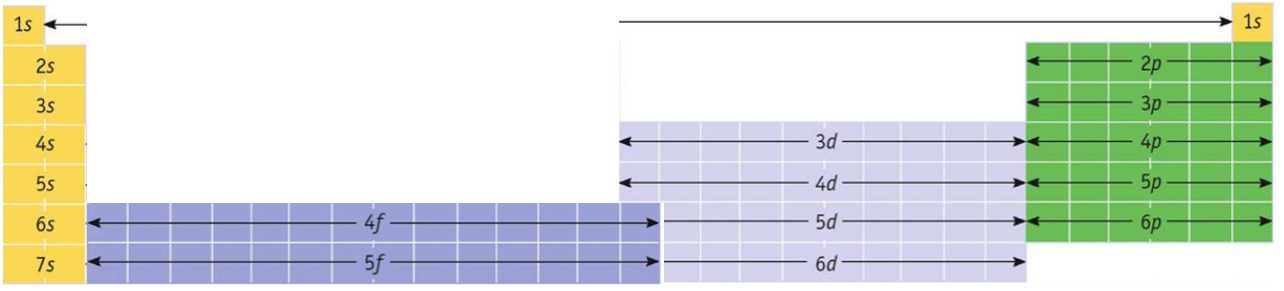


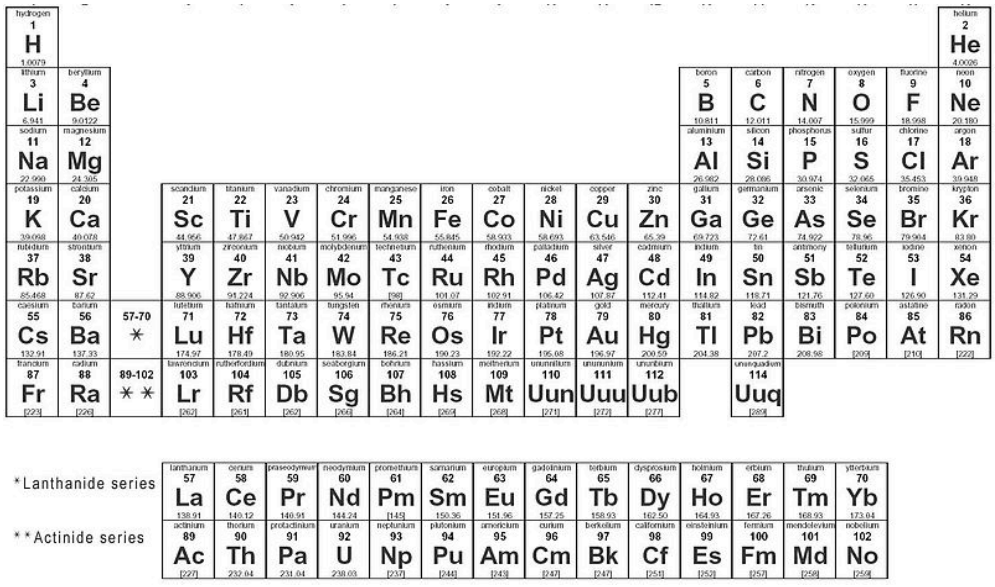
The Periodic Table

Why is it called a *periodic* table?

The properties of the elements in the table repeat in a "periodic" way (specific pattern).

Periodic law: There is a periodic repetition of chemical and physical properties of the elements when they are arranged by increasing atomic number

Metals, nonmetals, and metalloids



- stair steps separate metals from nonmetals
- metalloids line the stair-step line
  - > Semimetals: have properties in between metals and nonmetals

