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: Chapter Diary 5

Objective: To summarize concepts related to the Periodic table of elements

Directions: After reading Chapter diary 5 answer the questions in your workbook.

Lesson 2: The History of the Table &Chemical Periodicity

Objective: To relate the work of chemists to the modern periodic law and the repeating patterns in the periodic table

Watch the following clip:

[History & structure of the periodic table]

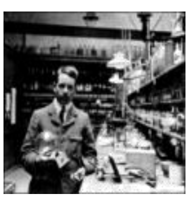
Dmitri Mendeleev (Russia)

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.

- 1^{st} 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)

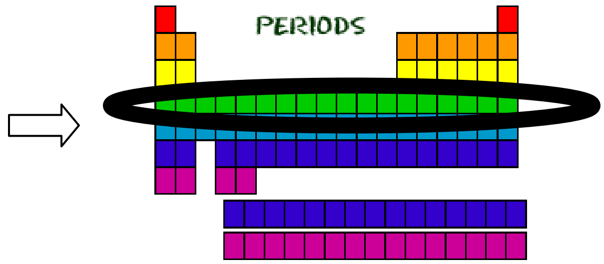
By 1907, when Mendeleev died, chemists were sure that iodine followed tellurium in the Periodic Table and that there was something odd about their relative atomic masses. However no-one was able to *measure* atomic number, it was just the position of an element in the Periodic Table sequence. For example 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**

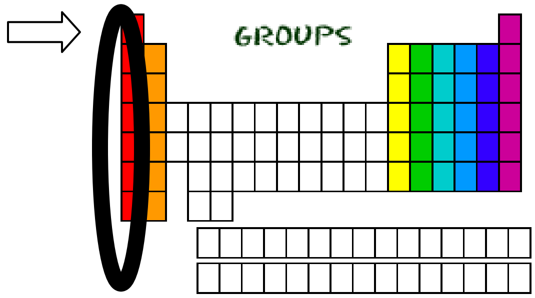
<u>How is the table actually arranged?</u>

<u>Periods</u> = **Horizontal rows** (run left to right) on Periodic Table



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

2. <u>Groups (Families)</u> = **<u>vertical columns</u>** (run up & down) on Periodic Table; each group contains the same number of **<u>valence electrons</u> 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 = **<u>1</u>**
Li = **<u>2-1</u>**
Na = **<u>2-8-1</u>**
K = **<u>2-8-8-1</u>**
Rb = **<u>2-8-18-8-1</u>**
Cs = **<u>2-8-18-18-8-1</u>**
*Fr = **<u>-18-32-18-8-1</u>**

What similarities can you observe within the above electron configurations?
All have **<u>1</u>** valence electron
Remember…
Group # = **<u>number of valence electrons</u>**
Period number = **<u>number of principle energy levels(shells)</u>**

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

- They have the same **<u>number of valence electrons</u>**
- Valence electrons affect **<u>Reactivity</u>**

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

- All atoms (except hydrogen and Helium) want **<u>8 electrons</u>** 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

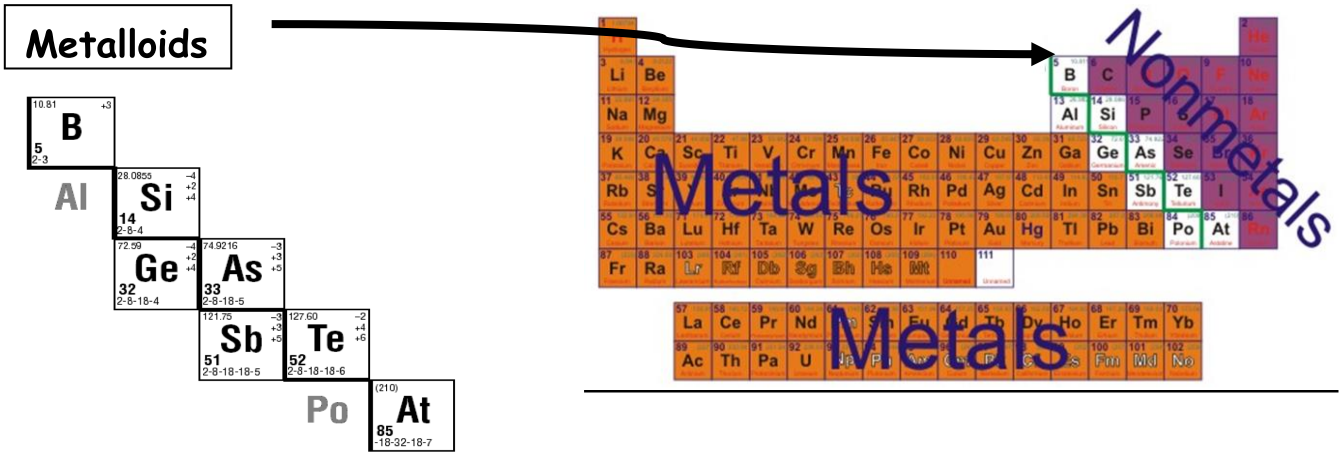
Atoms tend to gain, lose or share one or more of their valence electrons to achieve a filled outer electron shell

