Unit 9: CHEMICAL BONDING

Unit 9: Bonding:

1. Electronegativity
2. Intramolecular Bonding
3. Intermolecular Bonding
4. Drawing Lewis Structures
5. Lewis Structures for Polyatomic Ions
6. Exceptions to the Octet Rule
7. Resonance Structures
8. Molecular Shapes
9. Molecular Polarity

1. Electronegativity

Valence Electrons

- The outer electrons involved in bonding
- Number of valence electrons will be between 1 and 7

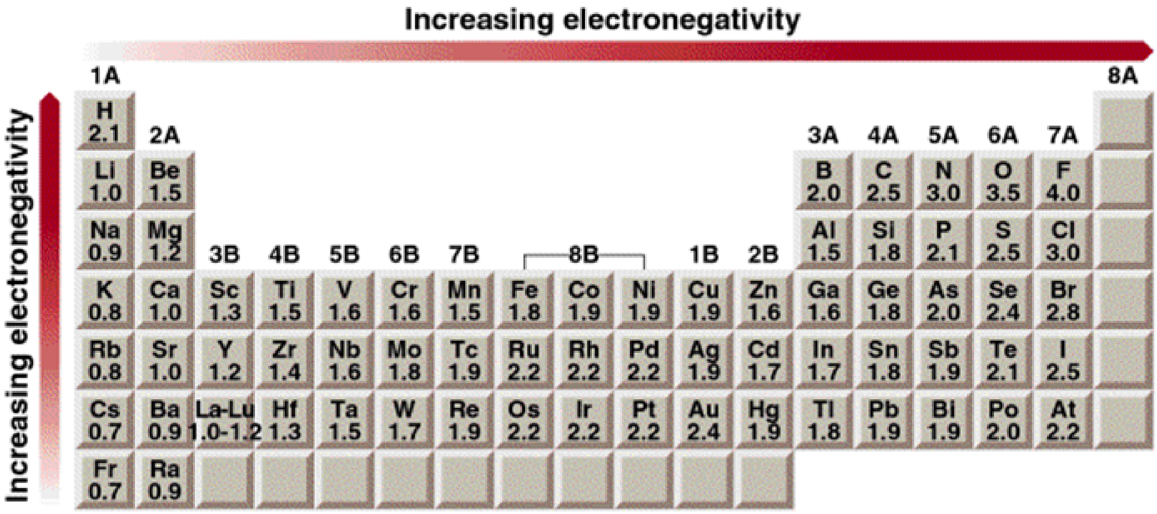
The representative metals are on the left side of the staircase, and most of them have 1, 2, or 3 valence electrons.
For metal atoms with 1, 2, or 3 valence electrons, it is easier to <u>give them away</u> in order to become isoelectronic with a noble gas, resulting in a + charge (cation).

In general, non-metals have 4-7 valence electrons, so it is easier in most cases to gain electrons to become isoelectronic with a noble gas.

Electronegativity

- The ability of an atom to attract and/or gain electrons
- elements on the left of the Periodic Table like to lose electrons to become isoelectronic with a Noble Gas, thus their ability to attract electrons is poor
- elements on the right of the Periodic Table like to gain electrons to become isoelectronic with a Noble Gas, thus their ability to attract electrons is quite strong

Electronegativities of Common Elements



- The most ***electronegative*** elements are found on the right (non-metals), but strongest on the top right (fluorine)

- ***Electropositive*** elements are very poor at gaining electrons (because they like to give up electrons instead (Groups I, II, III)
- These elements are found on the left (metals), but strongest on the bottom left (Francium).

Two Categories of Bonding

Intramolecular Bonding

- refers to attraction ions or atoms have for one another **WITHIN** a molecule

Intermolecular Bonding

- refers to the attraction **BETWEEN** molecules

Note: Intramolecular bonds are, in general, much stronger than intermolecular bonds

2. INTRAMOLECULAR BONDING

- Atoms combine in order to achieve a lower energy state (try to get full orbitals)

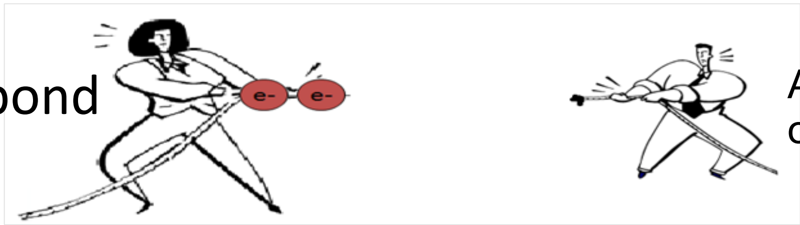
- We will focus on the representative elements (s & p blocks) *(bonding for transition metals is not covered in Chem 11)*

The Octet Rule

- Every representative element strives to achieve an s^2p^6 arrangement (full orbital) in order to be isoelectronic with a noble gas
- 2 + 6 = 8 which gives us the "*octet*" rule
- 8 valence electrons = 0 valence electrons
- Elements achieve an octet by chemical reactions and form one of 3 types of bonds:
- an ionic bond
- a polar covalent bond
- a non-polar covalent bond

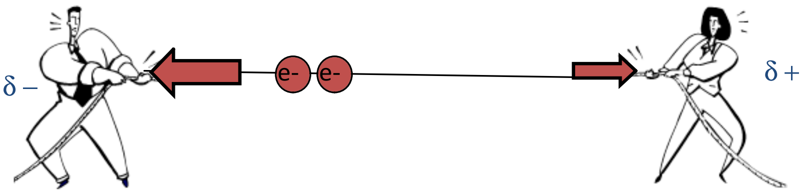
A tug of war between atoms

Ionic bond



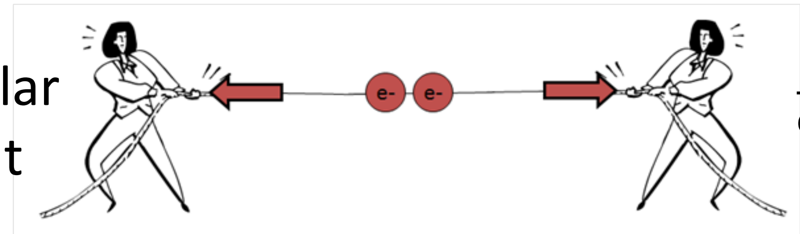
A complete transfer of electrons

Polar covalent bond



Unequal sharing of electrons

Non-polar covalent bond



Equal sharing of electrons

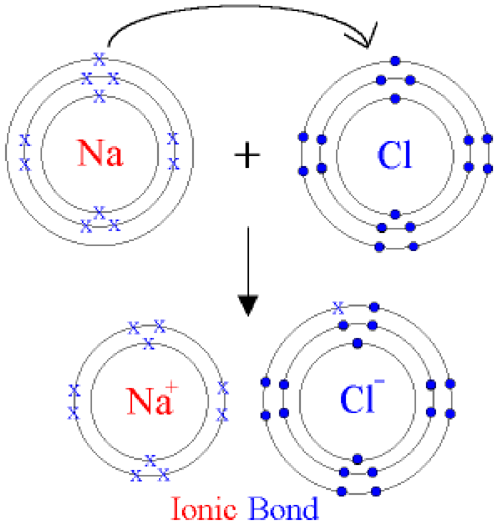
The type of bond formed depends on the electronegativity difference between the two atoms

A) Ionic Bonds

- Involves a metal and a non-metal
- Electronegativity difference ≥ 1.7

E.g. What happens when Na metal comes into contact with a Cl atom?

Na will give up an electron to Cl



Electronegativity calculation:

3.0 - 0.9 = 2.1 (greater than 1.7)
Cl - Na
Therefore, an IONIC bond
(a complete transfer of an electron)

$Na + Cl \longrightarrow Na^+ + Cl^-$

Now both Na^+ and Cl^- have an octet!

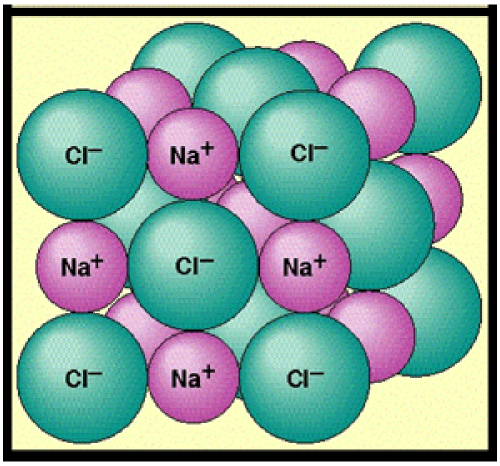
How does it happen?

- Since Na has a very low Ionization Energy (wants to give up an electron) and Cl has a very high electronegativity (wants to gain an electron) a complete transfer of electrons will take place

- Na^+ and Cl^- ions are attracted and stay beside each other due to electrostatic forces (+ - attraction)

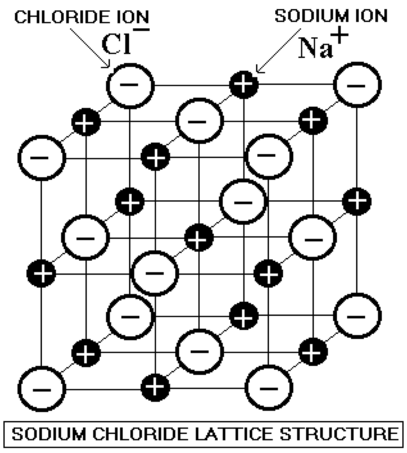
- Many metals and non-metals have an electronegativity difference greater than 1.7, resulting in ionic bonds (metal cations and non-metal anions)

- If this happens many times in the same vicinity, ions form a <u>crystal lattice</u> with positions of ions fixed due to electrostatic attraction



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

- Ionic solids are brittle and have very high melting points (difficult for ions to move due to strong electrostatic attraction)



- Each ionic solid (salt) has a unique crystal lattice arrangement

How does Ca bond with Br_2?

2.8 - 1.0 = 1.8 so ionic bond

Ca has 2 valence electrons to give up to be like a noble gas.

Each Br has 7 valence electrons and needs one more to be like a noble gas.

Therefore, one Ca will give up two electrons to become Ca^{2+} and this will cause two Br atoms to each gain one electron to become two Br^- ions

$$Ca + Br_2 \longrightarrow Ca^{2+} + 2Br^-$$

B) Non-Polar Covalent Bonds

- <u>equal sharing</u> of electrons to obtain octet
- Form between non-metal atoms, usually of the same element (diatomic molecules) eg.Cl_2, H_2, etc.
- Electronegativity difference of 0.0 - 0.2
- Each covalent bond is made up of two electrons, one from each atom

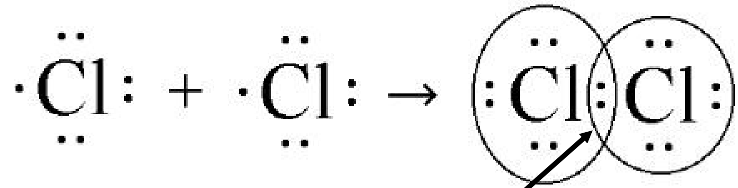
H•   •H

H : H

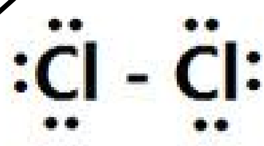
H–H

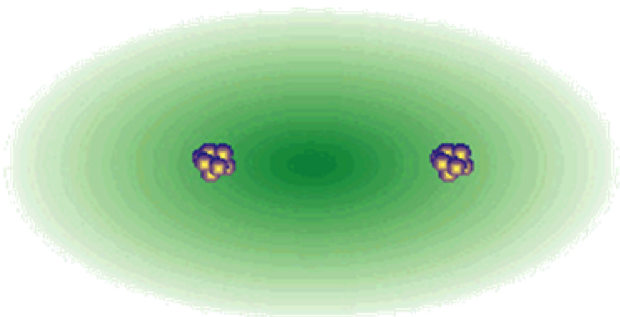
Hydrogen has 1 valence electron and can either give it up to have 0, or it can gain another to become isoelectronic with He ($1s^2$). So, if each H shares its electron with the other, they both have two electrons (like Helium). This is why H comes naturally as the diatomic molecule H_2

Cl_2: Each Cl atom has 7 valence electrons. If they can each gain 1 electron, they can have an octet like a noble gas (like Neon). Thus, they each share one of their electrons.



The two shared electrons are shared equally between the chlorines because they have equal electronegativities.





