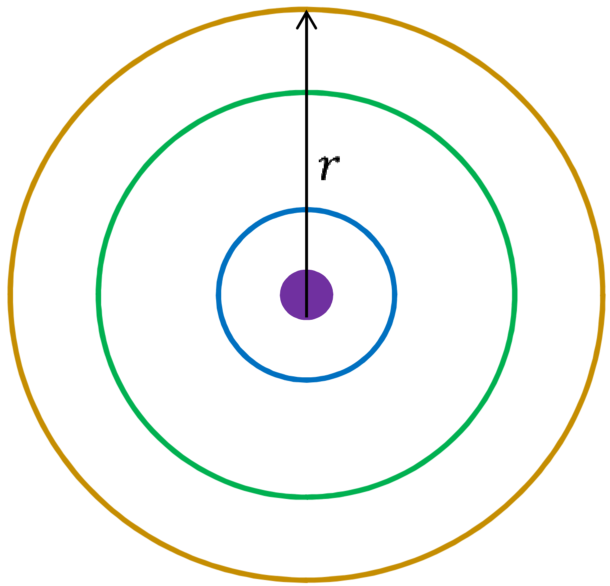
- Radius of Bohr Orbit: $r = n^2 a_0$

$a_0 \rightarrow$ Bohr radius (5.29×10^{-11} m)

Energy of an electron in a particular orbit:

$$E_n = -\frac{13.6}{n^2}\ (\text{eV})$$

$$E_1 = -13.6\ \text{eV}$$
$$E_2 = -3.4\ \text{eV}$$
$$E_3 = -1.5\ \text{eV}$$
$$E_4 = -0.85\ \text{eV}$$

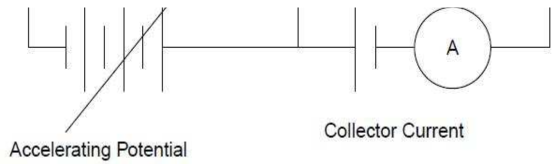


<u>Balmer Series:</u>

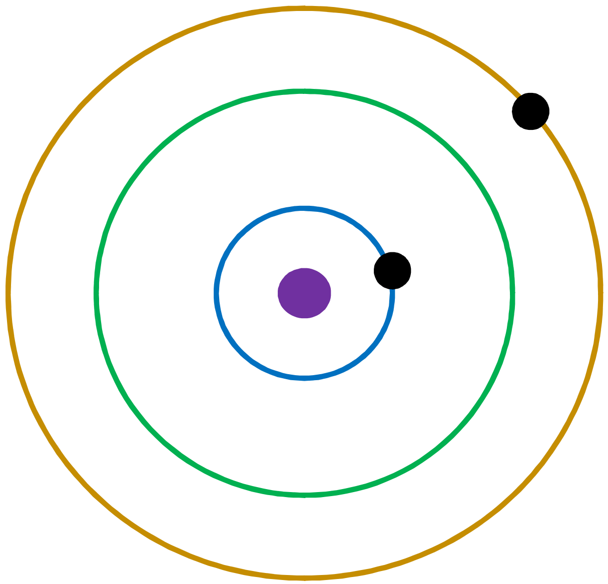
$$n_f = 2$$
$$n_i = 3 \Rightarrow E_3 - E_2 = 1.9\ \text{eV}$$
$$\lambda = 656.3\ \text{nm}$$
$$n_i = 4 \Rightarrow E_4 - E_2 = 2.55\ \text{eV}$$
$$\lambda = 486.1\ \text{nm}$$

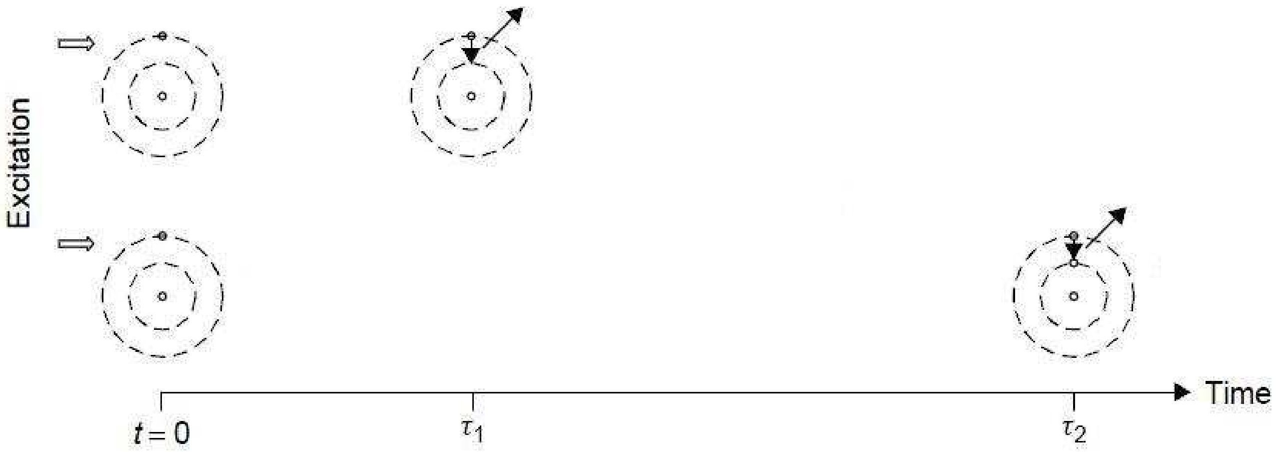
$n_f = 1 \Rightarrow$ Lyman Series
$n_f = 3 \Rightarrow$ Paschen Series
$n_f = 4 \Rightarrow$ Brackett Series
$n_f = 5 \Rightarrow$ Pfund Series

