Ms. Randall
Regents Chemistry
Unit 4: Periodic Table Unit Notes

Unit Objectives:

- Describe the origin of the periodic table
- State the modern periodic law
- Explain how an element's electron configuration is related to the element's placement within a period and a group on the periodic table.
- State the trends of the following properties within periods and groups of elements including:
    - Ionization energy
    - Electronegativity
    - Atomic Radius
    - Reactivity
    - Metallic/Nonmetallic character
- Identify and state the properties of the following groups in the periodic table:
    - Alkali metals
    - Alkaline earth metals
    - Halogens
    - Noble Gases
    - Transition elements
- Locate within the periodic table and state the properties of the metals, nonmetals, and metalloids (semi-metals)

<u>Define the following vocabulary:</u>

Allotrope	Nonmetals	Mendeleev
Ion	Metalloids	Period
Cation	Luster	Group(family)
Anion	Malleability	Alkali Metals
Electron	Ductility	Alkaline Earth Metals
Proton	Conductivity	Halogens
Neutron	Nonmetals	Atomic radius
Compound	Brittleness	Ionization energy
Element	Dull	Electronegativity
Valence electron	Non-conductor	Reactivity
Lewis Dot Diagram	Noble gas	Electron configuration
Metals	Periodic Law	

Lesson 1: The History of the Table and Chemical Periodicity

Objective:

- ***Explain how the periodic table was developed***
- ***Identify the differences between periods and groups***

Watch the following clip:

History & structure of the periodic table

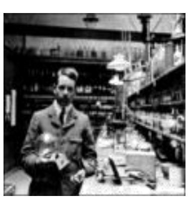
Dmitri Mendeleev (Russia)

Between 1868 and 1870, in the process of writing his book, *The Principles of Chemistry*, Mendeleev created a table or chart that listed the known elements according to increasing order of atomic weights. When he organized the table into horizontal rows, a pattern became apparent--but only if he left blanks in the table. If he did so, elements with similar chemical properties appeared at regular intervals--periodically--in vertical columns on the table.

Mendeleev was bold enough to suggest that new elements not yet discovered would be found to fill the blank places. He even went so far as to predict the properties of the missing elements. Although many scientists greeted Mendeleev's first table with skepticism, its predictive value soon became clear. The discovery of gallium in 1875, of scandium in 1879, and of germanium in 1886 supported the idea underlying Mendeleev's table. Each of the new elements displayed properties that accorded with those Mendeleev had predicted based on his realization that elements in the same column have similar chemical properties. The three new elements were respectively discovered by French, a Scandinavian, and a German scientist, each of whom named the element in honor of his country or region. (Gallia is Latin for France.) Discovery of a new element had become a matter of national pride--the rare kind of science that people could read about in newspapers, and that even politicians would mention.

- 1st chemist to arrange newly found elements into a table form/usable manner
- Elements arranged according to **Atomic mass**
- Resulted in **gaps** or periodic intervals being **out of order**

Periodic = cyclic; repeating patterns/cycles; similar to monthly/weekly calendar (days of the week)
Ex: tired on Mondays, happy on Fridays



Henry Moseley (England)
By 1907, when Mendeleev died, chemists were sure that iodine followed tellurium in the Periodic Table and that there was something odd about their relative atomic masses. However no-one was able to *measure* atomic number, it was just the position of an element in the Periodic Table sequence. For example lithium was known to be the third element but this number three was only because its properties meant that it slotted in between helium and beryllium. **Henry Moseley** found and measured a property linked to Periodic Table position. Hence atomic number became more meaningful and the three pairs of elements that seemed to be in the wrong order could be explained. Moseley used what was then brand-new technology in his experiments. A device now called an electron gun had just been developed. He used this to fire a stream of electrons (like machine gun bullets) at samples of different elements. He found that the elements gave off X-rays. (This is how the X-rays used in hospitals are produced.) Moseley measured the frequency of the X-rays given off by different elements. Each element gave a different frequency and he found that this frequency was

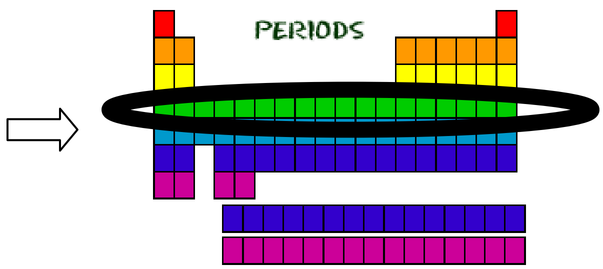
mathematically related to the position of the element in the Periodic Table – he could actually measure atomic number!

