

N18 – Atomic Structure and Periodicity

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



N18 – Atomic Structure

and Periodicity

Periodic Trends

Patterns work really well!

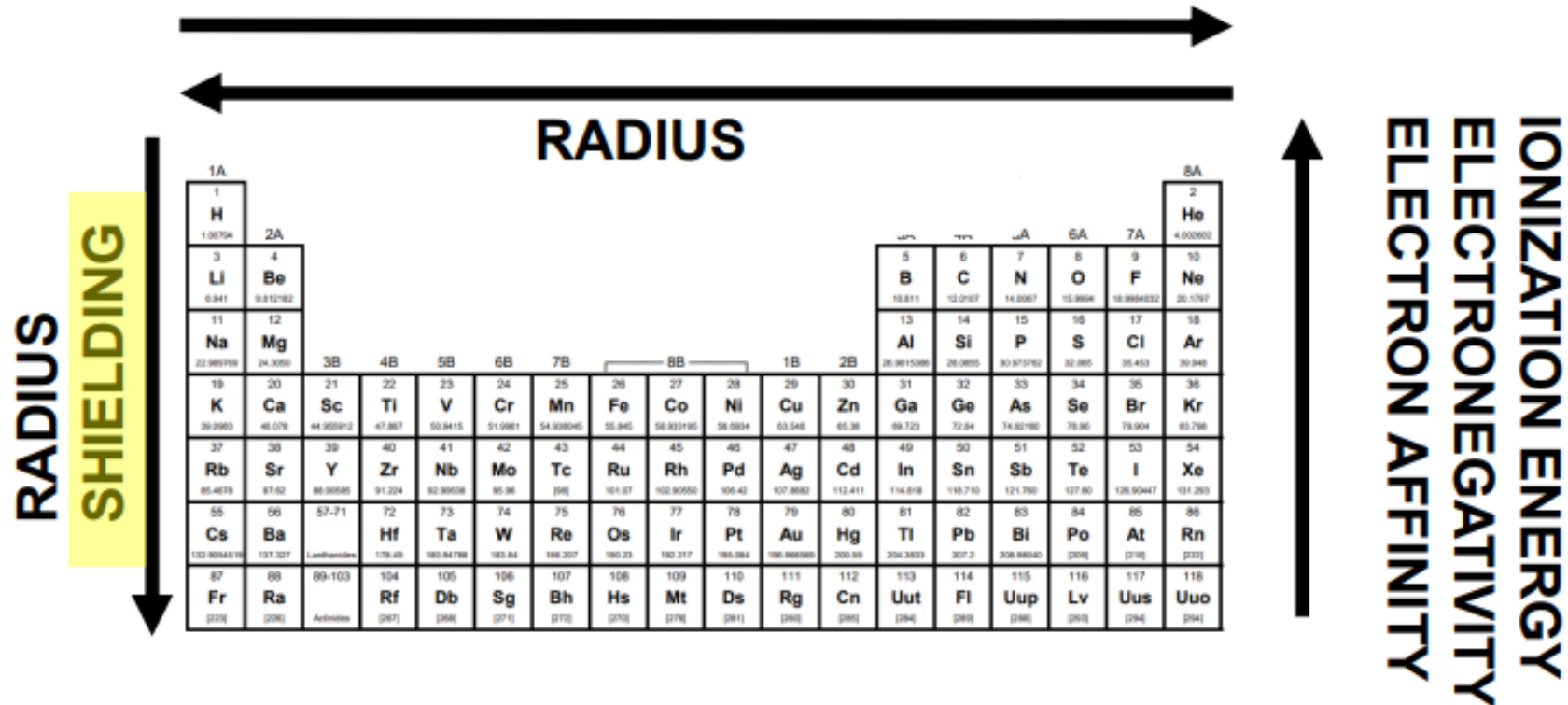
Mendeleev predicted the properties of lots of elements!

Gallium (eka-aluminum)			Germanium (eka-silicon)		
	Mendeleev's predicted properties	Actual properties		Mendeleev's predicted properties	Actual properties
Atomic mass	About 68 amu	69.72 amu	Atomic mass	About 72 amu	72.64 amu
Melting point	Low	29.8 °C	Density	5.5 g/cm ³	5.35 g/cm ³
Density	5.9 g/cm ³	5.90 g/cm ³	Formula of oxide	XO ₂	GeO ₂
Formula of oxide	X ₂ O ₃	Ga ₂ O ₃	Formula of chloride	XCl ₄	GeCl ₄
Formula of chloride	XCl ₃	GaCl ₃			

