Ms. Randall
Regents Chemistry
<u>Unit 5:</u> Bonding and Naming compounds Unit Notes

Unit Objectives:

- Differentiate compounds by their chemical and physical properties
- Categorize compounds as ionic and molecular (covalent) compounds or metallic
- Relate valence electrons to type of bond ex/ transferred from one atom to another (ionic); shared between atoms (covalent); mobile within a metal (metallic).
- Differentiate between single and multiple covalent bonds
- Determine molecular polarity by the shape and distribution of charge. Symmetrical (nonpolar) molecules include CO_2, CH_4, and diatomic elements. Asymmetrical (polar) molecules include HCl, NH_3, and H_2O.
- Relate type of ion to changes in atomic radius
- Compare energy changes that occur in the formation and breaking of chemical bonds.
- Describe the formation of a stable valence electron configuration by bonding with other atoms. Identify the Noble gases a having stable valence electron configurations that tend not to form bond.
- Explain physical properties of substances in terms of chemical bonds and intermolecular forces. These properties include conductivity, malleability, solubility, hardness, melting point, and boiling point.
- Draw Electron-dot diagrams (Lewis structures)to represent the valence electron arrangement in elements, compounds, and ions.
- Define electronegativity as how strongly an atom of an element attracts electrons in a chemical bond. Electronegativity values are assigned according to an arbitrary scale.
- Use the electronegativity difference between two bonded atoms to assess the degree of polarity in the bond.
- Define ionic compounds by as metals reacting with nonmetals .Nonmetals tend to react with other nonmetals to form molecular (covalent) compounds. Ionic compounds containing polyatomic ions have both ionic and covalent bonding.
- Determine the noble gas configuration an atom will achieve when bonding.
- Demonstrate bonding concepts, using Lewis dot structures, representing valence electrons: transferred (ionic bonding); shared (covalent bonding); in a stable octet.
- Distinguish between nonpolar and covalent bonds (two of the same nonmetals) and polar covalent bonds.

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

Ion
Ionic Bond
Stable Octet
Octet Rule
Diatomic Molecules
Electronegativity
Ionic bond
Covalent bond
Metallic bond
Network solid
Polar Covalent Bond
Non-Polar Covalent Bond
Intermolecular forces of attraction.
Van der Waals Forces
Dipole –Dipole
Hydrogen Bonding

Lesson 1: Chapter Diary 6

Objective: To summarize concepts related to chemical bonding

Directions: After reading Chapter diary 6 answer the questions in your workbook.

Lesson 2: Defining a Chemical Bond

Objective: Categorize compounds as ionic and molecular (covalent) compounds or metallic

A bond is an attraction between atoms that causes them to be held together. This **attraction** is between one atom's nucleus (protons) and another atom's **valence electrons**. Atoms bond together to get **8 valence electrons** to become **STABLE (Stable Octet**) Exception: Hydrogen can only have 2 (stable duet)

Chemical compounds are formed by the **joining of two or more atoms**. A **stable compound occurs** when the total energy of the combination has **lower energy** than the separated atoms. The bound state implies a net attractive force between the atoms ... a chemical bond. The two extreme cases of chemical bonds are:

Covalent bond: bond in which one or more pairs of electrons are shared by two atoms.

Ionic bond: bond in which one or more electrons from one atom are removed and attached to another atom, resulting in positive and negative ions which attract each other.

Another type of bond is a **metallic bond**

The formation of a Bond **is SPONTANEOUS.** Energy is **RELEASED** which creates **STABILITY.**

EXOTHERMIC = when energy is **RELEASED** as a product

A + B → C + energy

*NOTE: "energy" on **PRODUCTS** side of arrow

Ex: $CaCl_2 + NaHCO_3 + H_2O \rightarrow CaCO_3 + CO_2 +$ **HEAT**

The breaking of Bonds **is NOT spontaneous.** Energy is **ABSORBED.** It requires ENERGY to break a bond.

ENDOTHERMIC = when energy is **CONSUMED** or needed as an ingredient to fuel the process.

A + energy → B + C

*NOTE: "energy" on **REACTANT** side of arrow

Ex: an ice pack → chemicals combine, bonds are broken, and energy is consumed → you feel "cold" because you're losing heat to the ice pack

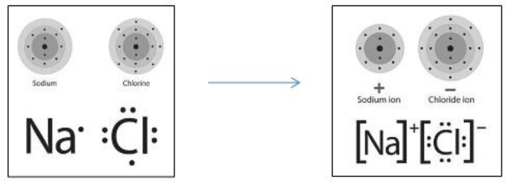
Remember "***BARF "***

- Breaking bonds Absorbs (requires) energy
- Release of energy when bonds are Formed

Watch the following video: [Ionic and molecular compounds]()

<u>There are THREE types of bonding:</u>

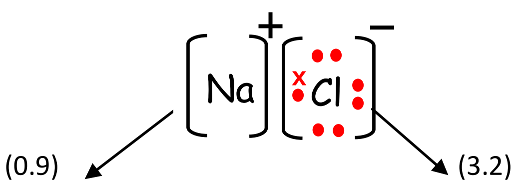
1. **<u>Ionic Bonds:</u> TRANSFER** of electron(s) from a **METAL** to a **NONMETAL**



<u>Physical Properties:</u>

- Strong bonds
- High m.p./b.p (melting point/ boiling point)
- **Electrolyte**-Compound that separates into ions in solution and is able to conduct electricity
- Hard/Brittle

Electronegativity Difference = 2.3

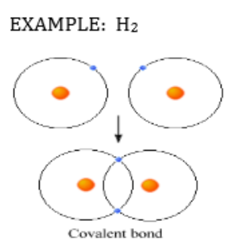


- The greater the ELECTRONEGATIVITY difference between two elements the greater the percent IONIC CHARACTER **(>1.7 EN difference)**

2. **<u>Covalent Bonds:</u>** Electrons **SHARED** between 2 or more **NONMETALS**

Molecule = A **COVALENTLY** bonded substance; always 2 or more **NONMETALS** bonded together

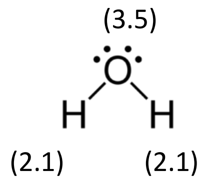
***MOLECULAR** compound = **COVALENT** compound



<u>Physical Properties:</u>

- Weaker bonds
- Low m.p./b.p. (melting point/ boiling point)
- Nonelectrolyte (does not conduct electric. current in solution)
- Soft

