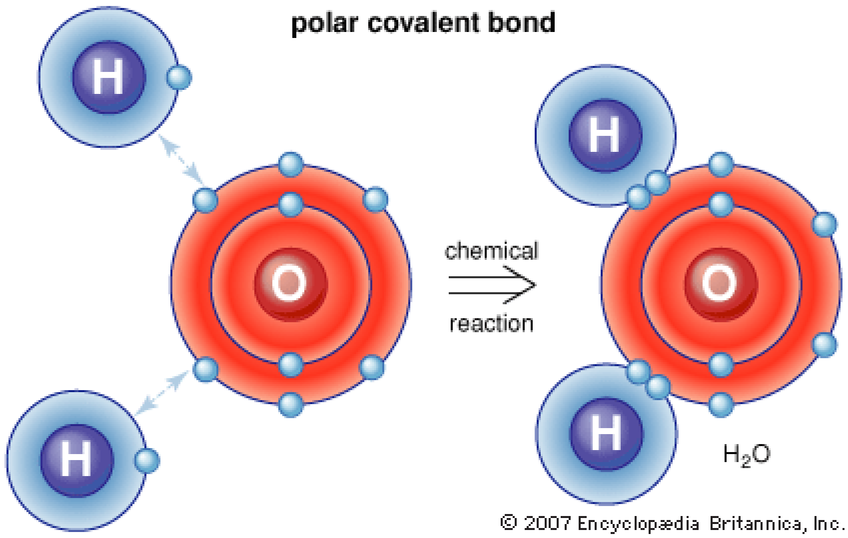
***STUDENT*** # UNIT 4 ***STUDENT***

# BON::DING
# & Naming Compounds




## Vocabulary:

1. Molecule
2. Compound
3. Bond
4. Octet Rule
5. Exothermic
6. Endothermic
7. Ionic Bond
8. Covalent Bond
9. Oxidation number
10. Polyatomic ions
11. Stock system
12. Binary compound
13. Ternary compound
14. Polar molecule
15. Nonpolar molecules
16. Intermolecular forces (IMF's)

**Review** – What is a compound? A substance that has atoms of 2 or more ______________________ chemically bonded together (Ex: NaCl)

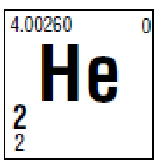
**BOND** = forces of __________________ between the protons (nucleus) of one atom and the electrons of another atom

Only ____________ electrons participate in bonding

Bonds are formed as a result of a ____________________________

***LAW of CONSERVATION of ENERGY/MASS/CHARGE*** = during a chemical reaction, energy, mass, and charge can ______ be created or destroyed (That means that all three are ______________________).

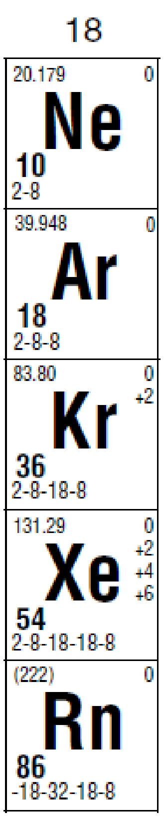
OCTET RULE → atoms bond together to get ___ valence electrons around them


Recall: Noble Gases already have 8 valence electrons → this is why they don't react/bond with other elements

- Which Noble Gas, He, Ne, Ar, or Kr would be most likely to form a bond with another element? (Hint: Look at your Periodic Tables) ______
- Why do you think that is?

  __________________________________________________

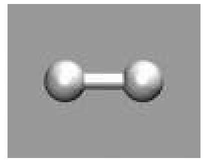



- Which element Na, O, F, P, or S would be most likely to form a bond with Kr? ______
- Why do you think that is?