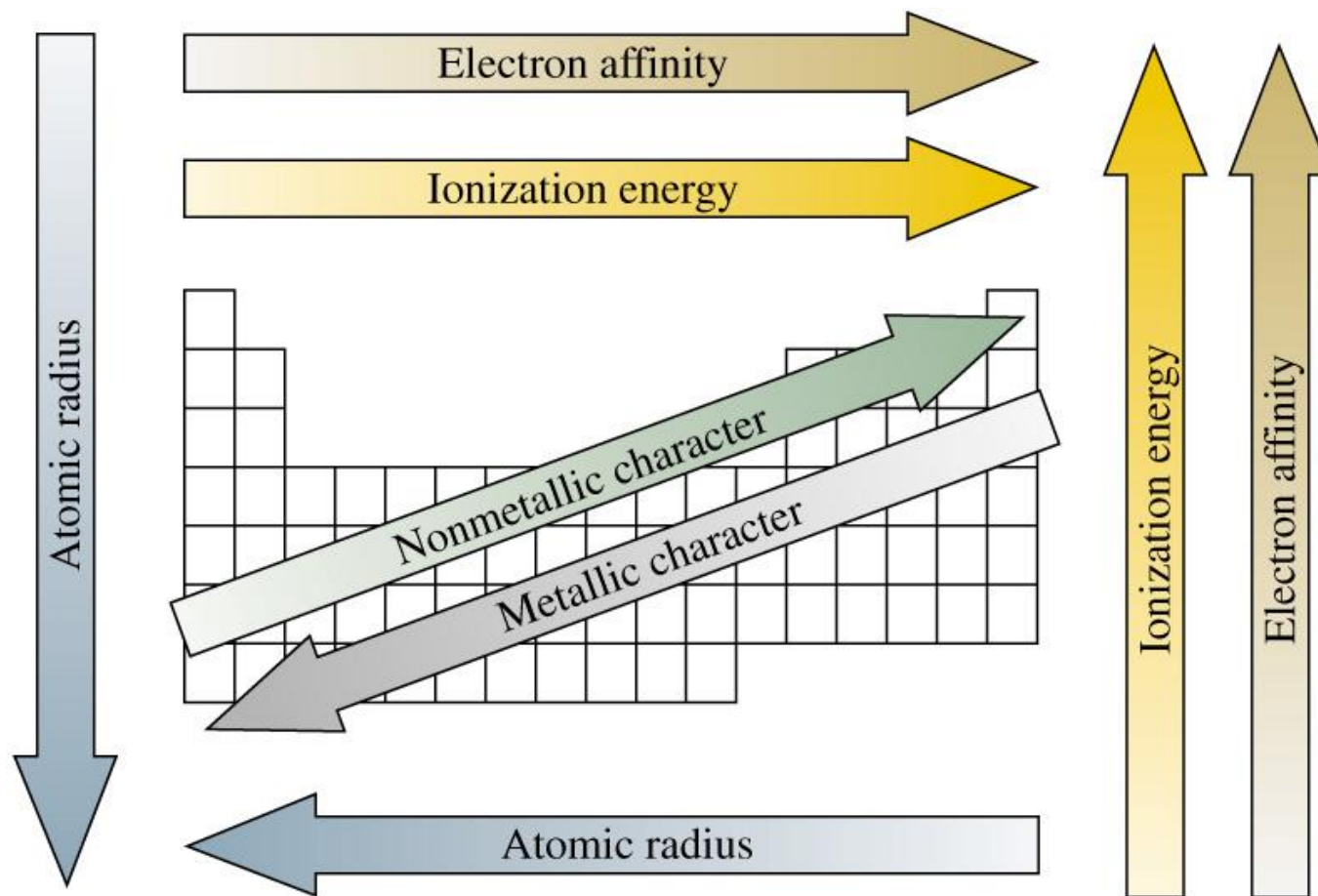
Summary of Periodic Trends

IONIZATION ENERGY
ELECTRONEGATIVITY
ELECTRON AFFINITY*

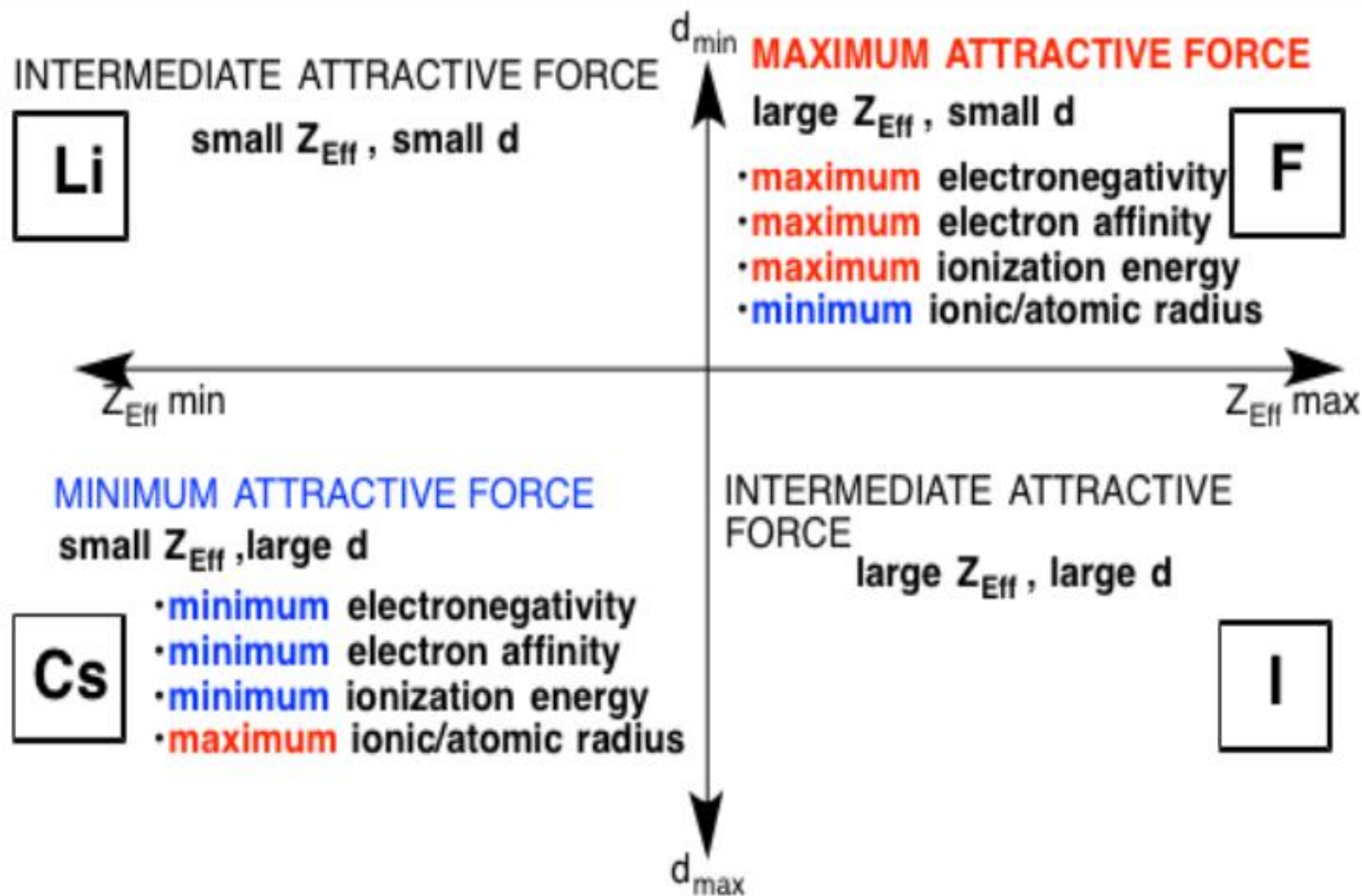
EFFECTIVE NUCLEAR CHARGE - Z_{EFF}



Summary of Periodic Trends

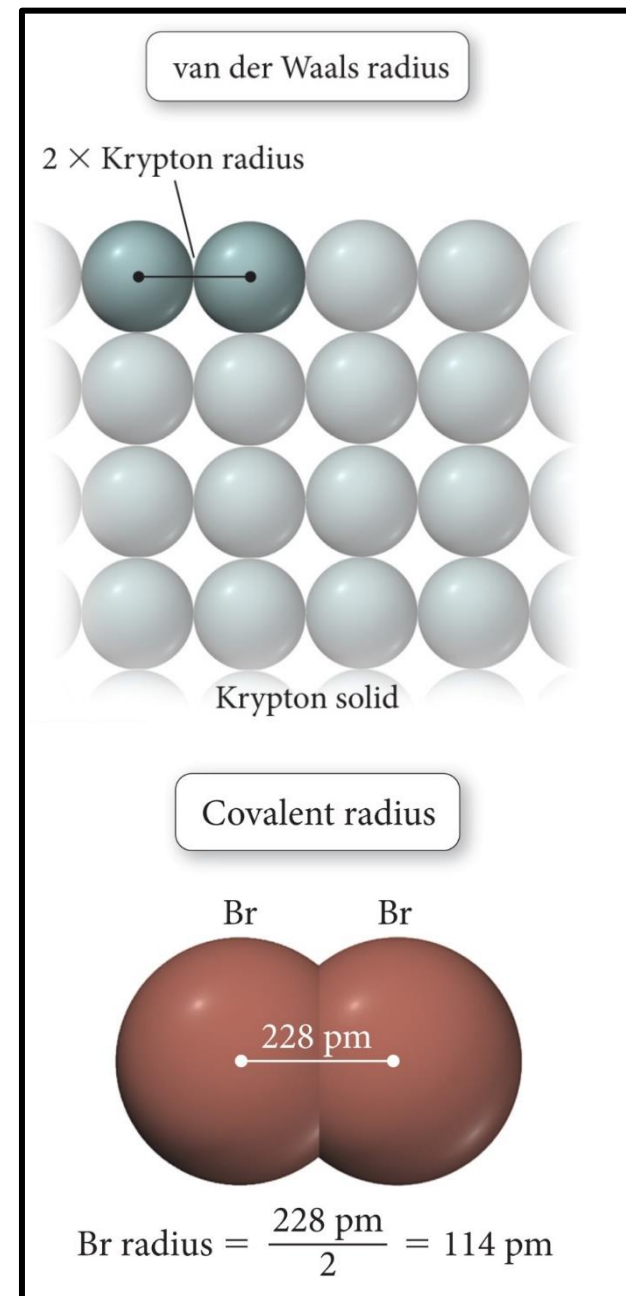


Summary of Periodic Trends



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

Radius – Quantum Mechanical Reason

