Reference Sheets for
Unit #4 – Periodic Table

Periodic Table Structure Info Sheet

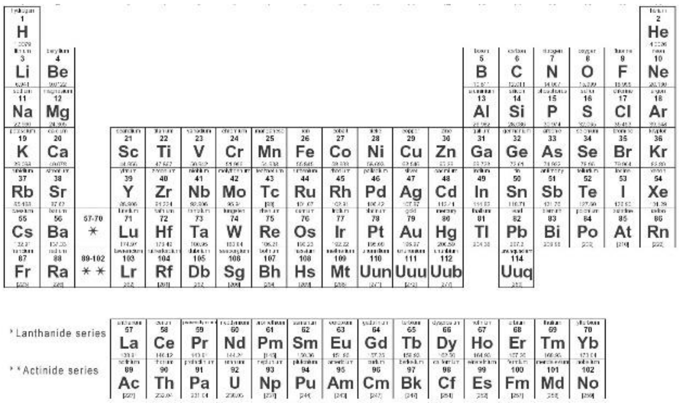
Periods (rows) →

Groups (columns) ↑

Mendeleev – Organized PT based on atomic masses & properties (almost right…)

Moseley – Organized PT based on atomic numbers (the way we do it now!)

Three classes of elements: Metals, non-metals, metalloids/semi-metals



Color code each class of element. Make a key here

Some videos about the structure & creation of the periodic table

- https://tinyurl.com/n4o9dns
- https://tinyurl.com/y7jtlkbw
- https://tinyurl.com/abq96op
- https://tinyurl.com/q2z47cl

Metal Properties:

Chemical Prop.	Physical Prop.
Few electrons in VALENCE shell (outer shell)	Ductile Malleable
Lose electrons easily	Good conductors
POSITIVE charge like Ca^{2+}	Shiny
Make Cations	Solid at room temp

Non-metal Properties:

Chemical Prop.	Physical Prop.
Almost full, or totally full valence shell	NOT Ductile NOT malleable
Tend to gain electrons	BAD conductors
NEGATIVE charge like N^{3-}	Mostly solid
Make ANIONS	Some are gas at room temp

Semi-metal Properties:

Chemical Prop.	Physical Prop.
Most have half full valence shell	Have properties of metals AND non-metals
Make anions OR cations depending on their environment	No way to know which properties of each

Things in the same period have:

Increasing atomic # and mass L→R

Same number of energy levels

Period 1 has 1 level

Period 2 has 2 levels etc…

Things in the same group have:

Increasing atomic # and mass ↓

Same number of valence electrons

Exceptions: d and f block

Similar physical and chemical properties

b/c they have same # of valence e^-s

Valence Electrons:

Outer electrons

Matches the “A” column number

1A has 1 v.e^-, 2A has 2v.e^-, etc.

d and f blocks don’t follow rules

Shielding and Z_{eff}:

Outer electrons have trouble “seeing” the protons in the nucleus – the nucleus is “shielded” by the electrons. You can calculate how much “shielding” there is by calculating the “Effective Nuclear Charge”

$$Z_{eff} = Z - S$$

Z_{eff} = effective nuclear charge

Z = atomic #

S = all non-valence electrons

Periodic Table Structure Info Sheet

- [ ] Alkali metals
- [ ] Alkaline earth metals
- [ ] Transition metals
- [ ] Rare earth metals
- [ ] Metalloids (semi-metals)
- [ ] Other metals
- [ ] Non-metals
- [ ] Halogens
- [ ] Noble Gases

