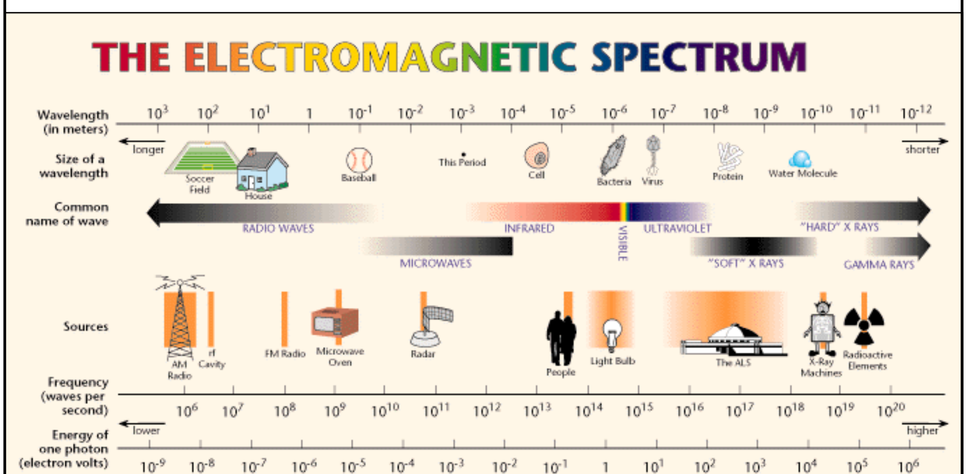
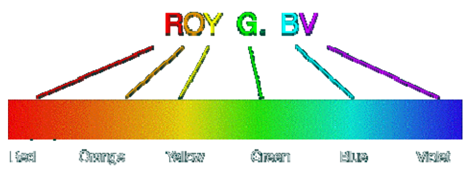
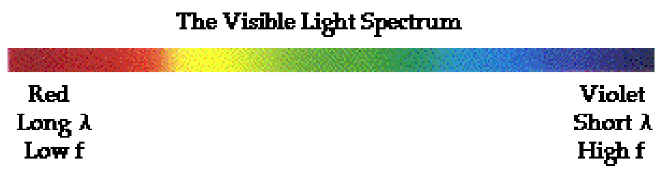
Atomic Structure

Electromagnetic Spectrum



Visible Spectrum





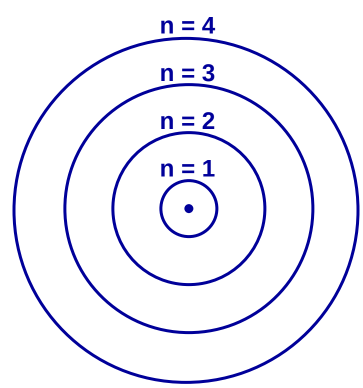
Line Spectrum

- Each element has its own unique line spectrum
- Each line is caused by an electron dropping from one energy level to another
- The color (frequency, wavelength, energy) of each line in the spectrum is determined by the difference in energy levels

Models of the Atom

Bohr's Model

- Electrons move around the nucleus in only certain allowed circular orbits
- Each orbit has a certain quantum number (energy number) associated with it.



- Energy levels could be calculated with a simple calculation using a constant determined by Rydberg
- Bohr's calculated values for the energy levels of Hydrogen matched experimental results
- Bohr was not able to explain why there should only be specific energy levels
- deBroglie proposed that maybe the electrons behaved like waves
- Davisson, Germer, & Thomson demonstrated this

However….