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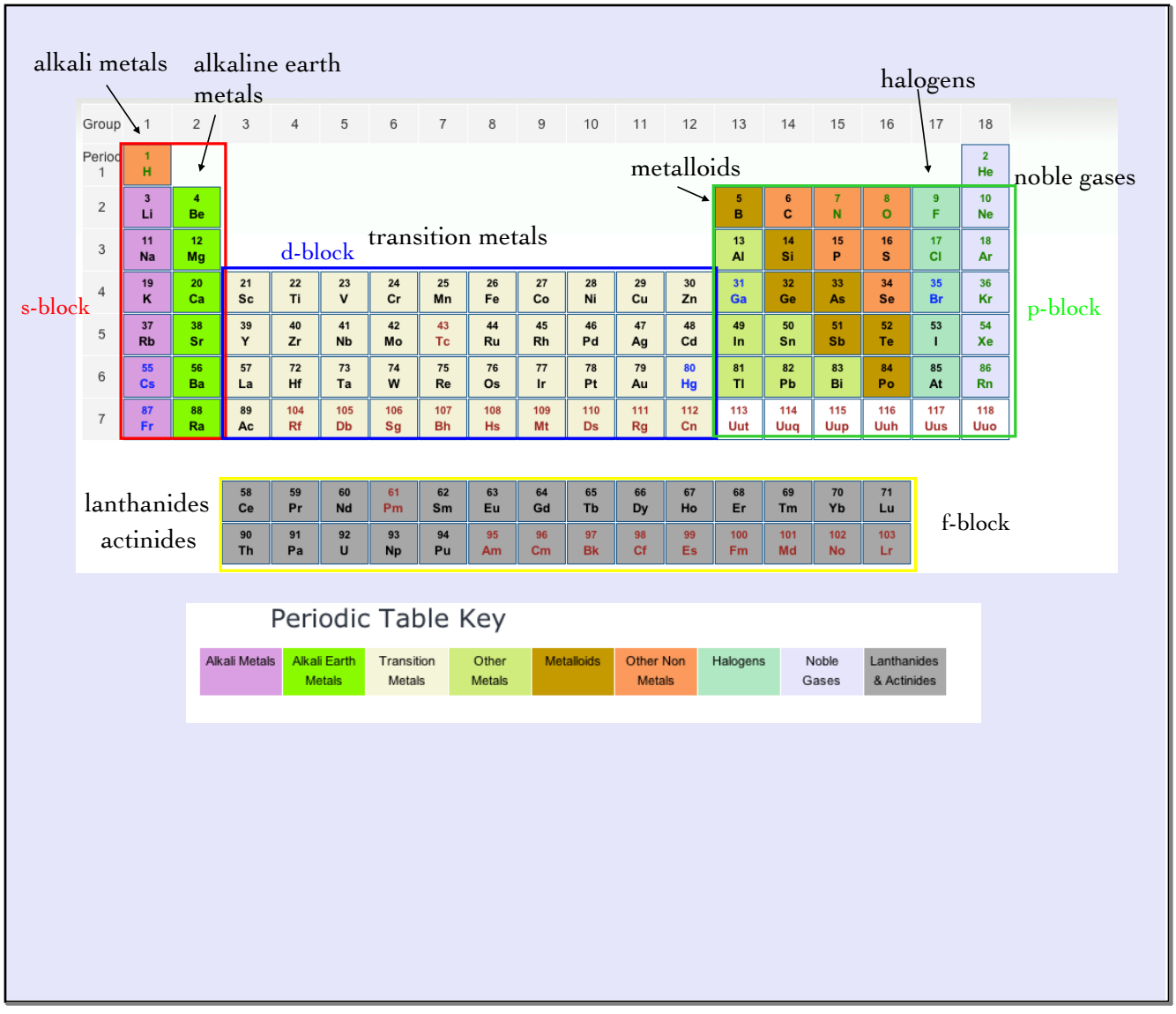


