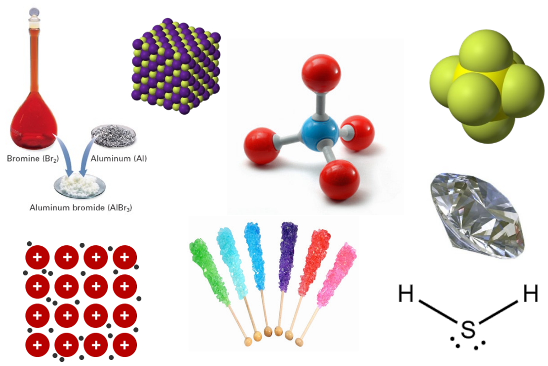
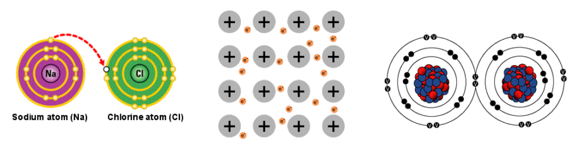
BONDING



Chemical Bond

Attraction that holds atoms together

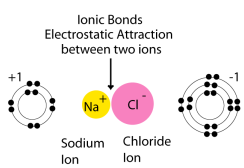
- Types include IONIC, METALLIC, or COVALENT
- Differences in electronegativity determine the bond type

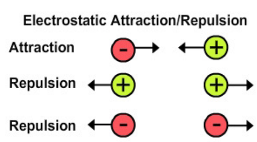


Electrostatic Attraction

Two objects near each other with different / opposite electrical charges

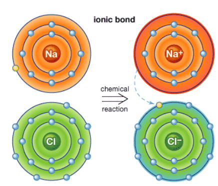
- Forces of attraction that allow atoms to bond… +/- ions attracted to each other in an ionic bond OR the positive nuclei attracted to the shared e- in covalent





Ionic Bond

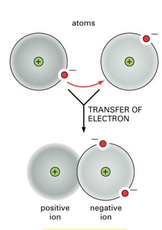
TRANSFER of electrons between atoms



- Each atom achieves a noble gas configuration (full valence shell)
- Usually between a METAL and a NONMETAL
- Formula Unit: lowest whole-number ratio of ions in an ionic compound (ex: NaCl or $MgCl_2$)

Ionic Bonding

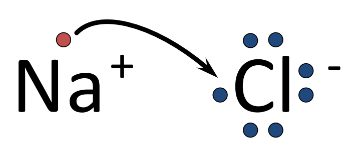
- How to show using Electron Dot Structures:
  1) Draw the dot structure for each element in the ionic compound
  2) Determine which element will lose e- and which will gain e- and how many
  3) Use arrows to show the e- being transferred to an empty space on the anion's dot structure
  4) Continue until each element has a complete octet… add more of each element as needed



Ionic Bonding

- **EXAMPLE:**

Show the electron dot structures for sodium and chlorine using arrows to indicate the transfer of e-.



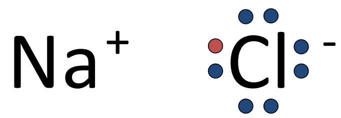
Wants to lose ONE electron!

Wants to gain ONE electron!

Ionic Bonding

• **EXAMPLE:**

Show the electron dot structures for sodium and chlorine using arrows to indicate the transfer of e-.

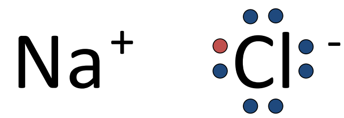


To write the formula unit for the compound: write the element symbol for the cation followed by the anion symbol… to indicate the # of each use subscripts (none needed if just one)!!

Ionic Bonding

• **EXAMPLE:**

Show the electron dot structures for sodium and chlorine using arrows to indicate the transfer of e-.

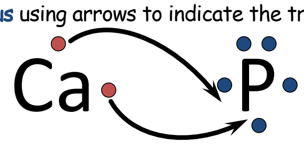


NaCl

Ionic Bonding

• **EXAMPLE:**

Show the electron dot structures for calcium and phosphorus using arrows to indicate the transfer of e-.