The Periodic Table of the Elements

Element name → Mercury
Atomic # → 80
Symbol → Hg
Avg. Mass → 200.59

1	2		3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
Hydrogen 1 H 1.01																		Helium 2 He 4.00
Lithium 3 Li 6.94	Beryllium 4 Be 9.01												Boron 5 B 10.81	Carbon 6 C 12.01	Nitrogen 7 N 14.01	Oxygen 8 O 16.00	Fluorine 9 F 19.00	Neon 10 Ne 20.18
Sodium 11 Na 22.99	Magnesium 12 Mg 24.31												Aluminum 13 Al 26.98	Silicon 14 Si 28.09	Phosphorus 15 P 30.97	Sulfur 16 S 32.07	Chlorine 17 Cl 35.45	Argon 18 Ar 39.95
Potassium 19 K 39.10	Calcium 20 Ca 40.08		Scandium 21 Sc 44.96	Titanium 22 Ti 47.88	Vanadium 23 V 50.94	Chromium 24 Cr 52.00	Manganese 25 Mn 54.94	Iron 26 Fe 55.85	Cobalt 27 Co 58.93	Nickel 28 Ni 58.69	Copper 29 Cu 63.55	Zinc 30 Zn 65.39	Gallium 31 Ga 69.72	Germanium 32 Ge 72.61	Arsenic 33 As 74.92	Selenium 34 Se 78.96	Bromine 35 Br 79.90	Krypton 36 Kr 83.80
Rubidium 37 Rb 85.47	Strontium 38 Sr 87.62		Yttrium 39 Y 88.91	Zirconium 40 Zr 91.22	Niobium 41 Nb 92.91	Molybdenum 42 Mo 95.94	Technetium 43 Tc (98)	Ruthenium 44 Ru 101.07	Rhodium 45 Rh 102.91	Palladium 46 Pd 106.42	Silver 47 Ag 107.87	Cadmium 48 Cd 112.41	Indium 49 In 114.82	Tin 50 Sn 118.71	Antimony 51 Sb 121.76	Tellurium 52 Te 127.60	Iodine 53 I 126.90	Xenon 54 Xe 131.29
Cesium 55 Cs 132.91	Barium 56 Ba 137.33	57-70 *	Lutetium 71 Lu 174.97	Hafnium 72 Hf 178.49	Tantalum 73 Ta 180.95	Tungsten 74 W 183.84	Rhenium 75 Re 186.21	Osmium 76 Os 190.23	Iridium 77 Ir 192.22	Platinum 78 Pt 195.08	Gold 79 Au 196.97	Mercury 80 Hg 200.59	Thallium 81 Tl 204.38	Lead 82 Pb 207.20	Bismuth 83 Bi 208.98	Polonium 84 Po (209)	Astatine 85 At (210)	Radon 86 Rn (222)
Francium 87 Fr (223)	Radium 88 Ra (226)	89-102 **	Lawrencium 103 Lr (262)	Rutherfordium 104 Rf (267)	Dubnium 105 Db (268)	Seaborgium 106 Sg (271)	Bohrium 107 Bh (272)	Hassium 108 Hs (270)	Meitnerium 109 Mt (276)	Darmstadtium 110 Ds (281)	Roentgenium 111 Rg (280)	Copernicium 112 Cn (285)	Ununtrium 113 Uut (284)	Ununquadium 114 Uuq (289)	Ununpentium 115 Uup (288)	Ununhexium 116 Uuh (293)	Ununseptium 117 Uus (294?)	Ununoctium 118 Uuo (294)

*lanthanides	Lanthanum 57 La 138.91	Cerium 58 Ce 140.12	Praseodymium 59 Pr 140.91	Neodymium 60 Nd 144.24	Promethium 61 Pm (145)	Samarium 62 Sm 150.36	Europium 63 Eu 151.97	Gadolinium 64 Gd 157.25	Terbium 65 Tb 158.93	Dysprosium 66 Dy 162.50	Holmium 67 Ho 164.93	Erbium 68 Er 167.26	Thulium 69 Tm 168.93	Ytterbium 70 Yb 173.04
**actinides	Actinium 89 Ac (227)	Thorium 90 Th 232.04	Protactinium 91 Pa 231.04	Uranium 92 U 238.03	Neptunium 93 Np (237)	Plutonium 94 Pu (244)	Americium 95 Am (243)	Curium 96 Cm (247)	Berkelium 97 Bk (247)	Californium 98 Cf (251)	Einsteinium 99 Es (252)	Fermium 100 Fm (257)	Mendelevium 101 Md (258)	Nobelium 102 No (259)