Lesson 3: Key to the Periodic Table

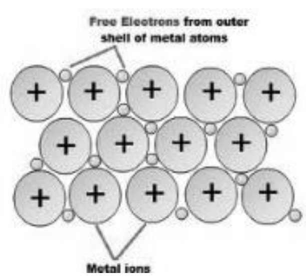
Objective: To define the location and compare and contrast the properties of metals, nonmetals, and metalloids.

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 determi 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**



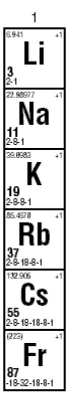
2. **Nonmetals:**
   - mostly **GASES** and **SOLIDS** @ STP—except **Br(l)**
   - **NOT** malleable/ductile; **BRITTLE** (shatter easily)
   - **LACK** luster (**DULL**)
   - **NON** or **POOR** conductors
   - like to **GAIN** e^- to form **NEGATIVE** ions

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

SUMMARY OF THE CATEGORIES OF ELEMENTS

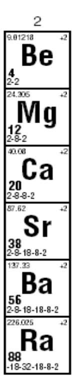
	Metals	Metalloids	Nonmetals
Phys. prop.	• malleable • ductile • shiny • excellent conductors (heat, electricity... MOBILE e-'s)	in-between	• brittle • dull • poor conductors (heat, electricity)
Chem. prop.	• lose e-'s • form + ions • low E.N. • low I.E.	B, Si, Ge, As, Sb, Te	• gain e-'s • form - ions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

The Groups in more detail



Group 1 → ALKALI METALS (FAMILY)

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



Group 2 → ALKALINE EARTH METALS (FAMILY)

- **All have 2 valence electrons**
- Prefer to LOSE their two electrons to become +2 ions
- **FAIRLY 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 $H_2(g)$, $H^+(aq)$ or $H^-(aq)$

The Lanthanide/Actinide Series

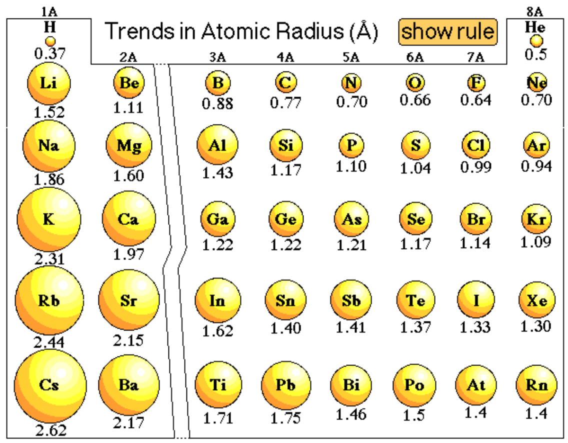
- Two rows on bottom of table (detached) – Elements 58 – 71 & 90 - 103
- Actually belong to the TRANSITION METALS

Lesson 4: Periodic Table Trends

Objective: To describe and explain the reason for periodic trends

1. 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

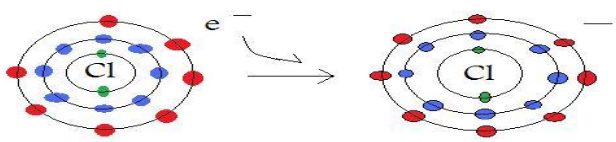
b. Going across a period, atomic radius DECREASES

Reasons:

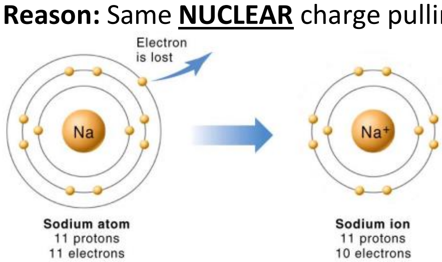
- **NUCLEUS** getting **HEAVIER** (more P & N)
- **NUCLEAR** charge is **INCREASING** due to more protons = greater pull on electrons
- e^- (remember they are very **LIGHT**) are being pulled in **TIGHTER**

