# N18 – Atomic Structure and Periodicity

## Periodic Trends

**Link to YouTube Presentation:** https://youtu.be/RwCDvBtAGbo

# N18 – Atomic Structure and Periodicity

## Periodic Trends

**Target:** I can describe and explain various patterns/ trends visible on the periodic table by using concepts such as shielding and nuclear attraction.

# Patterns work really well!

**Mendeleev predicted the properties of lots of elements!**

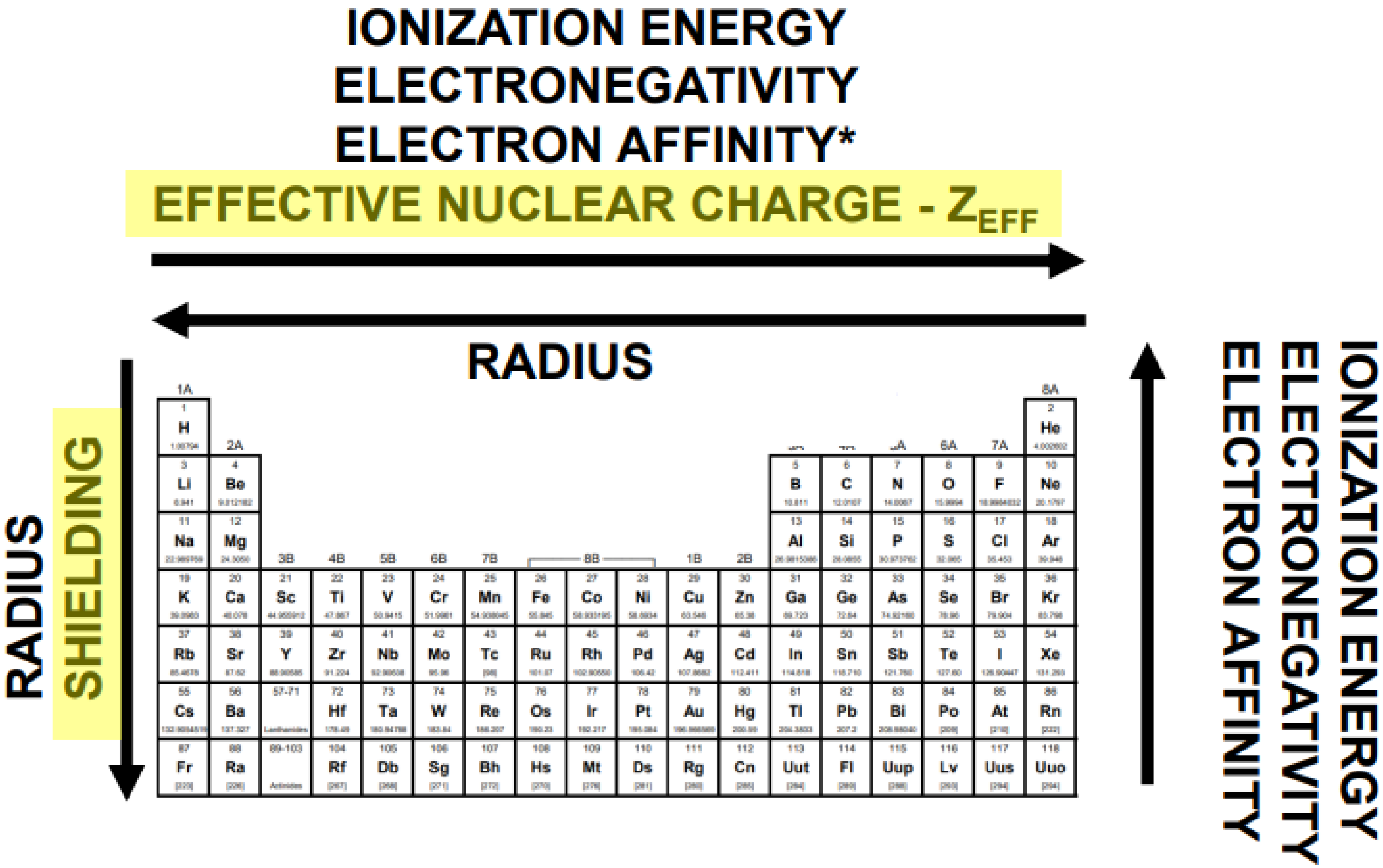
**Gallium (eka-aluminum)**

| | Mendeleev's predicted properties | Actual properties |
|---|---|---|
| Atomic mass | About 68 amu | 69.72 amu |
| Melting point | Low | 29.8 °C |
| Density | 5.9 $g/cm^3$ | 5.90 $g/cm^3$ |
| Formula of oxide | $X_2O_3$ | $Ga_2O_3$ |
| Formula of chloride | $XCl_3$ | $GaCl_3$ |

**Germanium (eka-silicon)**

| | Mendeleev's predicted properties | Actual properties |
|---|---|---|
| Atomic mass | About 72 amu | 72.64 amu |
| Density | 5.5 $g/cm^3$ | 5.35 $g/cm^3$ |
| Formula of oxide | $XO_2$ | $GeO_2$ |
| Formula of chloride | $XCl_4$ | $GeCl_4$ |