Frank-Hertz Experiment



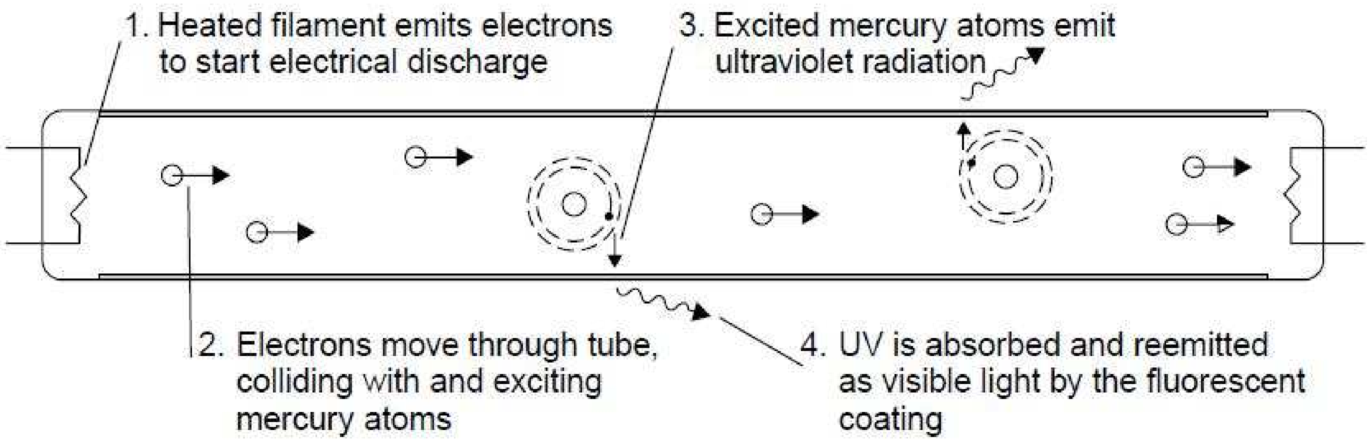
Spontaneous Lifetime

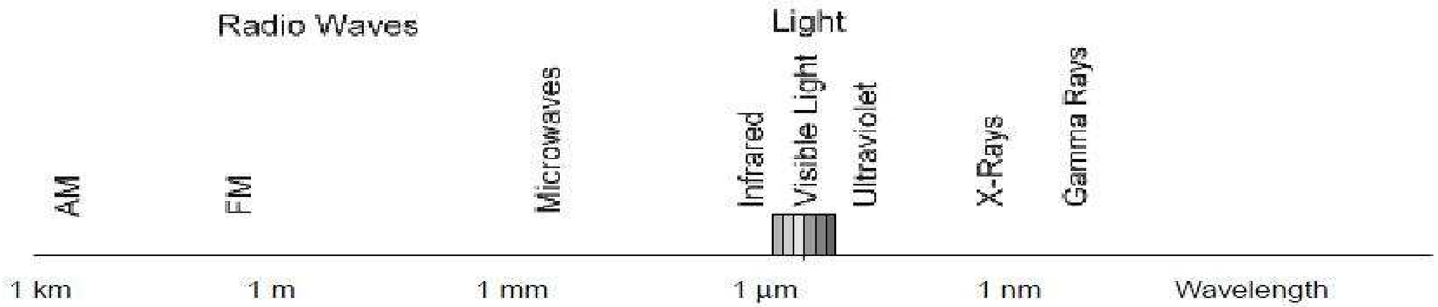




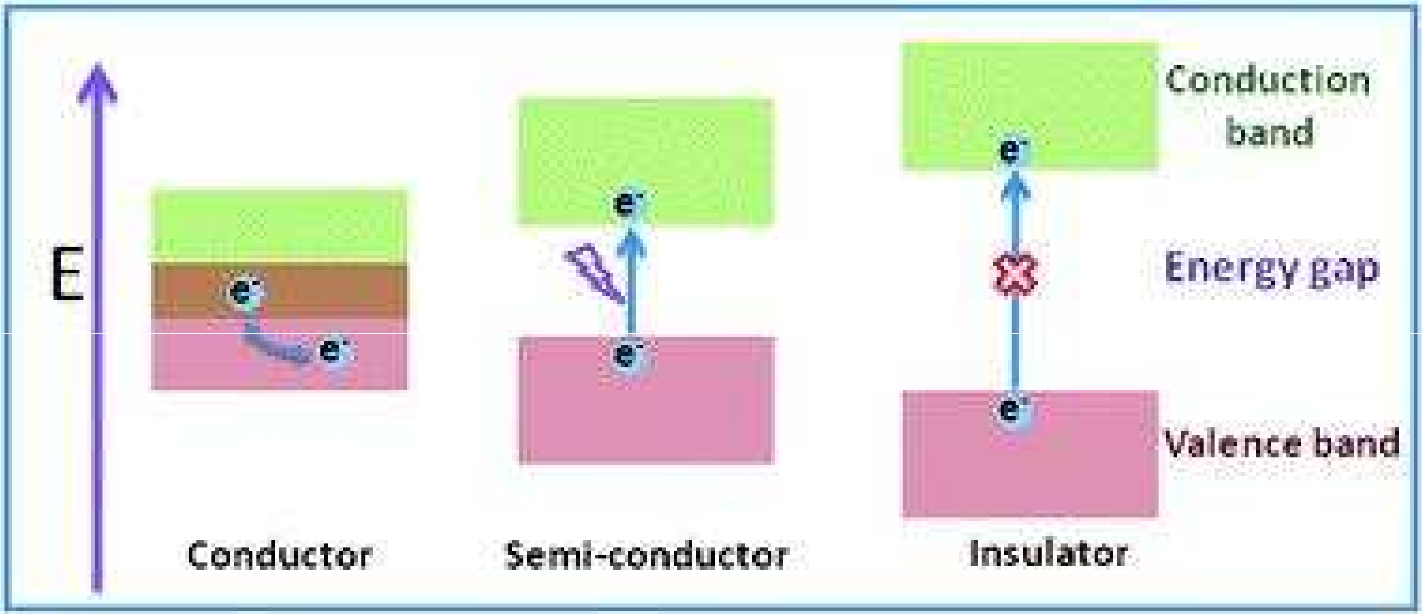
Fluorescence

- Emission of light by a substance that has absorbed light or other electromagnetic radiation.
- Fluorescent Tube: Used to convert otherwise useless ultraviolet emissions into useful visible light.



