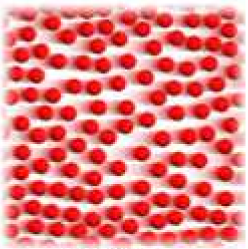
Light (Particle or Wave)?

Sir Isaac Newton (1642-1727)

- Corpuscular Theory of Light
- Classical Laws of Mechanics (3 Laws of Motion)
- “Greatest and most influential scientist to ever live”
- Einstein kept a framed picture of him

"Light is never known to follow crooked passages nor to bend into the shadow"



Particles

James Clerk Maxwell (1831-1879)

- Set of equations to describe Electricity, Magnetism and Optics.
- Kinetic Theory of Gases (Maxwell-Boltzmann Distribution)
- 3rd Greatest Scientist of All Time

Thomas Young (1773-1829)

- Double Slit Experiment
- Constructive and Destructive Inference
- Notable scientific contributions to vision, light, solid mechanics, energy, physiology, language, musical harmony, and Egyptology.

Young's Double Slit Experiment

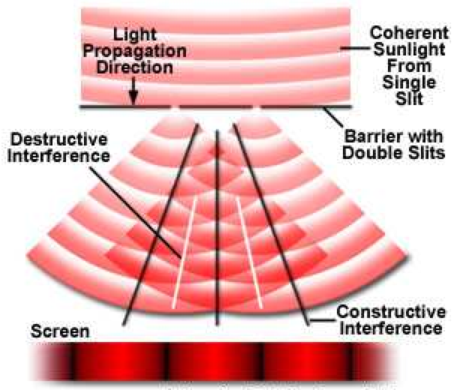


Figure 6

Particles and Waves Through Crossed Polarizers

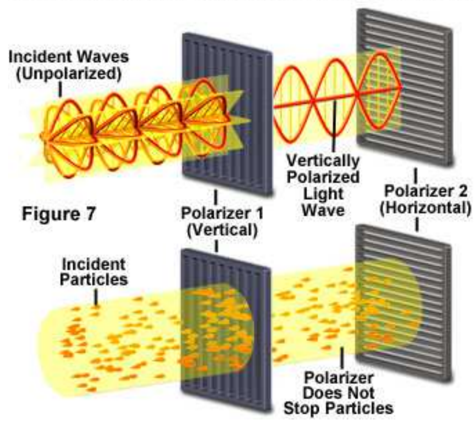
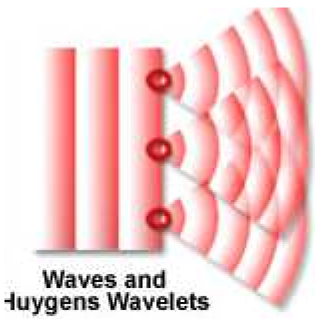


Figure 7

Christiaan Hugens (1629-1695)

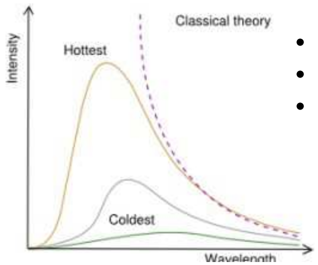
- Wave Theory of Light
- Invented Pendulum Clock,
- Probability Theory

Christiaan Huygens (1629-1695)



Waves and Huygens Wavelets

Planck
Black Body Radiation



- Classical Model Fails (UV Catastrophe)
- Energy (Light) is quantized
- No reason why its quantized

Einstein
Photoelectric Effect

- Classical Model Fails (Blue and Red light should eject same number of electrons if they have the same intensity (wavelength and frequency should have no effect)
- Energy (Light) is quantized in order to explain why a threshold exists
- No reason why its quantized

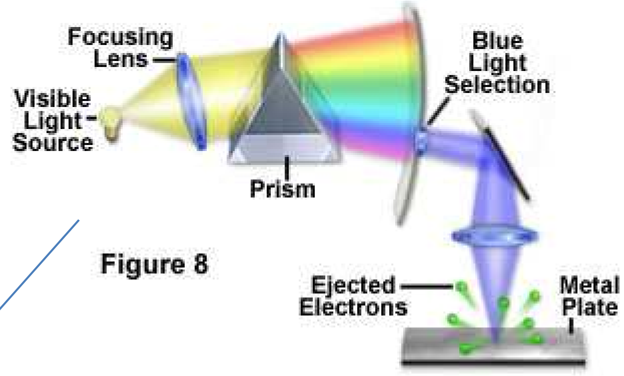
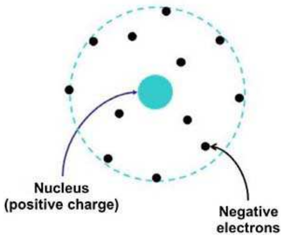