# Summary of Periodic Trends

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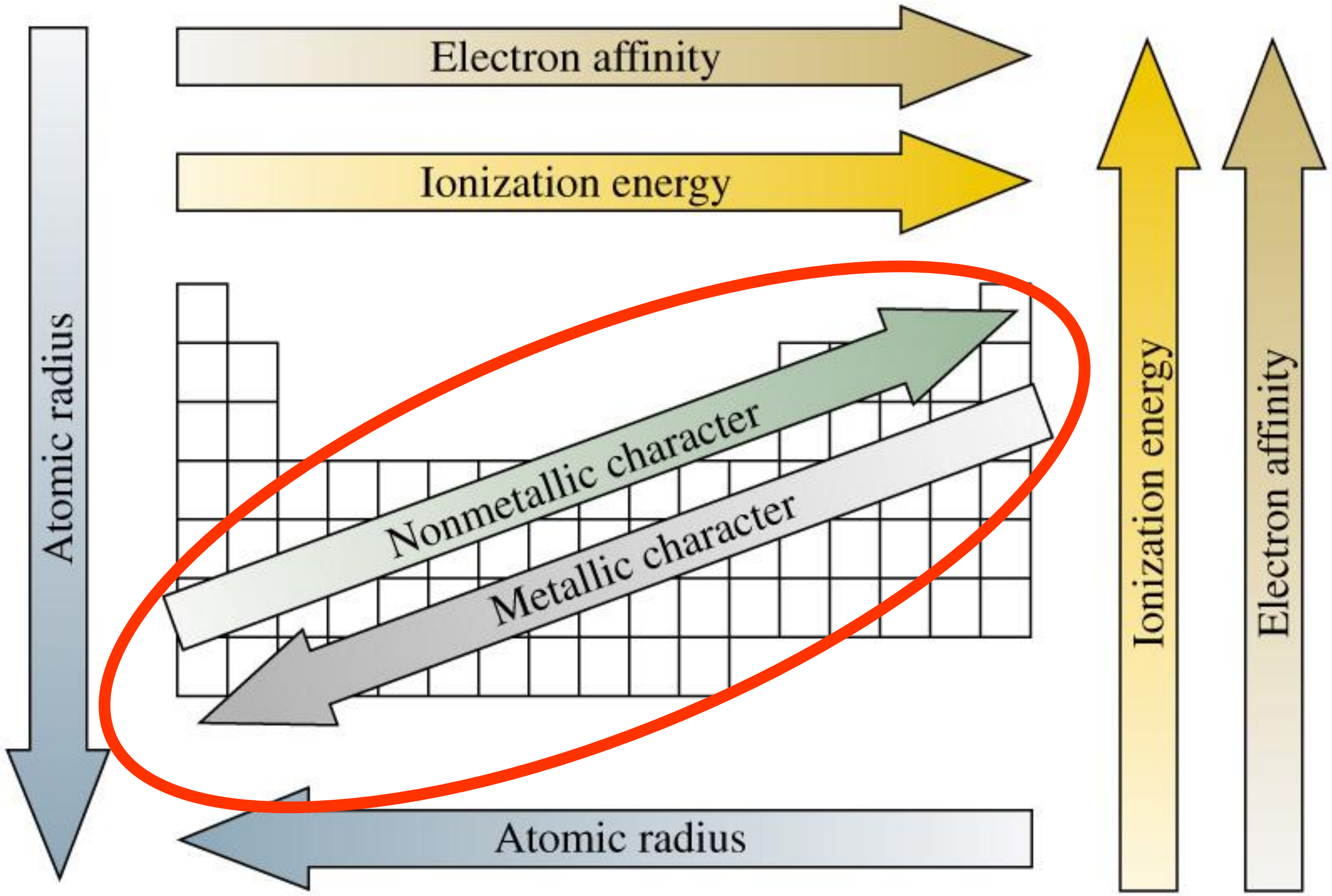

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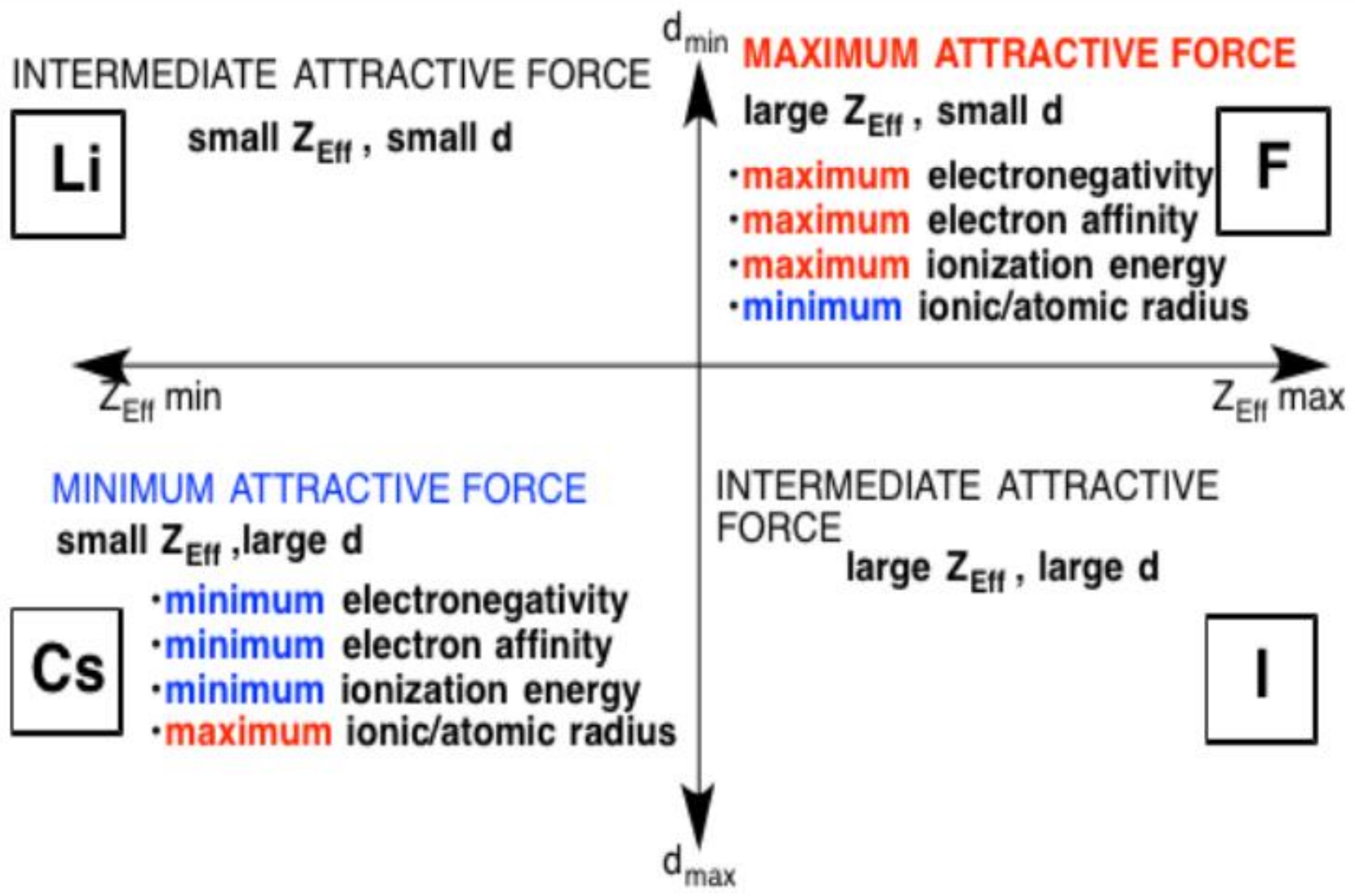
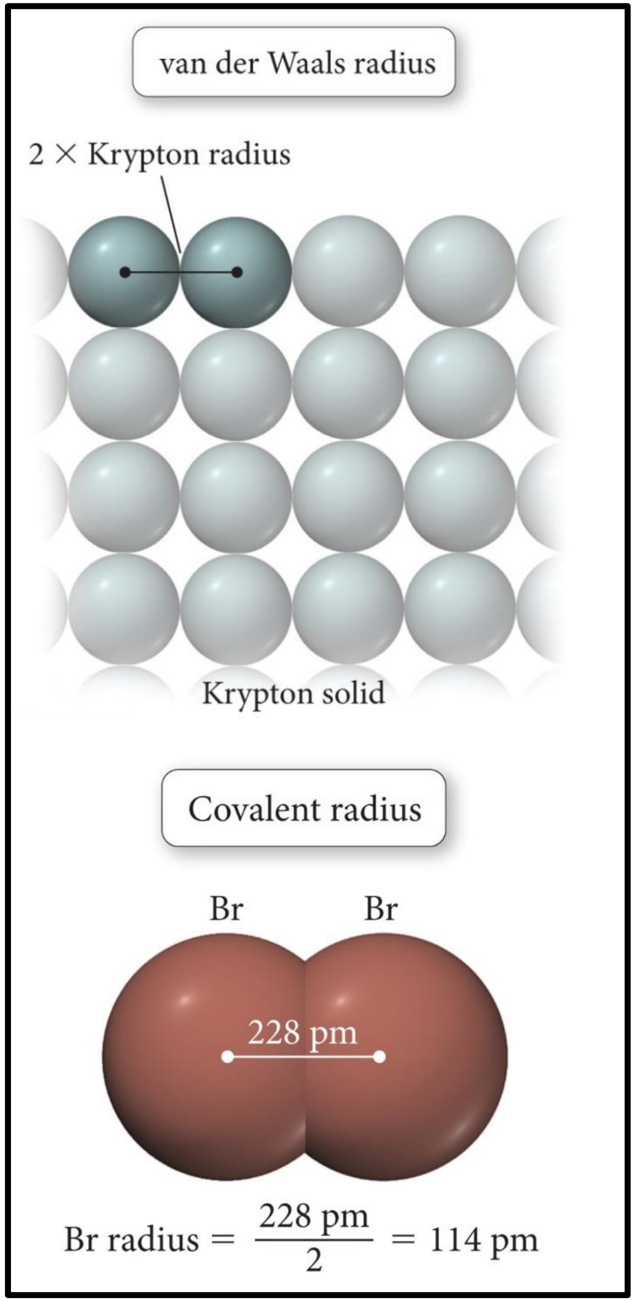