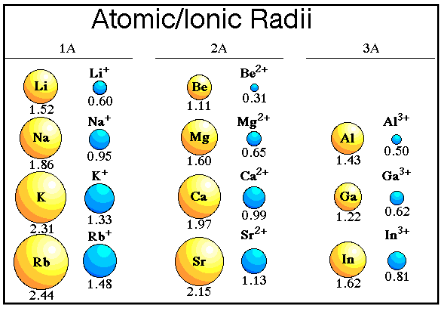
C. Ionic Radius (Atomic radius for ions):

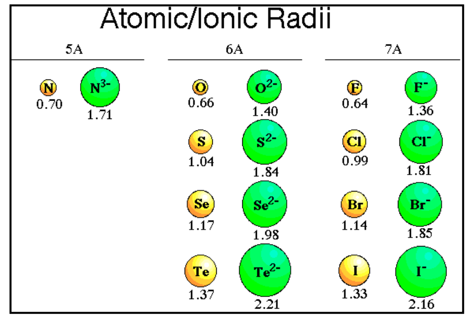
- If you **ADD** e^-, radius **INCREASES**
  Reason: Same **NUCLEAR** charge pulling on **MORE** e^- → nucleus has **LESS** pull on outermost e^-



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







2. 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)

X + energy → X⁺ + e⁻

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

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**

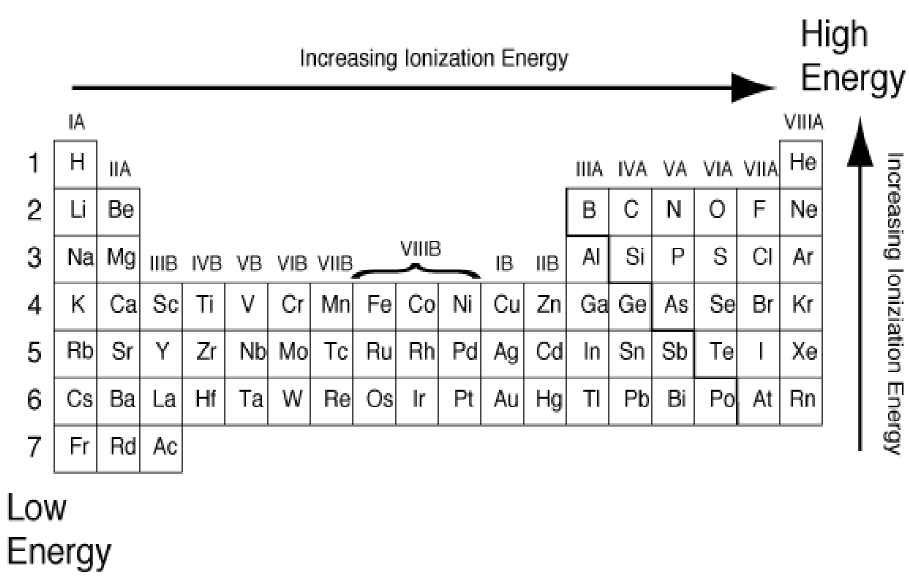
Reasons:

- 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. ***Electronegativity:***

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

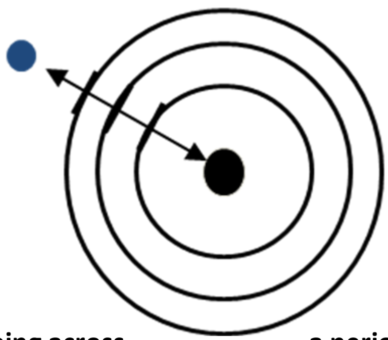
Electronegativity values range from 0.0 to 4.0

- **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)**

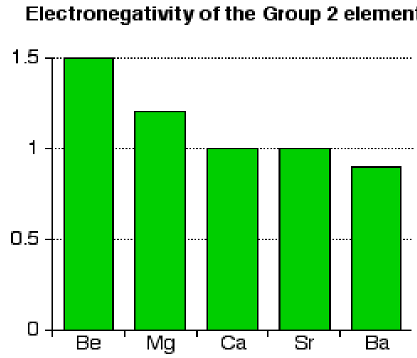
a. **Going down a group, electronegativity DECREASES**

Reasons:

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

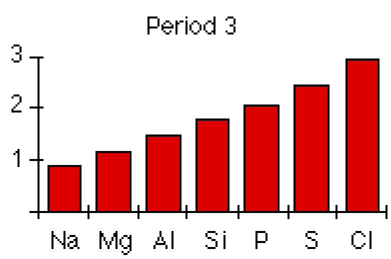


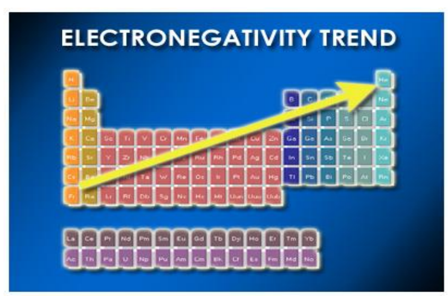
***Shielding**



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!)



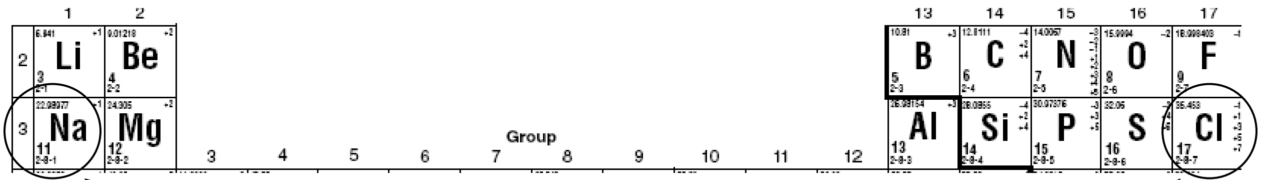


****YOU **NEED TO KNOW** THESE TRENDS BUT YOU **DO NOT** HAVE TO **MEMORIZE** THEM!!!!!
YOU CAN FIGURE THEM OUT USING YOUR **PERIODIC TABLE** AND **TABLE S** IN YOUR REFERENCE TABLE

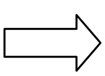
THE TRICK…..

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**

Practice: What is the trend in ionization energy as you go down a group?

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

<u>Metals:</u> (recall: the most reactive metal is **<u>FRANCIUM</u>**)

a. Going down a group, reactivity $\uparrow$ (for **<u>METALS</u>**)

Reason:

- Increased **<u>SHIELDING</u>** means **<u>VALENCE</u>** e^- are held less tightly → e^- **<u>LOST</u>** more easily

b. Going across a period, reactivity $\downarrow$ (for **<u>METALS</u>**)

Reasons:

- increased nuclear **<u>CHARGE</u>** and **<u>MASS</u>** pulls more tightly on tiny, negative e^- → **<u>HARDER</u>** to remove e^- *(magnet vs. car analogy)*

<u>Nonmetals:</u> (recall: the most reactive nonmetal is **<u>FLUORINE</u>**)

a. Going down a group, reactivity $\downarrow$ (for **<u>NONMETALS</u>**)

Reason:

- Increased **<u>SHIELDING</u>** → **<u>HARDER</u>** for nucleus to attract more valence e^-

b. Going across a period, reactivity $\uparrow$ (for **<u>NONMETALS</u>**)

Reason:

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

Lesson 5: Allotropes

Objective: To define and recognize an allotrope

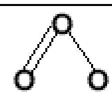
Allotropes are forms of the **same element** which exhibit **different physical properties**.
Elements such as carbon, oxygen, phosphorus, tin and sulfur, display allotropy.
The different physical properties displayed by allotropes of an element are explained by the fact that the atoms are arranged into 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:

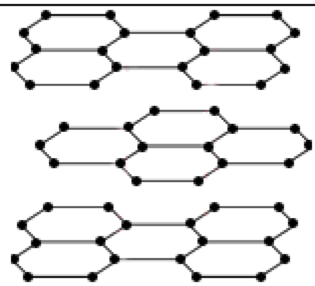
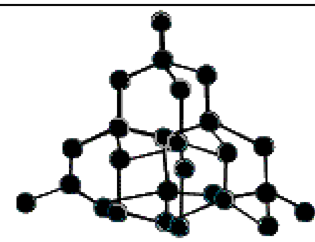
PROPERTY	OXYGEN (O_2)	OZONE (O_3)
Structure	**O=O** linear	 bent
Color	colorless gas pale blue liquid pale blue solid	pale blue gas deep blue liquid deep violet solid
Odor	odourless	sharp, pungent
Melting Point (^oC)	-219	-193
Boiling Point (^oC)	-183	-111
Density (20^oC)	1.3 g/L	2.0 g/L
Solubility in 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