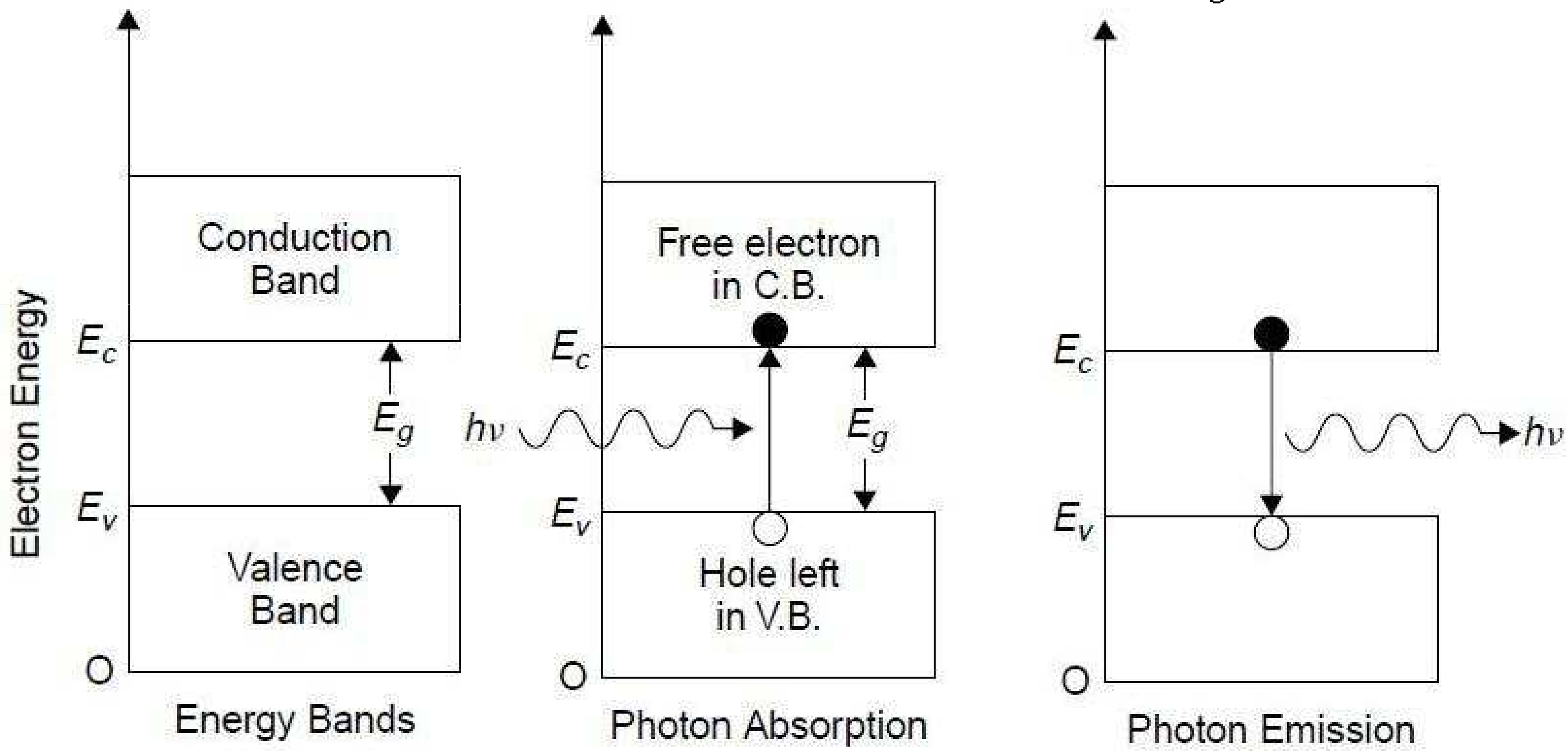


Types of Solids

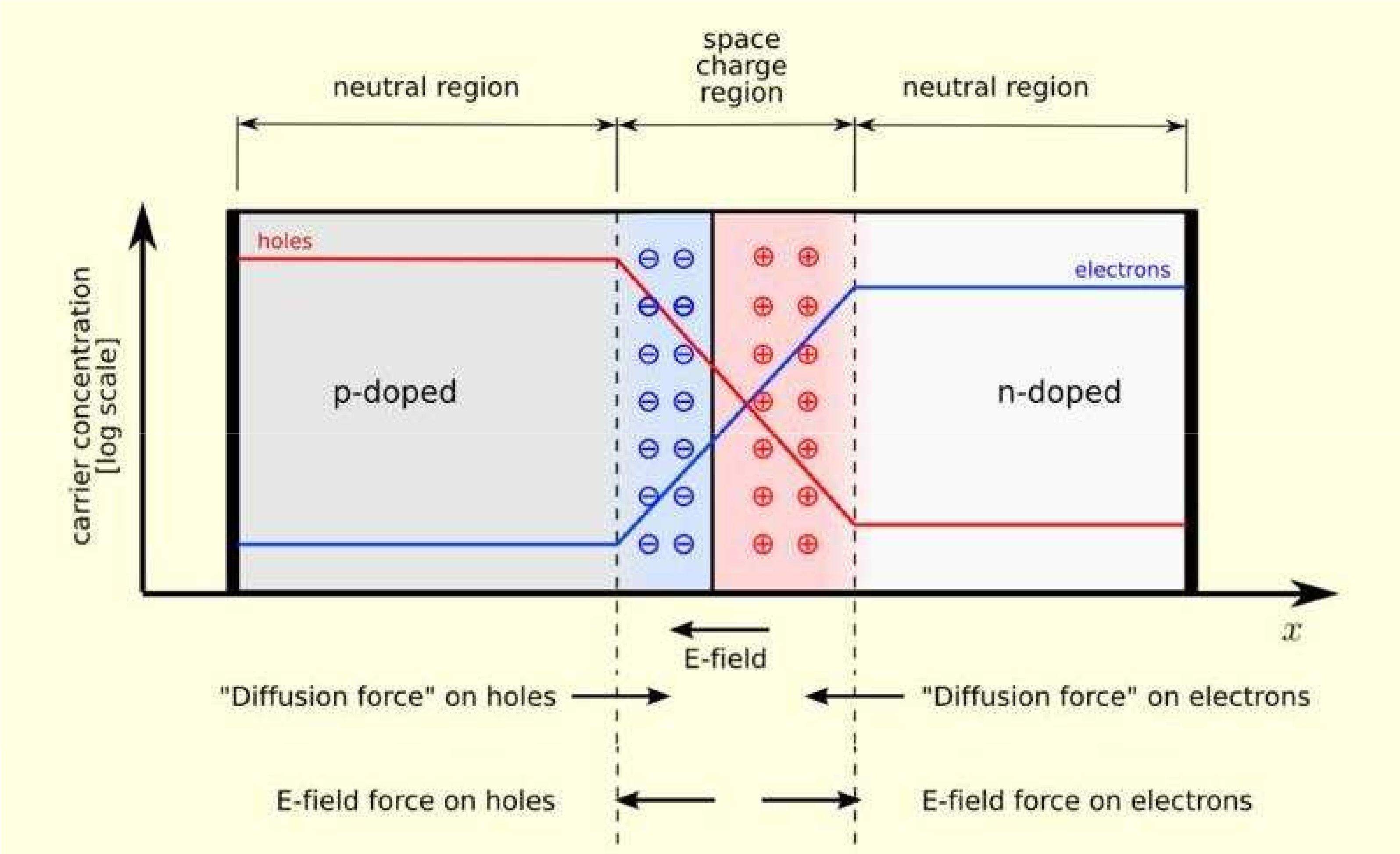


Electron-Hole Pair

Ideally, $h\nu = E_g$



p-n Junction



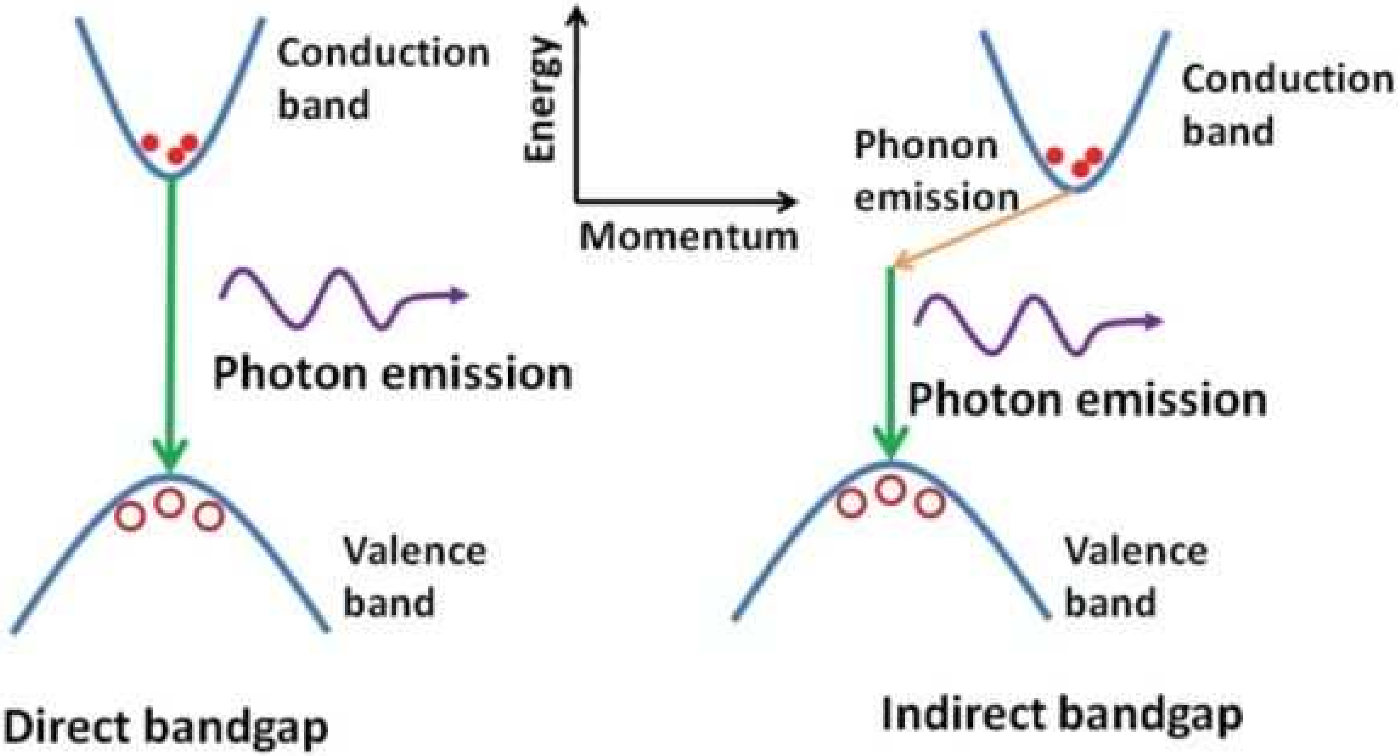
neutral region
space charge region
neutral region
holes
electrons
p-doped
n-doped
carrier concentration [log scale]
x
E-field
"Diffusion force" on holes
"Diffusion force" on electrons
E-field force on holes
E-field force on electrons

Types of Semiconductors

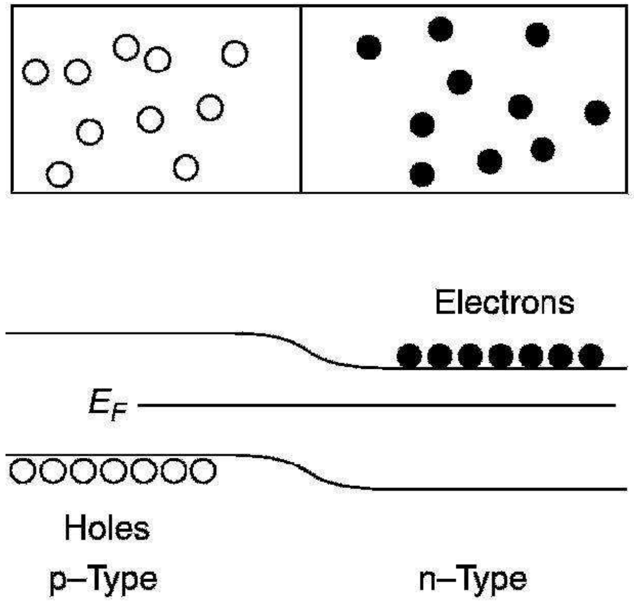
$$E = \frac{\hbar^2 k^2}{2m}$$

$$k = \frac{p}{\hbar}$$

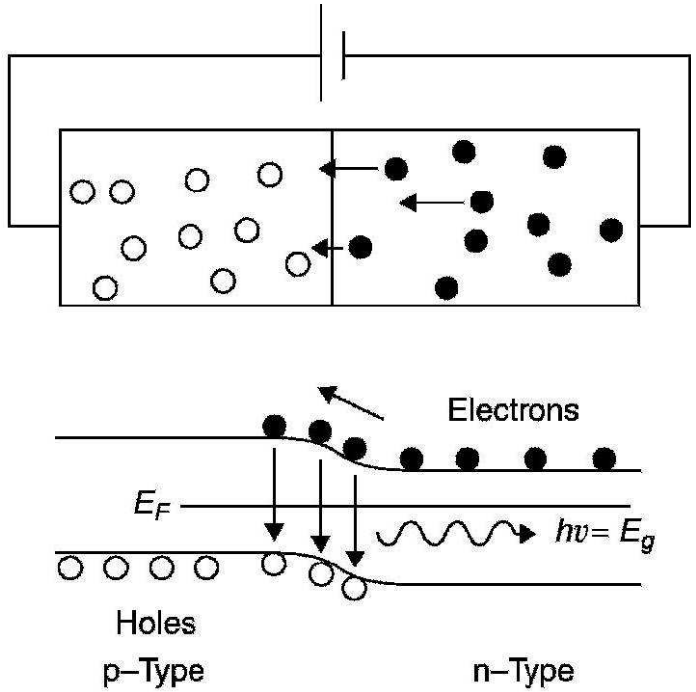
$E \rightarrow$ Energy of a particle
$k \rightarrow$ Wavenumber $(= 2\pi/\lambda)$
$m \rightarrow$ Mass of the particle
$p \rightarrow$ Momentum