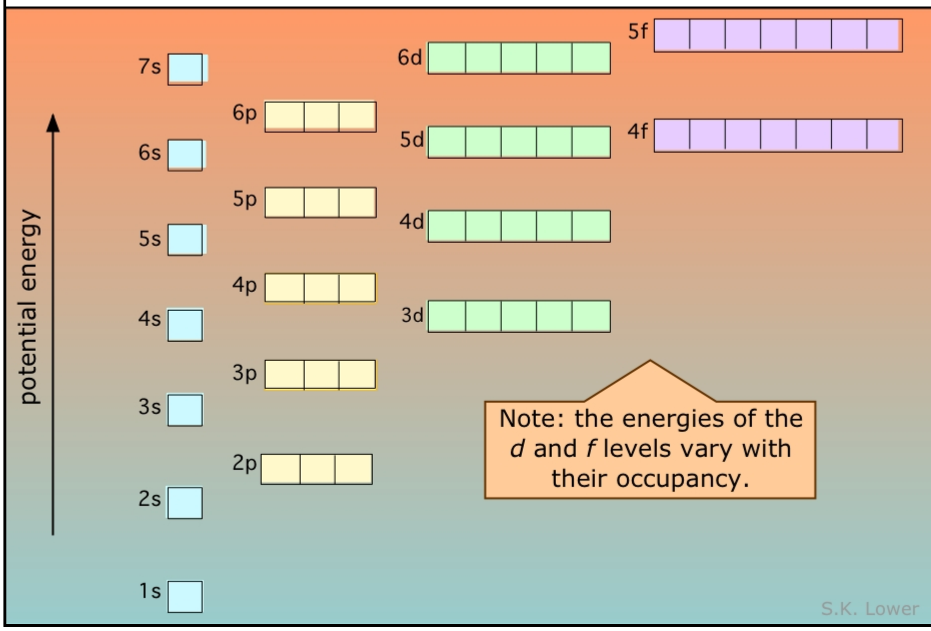
- Bohr's model only worked for the Hydrogen atom
- Schrödinger developed a new theory where the electrons are not in orbits but in probability regions known as orbitals

Orbitals

- An orbital is a mathematical function describing the standing wave that gives the probability of the electron manifesting itself at any given location in space.
- More commonly (and loosely) we use the word to describe the region of space occupied by an electron.

- Each kind of orbital is characterized by a set of quantum numbers n, l, and m.
- These relate, respectively, to the average distance of the electron from the nucleus, to the shape of the orbital, and to its orientation in space.

Diagram of Energy Levels



Aufbau Principle

- Each electron occupies the lowest energy orbital available
  - For example, all of the “2” orbits must be filled before electrons can go into the 3s orbital

Pauli Exclusion Principle

- A maximum of two electrons may occupy a single atomic orbital, BUT only if the electrons have opposite spins
  - The atomic orbital containing two electrons with opposite spins is written as ↑↓

Hund's Rule

- Single electrons with the same spin must occupy each equal-energy sublevel before additional electrons (with opposite spins) can occupy the same orbitals
  - For example, the "p" orbital must be filled in this order

1.	2.	3.
[↑] [ ] [ ]	[↑] [↑] [ ]	[↑] [↑] [↑]
4.	**5.**	**6.**
[↑↓] [↑] [↑]	[↑↓] [↑↓] [↑]	[↑↓] [↑↓] [↑↓]

