SCH 4U
J. Kropac

Explaining Periodic Trends

- In every trend we look at, there are two factors that are important to understand:

Shielding Effect

- Shielding effect: the number of electrons in full shells between the nucleus and the valence electrons.
- As you move from left to right across a period (→), shielding effect is constant.
- As you move down a group, the shielding effect increases.
  - There are more full electron shells, so atoms become larger and the valence electrons are further from the nucleus.
  - This effectively decreases the attraction between the electrons and the nucleus.

Net Nuclear Attraction

- Is calculated by taking the nuclear charge (Z = the number of protons, or atomic number) and subtracting the shielding effect
- Net nuclear attraction is a relative measure of the actual attraction between the nucleus and the valence electrons in an atom.
- As you move from left to right across a period (→), net nuclear attraction increases.

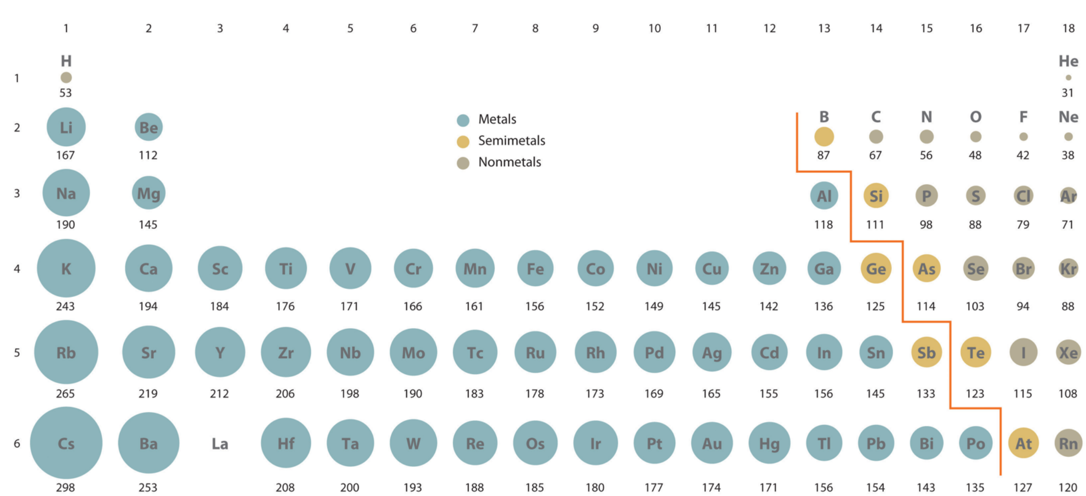
Trends across a period are due to net nuclear attraction.

Trends down a group are due to increasing shielding effect.

- There are five trends that we identify within the periodic table

A) Atomic Radius

- **Atomic radius (r):** half the distance between two nuclei of adjacent atoms.
- The size of an atom decreases going from left to right across a period.
  - This is because as you move down the period net nuclear attraction increases.
- The size of an atom increases as you go down a group.
  - This is because as you move down the groups, the number of occupied electron shells increases.



B) Ionic Radii

- The addition or removal of an electron from an atom results in an ion.
- Positive ions are always smaller than the neutral atom from which they are formed.
  - This is because nuclear attraction has increased.
- Negative ions are always larger than the neutral atom from which they are formed.
  - This is because nuclear attraction has decreased.

C) Electronegativity

- **Electronegativity**: The ability of an atom to attract an electron away from another atom.
  - Elements with high electronegativity have a strong tendency to gain an electron or electrons.
- Electronegativity increases going from left to right across a period.
  - This is because nuclear attraction is increasing.
- Electronegativity decreases going down a group.
  - This is because shielding increases down a group, making it easier to remove an electron.

D) Ionization Energy

- **Ionization Energy**: The amount of energy required to remove an electron from an atom.
  - The first ionization energy refers to the amount of energy required to remove the outermost electron.
  - The second ionization energy refers to the amount of energy required to remove the second outermost electron and so on.
  - The outermost electron is the easiest electron to remove. Removal of subsequent electrons requires more energy.
  - In terms of increasing energy: 1st Ionization Energy < 2nd Ionization Energy < 3rd Ionization Energy.
- Ionization Energy increases as you move left to right across a period.
  - This is because as the atom gets smaller, the valence electrons become closer to the nucleus. This increases nuclear attraction.
- Ionization Energy decreases as you move down a group.
  - This is because of shielding. The inner shells that are filled shield the outer shells from the positive charge of the nucleus, making outer electrons easier to remove.

E) Electron Affinity

- **Electron Affinity:** The amount of energy released when an electron is added to an element.
  - The higher an element's electron affinity the greater the attraction for an electron.
- Electron Affinity increase as you move left to right across a period.
  - This is because as you move across a period, net attraction increases.
- Electron Affinity decreases as you move down a group.
  - This is because as you move down a group, the attraction for electrons decreases.