Light Emitting Diode

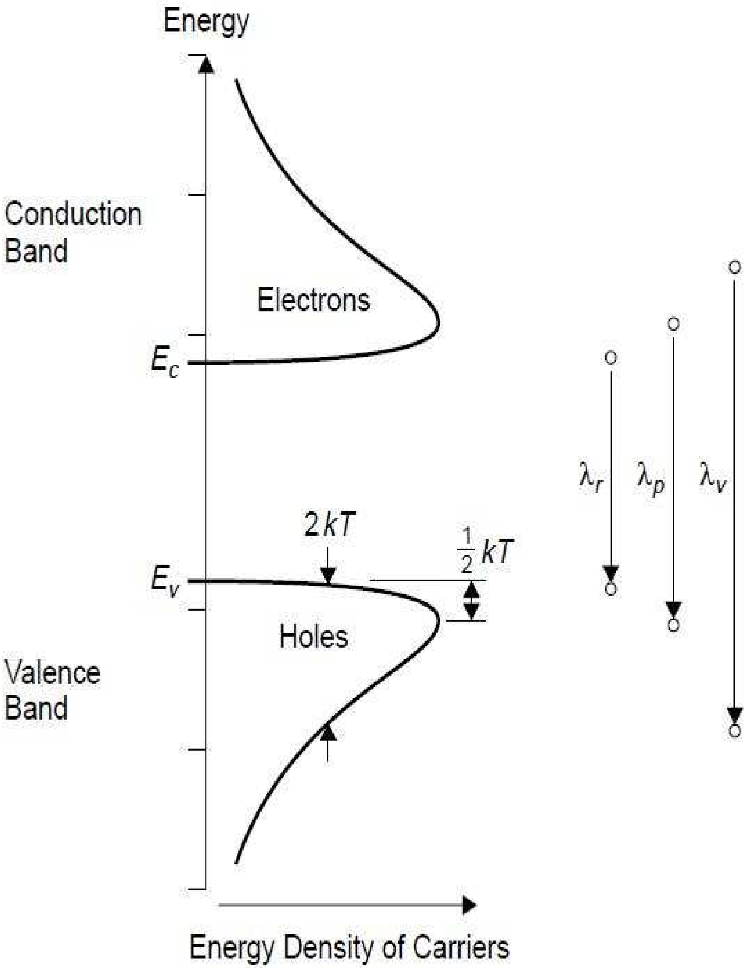


Equilibrium



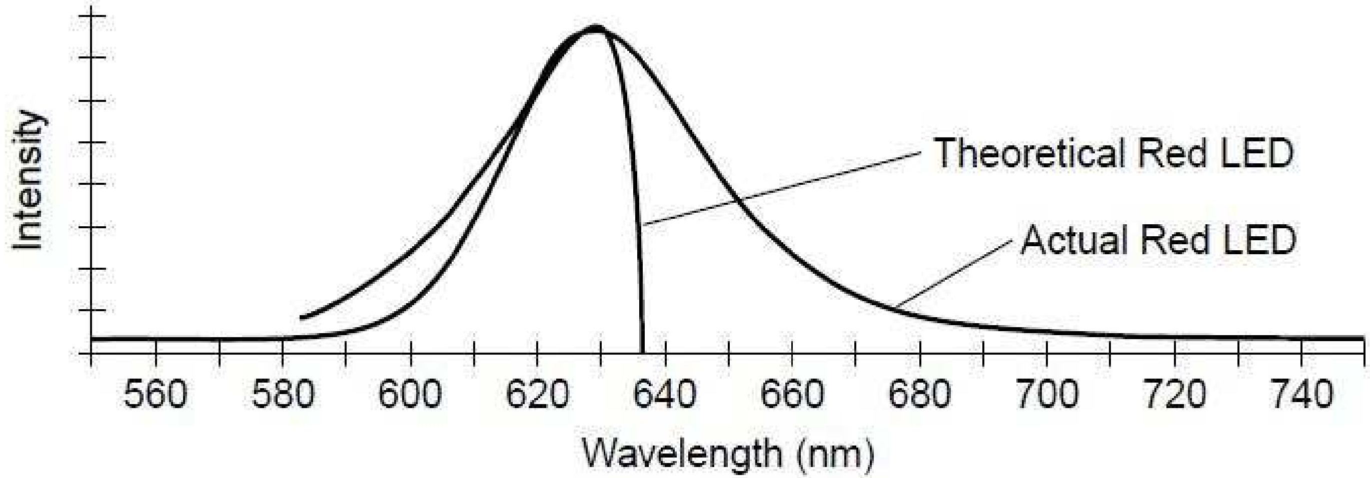
Forward Bias

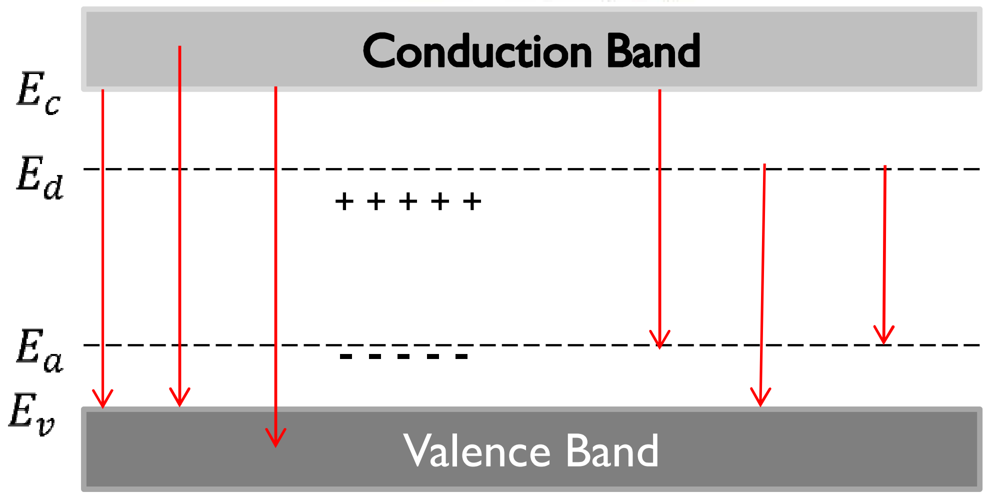
Light Emitting Diode (Ctd.)



- Emission wavelength should follow $hc/\lambda = E_g = E_c - E_v$.
- Peak occurs at $\left(E_c + \frac{1}{2}kt\right) - \left(E_v - \frac{1}{2}kt\right) = E_g + kT$.
- Peak emission wavelength, $\lambda = \frac{hc}{E_g + kT}$.
- Minimum energy difference = E_g.
- Maximum possible wavelength corresponds to E_g.

Light Emitting Diode (Ctd.)





Limitations of the Bohr Model

- Suitable for atoms with single valence electron (e. g. Hydrogen); inadequate for complex atoms such as neon (6 electrons in outer shell) or even helium (2 electrons in outer shell).
- Line spectra for such multiple electrons atoms cannot be explained.
- Angular momentum for ground state ($n = 1$) of hydrogen is $\frac{h}{2\pi}$.

  But experimentally it has been found to be zero.

