The Nature of Light and the Planetary Model

In this unit, you will be introduced to the dual nature of light, the quantum theory and Bohr's planetary atomic model. The planetary model was an improvement on the nuclear model and attempted to answer the question, "Where is the electron?"

The Duality of Light


Isaac Newton

In the 1700's there were two theories on the nature of light. One theory, proposed by Sir **Isaac Newton**, suggested that light consisted of a stream of tiny particles he called corpuscles. The other theory, proposed by **Christiaan Huygens**, explained that light consisted of waves. Who was right?

Well, most of the evidence at the time pointed to light being a wave and by the 1800's most scientists accepted the wave theory of light because of the work done by **Thomas Young** and **James Clerk Maxwell**. Both men were able to show that light diffracts, and diffraction can only be explained in terms of waves.

But, Maxwell also developed four equations that unified electricity, magnetism and light as expressions of the same phenomenon, electromagnetic radiation. It was these equations that solidified the idea that light consisted of waves. However, in the early 1900's the work of Max Planck, Arthur Compton, and Albert Einstein showed that light has particle properties as well. Today we accept a **wave-particle duality** concept of light.

Electromagnetic Radiation

When you light a fire in a fireplace you see and feel **electromagnetic radiation**. The fire gives off light (visible radiation) and heat (thermal radiation). Both types of radiation exist in the form of electromagnetic waves and each has energy particles. Therefore, we must understand particle and wave properties.

Light as a Wave

Phenomena such as colors in soap bubbles, oil film and the rainbow are best explained when light is thought of as a wave. So, let's look at the components of a wave.

An **electromagnetic wave** is composed of an electric wave and a magnetic wave traveling together at right angles to each other. The distance between adjacent maxima of the electromagnetic wave is the **wavelength** (λ) and half the distance from a maximum to a minimum is the **amplitude**. The number of cycles (maximum) that pass through a point in a given amount of time is known as the **frequency** (ν) and a **hertz** (Hz) is defined as a cycle per second (s^{-1}).

Draw and label a wave.

Unlike other types of waves, electromagnetic waves do not require a medium to travel and electromagnetic waves can move through a vacuum at **3.00 x 10^8 m/s**, "*the speed of light*".

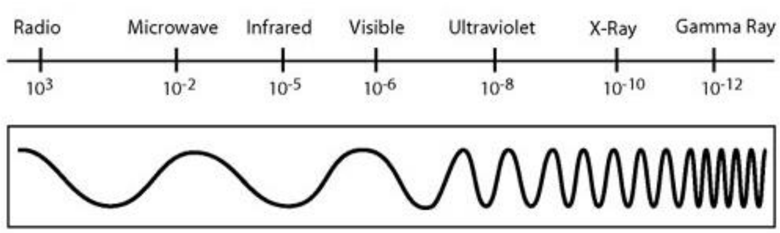
It is known that the speed of any wave is the product of the wavelength and frequency and since the speed of light is constant, the wavelength (λ) and frequency (ν) must be inversely proportional.

$$\lambda = \frac{c}{\nu}$$

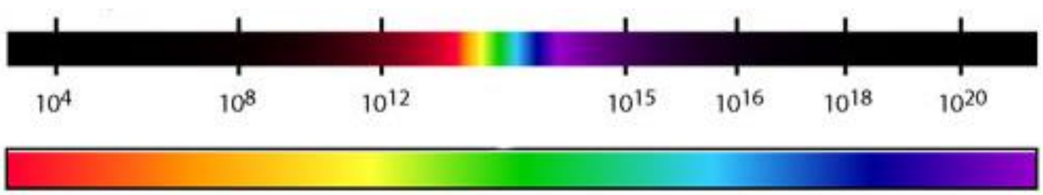
- **shorter** the wavelength the **higher** the frequency
- **longer** the wavelength the **lower** the frequency

The Spectrum

The range of electromagnetic radiation wavelengths is called the **electromagnetic spectrum** and ranges from very short wavelengths (cosmic rays) to very long wavelengths (thermal waves). Visible electromagnetic radiation (white light) is a very small region of the electromagnetic spectrum that spans from 750 nm to about 380 nm.



Visible light can be broken into its component wavelengths by passing it through a **prism.** The prism bends the light (**refraction**) as the light passes through and produces a complete array (no spaces) of colors (wavelengths) called a **continuous spectrum**.

