**Atomic Spectra and Energy Levels**

- Excited atoms emit light (neon signs, etc.)
- Emission from different elements is different colors.
- Emission of only certain wavelengths
- “Spectral lines”
- Existence of spectral lines implies “quantized energy levels.”

---

**Atomic Spectra: Emission**

- Excited Atom $\rightarrow$ De-excited Atom + Photon

\[ \Delta E_{\text{atom}} = E_{\text{photon}} = h\nu \]

---

**Atomic Spectra: Absorption**

- Atom + Photon $\rightarrow$ Excited Atom

\[ \Delta E_{\text{atom}} = E_{\text{photon}} = h\nu \]

---

**Atomic Spectra**

- Spectra reveal quantized energies.
- Conservation of Energy
  - Relates atom’s energy levels to photon’s wavelength
- Absorption & Emission
  - Same lines (from same element)
  - Usually see more lines in emission
A hypothetical atom has only 4 allowed energy levels.

The emission spectrum of this atom shows 6 lines at wavelengths of 100 nm, 114 nm, 150 nm, 300 nm, 480 nm, and 800 nm.

Atomic energy levels usually get closer together as energy increases.

Problem - Energy Levels

Draw an energy level diagram for this hypothetical atom.

Label the 4 states as $E_1$, $E_2$, $E_3$, and $E_4$, with $E_1 < E_2 < E_3 < E_4$.

Use arrows to show the 6 observed transitions, and the correct wavelengths.

Try to draw your diagram “to scale,” showing the spacings between levels.

4 Energy Levels

6 Transitions

Problem - Energy Levels

Convert to frequencies using: $v = c/\lambda$.

- $\lambda = 100$ nm $\Rightarrow v = 3 \times 10^{15}$ m s$^{-1}$
- $\lambda = 114$ nm $\Rightarrow v = 2.63 \times 10^{15}$ m s$^{-1}$
- $\lambda = 150$ nm $\Rightarrow v = 2 \times 10^{15}$ m s$^{-1}$
- $\lambda = 300$ nm $\Rightarrow v = 1 \times 10^{15}$ m s$^{-1}$
- $\lambda = 480$ nm $\Rightarrow v = 6.25 \times 10^{14}$ m s$^{-1}$
- $\lambda = 800$ nm $\Rightarrow v = 3.75 \times 10^{14}$ m s$^{-1}$

Remember, $E \propto v$ ($E = h v$)
Absorption and emission spectra exist for both atoms and molecules. Observed wavelengths are characteristic of a particular substance. Spectra are often used to identify unknown substances, in areas ranging from forensics to astrophysics.

Electron diffraction

Electron diffraction from aluminum foil. Pattern is similar to pattern obtained in X-ray (light) diffraction. Electrons have some wave-like properties.

Uncertainty principle

A curious result of wave mechanics. "Position and energy (or momentum) can not be specified simultaneously." Mathematically (one dimension):

\[(\Delta x)(\Delta p) > \hbar/4\pi\]

Electrons as delocalized waves rather than particles at a specific position.

Properties of Electrons

Like light, electrons can show properties of both waves and particles. Electrons bound in atoms can only have certain quantized energies. Electrons in atoms can best be described as "delocalized waves."

The position of an electron can not be specified. The more precisely its position is specified, the less we can know about its momentum. Electrons have magnetic properties. "Spin" 2 possible spin values; the image of a spinning particle is not literally true.
DeBroglie’s inspired guess

For light:
Momentum, \( p = (\text{Energy}/\text{velocity}) = E/c \)
using Planck’s formula, \( E = h\nu = (hc/\lambda) \)
or, combining equations: \( p = (h/\lambda) \)
DeBroglie assumed the same relationship for
free electrons and used the particle-like
definition of momentum, \( p = mu: \)
\[ p = mu = h/\lambda \quad \text{or} \quad \lambda = (h/mu) \]

Photons vs. Electrons

<table>
<thead>
<tr>
<th>Photons</th>
<th>Electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>Energy ( (E) )</td>
<td>( E = (1/2)mu^2 )</td>
</tr>
<tr>
<td>Wavelength ( (\lambda) )</td>
<td>( \lambda = (h/p) = (h/mu) )</td>
</tr>
<tr>
<td>Velocity ( c = 3 \times 10^8 \text{ m s}^{-1} )</td>
<td>( u = \sqrt{(2E/m)} = (h/m\lambda) )</td>
</tr>
<tr>
<td>( h = 6.626 \times 10^{-34} \text{ J s} )</td>
<td></td>
</tr>
</tbody>
</table>

Problems

- A common wavelength used in X-ray
  diffraction is \( \lambda_{Cu} = 1.54 \text{ Å} = .154 \text{ nm} \) (This
  is comparable to the lengths of chemical
  bonds.) Q: What is the energy of one mole
  of such photons?
  Answer: \( 7.77 \times 10^5 \text{ kJ/mol} \)
- What is the kinetic energy of one mole
  of electrons with the same wavelength?
  Answer: \( 6.12 \times 10^3 \text{ kJ/mol} \)

Electron Diffraction

Electron diffraction from aluminum foil.
Patterns similar to those in X-ray (light) diffraction.
⇒ Electrons have some wave-like properties.

Electrons

- Electrons show properties of waves AND
  particles. (Like light does!)
- Electrons in atoms best described as
  “delocalized waves”

Problems

- A common wavelength used in X-ray
  diffraction is \( \lambda_{Cu} = 1.54 \text{ Å} = .154 \text{ nm} \) (This
  is comparable to the lengths of chemical
  bonds.) Q: What is the energy of one mole
  of such photons?
- What is the kinetic energy of one mole
  of electrons with the same wavelength?
Functions for Electron in a Box

We can write a general equation for the allowed standing waves:

\[ \Psi_n = \sin \frac{n\pi x}{L} \]

Where \( n = 1, 2, 3 \ldots \)

\( n \) is called a quantum number

Standing Waves - Wavefunctions

(Electrons Confined to small spaces!)

<table>
<thead>
<tr>
<th>Amplitude</th>
<th>( n = 1 )</th>
<th>( n = 2 )</th>
<th>( n = 3 )</th>
</tr>
</thead>
</table>

Here \( n \) is a quantum number, identifying the allowed states.

Bohr “Orbits” and de Broglie waves (Discarded)

There are some animated standing waves at:
http://www.sengpielaudio.com/StandingWaves.htm