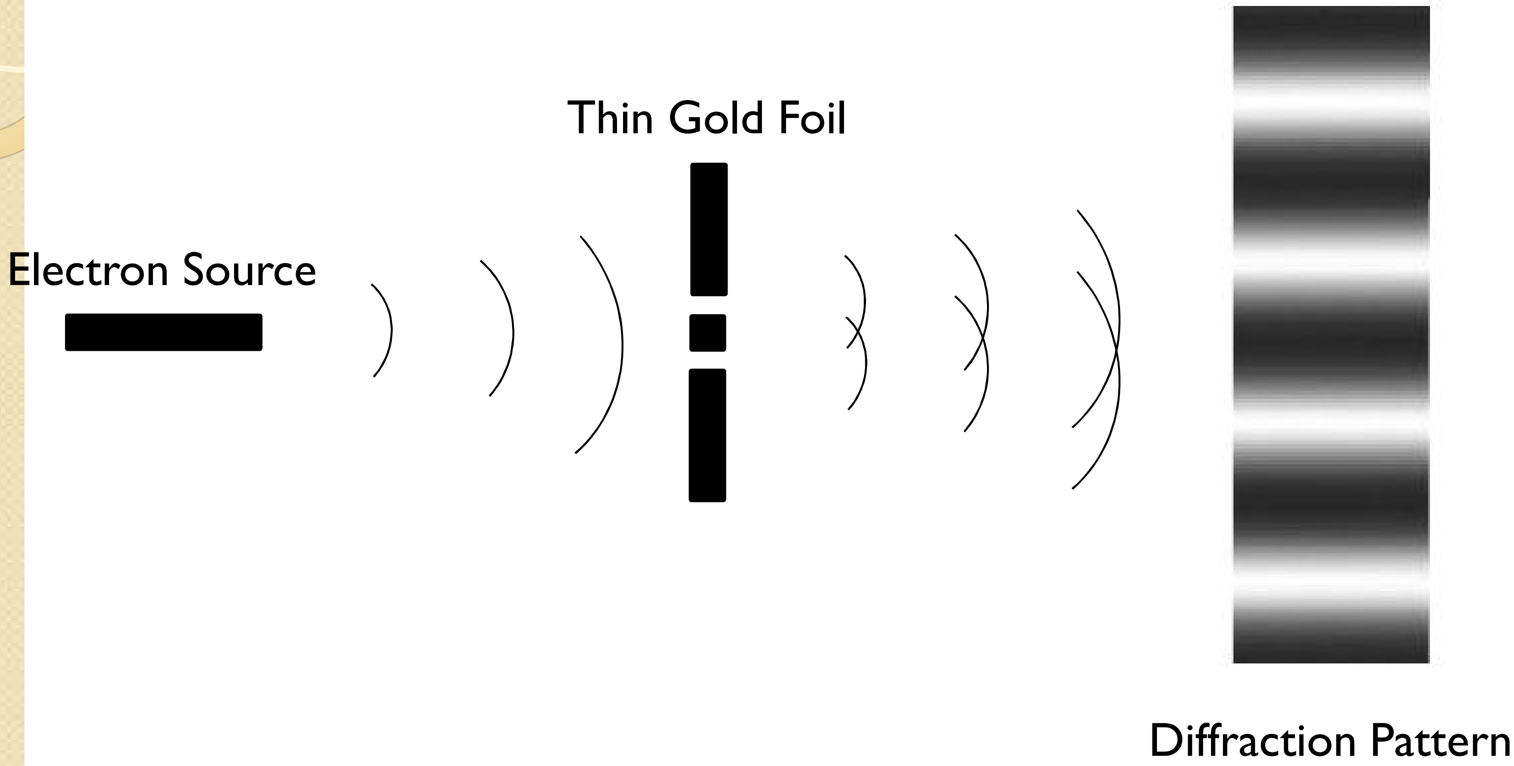
There must be something besides shell number, n.

Wave Particle Duality

- Planck-Einstein relation for photons: $E = h\nu$.
- Einstein's mass-to-energy equivalency: $E = mc^2$.
- Wavelength and frequency relation: $c = \nu\lambda$.
- Momentum, $p = mc$.
- Combining them altogether,

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

Evidence of Wave Properties in Electrons



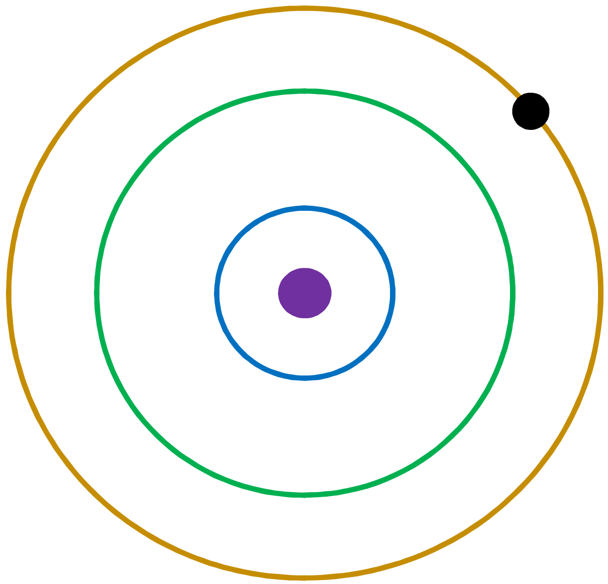
Angular Momentum

- Charged electron has two types of angular momentum:
  - Orbital
  - Spin
- Analysis of hyperfine structure of hydrogen line shows a series of very closely spaced lines in each line!
- Orbital Quantum Number:
  - Represents the magnitude of the orbital angular momentum.
  - Represented by l.
  - Values of $l = 0$ to $n - 1$.
  - Example:

$n = 2 \Rightarrow l = 0,1.$

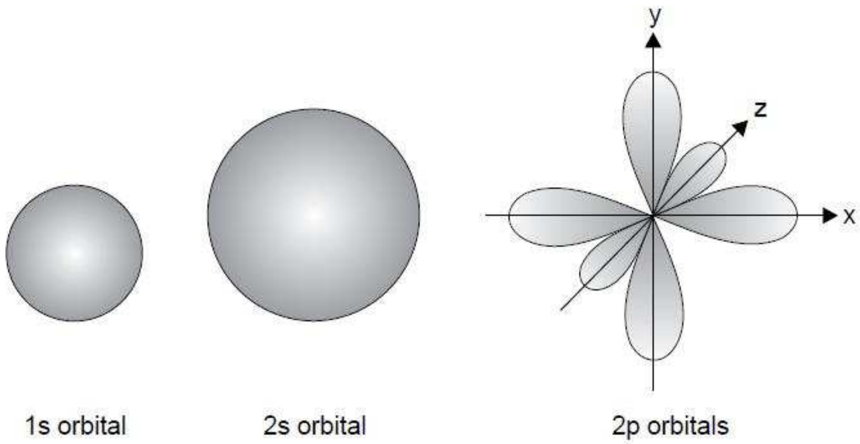
$l = 0 \Rightarrow$ Zero angular momentum; Circular orbit.

$l = 1 \Rightarrow$ Some discrete value of angular momentum; Elliptical orbit.



Angular Momentum (Ctd.)

Orbital Quantum Number	Description	Notation	Maximum Number of Electrons
$l = 0$	Sharp	s	2
$l = 1$	Principal	p	6
$l = 2$	Diffuse	d	10
$l = 3$	Fundamental	f	14

