Name:______________________________ Period:____ **Due Date: 1-18-2019** ______/ 100 Formative pts.

Chemistry Unit: Chemical Bonding (chapter 7 and 8) Notes

Topic-1: review
1. Valence electrons
2. Lewis dot structures
3. Electronegativity
4. Cations and Anions
5. Octet Rule

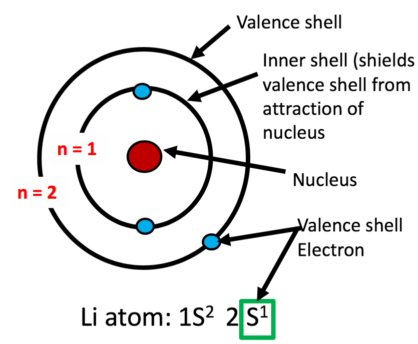
Topic-2: Chemical Bonds
1. Ionic Bonds
2. Covalent bonds
3. Metallic bonds

Topic-3: Inter-molecular Chemical Bonds
1. Van der Waalls attractions
2. Hydrogen bonds
3. Dipole-dipole bond

Topic-1: Review:

1. Valence electrons:

- The electrons in the outermost ____ of an atom
- **Valence** electrons are _________________or _________________during chemical reactions
- The _________ of valence electrons are directly responsible for the **chemical properties** and **reactivity of elements**
- **Periodic-Table-Groups** are based on ____________ electron numbers



2. Electron (Lewis) dot structures:

Lewis dot structures are _____________ that show _____________electrons of atoms.

Table-1: Lewis dot structures of atoms:

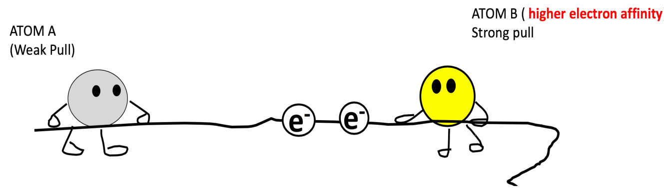
	Atomic number	Element	Electron configuration	# of Valence Electrons	Electron dot diagram
Period-1	1	**H**	$1s^1$	**1**	**H**
Energy level: n=1 Sub-levels: s	2	**He**	$1s^2$	**2**	
	3	**Li**	$1s^2$ $2s^1$	**1**	
Period-2	4	**Be**			
Energy level: n=2	5	**B**			
Sub-levels: s,p	6	**C**			
	7	**N**			
	8	**O**			
	9	**F**			
	10	**Ne**			
Period-3 Energy level: n=3	11	**Na**			
Sub-levels: s,p,d	12	**Mg**			

3. Electronegativity

- Electronegativity occurs when a one atom that __________ electrons with another is more attracted to these electrons due to its higher electron __________.
- Fluorine (F) is the _______ electronegative element on the periodic table. Followed by O, then N and Cl.
- In general, the electronegativity of an element ____________ as one goes up a group and left to right across a period.
- Note that Noble gases (group 8A) are not electronegative)

Example: In group 7A: Electronegativity ___________ in the order: I < Br < Cl < F.

In period 2: Li has the lowest and _______has the highest electronegativity



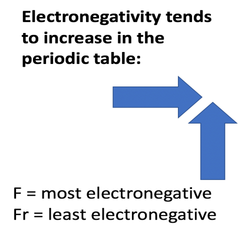


Table-2: Number of valence electrons in Noble gasses (group 8A):

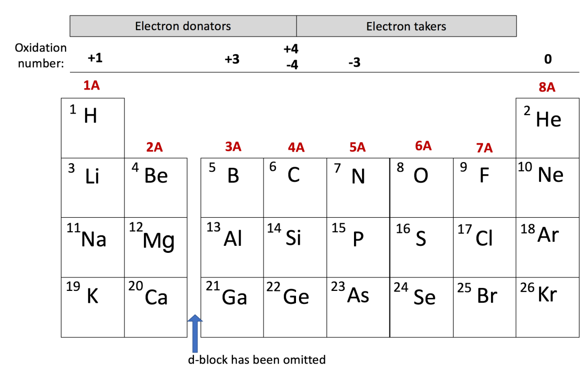
Atomic number	Symbol of noble gas	Electron configuration (underline the valence shell Orbitals)	Number of Valence electrons
10	Ne		
	Ar		
36			
	Xe		
86			