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⁻'s 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.
- 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 and therefore a smaller radius
 - yes this seems obvious...but if you want points be careful!

Radius – Quantum Mechanical Reason

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.

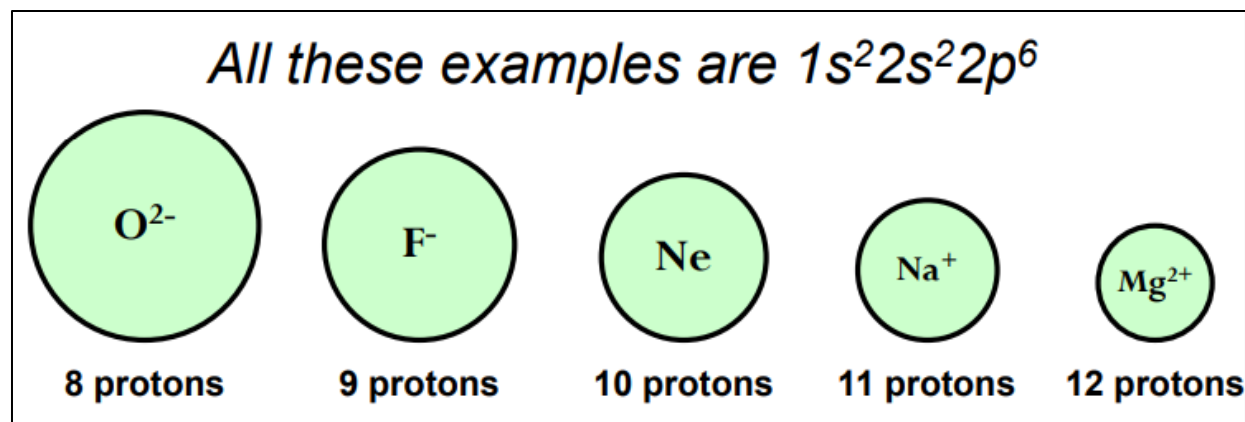
The stronger the nuclear attraction:

- The closer they are to the nucleus

Therefore: smaller radius







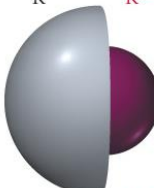




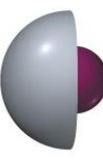
Ionic Radius Trend



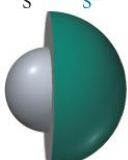

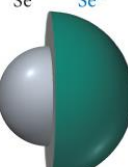
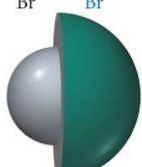


- Ions in same group have the same charge.
- Ion size increases down column.
 - Higher valence shell, larger
- Cations < neutral atoms
- Anions > neutral atoms.
- Cations < anions.
- Larger (+) charge = smaller cation
 - For isoelectronic species
 - **Isoelectronic** = same electron configuration
- Larger (-) charge = larger anion
 - For isoelectronic species



Ionic Radius Trend

Neutral
Cation
Anion

Group 1A	Group 2A	Group 3A
Li Li^+  152 60	Be Be^{2+}  112 31	B B^{3+}  85 23
Na Na^+  186 95	Mg Mg^{2+}  160 65	Al Al^{3+}  143 50
K K^+  227 133	Ca Ca^{2+}  197 99	Ga Ga^{3+}  135 62
Rb Rb^+  248 148	Sr Sr^{2+}  215 113	In In^{3+}  166 81

Group 6A	Group 7A
O O^{2-}  73 140	F F^-  72 136
S S^{2-}  103 184	Cl Cl^-  99 181
Se Se^{2-}  117 198	Br Br^-  114 195
Te Te^{2-}  143 221	I I^-  133 216

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



2nd Ionization – Energy to remove e⁻ from 1+ ion



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
- Increase in nuclear attraction
- Harder to take one away
- Increased IE

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)

Irregularities

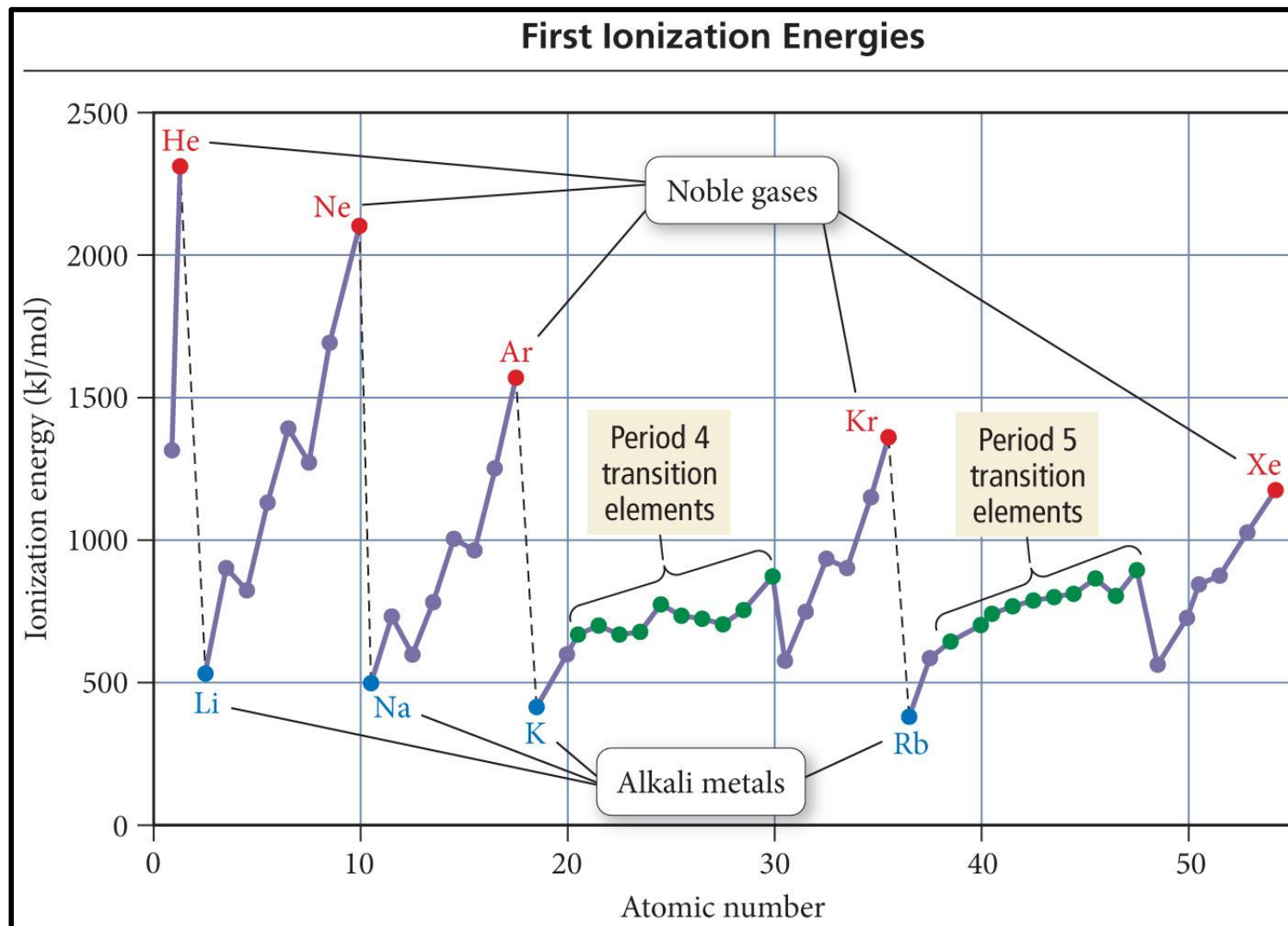
Half filled and totally filled sublevels (orbital set)