- Electrons are equally shared between atoms because electronegativity (pulling power) is identical (or almost!)
- Can have a single, double or triple bond depending on each atom’s needs to fill its octet (we'll learn more about this later in the unit)

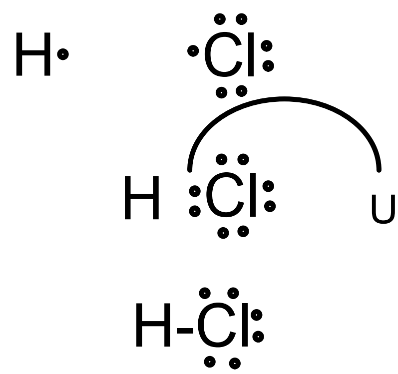
http://www.mhhe.com/physsci/chemistry/animations/chang_7e_esp/bom1s2_11.swf

C) Polar Covalent Bonds

- Usually form between non-metal atoms, but sometimes between a metal and a non-metal if the electronegativity difference is between 0.2 and 1.7
- Each atoms offers up an electron for sharing, but because one atom is significantly more electronegative that the other, the electrons are shared unequally (they are closer to the more electronegative atom)

- The atom that is pulling the shared electrons closer results in a 'partial' negative end and the other end is 'partially' positive
- The partial negative end of a polar bond is always at the atom with the highest electronegativity (and vice versa)
- Example: H and Cl

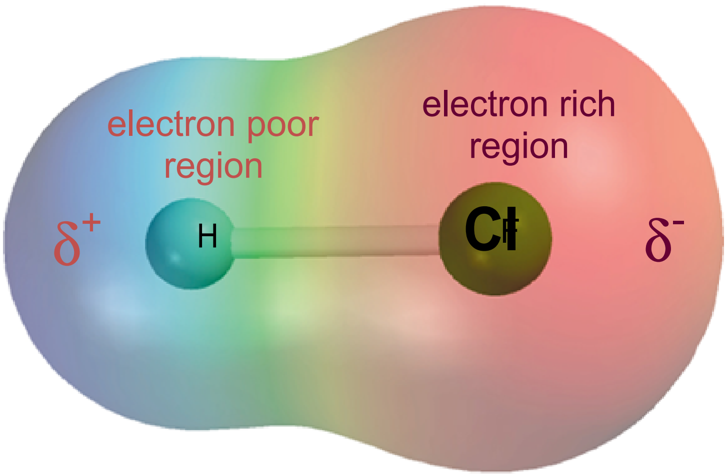
<u>Electronegativity Difference:</u>
3.0 - 2.1 = 0.9
Thus, polar covalent bond

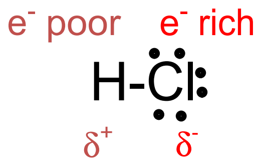


UNEQUAL SHARING

Polar Molecule

Slight charge at each end called a dipole (δ)





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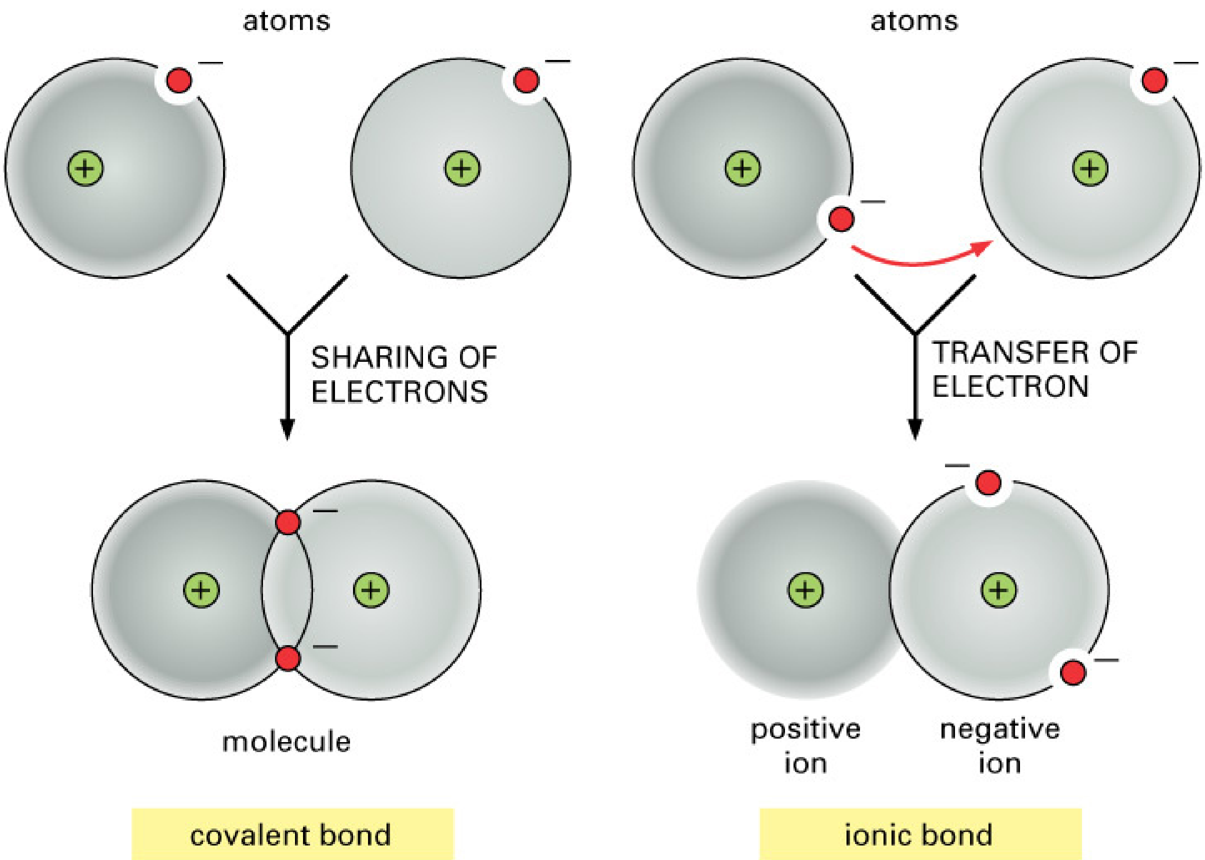
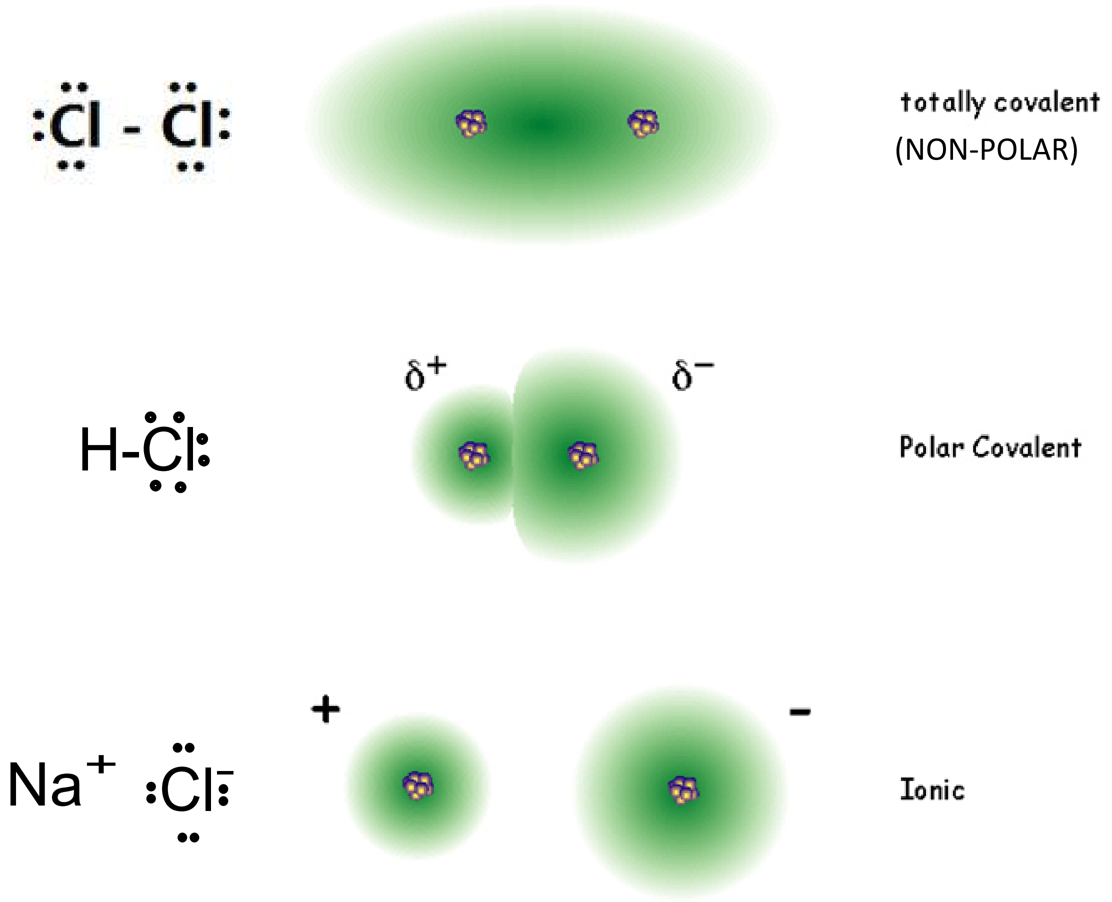
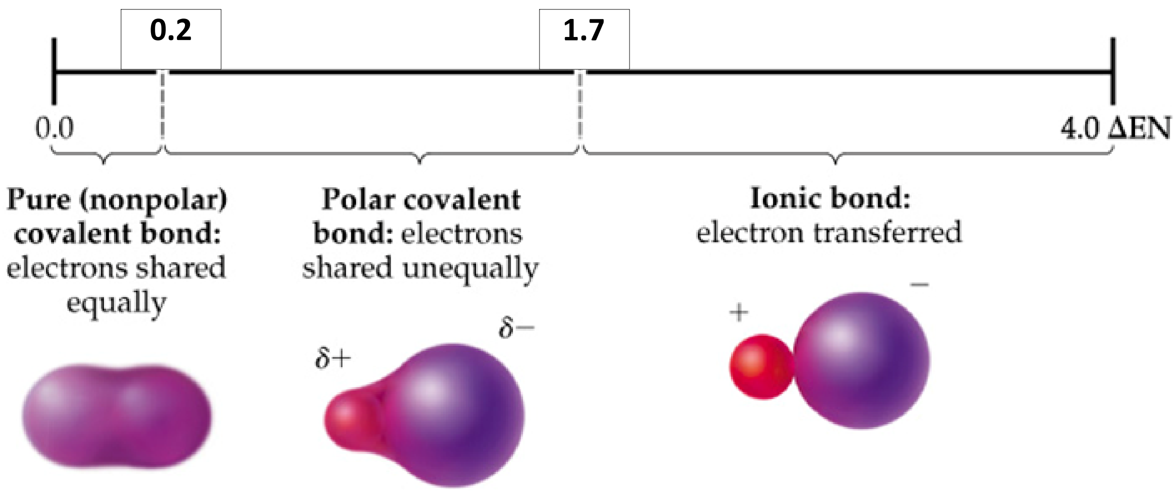


Figure 2.6 Essential Cell Biology, 2/e. (© 2004 Garland Science)


:Cl - Cl:
totally covalent
(NON-POLAR)
δ+
δ−
H-Cl:
Polar Covalent
+
-
Na+ :Cl:−
Ionic

http://chemsite.lsrhs.net/ChemicalBonds/images/custom_dipole2.swf



Electronegativities of Common Elements

Increasing electronegativity →

↑ Increasing electronegativity

1A	2A	3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	8A
H 2.1																	
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	
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 1.9	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	
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	
Cs 0.7	Ba 0.9	La-Lu 1.0-1.2	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	
Fr 0.7	Ra 0.9																

Examples

- What kind of intramolecular bond forms in each of the following?
- CS_2
- H_2O
- Al_2O_3
- AsH_3

HOMEWORK:

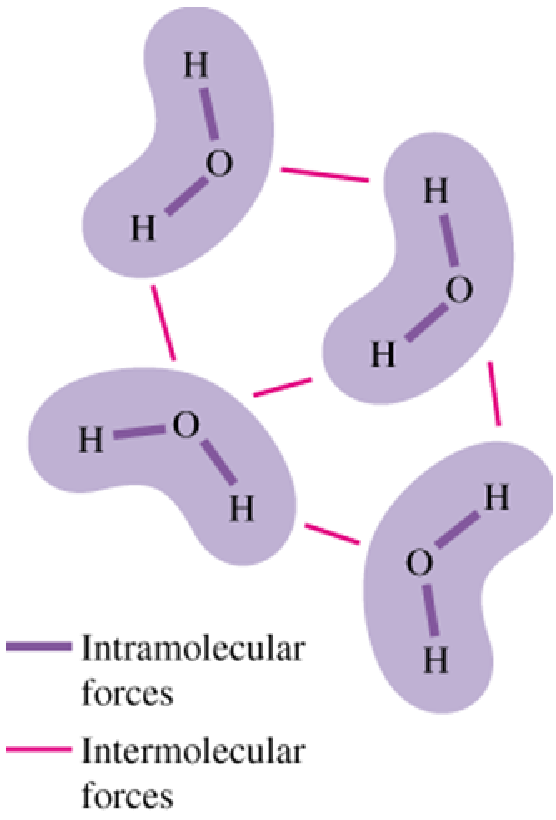
- Part 1 of the Intramolecular and Intermolecular Bonding Worksheet

3. Intermolecular Bonding

Recall:

Intramolecular bonding
- within a molecule
- holds atoms together as molecules
- holds ions together

Intermolecular bonding
- between molecules
- holds molecules to each other



Water striders use the strong intermolecular forces of water (creating surface tension) to stay on the surface.