- Arranged table by **Atomic number** (or # of protons) which proved to be much more effective
- How the **modern day** periodic table is arranged

Periodic Law = elements in periodic table are **periodic** functions of their **atomic number**

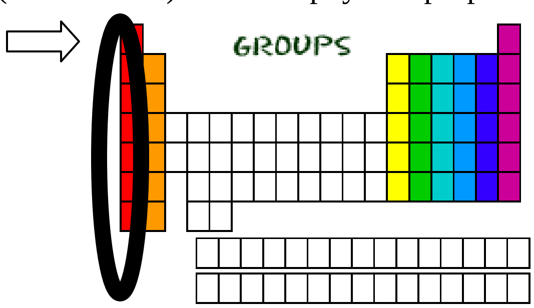
How is the table actually arranged?

Periods = **Horizontal rows** (run left to right) on Periodic Table



The period number tells us the number of **electron shells** (principal energy levels)
The properties of elements change drastically across a period from metals → metalloids/semi-metals → nonmetals.
The number of **valence electrons** increases from left to right (1 → 8)

2. **Groups (Families)** = **vertical columns** (run up & down) on Periodic Table; each group contains the same number of **valence electrons and similar** (not identical) chemical/physical properties.



Example: K in Group 1
Let's look at the electron configuration of elements in the same group…

H = **1**
Li = **2-1**
Na = **2-8-1**
K = **2-8-8-1**
Rb = **2-8-18-8-1**
Cs = **2-8-18-18-8-1**
*Fr = **-18-32-18-8-1**

What similarities can you observe within the above electron configurations?

All have **1** valence electron

Remember…

Group # = **number of valence electrons**

Period number = **number of principle energy levels(shells)**

Why do elements in the same group have similar chemical/physical properties?

- They have the same **number of valence electrons**
- Valence electrons affect **Reactivity**

Reactive elements can bond *easily with other elements. They have an* ***incomplete valence electron shell****.*

- All atoms (except hydrogen and Helium) want **8 electrons** in their valence shell (outermost energy level)

Most elements, except noble gases, combine to form compounds. Compounds are the result of the formation of chemical bonds between two or more different elements. In the formation of a chemical bond, atoms lose, gain or share valence electrons to complete their outer shell and attain a noble gas configuration. This tendency of atoms to have eight electrons in their outer shell is known as the ***octet rule***.

An **ion** (charged particle) can be produced when an **atom gains** or **losses** one or more **electrons**.

- A **cation** (**+ ion**) is formed when a neutral atom **loses an electron**. Metals tend to form cations.
- An **anion** (ion) is formed when a neutral atom **gains** an **electron**. Nonmetals tend to form anions.

Isoelectronic: *atoms or ions that have the* ***SAME*** *number of* ***ELECTRONS***

Ex: F^-, Ne, and Na^+ all have **10** electrons

Octet = full **valence shell** (8 electrons, except for period 1 elements….they need 2 to have a full valence shell