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

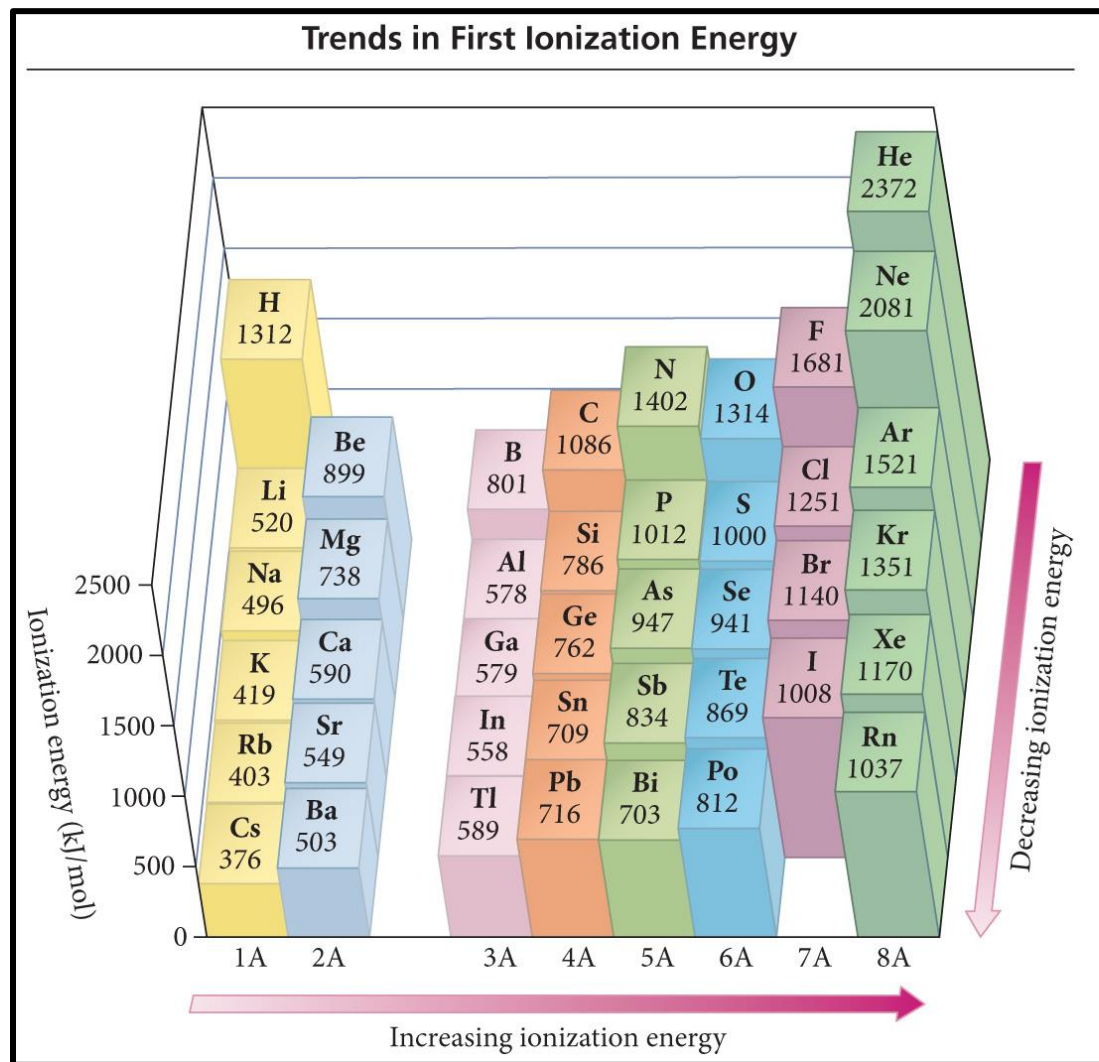
Moving to a p orbital (Mg → Al)

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

Ionization Energy (IE)



Ionization Energy (IE)