# Atomic Radius

- **Several ways to measure**
  - **Van der Waals radius = nonbonding**
  - **Covalent radius = bonding radius**

- **All give slightly different values**

- **Atomic radius is an average radius of an atom based on measuring large numbers of elements and compounds.**



# Atomic Radius

# Atomic Radius Trend

**KEY POINTS TO DESCRIBE GOING DOWN A GROUP:**

- Can ***NOT*** just say "because there is more shielding"
  – no vocab dropping!

- The size of an atom is related to the *distance* the valence electrons are from the nucleus and the *attractive forces.*

- You ***must*** specifically mention that the higher energy level is bigger volume and further away from nucleus.
  – yes this seems obvious…
  ***but*** if you want points be careful!

# Radius – Reasoning

**Increases down a group (top to bottom)**

**Moving down a group:**

- Adds a (principal) energy level.

**The larger the principal energy level an orbital is in:**

- The larger its volume.
- The farther the $e^-$ (most probable distance) is from nucleus.
- The less attraction it will have for the nucleus.
- The more shielding the valence electrons experience from inner core electrons.

**Therefore: The larger the radius**

# Atomic Radius Trend

**KEY POINTS TO DESCRIBE GOING ACROSS A PERIOD:**

- Can ***NOT*** just say "because there is greater effective nuclear charge"
  - – no vocab dropping!
- The size of an atom is related to the *distance* the valence electrons are from the nucleus and the *attractive forces.*
- As you go to the right there are more protons added **BUT shielding *doesn't* increase** since the e's are added to the same energy level.
- You ***must*** specifically mention that this results in greater nuclear attraction due to increased # of protons (nuclear charge) and therefore a smaller radius
  - – yes this seems obvious…
    ***but*** if you want points be careful!

# Radius – Reasoning

**Decreases Across a Period (Left to Right)**

**Going to the right:**

- Adds a proton each time
- No addition of shielding (adding e- to same energy level)

**Adding a proton with no increased shielding:**

- Increases effective nuclear charge on the valence e’s
- The stronger the attraction it will have for the nucleus.

**The stronger the nuclear attraction:**

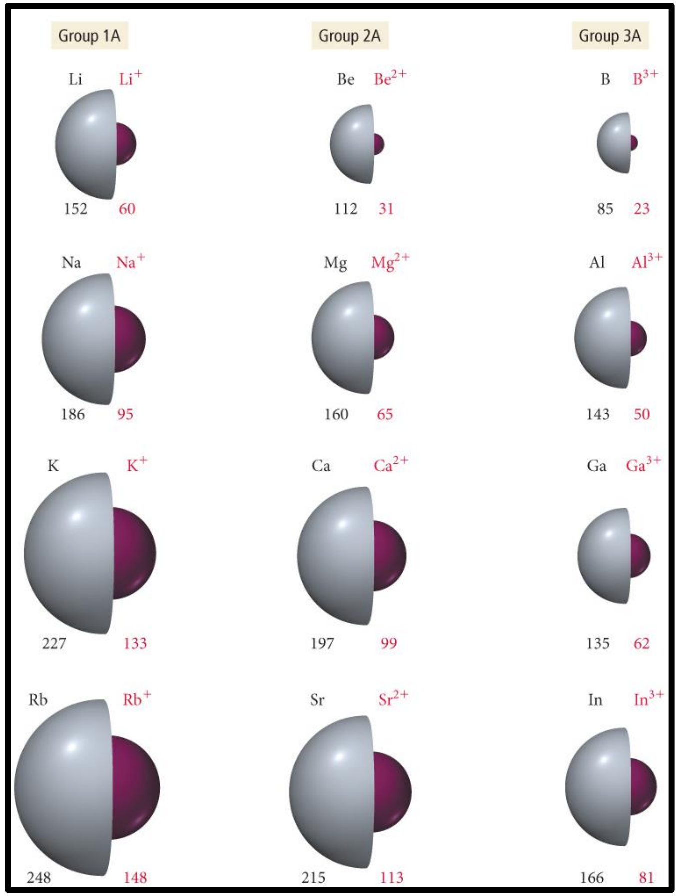
- The closer they are to the nucleus

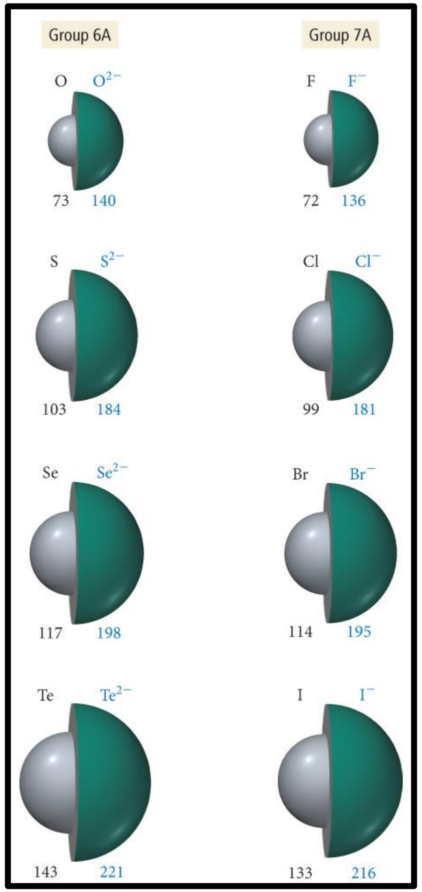
**Therefore: smaller radius**

# Ionic Radius Trend

- **Ions in same group have the same charge.**
- **Ion size increases down column.**
  - **Higher valence shell, larger radius**
- **Cations < neutral atoms**
- **Anion > neutral atoms.**