*******KNOW THIS!!!!*******

Octet Rule

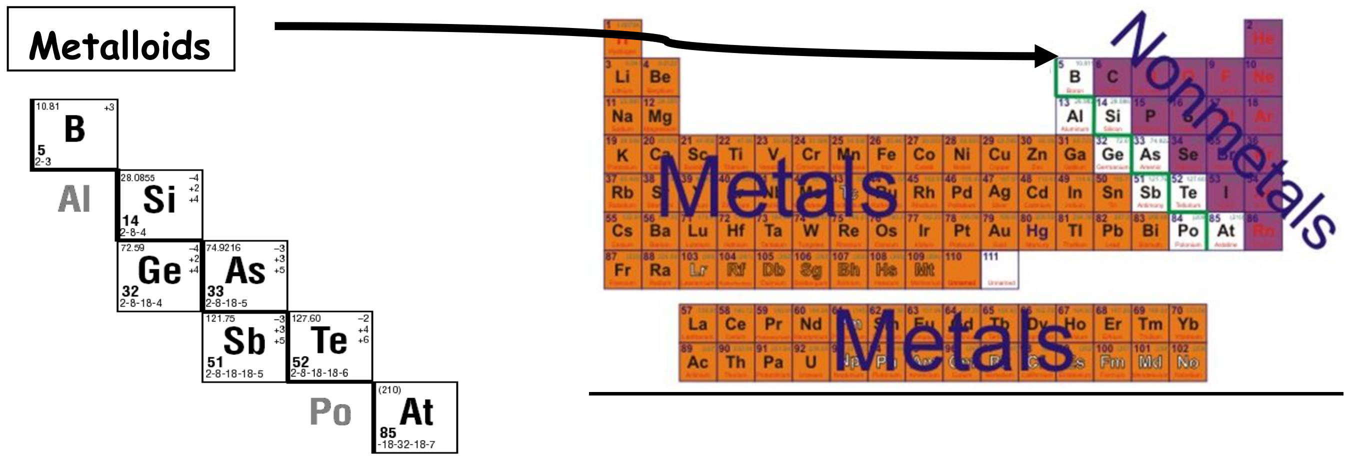
Atoms tend to gain, lose or share one or more of their valence electrons to achieve a filled outer electron shell

Lesson 2: Key to the Periodic Table

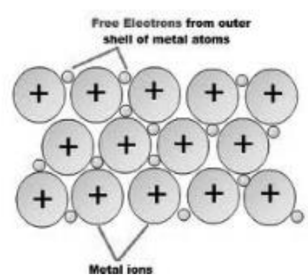
Objective:
- ***Differentiate between the different groups of elements***
- ***Identify the properties specific to each category of element***

Check this out! **Interactive Periodic Table**

The structure of the periodic table gives us a lot of information. Depending on an element's position in the table, we can determine a lot of its physical and chemical properties.



1. **Metals:**
    - all **SOLIDS** (except **Hg** $_{(l)}$)
    - **MALLEABLE** (can be hammered/molded into sheets)
    - **DUCTILE** (can be drawn/pulled into wire)
    - Have **LUSTER** (are shiny when polished)
    - Good **CONDUCTORS** (allow heat & electricity to flow through them)
        - due to "**sea of MOBILE** valence electrons"
    - Like to **LOSE** e^- to form **POSITIVE** ions….why? **TO HAVE A FULL VALENCE SHELL OF ELECTRONS**



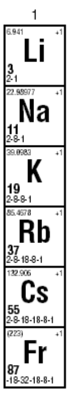
2. **Nonmetals:**
    - mostly **GASES** and **SOLIDS** @ STP—except **Br(l)**
    - **NOT** malleable/ductile; **BRITTLE** (shatter easily)
    - **LACK** luster (**DULL**)
    - **NON** or **POOR** conductors
    - like to **GAIN** e^- to form **NEGATIVE** ions

3. **Metalloids** (AKA semi-metals):
    - have properties of both **<u>METALS</u>** & **<u>NONMETALS</u>**
    - **<u>ALONG</u>** staircase (between **<u>METALS</u>** & **<u>NONMETALS</u>** on table)—above and below except **<u>Al & Po</u>**

SUMMARY OF THE CATEGORIES OF ELEMENTS

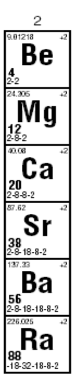
	Metals	Metalloids	Nonmetals
Phys. prop.	• malleable • ductile • shiny • excellent conductors (heat, electricity... MOBILE e-'s)	in-between	• brittle • dull • poor conductors (heat, electricity)
Chem. prop.	• lose e-'s • form + ions • low E.N. • low I.E.	B, Si, Ge, As, Sb, Te	• gain e-'s • form - ions • high E.N. • high I.E.