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

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

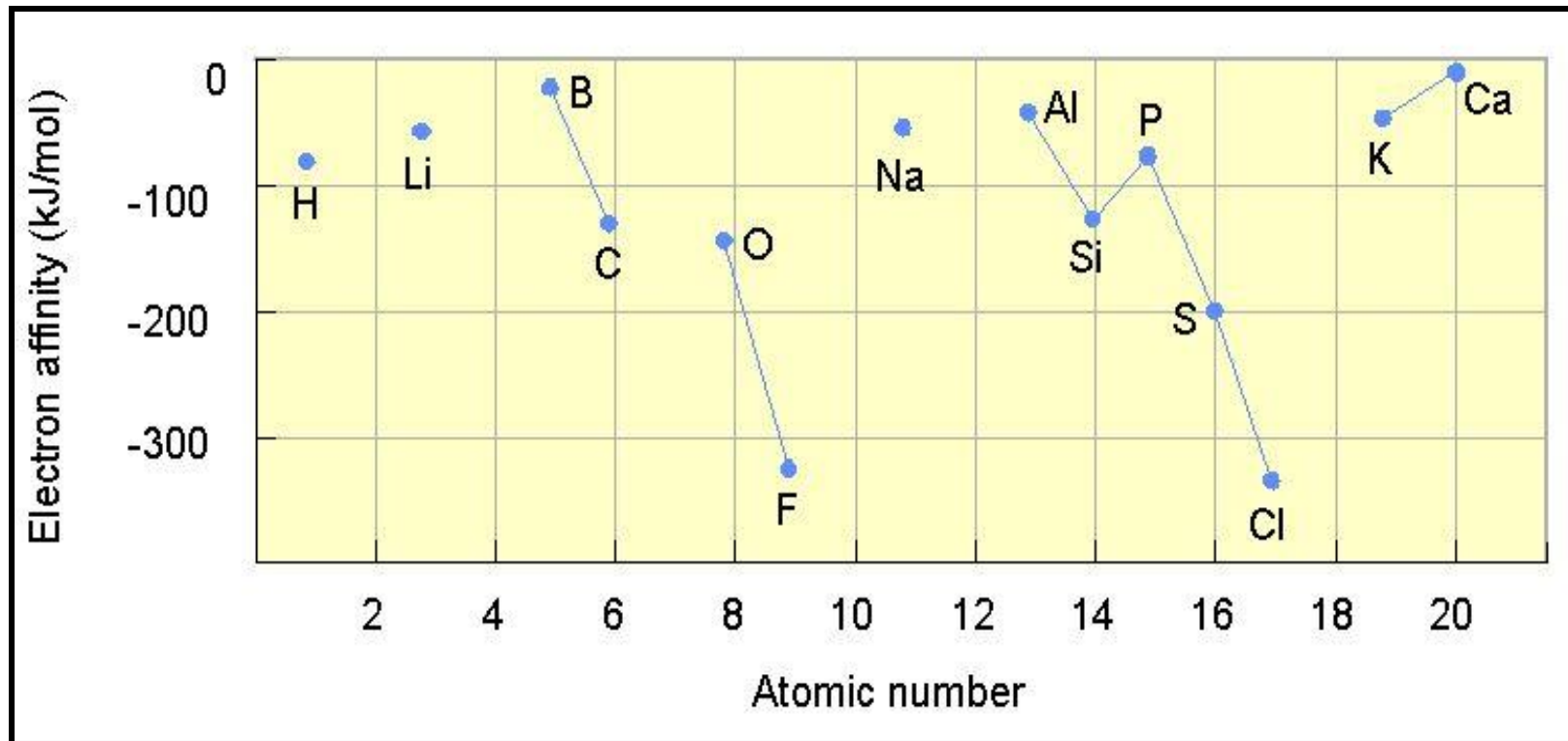
Highest EA in any period = halogen

Very irregular pattern compared to other PT Trends

Electron Affinity

Electron Affinities (kJ/mol)							
1A							8A
H -73							He >0
	2A	3A	4A	5A	6A	7A	
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

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

Increases across period (left to right)

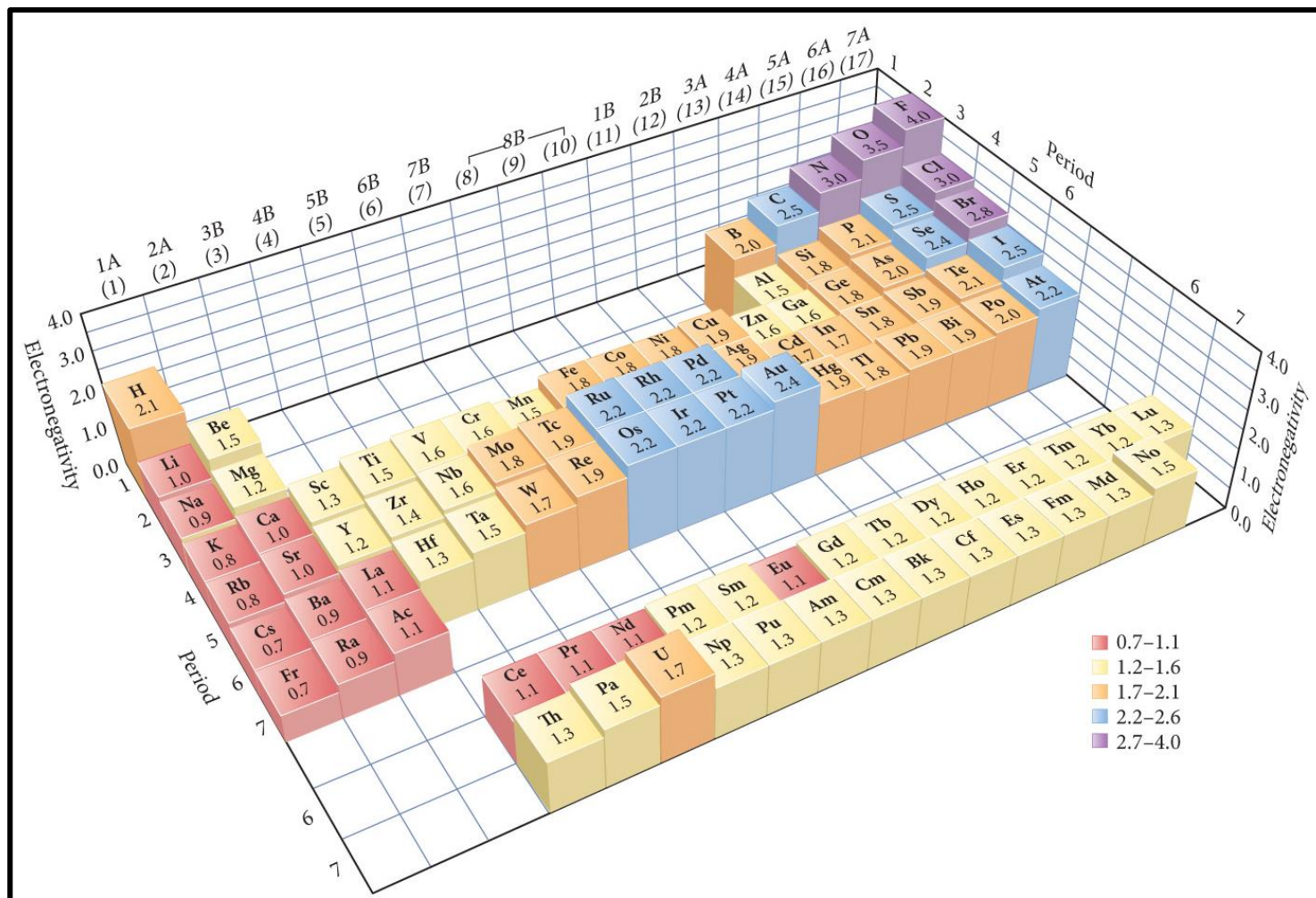
Decreases down group (top to bottom)

