History of the Atomic Model

John Dalton (1766-1844): Atomic Theory

Atomic Theory
(1) All matter is made of atoms- particles which are too small to see
(2) Each element has its own kind of atom
(3) Atoms cannot be created or destroyed
(4) Molecules are formed when atoms are combined (compounds are formed when elements are combined)

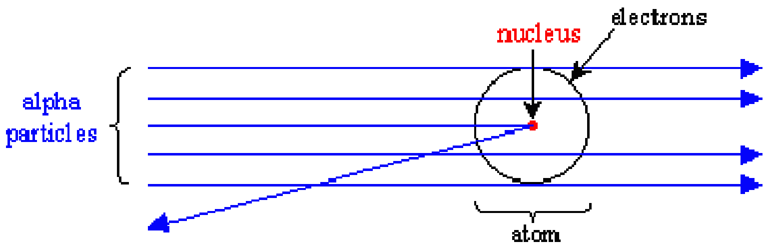
Joseph John Thomson (1856-1940): discovered the electron

"Plum Pudding Model"
An atom as lightweight negative particles (electrons) scattered throughout a heavy positive matrix (protons).

Ernest Rutherford (1871-1937): discovered the nucleus

"Nuclear Model"
An atom has a small positively charged nucleus consisting of protons which makes up the majority of the mass surrounded by a cloud of negatively charged electron electrons which take up most of the space.

Rutherford's Gold Foil Experiment
A layer of gold foil was bombarded by fast moving alpha particles (positively charged). It was expected that the alpha particles would pass through the foil. What was observed was that some of the alpha particles (1 in 8000) were actually sent backwards once they hit the foil. Rutherford proposed that the particles send backwards had hit a small, positively charged "nucleus".



James Chadwick (1891-1974): discovered the neutron, a neutral particle in the nucleus which contributes to the mass of an atom.

Properties of Subatomic Particles

Particle	Charge	Relative Mass	Location
Proton	+1	1	nucleus
Electron	-1	0.0001	orbiting nucleus
Neutron	0	1	nucleus

Atomic Configuration

8	-2
O	
Oxygen	
16.00	

Atomic Number: Different for each element. Gives the number of protons and for neutral (uncharged) atom, also gives the number of electrons.

$$Atomic\ mass = number\ of\ protons + number\ of\ neutrons$$

ex. complete the following table for **NEUTRAL** atoms

Atomic Number	Element	Symbol	Atomic Mass	Protons	Electrons	Neutrons
8						
	Calcium					
				30		

Isotopes: Different versions of atoms of the same element having the same number of protons but different numbers of neutrons (and therefore different atomic masses).

The atomic mass listed on the periodic table is actually the average atomic mass of all of the isotopes

$$Average\ Atomic\ Mass = (mass)(abundance) + (mass)(abundance)\ldots$$

ex. There are two isotopes of chlorine: chlorine-35 (75.8%) and chlorine-37 (24.2%)
(a) How many neutrons does each isotope have?

(b) What is the average atomic mass of chlorine?

Niels Bohr (1885-1962): discovered orbitals

"Orbital Model"
Electrons surrounding the nucleus are located in shells. Each shell contains a specific number of electrons and has a specific energy.

Shell	Number of Electrons
1st	2
2nd	8
3rd	8
4th	18

ex. Draw a Bohr Diagram for the following elements. Indicate the number of protons, electrons, and neutrons.

(a) O

(b) Cl

Elements in the periodic table are arranged into **groups** (vertical) and **rows** (horizontal)
Valence Electrons: Electrons in the outer shell. The number of valence electrons of an atoms is equal to the group number.
The number of orbitals of an atom is equal to the row number.

Ions: Charged atoms.
Atoms are most stable when they have a full valence shell (like the *noble gases*). Atoms can lose or gain electrons to establish a full valence shell. When electrons are lost, the atom becomes a positively charged **cation**. When electrons are gained, the atom becomes a negatively charged **anion**.

ex. Draw Bohr diagram of the atom and the ion for the following elements.