<u>The Groups in more detail</u>



Group 1 → <u>ALKALI METALS</u> (FAMILY)

- **All have <u>1</u> valence electron**
- Easily <u>LOSE</u> their one electron to become <u>+1</u> ions
- **React violently with water**
- <u>EXTREMELY</u> reactive → never found alone in nature
- Contains the <u>MOST</u> reactive metal: Probably <u>FRANCIUM (Fr)</u>, but it's so rare, we've got to go with <u>CESIUM (Cs)</u>



Group 2 → <u>ALKALINE EARTH METALS</u> (FAMILY)

- **All have <u>2</u> valence electrons**
- Prefer to <u>LOSE</u> their two electrons to become <u>+2</u> ions
- **<u>FAIRLY</u> reactive → never found alone in nature**

Groups 3-12 → TRANSITION METALS

- Found in the MIDDLE of the table (the D block)
- **Form COLORED IONS in solution** (ex: Cu is bright blue when dissolved in water)
- Tend to be UNPREDICTABLE → will lose electrons or gain them depending on what other METALS are present
- LEAST reactive group of metals

Groups 13-16 → BCNO groups (not a single group)

MISCELLANEOUS groups

- Metals, nonmetals, & metalloids found along the staircase (many different properties)

Group 17 → HALOGENS (FAMILIY)

- **7 valence electrons**
- Like to gain 1 electron to become ions with -1 charge (8 is great!)
- **Form SALTS/COMPOUNDS called HALIDES**
- Contains the most (RE)ACTIVE nonmetal: FLUORINE (F)
- All NONMETALS making up the group
- Three states of matter found in group: SOLID (s), LIQUID (l), GAS (g)
  Ex: Chlorine (g), Iodine(s), Bromine(l)

Group 18 → NOBLE GASES (FAMILY)

- **UNREACTIVE or INERT**
- **Have OCTET** (8 e- in valence shell/outer energy level)
- Most STABLE group; exist ALONE in nature. Non reactive!
- Exception to the OCTET is He (only has 2 valence e-)
- EVERYONE WANTS TO BE A NOBLE GAS & HAVE 8 ELECTRONS! 8 IS GREAT!
  Ex: Neon (Ne)

Hydrogen → Not officially part of a group

- Both a NONMETAL and a GAS
- Can be seen as H2(g), H+(aq) or H-(aq)

The Lanthanide/Actinide Series

- Two rows on bottom of table (detached) – Elements 58 – 71 & 90 - 103
- Actually belong to the TRANSITION METALS

Lesson 3: Periodic Trends

Objective:

- ***Describe the trend in atomic radius, ionization energy and electronegativity***
- ***Explain why this trend in atomic radius exists***

Periodic trends are specific patterns that are present in the periodic table that illustrate different aspects of a certain element, including its size and its electronic properties. Major periodic trends include: electronegativity, ionization energy, atomic radius, and metallic character. Periodic trends, arising from the arrangement of the periodic table, provide chemists with an invaluable tool to quickly predict an element's properties. These trends exist because of the similar atomic structure of the elements within their respective group families or periods, and because of the periodic nature of the elements.

1. Electronegativity Trends

Electronegativity can be understood as a chemical property describing an atom's ability to attract and bind with electrons. Because electronegativity is a qualitative property, there is no standardized method for calculating electronegativity. Electronegativity values for each element can be found on reference table S.

Electronegativity measures an atom's tendency to attract and form bonds with electrons. This property exists due to the electronic configuration of atoms. Most atoms follow the octet rule (having the valence, or outer, shell comprise of 8 electrons). Because elements on the left side of the periodic table have less than a half-full valence shell, the energy required to gain electrons is significantly higher compared with the energy required to lose electrons. As a result, the elements on the left side of the periodic table generally lose electrons when forming bonds. Conversely, elements on the right side of the periodic table are more energy-efficient in gaining electrons to create a complete valence shell of 8 electrons. The nature of electronegativity is effectively described thus: the more likely an atom is to gain electrons, the more likely that atom will pull electrons toward itself.

- **From left to right across a period of elements, electronegativity increases.** If the valence shell of an atom is less than half full, it requires less energy to lose an electron than to gain one. Conversely, if the valence shell is more than half full, it is easier to pull an electron into the valence shell than to donate one.
- **From top to bottom down a group, electronegativity decreases.** This is because atomic number increases down a group, and thus there is an increased distance between the valence electrons and nucleus, or a greater atomic radius.
- **Important exceptions of the above rules include the noble gases, lanthanides, and actinides.** The noble gases possess a complete valence shell and do not usually attract electrons. The lanthanides and actinides possess more complicated chemistry that does not generally follow any trends. Therefore, noble gases, lanthanides, and actinides do not have electronegativity values.
- **As for the transition metals, although they have electronegativity values, there is little variance among them across the period and up and down a group.** This is because their metallic properties affect their ability to attract electrons as easily as the other elements.