Electronegativity Difference = 1.4



- The closer the ELECTRONEGATIVITY difference is to zero, the greater the percent COVALENT CHARACTER

` **(< 1.7 EN difference)**

Watch the following video: Electronegativity

3. POLYATOMIC IONS

- Contain both Ionic and Covalent Bonds
- Located on Table E
- The atoms in polyatomic ions are held together by covalent bonds
- Form IONIC COMPOUNDS with other substances due to presence of ions

  Ex. NH_4Cl

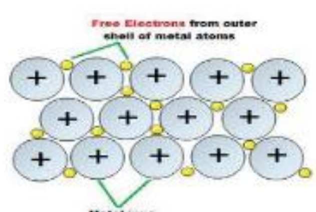
Formula	Name
H_3O^+	hydronium
Hg_2^{2+}	mercury(I)
NH_4^+	ammonium
$C_2H_3O_2^-$ / CH_3COO^-	acetate
CN^-	cyanide
CO_3^{2-}	carbonate
HCO_3^-	hydrogen carbonate
$C_2O_4^{2-}$	oxalate
ClO^-	hypochlorite
ClO_2^-	chlorite
ClO_3^-	chlorate
ClO_4^-	perchlorate

Formula	Name
CrO_4^{2-}	chromate
$Cr_2O_7^{2-}$	dichromate
MnO_4^-	permanganate
NO_2^-	nitrite
NO_3^-	nitrate
O_2^{2-}	peroxide
OH^-	hydroxide
PO_4^{3-}	phosphate
SCN^-	thiocyanate
SO_3^{2-}	sulfite
SO_4^{2-}	sulfate
HSO_4^-	hydrogen sulfate
$S_2O_3^{2-}$	thiosulfate

4. <u>Metallic bonds:</u>

Metals = **pure substances found to the left of the "staircase"**

- atoms of a metal do not bond w/ other metal atoms
- metals "share" a **"sea of MOBILE valence electrons"**
  - ➢ allows metals to conduct electric **CURRENT** (**WAVE** energy)



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Lesson 3: Ionic Bonding

Objective: To represent the transfer of electrons in an ionic bond with a Lewis diagram

Metals can be defined as elements that are **likely to lose electrons in bonding** situations. We can say that *the most metallic elements* are the ones that *are most likely to lose electrons*.

Nonmetals can be defined as elements that are **likely to gain electrons in bonding** situations. The *most nonmetallic elements* are the ones that are *most likely to gain electrons***.**

Sodium (2-8-1) has 1 electron more than a stable noble gas structure (2-8). If it gave away that electron it would become more stable. Chlorine (2-8-7) has 1 electron short of a stable noble gas structure (2-8-8). If it could gain an electron from somewhere it too would become more stable.
The answer is obvious. If a sodium atom gives an electron to a chlorine atom, both become more stable.

Na 2,8,1 $\longrightarrow$ Na^+ 2,8

Cl 2,8,7 $\qquad\quad$ Cl^- 2,8,8

The sodium has lost an electron, so it no longer has equal numbers of electrons and protons. Because it has one more proton than electron, it has a charge of 1+. If electrons are lost from an atom, positive ions are formed.
Positive ions are called ***cations.***

The chlorine has gained an electron, so it now has one more electron than proton. It therefore has a charge of 1-. If electrons are gained by an atom, negative ions are formed.
A negative ion is called an ***anion.***

The nature of the bond

The sodium ions and chloride ions are held together by the strong **electrostatic attractions** between the **positive** and **negative** charges.

The formula of sodium chloride
You need one sodium atom to provide the extra electron for one chlorine atom, so they combine together 1:1. The formula is therefore NaCl.

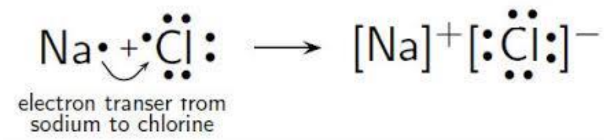
Ionic compounds are Hard CRYSTALLINE SOLIDS (not molecules). They have high melting and boiling points (strong attraction between ions). They are Electrolytes which means they can conduct electricity when DISSOLVED in WATER. Ions are charged particles that are free to move. They DO NOT conduct electricity as solids.

Make some Ionic Compounds:
http://www.learner.org/interactives/periodic/groups_interactive.html

Lewis dot diagrams give us a PICTURE of the ionic compound and shows the ARRANGEMENT of atoms in the compound. Within IONIC COMPOUNDS there is a TRANSFER or DONATION of electrons from the METAL to the NONMETAL.

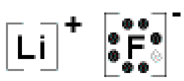
Steps for Drawing Ionic Bonding Lewis Dot Diagrams:

In an ionic bond, one atom loses all its outer electrons (leaving behind a filled inner shell) while another atom gains electron(s) to fill its valence shell. When you draw an ion, don't forget [ ] and a charge.
The two opposite ions attract each other.

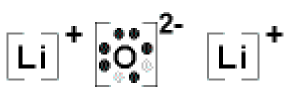


Look the metal has no valence electrons and the nonmetal is full. Always!!!!

Examples:

LiF==> 

or

Li_2O==> 

So far we have discussed ionic bonding and the compounds formed as a result. The ionic compounds we have discussed so far contain two elements. One is a metal, the other is a nonmetal. **Bi**nary is the Chemist's word for an ionic compound with **two** elements in it.

Example: $AlBr_3$ is a binary compound named Aluminum bromide.

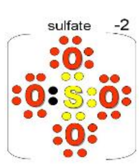
What about compounds with more than 2 elements?!

Ternary Ionic Compounds

- A ternary ionic formula contains three or more elements, but is still formed as the result of an attraction between ions of opposite charge.
- A ternary ionic compound involves a **polyatomic** ion.

Since this has three elements, it is possible to have both ionic and covalent bonds.

Examine the Lewis diagram for the "sulfate" polyatomic ion

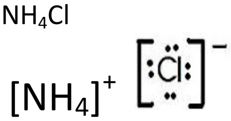


- Why this is called a "polyatomic ion?"
  - It's made of more than one atom but has a charge.
  - Notice the electrons are being shared to build the ion, however the ion is used to build an ionic compound.

Common Polyatomic Ions are located on Table E of the reference tables.
When we use polyatomic ions in an ionic compound we just write the ion in parentheses and its charge on the upper right hand corner to the outside as its Lewis dot diagram.