  __________________________________________________

## ENERGY & BONDING

Energy associated with bonding:



- ____________ **Bond formation** = Energy is ____________
  - go from ________ energy (unhappy atoms) to ________ energy (happy atoms)
  - creating a bond creates ____________

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

$$\boxed{A + B \rightarrow C + \text{energy}}$$

*NOTE: "energy" on ____________ side of arrow

Ex:

$CaCl_2 + NaHCO_3 + H_2O \rightarrow CaCO_3 + CO_2 +$ ____________

- **Breaking Bonds (NOT spontaneous)** = Energy is ____________
  - go from ________ energy (happy atoms) to ________ energy (unhappy atoms)
  - ripping two atoms apart takes ____________

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

$$\boxed{A + \text{energy} \rightarrow B + C}$$

*NOTE: "energy" on ____________ side of arrow

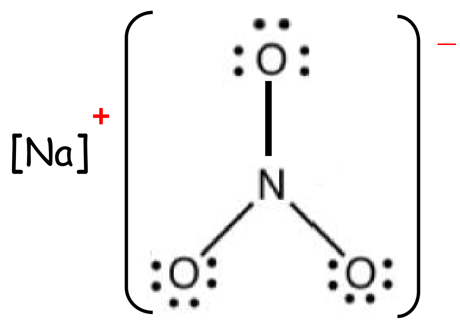
Ex: an ice pack → chemicals combine, bonds are broken, and energy is consumed → you feel "cold" because you're losing heat to the ice pack

# The Two Most Common Types of Compounds:

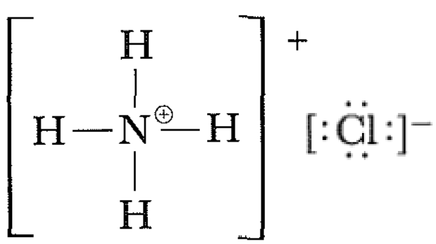
1. **IONIC** – compounds formed by the TRANSFER of ELECTRONS from one atom (or polyatom) to another

   a. METAL LOSES e⁻ to a NONMETAL
      Ex:

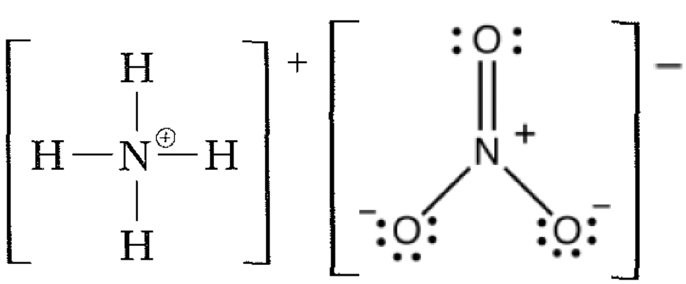
   b. METAL combines with a POLYATOMIC ION



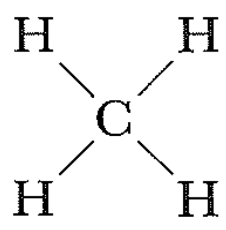
   c. POLYATOMIC ION combines with a NONMETAL



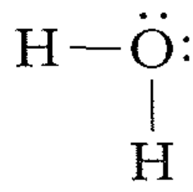
   d. POLYATOMIC ION combines with another POLYATOMIC ION



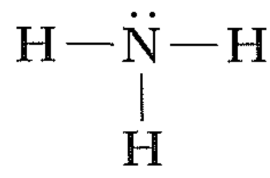
**2. COVALENT** – compounds formed when 2 or more nonmetals SHARE ELECTRONS



methane



water



ammonia

**<u>IONIC vs. COVALENT SUMMARY</u>:**

| NATURE of the ELEMENTS | CLASSIFICATION |
|---|---|
| **Metal + Nonmetal** | |
| Metal + Polyatomic Ion | |
| Nonmetal + Polyatomic Ion | |
| Polyatomic Ion + Polyatomic Ion | |
| **Nonmetal + Nonmetal** | |

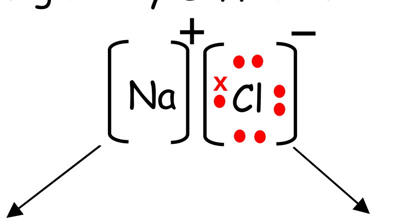
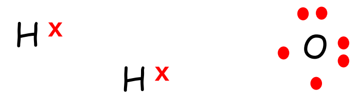
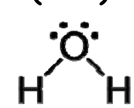
**METALS = pure substances found to the left of the "staircase"**

- atoms of a metal do not bond w/ other metal atoms
- metals "share" a "sea of ____________ valence electrons"
  - allows metals to conduct electric ______________ (________ energy)