In order to change a liquid to a gas, energy is required to overcome the intermolecular attractions. So to make tea or coffee, you need to decrease the intermolecular attractions of water.

DNA uses intermolecular forces to stay twisted in a small space. If it completely unwound, a human DNA would be over 3m long.

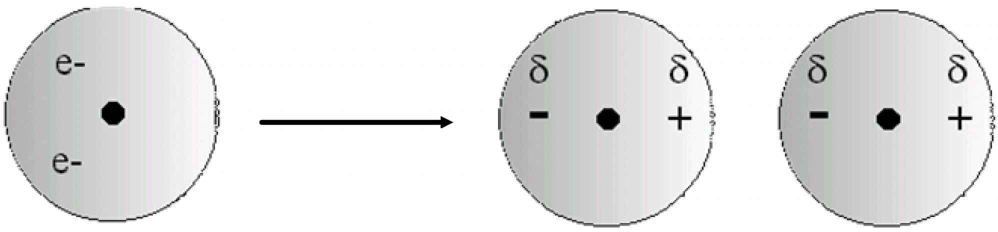
Intermolecular Attraction

- Much weaker than intramolecular bonds (ionic and covalent bonds)
- Intermolecular bonds become weaker as distance between molecules increases

- Three types:
- **London Dispersion Forces** - weakest type of intermolecular bonding
- **Dipole-to-Dipole Forces** in polar molecules
- **Hydrogen Bond** is a special type of very strong dipole-to-dipole forces (Water has very strong H-bonds)

London Dispersion Forces

- Temporary dipoles (partial positive or negative charges) that result from the random movement of the electrons around the particle
- A temporary dipole in one molecule will induce a temporary dipole in a neighbouring one. The two dipoles then attract each other
- Dispersion forces are temporary and work over very short distances



Instantaneous uneven distribution of electrons in He atom (dispersion force)

Creates a temporary dipole

Causes a neighbouring He atom to create a temporary dipole

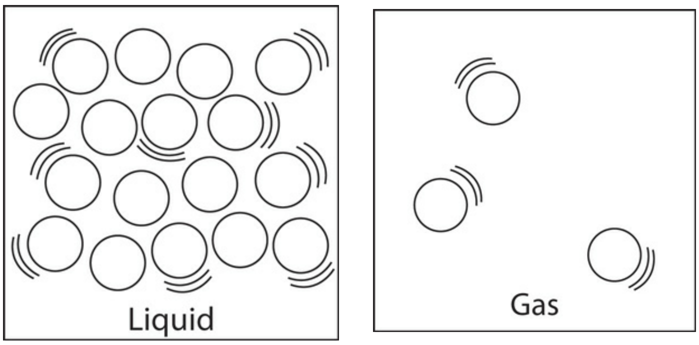
http://www.media.pearson.com.au/schools/cw/au_sch_derry_ibcsl_1/int/dispersionForces/f8hp/1111.html

- All molecules have them, but they are **overshadowed** by other, stronger intermolecular forces in polar molecules or ionic bonding in ionic compounds

- Important for noble gases and non-polar molecules, as it is their only form of intermolecular bonding
- Size of the London Force depends on the number of electrons in the atoms or molecules
- Large atoms and molecules have a larger electron cloud. Because of this they distort and polarize more easily, thus stronger dispersion forces

What is the result of dispersion forces?

- As atoms or molecules get larger, they have more electrons in their electron clouds, therefore can create stronger dispersion forces, creating a stronger attraction between the molecules



The weaker the dispersion forces, the easier the particles can become separated (the lower the temperature at which they will change state).

London Forces and Boiling Points

- Boiling points (C) of halogens (non-polar molecules)

smallest	F_2	Cl_2	Br_2	I_2	largest
	-188	-35	59	184	

- At room temperature, F_2 and Cl_2 are gases, Br_2 is a liquid, and I_2 is a solid

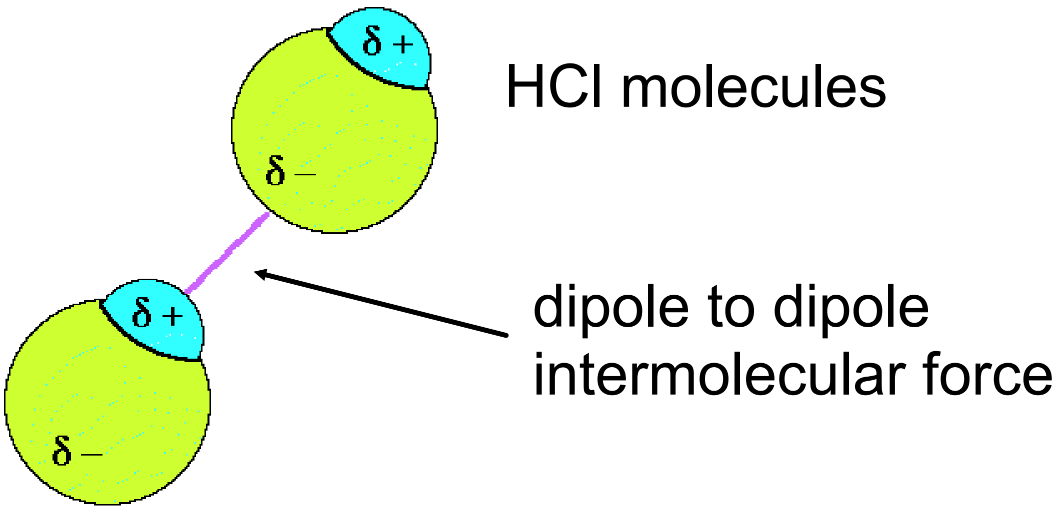
- Boiling points (C) of Noble gases

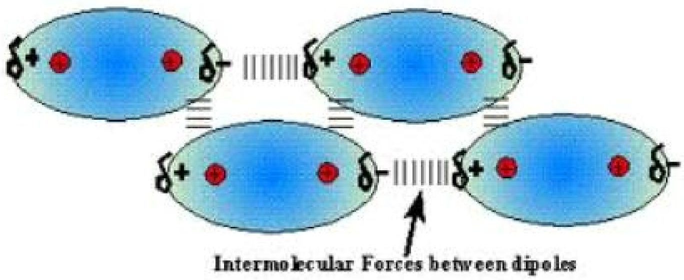
smallest	He	Ne	Ar	Kr	Xe	largest
	-269	-246	-186	-152	-108	

Dipole-to-Dipole Attractions

In polar covalent molecules, differences in electronegativity lead to unequal sharing of electrons which results in permanent dipoles (partial charges)

- the partial negative end of one molecule is attracted to partial positive end of another
- dipole-dipole forces are stronger than dispersion forces, and thus polar molecules have, in general, higher melting points & boiling points than non-polar molecules and atoms


δ +
δ –
δ +
δ –
HCl molecules
dipole to dipole
intermolecular force


δ+
δ-
Intermolecular Forces between dipoles

a) The interaction of two polar molecules

b) The interaction of many dipoles in a liquid

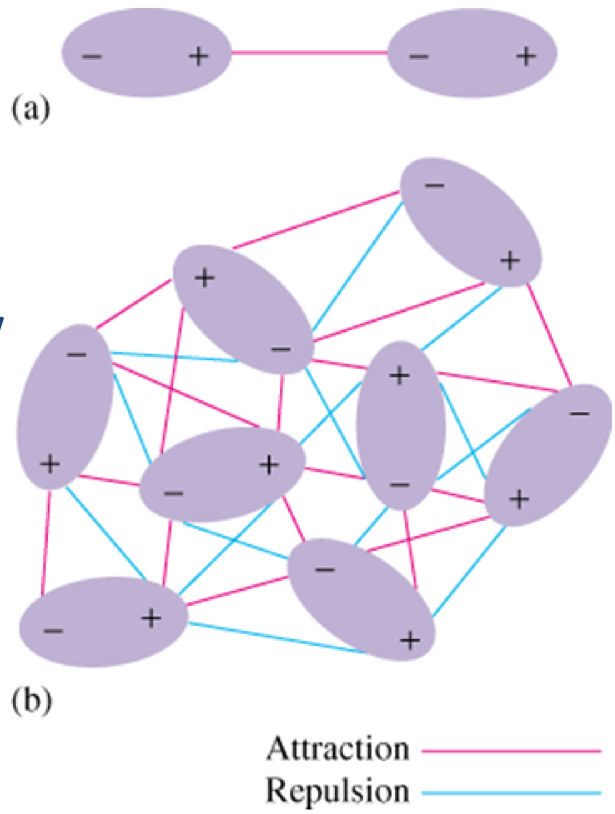


Table 11.2 Molecular Masses, Dipole Moments, and Boiling Points of Several Simple Organic Substances

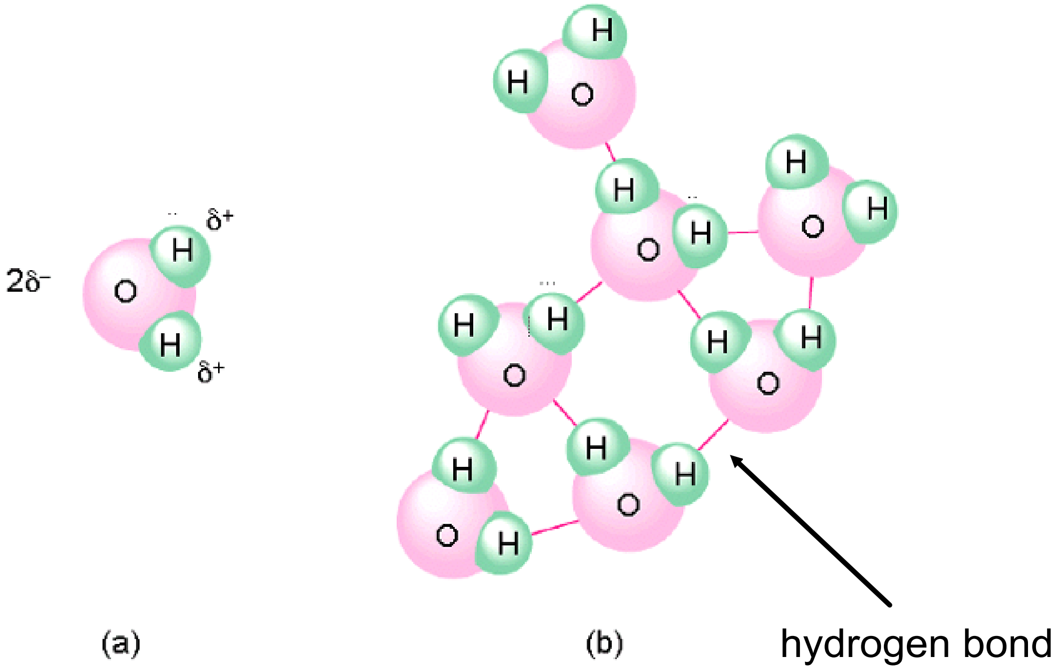
Substance	Molecular Weight (amu)	Dipole Moment, μ (D)	Boiling Point (K)
Propane, $CH_3CH_2CH_3$	44	0.1	231
Dimethyl ether, CH_3OCH_3	46	1.3	248
Methyl chloride, CH_3Cl	50	1.9	249
Acetaldehyde, CH_3CHO	44	2.7	294
Acetonitrile, CH_3CN	41	3.9	355

In general, the larger the dipole-dipole interaction, the higher the boiling point.

