Periodic Trends

		Direction of Increase	Left to Right Across Period	Down Group
A	Atomic Radii	← ↓	decreases	increases
I	Ionization Energy	↑ →	increases	decreases
M	Metallic Properties	← ↓	decreases	increases
E	Electronegativity	↑ →	increases	decreases
S	Shielding	↓	stays same	increases

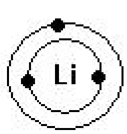
Forces Determining Trends:

- **Electromagnetic Force**: the positive protons inside the nucleus pull the negative electrons outside the nucleus in toward the center because of Coulomb's Law that “opposite attract”; this is stronger than electron-electron repulsion

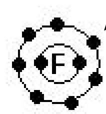
- **Electron-Electron Repulsion**: the electrons in the valence orbital and core orbitals will repel each other and try to spread out because Coulomb's Law says “likes repel”; this is weaker than the nuclear force

Atomic Radii – involves the size of the nucleus and electron cloud around the atom where the number of protons inside the atom equals the number of electrons outside the atom

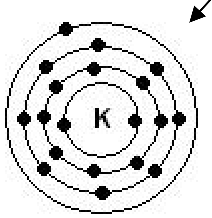
Lithium has 3 protons pulling in 3 electrons and only 2 energy levels.



Fluorine has 9 protons pulling in 9 electrons and only 2 energy levels.

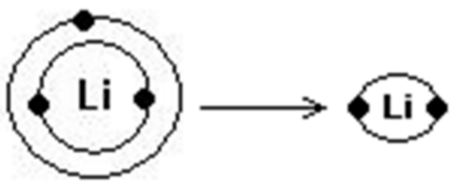
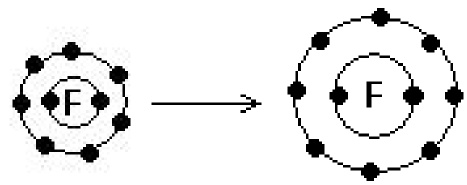


Potassium has 19 protons pulling in 19 electrons and 4 energy levels.

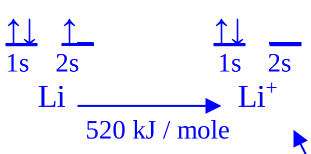


The nuclear force is pulling Fluorine’s electrons in closer because there are 9 protons in the nucleus compared with only 3 for Lithium. Potassium is bigger than Lithium and Fluorine because it is in a higher energy level and has more orbitals.

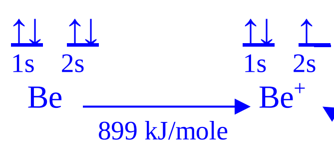
Ionic Radii - the size of the nucleus and electron cloud around the atom where the number of protons inside the atom DOES NOT equal the number of electrons outside the atom; the size pattern for Group A1 – A8 is Large, Medium, Small, Equal, Large, Medium, Small, N/A

Cation	**Anion**
Positive ion; electron is removed	Negative ion; electron is gained
	
Li → Li^{+}	F → F^{-}
Lithium atom has 3 electrons and 3 protons; Lithium ion has 2 electrons and 3 protons	Fluorine atom has 9 electrons and 9 protons; Fluorine ion has 10 electrons and 9 protons
Cations get smaller because there are more protons pulling inward on fewer electrons.	Anions get larger because there are less protons pulling inward on more electrons.
The nuclear force is constant and the electron-electron repulsion force is decreased	The nuclear force is constant and the electron-electron repulsion force is increased

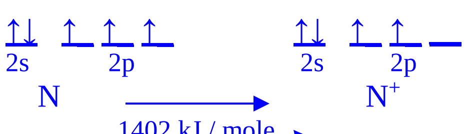
Ionization Energy – the energy required to ***remove*** an electron from an atom in the gas phase; because energy is required to ionize an atom, the process will always be endothermic (+); also removing an electron always makes the atom positive forming a ***cation***



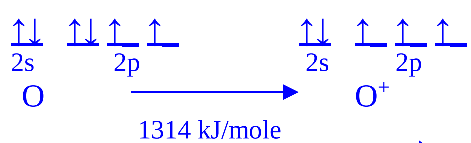
Li^{+} has a noble gas configuration, which is more stable. Therefore, not much energy is required to remove the $2s^1$ electron.