Example:

NH_4Cl



Lesson 4: Covalent Bonding **Date:________**

Objective: To form and name covalent (molecular) compounds

Covalent Bonding is the **sharing** of valence electrons between **nonmetals.** Molecules are electrically **neutral** just like atoms. Ratio depends on how many electrons are needed to complete the **octet.** Most **molecular** (covalent) compounds are in the liquid or gas state. They do not have as strong of a particle attraction as ionic substances! They are soft due to weaker particle attractions. They have low melting and boiling points and are poor conductors of electricity.

COVALENT
Substances

Nonmetals only

(Note: count *metalloids* like nonmetals)

Writing Covalent formulas:

1. LEAST electronegative element is written FIRST.
2. MOST electronegative element is written LAST.
3. SUBSCRIPTS tell you the PREFIX of each element in the formula. Pre-fixes can be found on the reference table Example: CO_2 = carbon *di*oixide)

# of atoms (subscript)	1	2	3	4	5	6	7	8	9	10
Prefix	mono	di	tri	tetra	penta	hexa	septa hepta	octo	nona	deca

Naming covalent formulas

1. Write name of first element in formula. (as on Table S)

2. Write root of second element in formula with –***ide*** ending.

3. Turn subscripts into Greek prefixes.

1 (or nothing written) = mono ← never use on first element
2 = di
3 = tri
4 = tetra
5 = penta

4. Diatomic elements: just write their names (no prefixes)

Examples:

Formula	**Name**	**Formula**	**Name**
PH_3	phosphorus trihydride	SF_4	sulfur tetrafluoride
PCl_3	phosphorus trichloride	F_2	fluorine
SO_2	sulfur dioxide	NH_3	nitrogen trihydride
N_2O_2	dinitrogen dioxide	CO	carbon monoxide

Summary:

COVALENT
Substances

Nonmetals only

(Note: count *metalloids* like nonmetals)

B, Si, Ge, As, Sb, Te

One word

Just write the element's name

(watch out for diatomic elements)

bromine, iodine, nitrogen, chlorine, hydrogen, oxygen, fluorine

Two words

The **Greek prefix** becomes the subscript on the element

mono = 1

di = 2

tri = 3

tetra = 4

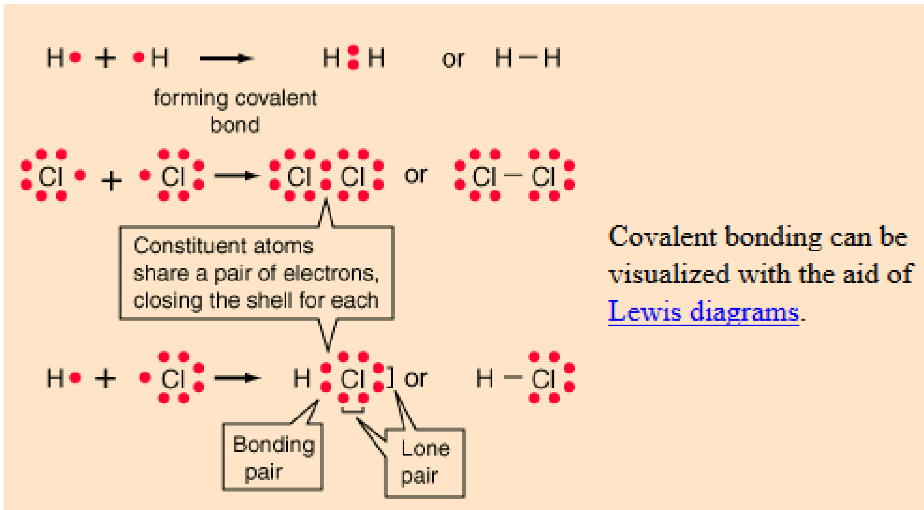
penta = 5

*If no prefix, assume *one*

Do not simplify the subscripts on covalent substances…keep the Greek prefix as is

<u>Lesson 5: Lewis Dot Diagrams for Covalent Compounds</u>

<u>Objective</u>: To draw Lewis strutures to represent molecules



A Covalent bond is the sharing of 2 electrons.

Covalent bonds share electrons in order to form a stable octet around each atom in the molecules. Hydrogen is the exception it only requires 2 electrons (a duet) to be stable.

How do we draw a covalent Lewis Dot Structure?

<u>Single Bonds-</u>Two electrons shared

1. Add up all the valance electrons of the atoms involved. ex CF_4

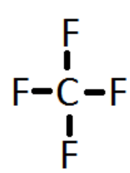
So C has 4 and F has 7 (x4 we have 4Fs) = 32 valence electrons

2. You need to pick the central atom. This is usually easy, this atom will be surrounded by the others. (Never H though since it can only form one bond.)

So C will be surrounded by F's.

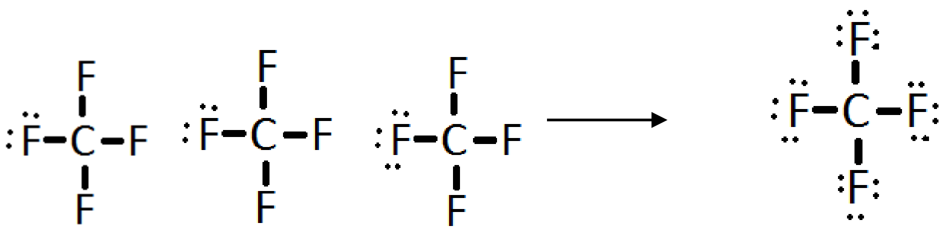
F
F C F
F

3. Now we create our skeleton structure by placing bonds in. A bond is a dash that represents 2 electrons.



We have now placed 8 electrons as 4 bonds. We have 32-8= 24 more to place.

4. Starting with the outer atoms add the remaining electrons in pairs until all the electrons have run out.



All 32 electrons are now in place, count the dots around each F. 6 dots and a bond (2 electrons) is 8. We have our octet.

The carbon has 4 bonds (2electrons) for its 8. DONE

<u>Double and Triple bonds-</u> four or six electrons being shared

Same rules apply until #4

1. Add up all the valance electrons of the atoms involved. ex CO_2

So C has 4 and O has 6 (x2) = 16 valence electrons

2. You need to pick the central atom. This is usually easy, this atom will be surrounded by the others. Never H.

So C will be surrounded by O's.