All noble gases have 8 valance electrons as their highest Principal Energy Level-SUB LEVELS (____ and ___) are completely filled with electrons. This is the reason for the Non-reactivity and stability of noble gases.

4. Cations and anions (oxidation numbers):

Group 1A – 4A atoms tend to form __________ by giving electrons to complete octet

Group 4A-7A Tend to form ____________ taking electrons to complete octet



5. Octet-Rule:

Atoms are always trying to get _____electrons in their outer shell (principal energy level) to be like the _______noble gas in the periodic table. This is because, having 8 electrons in outer shell is energetically very _______ for all atoms (except ___ and ___).

Table-3: atoms and their desired noble gas electron configuration based on OCTET-RULE

Element	Number of valence electrons	Desired noble gas electron configuration based on OCTET-RULE:
Na		
Mg		
C		
N		
Cl		

Name:______________________________ Period:____ **Date: 1-11-2019**

Topic-2: Bonds that Occur Between Atoms (Ionic, Metallic and Covalent Bonds)

1. Ionic Bonds:

- Ions with _______________ **charges** are **mutually attracted** to each other.
- Attractions between oppositely charged ions are called ______________ **attractions**.
- **Ionic bonds** are electrostatic attractions that occur between ions with opposite charge.
- Ionic bonds always occur between ________________ ions (cations) and **metal** ions (anions).
- All ionic compounds are ______________________ because their **total** __________ **and total negative charges add up to zero.**
- Ionic compounds arrange in a _________ ___________ called a _______ ____________ (like a lattice pie!)
- The energy needed to break this lattice is called ______________________.
- When the lattice is broken, the ions will become atoms with zero charge and the ionic bonds will be broken.

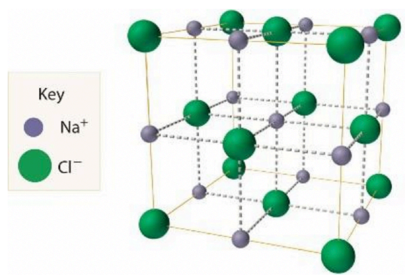
Properties of Ionic compounds

- Form **crystalline solids** at room temperature
- Generally have **high** ___________ **points**
- When dissolved in water, can **conduct** ___________ **and** form __________________.

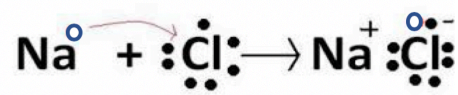
Questions:

a. What is an electric current? ____

b. Why do ionic compounds conduct electricity? ____



Lewis dot structure for NaCl:



Ex: Write the net ionic equation for the formation of Na^+ and Cl^- ions:

$Na - 1e^- \longrightarrow Na^+ + 1e^-$ (1 electron donated by sodium)

$Cl + 1e^- \longrightarrow Cl^-$ (Sodium electron accepted by Cl)

- Na is in group 1A = donate the valence electron 1 e -
- Cl is in Group 7A = accept this valence 1 e -
- So, when Na and Cl atoms form a chemical bond, Na will donate an electron to Cl.
- **Na donates 1** electron and become an **Na^+ ion** and Cl will **accept 1** electron and become a **Cl^- ion**

Crisscross-method of predicting Ionic compound formulas:

- **Step-1**: Write the name of the atoms side by side: _____________
- **Step-2**: Find out the number of valence electrons for each atom and the Charges made when they ionize (become ions). Ex: +2 or -1 etc. Write the Charge above the atoms (super-scrips): ____________
- **Step-3**: Now swap the charges as shown by the arrows:_______________
- **Step-4**: Then write the swapped ionic-charges as coefficient in front of the Ion: _____________

Note: Ones are not shown when writing formulas: ____________

This means, one Cl-ion combines with 3 Al-ions in the ionic compound. This will become clear to you once your draw the Lewis dot structures of the atoms and the compound.