According to these two general trends, the *most electronegative element is **fluorine***, with 4.0.

Summary:

- **DESIRE** to **GAIN** e^-
- **GREEDINESS** of an atom/ion for e^- (values for each element listed in Table S)

Electronegativity values range from 0.0 to 4.0

- **The MOST electronegative element on the Periodic table is FLUORINE (4.0)**
- The **LEAST** electronegative elements on the Periodic table are **CAESIUM (Cs)** or **FRANCIUM (Fr) (0.7)**

a. **Going down a group, electronegativity DECREASES**

Reasons:

- Add one energy level → **Inner** shells **SHIELD** the **NUCLEUS** from the **VALENCE** electrons
- Harder for **NUCLEUS** to attract additional e-

b. Going across a period, electronegativity INCREASES

Reason: heading across a period you are reaching the **OCTET** so desire to **GAIN** electrons increases (8 is great!)

2. Ionization Energy Trends

Ionization energy is the energy required to remove an electron from a neutral atom in its gaseous phase. Conceptually, ionization energy is the opposite of electronegativity. The lower this energy is, the more readily the atom becomes a cation. Therefore, the higher this energy is, the more unlikely it is the atom becomes a cation. Generally, elements on the right side of the periodic table have higher ionization energy because their valence shell is nearly filled. Elements on the left side of the periodic table have low ionization energies because of their willingness to lose electrons and become cations. Thus, ionization energy increases from left to right on the periodic table.

Another factor that affects ionization energy is *electron shielding*. Electron shielding describes the ability of an atom's inner electrons to shield its positively-charged nucleus from its valence electrons. When moving to the right of a period, the number of electrons increases and the strength of shielding increases. As a result, it is easier for valence shell electrons to ionize, and thus the ionization energy decreases down a group.

- The ionization energy of the elements within a period generally increases from left to right. This is due to valence shell stability.
- The ionization energy of the elements within a group generally decreases from top to bottom. This is due to electron shielding.
- The noble gases possess very high ionization energies because of their full valence shells. Note that helium has the highest ionization energy of all the elements.

Some elements have several ionization energies; these varying energies are referred to as the first ionization energy, the second ionization energy, third ionization energy, etc. The first ionization energy is the energy required to remove the outermost, or highest, energy electron, the second ionization energy is the energy required to remove any subsequent high-energy electron from a gaseous cation, etc.

Generally, any subsequent ionization energies (2nd, 3rd, etc.) follow the same periodic trend as the first ionization energy.

Summary:

Ionization Energy = amount of **ENERGY** needed to **REMOVE** the most **LOOSELY** bound e^- from and atom/ion in the **GAS** phase (values for each element listed in Table S)

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

Metals → like to lose e^- (to get full valence shell) → **LOW I.E.**
Nonmetals → like to gain e^- (to get full valence shell) → **HIGH I.E.**

a. Going **down a group, ionization energy DECREASES**

Reasons:

- Add one energy level → **INNER** shells **SHIELD** the **NUCLEUS** from the **VALENCE** electrons

b. **Going across a period, ionization energy INCREASES**

Reasons:

- e^- are being pulled **CLOSER** to the **NUCLEUS** (increased **NUCLEAR CHARGE**)
- more **ENERGY** needed to remove an e^-

3. Atomic Radius Trends

The **atomic radius** is one-half the distance between the nuclei of two atoms (just like a radius is half the diameter of a circle). This distance is measured in picometers. Atomic radius patterns are observed throughout the periodic table.

Atomic size gradually decreases from left to right across a period of elements. This is because, within a period or family of elements, all electrons are added to the same shell. However, at the same time, protons are being added to the nucleus, making it more positively charged. The effect of increasing proton number is greater than that of the increasing electron number; therefore, there is a greater nuclear attraction. This means that the nucleus attracts the electrons more strongly, pulling the atom's shell closer to the nucleus. The valence electrons are held closer towards the nucleus of the atom. As a result, the atomic radius decreases.