Neutral Cation Anion



# Ionic Radius Trend

Neutral Cation Anion

### Cations

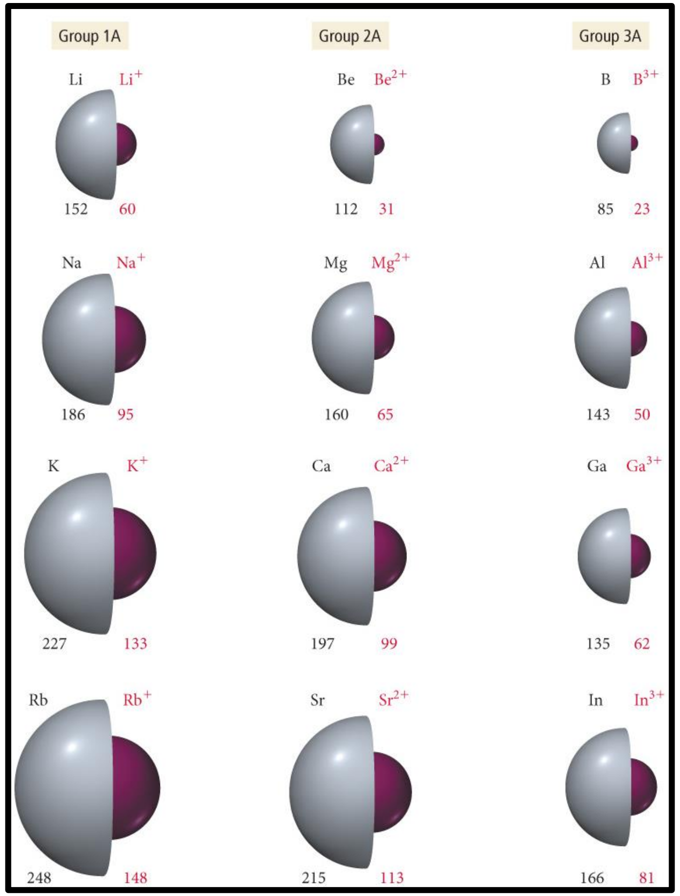
**Higher (+) charge, smaller radius**

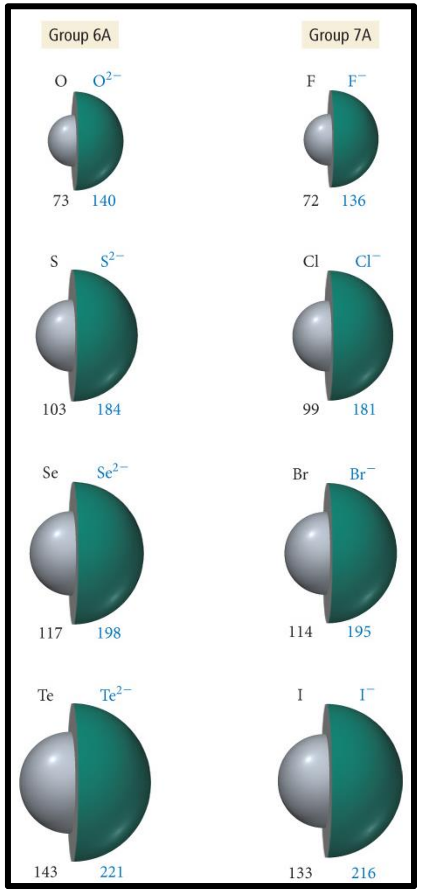
- $Mg^+$ radius > $Mg^{2+}$ radius

### Anion

**Higher (-) charge, larger radius**

- $O^-$ radius < $O^{2-}$ radius



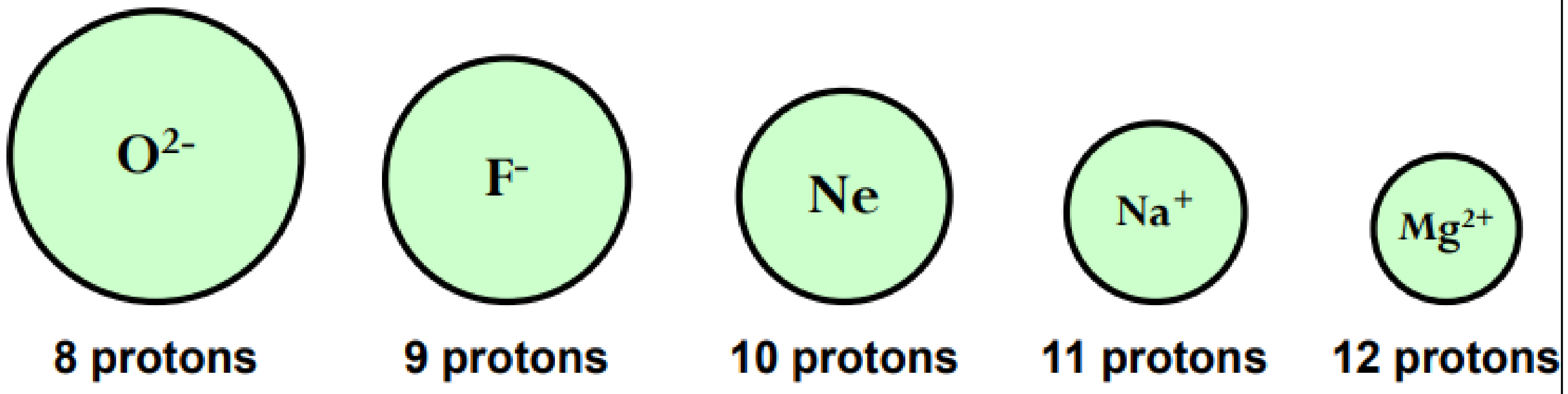
# Ionic Radius Trend

**Isoelectronic** = same electron configuration

**For isoelectronic species compare the # of protons to # e-**

- **More protons, more attraction, smaller radius**

*All these examples are $1s^2 2s^2 2p^6$*



# Radius

## Irregularities

**Radius doesn't change very much going across d-block**

- New electrons are being added to an INNER (core) energy level
    - Yes you are adding an extra proton to shrink radius
    - BUT you are also adding shielding!
    - Result is yes radius will decrease but not by as much as you might expect.