Examples

- H
  - $1s^1$
- He
  - $1s^2$
- Li
  - $1s^2 2s^1$
- C
  - $1s^2 2s^2 2p^2$

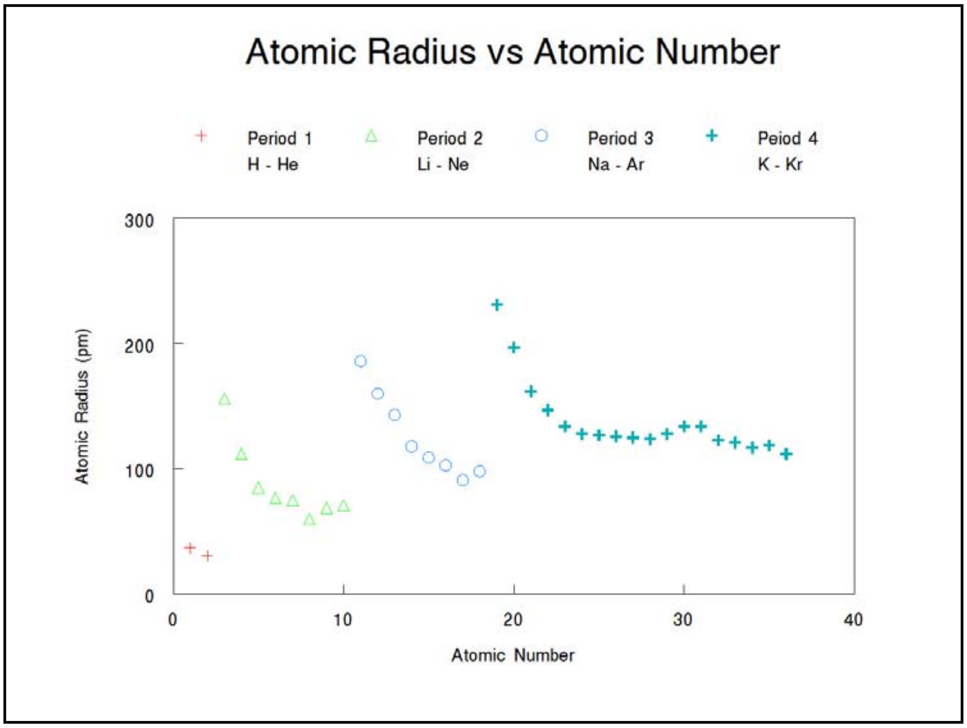
Noble Gas Notation

- Ne
  - $1s^2 2s^2 2p^6$
- Na
  - $1s^2 2s^2 2p^6 3s^1$
  - $[Ne]3s^1$

Periodic Trends

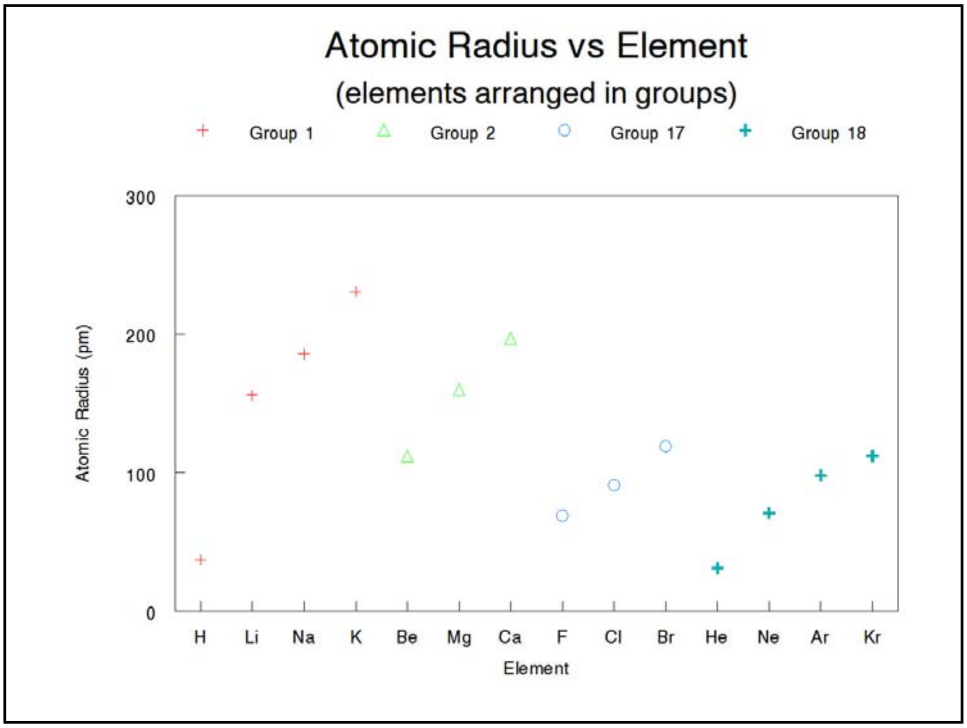
- Atomic radii
- Ionization energy
- Ionic radii
- Electronegativity

