Chem 103, Section F0F
Unit II - Quantum Theory and Atomic Structure
Lecture 6

- Light and other forms of electromagnetic radiation
- Light interacting with matter
- The properties of light and matter

Lecture 6 - Atomic Structure

- Reading in Silberberg
  - Chapter 7, Section 1 *The Nature of Light*
  - Chapter 7, Section 2 *Atomic Spectra*
  - Chapter 7, Section 3 *The Wave-Particle Duality of matter and Energy*

Lecture 6 - Introduction

We have seen how several models of nature that have been changed or overturned:

- Lavoisier's theory of combustion overturned the phlogiston model for fire.
- Dalton's established the modern model that the elements were made of indivisible atoms.
  - Dalton's model was altered with the discovery of subatomic particles.
- Rutherfords nuclear model of the atom overturned J. J. Thomson's "plum pudding" model of the atom

Lecture 6 - Introduction

Rutherford's model was almost immediately questioned because it could not explain why the electrons in an atom did not spiral down into the nucleus

- The subatomic particles seemed to be contradicting real-world experiences.

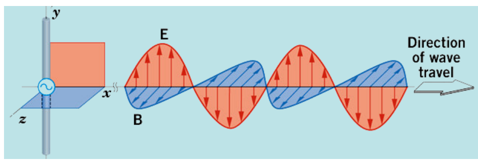
Lecture 6 - Introduction

Discoveries in the early 20th century lead to a blurring in the distinctions between particulate characteristics of matter and wave-like characteristics of energy

- These discoveries lead to a powerfull model for the electronic structure of atoms.
- This model allows us to predict much about the structure and properties of the atoms and of the compounds and systems they are comprise them

Lecture 6 - The Nature of Light

Electromagnetic radiation is a type of energy consisting of oscillating electric (E) and magnetic (B) fields, which move through space.



- Forms include:
  - Infrared radiation
  - Ultraviolet radiation
  - X-rays
  - Gamma-rays
  - Radio waves
  - Microwaves

Lecture 6 - The Nature of Light

Electromagnetic radiation is a type of energy consisting of oscillating electric (E) and magnetic (B) fields, which move through space.

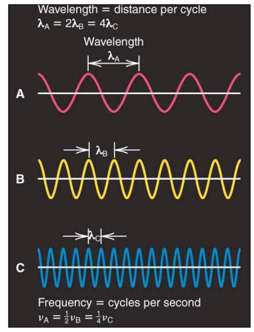
- The earliest characterization of electromagnetic radiation was based on its wave-like properties
  - Wavelength, λ (m)
  - Frequency, v (1/s)
  - Velocity, c = 2.998 x 10^8 m/s
  - Amplitude

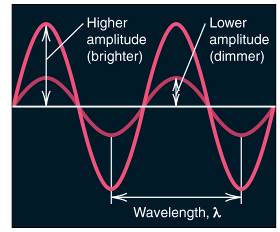
$$c = \lambda \times v$$

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

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

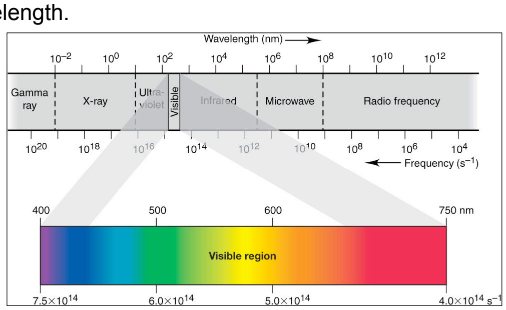
Lecture 6 - The Nature of Light





Lecture 6 - The Nature of Light

Using the wave-like description of electromagnetic radiation, the different forms are characterized by their frequency or wavelength.



Lecture 6 - Question

Today I noticed a maple tree outside of my office window had turned orange. It looks orange to me because it is reflecting orange light (λ=670 nm) while absorbing the other colors of visible light.