Down a group, atomic radius increases. The valence electrons occupy higher levels due to the increasing number of shells (principal energy levels). As a result, the valence electrons are further away from the nucleus as the number if shells increases. Electron shielding prevents these outer electrons from being attracted to the nucleus; thus, they are loosely held, and the resulting atomic radius is large.

- Atomic radius **decreases** from left to right within a period. This is caused by the **increase** in the number of protons and electrons across a period. One proton has a greater effect than one electron; thus, electrons are pulled towards the nucleus, resulting in a smaller radius.
- Atomic radius **increases** from top to bottom within a group. This is caused by electron shielding.

The **ionic radius** is the measure of an atom's ion in a crystal lattice. It is half the distance between two ions that are barely touching each other. Since the boundary of the electron shell of an atom is somewhat fuzzy, the ions are often treated as though they were solid spheres fixed in a lattice.

The ionic radius may be larger or smaller than the atomic radius (radius of a neutral atom of an element), depending on the electric charge of the ion.

Cations are typically smaller than neutral atoms because an electron is removed and the remaining electrons are more tightly drawn in toward the nucleus. An anion has an additional electron, which increases the size of the electron cloud and may make the ionic radius larger than the atomic radius.

Summary

Atomic Radius = ½ the distance between neighboring **NUCLEI** of a given **ELEMENT (**value listed on table S)

a. Going down a group, atomic radius INCREASES

Reasons:

- **MORE** orbitals/energy levels take up **MORE** space
- **SHIELDING** → electrons from inner energy levels shield/block valence electrons from the nuclear charge of the nucleus

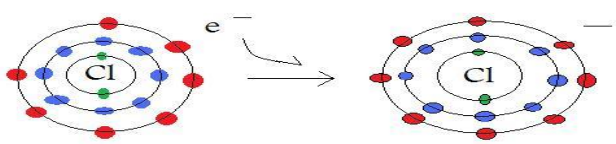
b. Going across a period, atomic radius DECREASES

Reasons:

- **NUCLEUS** getting **HEAVIER** (more P & N)
- **NUCLEAR** charge is **INCREASING** due to more protons = greater pull on electrons
- e^- (remember they are very **LIGHT**) are being pulled in **TIGHTER**

Ionic Radius **(Atomic radius for ions):**

- If you **ADD** e^-, radius **INCREASES**
  Reason: Same **NUCLEAR** charge pulling on **MORE** e^- → nucleus has **LESS** pull on outermost e^-



- If you **REMOVE** e^-, radius **DECREASES**
  Reason: Same **NUCLEAR** charge pulling on **LESS** e^- → nucleus pulls e^- **TIGHTER/CLOSER**

4. Metallic Character Trends

The **metallic character** of an element can be defined as how readily an atom can lose an electron. From right to left across a period, metallic character increases because the attraction between valence electron and the nucleus is weaker, enabling an easier loss of electrons. Metallic character increases as you move down a group because the atomic size is increasing. When the atomic size increases, the outer shells are farther away. The principle quantum number increases and average electron density moves farther from nucleus. The electrons of the valence shells have less attraction to the nucleus and, as a result, can lose electrons more readily. This causes an increase in metallic character.

- Metallic characteristics decrease from left to right across a period. This is caused by the decrease in radius of the atom that allows the outer electrons to ionize more readily.
- Metallic characteristics increase down a group. Electron shielding causes the atomic radius to increase thus the outer electrons ionizes more readily than electrons in smaller atoms.
- Metallic character relates to the ability to lose electrons, and nonmetallic character relates to the ability to gain electrons.

Another easier way to remember the trend of metallic character is that moving left and down toward the bottom-left corner of the periodic table, metallic character increases toward Groups 1 and 2, or the alkali and alkaline earth **metal groups**. Likewise, moving up and to the right to the upper-right corner of the periodic table, metallic character decreases because you are passing by to the right side of the staircase, which indicate the **nonmetals**. These include the Group 8, the **noble gases**, and other common gases such as oxygen and nitrogen.