(a) S

S^{2-}

(b) Mg

Mg^{2+}

ex. complete the following table for **IONS**

Atomic Number	**Element**	**Symbol**	**Atomic Mass**	**Protons**	**Electrons**	**Neutrons**
3						
	Phosphorus					
			78.96			
		Sr^{2+}				

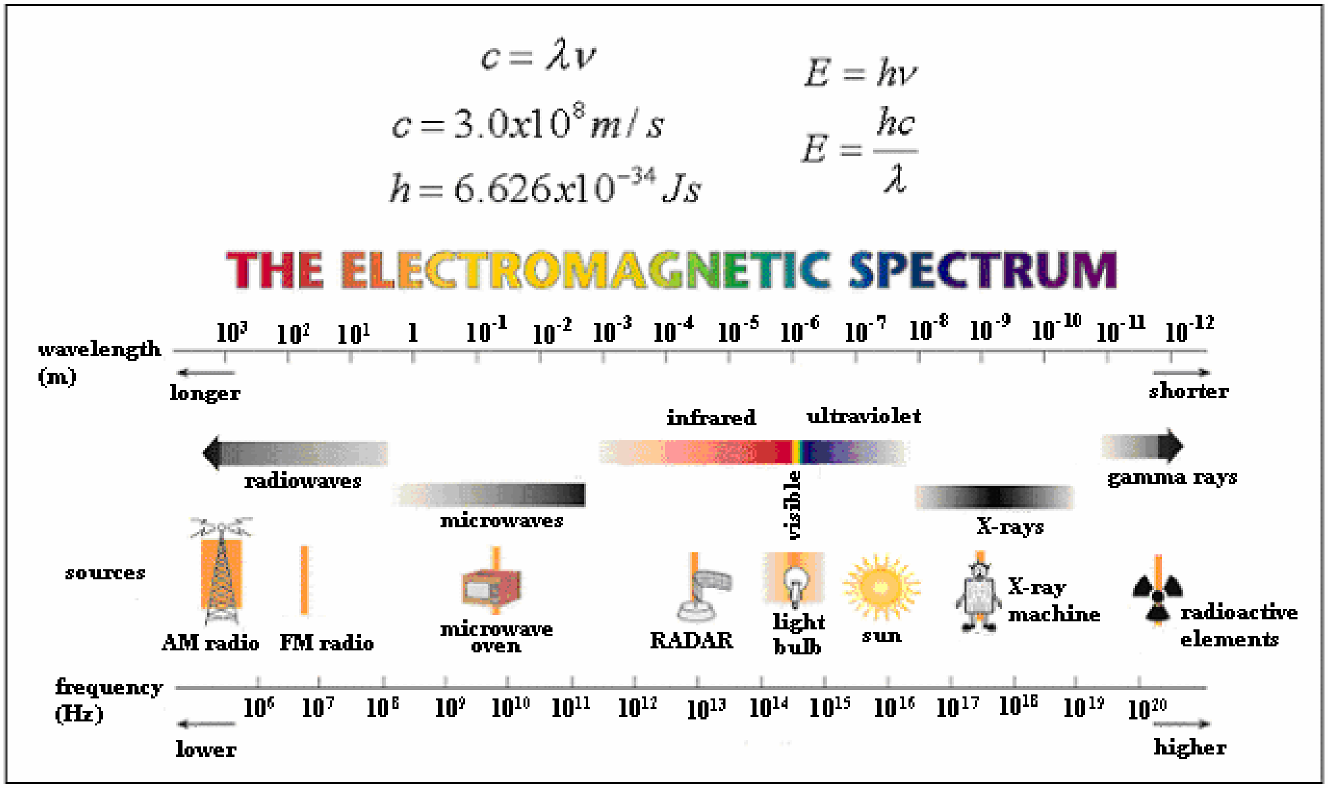
Properties of Light

(1) Light as a Wave

Light is an example of electromagnetic radiation.

Electromagnetic radiation: a form of energy that travels in waves at a speed of $3.0x10^8$ m/s (the speed if light, c).