# Ionization Energy

# Ionization Energy (IE)

**Ionization Energy is the minimum energy needed to remove an electron from an atom or ion**

- In the gas state
- Endothermic process – takes energy
- Valence electron easiest to remove, lowest IE

**1st Ionization Energy – Energy to remove $e^-$ from neutral atom**

**$M_{(g)} + IE_1 \rightarrow M^{1+}_{(g)} + 1\ e^-$**

**2nd Ionization – Energy to remove $e^-$ from 1+ ion**

**$M^{1+}_{(g)} + IE_2 \rightarrow M^{2+}_{(g)} + 1e^-$**

# Ionization Energy (IE)

## Decreases down a group

- Each time you go down you have another energy level
- Inner core electrons shield outer electrons
- Increased radius
- Decreased nuclear attraction
- Easier to take away an electron
- Decreased IE

# Ionization Energy (IE)

**Increases across a period (left to right)**

- Each time you go to the right you add a proton
- No significant increase in shielding b/c adding e- to same energy level – they do not shield as well as inner levels
- Therefore, the radius decreases
- Increase in nuclear attraction
- Harder to take one away
- Increased IE

# Ionization Energy (IE)

## Irregularities

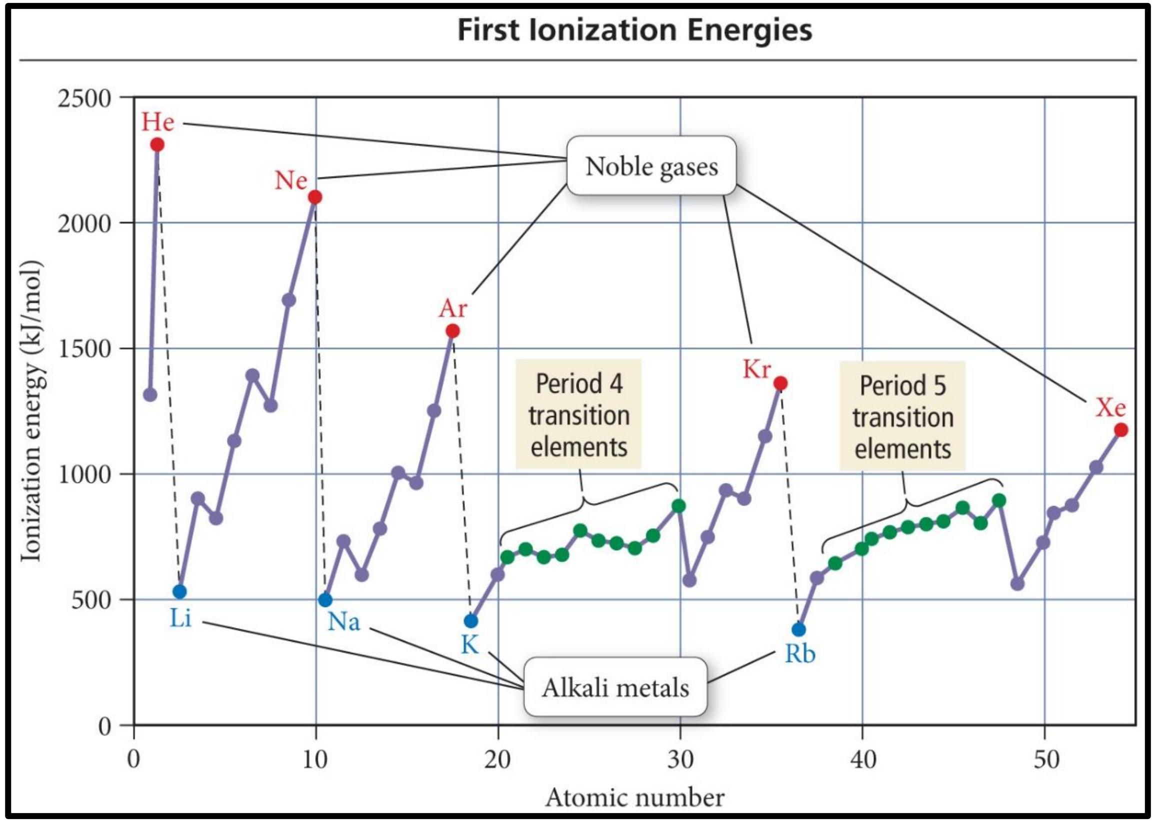
### Half filled and totally filled sublevels (orbital set)

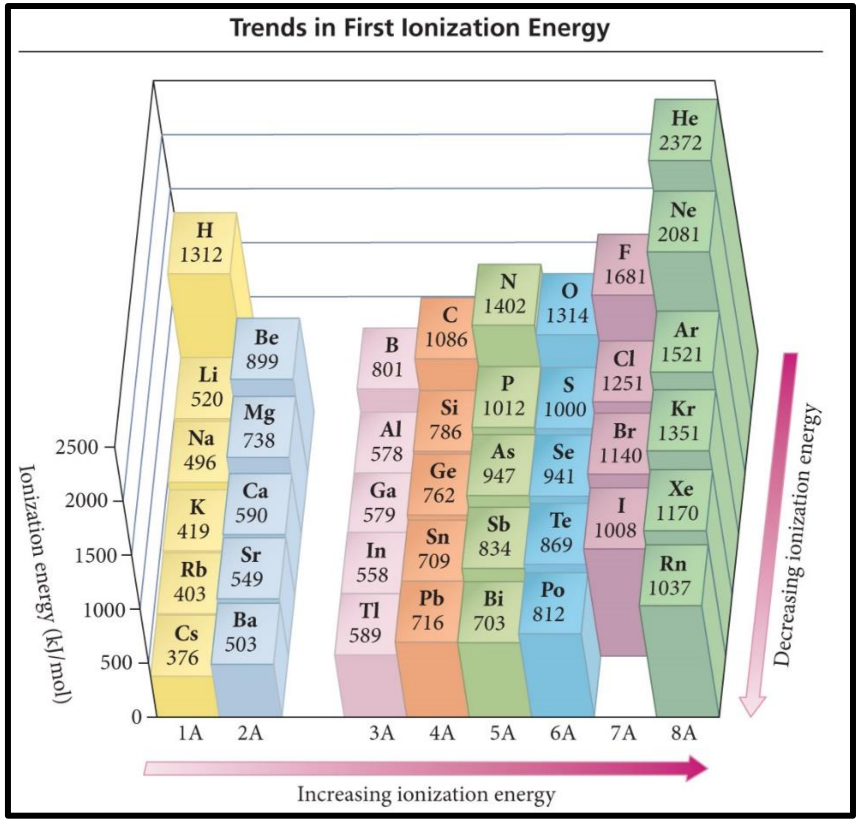
- Extra repulsions of electrons in paired orbitals
  - Makes it easier to remove an electron
  - Lower IE than expected

### Moving to a p orbital (Mg → Al)

- p orbital does not penetrate as much as an s orbital
  - Less nuclear attraction
  - Lower IE than expected

# Ionization Energy (IE)



# Ionization Energy (IE)

## Increases for successive $e^-$'s taken from same atom

- Each time you take one away, atom gets smaller.
- Smaller atom means greater nuclear attraction to valence e-
- Harder to take away another e-
- Increases IE

| Element | $IE_1$ | $IE_2$ | $IE_3$ | $IE_4$ |
|---|---|---|---|---|
| Na | 496 | 4560 | | |
| Mg | 738 | 1450 | 7730 | |
| Al | 578 | 1820 | 2750 | 11,600 |

# Successive Ionization Energies

## Large jump in IE shows when you begin removing core e$^-$'s

- Helps you figure out most likely charge on element
- The charge is the number of ionizations that happened BEFORE the large jump