- 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

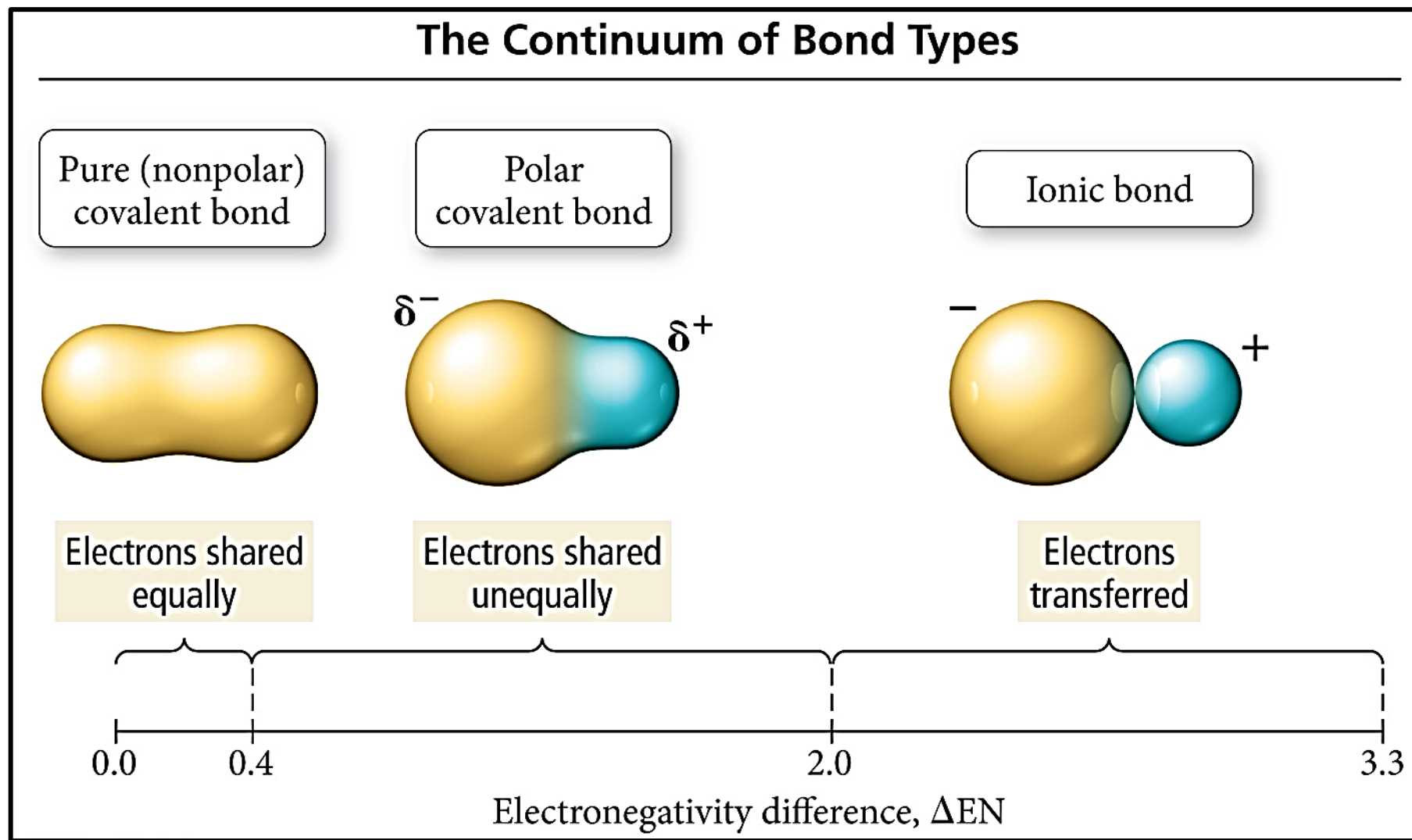
Polar Covalent

- 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 - μ , - 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