- Move left across period and down the group: increase metallic character (heading towards alkali and alkaline metals)
- Move right across period and up the group: decrease metallic character (heading towards nonmetals like noble gases)

Summary:

Reactivity = **ABILITY** or **TENDENCY** of an element to go through a **CHEMICAL** change (or **REACT** with another element) **(*Can NOT compare metals to nonmetals)**

Metals: (recall: the most reactive metal is **FRANCIUM**)

a. Going down a group, reactivity ↑ (for **METALS**)

Reason:

- Increased **SHIELDING** means **VALENCE** e^- are held less tightly $\rightarrow$ e^- **LOST** more easily

b. Going across a period, reactivity ↓ (for **METALS**)

Reasons:

- increased nuclear **CHARGE** and **MASS** pulls more tightly on tiny, negative e^- $\rightarrow$ **HARDER** to remove e^-
  (magnet vs. car analogy)

Nonmetals: (recall: the most reactive nonmetal is **FLUORINE**)

a. Going down a group, reactivity ↓ (for **NONMETALS**)

Reason:

- Increased **SHIELDING** → **HARDER** for nucleus to attract more valence e^-

b. Going across a period, reactivity ↑ (for **NONMETALS**)

Reason:

- Increased nuclear **CHARGE** and **MASS** → **EASIER** for nucleus to attract more valence e^-

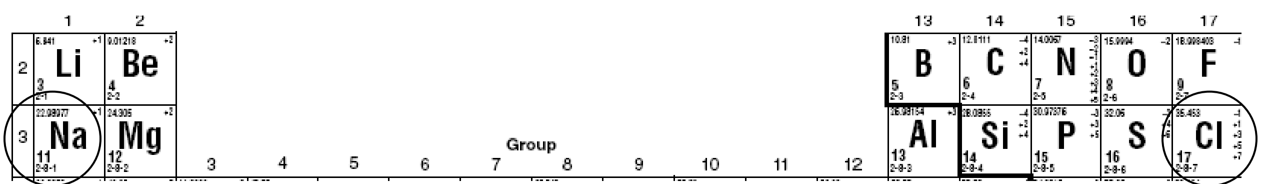
So how are you supposed to remember all of this????

****YOU **NEED TO KNOW** THESE TRENDS BUT YOU **DO NOT** HAVE TO **MEMORIZE** THEM!!!!! YOU CAN FIGURE THEM OUT USING YOUR **PERIODIC TABLE** AND **TABLE S** IN YOUR REFERENCE TABLE

THE TRICK…..

Example- If you are looking for the trend in electronegativity going across a group:

- Pick 1 element on the left side of the group and 1 element on the right
- Look up their values using table S
- If the values get larger than the trend is increasing; if smaller than trend is decreasing



Atomic Number	Symbol	Name	First Ionization Energy (kJ/mol)	Electro-negativity	Melting Point (K)	Boiling* Point (K)	Density** (g/cm^3)	Atomic Radius (pm)
1	H	hydrogen	1312	2.2	14	20.	0.000082	32
2	He	helium	2372	—	—	4	0.000164	37
3	Li	lithium	520.	1.0	454	1615	0.534	130.
4	Be	beryllium	900.	1.6	1560	2744	1.85	99
5	B	boron	801	2.0	2348	4273	2.34	84
6	C	carbon	1086	2.6	—	—	—	75
7	N	nitrogen	1402	3.0	63	77	0.001145	71
8	O	oxygen	1314	3.4	54	90.	0.001308	64
9	F	fluorine	1681	4.0	53	85	0.001553	60.
10	Ne	neon	2081	—	24	27	0.000825	62
11	Na	sodium	496	0.9	371	1156	0.97	160.
12	Mg	magnesium	738	1.3	923	1363	1.74	140.
13	Al	aluminum	578	1.6	933	2792	2.70	124
14	Si	silicon	787	1.9	1687	3538	2.3296	114
15	P	phosphorus (white)	1012	2.2	317	554	1.823	109
16	S	sulfur (monoclinic)	1000.	2.6	388	718	2.00	104
17	Cl	chlorine	1251	3.2	172	239	0.002898	100.