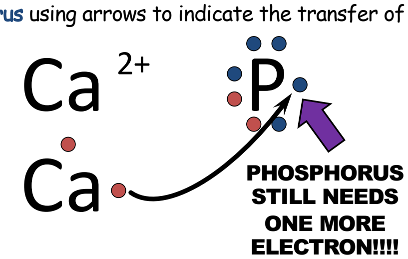
Wants to lose TWO electrons!

Wants to gain THREE electrons!

Ionic Bonding

• **EXAMPLE:**

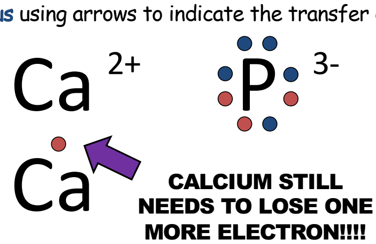
Show the electron dot structures for calcium and phosphorus using arrows to indicate the transfer of e-.



Ionic Bonding

• **EXAMPLE:**

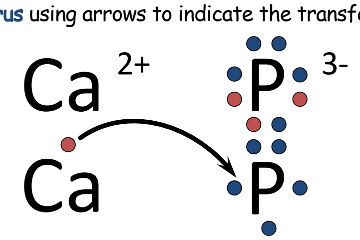
Show the electron dot structures for calcium and phosphorus using arrows to indicate the transfer of e-.



Ionic Bonding

• **EXAMPLE:**

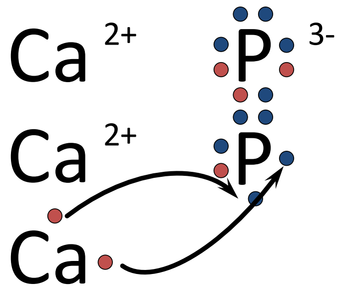
Show the electron dot structures for calcium and phosphorus using arrows to indicate the transfer of e-.



Ionic Bonding

- **EXAMPLE:**

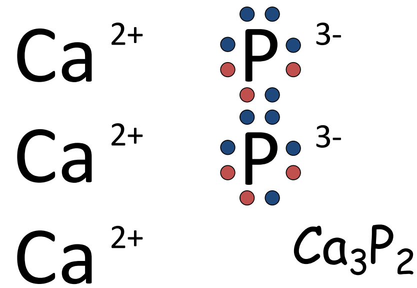
Show the electron dot structures for calcium and phosphorus using arrows to indicate the transfer of e-.



Ionic Bonding

- **EXAMPLE:**

Show the electron dot structures for calcium and phosphorus using arrows to indicate the transfer of e-.



Ionic Bonding

- **MORE EXAMPLES:**

Show the electron dot structures for each pair and show the transfer of electrons using arrows. Be sure to include the charge on each ion after the transfer and write the formula unit.

LITHIUM and FLUORINE

ALUMINUM and SULFUR

MAGNESIUM and CHLORINE

POTASSIUM and OXYGEN

Coulomb's Law

Determines the energy / strength of an ionic bond

$$F = \frac{kq_1q_2}{r^2}$$

r = distance between ions in nm (SIZE MATTERS!)

Q_1 and Q_2 = numerical ion charges (with signs)

F = electrical force between the atoms

Coulomb's Law

Determines the energy / strength of an ionic bond

Force of attraction between two oppositely charged particles is directly proportional to the MAGNITUDE of the charges and inversely proportional to the DISTANCE between those charges...

*Applies to PES!

GREATER charges = GREATER attraction
CLOSER (smaller ions) = GREATER attraction

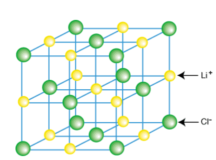
(Attraction → Opposites / Repulsion → Same)

*When calculated for an ionic compound, energy is NEGATIVE, indicating an attractive force that has a LOWER energy than the separated ions!!

Lattice Energy

Change in energy that takes place when separated GASEOUS ions are packed together to form an ionic solid... (tells how much energy needed to separate ions)

- Ionic compounds form solid crystals organized in a CRYSTAL LATTICE of alternating + and – ions

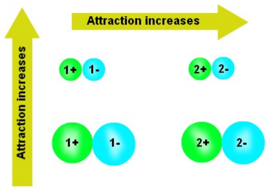


- Relates to Coulomb's Law: GREATER charges and SMALLER size mean HIGHER lattice energy (STRONGER BOND!)