Periodic Trends Info Sheet

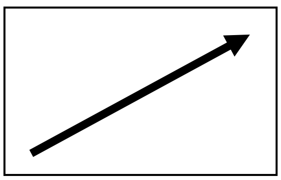
IONIZATION ENERGY

The minimum amount of energy required to *remove* the most loosely bound electron, the valence electron, of an isolated neutral gaseous atom to form a cation. It is quantitatively expressed in symbols as:

$$X + \text{energy} \rightarrow X^{+} + e^{-}$$

In other words…
How much energy does it take to take away an electron from an atom?

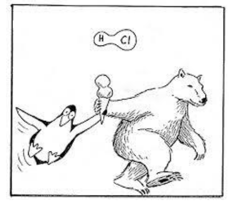
Trend
Increases →
Increases ↑



Highest = Fluorine only higher are Noble Gases (Rn)
Lowest = Francium

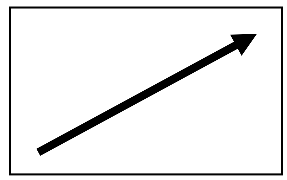
ELECTRONEGATIVITY

The tendency of an atom to attract a shared pair of electrons (or electron density) towards itself.



In other words…
How strong is the atom? How hard can it pull on electrons when sharing them with another atom?

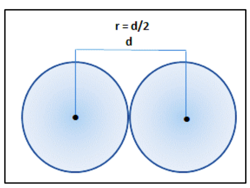
Trend
Increases →
Increases ↑



Highest = Fluorine
Lowest = Francium only lower are Noble Gases (He)

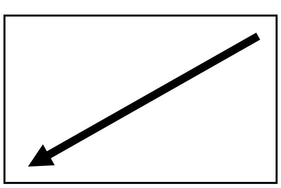
ATOMIC RADIUS

Usually the mean or typical distance from the center of the nucleus to the boundary of the surrounding cloud of electrons. Since the boundary is not a well-defined physical entity, there are various non-equivalent definitions of atomic radius.



In other words…
Half the diameter of an atom. Hard to measure because atoms do not have actual tangible edges. Lots of ways to measure it.

Trend
Increases ←
Increases ↓



Highest = Francium
Lowest = Helium

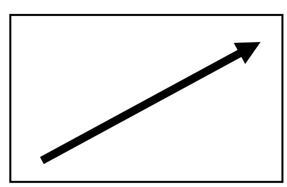
ELECTRON AFFINITY

The amount of energy released* when an electron is *added* to one mole of a neutral atom, or molecule, in the gaseous state to form a negative ion. Usually written as a ΔE value.

$$X + e^{-} \rightarrow X^{-} + \text{energy}$$

In other words…
How happy is the atom to gain that new electron? Really happy – much more stable, releases lots of energy. Not happy – not as much stability, doesn't release lots of energy.

Trend
Increases →
Increases ↑



Highest = Fluorine
Lowest = Francium only lower are Noble Gases (zero)

*Some electron affinities are actually positive ΔE – meaning energy is absorbed. They are not commonly talked about or used though.

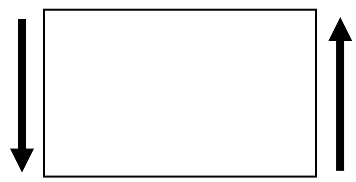
Periodic Trends Info Sheet

REACTIVITY

The impetus for which a chemical substance undergoes a chemical reaction, either by itself or with other materials, with an overall release of energy.

In other words…
How quickly, violently, readily, does an element undergo certain reactions. More reactivity means faster, more violent, easier reaction with lots of energy released.

