The Wave-Particle Duality of Light

- The above title simply refers to the fact that light can be thought of as either or a particle or a wave.
- This seems contradictory, because a particle such as a bullet does not appear to have wavelength.
- Also, waves are usually thought of as disturbances in some medium, or energy; a ray of light is not something that we usually consider to be made up of particles having mass, momentum, etc.

Light Is a Wave

- Light can be described as a wave because it exhibits many of the properties of waves: reflection, refraction (bending of a wave when it passes from one medium to another – a pencil that appears "broken" when you look at it in a glass of water), diffraction (spreading of a wave as it passes a barrier), and interference.
- Like any wave, the speed of light can be related to its frequency and wavelength by the formula

  Speed = (wavelength) x (frequency)

- In the case of light, this formula becomes

$$c = \lambda\nu = 2.998 \text{ x } 10^{8} \text{ m/s}$$

- This means that the frequency of light and its wavelength are inversely related
- It is still useful to think of light as a wave in many situations. Our view of light has been modified since 1900, but the "wave model" of light has not been entirely discarded.

Problems with the Wave Model of Light

- Certain behaviors of light could not be accounted for when light was considered to be only a wave phenomenon.
- Two of these problems were
  1. The spectrum of light emitted by a "blackbody"
  2. The photoelectric effect

Planck Explains Blackbody Radiation by "Quantizing" Energy

- A simple definition of a blackbody radiator is an object that has no color itself, and which reflects no colors or light from anything else. The light given off by a perfect blackbody would in theory only be the light due to the thing glowing – presumably in response to the heat or other energy that it was absorbing.
- Max Planck explained the spectrum of blackbody radiation by assuming that energy absorbed by the radiator was only accepted or emitted in tiny specific units or amounts.
- He called these amounts "**quanta**" of energy. His theory of energy absorption and emission was therefore called the **quantum theory.**
- This model of energy was successful in explaining the observations of blackbody radiation. However, it was troubling because it suggested that energy could not be absorbed or emitted in any old amount that you wanted to use or get.
- Planck summarized his relationship between energy and frequency of light with the following formula:

$$E = h\nu$$

- "*h*" is known as Planck's constant. It has a value of $h = 6.626 \times 10^{-34}$ Js
- This means that the energy absorbed or emitted by some atom may only be some multiple of *h*

Albert Einstein Explains the Photoelectric Effect by Quantizing Light

- By the early 1900's, scientists had known about a phenomenon known as the **photoelectric effect**.
- When light is shown on certain metals, electrons are ejected from the surface of the metal.
- This effect is the basis of solar cells that convert light into electricity
- When light of a fairly low frequency (and therefore, low E) was shone upon a metal, it would not cause an emission of electrons. Even if the brightness of such light was increased, the electrons in the metal would stay put.
- However, it was observed that even if dim light of a higher frequency was shone upon this same metal, then the electrons *were* ejected.
- Furthermore, brighter light that had this sufficient frequency (a.k.a. "threshold frequency") would only increase the number of electrons ejected, not the speed (energy) of the electrons.
- Only by increasing the frequency of the light could the energy of the ejected electrons be made greater.
- Albert Einstein concluded that light was also quantized. That is, he could explain the photoelectric effect by assuming that light itself traveled in tiny little bullets that he called "**photons**". These photons are simply quanta of light energy.

Niels Bohr Proposes an Explanation of Hydrogen's Emission Spectrum

- When a gas such as hydrogen is excited by energy (electricity, light, heat, etc.) it will give off a characteristic type of light.
- Although the light may look like a single color, it is actually many individual frequencies of light.
- The light can be separated into its constituent frequencies by passing it through a prism or grating (such as the glasses that we wore in class).
- This collection of "bands" of light that a substance emits when excited is called the **emission spectrum** of that substance.
- <u>Bohr could only explain the emission spectrum of the simplest atom: hydrogen.</u>
- Bohr suggested that the individual bands in hydrogen's emission spectrum were each due to a different electron transition within the hydrogen atom.
- A huge drop from a high orbit to a low orbit would cause a very energetic photon to be emitted – that is, the emitted photon would have a high frequency (nad would be, say, violet).
- A small drop from a kinda high orbit to a lower orbit would result in a relatively low frequency photon emission (e.g., red).
- Bohr's model depicted electrons as moving in fixed paths around the nucleus. That is, electrons could only jump into or out of these "orbits."
- Bohr's model could not account for the emission spectrum of any other element.
- Bohr's model could not account for why there were orbits, why there were a certain number of orbits, or why the orbits had the "observed" distances from the nucleus.

Louis DeBroglie Suggests That If Waves Can Act Like Particles, Then Perhaps Particles Such As Electrons Can Act Like Waves.

- Einstein suggested that waves such as light can act like particles
- DeBroglie suggested the converse: maybe electron (particles) can act like waves.
- If an electron in an atom had a wavelength, its wavelength would determine how far it could be from the nucleus.
- Only "routes" around the nucleus which were whole number multiples of the wavelength would be allowed "orbits"
- This explained the number and distances of Bohr's "orbits"

Heisenberg States His Uncertainty Principle

- Heisienberg noted that it was impossible to know both the position and the momentum (kind of like speed for our purposes) of an electron at any given point in time.
- This is known as the **Heisenberg Uncertainty Principle**
- Think of it like this: we bounce light off things in order to ‘see” those things. Since an electron is about as “big” as a photon, well, just by looking at an electron we will be knocking this electron away from where it was by bouncing photons off it.
- This led scientists to discard the idea of “orbits” where electrons are definitely going to be located.

Ernst Schrödinger Devises the Wave Equation for an Atom’s Electrons

- This Austrian physicist devised an equation which described the probability of finding an electron at a given point in a given time.
- We use his equation to make approximations of where electrons probably are in atoms
- These regions where electrons are likely to be found are called “orbitals”
- That equation is so dang hard to understand and use because it uses “scary” math like partial derivatives and Hamiltonian operators. But it’s not as bad as it looks. Oh yeah, it can only be solved **exactly** (no approximations) for a hydrogen atom! But really, even though his contributions are hard to express in terms of first-year chemistry, “Schrö” really is the man.
- If you get Schrödinger’s wave equation, “$\hat{H}\Psi = E\Psi$,” tattooed on your body, Dr. Mike Messina of UNCW will give you an “A” in Chemistry 420: Quantum Mechanics. Incidentally, this has been done. The student graduated last year. (She would have gotten an “A” anyway – she’s smart!)