Can be used to help determine properties of ionic compounds...

Lattice Energy

Energy needed to separate ions (how strong is the ionic bond)



Lattice Energy

Energy needed to separate ions (how strong is the ionic bond)

TABLE 6.3 Lattice Energies of Some Ionic Solids (kJ/mol)

	Anion				
Cation	F^-	Cl^-	Br^-	I^-	O^{2-}
Li^+	1036	853	807	757	2925
Na^+	923	787	747	704	2695
K^+	821	715	682	649	2360
Be^{2+}	3505	3020	2914	2800	4443
Mg^{2+}	2957	2524	2440	2327	3791
Ca^{2+}	2630	2258	2176	2074	3401
Al^{3+}	5215	5492	5361	5218	15,916

SMALLER ions of the same charge = STRONGER bond!

GREATER charge = STRONGER bond!

Lattice Energy

Energy needed to separate ions (how strong is the ionic bond)

Example: Formation of lithium fluoride

Process	Description	Energy Change (kJ)
$Li(s) \rightarrow Li(g)$	Sublimation energy	161
$Li(g) \rightarrow Li^+(g) + e^-$	Ionization energy	520
$1/2F_2 \rightarrow F(g)$	Bond energy (1/2 mole)	77
$F(g) + e^- \rightarrow F^-(g)$	Electron affinity	-328
$Li^+(g) + F^-(g) \rightarrow LiF(s)$	Lattice energy	-1047
$Li(s) + 1/2F_2(g) \rightarrow LiF(s)$	ΔH	-617

***Notice the "-" energy indicating the ions are attracted and the lattice energy is high because both atoms have a small atomic size!**

Lattice Energy

Calculate the ΔH_f for Sodium Chloride:

$Na(s) + \frac{1}{2} Cl_2(g) \rightarrow NaCl(s)$

Lattice Energy	-786 kJ/mol
Ionization Energy for Na	495 kJ/mol
Electron Affinity for Cl	-349 kJ/mol
Bond energy of Cl_2	239 kJ/mol
Enthalpy of sublimation for Na	109 kJ/mol

$Na(s) \rightarrow \cancel{Na}(g)$ + 109 kJ

$\cancel{Na}(g) \rightarrow \cancel{Na^+}(g) + \cancel{e^-}$ + 495 kJ

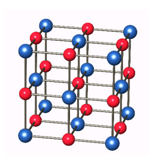
$½ Cl_2(g) \rightarrow \cancel{Cl}(g)$ + ½(239 kJ)

$\cancel{Cl}(g) + \cancel{e^-} \rightarrow \cancel{Cl^-}(g)$ - 349 kJ

$\cancel{Na^+}(g) + \cancel{Cl^-}(g) \rightarrow NaCl(s)$ -786 kJ

$Na(s) + ½ Cl_2(g) \rightarrow NaCl(s)$ -412 kJ/mol

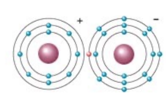
Ionic Compounds

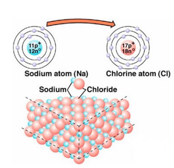


- Formed from METAL and NONMETAL ions bonded together
- Known as "SALTS"
- Formed from regular repeating arrangement of formula units

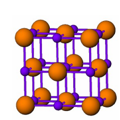
Properties of Ionic Compounds

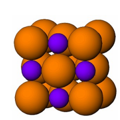
1) STRONGLY bonded: strong attraction between the ions