⇨ Trend is increasing

****This can be done for any trend going across a period or down a group

<u>Lesson 4: Allotropes</u>

Objective:

- ***Describe allotropes***
- ***Explain how they are used in technology***

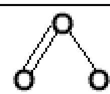
Allotropes are forms of the **same element** which exhibit **different physical properties**.
Elements such as carbon, oxygen, phosphorus, tin and sulfur, display allotropy.
The different physical properties displayed by allotropes of an element are explained by the fact that the atoms are arranged into molecules or crystals in different ways.
Some allotropes of an element may be more chemically stable than others.

Allotropes of Oxygen

There are two main allotropes of oxygen, molecular oxygen (O_2) and ozone (O_3).

Both allotropes of oxygen are made up only of oxygen atoms, but they differ in the arrangement of the oxygen atoms.
O_2 is a linear molecule while O_3 is a bent molecule.

O_2 and O_3 have different physical properties such as colour, odour, melting and boiling point, density and solubility.
Some properties of the allotropes of oxygen are shown below:

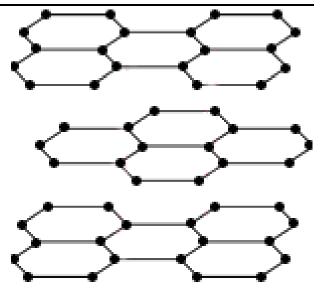
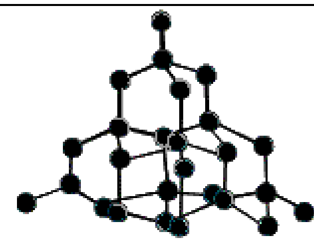
PROPERTY	OXYGEN (O_2)	OZONE (O_3)
Structure	**O=O** linear	 bent
Color	colorless gas pale blue liquid pale blue solid	pale blue gas deep blue liquid deep violet solid
Odor	odourless	sharp, pungent
Melting Point (^oC)	-219	-193
Boiling Point (^oC)	-183	-111
Density (20^oC)	1.3 g/L	2.0 g/L
Solubility in Water	slightly soluble	more soluble that O_2
Chemical Stability	stable	decomposes to O_2 easily
Uses	common oxidizer	sterilizing agent it is poisonous to many living things

Allotropes of Carbon

The two most common, naturally occurring allotropes of carbon are graphite and diamond.

Both graphite and diamond are made up of carbon atoms, but the arrangement of atoms is different in each allotrope which results in different physical properties.
In particular, the presence of delocalized electrons in the structure of graphite results in it being soft and a good electrical conductor whereas diamond is very hard and an electrical insulator.

Some properties of graphite and diamond are shown below:

PROPERTY	GRAPHITE	DIAMOND
Structure	 Each carbon atom is bonded to 3 other carbon atoms in layers with delocalized electrons between the layers.	 Each carbon atom is bonded to 4 other carbon atoms in a 3-dimensional covalent network. All valence electrons are used in bonding.
Color	black	colorless
Melting Point (K)	sublimes at ~3500	sublimes at ~4000
Electrical Conductivity	good delocalized electrons between the layers allow an electric current to pass through	poor (an insulator) no delocalized electrons to allow for the flow of electrical current
Hardness (Mohs Scale)	1-2 (soft) delocalized electrons allow the sheets to move over each other	10 (hardest known natural mineral)
Chemical Stability	stable	decomposes slowly over time
Uses	lubricant because it is soft	abrasive because it is so hard

Allotropes of Phosphorus

There are three allotropes of phosphorus; white, red and black.

Some properties of the allotropes of phosphorus are given below:

PROPERTY	WHITE PHOSPHORUS	RED PHOSPHORUS	BLACK PHOSPHORUS
Structure	P_4 molecules packed into a crystal	Chains of P_4 molecules polymer	Puckered layers of phosphorus atoms polymer
Color	white	red	black
Chemical Stability	least stable	intermediate stability	most stable

Allotropes of Sulfur

Sulfur has several allotropes.

α-sulfur forms yellow, rhombic crystals out of 8-membered rings of sulfur atoms (S_8).

γ-sulfur forms yellow, monoclinic, needle-like crystals out of 8-membered rings of sulfur atoms (S_8).

Plastic sulfur is yellow and made up of long chains of sulfur atoms. It reverts to S_8 rings in time.

Allotropes of Tin

There are three allotropes of tin:

Grey tin (α tin): a diamond-type lattice structure
White tin (β tin): body centered tetragonal structure
brittle tin: rhombic structure