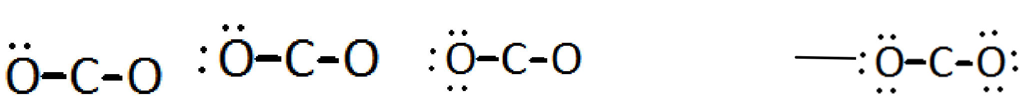
O C O

3. Now we create our skeleton structure by placing bonds in. A bond is a dash that represents 2 electrons.

O–C–O

We have now placed 4 electrons as 2 bonds. We have 16-4=12 more to place.

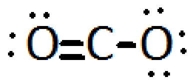
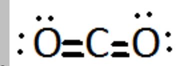
4. Starting with the outer atoms add the remaining electrons in pairs until all the electrons have run out.



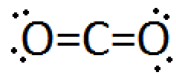
All 16 electrons are now in place, count the dots around each O. 6 dots and a bond (2 electrons) is 8. We have our octet.

The carbon has 2 bonds (2electrons) for its 4....?

We need 8, so move a pair of electrons from the O to between the C and O. It will share 2 pairs of electrons instead of 1. It now has a double bond instead of a single bond.

 carbon has 6 electrons, so move 2 from the other oxygen 

Carbon has 6 electrons, so move 2 from the other Oxygen. Now they all have an octet, it cleans up like this



<u>Make it symmetrical.</u>

When constructing the structures keep in mind the following:

- The dots surrounding the chemical symbol are the valence electrons, and each dash represents one covalent bond (consisting of two valence electrons)
- Hydrogen is always terminal in the structure
- The atom with the lowest ionization energy is typically the central atom in the structure
- The octet rule means there are 8 valence electrons around the atoms, but for hydrogen the maximum is 2 electrons

Lesson 6: Polar vs. Nonpolar **Date:______**

Objective: Distinguish between nonpolar and covalent bonds (two of the same nonmetals) and polar covalent bonds. To determine molecular polarity by the shape and distribution of charge.

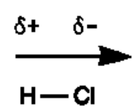
Depending on the relative electronegativity's of the two atoms sharing electrons, there may be partial transfer of electron density from one atom to the other. When the electronegativites are not equal, electrons are not shared equally and partial ionic charges develop. The greater the electronegativity difference, the more ionic the bond is. Bonds that are partly ionic are called ***polar covalent bonds***. ***Nonpolar covalent bonds***, with equal sharing of the bond electrons, arise when the electro negativities of the two atoms are equal.

Nonpolar Covalent Bond

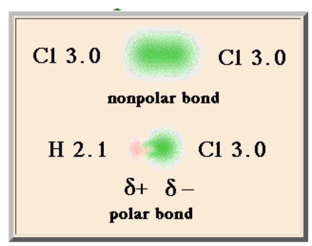
- A bond between 2 nonmetal atoms that have the same electronegativity and therefore have equal sharing of the bonding electron pair
- Example: In H-H each H atom has an electronegativity value of 2.1, therefore the covalent bond between them is considered nonpolar

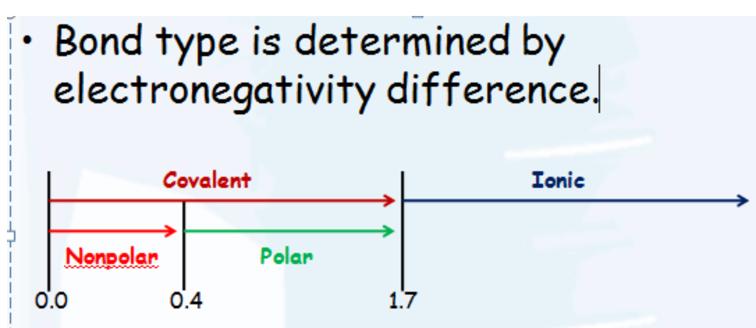
Polar Covalent Bond

- A bond between 2 nonmetal atoms that have different electronegativities and therefore have unequal sharing of the bonding electron pair
- Example: In H-Cl, the electronegativity of the Cl atom is 3.0, while that of the H atom is 2.1
- The result is a bond where the electron pair is displaced toward the more electronegative atom. This atom then obtains a partial-negative charge while the less electronegative atom has a partial-positive charge. This separation of charge or *bond dipole* can be illustrated using an arrow with the arrowhead directed toward the more electronegative atom.



The Greek letter *delta* indicates "partially".

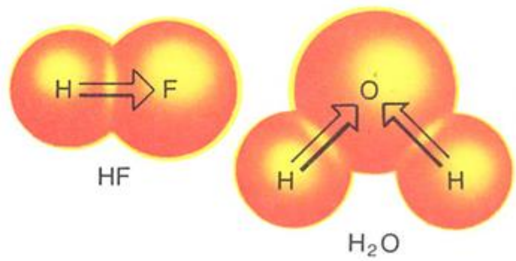




Molecule Polarity and Shape

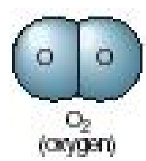
POLAR Molecules are ASYMETRICAL molecules that have tend to have an **UNEQUAL** sharing of **ELECTRONS.**
They are built out of two atoms that have differing electronegativites. There is sometimes unbonded e^- or lone pairs of electrons on the central atom. Doesn't pass the "mirror test"

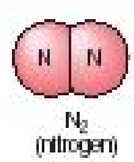
- Can't be folded to reflect itself



NONPOLAR Molecules are SYMMETRICAL molecules that tend to have

- **EQUAL** sharing of **ELECTRONS** . They are built out of two atoms that have similar electronegativites.
- There are **NO** unbonded e^- or **LONE** pairs around the **CENTRAL** atom
  - DOES pass the "mirror test"
  - CAN be folded to reflect itself





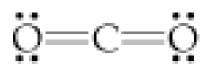
BEWARE! There are often POLAR BONDS inside NONPOLAR MOLECULES

VESPR -The Valence Shell Electron Pair Repulsion

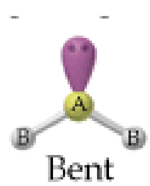
Valence electrons are arranged as far from one another as possible to minimize the repulsion between them (Like charges repel each other)

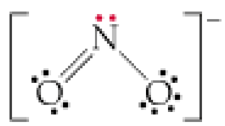
Possible Molecule shapes:

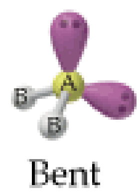
1. Linear (X_2 HX CO_2)