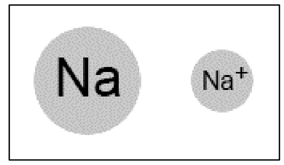
Trend
Metals
Increase ↓←
Non-metals
Increases ↑→



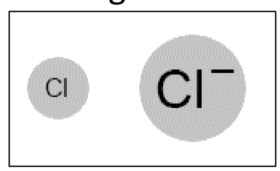
Highest metal = Francium
Highest non-metal = Fluorine
Lowest non-metal = Noble gases (He)

IONIC RADIUS

Cations:
Get smaller when they lose electrons



Anions:
Get larger when they gain electrons



EFFECTIVE NUCLEAR CHARGE - Zeff

The pull that the nucleus has on electrons

$$Zeff = Z - S$$

Z = # of protons
S = Core/Inner electrons (# of electrons in previous noble gas + any d or f electrons)

SUBSEQUENT IONIZATIONS

1st ionization energy
$X \rightarrow X^{+} + e^{-}$
2nd ionization energy
$X^{+} \rightarrow X^{2+} + e^{-}$
3rd ionization energy
$X^{2+} \rightarrow X^{3+} + e^{-}$

Ionization energy increases with each subsequent ionization because there is more attraction between the nucleus and the valence electrons each time you lower the number of valence electrons.

There is a huge leap in ionization energy once an atom loses all it's valence electrons because it now looks like a noble gas and really doesn't want to let go of any more electrons!

SHIELDING

When the inner/core electrons repel the valence electrons and prevents them from seeing the nucleus. Decreases how strongly the electrons are held onto by the nucleus.

BREAKS IN PATTERNS

There are *many* examples of elements that do not appear to follow the general trends typically described. This can be due to a variety of reasons. Here are two of many reasons why this can happen.

- Half-filled orbitals have slightly more stability than expected
  Example: p orbital set: ↑ ↑ ↑
- Unpairing an electron takes slightly less energy than expected b/c extra electron repulsions
  Example: p orbital set: ↑↓ ↑ ↑

Warning…

- Don’t over think this stuff.
- You can talk yourself into backwards answers.
- Focus on the fact that there are only a set number of trends to learn.
- Practice explaining each trend until you can do it in your sleep!
- There will ALWAYS be exceptions. Don’t worry about that – focus on the pattern and answer questions based on the patterns.

Warning…

- There is about to be a lot of notes because it takes a lot of words to explain
- You don’t need EVERY word written down to understand it.
- Focus on listening and understanding.
- You can add to your notes at home.
- Capture enough to pay attention, leave space to come back and add/annotate.

Warning…

<u>Make sure you capture:</u>

What

How

Why

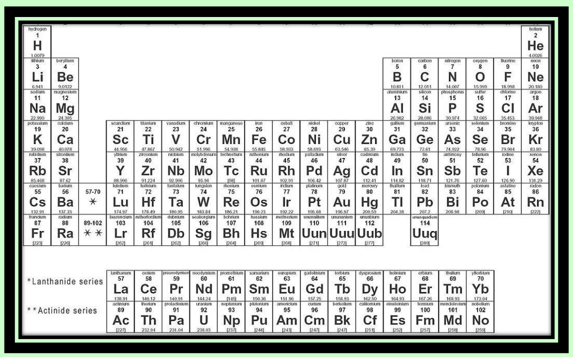
<u>Make sure you can tell me:</u>

What

How

Why

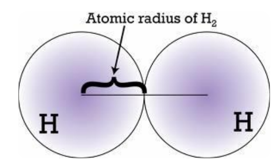
Periodic Trends

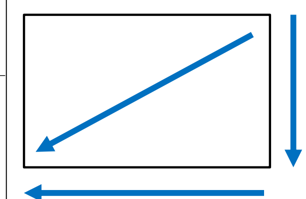


Atomic Radius

ATOMIC RADIUS

What	How
• ½ the distance between two bonded nuclei • Cant measure to the edge b/c orbitals aren’t tangible!	



