Lecture 19
Chapter 7  Sections 3 and 4

• Photoelectric effect
• Atomic energy levels
• Transitions – spectroscopy
• de Broglie
• Heisenberg
Announcements

• Two good-looking seminars Friday
  – Chemistry (you have info)
  – Bio – bioinformatics  SC1019 3:00

• Exam 2 isn’t too far off
  – Do lots of book problems
Light also is particle-like

- Light comes in bundles with specific amounts of energy, determined by their frequency (wavelength)
  - The bundles are called **photons**

\[ E_{\text{photon}} = h \nu_{\text{photon}} \]

\( h \) is Planck’s constant = 6.62606876 \( \times 10^{-34} \) J s

- These bundles have momentum, just like particles
- And they can interact with matter
- One example is the “Photoelectric Effect”
  - HUGE in early quantum mechanics
Your home microwave still produces E-M radiation with a frequency of $2.45 \times 10^9$ Hz. What is the energy in one photon of this radiation?

$$E_{\text{photon}} = h \nu_{\text{photon}}$$

| 25% | 1.  $8.08 \times 10^{-35}$ J |
| 25% | 2.  $1.62 \times 10^{-24}$ J |
| 25% | 3.  $1.84 \times 10^{32}$ J |
| 25% | 4.  $3.69 \times 10^{42}$ J |
The Photoelectric Effect

• Several phenomena can be observed:
  – Below a certain frequency, no electrons were observed, no matter what the intensity.
  – The energy of the ejected electrons increased linearly with the frequency of light \( E \propto \nu \)
  – The number of emitted electrons increased with light intensity.
  – All metals show the same pattern, but each metal has a different threshold frequency.
Behavior Explained by Some Guy Named Einstein
Nobel Prize for this work

- Light has two properties
  - Wave-like
  - Particle-like

- Intensity of light is the number of photons
  - Many photons \(\rightarrow\) high intensity

- Energy of a photon determined by the frequency
  - High frequency \(\rightarrow\) high energy

- The photon must have \(E\) above some threshold to eject an electron (the \(E_{\text{threshold}}\) is different for each metal)

- Each photon can cause zero or one electron to be ejected
  - Even if \(E\) is very high, still just one electron
  - But that electron will have huge kinetic energy
Binding and Kinetic Energy of Electron

- When a metal surface absorbs a photon, the energy is transferred to an electron.
- Some of the energy must be used to overcome the forces that bind the electron to the metal.
- The remainder is used to eject the electron and becomes kinetic energy.

Electron kinetic energy = Photon energy – Binding energy

\[ E_{\text{kinetic}}(\text{electron}) = h \nu - h \nu_0 \]
Early Quantum Mechanics

- Many huge discoveries at turn of century
  - Photoelectric effect just one of them
- Led to understanding that electrons in atoms have only certain specific energies they can exist at.
- For H, the electron can have the following potential energies:
  \[ E_n = -\frac{2.18 \times 10^{-18}}{n^2} J \]
  where \( n = 1,2,3,\ldots \)

- Note energies are negative – electron is bound
- These quantized energies lead to an energy level diagram
Quantized Energies

- Energy can be gained or lost in only specific amounts – quanta.
- **Ground state:** lowest energy state of an atom.
- **Excited state:** higher energy states – reached by absorbing light (usually).
- **Energy level diagram:** depicts the changes in energy of an atom.
- When an atom *emits* a photon (or radiates heat), it returns to a lower state.
- \( \Delta E = \pm h\nu_{\text{photon}} \)
For Hydrogen (Balmer)

\[ \nu = 3.29 \times 10^{15} \text{ s}^{-1} \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right) \]
de Broglie

- Recall light has both wave and particle properties
- de Broglie said that all things have both wave and particle properties.

\[ \lambda_{\text{particle}} = \frac{h}{mu} \]

- Where \( m \) is mass and \( u \) is velocity
- Wave properties of matter most easily observed for things with small mass (big wavelength)
Heisenberg

1. A particle occupies a particular location, but a wave has no exact position.
2. Because of their wave-like properties, electrons are always spread out in space.
3. As a result, the position of an electron cannot be precisely defined.
4. Therefore, electrons are delocalized, rather than pinpointed.

The Heisenberg Uncertainty Principle says that the more accurately we know position, the more uncertain we are about energy, and vice versa.

Since we know electron energies very well (quantum mechanics) we can’t know their positions precisely.
So instead, we identify “probable locations” of the electrons in an atom, not a exact locations.
Today

- Think about CAPA #11

By Friday

- Read Chapt 7 sections 5-6
- Work extra problems!
- Plan to attend seminars

Remember: You are done with the homework when you understand it!