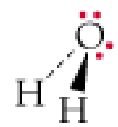


2. Bent (NO_2, H_2O)

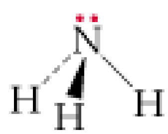




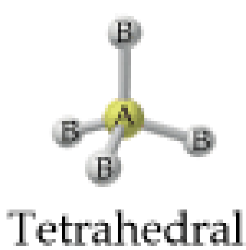


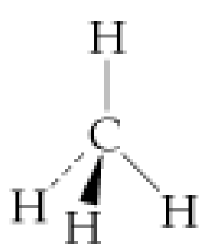


3. Pyramidal (NH_3)



4. Tetrahedral (CH_4 CCl_4)





Lesson 7: Intermolecular Forces of Attraction(IMF's)

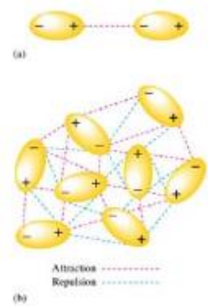
Objective: To describe the attractive forces between molecules

Intermolecular forces of attraction are attractive forces that only occur between COVALENT MOLECULES.
They are **WEAK** forces that act BETWEEN **MOLECULES** that hold molecules to EACH OTHER and only exist in GASEOUS and LIQUID states. They are considered WEAK forces because they are much weaker than CHEMICAL BONDS.

***REMEMBER: IMF's occur BETWEEN molecules, whereas BONDING occurs WITHIN molecules.**

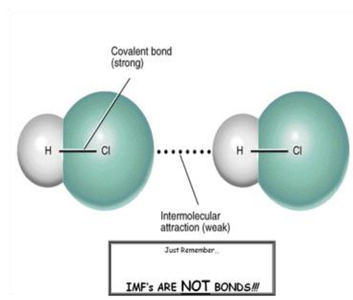
Why are they important? IMF's determine melting and boling points. The stronger the IMF, the higher the mp/bp.

THE TYPES OF INTERMOLECULAR FORCES IN ORDER OF DECREASING STRENGTH:



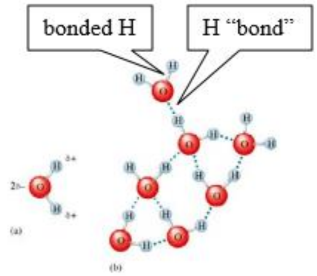
1. Dipole-dipole —the force of attraction that enables two polar molecules to attract one another.

Polar molecules are those which have an uneven charge distribution since their dipole moments do not cancel. Compounds exhibiting this type of IMF have higher melting and boiling points than those exhibiting weaker IMFs. Hydrochloric acid molecules are held to each other by this type of force. HCl—the chlorine pulls the electrons in the bond with greater force than hydrogen so the molecule is polar in terms of electron distribution. Two neighboring HCl molecules will align their oppositely charged ends and attract one another.



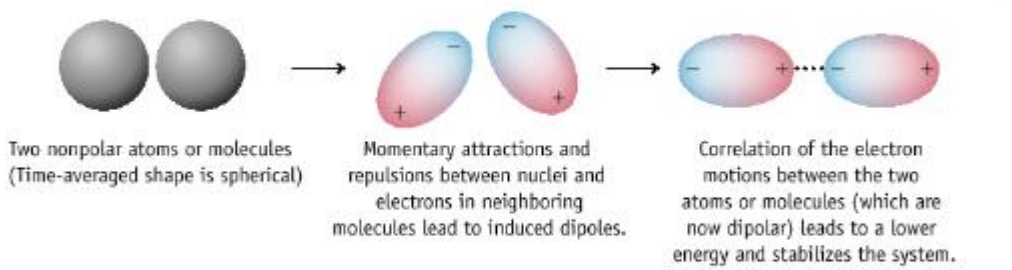
2. **Hydrogen bonding** —the force of attraction between the hydrogen atom of one molecule and an unshared electron pair on F, O, or N of a neighboring molecule (a special case of dipole-dipole).

This is the strongest IMF. Never confuse hydrogen bonding with bonded hydrogen. The unique physical properties of water are due to the fact that it exhibits hydrogen bonding. As a result of these attractions, water has a high boiling point, high specific heat, and many other unusual properties.

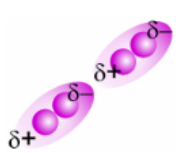


More info: http://programs.northlandcollege.edu/biology/Biology1111/animations/hydrogenbonds.html

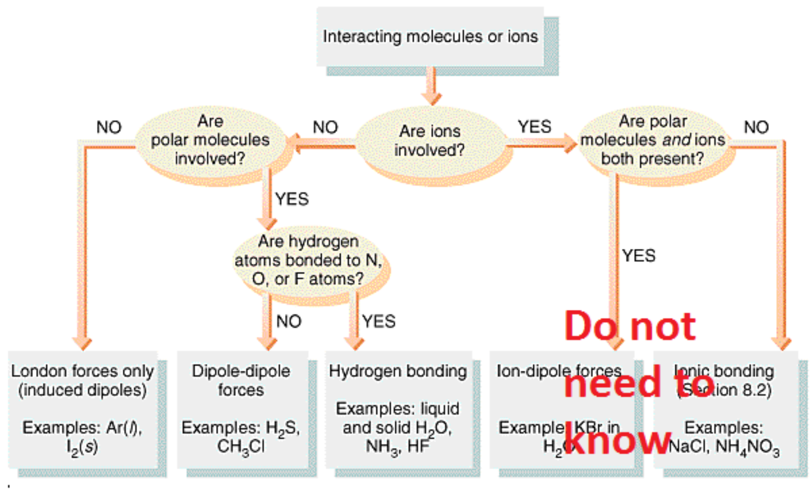
3. Van der waals forces AKA Induced dipole-induced dipole or London dispersion force—the force of attraction between two non polar molecules due to the fact that they can form temporary dipoles. Nonpolar molecules have no natural attraction for each other. This IMF is known by both names! Without these forces, we could not liquefy covalent gases or solidify covalent liquids.



To better understand **Van der waals forces**, examine the halogens. Halogens exist as diatomic molecules at room temperature and atmospheric conditions. F2and Cl2 are gases, Br2is a liquid and I2is a solid. Why? All of these molecules are completely nonpolar and according to theory, not attracted to each other, so one might predict they would all be gases at room temperature. ƒ Bromine exists as a liquid at room temperature simply because there is a greater attractive force between its molecules than between those of fluorine or chlorine.
Why?
Bromine is larger than fluorine or chlorine; it has more electrons and is thus more polarizable. Electrons are in constant motion so it is reasonable that they may occasionally "pile up" on one side of the molecule making a temporary negative pole on that end, leaving a temporary positive pole on the other end. This sets off a chain-reaction of sorts and this temporary dipole induces a dipole in its neighbors which induces a dipole in its neighbors and so on. Iodine is a solid since it is larger still, has even more electrons, is thus even more polarizable and the attractive forces are thus even greater. Note the two spheres representing I2 the diagram. Each iodine atom experiences a nonpolar covalent bond within the molecule. Be very clear that the IMF is between molecules of iodine, NOT atoms of iodine!