2) Solid CRYSTAL lattice structure



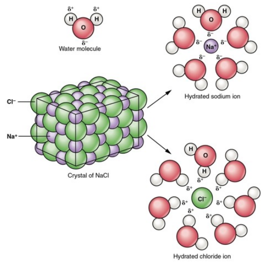


Properties of Ionic Compounds

magnesium chloride, $MgCl_2$, ionic bonding — MP = 708°C, BP = 1412°C

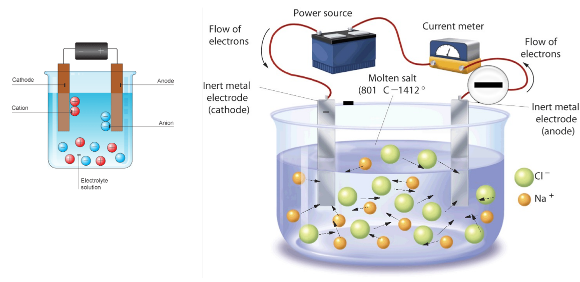
3) HIGH melting point and boiling point… due to the strong attractions between ions and the stable structure

4) SOLUBLE in water / DISSOCIATE (break apart into ions)



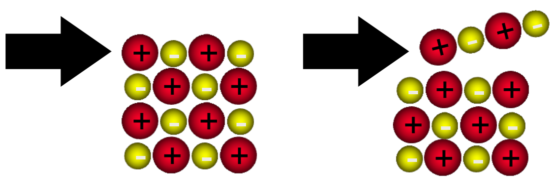
Properties of Ionic Compounds

5) Conduct ELECTRICITY when melted or dissolved in water (because they form ions)… too orderly when solid

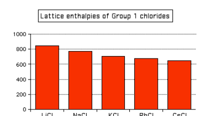


Properties of Ionic Compounds

6) Extremely BRITTLE… if you hit them hard enough they will shatter because they don't want to bend and there will be a strong repulsive force



Lattice Energy and Properties



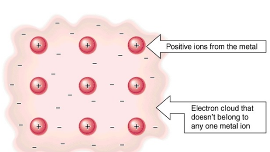
***Make sure you mention BOTH substances / ions in an AP answer!!**

- Which has a STRONGER bond? NaCl or KCl
- Which is more SOLUBLE in water? NaCl or $AlCl_3$

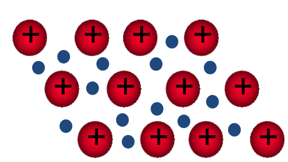
Metallic Bonds

Attraction of the FREE-FLOATING valence electrons for the positively charged metal ions

- Forces of attraction hold metals together
- Not ionic
- Have similar properties to ionic compounds
- Metals weakly hold on to their valence e-
- Positive ions (cations) are floating in a "SEA OF ELECTRONS"

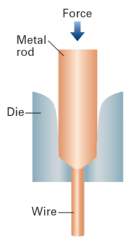


"Sea of Electrons"



ELECTRONS ARE FREE TO MOVE!!

Properties of Metals



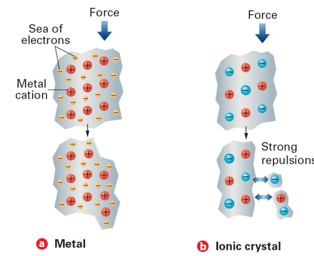
1) Malleable

2) Ductile

3) Conductive

All due to the fact that their valence electrons are MOBILE!

Properties of Metals



Electrons allow the atoms of metals to SLIDE like grease!

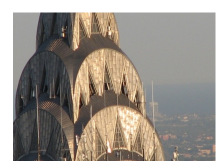
Alloys

Mixtures of two or more elements with at least one being a metal

- Formed by melting a mix of ingredients and then cooling
- Examples: brass (Cu and Zn), bronze (Cu and Sn), and steel (Fe, C, etc.)

Why Make Alloys?

- Properties are SUPERIOR to individual elements
- Sterling Silver (92.5% Ag, 7.5% Cu): harder, more durable than pure Ag / soft enough to work
- Steels: corrosion resistant, ductile, hard, tough, and cost effective



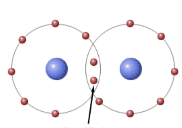
***MORE TO COME IN SECTION 3!**

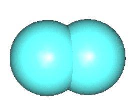
Covalent Bond

SHARING of electrons between atoms

- Involves two NONMETALS
- Known as covalent or molecular compounds
- MOLECULE: group of atoms joined by a covalent bond
- DIATOMIC MOLECULES: elements that cannot exist as single atoms