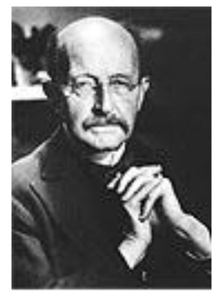


All light can be separated into component wavelengths using a **spectrometer,** but not all radiation creates a continuous spectrum. Many types of radiation have certain wavelengths missing and create a line spectrum. A **line spectrum** is a spectrum with bright lines appearing at certain wavelengths only (spaces)**.**

Light as a Particle

Though the wave theory of light seemed to answer many of the questions concerning light, there were certain phenomena that could not be explained by this idea. Phenomena such as black body radiation, the photoelectric effect and the Compton Effect pointed to the possibility of light being a particle.

Then in 1900, a German physicist, **Max Planck** while studying the effects of light given off by glowing matter suggested that atoms might behave like little oscillators and that the energy from the oscillating might be related to the frequency of the light emitted by glowing matter.


Max Planck

By comparing the frequencies and energies of different wavelengths of light, Planck not only showed that the frequencies and energies were related but that they were directly proportional and constant. And the calculation for the value of this proportionality constant is now known as Planck's constant. ($\boldsymbol{h}$ = 6.626 x 10^{-34} J·s)

The implications of this proposal were unsettling even for Planck. His discovery meant that energy was not a continuous flow, but was discrete like matter. This idea went against all classical physics but Planck was a fierce experimentalist and believed that facts are facts. So, in late 1900, Planck presented Planck's Hypothesis which stated that energy was not continuous and could only be emitted in discrete amounts (quantized) that were multiples of an elementary unit.

Planck developed an equation to support his hypothesis using the data he collected from studying the frequency and energy of various wavelengths of light.

$\Delta E = h\nu$ **Planck's Equation**

However, Planck's revolutionary idea was not well accepted, until a young Swiss patent clerk successfully used the quantum theory in his explanation of the photoelectric effect.

"... an act of despair... I was ready to sacrifice any of my previous convictions ... After all, even good scientists should be skeptical. Even of their own work." – Max Planck

The Photoelectric Effect

Albert Einstein

The **photoelectric effect** was first described by **Heinrich Hertz** in 1887. Hertz noticed that when certain short wavelengths of light were shone onto a shiny metal surface in a vacuum tube, electrons were emitted. It was also noted that the intensity of the light didn't matter if the energy was high enough (short wavelength). These observations lead to the question, how can the electrons be moved, if light is a wave only?

In 1905, **Albert Einstein** used Planck's constant to help explain the photoelectric effect and show that electromagnetic radiation has particle as well as wave properties. Starting with his own equation and then substituting Planck's equation in for energy, Einstein was able to show that a quantum of energy has mass. In fact, the higher the energy the greater the mass and the more it was like a particle.

$$E_{photon} = \frac{hc}{\lambda} \quad \text{and} \quad m = \frac{E}{c^2} \quad \therefore \quad m = \frac{h}{\lambda c}$$

The Quantum Theory

Quanta, now called *photons*, give light its particle properties. The photon is a specific "hunk" of energy directly proportional to its frequency, inversely proportional to its wavelength and can only be absorbed or released in whole integer amounts (quantized).

When the energy is high the wavelength is short and the photon acts as a particle, but when the energy is low the wavelengths are long and the photon acts as a wave. Therefore, light has a **wave - particle duality**.

The Hydrogen Spectrum

The spectroscope is an instrument that uses a prism to separate light into its component wavelengths. In 1860 **Robert Bunsen** and **Gustav Kirchhoff** used a **spectroscope** to study the light given off by heated objects. They discovered that each element produced a **line spectrum** with its own characteristic pattern, much like a fingerprint. Since, hydrogen was the simplest; scientists studied its line spectrum extensively.