**Click here to see animation: [http://www.ausetute.com.au/metallic.html](http://www.ausetute.com.au/metallic.html)**

# IONIC vs. COVALENT CONTINUED:

(http://www.mhhe.com/physsci/chemistry/animations/chang_7e_esp/bom1s2_11.swf)

<table>
<tr><th>IONIC</th><th>COVALENT</th></tr>
<tr>
<td>
Electron(s) ____________________<br>
<br>
(click the link below for the animation)<br>
(http://www.visionlearning.com/library/flash_viewer.php?oid=1349&mid=55)<br><br>
Physical Properties:
<ul>
<li>High m.p./b.p (melting point/ boiling point) due to strong BONDS</li>
<li>Electrolyte (conducts electric current in solution)</li>
<li>Hard</li>
</ul>
Electronegativity Difference = 2.3<br>
<br>
(0.9) (3.2)
</td>
<td>
Electrons ____________________<br>
<br>
(click the link below for the animation)<br>
http://web.visionlearning.com/custom/chemistry/animations/CHE1.7-an-H2Obond.shtml<br><br>
Physical Properties:
<ul>
<li>Low m.p./b.p. (melting point/ boiling point) due to weak IMF's</li>
<li>Nonelectrolyte (does not conduct elec. current in solution)</li>
<li>Soft</li>
</ul>
Electronegativity Difference = 1.4<br>
(3.5)<br>
<br>
(2.1) (2.1)
</td>
</tr>
</table>

- The greater the ______________________ difference between two elements the greater the percent ______________________
- The closer the ______________________ difference is to zero, the greater the percent ______________________

**Electronegativity Scale**

0.0 ------------------------------------------------- 4.0

** We can classify a compound as ionic or covalent in one of two ways:

1) Look at the elements in the compound (TRUMP CARD):
   - METAL and a NONMETAL = ____________________________
   - Two NONMETALS = ______________________________
2) Do the math:
   - The GREATER the electronegativity difference between the elements, the GREATER the ______________________________
   - The CLOSER the electronegativity difference between the elements, the GREATER the ________________________________

**Practice:** Find the electronegativity difference for each of the following compounds and then state whether each is ionic or covalent.

| Compound | Electronegativity difference | Ionic or covalent? |
|---|---|---|
| 1. $H_2O$ | | |
| 2. CaO | | |
| 3. NaCl | | |
| ***4. $FeCl_3$*** | | |
| 5. $H_2O_2$ | | |
| 6. $P_2O_6$ | | |
| ***7. LiH*** | | |
| 8. $CH_4$ | | |

# I. NAMING AND WRITING CHEMICAL FORMULAS:

Each element within a compound has its own "charge." We call these charges ____________________________.

| | |
|---|---|
| 12.0111 | –4<br>+2<br>+4 |
| **C** | |
| **6**<br>2-4 | |

Most common oxidation #'s

## A. Rules for assigning OXIDATION STATES (numbers):

**1)** ______________________ (elements not bonded to any other type of element) have an oxidation number of __________. This includes any formula that has *only* one chemical symbol in it (single elements & diatomic elements).

Examples: ____________________________

**2)** In ________________ (remember, they are neutral and have 2+ different elements bonded together), the sum of the ______________ must add up to _________ so the ions within a compound have oxidation numbers equal to their oxidation # found on periodic table/individual charges.

Ex: $NaCl$ Ex: $Mg_3N_2$

Ex: $HNO_3$

*We almost always write the (+) element first and the (-) element last in a compound formula.

EXAMPLE: $H\ Cl$

EXCEPTION to this rule: $N\ H_3$

3) **GROUP 1 METALS** always have an oxidation number of ____ when in a compound (bonded to another species).
**GROUP 2 METALS** always therefore have a ____ oxidation number when located within a compound.

Ex: $Na\ Cl$ $Ca\ Cl_2$

4) **FLUORINE** is always a ____ in compounds. The other **HALOGENS** (ex: Cl, Br) are also ____ as long as they are the most electronegative element in the compound.

Ex: $Na\ F$ $H\ Br$ $Li\ Cl$

5) **HYDROGEN** is a ____ in compounds unless it is combined with a metal (and is at the end of the formula), then it is ____.