Hydrogen Bonding

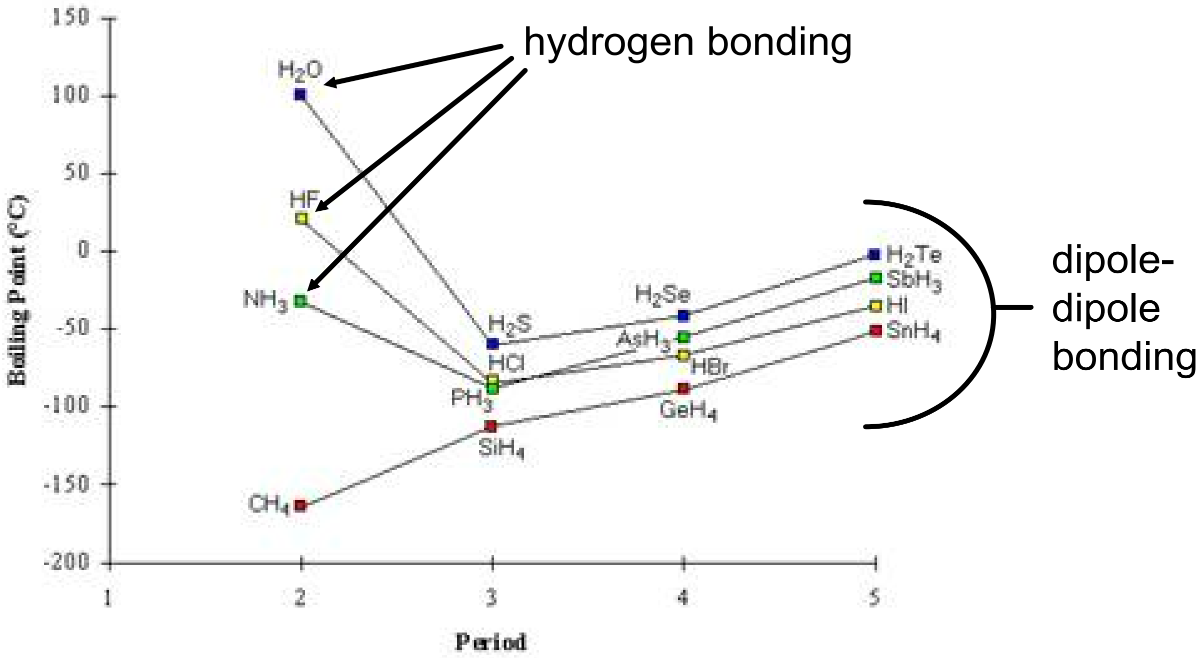
- a special type of dipole-dipole intermolecular bond
- Strongest intermolecular bond http://www.youtube.com/watch?v=aH2IbYs_XjY&feature=related
- Occurs in molecules that have **N-H**, **O-H** or **H-F** polar covalent intramolecular bonds (Hydrogen bonds are ***FON*** to learn!)
- permanent dipoles created are especially powerful as the hydrogen nucleus is essentially naked (due to the more electronegative F, O, or N pulling hydrogen's only electron away - unequal sharing)
- the naked hydrogen (partial positive charge) creates an intermolecular hydrogen bond with a neighbouring partial negative charge

(a) The polar water molecule
(b) Hydrogen bonding among water



Molecules with hydrogen bonding have:

- unusually high melting and boiling points



- unusually high solubility in water
- since water has hydrogen bonds, it likes to interact with other substances that have hydrogen bonds

- water has strong surface tension due to hydrogen bonds (water striders)

http://www.kentchemistry.com/links/bonding/Hbonding.htm
bottom left video

http://www.kentchemistry.com/links/bonding/bondingflashes/bond_types.swf

Deciding Which Intermolecular Bonds Exist in Certain Compounds

- Using the electronegativity table, decide which intramolecular bond exists within your molecule.
- If the intramolecular bond is:

a. **Ionic**, then the intermolecular bonds are also **ionic** (crystal lattice), however ionic bonds are all considered intramolecular bonds

b. **Non-polar** covalent: then the intermolecular bonds are London Dispersion Forces.

c. **Polar covalent:** then the intermolecular bonds are Dipole-Dipole Forces, unless the intramolecular bonds are between H-O, H-F, or H-N, in which case they are Hydrogen Bonds

Try these:

- HCl
  Intra: Polar Covalent
  Inter: Dipole-Dipole
- I_2
  Intra: Non-polar Covalent
  Inter: London Forces
- NH_3
  Intra: Polar Covalent
  Inter: Hydrogen
- NaBr
  Intra: Ionic
  Inter: Ionic

http://www.wwnorton.com/college/chemistry/gilbert2/tutorials/interface.asp?chapter=chapter_10&folder=intermolecular_forces

http://www.youtube.com/watch?v=S8QsLUO_tgQ

Table 14-2

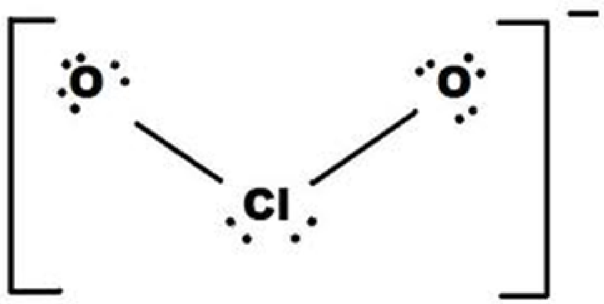
Weak Forces Summary

	Type of Force	Substances Exhibiting Force	Source of the Force	Properties Due to the Force	Example
INTERMOLECULAR BONDS Molecules contain covalently bonded atoms	**Dipole**	Polar covalent molecules	Electric attraction between dipoles resulting from polar bonds	Substances have higher boiling and melting points than those having nonpolar molecules of similar size $100^\circ < mp < 600^\circ$	ICl, SO_2, $BiBr_3$, AlI_3, SeO_3
	Dispersion Forces	Nonpolar molecules	Weak electric fluctuations which destroy spherical symmetry of electronic fields about atoms	Substances have low melting and boiling points	Cl_2, CH_4, N_2, O_2, F_2, Br_2, Cl_2, He, Ar

HOMEWORK:

- Complete Part 2 of the Intramolecular and Intermolecular Bonding Worksheet
-
- Hebden p. 180-182 #73, 76, 80, 81, 83

4.Drawing Lewis Structures

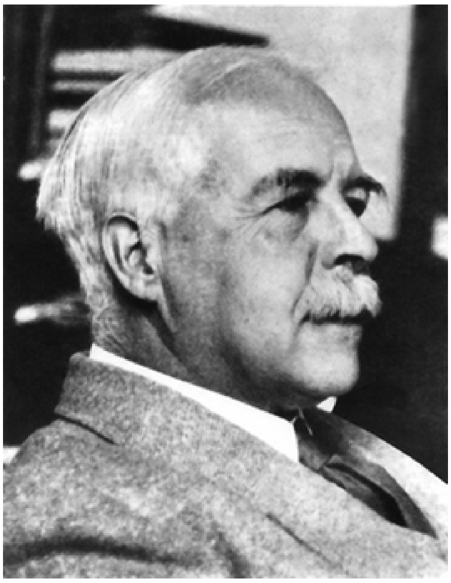


Lewis Structures

- Diagrams that show how valence electrons are distributed in an atom, ion or molecule
- Also called electron dot diagrams
- Each valence electron is represented as a dot
- Electron pairs shared between atoms form bonds
- Unshared pairs are called "lone pairs"

The Lone Ranger

Lewis Bonding



- Valence e^- are the players in bonding
- Valence e^- transfer leads to *ionic bonds.*
- Sharing of valence e^- leads to *covalent bonds.*
- e^- are transferred or shared to give each atom a noble gas configuration (a full shell)
- *the octet.*

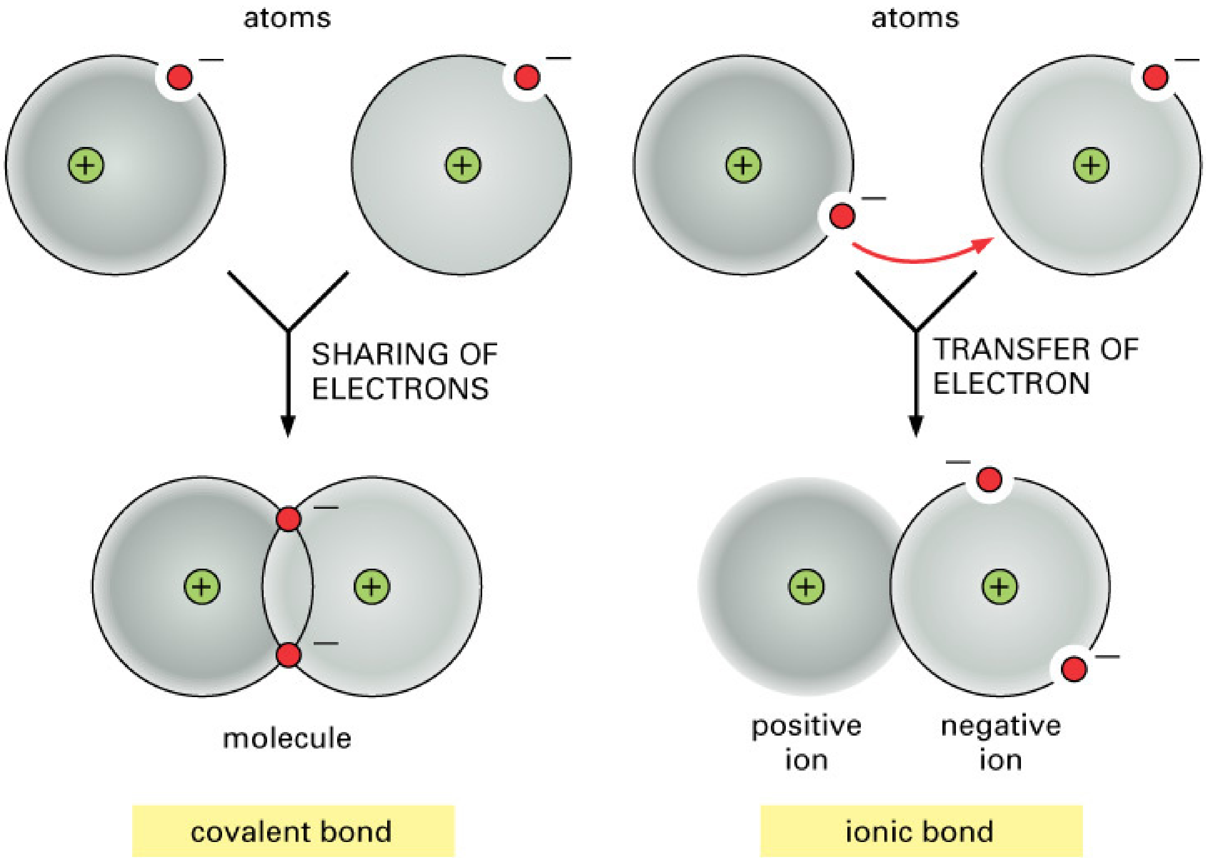


Figure 2.6 Essential Cell Biology, 2/e. (© 2004 Garland Science)

The Octet Rule

Atoms tend to gain, lose, or share electrons until they have eight valence electrons (and s^2p^6 arrangement).