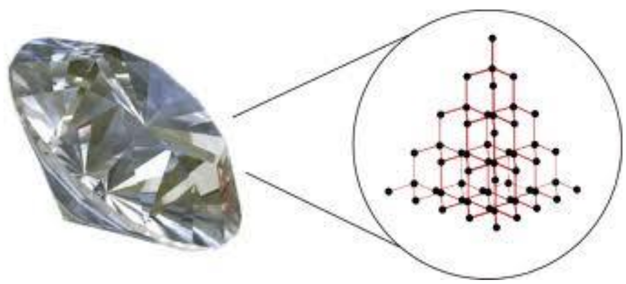
****** Ignore Ion-dipole forces in diagram below***



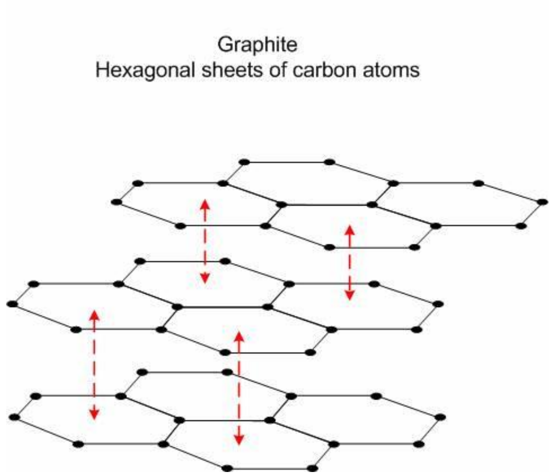
Interacting molecules or ions
Are ions involved?
NO
YES
Are polar molecules involved?
NO
YES
Are polar molecules and ions both present?
NO
YES
Are hydrogen atoms bonded to N, O, or F atoms?
NO
YES
London forces only (induced dipoles)
Examples: Ar(l), $I_2(s)$
Dipole-dipole forces
Examples: H_2S, CH_3Cl
Hydrogen bonding
Examples: liquid and solid H_2O, NH_3, HF
Ion-dipole forces
Example: KBr in H_2O
Ionic bonding (Section 8.2)
Examples: NaCl, NH_4NO_3
Do not need to know

Supplemental Lesson: Network Solids

The best example of a network solid is a “diamond”. Look at the model of a diamond below. Note that the carbon atoms are bonded together with **covalent bonds**. The basic building unit is an **atom** of carbon. The structure has a very definite tetrahedral crystal shape, because these atoms are arranged and **held rigidly in a fixed pattern**. A diamond is very hard (a “10” on the Moh’s Scale of Hardness…the highest value possible). In order to scratch a diamond, you must break 1000’s of very strong covalent bonds! Similarly, to melt (or boil) a network solid, like a diamond, you must break 1000’s of these covalent bonds. This involves considerable energy and is the reason for their high melting points. It is because of these high temperatures and their hardness that network solids are frequently used in industry as “abrasives” (on sandpaper and on the tips of drills for cutting tools). You don’t have to worry about them melting if it gets too hot from friction or being scratched and dulled when contacting most other surfaces. Network solids have the type of properties you would expect from atoms being held together via strong covalent bonds, e.g. diamonds. They have very high melting points and are practically insoluble; are mostly nonconductors (no free electrons or ions); and they are very brittle (atoms must maintain a fixed crystal structure, if they are pushed too close together …. they repel).



Graphite is also shown below. Note that it is also pure carbon, like a diamond. However, the covalent bonds only attach carbon atoms in 2 directions, not 3 like diamonds. The **dashed lines** between the layers of covalently bonded carbon atoms represent weak **Van der Waal forces**. Graphite is a 2 dimensional network solid. The strong covalent bonds only go in “plates”, in 2 directions. The “plates” are connected via weak VDW forces. Graphite STILL has a high melting point. – You must weaken/break all of its bonds (VDW and Covalent) to melt it. But since the weaker VDW forces are present and break easily, graphite is often used as a “dry lubricant”. If you squirt graphite dust into a lock, it will lodge between the lock’s moving metallic parts. When you put in a key and turn, the graphite structure will break apart between its “plates.” VDW’s break and make it turn more easily. Graphite also has free, “delocalized” electrons (it is a resonant structure… there really are no double bonds present, but free electrons) thus…graphite is a network solid that is capable of conducting electricity. This is not characteristic of most network solids. Of course, pencil lead is graphite. What bonds break when you write??? Silicon bonds like carbon to form a network structure. The computer industry depends on “silicon chips” ,which are made conductive by placing impurities in their structure. These then provide for free electrons and allow the chip to do its job. But, pure silicon does not conduct.



Many network solids are composed of various combinations of relatively few elements on the periodic table. The elements B, C, Al, Si are found in many network solids. They can be pure or combine with one another or combine with elements near them. For example, SiO_2, quartz is an example of a network solid. Corundum, Al_2O_3, is a network and a common abrasive used on “sandpaper”. Many gem stones, like diamond, are network solids. Emerald is made of the mineral “beryl”. Its formula is $Be_3Al_2\ (Si_6O_{18})$. Ruby is a form of corundum.

1. What is a Network solid? Give an example

2. What are some physical properties of Network solids?

<u>Review of Naming Ionic Compounds</u>

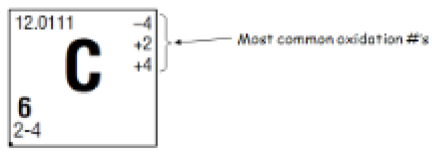
RECALL.... Example $2Ca(NO_3)_2$

Coefficient= in front of Formula
Subscript=small # after an atom
Subscript after ( ) = multiply everything inside by that #
# of atoms of each substance in the formula above= 2Ca, 2N, 6O

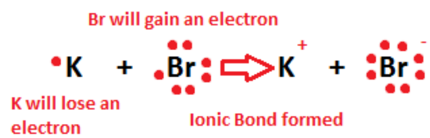
<u>OXIDATION NUMBERS:</u>

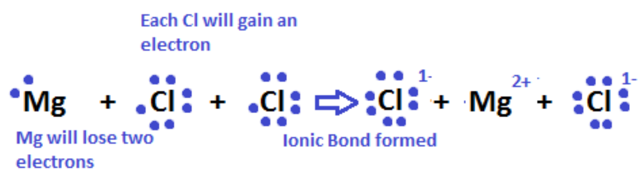
What are oxidation numbers and where are they located?