Ex: $H\ Cl$ $Li\ H$

6) **OXYGEN** is **USUALLY** ____ in compounds.

Ex: $H_2O$ $C\ O_2$

When combined with fluorine (F), which is more electronegative, oxygen is ____.

Ex: $O\ F_2$

When in a **PEROXIDE** oxygen is ____. A peroxide is a compound that has a formula of ____________.

Ex: $H_2O_2$ $Na_2O_2$

**7)** The sum of the oxidation numbers in polyatomic ions must equal the charge on the ________ (see Table E).

Ex: $(Cr_2O_7)^{-2}$ **Cr:**
**O:** ______________

*Compounds containing ______________________________ have ______________________ and ____________________ bonds.

**Practice:** Identify and list the polyatomic ions in the following compounds (use Table E). We'll do the first one together as an example.

| | |
|---|---|
| $(NH_4)_3PO_4$ | |
| $NaHCO_3$ | |
| $NH_4NO_3$ | |
| $H_2SO_4$ | |
| $KNO_3$ | |

Ex 1: $OH^-$

- Oxidation number of Oxygen is ALWAYS _____ (except in peroxides)
- Therefore H must have a charge of _____

***Note: Must add up to –1 (that is the charge on $OH^-$)*

Ex 2: $PO_4^{3-}$

- Oxidation number of O = _____ (4 x (_____) = -8 total)
- Oxidation number of P = _____ (is that one of the choices on the table?)

## B. IUPAC NAMING RULES for IONIC COMPOUNDS:

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

- Compounds have a __________ name and a ____________ name
- There is a SYSTEMATIC method for naming ionic compounds

**FORMULA → CHEMICAL NAME**

1. Name the ____________ element ________ (or the METAL)
2. Name the ____________ element ___________ (or the NONMETAL)
   - Replace the ending of the negative element with "IDE"

     Ex: fluorine → _____________

     oxygen → _____________

Examples:

$NaCl$ → ______________________ (table salt)

$K_2O$ → ______________________

$CaBr_2$ → ______________________

$Al_2S_3$ → ______________________

$LiI$ → ______________________

**CHEMICAL NAME → FORMULA**

Examples → lithium bromide → ______________________

Sodium fluoride → ______________________

Magnesium chloride → ______________________

Potassium iodide → ______________________

3. POLYATOMIC IONS are named exactly as they are seen on TABLE E

   Examples: $NaOH$ → ____________________________

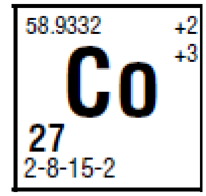
   $KNO_3$ → ____________________________

   Ammonium hydroxide → ____________

   Calcium phosphate → ____________

4. TRANSITION METALS tend to have more than one oxidation number so you must use a ________________ to indicate their oxidation number within a compound. The roman numeral appears in parentheses ________ the element symbol (STOCK SYSTEM)

Example: Cobalt chloride ***could*** have a formula of $CoCl_2$ or $CoCl_3$ since cobalt can have an oxidation number of +2 or +3.




$CoCl_2$ $CoCl_3$

Name: ____________________ Name: ____________________

Write the formula for the following

1) zinc oxide = ______________________

2) iron(II) chloride = ______________________

3) mercury(I) sulfide = ______________________

## C. IUPAC NAMING RULES for COVALENT COMPOUNDS:

**Recall: Covalent = _______________ + _______________**

**Binary Compound** = ___ elements bonded together (Ex: $CO_2$)
**Ternary Compound** = ___ elements bonded together (Ex: $C_6H_{12}O_6$)

The procedure for naming covalent compounds is very similar to the procedure for naming ionic compounds. The only difference however, is that you use ________________ to designate how many of each element you have in the covalent compound. The following are the prefixes you will use:

| # 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 |

*One exception: Drop the MONO prefix if there is only one atom of the _________ element in the compound name
**Final O's or A's of prefix are dropped when an element begins with a VOWEL (Ex: Carbon monoxide)

### RULES for Writing Chemical Formulas of Covalent Compounds:

1. _________ electronegative element is written _________.
2. _________ electronegative element is written _________.
3. ___________________ tell you the _________ of each element in the formula (Example: $CO_2$ = carbon ***di***oixide)