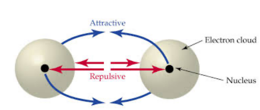
Ex: H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2





Bond Length

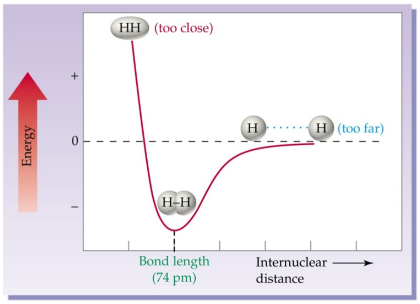
Distance where energy is at a MINIMUM



- When two atoms approach, two "bad" things occur… electron/electron repulsion and proton/proton repulsion
- One "good" thing occurs… proton/electron attraction
- When ATTRACTIVE forces offset REPULSIVE forces, energy decreases and a bond is formed… Always looking for LOWEST energy!

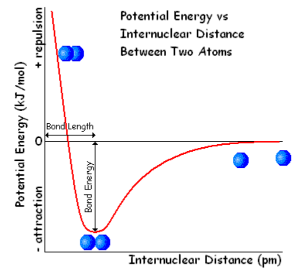
Bond Length

Distance where energy is at a MINIMUM



Bond Length

Distance where energy is at a MINIMUM

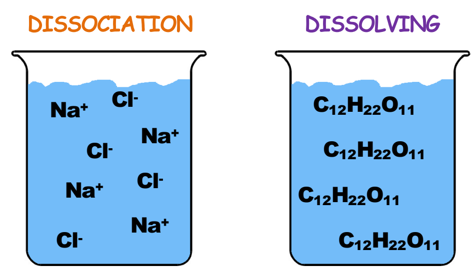


*Note how to find the bond length and energy from the graph!!

Properties of Covalent Compounds

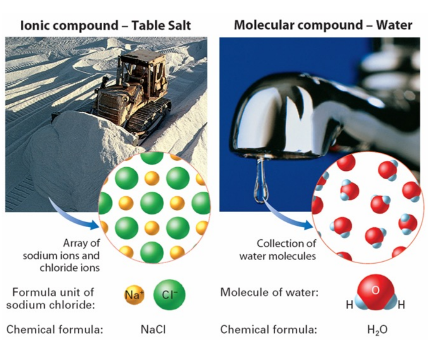
1) Low melting and boiling points
2) Don't usually conduct electricity
3) Not usually soluble in water (some dissolve but don't dissociate… SUGAR!)
4) Don't form crystals (most liquids or gases at room temperature)

Dissociation vs. Dissolving



Salt = IONIC **Sugar = COVALENT**

Properties of Covalent Compounds



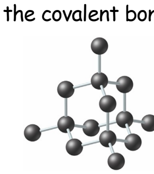
Attractions and Properties

- Wide range of physical properties among covalent compounds due to varying intermolecular attractions
- NETWORK SOLIDS: very stable structure consisting of all atoms being covalently bonded to each other

  -High melting points (1000°C or higher)

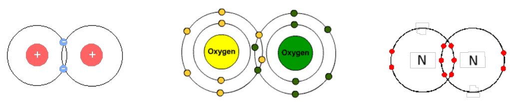
  -To melt, ALL the covalent bonds need to be broken

  EX: Diamonds
  SiO_2
  SiC



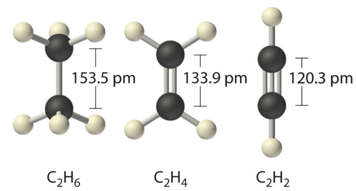
Types of Covalent Bonding

- There are three types of covalent bonds:
  1) SINGLE: sharing of only one pair of e- (2 total)
  2) DOUBLE: sharing of two pairs of e- (4 total)
  3) TRIPLE: sharing of three pairs of e- (6 total)



Remember the octet rule...
EIGHT electrons are needed!!

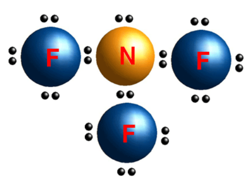
Bond Length



As the number of bonds between two atoms INCREASE, the bond length DECREASES... the bonds become SHORTER and STRONGER!!