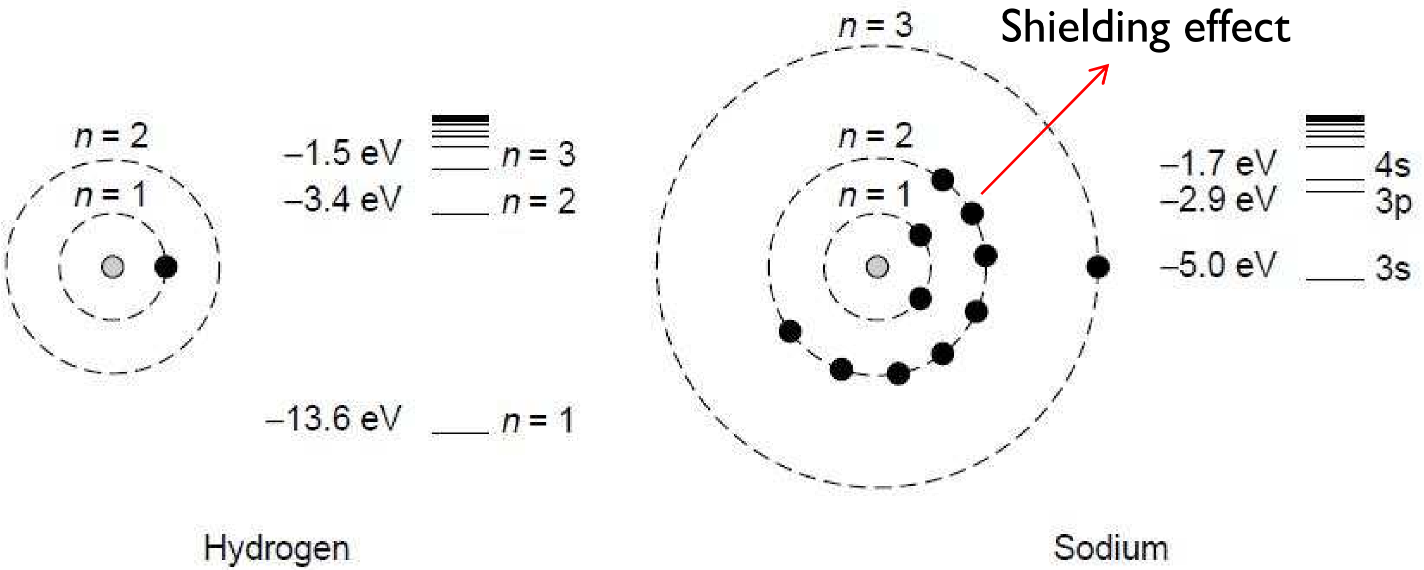


Electron Configuration Example: Na(11) = $1s^2 2s^2 2p^6 3s^1$

Transitions can occur between any energy state to yield photon emission.
Each level of n will now have n possible values of l.

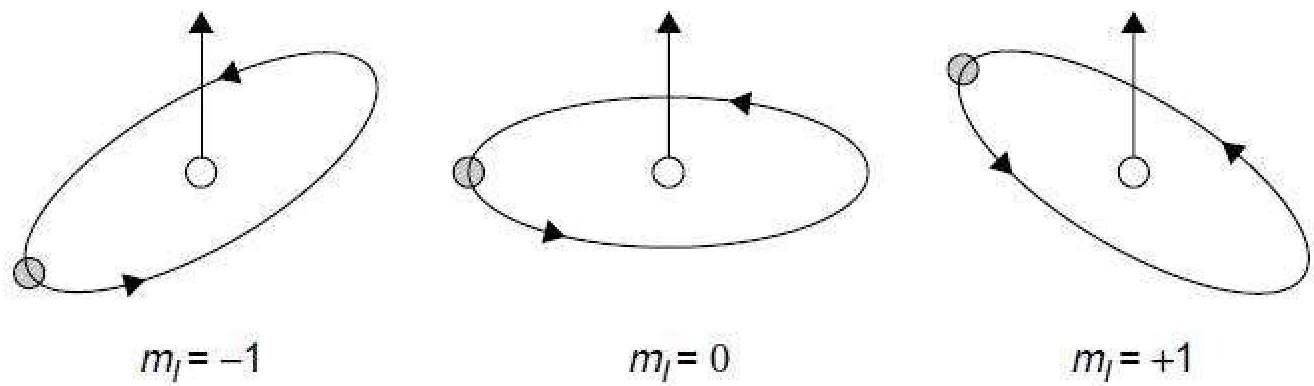
Angular Momentum (Ctd.)

- Bohr theory: all electrons in an $n = 2$ state have exactly the same energy, regardless of orbital configuration.
- But, electrons in various orbital configurations do indeed have different energies.

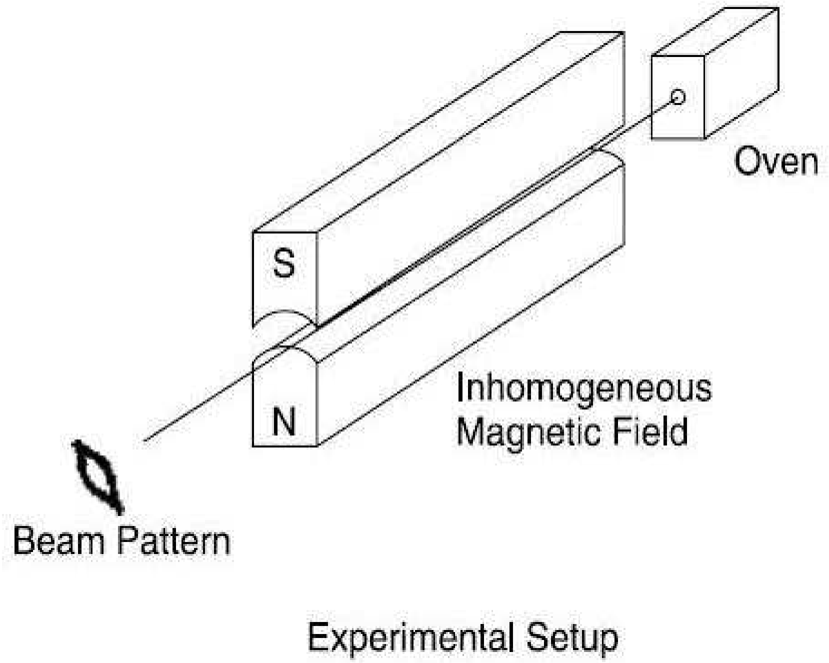


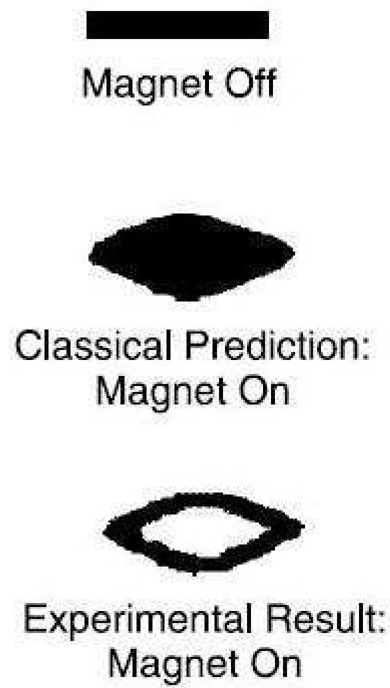
Magnetic Quantum Number

- An electron: analogous to a current loop and will exhibit a magnetic moment.
- Denoted by: m.
- Represents the direction of angular momentum of an electron.
- Values of $m =$ all integers from $-l$ to $+l$. (Example: $l = 1 \Rightarrow m = -1,0,1$)
- Can be thought of as the three-dimensional tilt of an elliptical orbit.