Name the following covalent compounds:

| Chemical Formula | Chemical Name |
|---|---|
| $CO$ | |
| $CO_2$ | |
| $N_2O$ | |
| $N_2O_5$ | |
| $CCl_4$ | |
| $SF_6$ | |

Give the formulas for the following compounds:

| Chemical Name | Chemical Formula |
|---|---|
| Nitrogen monoxide | |
| Carbon Tetrafluoride | |
| Bromine pentachloride | |
| Dinitrogen pentasulfide | |
| Sulfur tetrabromide | |

# II. LEWIS DOT DIAGRAMS for IONIC COMPOUNDS

- Gives us a ________________ of the compound
- Shows us the ______________________ 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:**

1) Draw all elements (of the given compound) and their individual Lewis Dot Diagrams
   Example: NaCl

2) Draw ELECTRONS going from the METAL to the NONMETAL. Your goal is to match up all UNPAIRED ELECTRONS (there should be none left unpaired when done).

3) Redraw your now bonded compound with dots and appropriate new CHARGES.

Draw the Lewis Dot Diagram for the following ionic compounds:

| | | |
|---|---|---|
| 1. KF | | |
| 2. BaS | | |
| 3. $AlBr_3$ | | |

# III. LEWIS DOT DIAGRAMS FOR COVALENT COMPOUNDS

http://www.d.umn.edu/~pkiprof/ChemWebV2/VSEPR/index.html

1) Write the ELEMENT SYMBOLS and draw their VALENCE ELECTRONS. If there are more than TWO atoms, place the LEAST ELECTRONEGATIVE in the CENTER.

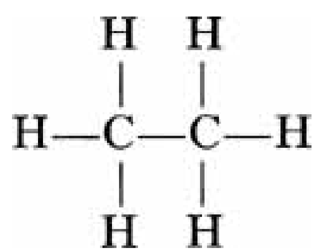
   Example: $H_2O$

2) Draw DASH/LINES (between the elements) connecting all lone electrons—your goal is to "pair up" ALL unpaired electrons

   Example (cont'd):

- ONE DASH/LINE is called a ________________________

  EX: C - C

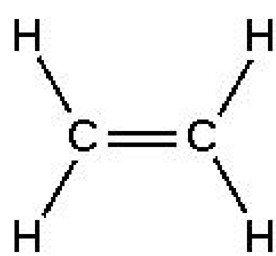


  # of $e^-$ being shared between carbons = _____

  # of $e^-$ pairs being shared between carbons = _____

- TWO DASH/LINES is called a ________________________

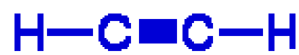
  EX: C = C



  # of $e^-$ being shared between carbons = _____

  # of $e^-$ pairs being shared between carbons = _____

- THREE DASH/LINES is called a ______________________

EX: $C \equiv C$ 

# of $e^-$ being shared between carbons = _____

# of $e^-$ pairs being shared between carbons = _____

Draw the Lewis Dot Diagram for the following covalent compounds:

| | | |
|---|---|---|
| 1. HCl | | |
| 2. $N_2$ | | |
| 3. $CCl_4$ | | |
| 4. $NH_3$ | | |

# IV. POLAR vs. NONPOLAR MOLECULES:

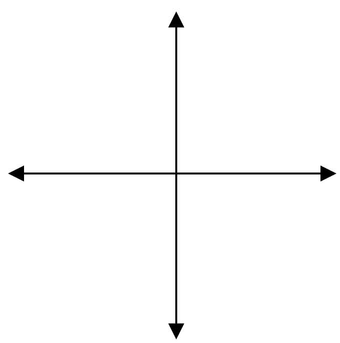
**Molecule** = A ________________ bonded substance; always 2 or more ________________ bonded together