Trends in Atomic Radii



Periodic Trends in Atomic Radii

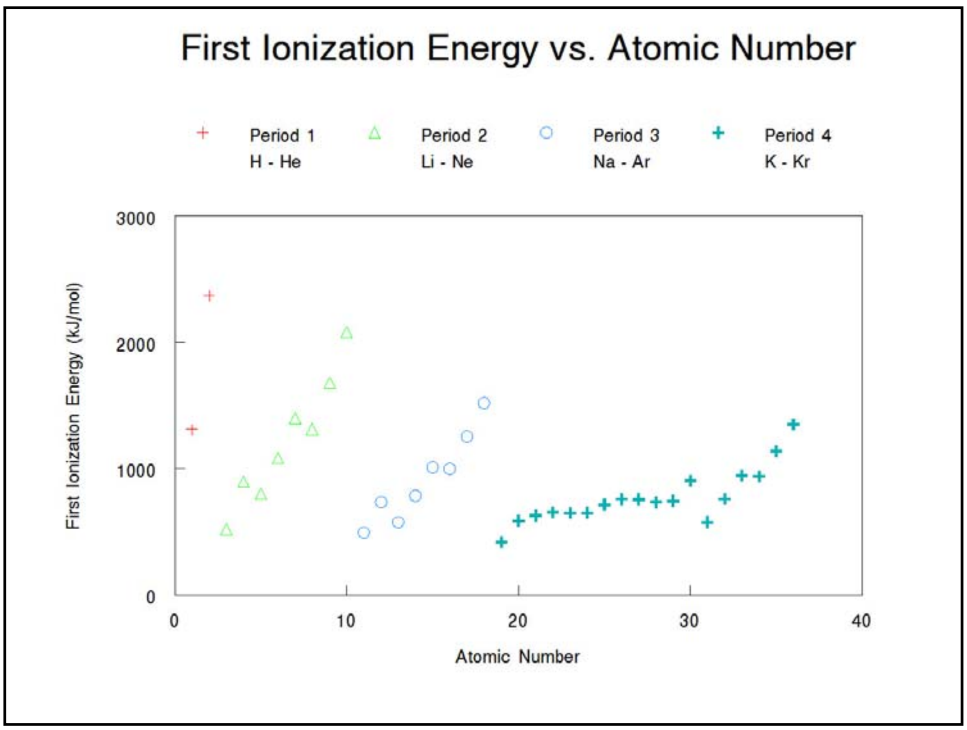
- The atomic radii generally decrease as you move across a period
  - Since each additional electron is added to the same principal energy level, the additional electrons are not shielded from the increasingly positive nucleus.
  - The increased nuclear charge pulls the valence electrons closer to the nucleus reducing the atomic radius.



Group Trends in Atomic Radii

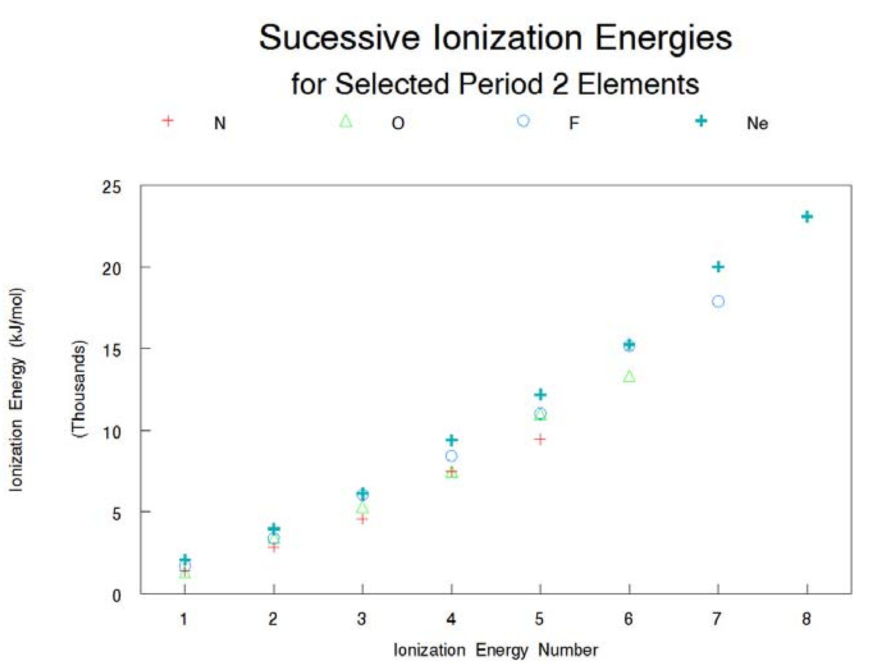
- The atomic radii generally increase as you move down a group.
  - As you move down a group the outermost orbital increases in size shielding the valence electrons from the pull of the nucleus.