Metals, nonmetals, and metalloids

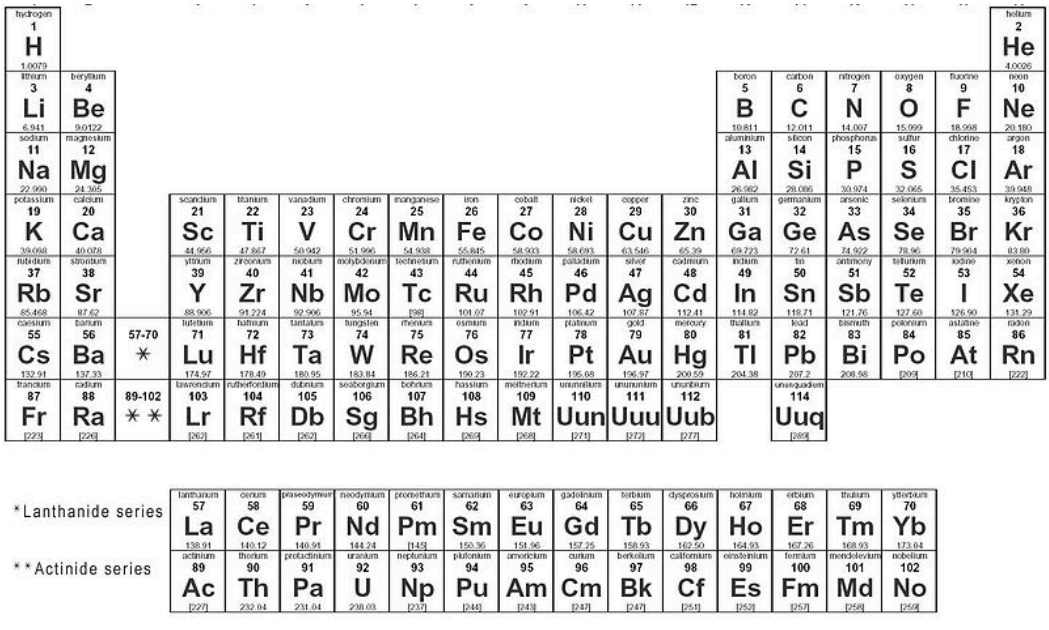
Metals	Metalloids	Nonmetals
Examples:	Examples:	Examples:

Periodic Table

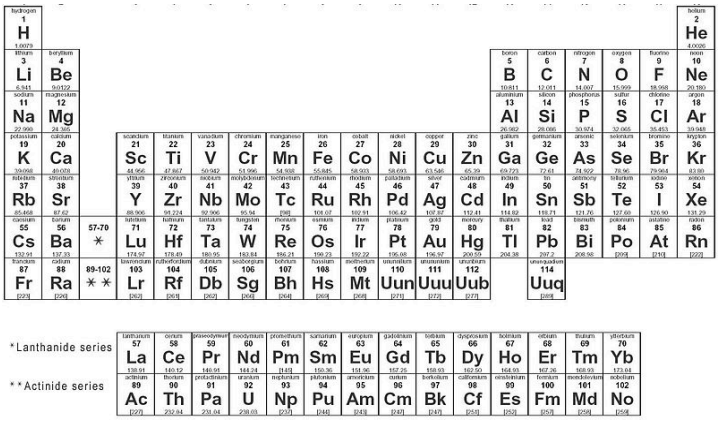
Periods = rows

- 7 periods
- period equals the highest principal energy level and therefore, the principal energy level of the valence electrons.

Examples: Li, Na, K



Periodic Table



Groups = columns

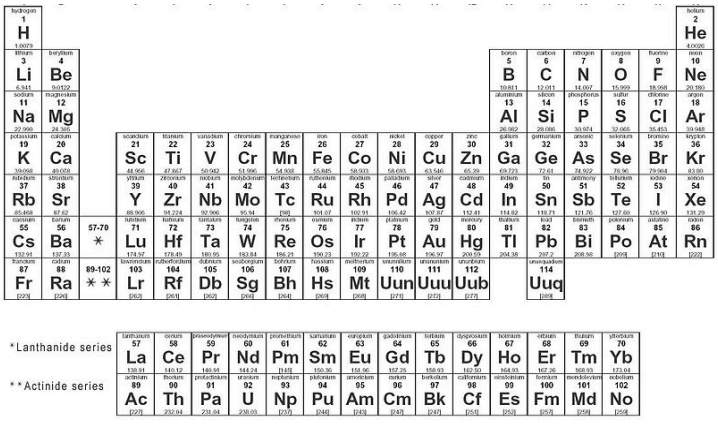
- Elements in the same group have the same number of valence electrons.
- Number of groups in a block (s, p, d, f) corresponds to the maximum number of electrons that can occupy that sublevel.
- 18 labeled groups