- **Step-5**: Now draw the combined electron dot structure for the compound and determine how many of each ion is found in a unit of the compound.

- *Note: Since each Cl atom needs only ____ electron and Al atoms has ___ valence electrons to donate, ___ Cl ions will come bine with a single Al atom to take its ___ valence electrons and complete their octets!*

Name__________________ Class period: ___ score: ____/100 (**submissions after 1-28-19 will NOT be graded**)

Chemical Bonding Homework-1 (assigned: 1-11-2019): Due date 1-14-19

FYI: Electron Filling Order: 1s, 2s 2p, 3s 3p, 4s 3d 4p, 5s 4d 5p, 6s 4f 5d 6p ,7s 4f 6d 7p

1. What are valence electrons? ____
____
2. If you need to find the valence electrons of an element in the periodic table which would you use? The group number, atomic number, atomic mass, period number (circle the correct answers)
3. What is a Lewis dot structure? ____
____
4. Draw the electron dot structures of the following atoms:
   a) Calcium
   b) Iodine
   c) Rubidium
   d) Carbon
   e) Oxygen
   f) chlorine
5. Explain the octet rule: ____
____
6. True or false? Metallic atoms tend to lose electrons and form cations while non-metal ions tend to gain electrons and become anions.
7. Explain why anions have a negative charge:____
____
8. Write the electron configuration of the following metal atoms and Circle the electrons LOST when these metals form cations:

Element Symbol	Total electrons	Electron configuration	Orbital box diagram
Mg			
Al			
K			

9. Match the correct noble gas or ion with the electron configurations below using arrows:

1) Ar — a. $1S^2$
2) Ca^{2+}
3) Cl^- — b. $1S^2 2S^2 2p^6$
4) Li+
5) Se^{2-} — c. $1S^2 2S^2 2p^6 3S^2 3p^6$
6) He
7) Ne — d. $1S^2 2S^2 2p^6 3S^2 3p^6 3d^{10} 4S^2 4p^6$

10. Draw the orbital box diagrams of the following:

1) Ar
2) Cl^-
3) Li^+
4) He
5) Al^{3+}

11. What is an ionic bond: ____

____

12. What are the properties of ionic compounds?

____

____

13. Predict the formulas of the ionic compounds formed by the atoms below:

1) *Na and Cl* ___________
2) *Na and S*___________
3) *Na and P*__________
4) *Na and C*__________
5) *Mg and Cl*__________
6) *Mg and S*___________
7) *Mg and P*___________
8) *Mg and C*___________
9) *Al and Cl*___________

