• Atomic Structure
Nucleus

– Nucleus has protons and neutrons.
– Protons are + charges
– Neutrons have no charge

– 2000 times more massive > electron
• Why is an atom electrically neutral?
Rutherford’s alpha particle scattering Scattering.

- What is alpha particle?
  - Bare nucleus of He
  - Has 2 protons and 2 neutrons
- Measurement of scattering angle
Scattering of Alpha Particles
• The Quantum Concept
• Light Photons $\rightarrow$ Quanta
  • $E = hf$
    – where $E =$ energy
    – $h =$ Plank’s constant $= 6.63 \times 10^{-34}$ Js
    – $f =$ frequency
Light from solids, liquids, dense gases

Continuum Spectrum of Colors

Light from gas

Line Spectra
Line Spectrum of Hydrogen

\[ \frac{1}{\lambda} = R \left( \frac{1}{2^2} - \frac{1}{n^2} \right) \]

Rydberg’s Const

\[ R = 1.097 \times 10^7 \frac{1}{m} \]

Violet \( (n=6) \)

Violet \( (n=5) \)

Blue-green \( (n=4) \)

Red \( (n=3) \)

Ultraviolet series

Visible (Balmer series)

Infrared series
• **Bohr’s Theory**
  
  – **Allowed Orbitals**
    • An electron can only orbit around an atom in specific orbits
  
  – **Radiation less Orbits**
    • An electron in an allowed orbit does not emit radiant energy as long as it remains in the orbit.
  
  – **Quantum Leaps**
    • An electron gains or loses energy only by moving from one allowed orbit to another.
    • The lowest energy state $\rightarrow$ ground state
    • Higher states $\rightarrow$ excited states
Energy of Orbits

\[ E_n = \frac{E_L}{n^2} \]

\( E_L \) = Lower energy level of any two electron orbits
\( E_n \) = Higher energy level of any two electron orbits
\( n \) = number of the higher electron orbit

Energy of each orbit is quantized
### Energy Level Diagram for Hydrogen Atom

<table>
<thead>
<tr>
<th>$n$</th>
<th>Energy (J)</th>
<th>Energy (eV)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>$-2.18 \times 10^{-18}$</td>
<td>$-13.6$</td>
</tr>
<tr>
<td>2</td>
<td>$-5.44 \times 10^{-19}$</td>
<td>$-3.40$</td>
</tr>
<tr>
<td>3</td>
<td>$-2.42 \times 10^{-19}$</td>
<td>$-1.51$</td>
</tr>
<tr>
<td>4</td>
<td>$-1.36 \times 10^{-19}$</td>
<td>$-0.850$</td>
</tr>
<tr>
<td>5</td>
<td>$-8.70 \times 10^{-20}$</td>
<td>$-0.544$</td>
</tr>
<tr>
<td>6</td>
<td>$-6.05 \times 10^{-20}$</td>
<td>$-0.377$</td>
</tr>
</tbody>
</table>

$1\text{eV} = 1.6 \times 10^{-19} \text{J}$

- Violet ($7.3 \times 10^{14} \text{Hz}$)
- Violet ($6.9 \times 10^{14} \text{Hz}$)
- Blue-green ($6.2 \times 10^{14} \text{Hz}$)
- Red ($4.6 \times 10^{14} \text{Hz}$)
Quantum Jump

\[ E_H - E_L = hf \]

- \( E_H \) = frequency
- \( E_L \) = Planck's constant

fff
Fluorescent Bulbs
Calculate the energy for levels $n=2$, $n=4$
Calculate the frequency of the photon emitted when an electron drops from $n = 3$ to the ground state. ($1 \text{ eV} = 1.6 \times 10^{-19} \text{ J}$)

$n = 4$

$n = 3$ \hspace{2cm} $E_{n=3} = -1.5 \text{ eV}$

$n = 2$

$n = 1$ \hspace{2cm} $E_{n=1} = -13.6 \text{ eV}$
a)

\[
E_n = \frac{E_L}{n^2}
\]

\[
E_2 = \frac{E_L}{2^2} = -\frac{13.6eV}{4} = -3.4eV
\]

\[
E_2 = -3.4eV \times \left( \frac{1.6 \times 10^{-19} J}{1eV} \right) = -5.44 \times 10^{-19} J
\]

\[
E_4 = \frac{E_L}{4^2} = -\frac{13.6eV}{16} = -0.85eV
\]

\[
E_4 = -0.85eV \times \left( \frac{1.6 \times 10^{-19} J}{1eV} \right) = -1.36 \times 10^{-19} J
\]
b) \[ E_H - E_L = hf \]
\[ E_3 - E_1 = -1.5eV - (-13.6eV) \]
\[ E_3 - E_1 = -1.5eV + 13.6eV \]
\[ E_3 - E_1 = 12.1eV = 1.94 \times 10^{-18} J \]
\[ E_3 - E_1 = hf \]
\[ 1.94 \times 10^{-19} J = \left(6.63 \times 10^{-34} J.s\right) \times f \]
\[
\therefore f = \frac{1.94 \times 10^{-18} J}{6.63 \times 10^{-34} J.s} = 2.92 \times 10^{15} \text{ Hz} \]