Li [He] $2s^1$

Na [Ne] $3s^1$

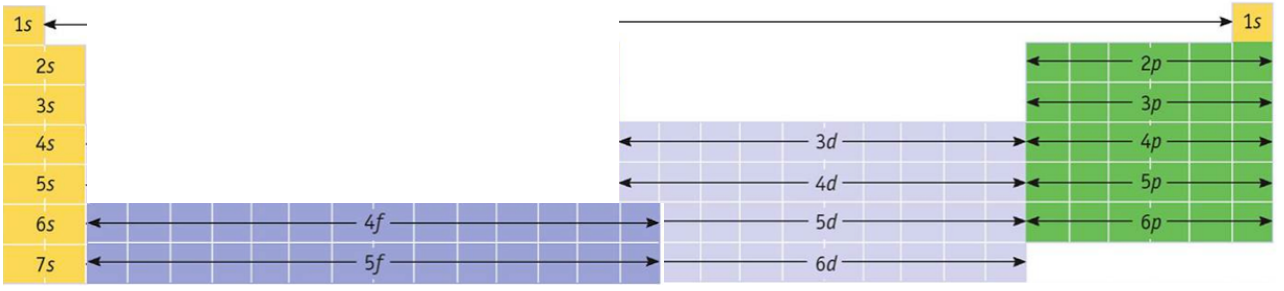
K [Ar] $4s^1$

Periodic Table



Groups = columns

- group 1 = alkali metals
- group 2 = alkaline earth metals
- group 17 = halogens
- group 18 = noble gases
- groups 3-12 = transition metals
- lanthanide and actinide series = "inner transition metals"



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Periodic Law

- Physical properties of the alkaline metals

- Physical properties of the halogens

Periodic Trends: chemical properties of elements are determined by the # of valence electrons.

- Properties are periodic because the number of valence electrons is periodic.

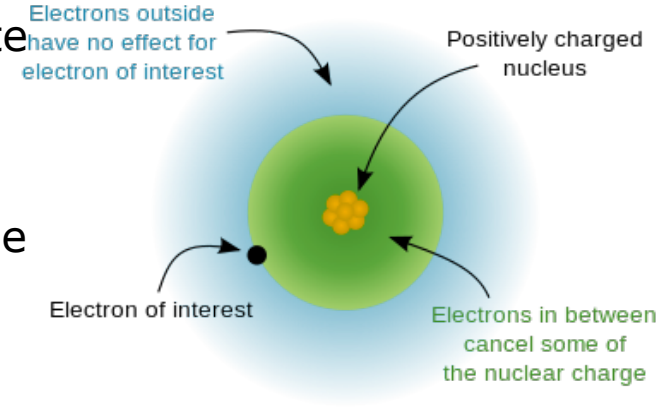
1. Atomic Radius
2. Ionic Radius
3. Ionization energy
4. ~~Electron Affinity~~
5. Electronegativity

Before we begin looking at the different periodic trends, we need to talk about a concept called **electron shielding effect**

Electrons and protons are attracted to one another because they have opposite charges.

Core electrons block some of the nuclear charge of the nucleus from the valence electrons.

Effective nuclear charge (Z_{eff}) is the charge felt by the valence electrons after you have taken into account the number of **shielding electrons** that surround the nucleus.



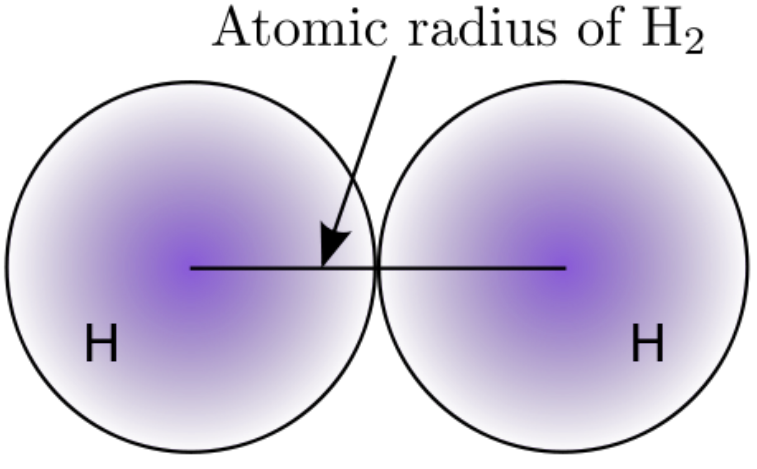
http://en.wikipedia.org/wiki/Effective_nuclear_charge

$$Z_{eff} = Z - \#\ core\ e^-$$

1. Atomic Radius

The electron cloud surrounding a nucleus is based on *probability* and does not have a clearly defined edge.