10) *Al and S*___________

11) *Al and P*___________

12) *Al and C*___________

13) *C and Cl* __________

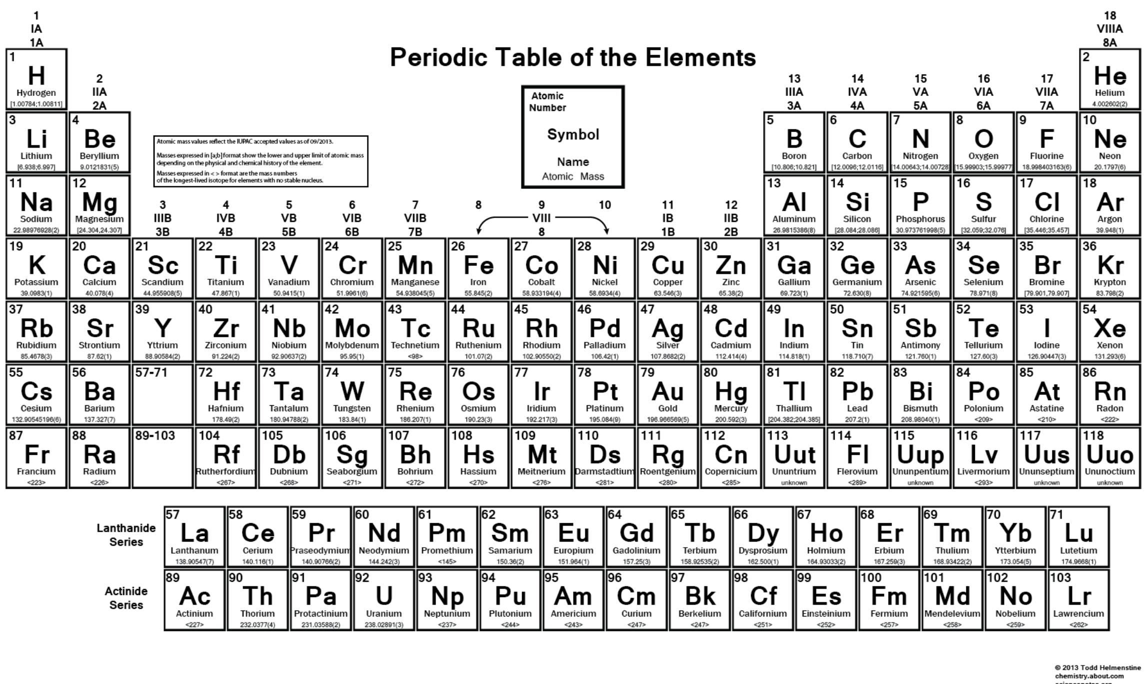
14) *C and S* __________

15) *C and P*__________

16) *Mg and C*_________

17) Potassium and Oxygen: _____________

18) Magnesium and Nitrogen:____________



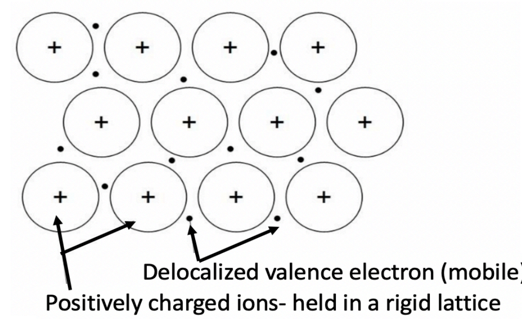
2. Metallic Bonds

What is a metallic bond? The attraction between metal cations and the surrounding sea of delocalized valence electrons is called metallic bond.

Physical Properties of Metals:

- Metals form ______________ cubic, rigid crystals
- Metals are good ___________ conductors as they have readily mobile electrons to produce a current
- The __________ -electron mobility is also why metals are _____________ (hammered into shapes) and ______________(stretched into wires).

Metal atomic structure:



Inside metals, positively charged ions (cations) swim in a _________ of delocalized valence electrons. These electrons are __________ by all neighboring atoms and are free to ________ around.

Metal Alloys:

What is a metal Alloy? ____
____

Metal alloys can be of two types:

1. **Substitution alloy** = another metal atom of _________ ________ is substituted in the space of a former metal atom (made up of 2 or more metals)
2. **Interstitial alloy:** made up of metal atoms and another __________ metal or non-metal atom that is in the ___________ ___________ metal atoms.

Examples of alloys:

1.___________: Cu + Tin + Al /Mn /Zn sterling silver,
2.Cast iron: Fe + ____+ Si
3.Steel: Fe +_____

Q: Why is steel stronger than cast iron? Explain your answer
____
____

3. Covalent Bonds [co-valent = shared- valence]

What is a covalent bond? Covalent bonds are _________________ bonds formed when atoms ____________ electrons. All compounds that make molecules are formed by___________ bonding of the atoms within.

