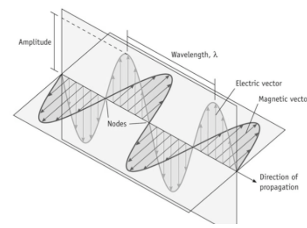
Electronic Structure of the Atom

An Introduction – Chapter 7

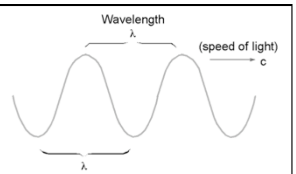
Electromagnetic Radiation

- Visible light is energy. It is one form of ***Electromagnetic (EM) Radiation***.
- The electromagnetic spectrum includes many different classifications of radiation, including gamma rays, UV rays, microwaves, and radio waves.



- **Light exhibits the properties of waves.**

Consider a wave:



- **Wavelength (λ)** - The distance from one wave crest to the next wave crest.
  - Measured in m, μm, nm, Å.
- **Frequency (ν)** - The number of wave cycles that pass a fixed point in a given period of time.
  - Measured in cycles per second (1/s) which is also called Hertz.

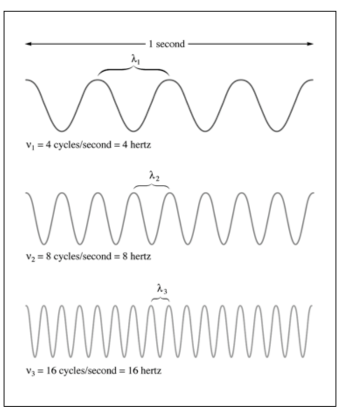
- All EM radiation travels at the **speed of light, c = 3.00 x 10^8 m/s** (2.99792 x 10^8 m/s).
- *Wavelength and frequency are related by the speed of light:* **$c = \lambda \nu$**
- The speed of light (**c**) is a constant.
- Wavelength and the frequency are ***inversely proportional:*** *as the wavelength increases, the frequency decreases.*

Long wavelength → small frequency

Short wavelength → high frequency

The nature of waves.

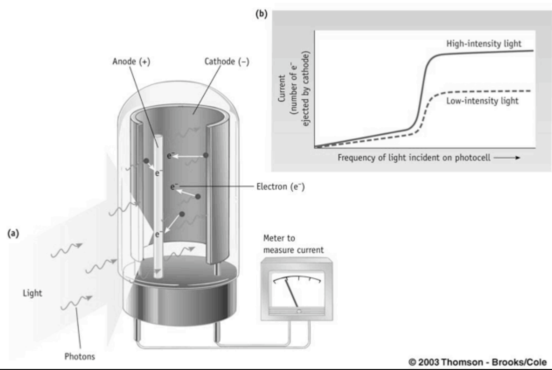
Note that the radiation with the shortest wavelength has the highest frequency.



Wave Calculations:

- Frequency is measured in:
  cycles/second = 1/s = s^{-1} = Hertz = Hz
  or some multiple (*eg.* kHz or MHz)
- Wavelength is measured in meters (m) or some multiple (*eg.* nm or μm).
- Example problems:
  1) What is the frequency of light with a wavelength of 535 nm?
  2) What is the wavelength of a radio broadcasting at 91.7 MHz?

Photoelectric Effect



The Photoelectric Effect

- The photoelectric effect occurs when light strikes the surface of a metal and electrons are ejected those electrons then hit a metal target and can comple an electrical circuit.
- Scientists made the following observations regarding the photoelectric effect:
  - Light of a minimum frequency (maximum wavelength) was required for electrons to be ejected.
  - Electricity would not flow if light of a lower frequency was used, no matter how bright the light.
  - As long as the light was of the threshold frequency or greater, the brighter the light, the more electrons that were ejected from the metal.

EINSTEIN and the Photoelectric Effect

- The photoelectric effect could not be fully explained by the wave theory of light.
- Einstein proposed:
  - Only a "packet" or photon of energy that had some minimum quantity of energy could cause an electron to be ejected from an atom.
  - The minimum energy corresponded to a minimum or “critical” or “threshold” frequency (higher frequency means higher energy).
  - The greater the intensity of light above the critical frequency, the greater the number of photons, therefore the more electrons ejected.