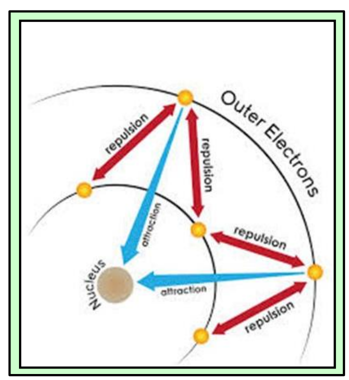


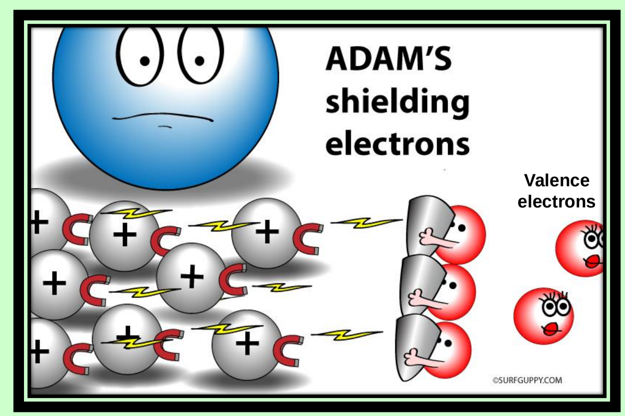
ATOMIC RADIUS

Why	
INCREASES DOWN	**DECREASES TO RIGHT**
• Adding energy levels Smaller “effective nuclear charge”	• Adding a proton = bigger change than adding an e-
• Inner e- keep valence e- from “feeling” the nucleus	• More protons pull the valence electrons in closer - “Greater Effective Nuclear Charge”
• Outer e-s are not as pulled in by the protons in the nucleus – there is more “shielding” by the inner electrons	• No increase in shielding b/c no new energy levels

<u>Effective Nuclear Charge (Z_{eff})</u>
The relative attraction the valence electrons have for the protons in the nucleus

<u>Shielding Effect</u>
The inner shell electrons repel the outer valence electrons – keeps the valence e- from “feeling” the nucleus. More energy levels - lowers the effective nuclear charge





Calculating Effective Nuclear Charge

The relative attraction the valence electrons have for the protons in the nucleus

$$Z_{eff} = Z - S$$

Z = nuclear attraction = # protons
S = the core/inner e- shielding the valence e-'s
= the total number of e- minus the e- in the highest occupied s and p energy levels
= *(# of e- in previous noble gas + any d or f e-'s past the noble gas in the element)*

Calculating Effective Nuclear Charge

$Z_{eff} = Z - S$

Magnesium
Z = 12 protons
S = Previous noble gas
= Neon = 10 electrons
Zeff = 12 - 10 = 2

Aluminum
Z = 13 protons
S = Previous noble gas
= Neon = 10 electrons
Zeff = 13 – 10 = 3

Aluminum is smaller – valence electrons are pulled in harder by the nucleus

IONIC RADIUS

What	*How*
The radius of an ion	Cation – always smaller
Cation – lost electrons Anion – gained electrons	Anion – always bigger

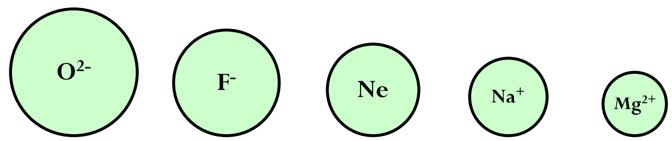
IONIC RADIUS

Why	
CATION SMALLER • Reduced repulsion between electrons • If you lose enough electrons you even drop down an energy level! Much smaller!	**ANION LARGER** • Extra valence electrons repel each other a bit more so it gets larger.