In 1885 **Johannes Balmer** developed a mathematical relationship between the hydrogen spectral lines in the visible region. This series of lines came to be called the Balmer series. Later, **Johannes Rydberg** put this relationship in a more convenient formula shown below. Where **Z** is the atomic number, **n** is the energy level and **R** is the Rydberg constant with a value of **1.097×10^{7} m^{-1}**.

$$\frac{1}{\lambda} = R\left(\frac{Z^2}{n^2}\right) \qquad \text{or} \qquad \frac{1}{\lambda} = R\left(\frac{Z^2}{n_2^2} - \frac{Z^2}{n_1^2}\right) \quad (n_1 < n_2)$$

Rydberg's equation could account for the visible spectral lines and the lines in the ultraviolet and infrared regions as well. The ultraviolet series was discovered by Theodore Lyman in 1906 and is known as the Lyman series. Then in 1908, the infrared series was discovered by Friedrich Paschen and is called the Paschen series.

Hydrogen Spectrum

Bohr's Conundrum

According to Newtonian physics, if a negatively charged particle (electron) moves about a positively charged particle (nucleus) energy will be released. Consequently, the electron would lose energy, slow down and move closer to the nucleus until it collides with the nucleus and destroys the atom.

But, this doesn't happen and the Rutherford model can't account for this phenomenon. Therefore, a new model is needed, a model that includes the new information from both the quantum theory and the hydrogen spectrum.

Energy Levels

In 1913 Danish physicist, **Niels Bohr**, suggested applying the quantum theory to the Rutherford model to explain why electrons don't spiral into the nucleus and that the Balmer-Rydberg equation showed the location of the electrons with respect to the nucleus.

Niels Bohr

Bohr's new model proposed that electrons are in fixed energy levels (n) he called orbits and the location of these orbits is restricted to certain radii measured from the nucleus. The energy of these orbits is quantized, and electrons must absorb or release energy at certain wavelengths to move between energy levels.

The lowest energy level of an electron is the **ground state** and the higher energy levels are **excited states**. If an electron absorbs photons of a certain wavelength, the electron would be **excited** to a higher energy level inversely proportional to that wavelength. But, the lowest energy level possible would be the most stable so the electron would immediately release the photons enabling the electron to return to its ground state.

If the quantum theory is true, particles that radiate energy must radiate energy as whole photons (quanta) only. There could be no in between energy levels for the electron. It would have to move from energy level to energy level completely or not at all. This would produce a line spectrum with dark portions instead of a continuous spectrum. Bohr's application of the quantum theory to the electrons in an atom solved an age-old problem concerning the line spectra of the elements.

Bohr's Calculations

Bohr's work led him to an equation that allowed him to calculate the potential energy of an electron (energy level). Also, Bohr mathematically confirmed the value of the Rydberg equation.

$$E_n = \frac{Rhc}{n^2} \quad \text{(n = energy level)}$$

Then using this equation, the Rydberg equation, combined with Planck's and Einstein's work, Bohr developed an equation that calculated the energy of a photon and allowed him to arrange the hydrogen orbits and account for the Lyman, Balmer and Paschen series of spectral lines.

$$E = 2.180 \times 10^{-18}\,\mathrm{J}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) \qquad \text{or} \quad \frac{1}{\lambda} = \frac{2.180 \times 10^{-18}\,\mathrm{J}}{hc}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

Bohr's equation for the hydrogen atom

The Planetary Model

The planetary model of the atom states that an atom is mostly space with a very dense, small, centrally located, positively charged nucleus surrounded by negatively charged electrons in energy levels (orbits) in the atomic space.

Think It Through

- Explain wave-particle duality.
- What is electromagnetic radiation? the speed of light?
- Describe the Quantum Theory.
- What is the photoelectric effect?
- Explain the relationship between wavelength and frequency.
- Explain the relationship between frequency and energy.
- What is a line spectrum? a continuous spectrum?
- Describe the Planetary Model.

Vocabulary

Frequency (ν)	amplitude	wavelength (λ)	continuous spectrum
quantized	quantum (a)	energy level	electromagnetic
excited state	photon	ground state	line spectrum
spectroscope	orbit	spectrum	Planck's Constant

Ideas

electromagnetic radiation
wave-particle duality
photoelectric effect
quantum theory
energy levels

People

Max Planck	Christiaan Huygens
Niels Bohr	Johannes Rydberg
Isaac Newton	James Clerk Maxwell
Johann Balmer	Albert Einstein

Video Toolbox

Understanding Light as a Wave
The Electromagnetic Spectrum
The Photon and Light as a Particle
The Planetary Model

"Excellence is the gradual result of always striving to do better."
- Pat Riley