Remember Be is a Group 2 element, so it wants to have a charge of 2+. So when only 1 electron is removed, the ion is not stable and requires a lot more energy.



Notice Nitrogen forms a N^{+} cation, not a N^{-3} anion. The configuration for N^{+} is becoming more unstable so the energy required is higher.



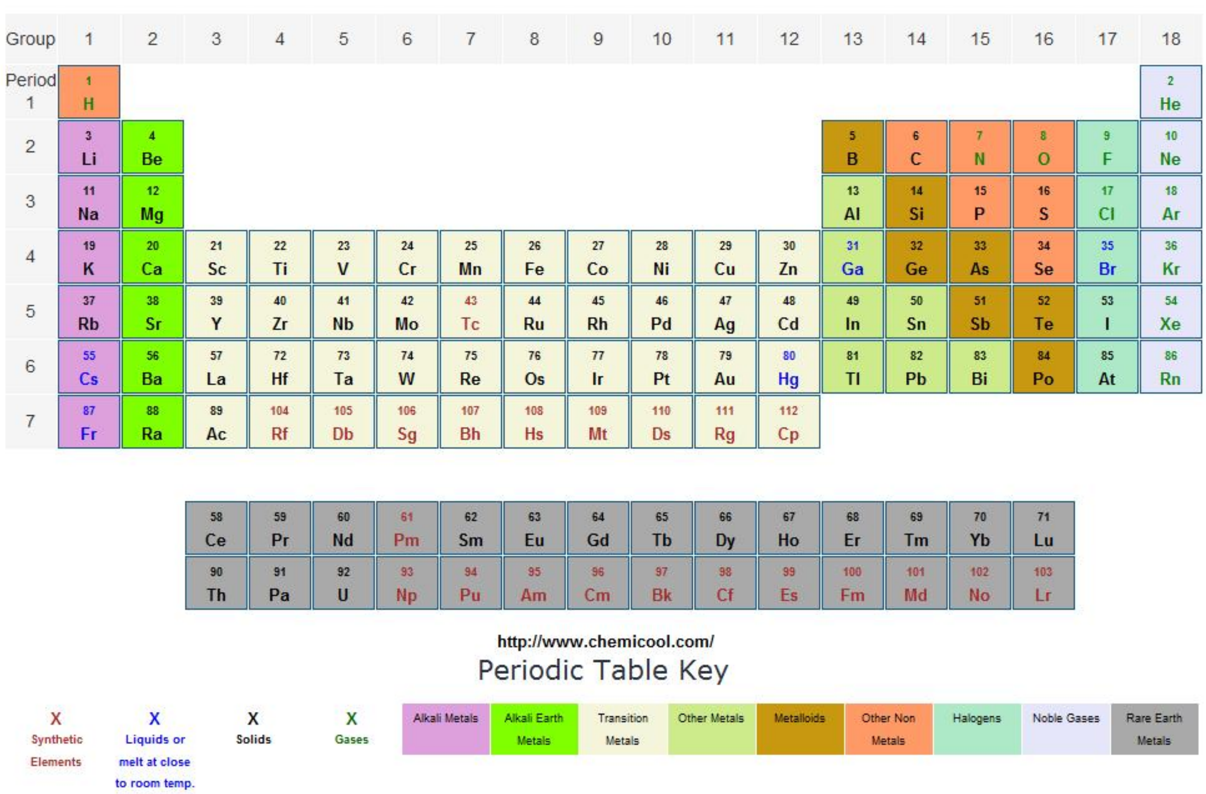
The electron configuration for O^{+} is actually more stable than for N^{+} because all orbitals have at least one electron; therefore, the energy is slightly less.

	Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7	Group 8
n = 2	**Li+** 520 kJ/mol	**Be+** 899 kJ/mol	**B+** 801 kJ/mol	**C+** 1086 kJ/mol	**N+** 1402 kJ/mol	**O+** 1314 kJ/mol	**F+** 1681 kJ/mol	**Ne+** 2081 kJ/mol
n = 3	**Na^+** 496 kJ/mol							
n = 4	**K^+** 419 kJ/mol							
n = 5	**Rb^+** 403 kJ/mol							
n = 6	**Cs^+** 377 kJ/mol							

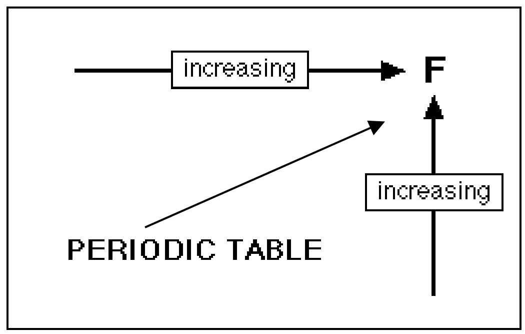
For Group 3 and Group 6 elements, more stable electron configurations are formed upon ionization; therefore, a slight dip in energy is seen in the upward trend.

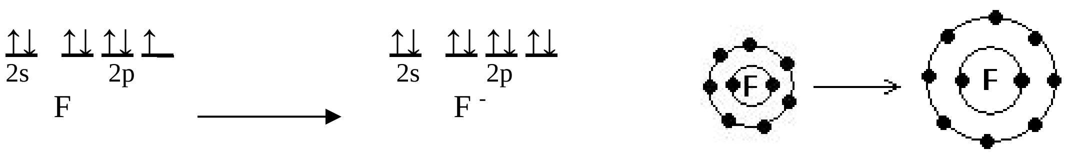
For periods n = 4 and higher, the energy difference within each group is noticed less and less. This is because the inner orbitals shield the outer valence level causing the nuclear force to be less intense. Therefore, it is easier and easier to remove an electron from the valence orbital as you go down a group.

Metallic Properties – elements characterized by tendency to give up electrons and by good thermal and electrical conductivity; as you down a group, a metals reactivity increases therefore increasing its metallic property; Group 1 = Alkali metals, Group 2 = Alkaline Earth metals, Group 3 through Group 12 = Transition metals; the right side of the periodic table is divided into metals, metalloids, and non-metals.



Electronegativity – a measure of the tendency of an atom to ***attract*** a bonding pair of electrons; gaining an electron always forms an *anion*; according to the Pauling scale, Fluorine (the most electronegative element) is assigned a value of 4.0, and values range down to cesium and francium which are the least electronegative at 0.7; if two atoms bond together, the difference in their electronegativities determines the type of bond (non-polar covalent < 0.4, polar covalent 0.5 – 1.6, and ionic > 1.6)





Fluorine is the most electronegative atom because its configuration becomes a complete octet when the electron is obtained and it has the least amount of shielding between the nucleus and outer valence orbital. The table above shows the trend of increasing Electronegativity.

Note: **Electron Affinity** is the energy required for the Electronegativity process in which the electron is assigned to the lowest energy unoccupied valence orbital of a gaseous atom. When a stable anion is formed, energy is released (exothermic; -). When an atom has no affinity for an electron (the configuration is already complete), an unstable anion is formed causing energy to be absorbed (endothermic; +). Electron Affinity is therefore positive for Groups 2 and 18.

Shielding – the inner orbitals between the nucleus and the valence orbital; as the energy level gets higher, shielding increases because there are more inner orbitals; shielding directly effects both the nuclear force and the electron-electron repulsion force and therefore affects all the previously mentioned trends

- **Nuclear Force** – as shielding increases, the positive charges inside the nucleus cannot pull as tightly on the outer valence orbital electrons, therefore, decreasing the nuclear force
- **Electron-Electron Repulsion** – as shielding increases, the inner orbitals block the nuclear force allowing electron-electron repulsion to repel electrons further away from each other.