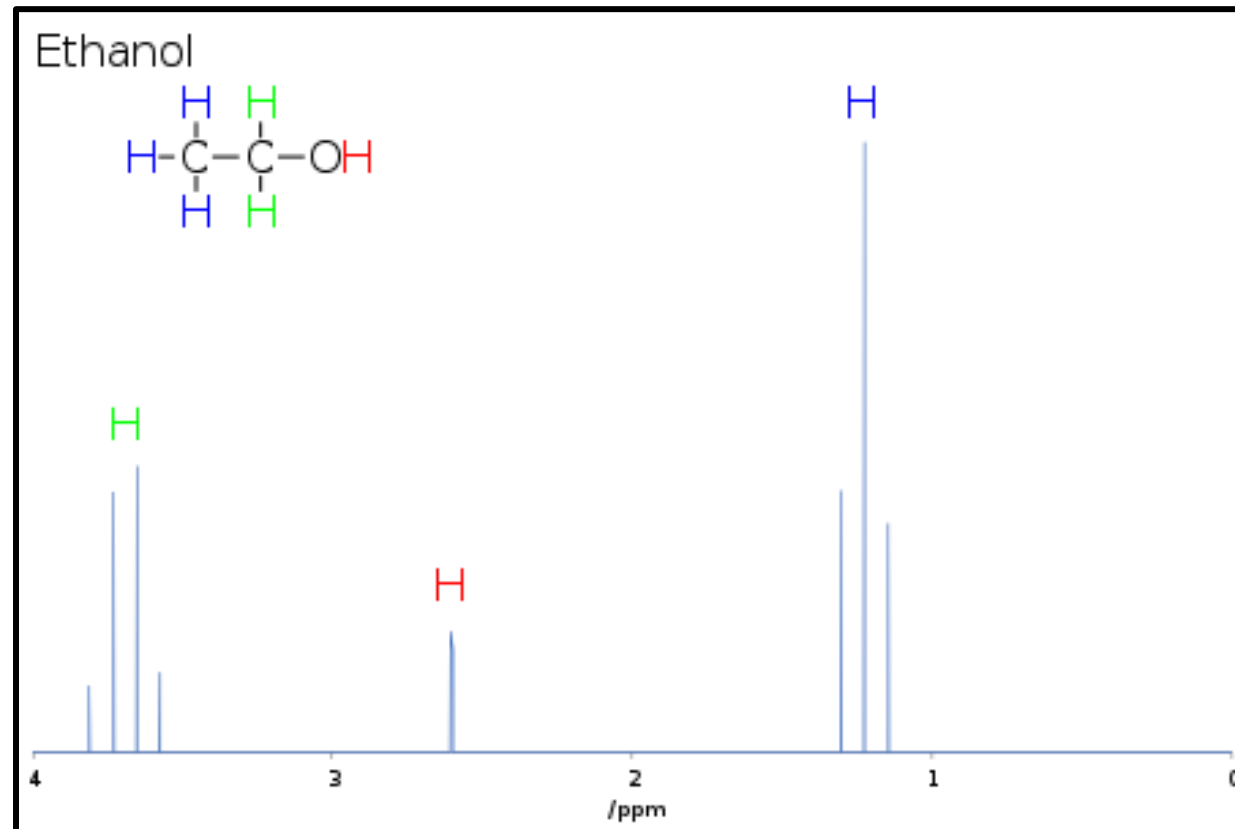
- Result of all paired electrons in orbitals
- Slightly repelled by a magnetic field

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

Why care about Paramagnetism?

NMR Machine – Helps determine the structure of molecules

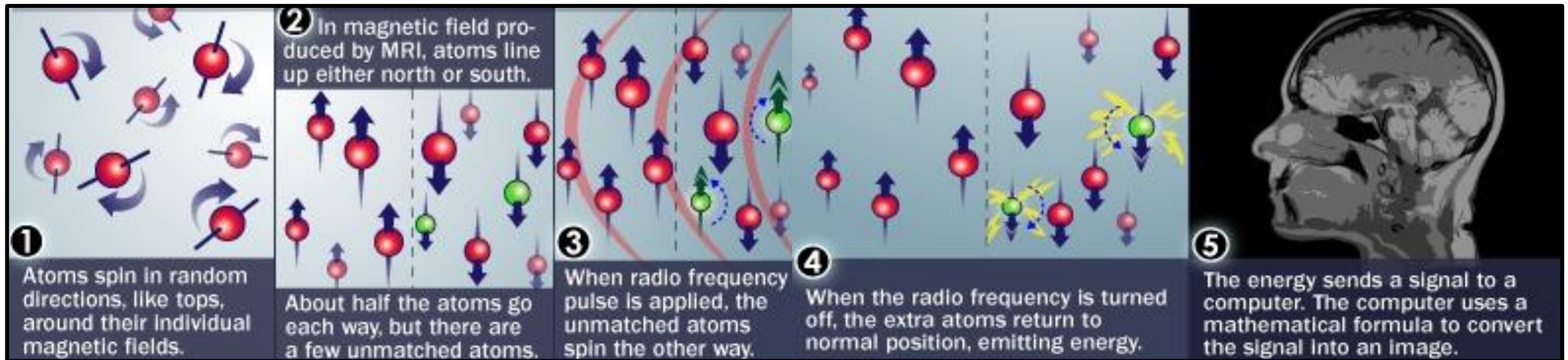
MRI – Applied to images of the body



Why care about Paramagnetism?

NMR Machine – Helps determine the structure of molecules

MRI – Applied to images of the body



Link to YouTube Presentation

<https://youtu.be/RwCDvBtAGbo>