**TABLE 8.1 Successive Values of Ionization Energies for the Elements Sodium through Argon (kJ/mol)**

| Element | $IE_1$ | $IE_2$ | $IE_3$ | $IE_4$ | $IE_5$ | $IE_6$ | $IE_7$ |
|---|---|---|---|---|---|---|---|
| Na | 496 | 4560 | | | | | |
| Mg | 738 | 1450 | 7730 | | | | |
| Al | 578 | 1820 | 2750 | 11,600 | | | |
| Si | 786 | 1580 | 3230 | 4360 | 16,100 | | |
| P | 1012 | 1900 | 2910 | 4960 | 6270 | 22,200 | |
| S | 1000 | 2250 | 3360 | 4560 | 7010 | 8500 | 27,100 |
| Cl | 1251 | 2300 | 3820 | 5160 | 6540 | 9460 | 11,000 |
| Ar | 1521 | 2670 | 3930 | 5770 | 7240 | 8780 | 12,000 |

Core electrons

# Electron Affinity

# Electron Affinity

**Electron Affinity – Δ in energy when neutral atom gains $e^-$**

- Gas state
- *Usually* energy is released (exothermic, negative value)

$$M_{(g)} + 1e^- \rightarrow M^{1-}_{(g)} + EA$$

- Some alkali metals and all noble gases are endothermic
- More energy released, the larger the electron affinity (larger negative = larger EA)

# Electron Affinity

**Alkali metals decrease electron affinity down the column.**

- But not all groups do
- Generally irregular increase in EA from second period to third period

**“Generally” increases across period**

- Becomes more negative from left to right
- Not absolute
- Group 5A often lower EA than expected - extra electron must pair
- Groups 2A and 8A generally very low EA because added electron goes into higher energy level or sublevel

**Highest EA in any period = halogen**

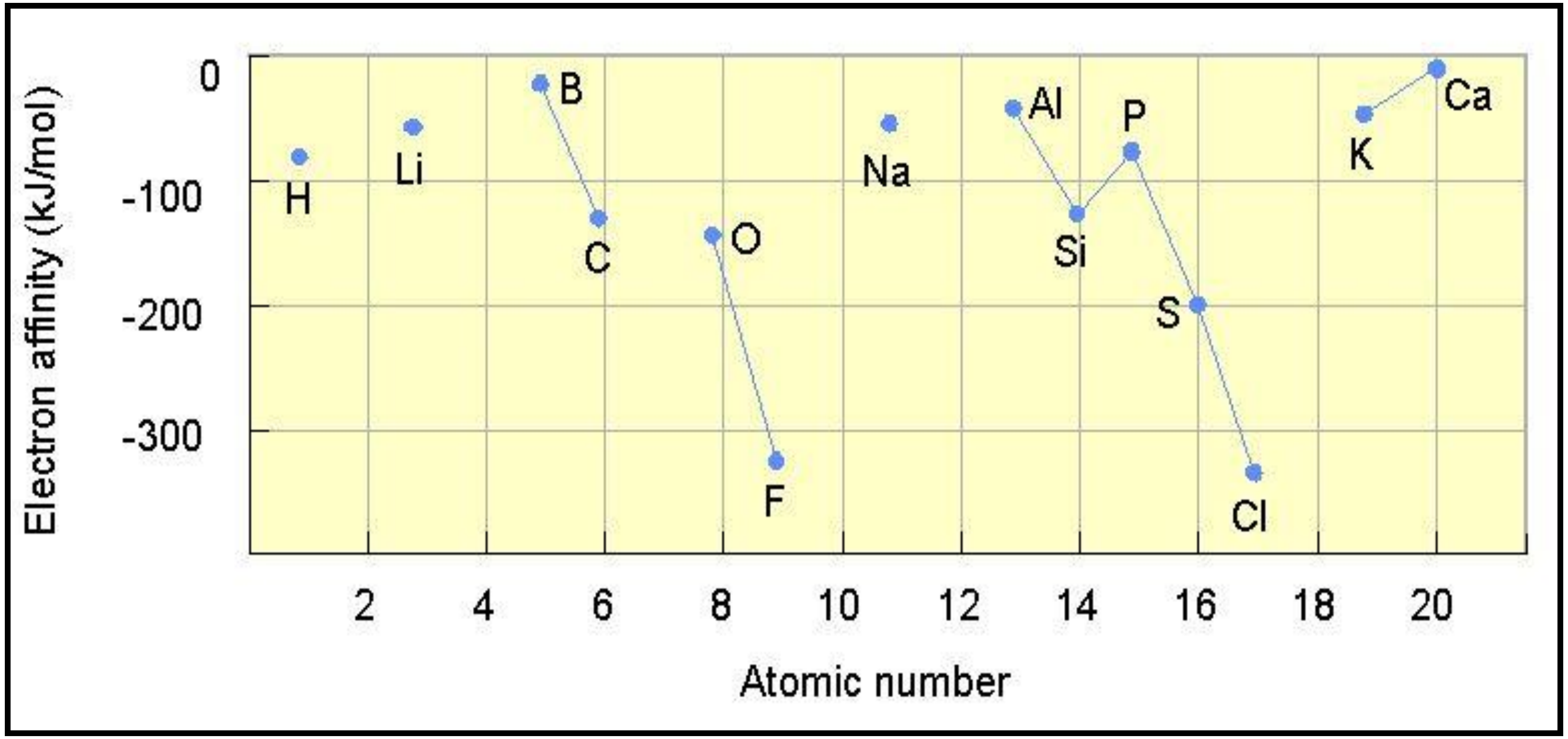
**Very irregular pattern compared to other PT Trends**

# Electron Affinity

**Electron Affinities (kJ/mol)**

| 1A | 2A | 3A | 4A | 5A | 6A | 7A | 8A |
|---|---|---|---|---|---|---|---|
| H −73 | | | | | | | He >0 |
| Li −60 | Be >0 | B −27 | C −122 | N >0 | O −141 | F −328 | Ne >0 |
| Na −53 | Mg >0 | Al −43 | Si −134 | P −72 | S −200 | Cl −349 | Ar >0 |
| K −48 | Ca −2 | Ga −30 | Ge −119 | As −78 | Se −195 | Br −325 | Kr >0 |
| Rb −47 | Sr −5 | In −30 | Sn −107 | Sb −103 | Te −190 | I −295 | Xe >0 |