- In covalent bonds, electrons are usually shared between 2 or more _____________ atoms so that each can obtain the noble gas electron configuration.
- There can be single, double or triple covalent bonds between two atoms
- Based on the _______________________ values of each atom in a covalent bond, the resulting molecule may have a dipole (polarity) based on negative and positive charged regions of the same molecule. [Polarity is explained with the symbol 'delta' = _____ or ____]
- Covalent bonds and ionic bonds exist in a____________ (continuum) going from zero to extreme electronegativity.

Physical Properties of covalent molecules:

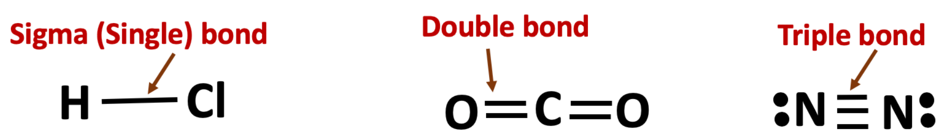
- Bond strength varies depending on type of bond and atoms involved in bonding
- Can be found in Solid, liquid or gas state depending on the compound
- Poor electrical conductors (as they do not have readily mobile electrons to produce a current)
- Can be colorless or have colors
- Low melting point and boiling point
- Do not conduct heat well

Types of covalent bonds:

1. Single bonds [formed by sharing __ electron-pair = ____ electrons between 2 atoms]
2. Double bonds [formed by sharing __ electron-pairs = ___electrons between 2 atoms]
3. Triple bonds [formed by sharing ____electron-pair = ____ electrons between 2 atoms]

Note:

- Each bond is represented as a straight ______. A single bond will have______ line, double will have 2 lines and a triple bond will have _____ lines.
- ________ bonds are the strongest and single bonds are the weakest and ______ to break.
- Bond distance is shortest in triple bonds and longest in single bonds. See figure below:



Draw the Lewis dot structures of the following covalent compounds/molecules:

1. HCL

2. CO_2 (carbon dioxide)

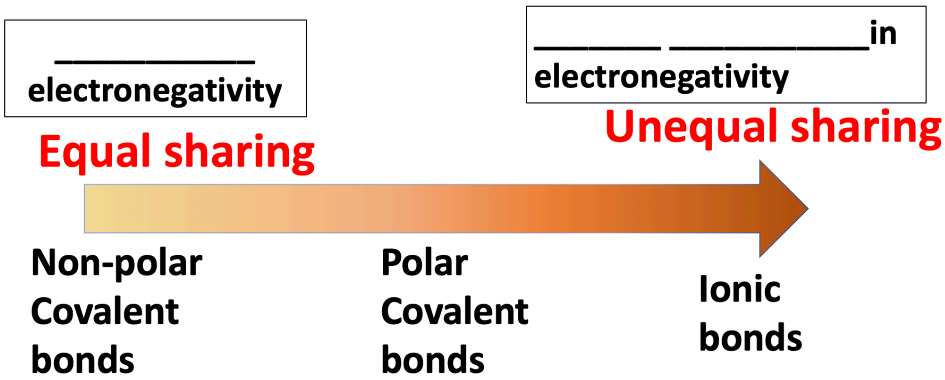
3. N_2 (nitrogen gas)

4. H_2 (hydrogen gas)

5. O_2 oxygen gas

6. H_2O (Water)

Sale of valence electron sharing in covalent bonds:



<u>Polar covalent bonds</u>

Ex: water molecule _______

H_2O

- Oxygen is much more ______________ than Hydrogen, so Oxygen pulls the shared electrons towards itself
- This greater electronegativity of oxygen compared to hydrogen in the covalent bonds of water molecules result in a _________or <u>polarity.</u> The result is a shifting of the electron density towards Oxygen.
- H gets a small positive charge Delta + (___) and Oxygen gets a ______________ charge (____) as electrons are less equally shared between O and H.

<u>Non-polar covalent bonds:</u>

Ex: N_2

- If a molecule if formed by the same kind of atom or atoms with similar electronegativity then the covalent bond is _____________ (Ex: O_2, N_2, Cl_2, O_3, Br_2)
- Carbon and __________ makes non polar covalent molecules such as hydrocarbons (CH_4, C_2H_6, C_3H_8 and so on.)