- Atomic radius = half the distance between nuclei of adjacent identical atoms
  - > Adjacent nuclei in a crystal
  - > Nuclei of two atoms bonded together



http://en.wikibooks.org/wiki/High_School_Chemistry/Atomic_Size

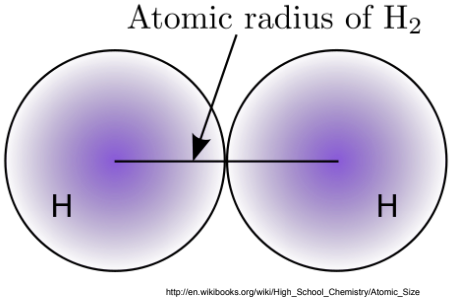
1. Atomic Radius

Atomic radius is determined by the strength of attraction between the valence electrons and the nucleus

Coulombs Law: $E = k \frac{q_1 \, q_2}{d}$

Greater charge = greater attraction

Greater distance = weaker attraction



http://en.wikibooks.org/wiki/High_School_Chemistry/Atomic_Size

Atomic Radius

Two factors determine the atomic radius:

① the principal ~~quantum number~~ energy level (n)

② effective nuclear charge which is the amount of charge felt by the valence e- after considering the shielding e- (Core e-)

(Z_{eff})

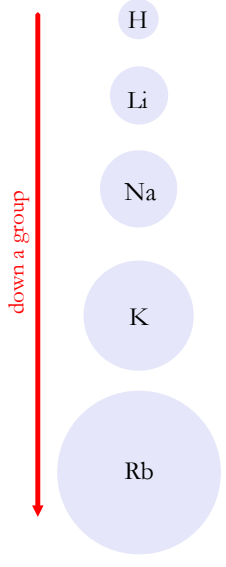
Z - #core e-

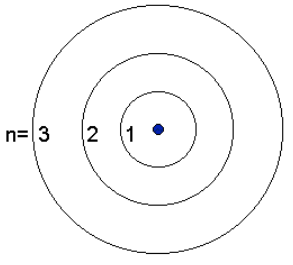
$E \propto \frac{Q_1 Q_2}{d}$

energy to remove e- according to Coulombs law

Atomic Radius

The larger the principal quantum ~~number~~ energy level (n), the larger the region of space occupied by the electrons and the farther from the nucleus the e- are found.