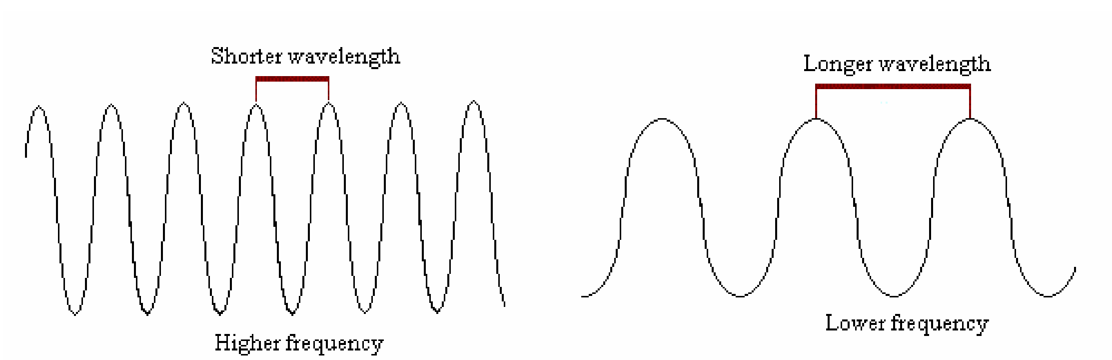
All forms of electromagnetic radiation are part of the electromagnetic spectrum.



Each type of electromagnetic radiation has a specific wavelength and frequency.

Wavelength (λ): the distance between peaks of a wave (measured in meters, m)

Frequency (v): the number of waves that pass a given point in a second (measured in Hertz, Hz)



$$c = \lambda \nu$$

$$c = 3.0x10^8 m/s$$
$$\lambda = wavelength\,(m)$$
$$\nu = frequency\,(Hz)$$

ex. A radio station broadcasts at a frequency of 106.7 MHz; what is the wavelength?

ex. An X-ray has a wavelength of 1.50 x 10^{-9} m; what is the frequency?

(2) Light at as Particle

Max Planck (1858-1947): the energy of electromagnetic radiation depends on the wavelength/frequency.

Albert Einstein (1879-1955): light is a stream of particles called photons. Photons have a specific energy depending on their wavelength/frequency. Photons can be thought of as massless "packets of energy".

The Photoelectric Effect

When light of a certain wavelength/frequency is shined on the surface of a given metal, the metal will emit electrons. The photons of light must have enough energy (high frequency/low wavelength) to "knock" an electron off the surface of the metal.

$$E = h\nu \qquad E = \frac{hc}{\lambda}$$

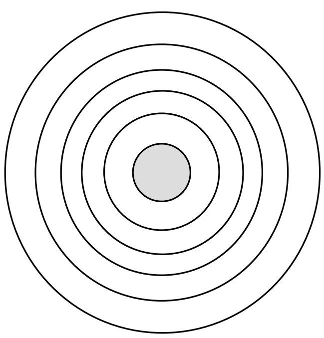
$E = Energy\,(J)$
$h = 6.626x10^{-34}\,Js$
$c = 3.0x10^{8}\,m/s$
$\lambda = wavelength\,(m)$
$\nu = frequency\,(Hz)$

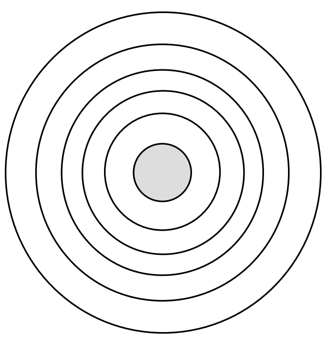
ex. A microwave has a wavelength of 2.45 cm; what is the energy?

ex. A gamma ray has a frequency of 9.4 x 10^{20} Hz; what is the energy?

Atomic Spectrum
Spectrum: a pattern of light given off by an atom.

When electricity is applied to an atom, it changes from being in its lowest energy or "ground" state to being in a higher energy "excited" state. After time, the atom will return to it its ground state (or a lower energy excited state) and give off light of a certain wavelength. The light is then sent through a prism and separated into its component colours to produce a line emission spectrum. The colours/wavelengths produced correspond to the energy differences between the orbitals of the atom. Each atom has a unique spectrum because the energy spacings of orbitals are different.