The Stern-Gerlach Experiment

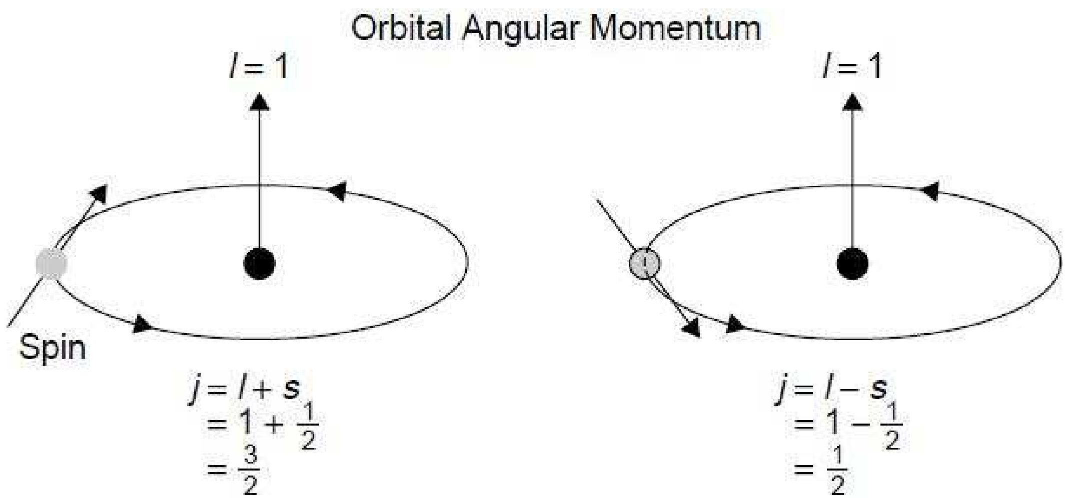




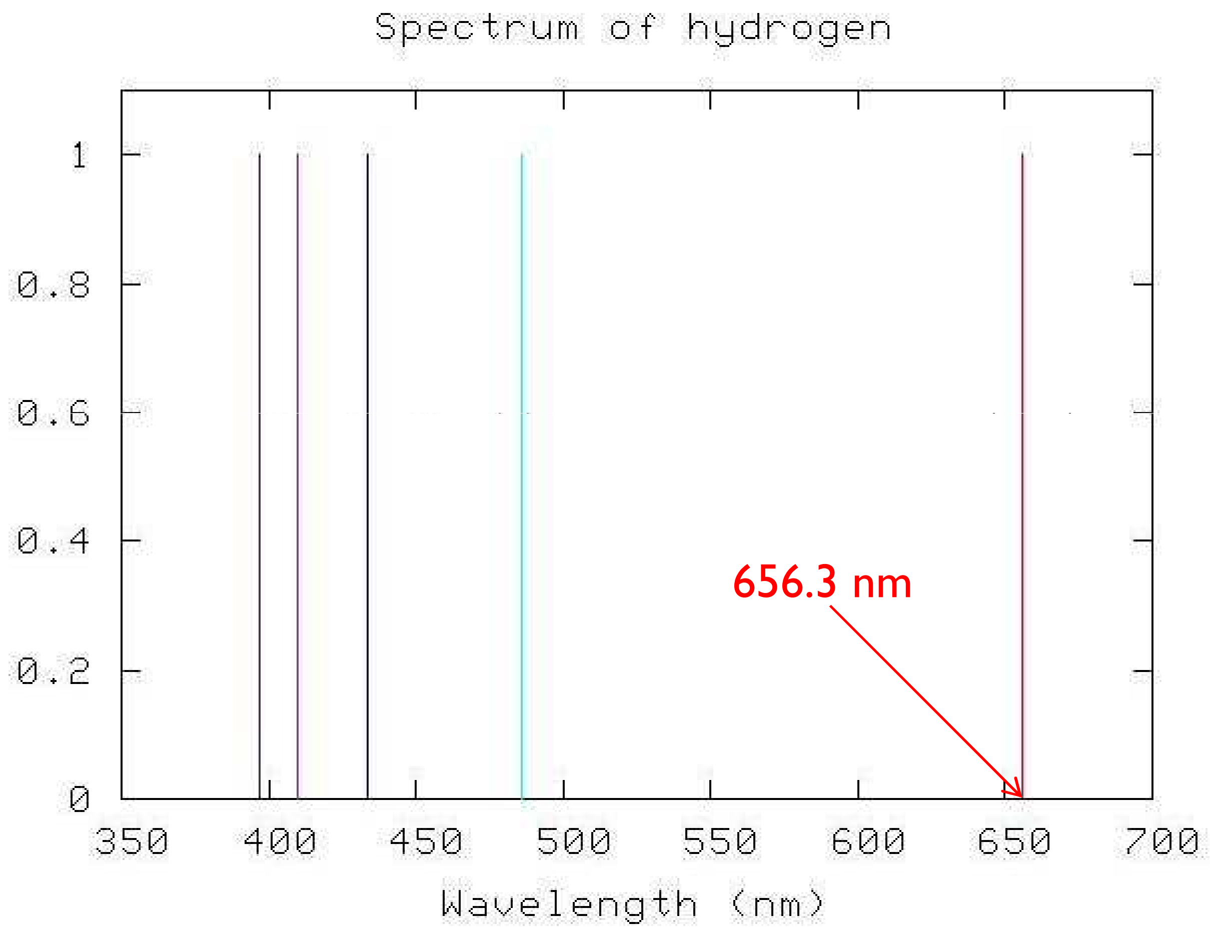
- Neutral silver atoms emerging from a hot oven.
- Deflection via a magnetic field.
- Target: a photographic plate.

Spin Quantum Number

- Spin: rotation of a particle around some axis.
- Denoted by s.
- Two possible values: $-\frac{1}{2}$ and $+\frac{1}{2}$.
- When s is added and subtracted from l, the effects of spin on energy levels can be seen. --- $l - s$ coupling