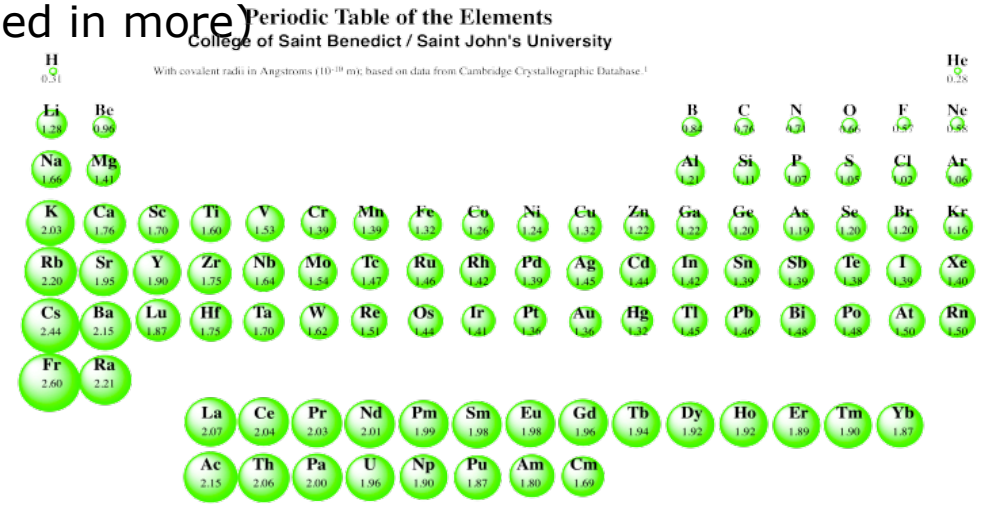
1. Atomic Radius

Moving down a group: Atomic radius increases

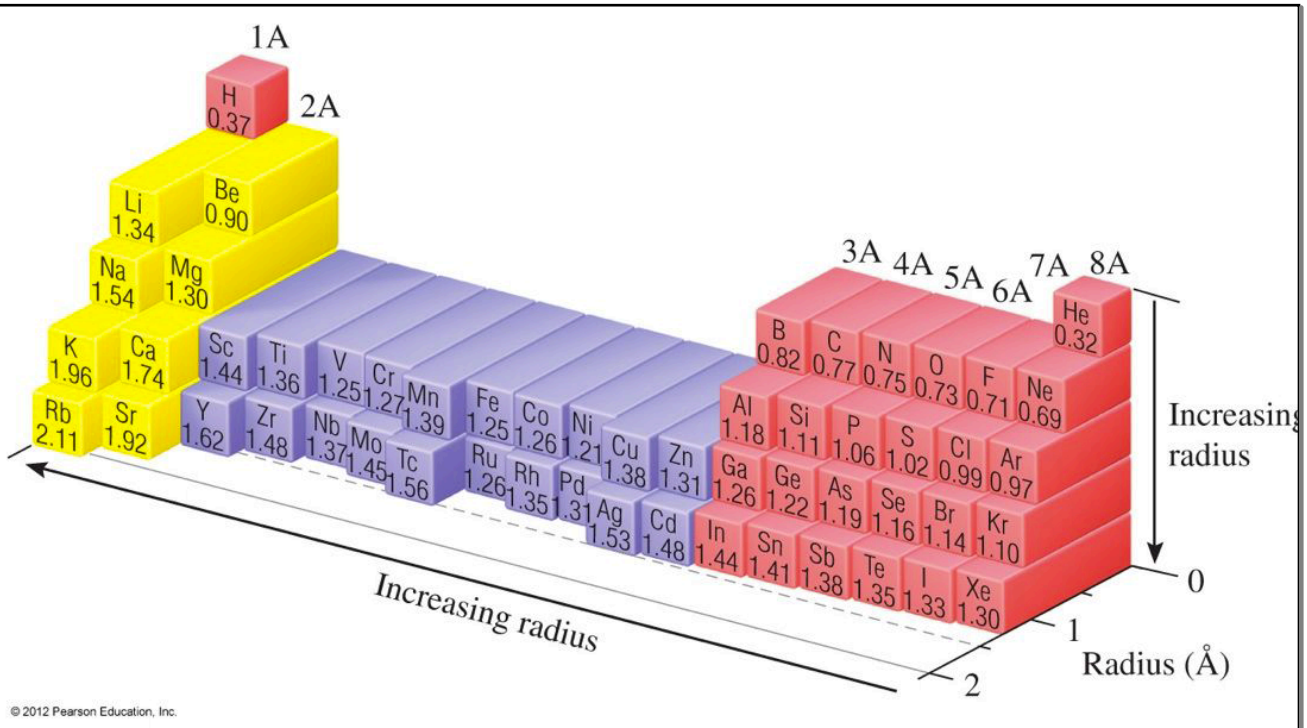
- Principal energy levels increase = increased orbital size, electrons are farther from nucleus

Moving across a period: Atomic radius decreases

- Principal energy level remains the same
- # of protons increases -> Z_{eff} increases (electrons pulled in more)



http://employees.csbsju.edu/cschaller/Principles%20Chem/atoms/atomperiodic.htm



Order the following atoms from smallest to biggest atomic radius:

Ga, Se, K

I, F, Cl

2. Ionic Radius

Ion = atom that has gained or lost electrons

- #p ≠ # e-
- Has a positive or negative charge

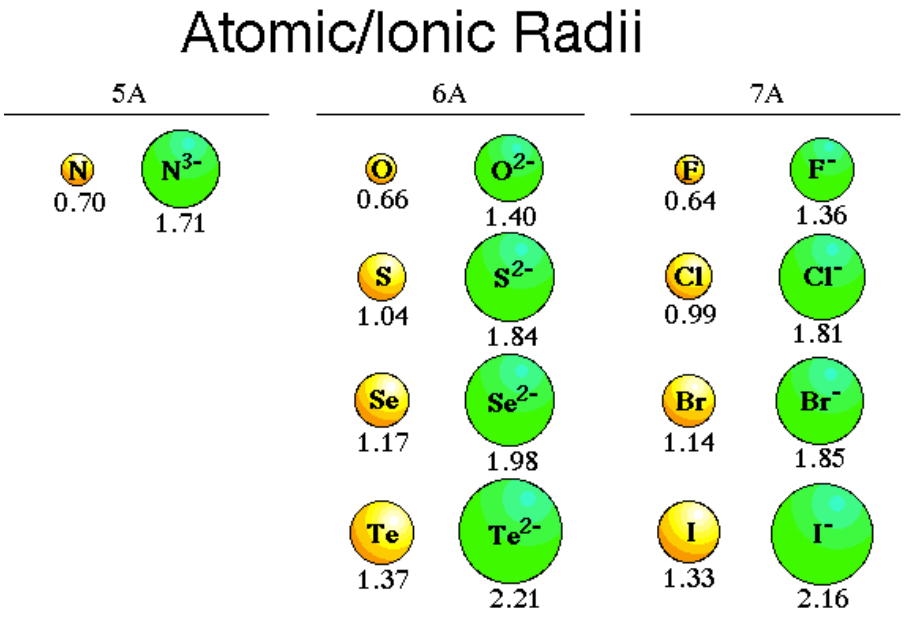
2. Ionic Radius

When an atom gains e- (anion): radius increases

- # protons same, but electron cloud increases (electrons repel each other)

When an atom loses e- (cation): radius decreases

- Loss of valence electron causes electron cloud to be pulled tighter by protons.



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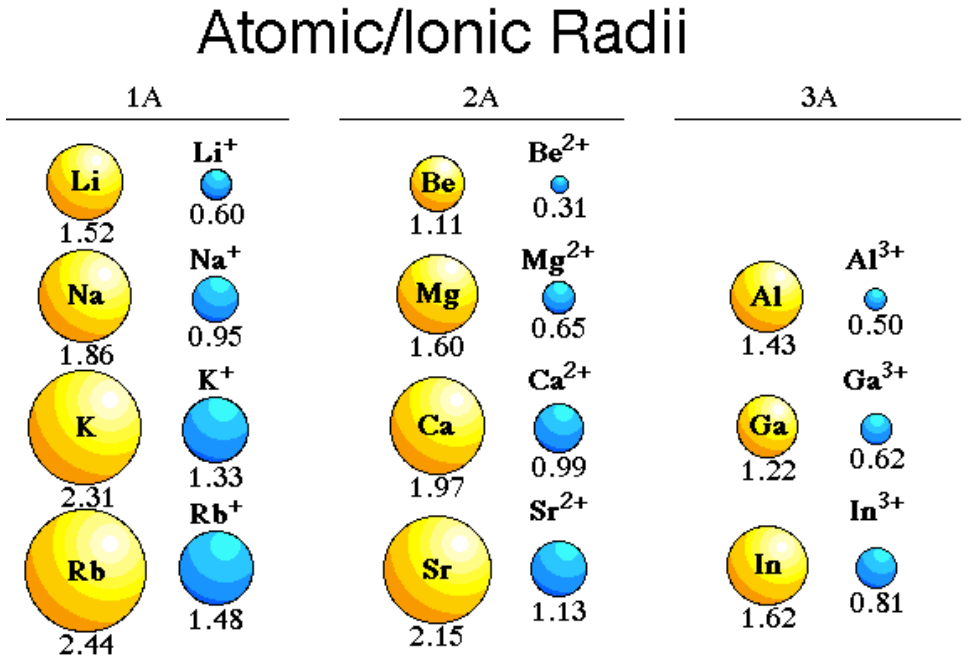

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- Loss of valence electron causes electron cloud to be pulled tighter by protons.

Moving down a group: ionic radii increases

Moving across a period: ionic radii of cations decreases, ionic radii of anions decreases

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