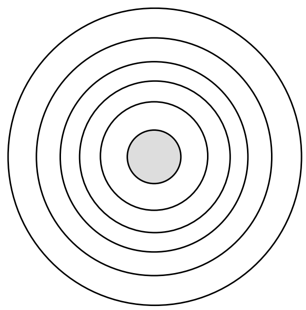


ex. In a hydrogen atom, the light emitted when an electron falls from the fourth energy level to the second energy level is 486 nm and the light emitted when an electron falls from the fifth energy level to the second energy level is 434 nm.

(a) Determine the frequency of each of these wavelengths.

(b) Determine the energy of each of these wavelengths.

(c) Draw a diagram to show the transitions taking place.



(d) What energy radiation would be emitted when an electron falls from the fifth energy level to the fourth energy level?

(e) What is the wavelength and frequency of this emitted radiation? Does this radiation fall into the visible range (380 nm – 740 nm)?

Bohr's model of the atom did a good job at explaining the spectra of smaller atoms but when the spectra of larger atoms were examined, they were found to be much more complex. This indicated that the orbitals were much more complex.

Quantum Mechanics

Werner Heisenberg (1901-1976) and Erwin Schrodinger (1887-1961): used mathematics to develop the nature of the orbitals. Orbitals are now thought to be three dimensional regions of space where electrons are located. Each orbital is defined by four quantum numbers.

Quantum Numbers
(1) Principal Quantum number, n
Gives the main energy level of an orbital and indicates how close the orbital is to the nucleus. Higher values of n indicate farther orbitals which are higher in energy.

(2) Angular Momentum Quantum number, l
Gives the shape of the orbital.

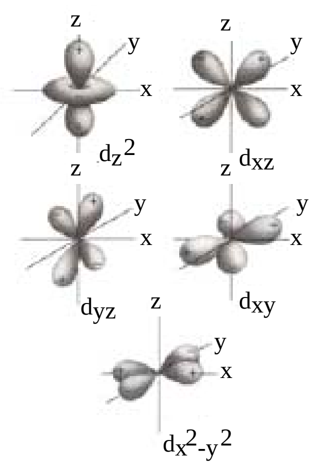
l	**Type of Orbital**
$l = 0$	s
$l = 1$	p
$l = 2$	d
$l = 3$	f

(3) Magnetic Quantum Number, m_l
Gives the orientation of the orbital.

(4) Spin Quantum Number, Ms

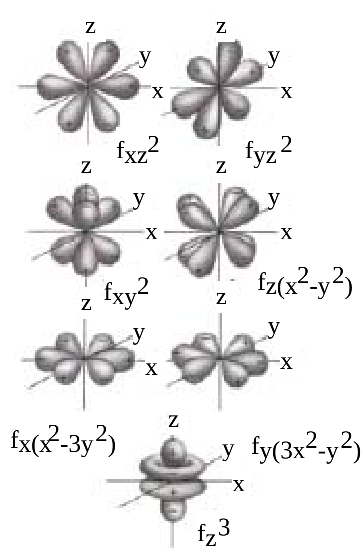
Gives the "spin" of an electron. Electrons can have a spin of $+\frac{1}{2}$ or $-\frac{1}{2}$. Each orbital can hold two electrons of opposite spin.