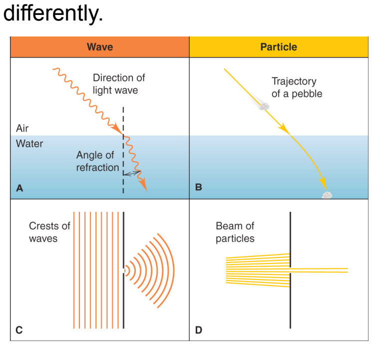
What is the frequency of the light being reflected by this tree?

Lecture 6 - The Nature of Light

Based on classical observations, energy (waves) and matter (particles) behave differently.

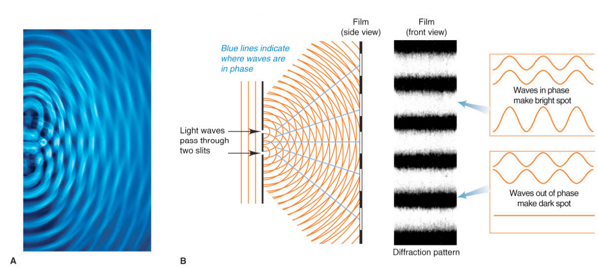
Refraction

Diffraction



Lecture 6 - The Nature of Light

Waves, but not particles, exhibit interference.



Lecture 6 - The Nature of Light

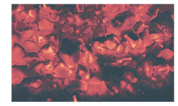
During the early 20th century, there were several observations that confounded the the wave-like description of “light”

- Blackbody radiation
  - The colors of light emitted by an hot object
- The Photoelectric Effect
  - The behavior of electrons ejected by light from a metal
- Atomic Emission Spectra
  - The wavelengths of light emitted by excited atoms of an element

Lecture 6 - The Nature of Light

Blackbody radiation

- When an object is heated it emits light
- As the object is heated to a higher temperature, the distribution of wavelengths of the light emitted changes.
- The distribution of wavelengths is not consistent with a model that allows the light energy to have any value.



A Smoldering coal

B Electric heating element

C Lightbulb filament

Lecture 6 - The Nature of Light

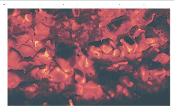
Max Planck proposed that this meant that light energy could only be emitted with energies of fixed value

- He proposed that the light energy is emitted (or absorbed) in packets (**quanta**) of fixed energies, with values directly proportional to the frequency of the light being emitted (or absorbed):

$E = h\nu$ (The energy, *E*, of 1 quantum of light with frequency *v*.)

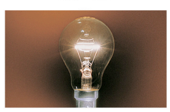
$h = 6.6261 \times 10^{-34}$ J•s (Planck's constant)

Max Planck received the Nobel Prize in Physics for this work in 1918



A Smoldering coal

B Electric heating element

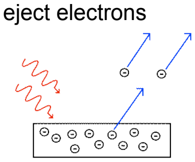


C Lightbulb filament

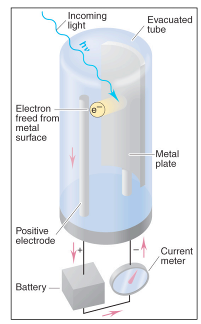
Lecture 6 - The Nature of Light

The Photoelectric Effect

- When light strikes a metal surface it can eject electrons



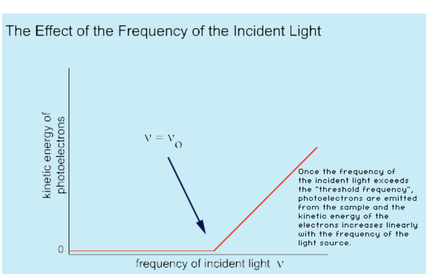
- The ejected electrons can be measured as a current using a vacuum tube



vacuum tube

Lecture 6 - The Nature of Light

The Photoelectric Effect



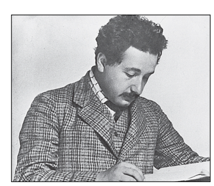
View Simulation

Lecture 6 - The Nature of Light

The Photoelectric Effect

- Albert Einstein (1879-1955) published an explanation for the photoelectric effect in 1905.
  - He explained that the energy of the light striking the metal plate was not dependent on the the intensity of the light, but rather, it frequency.
  - The light was behaving like it was made up of particles (**photons)**
  - The energy The energy of a photon is given by Planck’s equation, and is proportional to its frequency.

$$E_{\text{photon}} = h\nu = \Delta E_{\text{atom}}$$



Einstein received the Nobel Prize in Physics for this work in 1921.