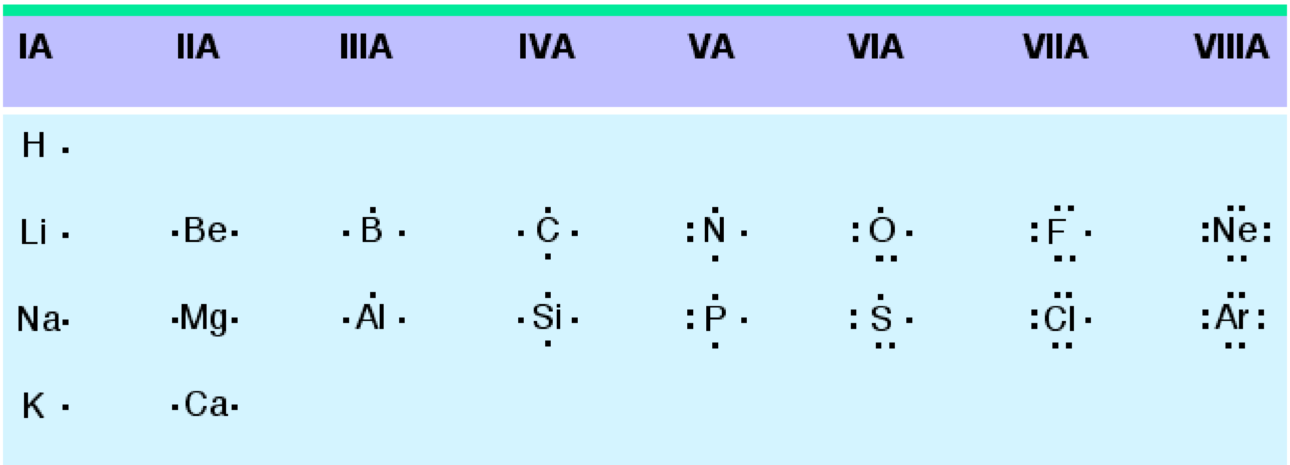
Hydrogen is an exception. It likes to gain one electron to have 2 electrons (like Helium)

- **Recall how to determine the valence electron for the elements based on position in the periodic table.**

hydrogen 1 H 1.0079																		helium 2 He 4.0026
lithium 3 Li 6.941	beryllium 4 Be 9.0122												boron 5 B 10.811	carbon 6 C 12.011	nitrogen 7 N 14.007	oxygen 8 O 15.999	fluorine 9 F 18.998	neon 10 Ne 20.180
sodium 11 Na 22.990	magnesium 12 Mg 24.305												aluminium 13 Al 26.982	silicon 14 Si 28.086	phosphorus 15 P 30.974	sulfur 16 S 32.065	chlorine 17 Cl 35.453	argon 18 Ar 39.948
potassium 19 K 39.098	calcium 20 Ca 40.078		scandium 21 Sc 44.956	titanium 22 Ti 47.867	vanadium 23 V 50.942	chromium 24 Cr 51.996	manganese 25 Mn 54.938	iron 26 Fe 55.845	cobalt 27 Co 58.933	nickel 28 Ni 58.693	copper 29 Cu 63.546	zinc 30 Zn 65.39	gallium 31 Ga 69.723	germanium 32 Ge 72.61	arsenic 33 As 74.922	selenium 34 Se 78.96	bromine 35 Br 79.904	krypton 36 Kr 83.80
rubidium 37 Rb 85.468	strontium 38 Sr 87.62		yttrium 39 Y 88.906	zirconium 40 Zr 91.224	niobium 41 Nb 92.906	molybdenum 42 Mo 95.94	technetium 43 Tc [98]	ruthenium 44 Ru 101.07	rhodium 45 Rh 102.91	palladium 46 Pd 106.42	silver 47 Ag 107.87	cadmium 48 Cd 112.41	indium 49 In 114.82	tin 50 Sn 118.71	antimony 51 Sb 121.76	tellurium 52 Te 127.60	iodine 53 I 126.90	xenon 54 Xe 131.29
caesium 55 Cs 132.91	barium 56 Ba 137.33	57-70 *	lutetium 71 Lu 174.97	hafnium 72 Hf 178.49	tantalum 73 Ta 180.95	tungsten 74 W 183.84	rhenium 75 Re 186.21	osmium 76 Os 190.23	iridium 77 Ir 192.22	platinum 78 Pt 195.08	gold 79 Au 196.97	mercury 80 Hg 200.59	thallium 81 Tl 204.38	lead 82 Pb 207.2	bismuth 83 Bi 208.98	polonium 84 Po [209]	astatine 85 At [210]	radon 86 Rn [222]
francium 87 Fr [223]	radium 88 Ra [226]	89-102 * *	lawrencium 103 Lr [262]	rutherfordium 104 Rf [261]	dubnium 105 Db [262]	seaborgium 106 Sg [266]	bohrium 107 Bh [264]	hassium 108 Hs [269]	meitnerium 109 Mt [268]	ununnilium 110 Uun [271]	unununium 111 Uuu [272]	ununbium 112 Uub [277]		ununquadium 114 Uuq [289]				

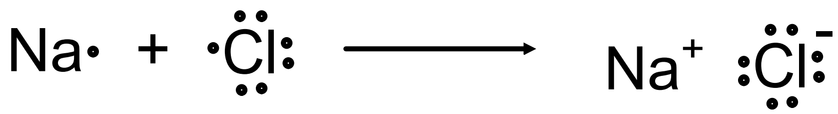
*Lanthanide series	lanthanum 57 La 138.91	cerium 58 Ce 140.12	praseodymium 59 Pr 140.91	neodymium 60 Nd 144.24	promethium 61 Pm [145]	samarium 62 Sm 150.36	europium 63 Eu 151.96	gadolinium 64 Gd 157.25	terbium 65 Tb 158.93	dysprosium 66 Dy 162.50	holmium 67 Ho 164.93	erbium 68 Er 167.26	thulium 69 Tm 168.93	ytterbium 70 Yb 173.04
* *Actinide series	actinium 89 Ac [227]	thorium 90 Th 232.04	protactinium 91 Pa 231.04	uranium 92 U 238.03	neptunium 93 Np [237]	plutonium 94 Pu [244]	americium 95 Am [243]	curium 96 Cm [247]	berkelium 97 Bk [247]	californium 98 Cf [251]	einsteinium 99 Es [252]	fermium 100 Fm [257]	mendelevium 101 Md [258]	nobelium 102 No [259]

- Lewis Dot Diagrams represent the valence electrons, as they are the electrons that are available in a reaction.



Ionic vs Covalent Compounds

- Formation of sodium chloride:



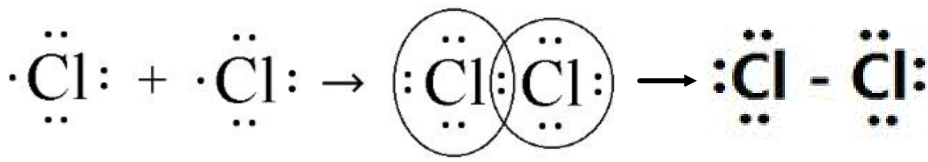
- Formation of hydrogen chloride:

A metal and a nonmetal transfer electrons to form an **ionic** compound. Two nonmetals share electrons to form a **covalent** compound.

Lewis Structures For **Covalent Compounds**

- A valid Lewis structure should have an octet for each atom except hydrogen (which should have a duet)

H_2 H• •H ⟶ H:H ⟶ H–H

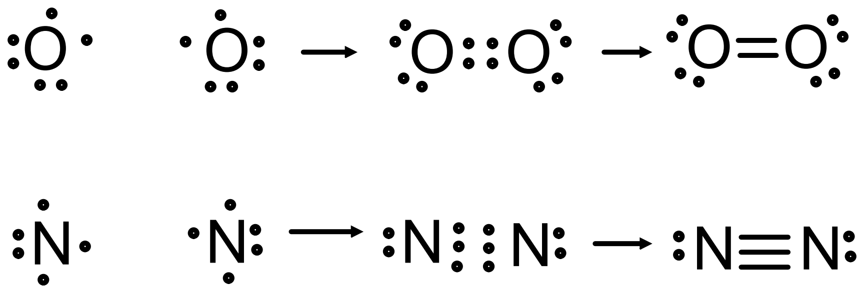
Cl_2 

Double and Triple Bonds

- Atoms can share four electrons to form a <u>double bond</u> or six electrons to form a <u>triple bond</u>.

O_2

N_2



- Triple bonds are the shortest and strongest whereas single bonds are the longest and weakest

How to draw a Lewis Structure

We'll use NCl_3 as an example

- 1. Find the number of valence electrons you already

Have. To do this, determine the number of valence electrons for each element in the structure.

N has 5 valence, and each Cl has 7, so 5 + 7(3) = 26 e^-

- 2. Find the total number of electrons **Needed** by all elements in the structure to satisfy the OCTET rule. Every element needs 8 (except hydrogen which needs 2).

N needs 8, so does each Cl, so 8(4) = 32 e^-

How to draw a Lewis Structure

- 3. Find the total number of electrons **Shared** between all elements. This is found by Need – Have.

For NCl_3, 32 - 26 = 6 e^-

- 4. Find the number of **Bonds** by dividing Shared by 2.

For NCl_3, 6 divided by 2 equals <u>3 bonds</u>

Other helpful hints:

A. Connectivity - From the Chemical Formula, determine the atom connectivity for the structure.

i. Given a chemical formula, AB_n, A is the central atom and B surrounds the A atom.

i.e., In NH_3, NCl_3 & NO_2 the N is the central atom in the structure

ii. H and F are never central atoms.

So, in NCl_3, N is central and the Cl atoms are terminal.

B. You can calculate the number of Non - Bonding Electrons by

Have – Shared

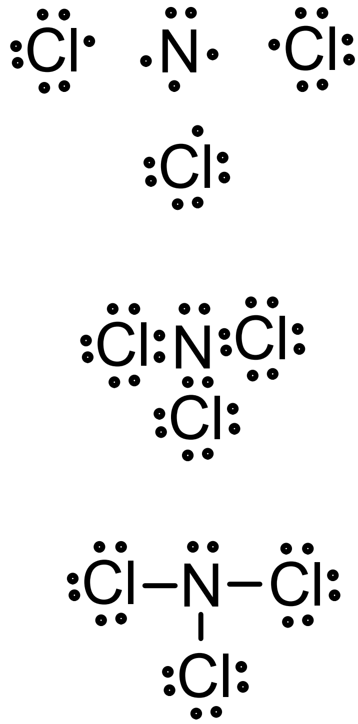
In NCl_3, nbe = 26 - 6 = 20 e^-

C. The number of bonds each particular atom tends to make depends on the number of its valence electrons. See next slide.

Family		→	# Covalent Bonds*
Halogens F, Br, Cl, I	X	→	1 bond often
Calcogens O, S	O	→	2 bond often
Nitrogen N, P	N	→	3 bond often
Carbon C, Si	C	→	4 bond always

So, N likes to make 3 bonds, and each Cl 1 bond. Thus the central N will have a single bond to each of the Cl atoms.

Example: NCl_3

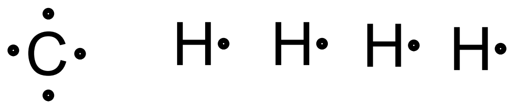


h: 26
n: 32
s: 6
b: 3
nbe: 20

central: N (3 bonds)
terminal: Cl (1 bond)

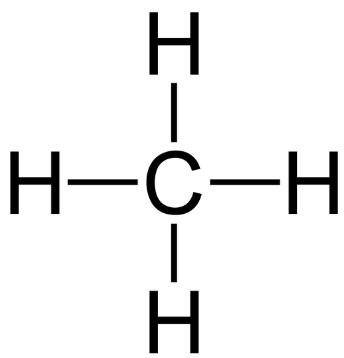
Using the hnsb System

CH_4



h: 8

n: 16



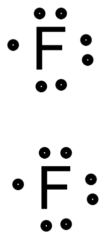
s: 8

b: 4

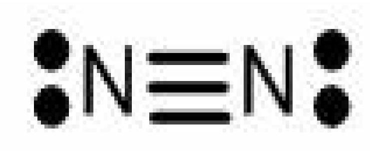
nbe: 0

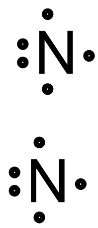
bonds: C makes 4 bonds and each H makes 1 bond

F_2
h: 14
n: 16
s: 2
b: 1
nbe: 12
bonds: F 1

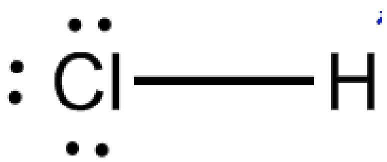


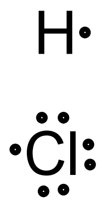
N_2
h: 10
n: 16
s: 6
b: 3
nbe: 4
bonds: N 3



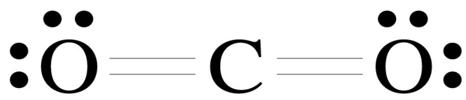


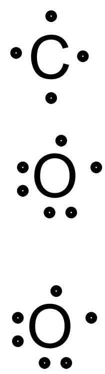
HCl
h: 8
n: 10
s: 2
b: 1
nbe: 6
bonds: H 1, Cl 1