Wave – particle duality of EM radiation

- Max Planck theorized that energy transitions within atoms are quantized – only specific amounts of energy are allowed.
- He reasoned that packets of electromagnetic energy were absorbed or emitted when these transitions occurred. He called these packets of energy quanta.
- A photon and a quantum (plural = quanta) are the same thing.
- While light exhibits the properties of waves, it also exhibits the properties of particles (matter).
- Light (and other EM radiation) can be thought of as traveling in packets of energy called ***photons***. Photons are quantized – they transmit a specific amount of energy.

Planck’s Equation:

- Planck related the frequencies of EM radiation to the energies of vibrational transitions in matter to give the equation:

$$E = h\,\nu$$

where ***h = Planck’s constant = 6.626 x 10^{-34} J•s***

This equation represents the energy of a SINGLE photon of the given frequency.

- Energy and frequency are *directly proportional*:
  - High frequency = short wavelength = high energy
  - Low frequency = long wavelength = low energy

Energy and EM radiation

1. What energy will a photon of light with a wavelength of 235 nm (UV radiation) transmit?
2. What energy will a mole of photons with a frequency of 5.66 GHz transmit?
3. What is the maximum wavelength of EM radiation that can cause an electron to be ejected from the surface of Magnesium metal, which has an ionization energy of 738 kJ/mol?

The Bohr Model of the Atom

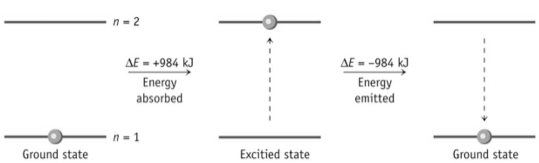
- **Neils Bohr** proposed the following model of the atom, which is call the ***Bohr model***, or ***planetary model*** of the atom.
- Though we no longer think of the atoms in true orbit around the nucleus, the quantized nature of Bohr's atom is useful in understanding electronic transitions in the atom.

Bohr Model (continued)

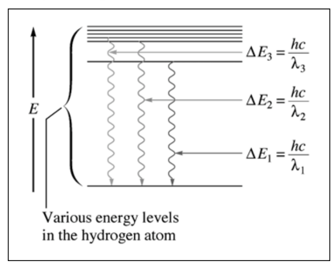
- The nucleus is at the center of the atom.
- Electrons reside in fixed ***energy levels***, or distances from the nucleus. Electrons cannot be in between energy levels.
- The greater the energy difference between energy levels, the greater the energy of the photon emitted or absorbed when the electron moves between levels.

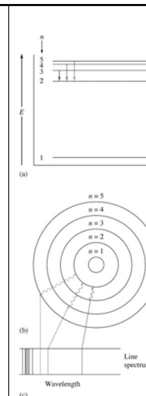
Bohr model and electronic transitions

- **Absorption:** When an electron in its lowest possible energy level, or **ground state** gains energy by absorbing a photon of light, it moves to a higher energy level, or **excited state**.
- **Emission:** When an electron moves from a high energy level to a lower energy level, it will give off a photon of light.



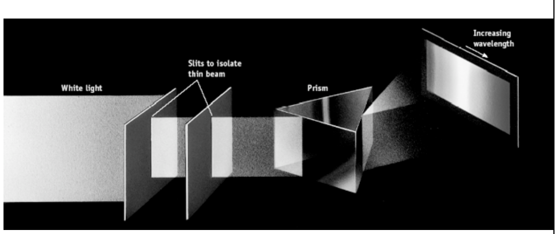
Electronic transitions in the Bohr model for the hydrogen atom.





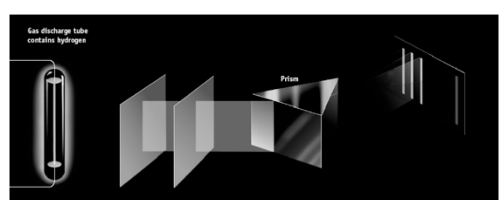
White light spectrum

- White light (eg. Sunlight) can be separated into its components by refraction through a prism to give a *continuous spectrum* (a rainbow).