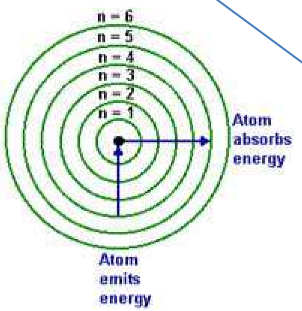
Figure 8

Rutherford Model



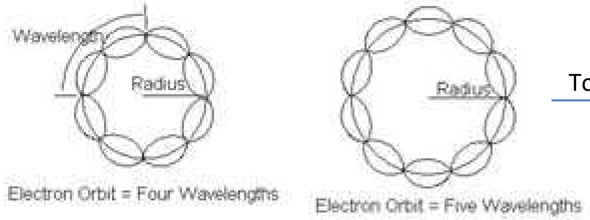
- Electrons are allowed to be anywhere
- Predicts continuous or rainbow like spectrum
- Flaw: Fails to explain Line Spectra

Bohr Model



- Electrons are in quantized orbitals
- Predicts line spectrum
- Flaws (1) No reason its quantize
  (2) Does not work for > 1 e

De Broglie Model



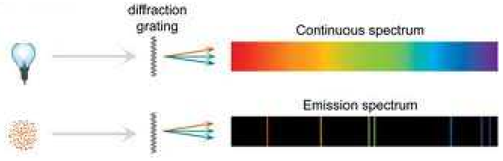
To Heisenberg

- If light can behave like a wave and a particle than an electron can behave like a particle and a wave.
- Explained Quantization of Electron Orbitals therefore fixing Flaw (1).
- Model still fails for > 1 e

Line Spectra

Normal Light
Rutherford Model

Actual Experiment
Pure Elements



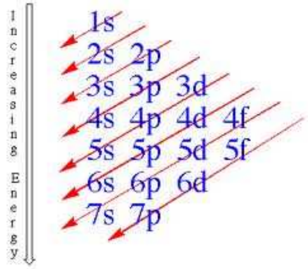
Pauli Exclusion Principle

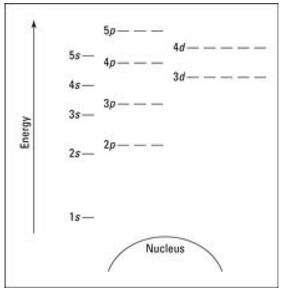
- No two electrons in an atom can have the same 4 QN's

Hund's Rule

- Two or more orbitals with the same energy, electrons will fill each until ½ full.
- Electrons in partially filled orbitals have the same spin QN (spin aligned.

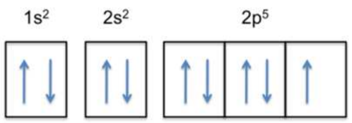
Energy Level Filling Diagrams





Electron Configurations

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$



Heisenberg Uncertainty Principle

From De Broglie

- It is impossible to know precisely where an electron is and what path it follows.

$$(\Delta x)(\Delta mv) \geq \frac{h}{4\pi}$$

- Uncertainty in the position of the electron is larger than the size of the atom → orbitals can't exist
- No model

Schrodinger Equation

- Statistical model using wave equations to describe the location of the electron.

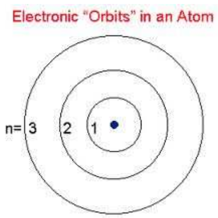
$$H\varphi = E\varphi$$

- Does not violate Heisenberg
- 4 Principle Quantum Numbers result

To Periodic Trends

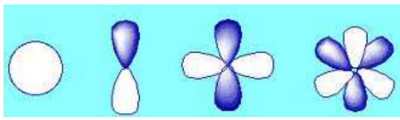
Principle Quantum Number (n)

- 1, 2, 3 … n (n = integer)
- Distance from the nuclei



Angular Momentum Quantum Number (ℓ)