Lecture 6 - The Nature of Light

The combined explanations for blackbody radiation (Max Planck) and the photoelectric effect (Albert Einstein) showed that in addition to its wave-like behavior, light also had a particle like behavior

Wave-like Behaviors	*Particle-like Behaviors*
Refractions	*Light energy can only be absorbed or emitted by matter in discrete packets (quanta/photons)*
Diffraction	
Interference	

Lecture 6 - Question

Today I noticed a maple tree outside of my office window had turned orange. It looks orange to me because it is reflecting orange light (λ=670 nm) while absorbing the other colors of visible light.

What is the energy of a photon of orange light?

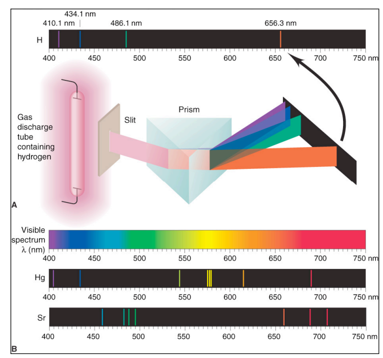
Lecture 6 - Atomic Spectra

Atomic spectra

- When atoms are excited they gain energy)
  - They can then release that energy as light.
- A spectrum is a plot of the wavelengths (or frequencies) of the the light that is emitted.

Lecture 6 - Atomic Spectra

The hydrogen gas in the tube is excited by applying a high voltage.

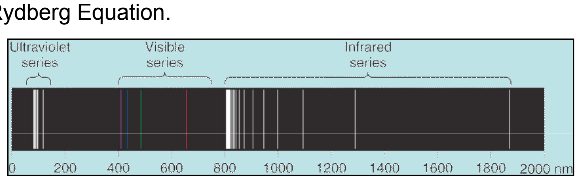


The different wavelengths of light emitted by the excited hydrogen atoms are dispersed with a prism.

The atoms of different elements produce different atomic spectra

Lecture 6 - Atomic Spectra

The “lines” in the hydrogen atomic spectrum can be fit to the Rydberg Equation.



$$\frac{1}{\lambda} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

where

R is the Rydberg constant ($1.097 \times 10^7\ \text{m}^{-1}$)

n_1 and n_2 are positive integers with $n_2 > n_1$

Lecture 6 - Atomic Spectra

Niels Bohr (1885-1962) combined the ideas of Planck and Einstein, along with the atomic spectrum for the hydrogen atom.

- He proposed that light is emitted from an atom when an electron in one energy state (initial) changes to a new, lower energy state (final).

$$\Delta E = E_{\text{final}} - E_{\text{initial}}$$

- The energy of the photon emitted is equal to this change in energy

$$E_{\text{photon}} = \Delta E = h\nu$$

Lecture 6 - Atomic Spectra

Niels Bohr (1885-1962) combined the ideas of Planck and Einstein, along with the atomic spectrum for the hydrogen atom.