Draw C_2H_8

Co-ordinate Covalent bonds:

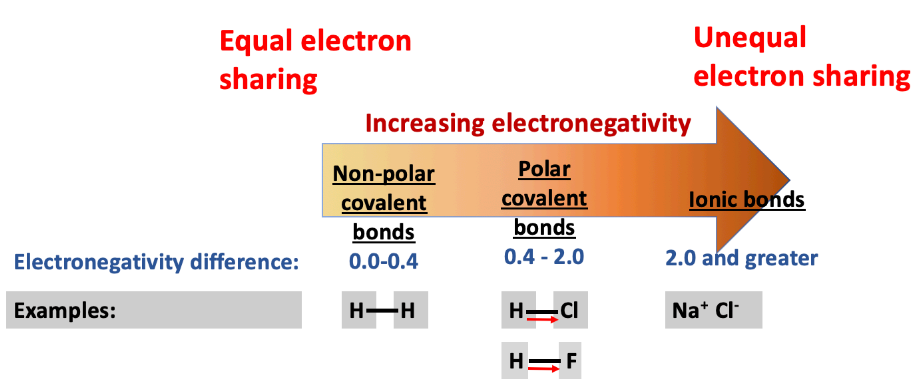
In typical covalent bond, each atom in the bond provides an electron for a single bond. But in coordinate covalent bonds this is not the case. Here, one atom contributes both of the shared electrons for one covalent bond.

Q: How are covalent compounds (molecules) different from Ionic compounds?

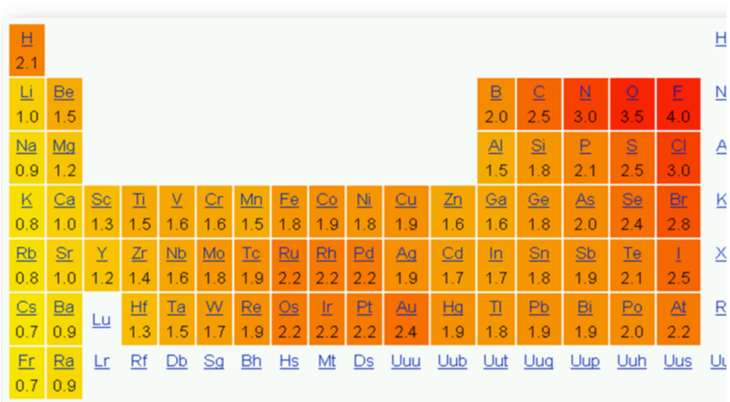
____

____

Electronegativity difference and Bond-Types:



Periodic Table of Electronegativities



H 2.1																	H
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	N
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	A
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	K
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	X
Cs 0.7	Ba 0.9	Lu	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2	R
Fr 0.7	Ra 0.9	Lr	Rf	Db	Sg	Bh	Hs	Mt	Ds	Uuu	Uub	Uut	Uuq	Uup	Uuh	Uus	Uu

Calculate the electronegativity of

1. F:________
2. H:_______
3. C:_______
4. H-F:______
5. C-C:_____
6. C-H:______

Ch7 and 8 notes contd. Name: _________________________ Period: ____ Date: 1-18-2019

Properties of chemical bonds:

	Property	Meaning
1	Bond Length	Distance between the nuclei of ________ bonded atoms (unit = picometers =pm)
2	Bond angle	Angles of any two bonds around an ________
3	Bond energy (bond dissociation energy)	Energy required to break a bond. (unit = Kilo ________ per mole = KJ/mol)

Bond	Length	Energy
C – C	154	348
C = C	134	614
C ≅ C	120	839
O – O	148	145
O = O	121	498

Compound	Bond	length (pm)	energy (kJ/mol)
H_2	H – H	74	436
HF	F – H	92	565
H_2O	O – H	96	464
NH_3	N – H	101	389
CH_4	C – H	109	414