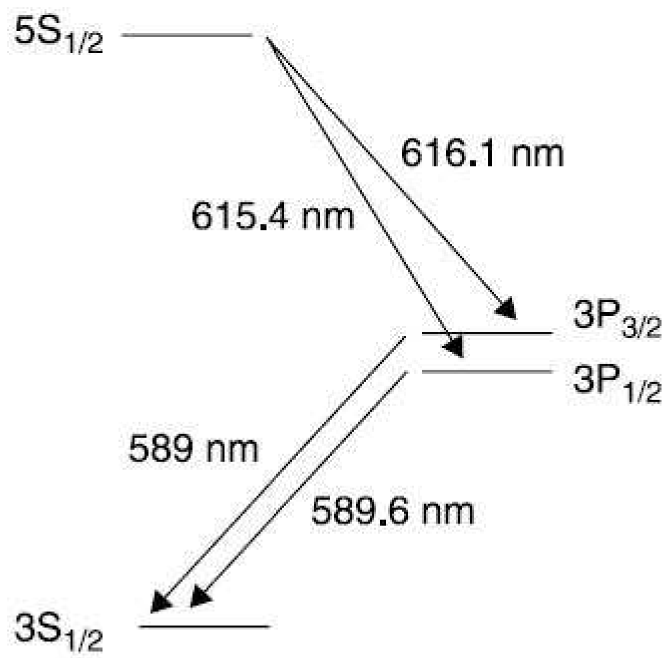


Effect of Spin



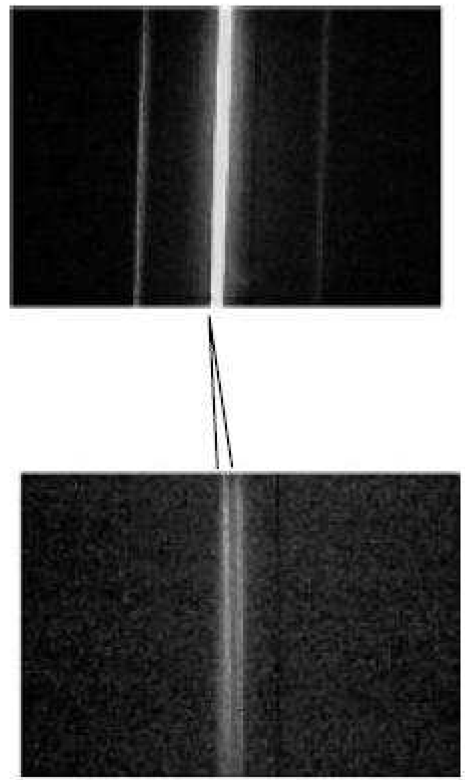
The Sodium Spectrum

Na(11) = $1s^2 2s^2 2p^6 3s^1$



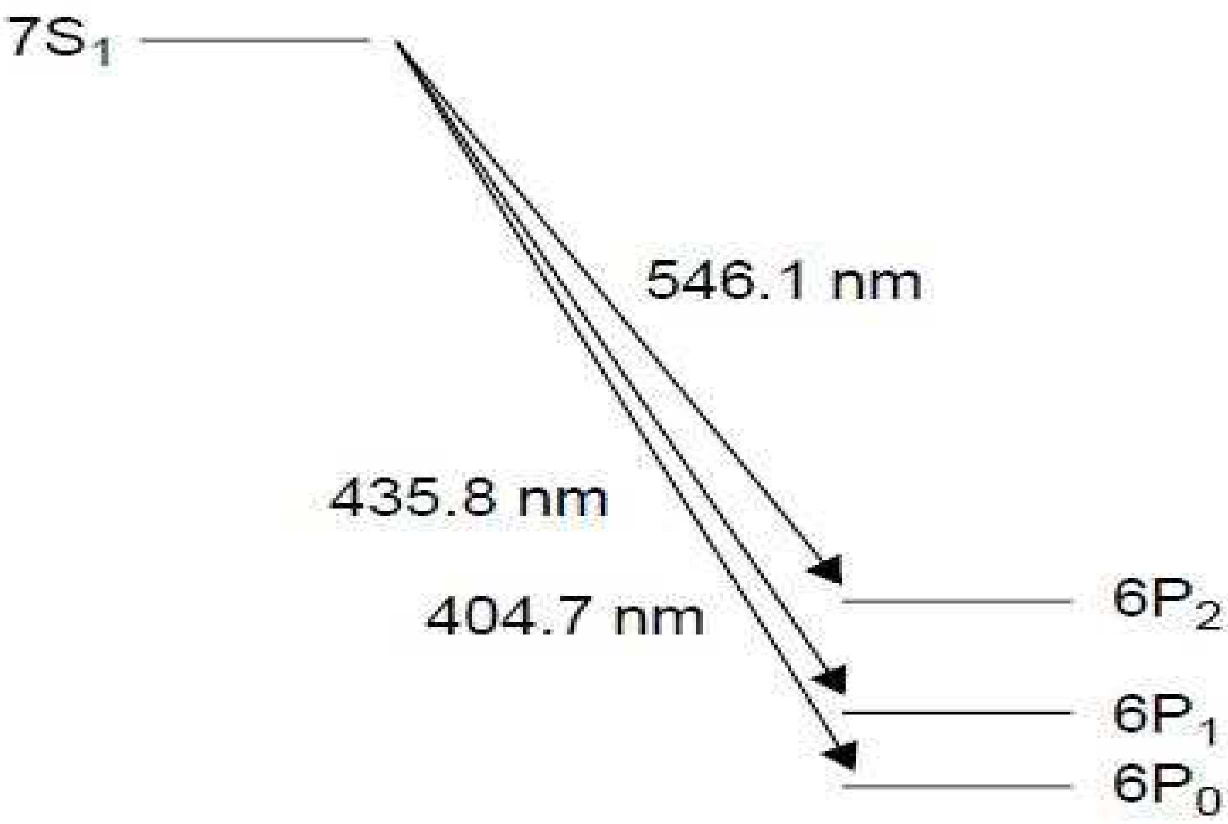
Energy Levels in Sodium

Sodium Spectrum



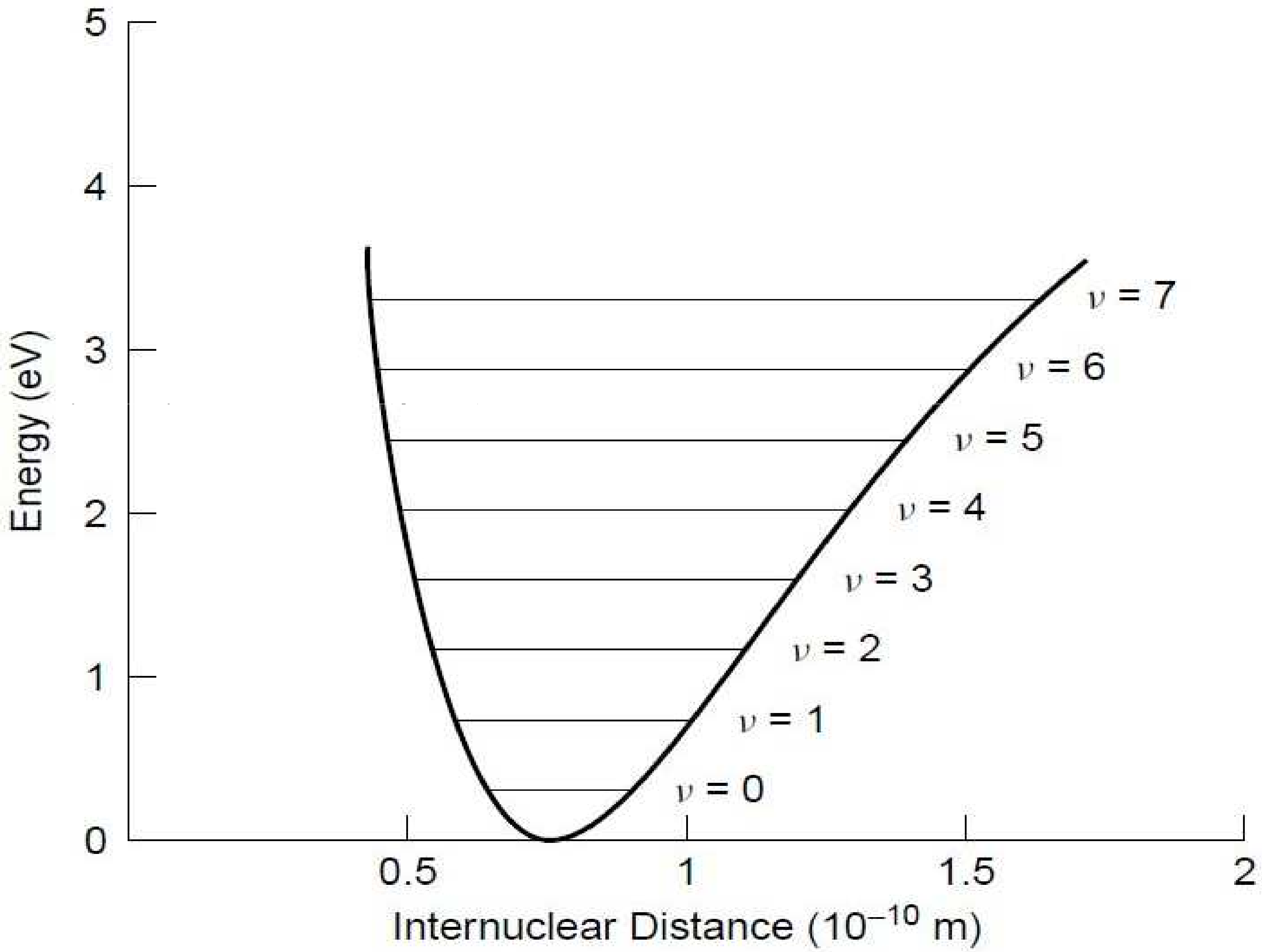
D Lines

The Mercury Spectrum



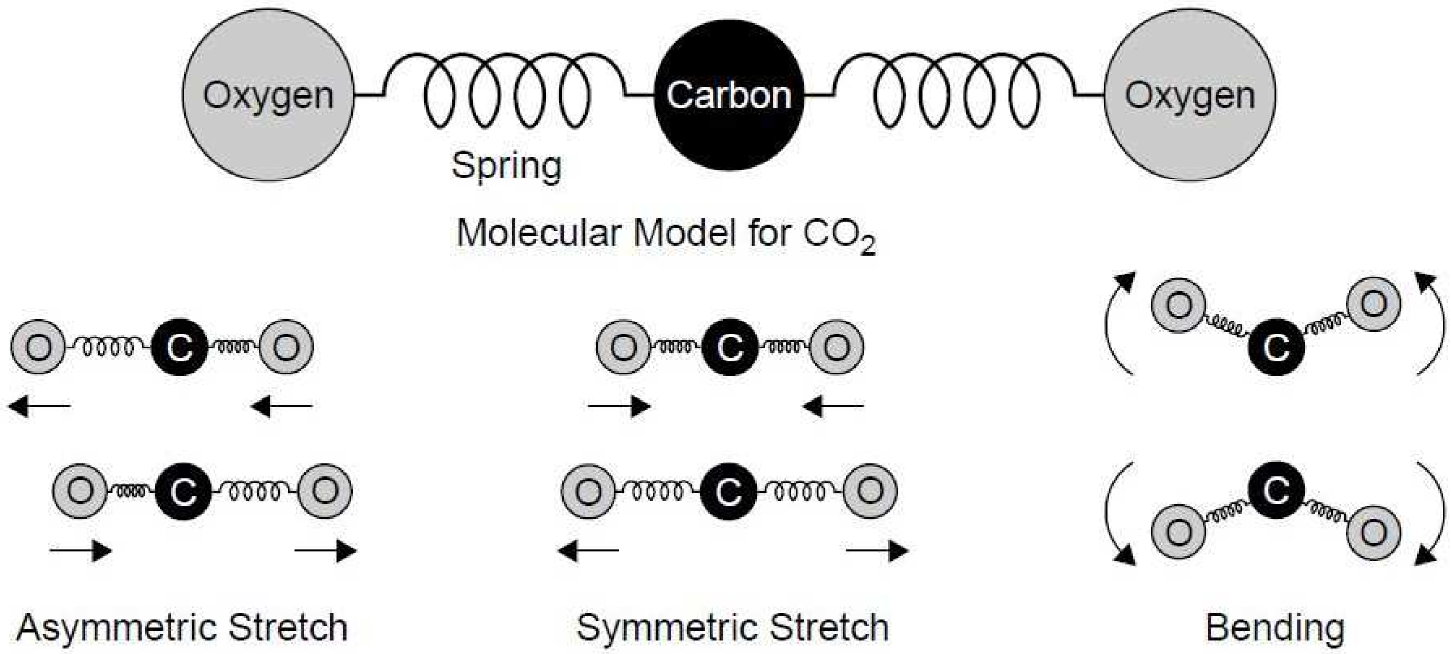
Energy Levels in Mercury

Hydrogen Molecule



Carbon Dioxide Molecule

- A single carbon atom chemically bonded to two oxygen atoms with double covalent bonds acting as springs leading to vibrational modes.



Thank You...