# Electron Affinity



# Electronegativity and Polarity

# Electronegativity

**The ability of an atom to attract bonding electrons to itself is called electronegativity.**

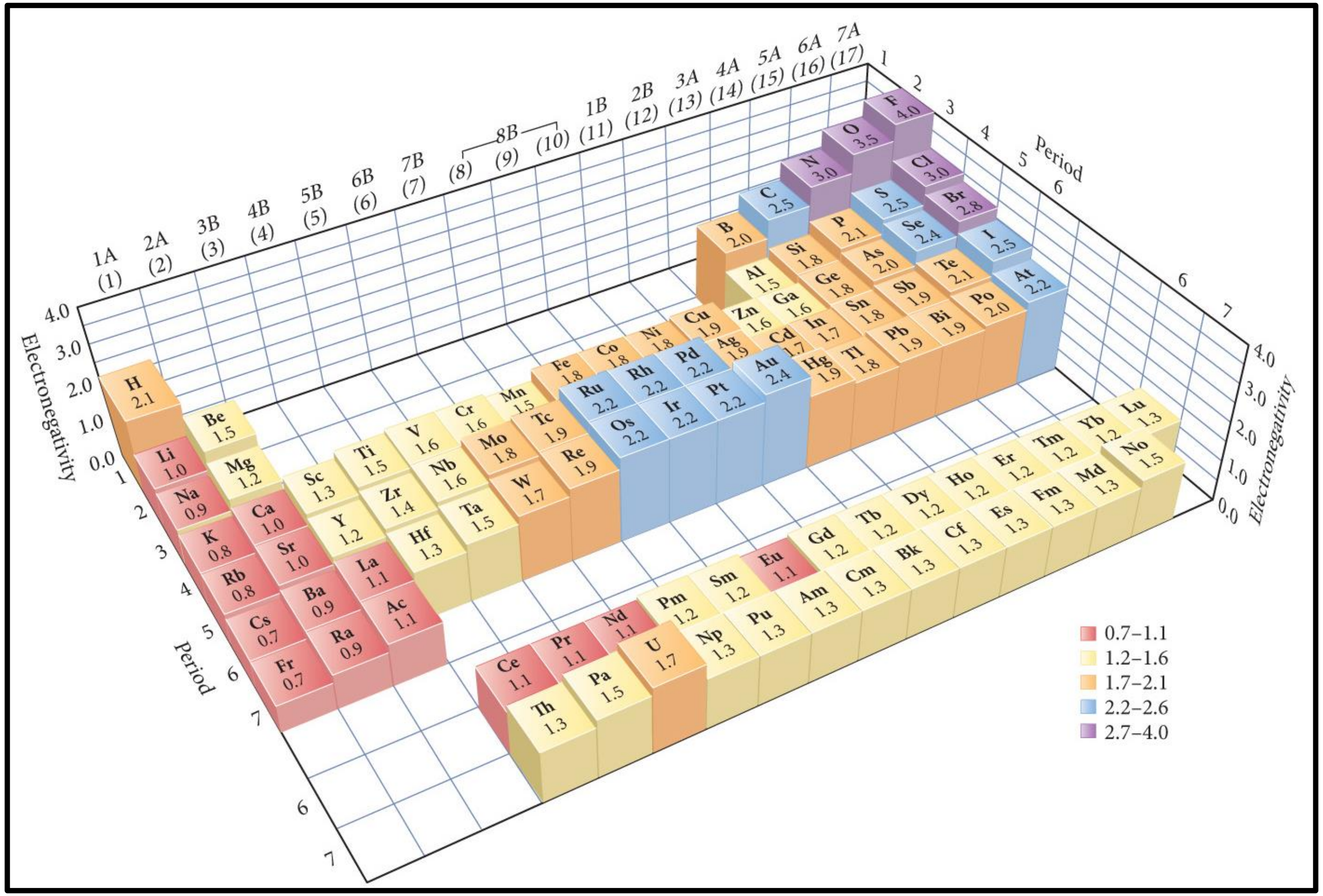
**Increases across period (left to right)**
**Decreases down group (top to bottom)**

- Fluorine - most electronegative
- Francium - least electronegative
- Noble gas atoms are not assigned values.
- Opposite of atomic size trend.

**The larger the difference in electronegativity, the more polar the bond.**

- Negative end toward more electronegative atom.

# Electronegativity



# Electronegativity Difference & Bond Type

## Pure Covalent

- Difference in electronegativity between bonded atoms is **0**
- Equal sharing

## Nonpolar Covalent

- Difference in electronegativity is **0.1 to 0.4**

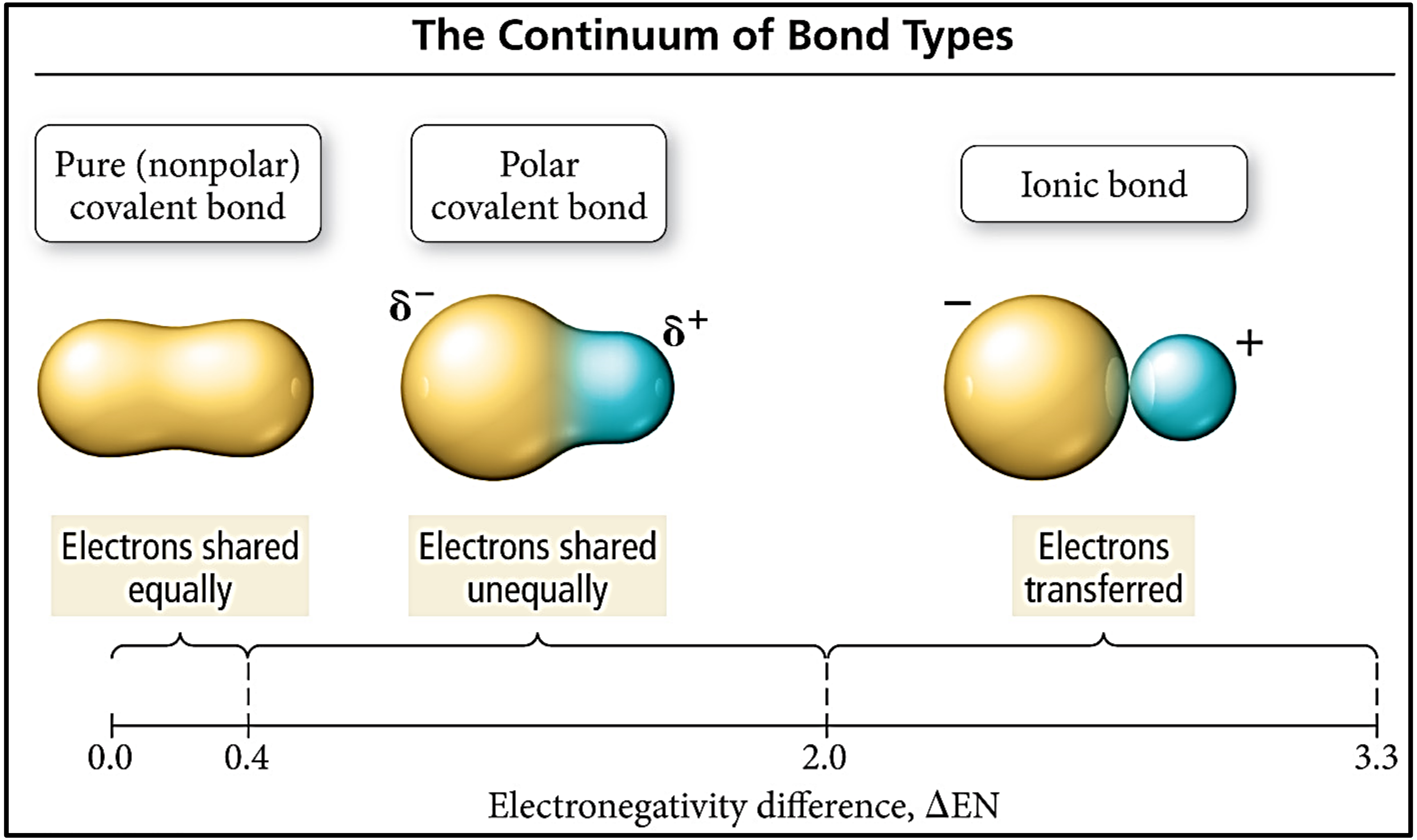
## Polar Covalent

- Difference in electronegativity is **0.5 to 1.9**

## Ionic

- Difference in electronegativity is **larger than or equal to 2.0**

# Electronegativity Difference & Bond Type




# Electronegativity Difference & Bond Type

**TABLE 9.1 The Effect of Electronegativity Difference on Bond Type**

| Electronegativity Difference (ΔEN) | Bond Type | Example |
| --- | --- | --- |
| Small (0–0.4) | Covalent | $Cl_2$ |
| Intermediate (0.4–2.0) | Polar covalent | HCl |
| Large (2.0+) | Ionic | NaCl |