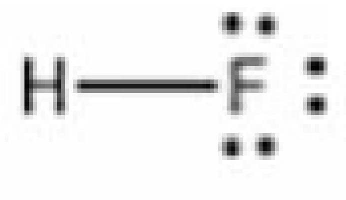


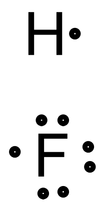
CO_2
h: 16
n: 24
s: 8
b: 4
nbe: 8
bonds: C 4, O 2



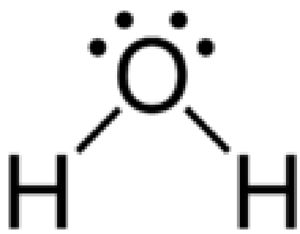


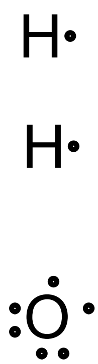
HF
h: 8
n: 10
s: 2
b: 1
nbe: 6
bonds: H 1, F 1



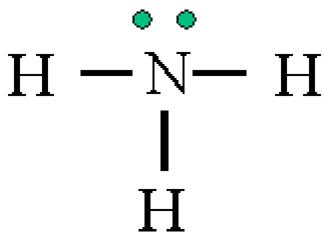


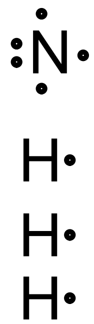
H_2O
h: 8
n: 12
s: 4
b: 2
nbe: 4
bonds: H 1, O 2



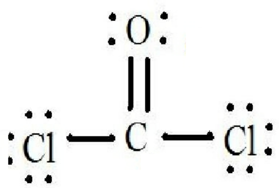


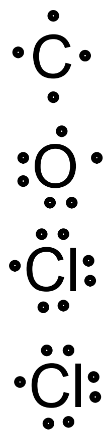
NH_{3-}
h: 8
n:14
s: 6
b: 3
nbe: 2
bonds: N 3, H 1





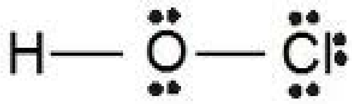
$COCl_2$
h: 24
n: 32
s: 8
b: 4
nbe: 16
bonds: C 4, O 2, Cl 1

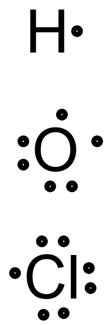




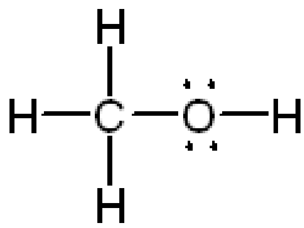
Sometimes, the way a formula is written, gives some information as to how the atoms are connected.

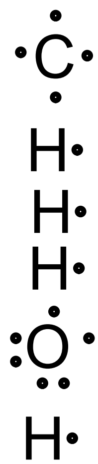
HOCl
h: 14
n: 18
s: 4
b: 2
nbe: 10
bonds: H 1, O 2, Cl 1





CH_3OH
h: 14
n: 24
s: 10
b: 5
nbe: 4
bonds: C 4, H 1, O 2



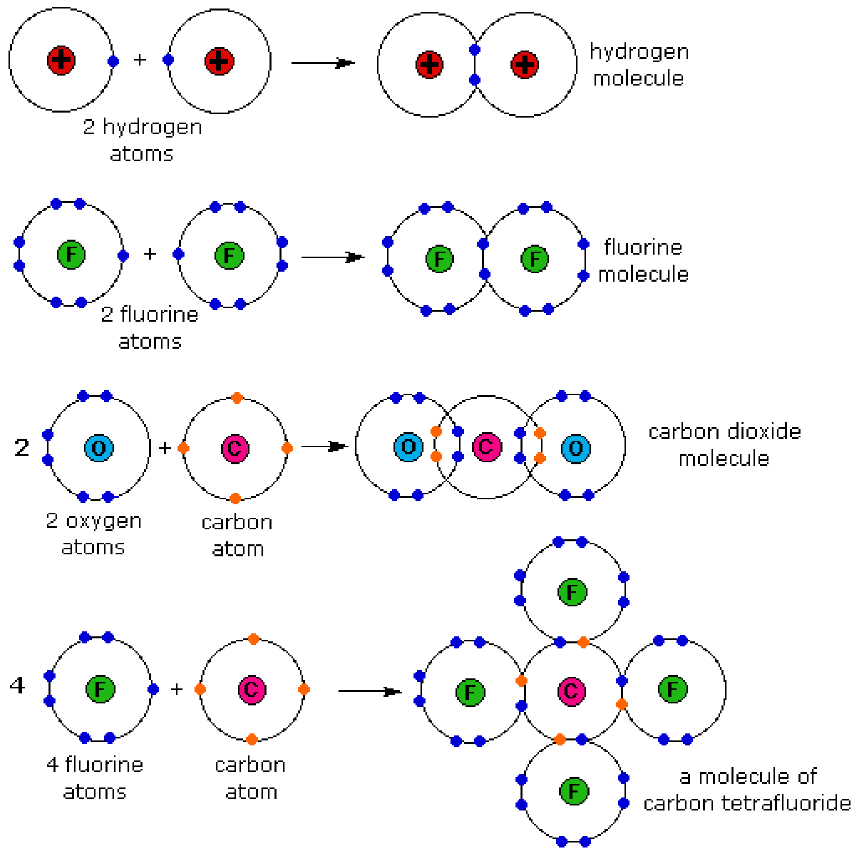


<u>HOMEWORK:</u>

Lewis Structures Worksheet - Part 1, Set A only

5. Lewis Structures for Polyatomic Ions

Covalent Bonding Pictoral Summary



Charged Polyatomic Groups

- For negatively charged ions, add the correct amount of electrons to your Have group
- For positively charged ions, subtract the correct amount of electrons from your Have group
- Follow the same rules as before (note: you will always Have an **even number of electrons**)

For neutral structures, C will make 4 bonds, N will make 3, O will make 2, and the halogens & hydrogen will make 1.
For polyatomic ions, the atom's first choice would be to keep to the # of bonds listed above, but:

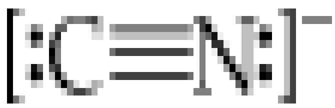
- C may sometimes only make 3 bonds
- N can make 2, 3, or 4 bonds
- O can make 1, 2, or 3 bonds
- Cl, Br, I can make 1 or 3 bonds

CN^-

h: 4 + 5 + 1 = 10

n: 16

s: 6



b: 3

nbe: 4

For polyatomic ions, when finished drawing the Lewis structure, but square brackets around it, and the ion charge in the top right corner outside of the brackets.

H_3O^+

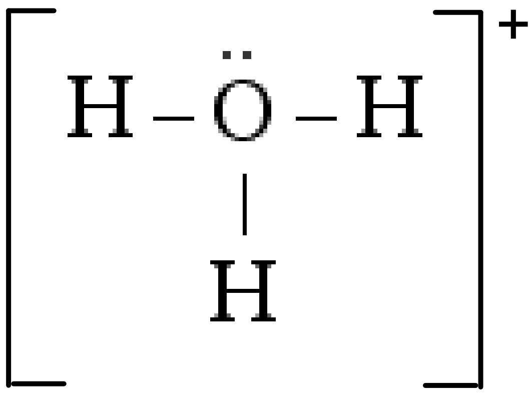
h: 6+3-1 = 8

n: 14

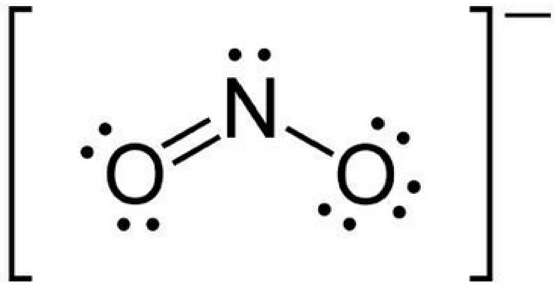
s: 6

b: 3

nbe: 2



NO_2^-
h: 5+6+6+1 = 18
n: 24
s: 6
b: 3
nbe: 12



HOMEWORK:

Lewis Structures Worksheet - Part 1 Set B only

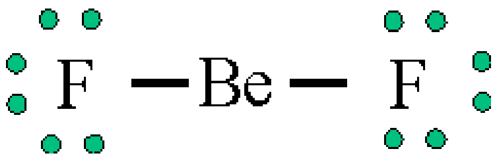
6. Exceptions to the Octet Rule

- There are numerous exceptions to the octet rule.
- Deficiencies:
- Be can be stable with only 4 valence electrons
- B can be stable with only 6 valence electrons

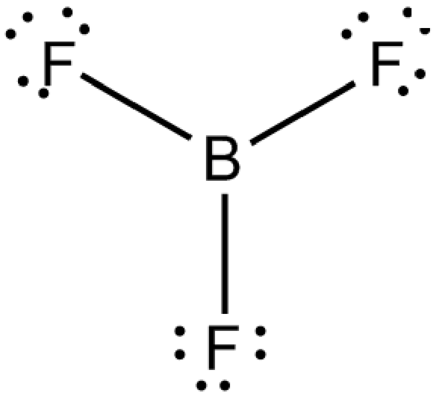
- Valence Shell Expansion:
- P and Cl can be stable with 10 valence electrons
- S can be stable with 12 valence electrons

The hnsb system does not work for these exceptions!

BeF_2



BF_3

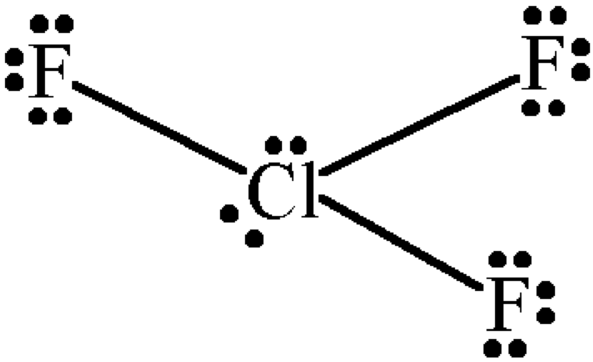


Valence Shell Expansion

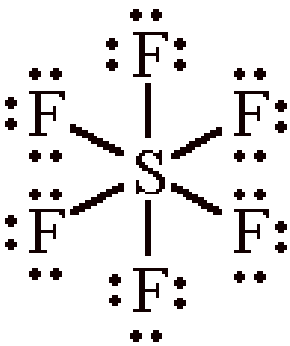
For third-row elements P, Cl and S, the energetic proximity of the 3d orbitals allows for the participation of these orbitals in bonding. When this occurs, more than 8 electrons can surround a third-row element.

- Cl and P can have $10e^-$ around it
- S can have $12e^-$

ClF_3



SF_6

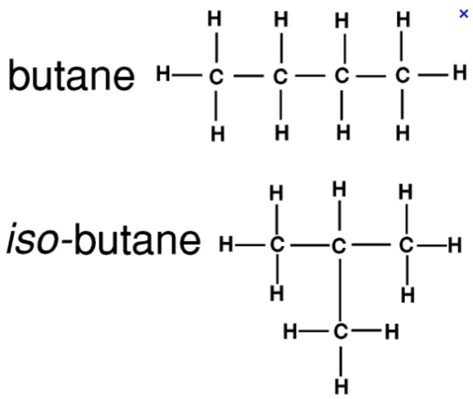


7. Resonance Structures

We have assumed up to this point that there is only valid Lewis structure for each formula.
There are systems for which more than one Lewis structure is possible:

- Different atomic linkages: Structural Isomers

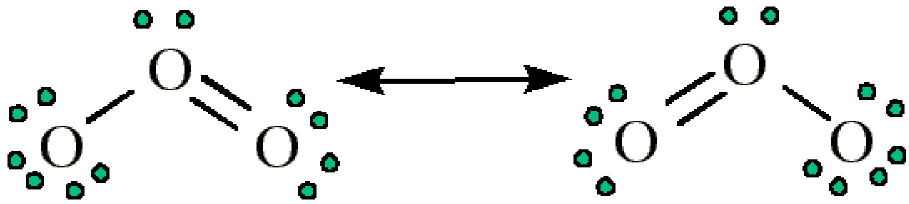
ex: C_4H_{10}



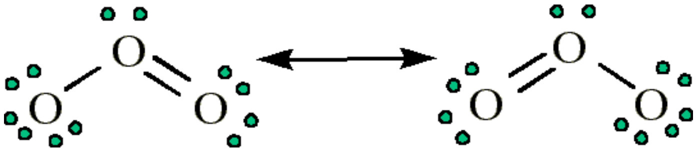
- Same atomic linkages, different bonding: Resonance