They are for keeping track of the movement of electrons and are found in the upper left hand of the element's box.



Let's build and name some compounds!



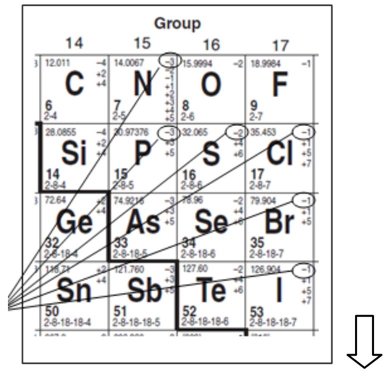


IUPAC = UNIVERSAL way of naming compounds; stands for "International Union of Pure and Applied Chemistry"

Compounds have a COMMON **name and a** CHEMICAL **name. There is a SYSTEMATIC method for naming ionic compounds.** When you are asked to name a compound, it will often tell you to name it according to IUPAC guidelines. This is just something the Regents throws in to confuse you… don't let it! IUPAC just stands for International Union of Pure and Applied Chemistry, and basically, these are the people who make all the rules for naming. So, if a question asks you to name something according to IUPAC, just name it the same way you always do!

<u>Writing Ionic formulas</u>

1. Write the symbols of the metal and nonmetal ions or polyatomic ion if there is one…ALWAYS CHECK <u>TABLE E</u> FIRST!
2. Put any polyatomic ions in ( ).



3. Write the NEGATIVE (nonmetal) ion's charge. The negative ion is always LAST in the formula Use the first charge listed on the P.T.

4. If there is a <u>Roman numeral in ( )</u>, this is the <u>charge on the metal ion</u>. If no Roman numeral, look up the charge of the metal ion in the P.T.

5. **CRISS-CROSS** and reduce the charges (oxidation numbers), turning them into $_{\text{subscripts.}}$ (Subscripts tell the <u>*ratio*</u> of ions in the substance.) Do not write 1's

6. Check: make sure the positive and negative charges cancel each other by multiplying charges & $_{\text{subscripts.}}$

7. Re-write the formula with only symbols and $_{\text{subscripts.}}$

IONIC
Compounds

METAL + nonmetal
or

has a **polyatomic ion**
(<u>check Table E</u>!)

Examples:

Name	Formula		Name	Formula
sodium chloride	NaCl		cobalt(II) chloride	$CaCl_2$
potassium oxide	K_2O		nickel(II) oxide	NiO
calcium dichromate	$CaCr_2O_7$		tin(II) sulfide	SnS
ammonium hydroxide	NH_4OH		tin(IV) sulfide	SnS_2
magnesium acetate	$Mg(C_2H_3O_2)_2$		iron(III) sulfate	$Fe_2(SO_4)_3$

Writing Ionic names

1. Look on the Periodic Table to see how many oxidation states are listed for the metal in the formula.
   - One oxidation state: NO ROMAN NUMERAL
     (skip to bottom of page for examples)
   - Two or more oxidation states: a Roman numeral in ( ) is needed to tell the CHARGE OF ONE METAL ION
     (continue on to Steps 2 & 3)

2. To figure out the charge of the metal ion, write the NEGATIVE ion's charge (the top one listed on the P.T., or if it's polyatomic the charge on Table E), and multiply the negative charge by the $_{\text{subscript}}$ for that ion to get the total negative charge in the compound.

3. Since compounds are *neutral*, the positive charge for the metal ion(s) must equal the total negative charge found in step 2. Divide the total negative charge by the $_{\text{subscript}}$ on the metal ion to get the positive charge of ONE metal ion. This goes as a Roman numeral in ().

2 elements in formula

metal name + nonmetal root with **–ide** *ending*

3 or more elements in formula

(has a **polyatomic ion**)

metal name + polatomic ion name on Table E

Examples:

Formula	**Name**
NaCl	sodium chloride
CaO	calcium oxide
FeO	iron(II) oxide
Fe_2O_3	iron(III) oxide
$Ca(NO_3)_2$	calcium nitrate
NH_4Cl	ammonium chloride
$FeSO_4$	iron(II) sulfate
$Fe_2(SO_4)_3$	iron(III) sulfate
$NaHCO_3$	sodium hydrogen carbonate

STUDY GUIDE BONDING:

Make a bond: **release Energy (exothermic)**, bonded atoms have **lower potential energy.** Greater energy released, more stable bond.

$$A + B \rightarrow AB + \text{energy}$$

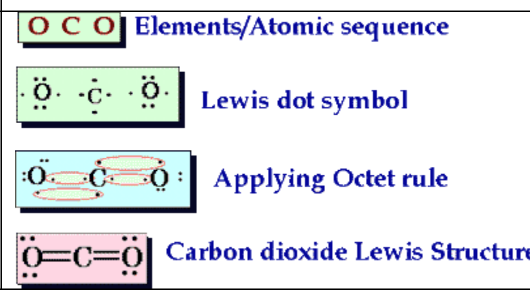
Break a bond: absorb Energy (endothermic), free atoms have **higher potential energy.**

$$AB + \text{energy} \rightarrow A + B$$

The Octet Rule: Atoms want a complete octet of eight valence electrons in order to be stable. To reach this goal of 8 valence electrons, atoms react by gaining, losing, or sharing electrons. Their goal is to have the **electron configuration of a noble gas.**

Bond: Attraction of positive nucleus of one atom for the negatively charged electrons of the other atom

Types of Bonds:

Metallic	Ionic	Covalent
Metals "**sea of mobile valence electrons**" Properties: - hard - always good conductors, because of **mobile valence electrons** - high melting point - malleable (made into sheets) - ductile (made into wires)	Metal + Nonmetal **Transfer of electrons** from metal to nonmetal **Electronegativity difference** is **> 1.7** Properties: - hard - good conductor as Liquid or (aq) ONLY because of **free moving ions** - high melting and boiling point	Nonmetal + Nonmetal **Share electrons** To achieve a stable arrangement of electrons Properties -Soft - poor conductors of heat and electricity because no charged mobile particles - low melting and boiling point
Drawing Lewis Dot Diagrams	NaCl → Na^+ + [:Cl:]$^-$	

Types of Covalent Bonds/molecular bonds:

Nonpolar Bonds: there is an **equal sharing** of electrons, **no pull** on electrons. Formed when atoms in the bond have the same electronegativity. All diatomic molecules have nonpolar **bonds** (H_2 O_2 N_2 Cl_2 Br_2 I_2 F_2)
Electronegativity Difference between atoms in the bond is equal to **ZERO**!