Trends in Ionization Energy



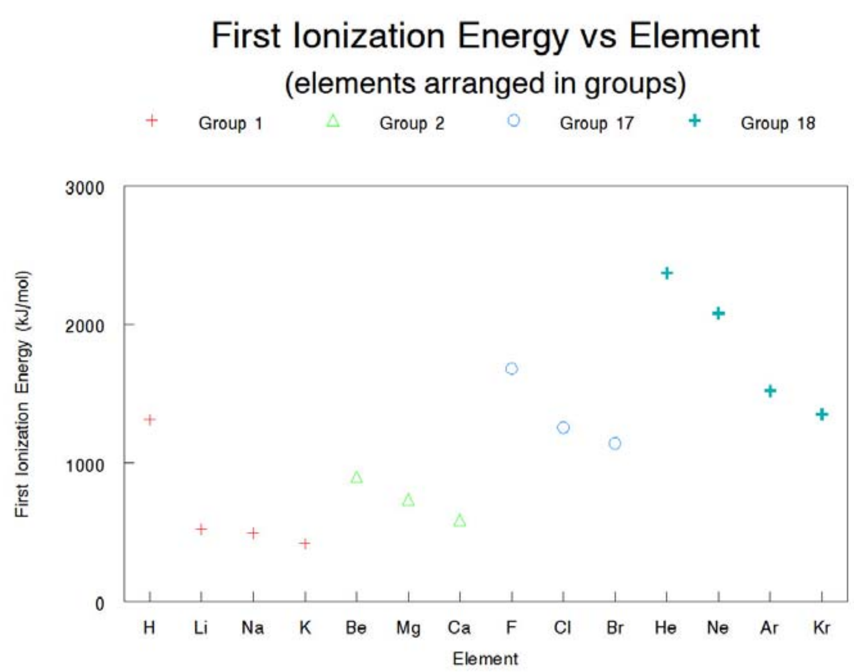
Periodic Trends in First Ionization Energies

- As you move across a period, the first ionization energy generally increases.
  - As you move across the row it becomes increasingly harder to remove a valence electron from the atom.
  - the increased nuclear charge of each successive element produces an increased hold on the valence electrons thereby increasing the ionization energies



Period Trends in Successive Ionization Energies

- The energy required for each successive ionization energy increases as you move across a period
  - The primary reason for this is that the increase in positive charge binds the electrons more strongly



Group Trends in Ionization Energies

- The ionization energies decrease as you move down a group
  - The increasing atomic size pushes the valence electrons further away from the nucleus
  - Consequently it takes less energy to remove the electron because the strength of attraction is less

Trends in Ionic Radii

- Atoms can gain or lose one or more electrons to form ions
- When atoms lose electrons and form positively charged ions, they always become smaller
  - The lost electron will almost always be a valence electron, leaving an empty outer orbital resulting in a smaller radius
  - The electrostatic repulsion between the now fewer remaining electrons is smaller, allowing them to be pulled closer to the nucleus

- When atoms gain electrons and form negatively charged ions, they always become larger
  - The addition of an electron to an atom increases the electrostatic repulsion between the atom’s outer electrons, forcing them to move farther apart

Periodic Trends in Ionic Radii

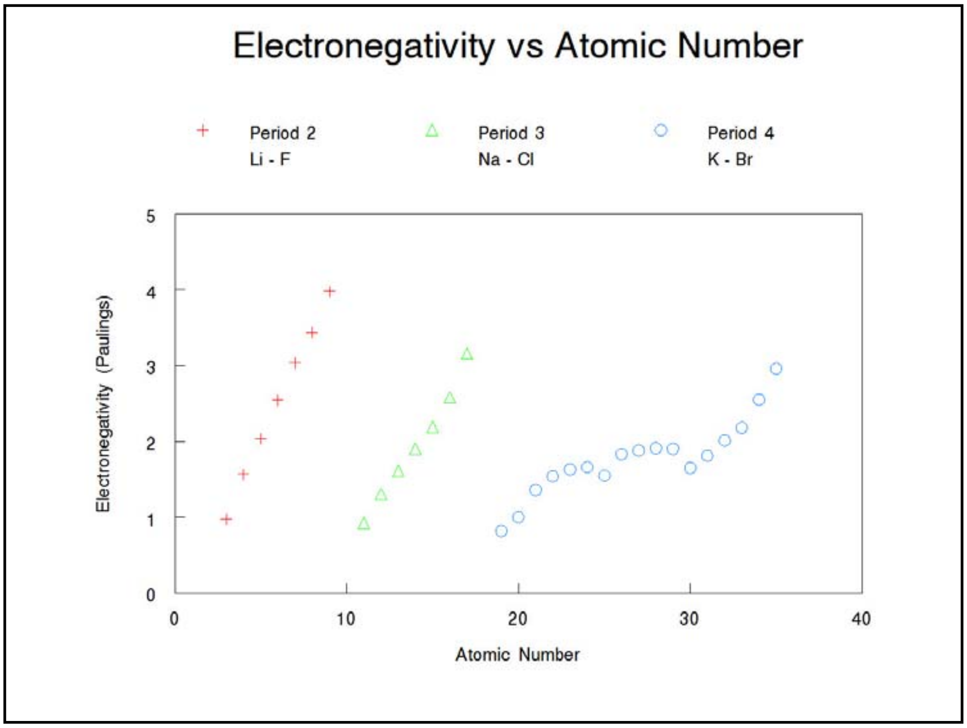
- The size of positive ions decrease as you move left to right across a period
  - For example: Li^{+}, Be^{2+}, B^{3+}, C^{4+}
- The size of negative ions decrease as you move left to right across a period
  - For example: N^{3-}, O^{2-}, F^{-}