- The atomic spectra indicated that the electron in a hydrogen atom can have only certain energy values.
  - These energy values are called **stationary states.**
  - An atom does not radiate light energy when in a stationary state.
- An atom radiates light energy when it changes from one stationary state to another.
  - The difference in energies between these states is given by the Rydberg equation

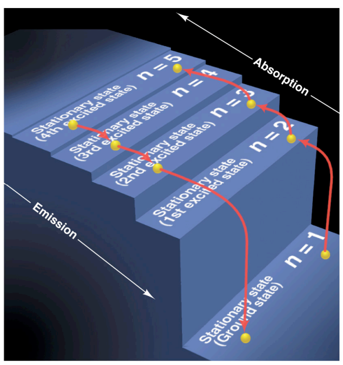
Lecture 6 - Atomic Spectra

Combining Planck's equation with the Rydberg equation gives:

$$\begin{aligned}
\Delta E &= E_{\text{final}} - E_{\text{initial}} \\
&= E_{\text{photon}} \\
&= h\nu \quad \text{(Planck's equation)} \\
&= \frac{hc}{\lambda} \\
&= hcR\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) \quad \text{(Rydberg equation)} \\
&= -hcR\left(\frac{1}{n_{\text{final}}^2} - \frac{1}{n_{\text{initial}}^2}\right) \quad (\Delta E \text{ is negative for emissions}) \\
&= -\left(6.626 \text{ x } 10^{-34}\ \text{J}\bullet\text{s}\right)\left(2.998 \text{ x } 10^{8}\ \frac{\text{m}}{\text{s}}\right)\left(1.097 \text{ x } 10^{-18}\ \frac{1}{\text{m}}\right)\left(\frac{1}{n_{\text{final}}^2} - \frac{1}{n_{\text{initial}}^2}\right) \\
\Delta E &= -\left(2.18 \text{ x } 10^{-18}\ \text{J}\right)\left(\frac{1}{n_{\text{final}}^2} - \frac{1}{n_{\text{initial}}^2}\right)
\end{aligned}$$

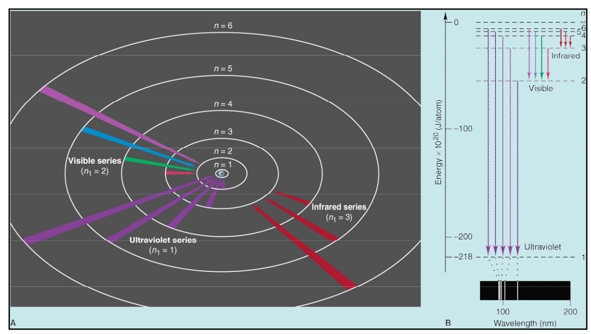
Lecture 6 - Atomic Spectra

$$\Delta E = -\left(2.18 \text{ x } 10^{-18}\ \text{J}\right)\left(\frac{1}{n_{\text{final}}^2} - \frac{1}{n_{\text{initial}}^2}\right)$$



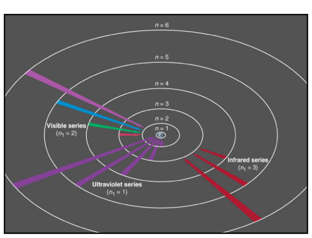
Lecture 6 - Atomic Spectra

Bohr's model of the hydrogen atom:



Lecture 6 - Atomic Spectra

Bohr received the 1922 Nobel Prize in Physics for this work



Lecture 6 - Atomic Spectra

While Bohr's model will work with other atoms that contain just a single electron

- He^{+}, Li^{2+} and Be^{3+}

Bohr's model does not work with atoms containing more than one electron.

Lecture 6 - Question

Use the Rydberg equation to determine the wavelength in nm of light that is emitted from a hydrogen atom when an electron drops from the n=4 to the n=2 energy level.

What color will this look like to your eyes?

What is the energy of the photon, E_{photon}, that was emitted?

What is the energy for a mole of these photons?

What is the energy difference, ΔE, for the transition?

(See Sample Problem 7.3)

Lecture 6 - Wave-Particle Duality

The Wave-Particle Duality of Matter and Energy

- It turns out both light energy and particles have both wave-like and particle-like behaviors.