Polar Bonds: there is **unequal sharing** of electrons, **pull** on electrons. Found when atoms in the bond have a **different electronegativity**. The element with a **higher electronegativity** has a greater attraction for electrons and ends up with a **partial negative charge**. The other end of the polar covalent bond, with a lower electronegativity acquires a partially positive charge. **Electronegativity difference** is between **0.4 - 1.7**

Types of Molecules:

Nonpolar Molecules:

1- All molecules that are made up of **nonpolar bonds only** are nonpolar molecules.(example all diatomic molecules)
2- Compounds that are symmetrical: have identical parts on each side of it's axis are nonpolar
   Memorize: CH_4, CO_2,
3- Are soluble in nonpolar solvents "like dissolves like"

Polar Molecules:

1- molecules made up of **polar bonds** that are asymmetrical: lack identical parts on each side of it's axis
   Memorize: H_2O, CH_3Cl , NH_3, H_2S
2- causes molecules to have **dipole-dipole attractions**
3- are soluble in polar solvents "like dissolves like"

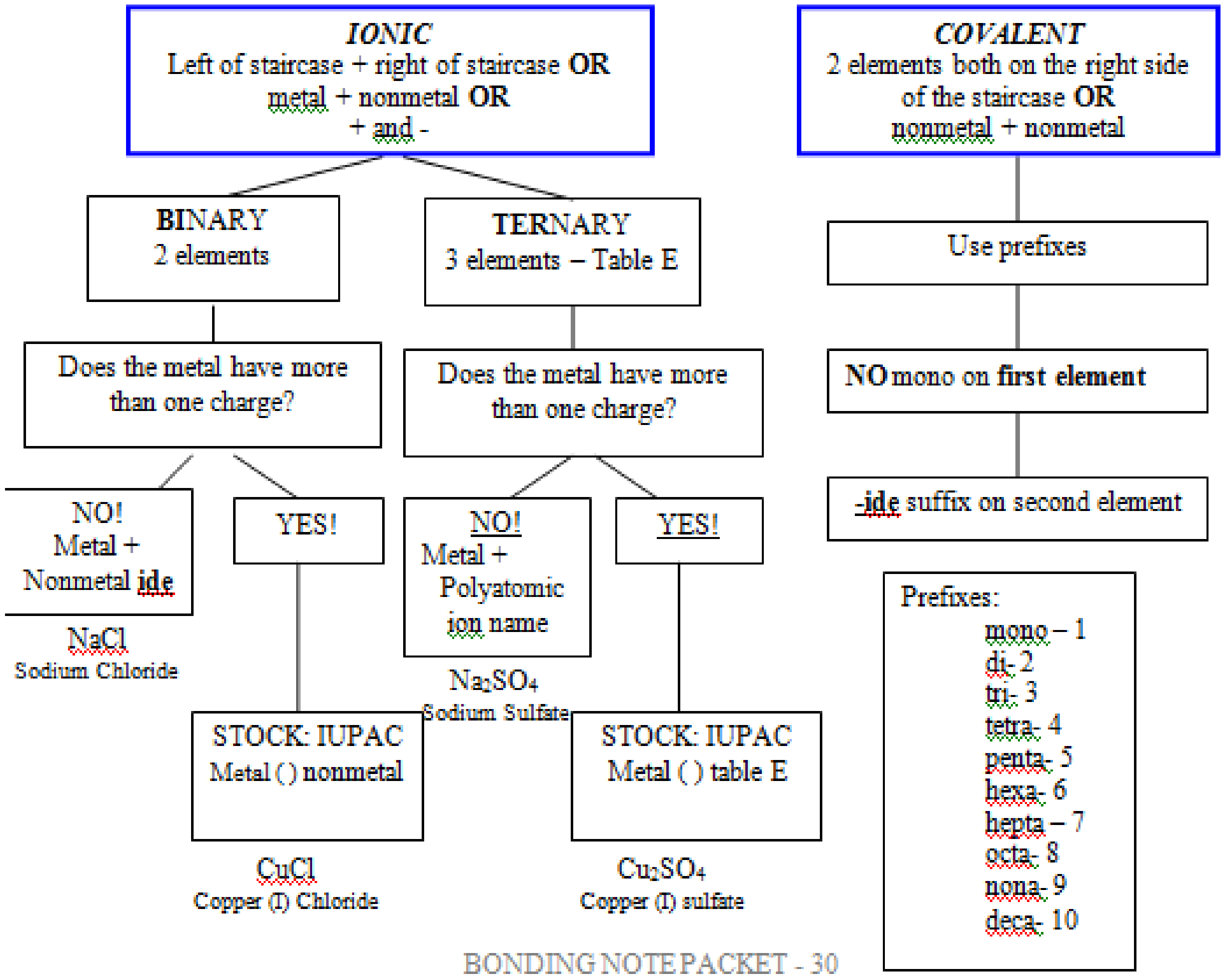
Single: One bond (pair of electrons) **between** atoms, **share 2 electrons**,
Double: Two bonds ((pairs of electrons) **between** atoms, **share 4 electrons**,
Triple: **Three bonds** (pairs of electrons) between atoms, **share 6 electrons**.

Intermolecular Forces; Hydrogen Bonds:
Intermolecular attraction between molecules containing

Hydrogen + F, O, or N.

Hydrogen in one molecule and an F, O, or N in another molecule. This type of intermolecular force is **responsible for** the extremely **high boiling point** of H_2O, NH_3 and HF; which all have strong hydrogen bonds.

STUDY GUIDE: NAMING



<u>For IUPAC naming:</u>**
When you change from a name to a chemical formula, you MUST write **oxidation numbers first**, and then criss-cross to find your neutral Chemical Formula!

<u>Roman numerals:</u>

1. I	4. IV	7. VII
2. II	5. V	8. VIII
3. III	6. VI	9. IX