Group Trends in Ionic Radii

- The ionic radii of all ions increases as you move down a group
  - As you move down a group, an ion’s outer electrons are in higher principal energy levels, resulting in a gradual increase in ionic size

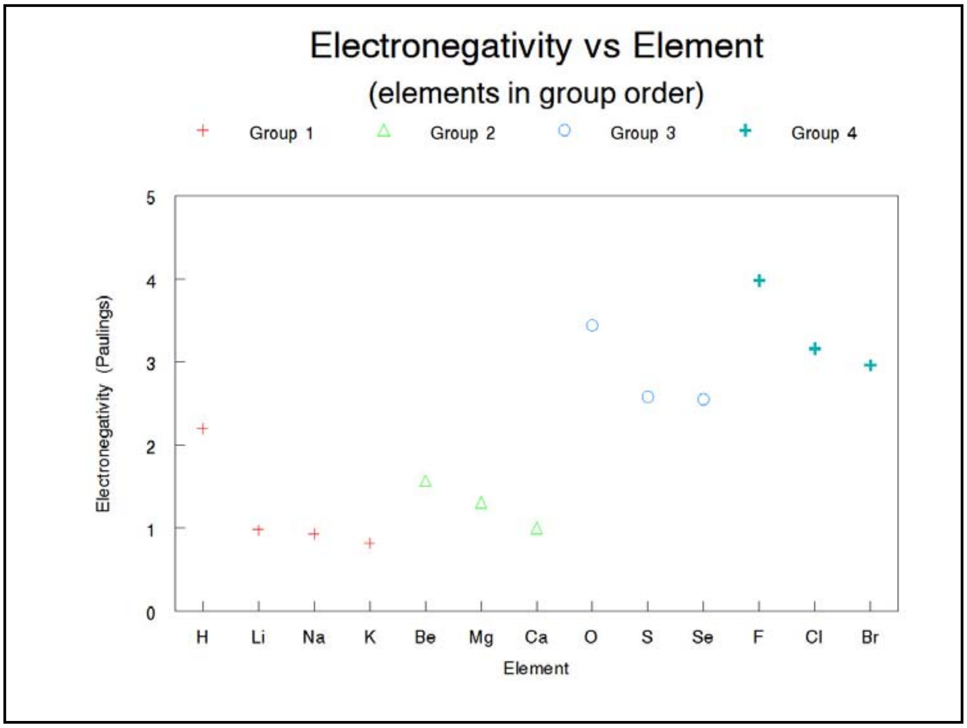
Trends in Electronegativity

- The electronegativity of an element indicates the relative ability of its atoms to attract electrons in a chemical bond
- The values are calculated based on a number of factors and expressed as a value of 3.98 or less
- Electronegativity is measured in arbitrary units called Paulings



Periodic Trends in Electronegativity

- Electronegativity generally increases as you move left-to-right across a period



Group Trends in Electronegativity

- Electronegativity generally decreases as you move down a group

Another Note about Electronegativity

- Because electronegativity is a measure of the ability of an atom to attract electrons, it can be used to predict the type of bond that will form between two atoms
- The type of bond can be predicted by the difference in electronegativity between the two elements

Electronegativity Differences and Predicted Bond Character

Electronegativity Difference	Predicted Bond Type	Example
0 – 0.4	Non-polar covalent	O_2 (3.44-3.44=0)
0.4 – 1.0	Moderately polar covalent	SCl_2 (3.16-2.58=0.58)
1.0 – 2.0	Very polar covalent	CaS (2.58-1.00=1.58)
≥2.0	Ionic	KCl (3.16-0.82=2.34)