*________________ compound = ________________ compound

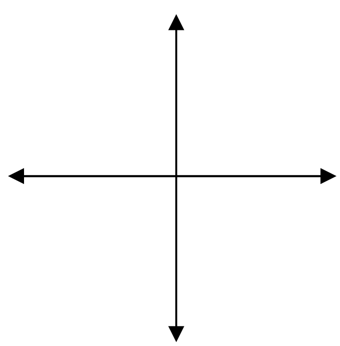
A.) **POLAR Molecules:**

- ________________ molecules
- ________________ sharing of ________________
  - Doesn't pass the "mirror test"
  - Can't be folded to reflect itself
- 2 atoms → different elements/electronegativities
- More than 2 atoms → unbonded $e^-$ or ________ pairs around the ________________ atom

Ex: HCl



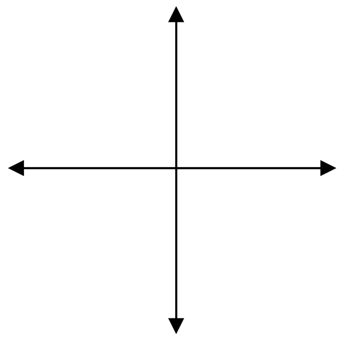
Ex: $H_2O$



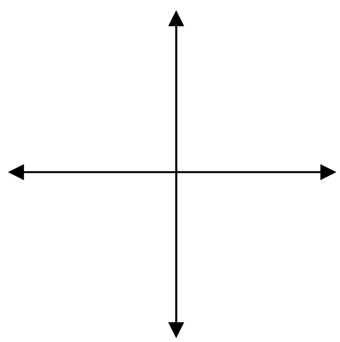
B.) **NONPOLAR Molecules:**

- ____________________ molecules
- __________ sharing of ___________________ *or*
  - ➢ DOES pass the "mirror test"
  - ➢ CAN be folded to reflect itself
- 2 Atoms → same element/electronegativities
- More than 2 atoms → ____ unbonded $e^-$ or ________ pairs around the _____________ atom

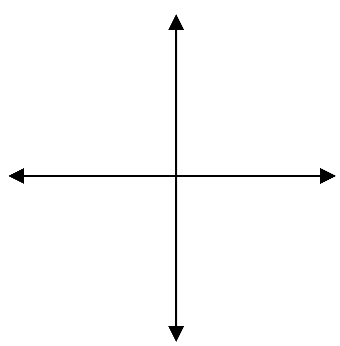
Ex: $Cl_2$



Ex: $CO_2$



Ex: $CCl_4$

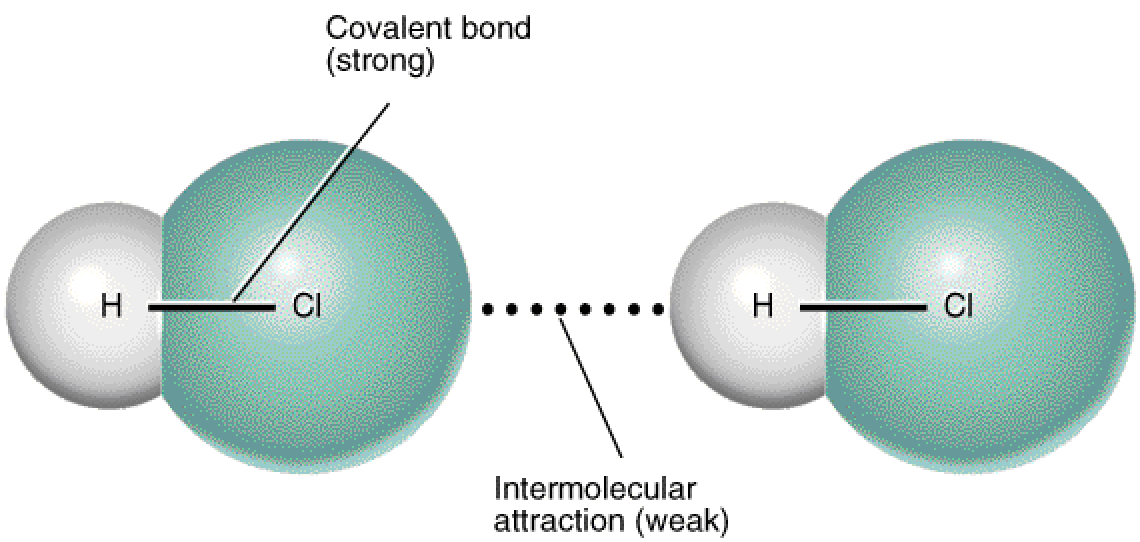


**BEWARE!** There are often POLAR BONDS inside NONPOLAR MOLECULES (look back at the previous 2 examples)

C. **Intermolecular forces, A.K.A. IMF's**

- ONLY IN COVALENT MOLECULES, NEVER IONIC COMPOUNDS!
- ________ forces that act BETWEEN ________________ that hold molecules to EACH OTHER
- Only exist in _______________ & ________________ states
- Called WEAK forces because they are much weaker than CHEMICAL BONDS