Resonance Structures

O_3
h: 18
n: 24
s: 6
b: 3
nbe: 12

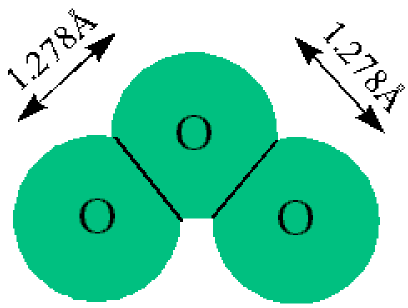


- In this example, O_3 has two resonance structures:

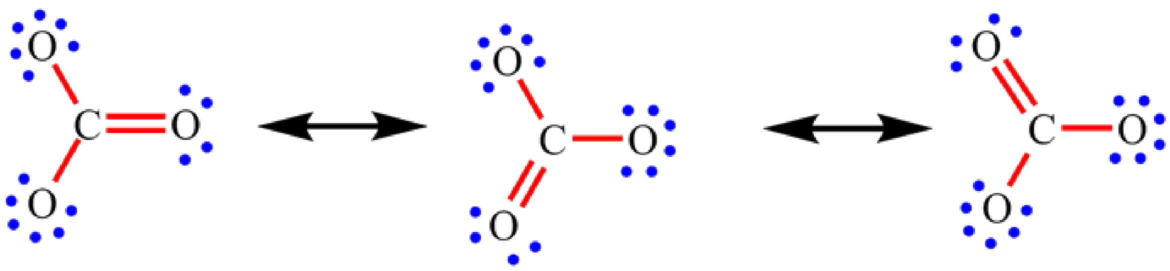


Each structure is perfectly valid, so conceptually, we think of the bonding being an average of these two structures, as the bond lengths are actually the same (see picture below). Electrons are delocalized between the oxygen atoms such that on average, the bond strength is equivalent to 1.5 O-O bonds.

http://www.wwnorton.com/college/chemistry/gilbert2/tutorials/interface.asp?chapter=chapter_08&folder=resonance

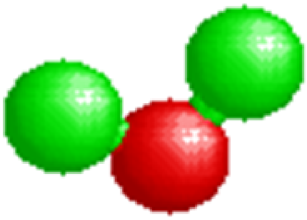


CO_3^{2-}
h: 24
n: 32
s: 8
b: 4
nbe: 16

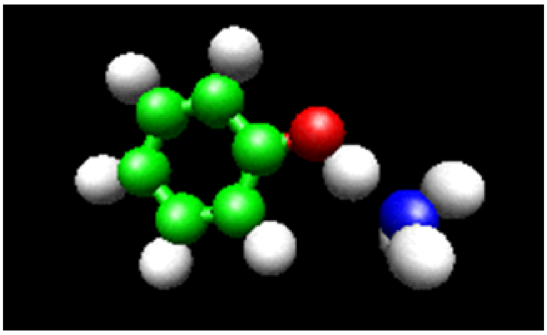


HOMEWORK:

Lewis Structures Worksheet - Part 1 Set C only



8. Molecular Shapes

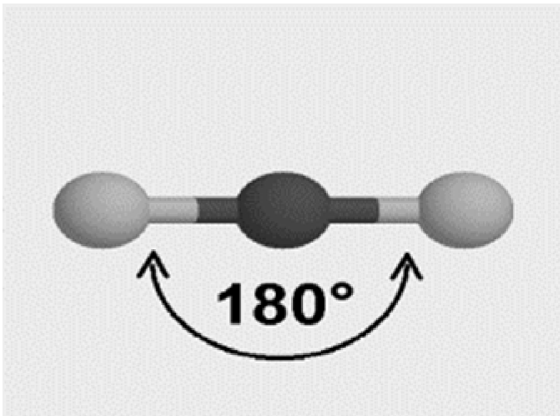


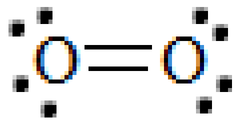
How do you determine shape?

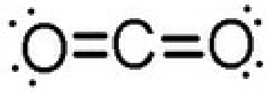
- Both bonding pairs and lone pairs need their space around the atom - VSEPR
- Valence Shell Electron Pair Repulsion
- Electron pairs repel each other -- maximize distance from other pairs
- Shape depends on arrangement of bonds and lone pairs around central atom
- Draw Lewis Structure and then use it to determine the 3D shape

Linear

- Bond angle of 180°
- All 2-atom molecules are linear
- 3-atom molecules with no lone pairs on central atom
- E.g. CO, H_2, CO_2,
- $BeCl_2$ (doesn't obey octet rule)





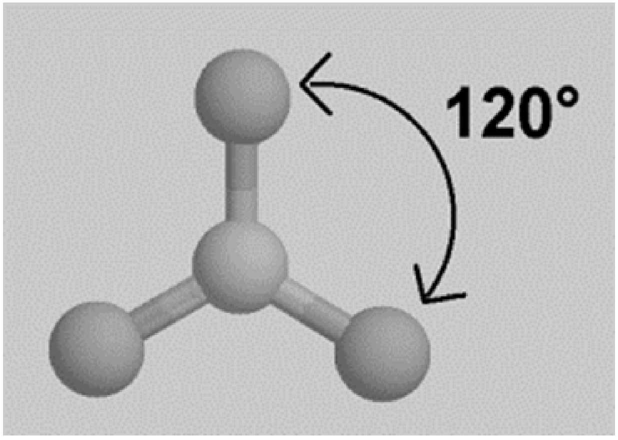


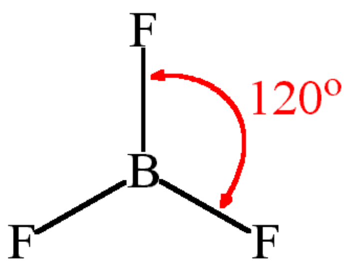
http://www.dlt.ncssm.edu/core/Chapter9-Bonding_and_Geometry/Chapter9-Animations/VSEPR/Linear.html

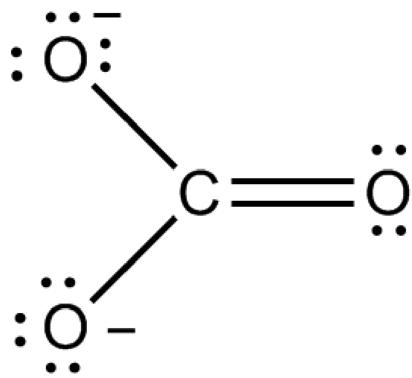
im in ur atmosphere, warmin ur gl0bes

Trigonal Planar

- Bond angle 120^{o}
- 3 atoms surrounding central atom with no lone pairs
- BF_3 (does not obey octet rule)
- CO_3^{2-}



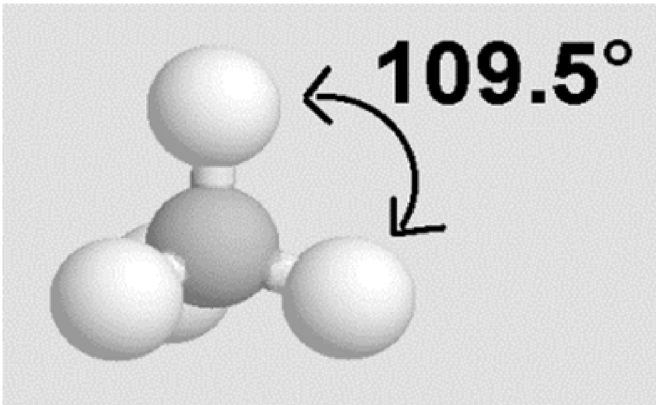


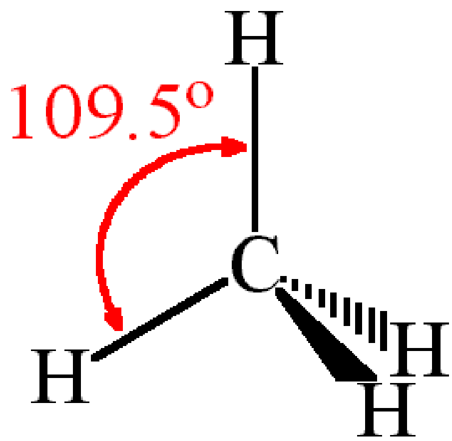


Tetrahedral

- Bond angle 109.5°
- When 4 atoms surround a central atoms which has no lone pairs
- E.g. CH_4, CCl_4

http://www.dlt.ncssm.edu/core/Chapter9-Bonding_and_Geometry/Chapter9-Animations/VSEPR/CH4.html

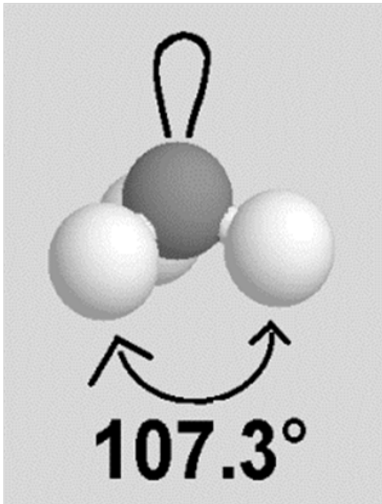


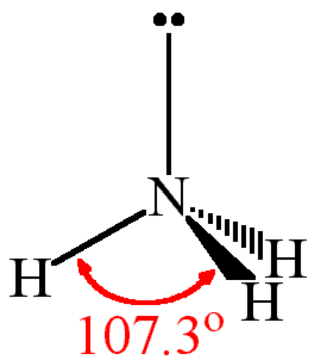


Trigonal Pyramidal

- Bond angle 107.3°
- lone pairs demand a little more space than a bond
- 3 atoms surround central atom with one lone pair
- E.g. NH_3, AsH_3, H_3O^+

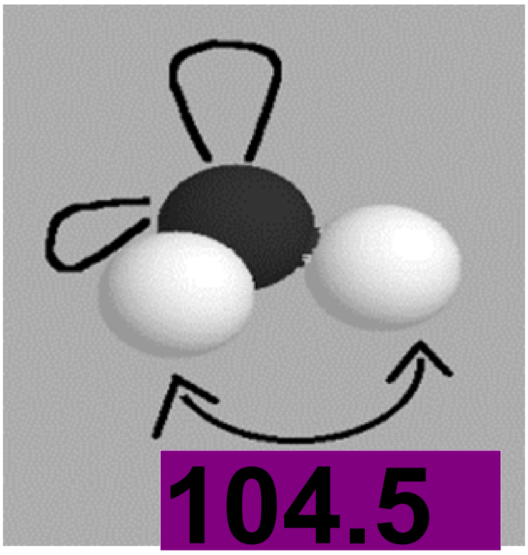
http://www.dlt.ncssm.edu/core/Chapter9-Bonding_and_Geometry/Chapter9-Animations/VSEPR/NH3.html





Angular (Bent)

- Bond angle 104.5^{o}
- 3 atoms present, with either 1 or 2 lone pairs on central atom
- E.g. H_2O, ClO_2^-, O_3



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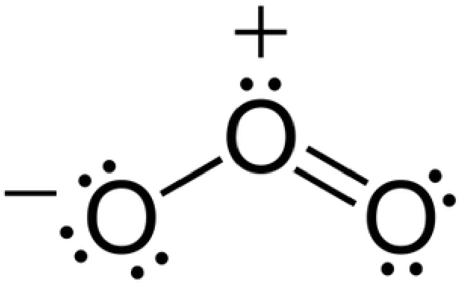
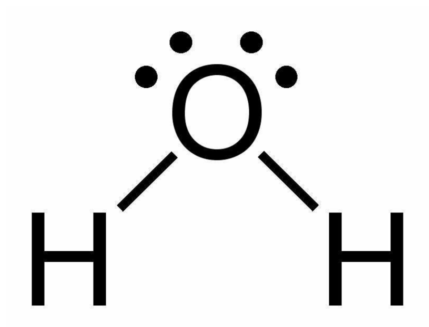


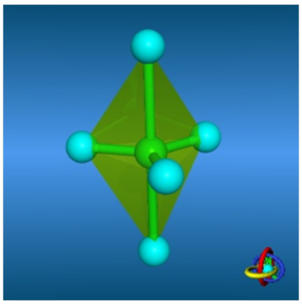
http://www.dlt.ncssm.edu/core/Chapter9-Bonding_and_Geometry/Chapter9-Animations/VSEPR/Bent_-109.html

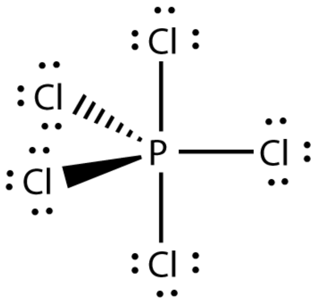


Other Shapes

Trigonal bipyramidal:

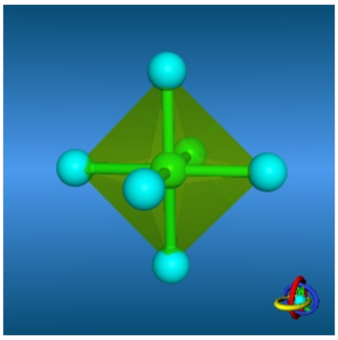
- E.g. PCl_5

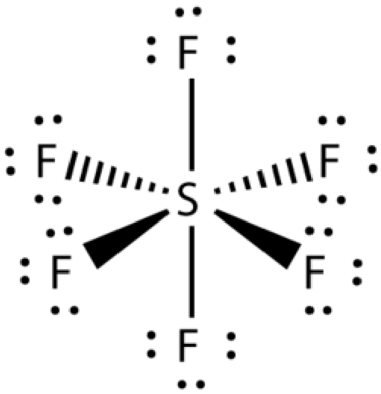




Octahedral:

- E.g. SF_6





http://www.dlt.ncssm.edu/core/Chapter9-Bonding_and_Geometry/Chapter9-Animations/VSEPR/Octahedral.html

Predict the shape of the following:

- H_2S
- HCl
- SiH_4
- SO_2

- HCN
- PCl_3
- CO_3^{2-}
- NH_4^+

Predict the shape of the following:

- H_2S Angular
- HCl Linear
- SiH_4 Tetrahedral
- SO_2 Angular
- HCN Linear
- PCl_3 Trigonal Pyramidal
- CO_3^{2-} Trigonal Planar
- NH_4^+ Tetrahedral

TABLE 9.2 Electron-Pair Geometries and Molecular Shapes for Molecules with Two, Three, and Four Electron Domains About the Central Atom

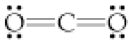
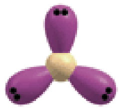
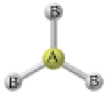
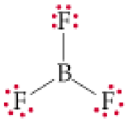
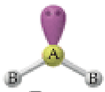
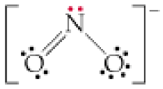
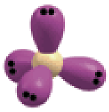
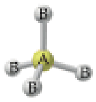
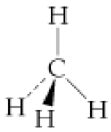
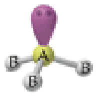
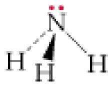
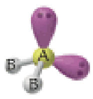
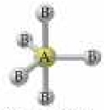
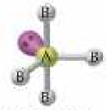
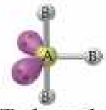
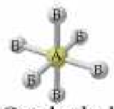
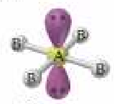
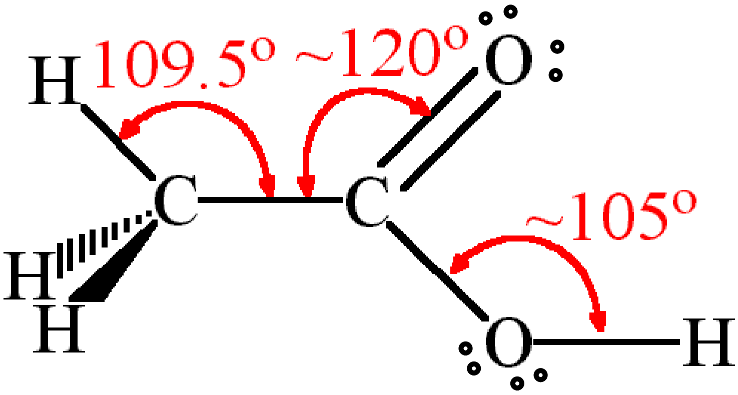
Total Electron Domains	Electron-Domain Geometry	Bonding Domains	Non-bonding Domains	Molecular Geometry	Example
2 pairs	 Linear	2	0	 Linear	
3 pairs	 Trigonal planar	3	0	 Trigonal planar	
		2	1	 Bent	
4 pairs	 Tetrahedral	4	0	 Tetrahedral	
		3	1	 Trigonal pyramidal	
		2	2	 Bent	

TABLE 9.3 Electron Pair Geometries and Molecular Shapes for Molecules with Five and Six Electron Pairs Domains About the Central Atom

Number of Electron Domains	Electron-Domain Geometry	Bonding Domains	Nonbonding Domains	Molecular Geometry	Example
5 domains	 Trigonal bipyramidal	5	0	 Trigonal bipyramidal	PCl_5
		4	1	 Seesaw	SF_4
		3	2	 T-shaped	ClF_3
		2	3	 Linear	XeF_2
6 domains	 Octahedral	6	0	 Octahedral	SF_6
		5	1	 Square pyramidal	BrF_5
		4	2	 Square planar	XeF_4

Hybrid Shapes:
Some structures are a hybrid of 2 or more shapes

CH_3COOH:



- tetrahedral
- trigonal planar
- bent

HOMEWORK:

Lewis Structures Worksheet - Part 2 Sets A, B, & C

9. Molecular Polarity

- Earlier this chapter we predicted the polarity of a bond (ionic, polar, or non-polar) using the electronegativity difference between the two elements
- Ionic bonds result in full charge (transfer)
- Polar Covalent bonds result in partial charges – *dipoles* (unequal sharing)
- Non-polar Covalent bonds result in no charges (equal sharing)

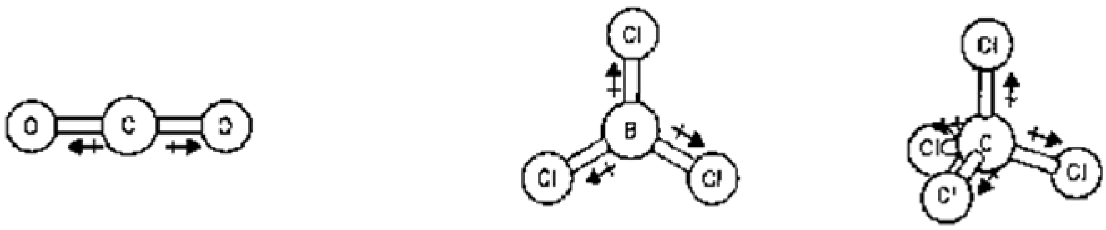
- To determine the polarity of the entire molecule, we must know the **shape of the molecule** and the **net result of all of the bond polarities.**
- **Two results possible**:
- If there is no charge concentration in any one direction in the molecule, then **NO net dipole** exists, resulting in a **non-polar** molecule
- If there is a concentration of +/- charge on any side of the molecule, then a **net dipole** exists, resulting in a **polar** molecule

- The method uses arrows to show bond dipoles

$$\delta+ \quad \longmapsto \quad \delta-$$

- Are the following molecules polar or non-polar? CO_2, BCl_3, CCl_4

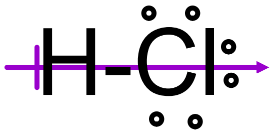
CO_2, BCl_3, and CCl_4 are all non-polar molecules



In all cases, the dipoles cancel out, due to the symmetry of the molecule. Thus, there is no net dipole.

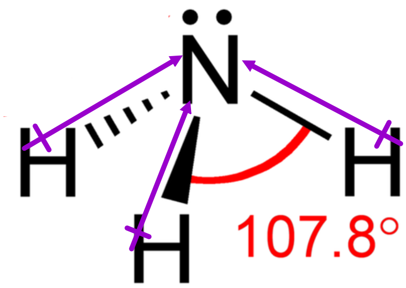
Are the following molecules polar or non-polar? HCl, NH_3, H_2O

HCl:



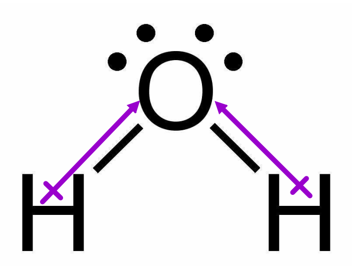
polar

NH_3:



polar

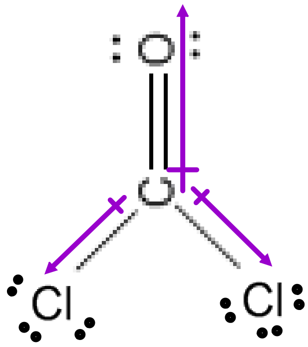
H_2O:



polar

Polar or non-polar?

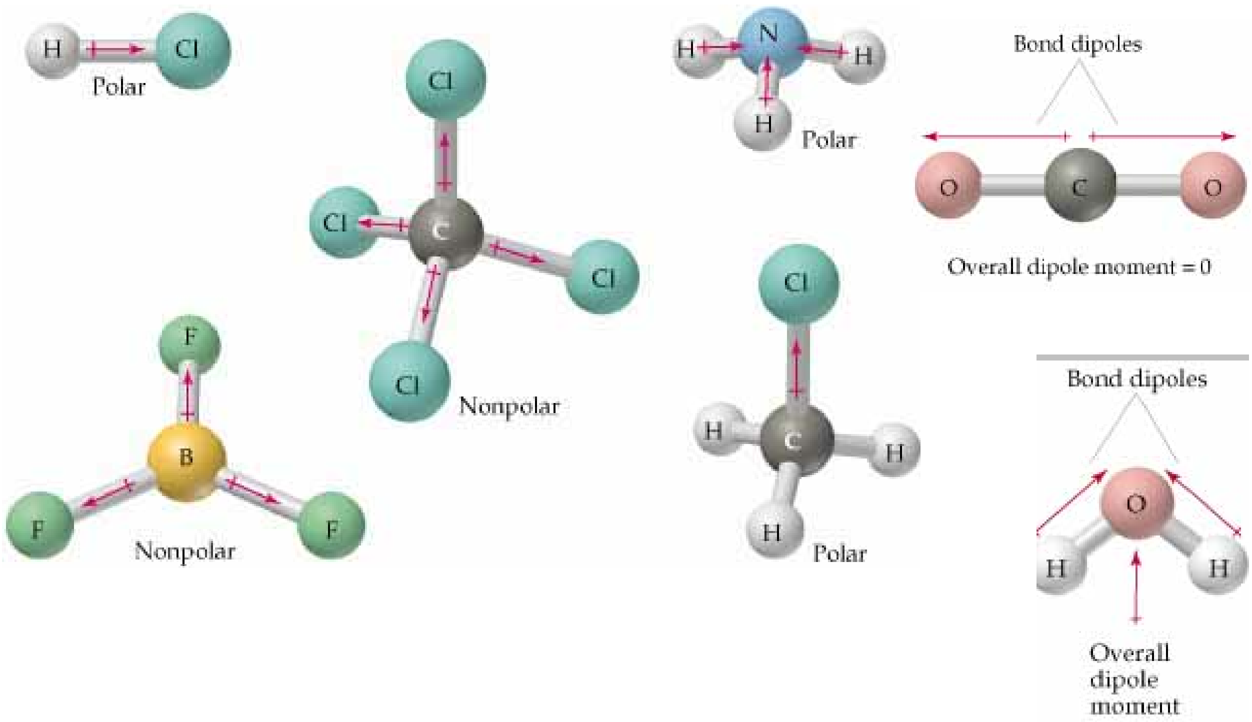
$COCl_2$



POLAR!

All of the particles we've investigated so far are molecules (neutral structures).

Ions have a charge, and are therefore polar molecules with no further investigating needed.



SNAP--> Symmetrical Nonpolar Asymmetrical Polar

Polar particles dissolve in polar liquids.
For example, the polar molecule CH_3COOH (acetic acid) dissolves in water (which is polar) to make vinegar. But non-polar canola oil does not dissolve in water.

Non-polar molecules dissolve in non-polar liquids.

Polar liquids are miscible (mix with) other polar liquids, but are immiscible (don't mix with) non-polar liquids.

http://www.dlt.ncssm.edu/core/Chapter10-Intermolecular_Forces/Chapter10-Animations/Polar_vs_Nonpolar.html

	H_2O polar	C_6H_{12} non-polar
$CuCl_2$ polar	$CuCl_2$ dissolves in H_2O	$CuCl_2$ does not dissolve in C_6H_{12}
I_2 non-polar	I_2 does not dissolve in H_2O	I_2 does dissolve in C_6H_{12}

Solids: $CuCl_2$
I_2

Liquids: H_2O
C_6H_{12}

HOMEWORK:

Lewis Structures Worksheet - Part 3 Sets A, B, & C