WHY?: More electrons between the nuclei which DECREASES nuclear repulsions and INCREASES nucleus to electron attractions!!

Molecular Dot Structures



- Visual representation of how the atoms are bonded together in a molecule
- Shows valence e- as dots... can see type of covalent bond as well (sharing of TWO dots = a SINGLE bond)

Molecular Dot Structures

- How to draw a structure:

1) Determine the TOTAL number of valence e- available for the entire molecule from the amount each element has... This is the # of e- that MUST be in the final structure

2) Position any C or N atoms in the center... With atoms other than C or N, put the least electronegative atom in the center (usually the atom that there is LESS of in the formula)... H never in center unless it's the only element

3) Whatever atoms are remaining in the formula go around the "center" (evenly distributed when possible)

Molecular Dot Structures

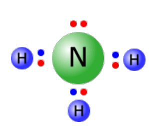
- How to draw a structure:

4) Any atoms next to the center need to SHARE at least a pair of electrons between the atoms (covalent bonding)... Insert pairs of e- between bonded atoms

5) Give each atom a complete octet by adding UNSHARED (lone) pairs until each atom is "happy" with its number

6) Total the e- in the molecule and see if total from step #1 is reached... If so, then done

7) If OVER the #, add double or triple bonds and remove pairs... If UNDER the #, add lone pairs (extra) to the center atom



DO NOT overfill an atom's valence shell... Ex: H only wants two electrons!!

HONC 1234

Certain elements USUALLY have a specific number of bonds (shared pairs) in a molecule!!!

Hydrogen – 1 bond

Oxygen – 2 bonds

Nitrogen – 3 bonds

Carbon – 4 bonds

Atoms still need a FULL valence!!

HONC IF YOU LOVE CHEMISTRY!!

Molecular Dot Structures

- **EXAMPLE:**

Write the dot structure for: NH_3

Molecular Dot Structures

- **EXAMPLE:**

Write the dot structure for: H_2O

Molecular Dot Structures

- **EXAMPLE:**

Write the dot structure for: CO_2

Molecular Dot Structures

- **EXAMPLES:**

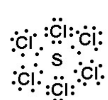
Write the dot structure for:

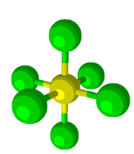
Write the dot structure for:

Exceptions to the Octet Rule

If the number of electrons doesn't work out…

1) H - full valence shell is 2 e-
2) B - **generally** satisfied with 6 valence e-
3) Be - **generally** satisfied with 4 valence e-
4) N - **can** be satisfied with 7 valence e-
5) As, S, I, Se, P, etc. can expand their octet to have 10, 12, or 14 valence e-

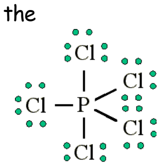




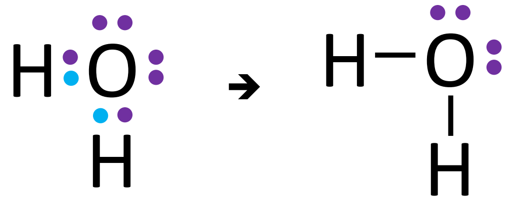
Exceptions to the Octet Rule

Why are some elements able to EXPAND their octet and take more than EIGHT?

- Can occur when an element has an EMPTY d-orbital… Electrons get put in there!
- Notice this only occurs in elements in the 3rd period or higher… WHY?
- Ex: Phosphorus



Structural Formulas



- Uses a dash (-) to represent a covalent bond
- Unshared pairs still shown as dots

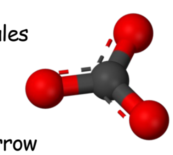
Structural Formulas

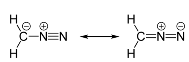
- Uses a dash (-) to represent a covalent bond
- Unshared pairs still shown as dots

Resonance Structures

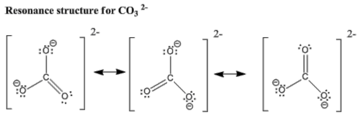
When more than one valid dot structure is possible for the same substance

- Found in many DOUBLE bonded molecules
- Ex: CO_3^{2-} or C_6H_6
- Each MUST be shown in brackets with charge in upper right… include double arrow as well!