- 0, 1, 2, 3 … (n-1)
- 3D - shape of an orbital



$\ell = 0$ s - orbital

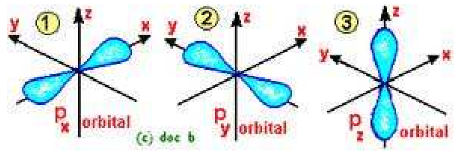
$\ell = 1$ p - orbital

$\ell = 2$ d - orbital

$\ell = 3$ f - orbital

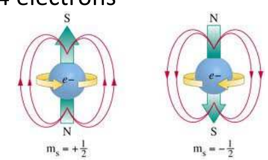
Magnetic Quantum Number (m_ℓ)

- $0, \pm 1, \pm 2 \;\ldots\; \pm \ell$
- Orientation of the orbital
- s = 1 orbital
- p = 3 orbitals
- d = 5 orbitals
- f = 7 orbitals

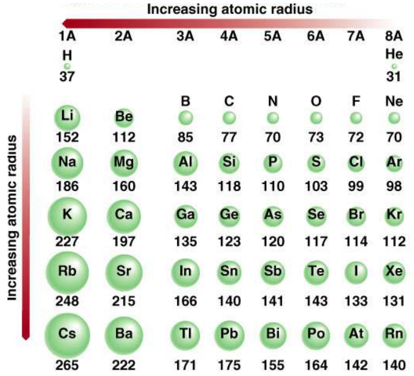


Electron Spin Quantum Number (m_s)

- $\mathcal{M}_s = \pm {}^1/_2$
- Orbitals can hold 2 electrons ∴
- s = 2 electrons
- p = 6 electrons
- d = 10 electrons
- f = 14 electrons

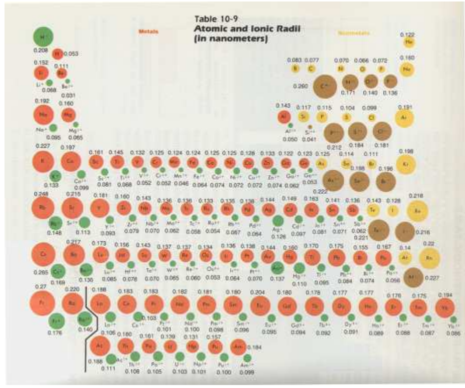


Atomic Radius



- Columns: Each step down is a new orbital (bigger n)
- Rows: $Z_{eff} = Z_{actual} - E.S.$
  - Z_{actual} ↑ *across a row*
  - $E.S.$ ~ same across a row
  - ∴ Z_{eff} ↑ and atom shrinks

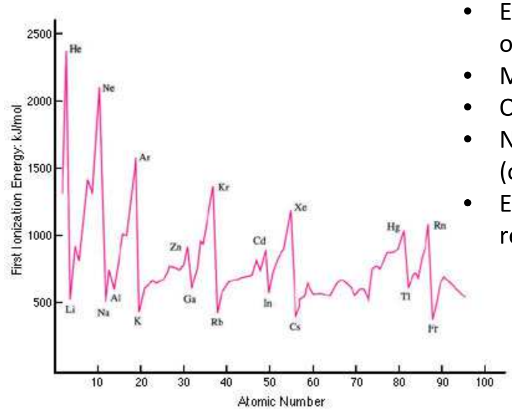
Ionic Radius



Cations are smaller: p/e ratio ↑ ∴ radius ↓

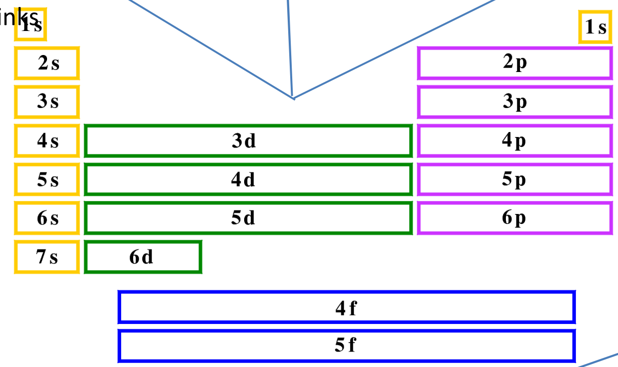
Anions are bigger: p/e ratio ↓ ∴ radius ↑