Isoelectric Species

Atoms/Ions that have the same number of e-

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

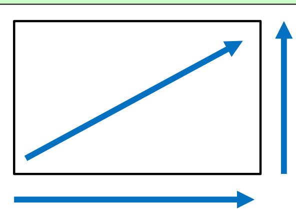


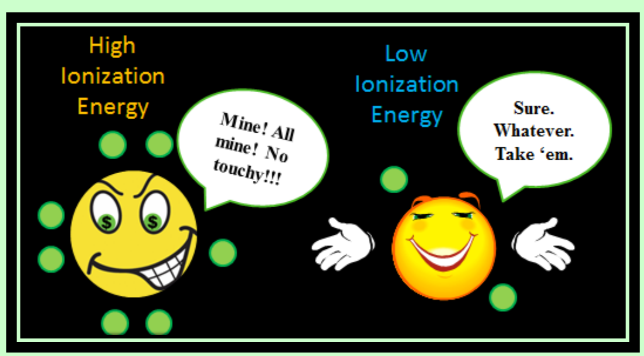
8 protons | 9 protons | 10 protons | 11 protons | 12 protons

Increased protons can pull harder on the valence electrons – greater effective nuclear charge – so the radius is smaller even though they have the same number of electrons and energy levels

Ionization Energy

IONIZATION ENERGY

What	*How*
The energy required to remove on electron from a neutral atom of an element	 Noble Gases are HIGHEST! They REALLY don't want to let go of an e-



IONIZATION ENERGY

Why	
DECREASES DOWN • Electrons are further from nucleus in higher energy levels • Increased shielding from core e- causes nucleus to not pull as hard on valence e- • Lower effective nuclear charge –so they are more easily removed	**INCREASES TO RIGHT** • Closer to having a full stable valence shell • Increased effective nuclear charge means nucleus is pulling harder on the valence e- so they are harder to remove

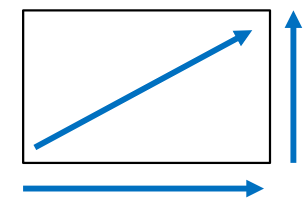
Subsequent Ionizations

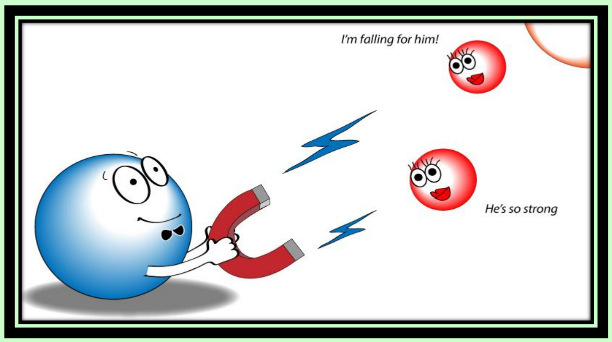
Every time you take an e- away it gets harder to take the next one. Radius is getting smaller, so nucleus can pull harder on the valence - harder to remove the next one. HUGE LEAP in I.E. once it's achieved noble gas configuration – why would it want to lose another one?!

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

Electronegatvity

ELECTRONEGATIVITY

What	How
A measure of the ability of an atom in a chemical compound to attract electrons from another atom in the compound *How strongly can one atom pull on the electrons being shared in a bond.*	 Noble Gases are LOWEST! They DON'T CARE about attracting electrons!



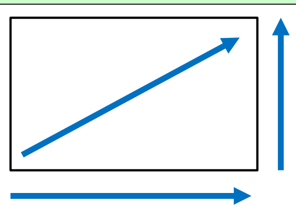
ELECTRONEGATIVITY

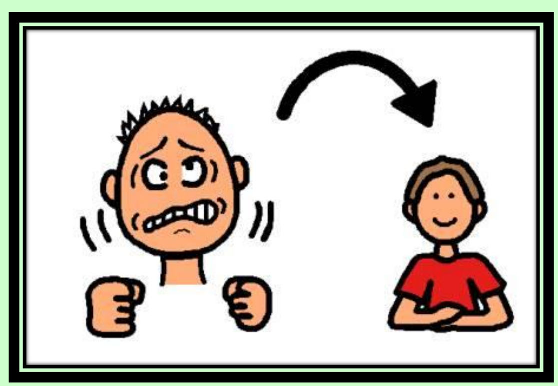
Why

DECREASES DOWN	INCREASES TO RIGHT
• e- are further from nucleus in higher energy levels • Increased shielding from core e- causes nucleus to not pull as hard on valence e- • So nucleus doesn't pull as hard on the bonding e-'s from another atom	• Smaller radius, increased effective nuclear charge • Nucleus is pulling harder on the valence electrons – which is where the bonding is occurring.