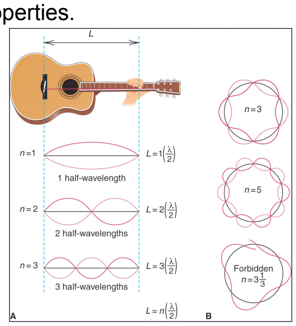
As an addendum to his 1905 paper on special relativity, Albert Einstein proposed an equivalency between energy and mass in his now famous equation:

$$E = mc^2$$

Lecture 6 - Wave-Particle Duality

After Planck and Einstein showed that light energy has a particle-like behavior, other scientists began to ask whether matter might have wave-like properties.

- Louis de Broglie (1894-1987) reasoned that if electrons have wave-like properties, that this might explain why the electrons in Bohr's hydrogen atom can have only certain fixed values.
  - Analogous to a plucked guitar string that can vibrate at only certain fixed frequencies.



Lecture 6 - Wave-Particle Duality

de Broglie combined Einstein's equation ($E = mc^2$) with Planck's equation ($E = h\nu = hc/\lambda$):

hc/λ = mc2, h/λ = mc, λ = h/mc

$$\frac{hc}{\lambda} = mc^2$$

$$\frac{h}{\lambda} = mc$$

$$\lambda = \frac{h}{mc}$$

For a particle that travels at a velocity, *u*, less than the speed of light

$$\lambda = \frac{h}{mu}$$

Lecture 6 - Wave-Particle Duality

de Broglie originally presented his theory in his one-page Ph.D. thesis.

- For this work he received the 1929 Nobel Prize in Physics

$$\lambda = \frac{h}{mu}$$

Lecture 6 - Wave-Particle Duality

de Broglie's equation can be used the determine the wavelengths for objects with mass.

$$\lambda = \frac{h}{mu}$$

Table 7.1 The de Broglie Wavelengths of Several Objects

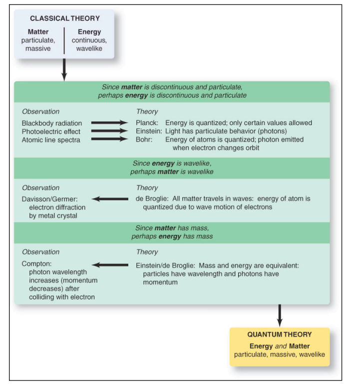
Substance	Mass (g)	Speed (m/s)	λ (m)
Slow electron	9×10^{-28}	1.0	7×10^{-4}
Fast electron	9×10^{-28}	5.9×10^{6}	1×10^{-10}
Alpha particle	6.6×10^{-24}	1.5×10^{7}	7×10^{-15}
One-gram mass	1.0	0.01	7×10^{-29}
Baseball	142	25.0	2×10^{-34}
Earth	6.0×10^{27}	3.0×10^{4}	4×10^{-63}

Lecture 6 - Wave-Particle Duality

For small particles, such as electrons, which are traveling at high velocities, this wave-like behavior can be seen.

- Davisson and Germer did this in 1927 by diffracting electrons from a nickel crystal
- On the flip side, Arthur Compton demonstrated in 1923, that photons have momentum, $p = mu$

Lecture 6 - Wave-Particle Duality



Lecture 6 - Wave-Particle Duality

Consequently, both mass and light energy have both wave-like and particle-like properties.

- These discoveries lead to a new field of physics called **quantum mechanics**.

One corollary of quantum mechanics is embodied in Heisenberg's uncertainty principle:

$$\Delta x \bullet m\Delta u \leq \frac{h}{4\pi}$$

- which says that we cannot know with certainty both a particle's position, *x*, and its velocity, *u*.

Lecture 6 - Wave-Particle Duality

For large objects this uncertainty is insignificant, but for small particles, such as electrons, it is significant.

- For this work, Werner Heisenberg (1901-1976) received the 1932 Nobel Prize in Physics

Lecture 6 - Question

How fast must a 140 g baseball travel in order to have a de Broglie wavelength that is equal to that of an x-ray photon with λ = 100. pm?

Unit II - Up Next

Unit II - The Elements and the Structure of Their Atoms

- The quantum-mechanical model of the atom

The End