D Orbitals

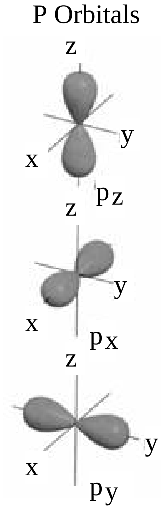


S Orbital

F Orbitals



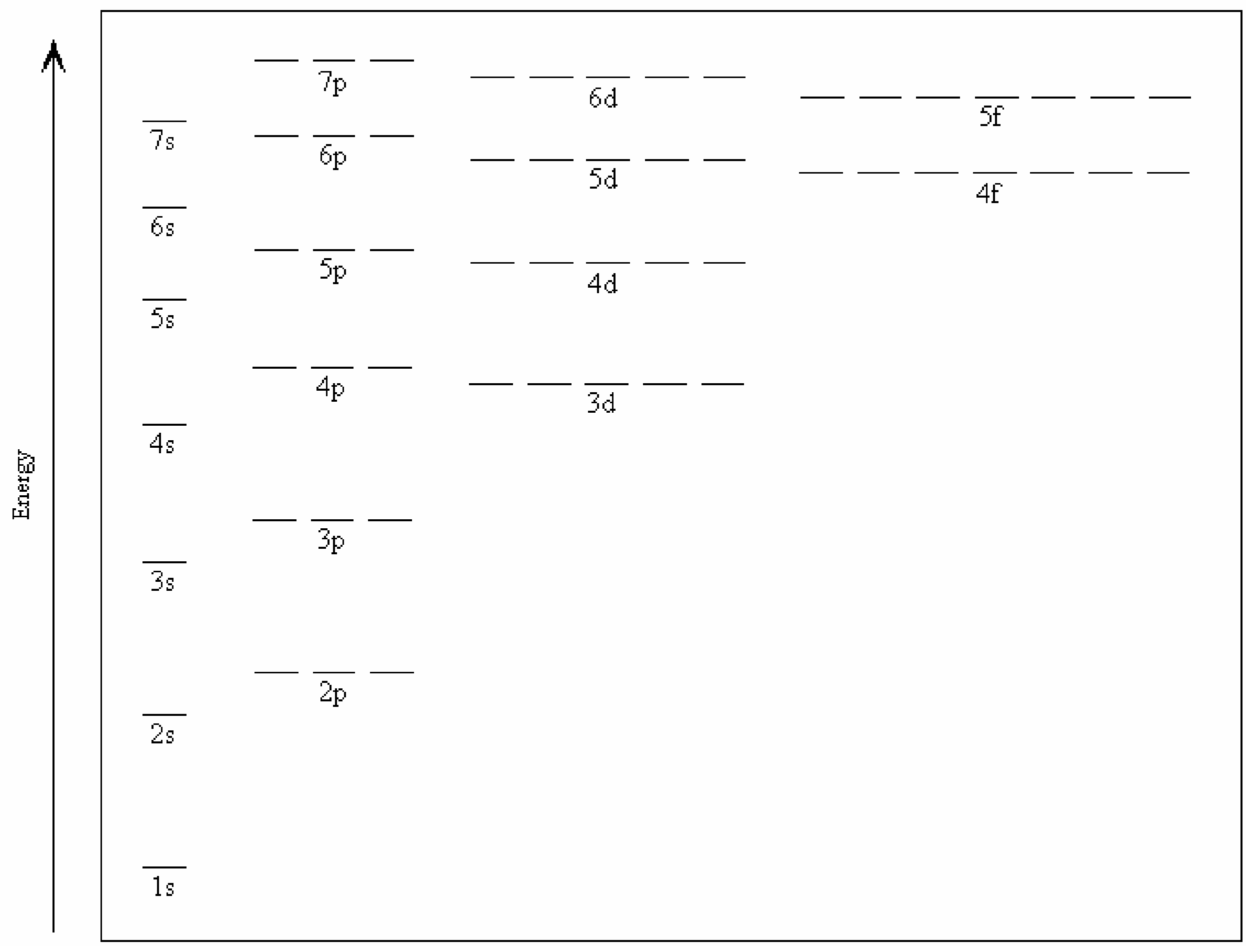
P Orbitals



Orbital Notation

(1) **Aufbau Principal**: an electron will occupy the lowest energy orbital available.
(2) **Pauli Exclusion Principal**: no two electrons in an atom can have the same set of quantum numbers.
(3) **Hund's Rule**: orbitals of equal energy are each occupied by one electron, with each electron having the same spin, before they are paired with electron having opposite spin.

Energy Level Diagram



Excited States
In an excited state, one or more electrons move to a higher energy orbital.
There are many possible excited states for a given atom.

ex. ground state of carbon: $1s^2\ 2s^2\ 2p^2$
possible excited states of carbon:
$1s^2\ 2s^2\ 2p^1\ 3s^1$ or $1s^2\ 2s^2\ 2p^1\ 3p^1$

Noble Gas Notation
Gives the noble gas that comes before the element and electrons that follow the noble gas.

ex. Write the noble gas notation for oxygen

ex. Write the noble gas notation for calcium

Ions
Electrons are gained (anions) or lost (cations) to form a charged ion.

Write the electron configuration notation and the noble gas notation for the following ions.

(1) P^{3-}

(2) Mg^{2+}

(3) Fe^{3+}