Introduction

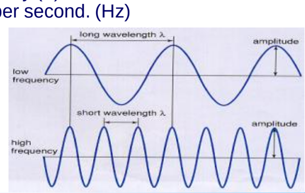
Much of the information we know about electrons comes from studies of interactions of light and matter.

In the early 1900's, scientists discovered that light has properties of both a wave and matter.

Electromagnetic Waves

Light, X-rays, gamma rays and radio waves are all examples of electromagnetic radiation. Waves can be described by five properties:

1) Amplitude - the height of the wave. (m)
2) Wavelength (λ)- A crest and trough of a wave (m)
3) Frequency (ν) - how fast the wave oscillates up and down per second. (Hz)



Electromagnetic Waves

More Properties:

4) Speed (c)- electromagnetic radiation all moves at a constant speed 3 x10^8 m/s.
5) Energy - amount of work done by the wave (J)

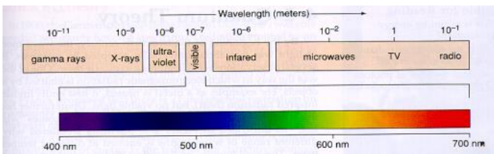
From these properties, we can determine the wavelength and the frequency using the equation:

$$c = \lambda\nu$$

Types of EM Radiation

There are 7 types of electromagnetic radiation:

1) Radio waves - lowest E 2) Microwaves
3) Infrared (IR) - heat 4) Visible Light - only type visible
5) Ultraviolet (UV) 6) X-rays
7) Gamma Rays - highest E



Planck's Theory

As wavelength increases, the ability of the radiation to penetrate bodies decreases.

Planck proposed that energy must be released in "chunks" of some minimum size. All energy is a multiple of this minimum size, or quantized.

Planck's theory - Energy must be proportional to the frequency:

$$E = h\nu$$

where $h = 6.63 \times 10^{-34}$ J•s

The Photoelectric Effect

When light of sufficient energy is shined on a clean metal surface, the surface will emit electrons.

Einstein proposed that light energy packets, or photons, hit the electrons of the atoms of the metal, and gave them sufficient energy to remove the electron.

The Bohr Model of the Atom

Bohr developed Rutherford's solar system model of the atom to include energy levels (designated by a quantum number, n).

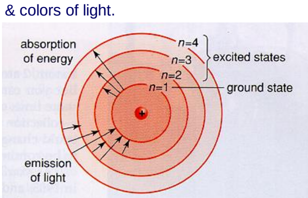
He stated that all electrons of the atom will fill the energy level orbital closest to the nucleus before moving to the next higher orbital. The ground state was the lowest energy level designated by n = 1.

Bohr Model

Bohr's theory also provided for a jumping of electrons to an excited state. When this happens, energy is absorbed by the atom, and the electron moves to a higher orbital.

When the electron falls back to its original energy level, it gives off quantized energy ($E = h\nu$), in the form of electromagnetic radiation.

Depending on the drop, the same atom can give off many different types & colors of light.



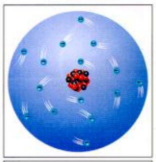
De Broglie & Heisenberg

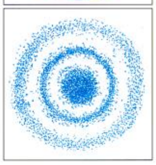
De Broglie proposed that if waves could behave as a stream of particles, then matter can behave as a wave. He concluded the wave-particle duality could be written as

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

Heisenberg extended this theory to include that if matter acted as a wave, it would be unable to determine a small object's (like an electron's) position and movement at the same time. This led us to the cloud model of the atom.

Electron Clouds





There are distinct probabilities where an electron can be found in an atom. It is in the form of an electron cloud around the nucleus.

Within this electron cloud, there are regions in which electrons are found, called an orbital.

An orbital is a region around the nucleus where an electron of given energy is found. Orbitals have characteristic shapes and sizes.

Orbitals and Energy

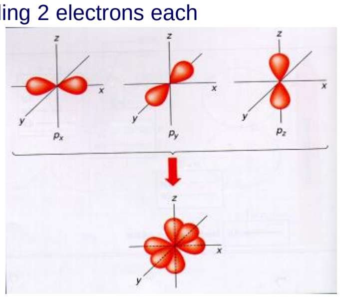
Orbitals tell the region of where 90% of the electrons are going to be.

There are four types of orbitals, designated s, p, d and f.

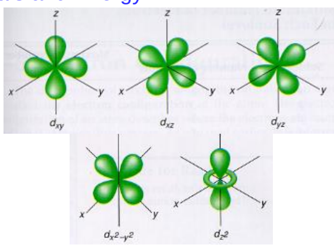
S-orbitals hold 2 e- and are shaped as shown here

Orbitals and Energy

p-orbitals hold 6 electrons, with each sub-orbital holding 2 electrons each

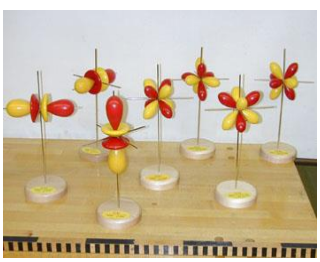


Orbitals and Energy



d-orbitals hold 10 electrons, with each of the 5 sub-orbitals holding 2 electrons each

Orbitals & Energy



f-orbitals hold 14 electrons, with each of the 7 sub-orbitals holding 2 electrons each

Orbital Diagrams

The distribution of electrons among the orbitals is called the orbital diagram of the atom.

These are determined by distributing the electrons in the orbital levels and sublevels based on a set of principles.

Principles of Orbital Diagrams

1) The Aufbau Principle - electrons are added one at a time to the lowest energy level possible. (Fill lowest energy level first, then move to next higher)
2) Pauli Exclusion Principle - an orbital can hold a maximum of 2 electrons. Electrons within an orbital must spin in opposite directions. (Electrons must move in opposite directions, one up, one down.)

Principles of Orbital Diagrams

3) Hund's Rule - electrons occupy equal energy levels so that a maximum number of unpaired electrons result.

(Electrons fill in all up in a p, d, or f orbital, then a second in each orbital)

Exceptions to the Aufbau Principle

When the Aufbau Principle is applied to all elements, certain elements do not agree experimentally with Aufbau.

Due to subtle electron interactions, a new configuration is formed which allows the atom to be more stable.

The exception is applied when the configuration ends in any of the following ways: d^4, d^9, f^6, or f^{13}

Shortcuts for Configurations

It is allowed (in order to save time) to abbreviate the configuration by listing the noble gas prior to the element [in parenthesis], then finishing off the configuration from there.

Ex. Nb - [Kr] $5s^2 4d^3$