****Bond length decreases when number of bonds increase**

****Bond Strength increases when number of bonds increase**

Periodic Table of Electronegativities

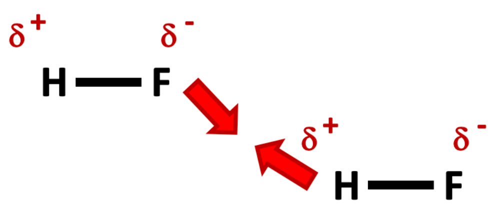
H 2.1																	He
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe
Cs 0.7	Ba 0.9	Lu	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2	Rn
Fr 0.7	Ra 0.9	Lr	Rf	Db	Sg	Bh	Hs	Mt	Ds	Uuu	Uub	Uut	Uuq	Uup	Uuh	Uus	Uuo

Topic 3: INTER-MOLECULAR BONDS

1. Dipole-Dipole bonds (attractions)

When two polar molecules get near each other, the positive end (δ^+) of one attracts the negative (δ^-) end of the other. This attraction between the two dipoles is known as a dipole-dipole bond

Ex: HF

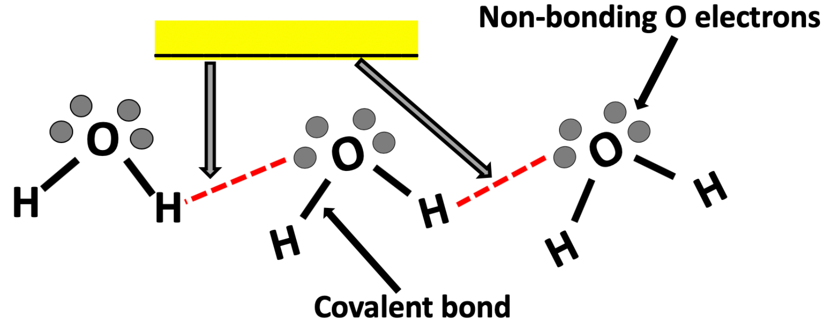


- Because F is very _______ __________, it pulls the shared electrons towards it causing a partial negative charge. This leads to a partial positive charge on the side of H leading to a **dipole** (opposite charges) in the atom.
- Between HF molecules, the _________ **charged ends attract** as shown in thick arrows. These are called dipole-dipole attractions.

2. Hydrogen bonds:

Hydrogen bonds: formed between non-bonding electrons of O, N, F atoms of one molecule and ______ atoms of another molecule

Ex: Water molecules



3. Van der Waal's attractions

- Electrons in atoms are in constant motion, and there are brief periods of time in which they are slightly to one side of the molecule or the other. During this time, the molecule has a very slight, ________ _________.
- If another, similar, temporary dipole exists in a ______________ **molecule**, the **weak _________ attraction** causes the **two dipoles to persist for slightly longer** than they otherwise would, and the two molecules remain attracted for a short period of time. Such forces are called **van der Waal's forces.**