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




http://www.northland.cc.mn.us/biology/Biology1111/animations/hydrogenbonds.html

Just Remember…

**IMF's ARE <u>NOT</u> BONDS!!!**

| Type of IMF | Description/Example(s) |
| --- | --- |
| **London dispersion forces (LDF's)** | • Weakest of all the IMF's<br>• Only important for NONPOLAR molecules<br>• ______________________________<br>• Electron-electron repulsion creates BRIEF DIPOLES in atoms/molecules<br>http://antoine.frostburg.edu/chem/senese/101/liquids/faq/h-bonding-vs-london-forces.shtml |
| **Dipole (dipole-dipole)** | • Molecules such as HCl have both POSITIVE and a NEGATIVE ends, or POLES<br>• Two poles = ____________<br>• Results from an UNEQUAL/ASYMMETRICAL sharing of electrons<br>• DIPOLE-DIPOLE = two molecules with permanent dipoles are attracted to one another ______________________<br>• DIPOLE MOMENT = measure of the _____________ of the dipole within a molecule (POLARITY)<br>• The GREATER the difference in ELECTRONEGATIVITY between atoms, the GREATER the POLARITY/DIPOLE MOMENT<br>• The HIGHER the dipole moment, the STRONGER the intermolecular forces (IMF's)<br>• The stronger the IMF's, the higher the m.p. and b.p.<br>http://chemmovies.unl.edu/ChemAnime/DIPOLED/DIPOLED.html |
| **Hydrogen Bonds** | • Specific type of ___________ interaction<br>• In a POLAR BOND, hydrogen is basically reduced to a BARE PROTON w/ almost no ATOMIC RADIUS<br>• _______________ of all IMF's by far<br>• Only occur in molecules containing _____________ AND _____________, _____________, or _____________<br>http://programs.northlandcollege.edu/biology/Biology1111/animations/hydrogenbonds.html |

## UNIT OBJECTIVES:

- ✓ Compounds can be differentiated by their chemical and physical properties
- ✓ Two major categories of compounds are ionic and molecular (covalent) compounds.
- ✓ Chemical bonds are formed when valence electrons are: transferred from one atom to another (ionic); shared between atoms (covalent); mobile within a metal (metallic).
- ✓ In a multiple covalent bond, more than one pair of electrons is shared between two atoms. Unsaturated organic compounds contain at least one double or triple bond.
- ✓ Molecular polarity can be determined by the shape and distribution of that charge. Symmetrical (nonpolar) molecules include $CO_2$, $CH_4$, and diatomic elements. Asymmetrical (polar) molecules include $HCl$, $NH_3$, and $H_2O$.
- ✓ When an atom gains one or more electrons, it becomes a negative ion and its radius increases. When an atom loses one or more electrons, it becomes a positive ion and its radius decreases.
- ✓ When a bond is broken, energy is absorbed. When a bond is formed, energy is released.
- ✓ Atoms attain a stable valence electron configuration by bonding with other atoms. Noble gases have stable valence electron configurations and tend not to bond.
- ✓ Physical properties of substances can be explained in terms of chemical bonds and intermolecular forces. These properties include conductivity, malleability, solubility, hardness, melting point, and boiling point.
- ✓ Electron-dot diagrams (Lewis structures) can represent the valence electron arrangement in elements, compounds, and ions.
- ✓ Electronegativity indicates how strongly an atom of an element attracts electrons in a chemical bond. Electronegativity values are assigned according to an arbitrary scale.
- ✓ The electronegativity difference between two bonded atoms is used to assess the degree of polarity in the bond.
- ✓ Metals tend to react with nonmetals to form ionic compounds. 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.