# Bond Dipole Moments

**Dipole – A substance with a partial (+) and partial (-) end**

**Dipole moment - $\mu$, - a measure of bond polarity.**

- Directly proportional to the size of the partial charges and directly proportional to the distance between them.

$$\mu = (q)(r)$$

# Magnetic Properties

**Paramagnetic – Atom or ion with a net magnetic field**

- Result of unpaired electrons in orbitals
- Will be weakly attracted to a magnetic field

**Diamagnetic – Atom or ion with no magnetic field**

- Result of all paired electrons in orbitals
- Slightly repelled by a magnetic field

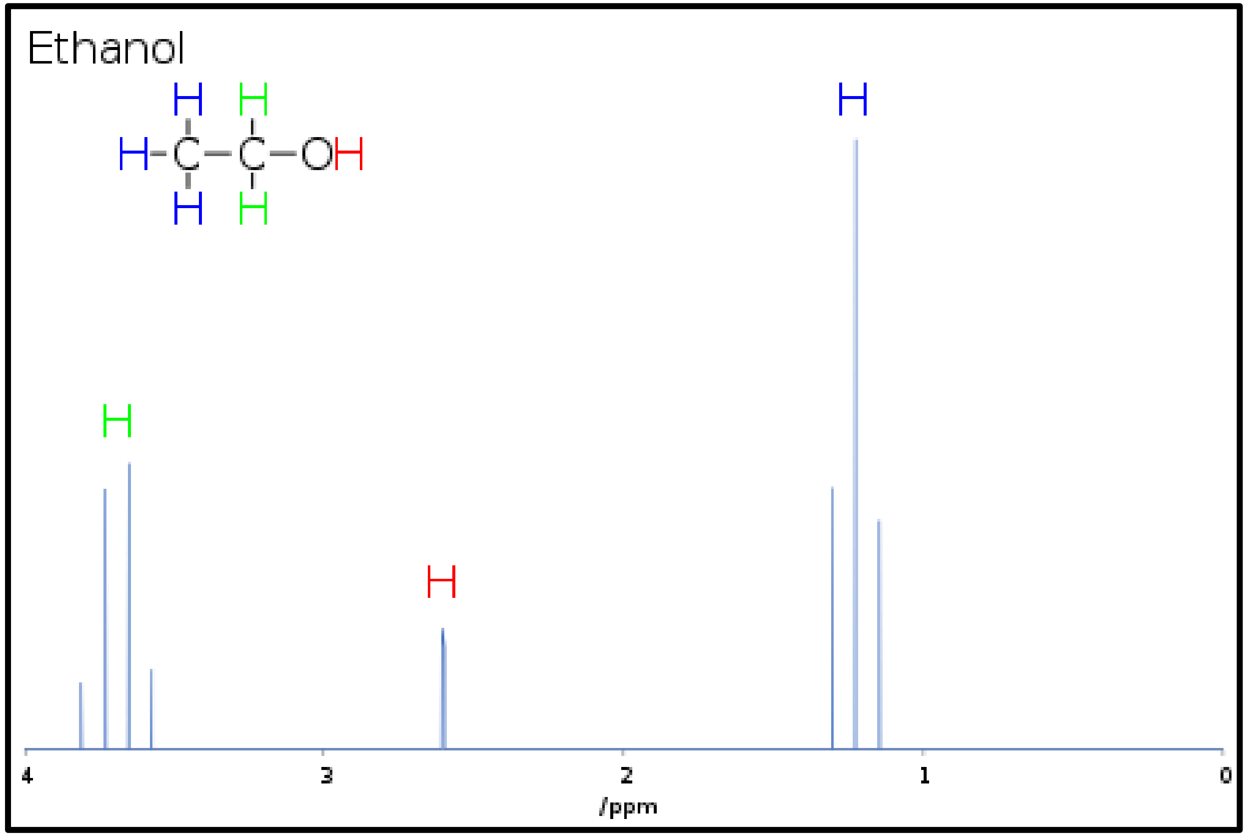
**Ferromagnetic – Group of atoms in a solid crystal or lattice that keeps its magnetism even when there is no magnetic field applied**

# Electron Configuration and Magnetism

# Why care about Paramagnetism?

**NMR Machine – Helps determine the structure of molecules**

**MRI – Applied to images of the body**



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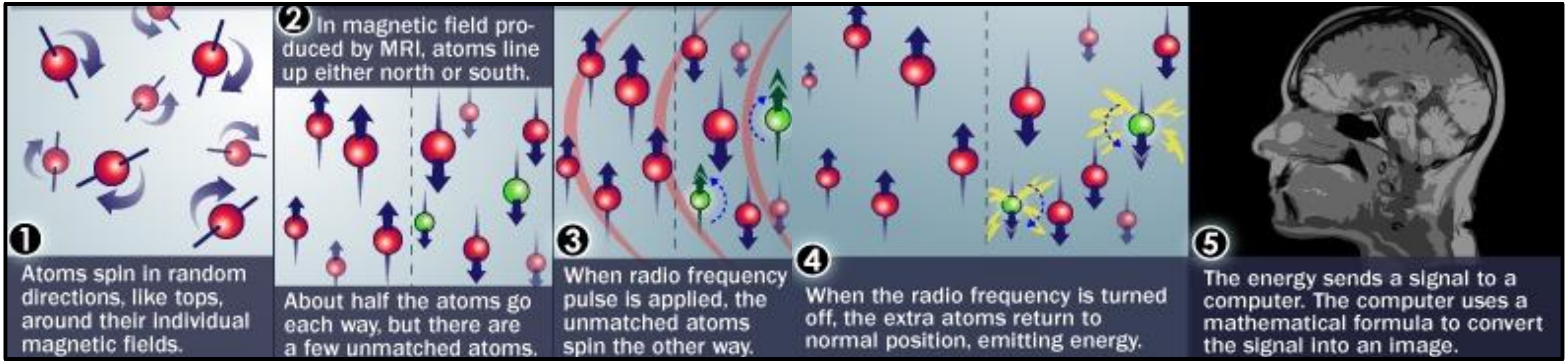

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