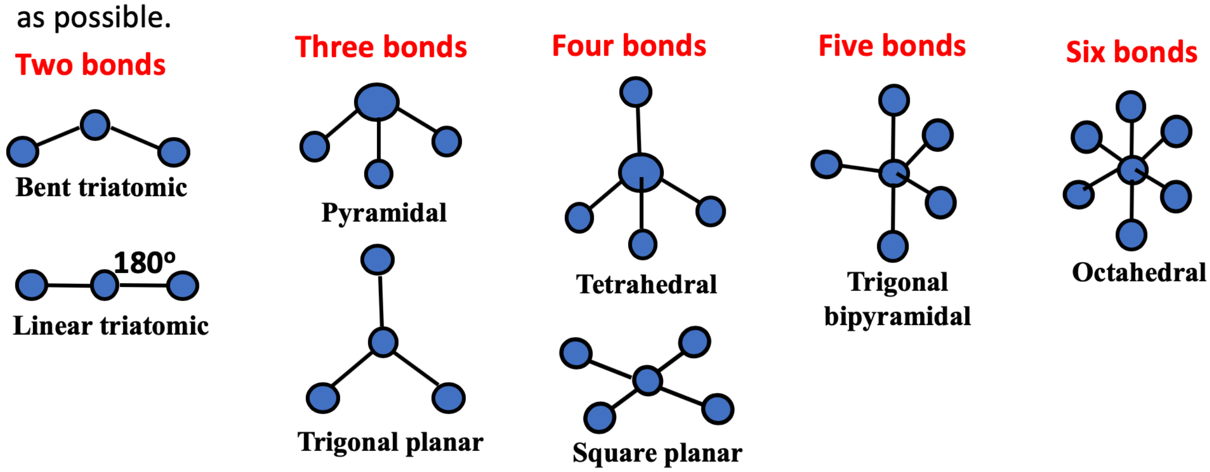
- **So, ________ compounds have these!**

Ch7 and 8 notes contd. Name: ________________________ Period: ____ Date: 1-18-2019

Topic 4: MOLECULAR GEOMETRY

VSPR: Valence-Shell Electron Pair Repulsion theory:

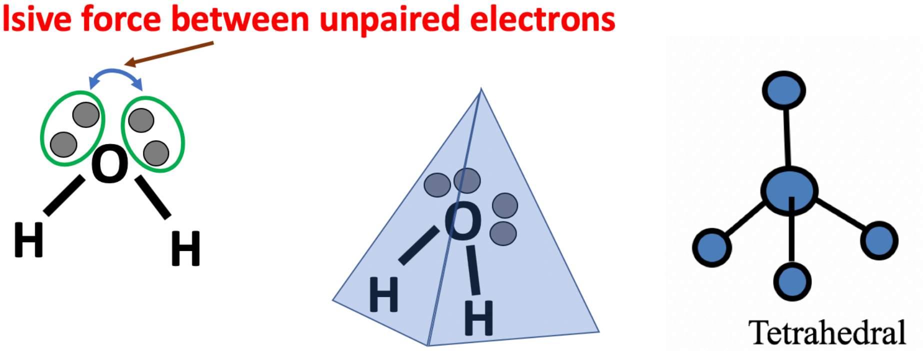
Repulsion between __________ _________ **electron pairs** cause molecular shapes to change so that valence electron pairs repel each bother and want to stay as far apart as possible.



Example: Water molecule:

Repulsion between ____________ electrons also effect 3D molecular shapes: Ex. H_2O
In water, the two lone pair electron repulsion and the electron clouds of these electron pairs cause the molecule to have a______________rather than a triatomic shape.

Repulsive force between unpaired electrons



1-18-2019 (HW2) HOME WORK QUESTIONs Due date: 1-22-19 (will NOT be graded 2 weeks post due date)

1. How are covalent compounds (molecules) different from Ionic compounds?
2. What is the molecular geometries of the following molecules?

   a. H_2O

   b. CO_2

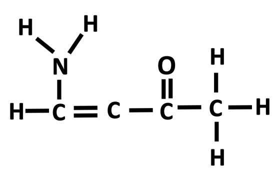
   c. NH_3

3. Draw the Lewis dot structures of the following compounds:

Name	Chemical formula
Hydrogen peroxide	H_2O_2
Hydrochloric acid	HCl
Sulfur trioxide	SO_3
Hydrogen cyanide	HCN
Nitrous Oxide	N_2O
Nitrogen dioxide	NO_2
Ethane	C_2H_2

4. Explain why all molecules have van der Waals forces:

5. Consider the following molecule:



   a. What are the chemical bonds formed between atoms in the following molecule?

   b. Which types of the chemical bonds formed between these molecules? Draw the two molecules and indicate bonds with arrows

   c. What are chemical bonds formed between these molecules and H-Br molecules? Draw the two molecules and indicate bonds with arrows