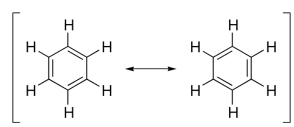


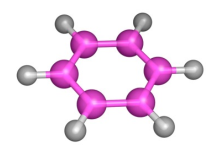
Resonance structure for CO_3^{2-}

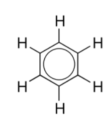


Resonance Structures

When more than one valid dot structure is possible for the same substance







Experiments show that the molecules with resonance have EQUAL bond lengths and strengths!

Bond Order

Used to determine the stability of a bond

- HIGHER bond order means MORE STABLE bond due to increased attraction between electrons, causing atoms to be held together tighter
- **Steps for Bond Order between two atoms:**
  1) Draw the molecular dot structure
  2) Count the # of bonds between the two atoms (single= 1, double= 2, triple= 3)
  3) Since there is only one type of bond between the two atoms, the bond order is the same as the # of bonds

Example:

:C≡O:

Bond Order =

Bond Order

Used to determine the stability of a bond

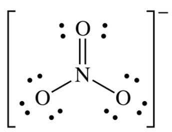
- HIGHER bond order means MORE STABLE bond due to increased attraction between electrons, causing atoms to be held together tighter
- **Steps for Bond Order for a molecule:**
  1) Draw the molecular dot structure
  2) Total the # of bonds between ALL the bonding groups (single= 1, double= 2, triple= 3)
  3) Divide total # of bonds by the # of bonding groups (attached atoms)

Bond Order

Used to determine the stability of a bond

- HIGHER bond order means MORE STABLE bond due to increased attraction between electrons, causing atoms to be held together tighter
- **EXAMPLE:**

  Determine the bond order for the nitrate ion (NO_3^-).



Bond Order =

Formal Charge

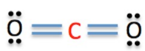
Tells which atom carries the charge in nonequivalent (more than one possibility) Lewis structures… assumes perfect sharing of electrons

$$\text{FC} = \underset{\substack{\uparrow \\ \text{Group \# (Ex: 4A)}}}{\begin{matrix}\text{\# Valence e}^- \\ \text{in Free Atom}\end{matrix}} - \begin{matrix}\text{Total \#} \\ \text{Nonbonding e}^-\end{matrix} - \underset{\substack{\uparrow \\ \text{Or just \# of bonds}}}{\frac{\text{Total \# Bonding e}^-}{2}}$$

- Sum of <u>all</u> FC MUST EQUAL the ion's charge (if no charge, then should add up to zero)

Formal Charge

Tells which atom carries the charge in nonequivalent Lewis structures… assumes perfect sharing of electrons



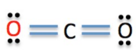
$Formal\ Charge = VE - NE - \frac{BE}{2}$

$VE = 4$
$NE = 0$
$BE = 8$

We now substitute the values and we have:

$Formal\ Charge = 4 - 0 - \frac{8}{2} = 0$

$Formal\ Charge\ of\ Carbon = 0$



$Formal\ Charge = VE - NE - \frac{BE}{2}$

$VE = 6$
$NE = 4$
$BE = 4$

We now substitute the values and we have:

$Formal\ Charge = 6 - 4 - \frac{4}{2} = 0$

$Formal\ Charge\ of\ Oxygen\ (left) = 0$

Formal Charge

Tells which atom carries the charge in nonequivalent Lewis structures… assumes perfect sharing of electrons

$$\text{FC} = \underset{\substack{\uparrow \\ \text{Group \# (Ex: 4A)}}}{\begin{matrix}\text{\# Valence e}^- \\ \text{in Free Atom}\end{matrix}} - \begin{matrix}\text{Total \#} \\ \text{Nonbonding e}^-\end{matrix} - \underset{\substack{\uparrow \\ \text{Or just \# of bonds}}}{\frac{\text{Total \# Bonding e}^-}{2}}$$

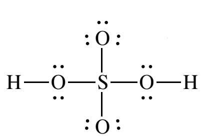
- If MORE than one possible structure could be drawn…
  1) Want structure with the LOWEST sum when the absolute values of all the FCs are added
  2) If still more than one, structure with the "-" FC on the MOST ELECTRONEGATIVE element is the best

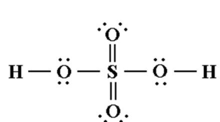
Formal Charge

Tells which atom carries the charge in nonequivalent Lewis structures… assumes perfect sharing of electrons

- **EXAMPLE:**

  Which is the correct structure for sulfuric acid?