Ionization Energy



- Energy required to remove the outermost electron from an atom.
- M + IE → $M^+ + e^-$
- Opposite trends as atomic radius
- Noble Gases are very stable (octets)
- Extra large amount of energy to remove an e^- from a noble gas

From Schrodinger Equations



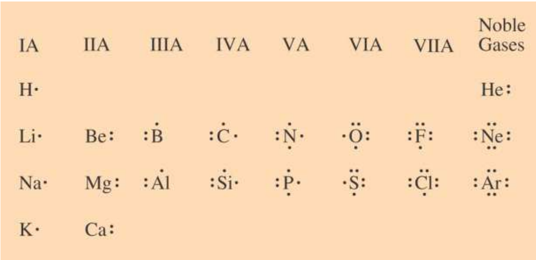
Valence Electrons

- Electrons in the outermost orbital (n)of an atom.
- s and p electrons with highest n
- Responsible for the formation of ions and molecules

To Lewis Structures

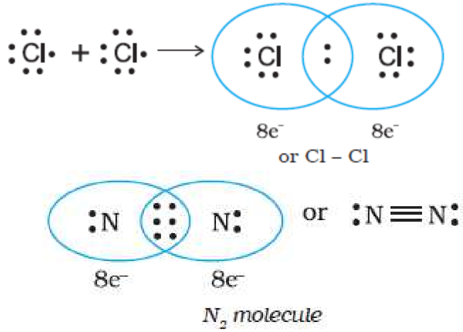
Ionic Compounds

- Complete octets by gaining or losing electrons



Molecular Compounds

- Complete octets by sharing electrons



From Valence Electrons

Lewis Dot Structures

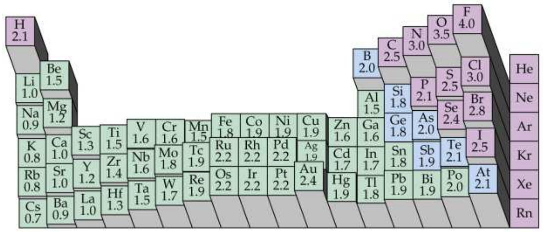
- Represents bonding between ionic and molecular compounds in terms of valence electrons and the octet rule

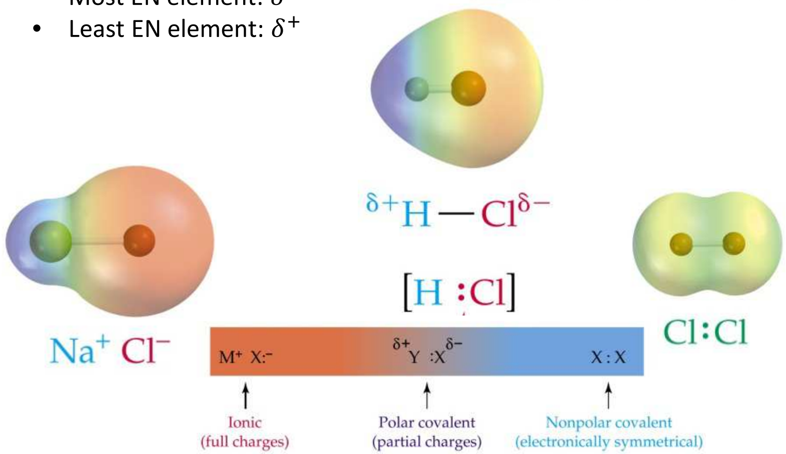
VSEPR Theory

- Valence Shell Electron Pair Repulsion Theory
- Shape of molecules is determined by the repulsive forces between electrons. Molecules minimize this by keeping the electrons as far apart as possible.

Electronegativity

- Attractive force that an atom has for shared electrons in a covalent bond.
- Most EN element: δ^-
- Least EN element: δ^+





Electron Groups Total	Electron Groups Bonding	Electron Groups Lone	Arrangement of Groups	Molecular Shape	Example
2	2	0	Linear	Linear	CO_2
3	3	0	Trigonal planar	Trigonal planar	NO_3^-
	2	1		Bent (or angular)	O_3 (Lone pair)
4	4	0	Tetrahedral	Tetrahedral	CH_4
	3	1		Trigonal pyramidal	PF_3 (Lone pair)
	2	2		Bent (or angular)	H_2O (Lone pairs)