Electron Affinity

ELECTRON AFFINITY

What	How
How much energy is released when the atom gains an electron to make a negative ion. *How much stability does it gain once it is an anion. More energy released – more stable.*	 Noble Gases are LOWEST! They DON'T CARE about attracting electrons!



ELECTRON AFFINITY

Why

DECREASES DOWN	INCREASES TO RIGHT
• Electrons are further from nucleus in higher energy levels • Increased shielding from core e-'s causes the nucleus to not pull as hard on valence e-'s • So atom doesn't notice as much if it gains an electron – doesn't gain much stability	• Closer to filling valence shell – noble gas configuration is most stable

Reactivity

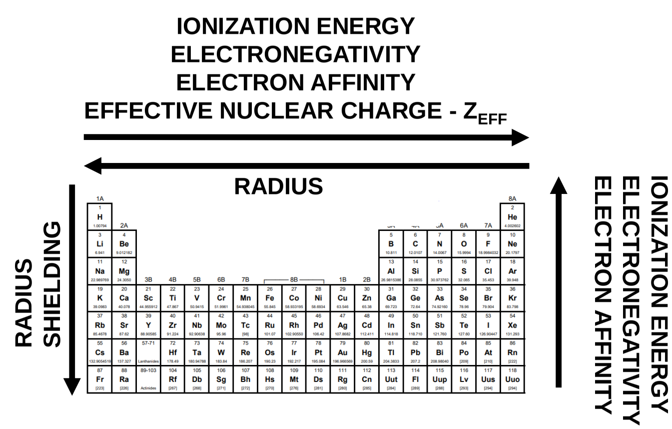
REACTIVITY

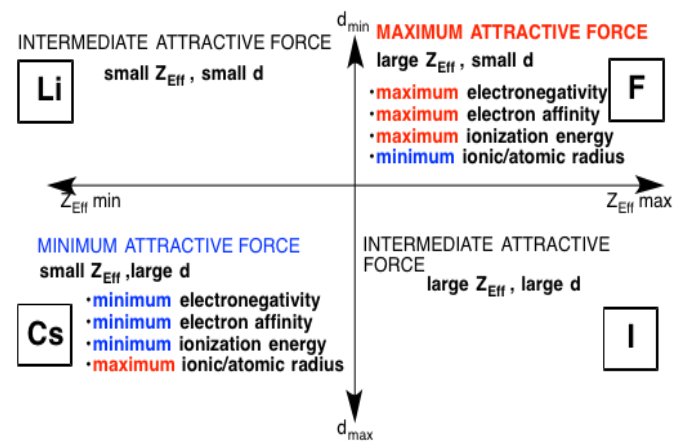
What	How
Elements in the same group have similar types of behaviors <u>because they have the same number of valence e-</u> BUT The MAGNITUDE of their reactions changes!	

REACTIVITY

Why	
METALS INCREASE DOWN • Larger radius and increased shielding means lower ionization energy so it is easier to remove electrons	**NON-METALS INCREASE UP** • Smaller radius and greater effective nuclear charge means higher electronegativity and electron affinity so it can attract an electron easier

Summary





Brainiac Video – note: they augmented the reactions, but it is such a fun, silly, memorable video I think it is still worth watching ☺

Disposal of Sodium – old footage from WWII. Neat to see such old footage and how they actually disposed of the sodium after the war!

Quick summary. Also has a quick but good explanation of some <u>exceptions</u> to the trends

https://www.youtube.com/watch?v=hePb00CqvP0