Gas Emission Spectrum

- When atoms of an element are excited to high energy states by electricity, they will eventually **relax** to lower energy states, emitting photons.
- The light can be resolved into its constituent wavelength by means of a prism:



Emission spectrum

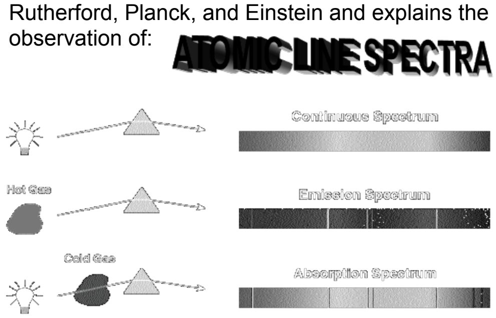
- When an element is subjected to high voltage, electrons are excited from the ground state to an excited state.
- When the electrons ***relax*** to lower energy levels, a photon equivalent to the energy transition is emitted.
- A bright line is seen in the spectrum at the wavelength corresponding to the energy of the photon.

Absorption spectrum

- When white light passes through an element, photons of a particular energy will match the energy difference between two electron energy levels.
- Absorption of those photons will cause an electron to move to from a lower energy state to a higher energy state.
- Because the photons of a particular energy are all absorbed, a particular wavelength of light is not seen in the spectrum.

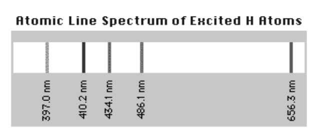
Bohr Model (continued)

- Bohr's model of the atom incorporates the ideas of Rutherford, Planck, and Einstein and explains the observation of:



Hydrogen emission spectrum – Balmer Series

- Below are the lines observed for Hydrogen transitions in the VISIBLE region of the electromagnetic spectrum.
- The shorter wavelength lines correspond to larger energy transitions in the Hydrogen atom.



The HYDROGEN atom and the BOHR model

- Energy and wavelength are inversely proportional and described by the relationship: $E = \frac{hc}{\lambda}$
- The energy of a particular energy level (**E_n**) *in a hydrogen atom* can be defined as the energy (of a photon) released when an electron moves from an infinite energy level to a given level (**n**): $E_n = -\frac{Rhc}{n^2}$

where R = Rydberg constant = 1.0974 x 10^7 m^{-1} and n = principal quantum number.

*Note: the constant term **Rhc = 2.18x10^{-18} J***

Energy of a Hydrogen energy level:

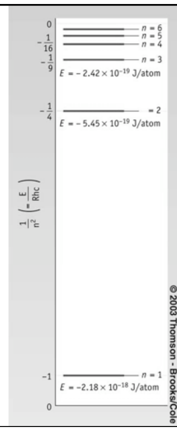
- Calculate the energy of the 4th energy level in the hydrogen atom.

-1.36 x 10^{-19} J

- Note: The value for the energy level is negative, because energy is released when an electron moves from infinity to an energy level within the atom.

Hydrogen Energy Levels

- The energy difference between levels DECREASES as the energy level (n) increases.
- The energy of each level is a NEGATIVE value, as it represents the energy RELEASED when an electron moves to that energy level from n = ∞



Electronic transitions in the H atom:

- The energy of a photon emitted or absorbed equals the energy change when an electron moves between energy levels:

$E_{photon} = |\Delta E|$

Sign indicates photon absorbed (+) or released (-).

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

Substituting:

$$\Delta E = -Rhc\left(\frac{1}{n^2_{final}} - \frac{1}{n^2_{initial}}\right)$$

Rydberg Equation

- Emission line wavelengths can also be calculated directly using the Rydberg equation in its original form:

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

where $n_2 > n_1$

Examples of energy transitions:

- Calculate the energy of a photon for the transition from the 3rd energy level to the 6th energy level of Hydrogen in J/atom and kJ/mol.
- Calculate the wavelength of light emitted when an electron in the 4th energy level of the Hydrogen atom relaxes to the 2nd energy level. What color line will be observed in the spectrum?