Chapter 6. Electronic Structure of Atoms

- ## 6.1 The Wave Nature of Light

- The **electronic structure** of an atom refers to the arrangement of electrons.

We are going to use light as a tool to study the ATOM

- Visible light is a form of **electromagnetic radiation** or *radiant energy*.

- Radiation carries energy through space.

- Electromagnetic radiation is characterized by its wave nature.

- All waves have a characteristic **wavelength**, λ (lambda), and amplitude, *A*.

- The **frequency**, ν (nu), of a wave is the number of cycles that pass a point in one second.

- The units of ν are *hertz* (1 Hz = 1 s^{-1}).

- The speed of a wave is given by its frequency multiplied by its wavelength.

- For light, speed, $c = \lambda\nu$,

- Electromagnetic radiation moves through a vacuum with a speed of 3.00×10^8 m/s.

- Electromagnetic waves have characteristic wavelengths and frequencies.

- The *electromagnetic spectrum* is a display of the various types of electromagnetic radiation arranged in order of increasing wavelength.

• Example: visible radiation has wavelengths between 400 nm (violet) and 750 nm (red).

Laser Pointer Demo

Do a calculation of wavelength and frequency.

520 nm -→ 5.76x10^14

6.2 Quantized Energy and Photons

• Some phenomena can't be explained using a wave model of light:

Hot Objects and the Quantization of Energy

• Heated solids emit radiation (black body radiation) • The wavelength distribution depends on the temperature (i.e., "red hot" objects are cooler than"white hot" objects).

• Planck investigated black body radiation.

• He proposed that energy can only be absorbed or released from atoms in certain amounts.

• These amounts are called quanta.

• A **quantum** is the smallest amount of energy that can be emitted or absorbed as electromagnetic radiation.

• The relationship between energy and frequency is: $E = h\nu$

• where h is **Planck's constant** (6.626×10^{-34} J-s).

Calculate the energy of the Green Laser Pointer.

Watt = J/sec, laser pointer is 30 mW, how many photons in 10 seconds?

30 mW * 10 sec = 300 mJ = 0.3 J.

One photon E = 6.626 × 10^{-34} J-s * 5.76e14 = 3.8x10^-19 J/photon

photons = 0.3 J/3.8x10^-19 J/photon = 7.8e17 photons

The Photoelectric Effect and Photons

• The **photoelectric effect** provides evidence for the particle nature of light. • It also provides evidence for quantization.

• Einstein assumed that light traveled in energy packets called **photons**.

• The energy of one photon is $E = h\nu$.

• Light shining on the surface of a metal can cause electrons to be ejected from the metal.

• The electrons will only be ejected if the photons have sufficient energy (*work function*):

• Above the threshold frequency, the excess energy appears as the kinetic energy of the ejected electrons.

Show slide and typical data. Do a binding energy calculation.

Sodium BE = 2.98x10-19J

BE = hc/λ = 6.6e-34*3.0e8/ λ

λ = 667 nm

Ephoton = BE + KE

• Light has wave-like AND particle-like properties. Waves can pile up, particles can't!

Line Spectra and the Bohr Model Line Spectra

- Radiation composed of only one wavelength is called *monochromatic*.
- Radiation that spans a whole array of different wavelengths is called *continuous*.
- When radiation from a light source, such as a lightbulb, is separated into its different wavelength components, a **spectrum** is produced.

• White light can be separated into a **continuous spectrum** of colors.

• A rainbow is a continuous spectrum of light produced by the dispersal of sunlight by raindrops or mist.

• On the continuous spectrum there are no dark spots, which would correspond to different lines.

• Not all radiation is continuous.

Show a gas cell in the spectrograph

- A gas placed in a partially evacuated tube and subjected to a high voltage produces single colors of light.
- The spectrum that we see contains radiation of only specific wavelengths; this is called a **line spectrum**.

- **Bohr's Model**

- Rutherford assumed that electrons orbited the nucleus analogous to planets orbiting the sun.
- However, a charged particle moving in a circular path should lose energy.
- This means that the atom should be unstable according to Rutherford's theory.
- Bohr noted the line spectra of certain elements and assumed that

electrons were confined to specific energy states. These were called **orbits**. **Show Bohr ppt**

- Bohr's model is based on three postulates:

- Only **orbits** of specific radii, corresponding to certain definite energies, are permitted for electrons in an atom.

- An electron in a permitted orbit has a specific energy and is an "allowed" energy state.

- Energy is only emitted or absorbed by an electron as it moves from one allowed energy state to another.

- The energy is gained or lost as a photon.

The Energy States of the Hydrogen Atom

- Colors from excited gases arise because electrons move between energy states in the atom.

- Since the energy states are quantized, the light emitted from excited atoms must be quantized and appear as line spectra.

- Bohr showed mathematically that

Equation:

1/ λ = Rh (1/n1^2 – 1/n2^2)

- where *n* is the *principal quantum number* (i.e., $n = 1, 2, 3, \ldots \infty$) and R_H is the Rydberg constant. 1.097x10^7 m-1

- The product $hcR_H = 2.18 \times 10^{-18}$ J.

  - E = 2.18x10^-18 (1/n1^2 – 1/n2^2) J

- The first orbit in the Bohr model has $n = 1$ and is closest to the nucleus.

- The furthest orbit in the Bohr model has $n = \infty$ and corresponds to $E = 0$.
- Electrons in the Bohr model can only move between orbits by absorbing and emitting energy in quanta ($E = h\nu$).
- The **ground state** is the lowest energy state.
- An electron in a higher energy state is said to be in an **excited state**.
- The amount of energy absorbed or emitted by moving between states is given by

Limitations of the Bohr Model

- It cannot explain the spectra of atoms other than hydrogen. – **Ultraviolet Catastrophe**
- Electrons do not move about the nucleus in circular orbits.
- However, the model introduces two important ideas:
  - The energy of an electron is quantized: electrons exist only in certain energy levels described by quantum numbers.
  - Energy gain or loss is involved in moving an electron from one energy level to another.
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The Wave Behavior of Matter

- Knowing that light has a particle nature, it seems reasonable to ask whether matter has a wave nature.
- This question was answered by Louis deBroglie.

- Using Einstein's and Planck's equations, deBroglie derived:
- $\lambda = h/m\upsilon$
- The **momentum**, $m\nu$, is a particle property, whereas λ is a wave property.
- **Matter waves are the term used to describe wave characteristics of material particles.**
- Therefore, in one equation deBroglie summarized the concepts of waves and particles as they apply to low-mass, high-speed objects.

λ = h/mv

Velectron = 6x10^6 m/s

Mass = 9.11 x 10^-34 kg

H = 6.626 x 10^-34 J sec

λ = h/mv = 1.2 x 10^-10 J*s^2/m kg (J = kg m/s^2) so … meters

λ = h/mv = 0.12 nm or 1.2 A

Electron wavelengths are of the same order as atom size!

- As a consequence of deBroglie's discovery, we now have techniques such as X-ray diffraction and electron microscopy to study small objects.