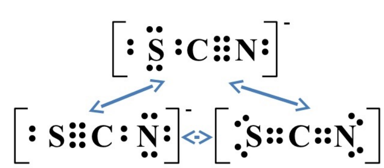


Formal Charge

Tells which atom carries the charge in nonequivalent Lewis structures… assumes perfect sharing of electrons

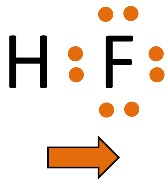
- **EXAMPLE:**

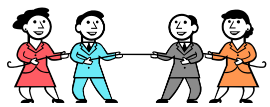
  Which is the correct structure for SCN^-?



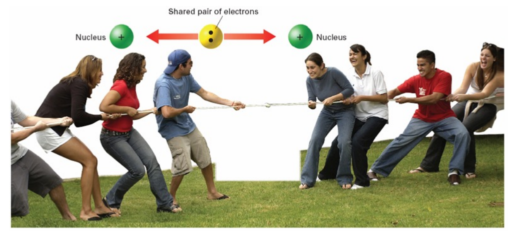
Bond Polarity

- Covalent bonds involve the sharing of electrons but not all atoms share equally!
- Electrons are in a "TUG OF WAR" between atoms





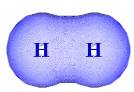
Bond Polarity



Nonpolar Covalent Bonds

EQUAL sharing of electrons

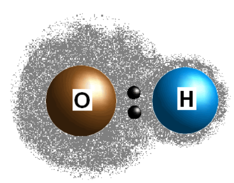
- Compounds NOT soluble in water ("like dissolves like")
- LOWER boiling and melting point than polar compounds



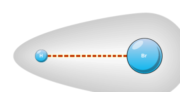
Cl_2

Polar Covalent Bonds

Electrons are NOT shared equally (one atom pulls the e- more)



- Compounds are SOLUBLE in water ("like dissolves like")
- HIGHER boiling and melting point than nonpolar compounds (but still low)



Polarity

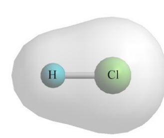
- Difference in ELECTRONEGATIVITY (**ability of an atom to pull e- towards itself**) between the atoms determines what type of bond forms…

Electronegativity Difference	Type of Bond	Example
0.0 - 0.4	Nonpolar covalent	S-Se 2.5 - 2.4 = 0.1
0.5 - 2.0	Polar covalent	Cl-P 3.5 - 2.1 = 1.4
> 2.0	Ionic	Na-F 4.0 - 0.9 = 3.1

Polarity

- **EXAMPLE:**

HCl

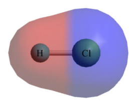


- H = electronegativity of 2.1
- Cl = electronegativity of 3.0
- POLAR BOND → difference is 0.9
- Cl "hogs" shared e- giving it a slight negative charge
- H has shared e- less time giving it a slight positive charge

Partial Charges

Slight regions of charge that form due to e- "hogging"

- Much less than a 1+ or 1- in ionic bonding
- **DELTA** denotes partial charges (δ^+, δ^-)
- Written as:



$$\overset{\delta+}{\mathrm{H}}-\overset{\delta-}{\mathrm{Cl}} \quad \text{or} \quad \overset{\longmapsto}{\mathrm{H}-\mathrm{Cl}}$$

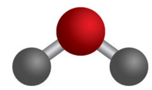
Having a polar bond and the right shape make HCl a POLAR MOLECULE and it is a DIPOLE (molecule with two poles)!!

Polar Molecules

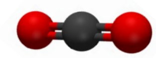
- For a molecule to be POLAR it must have:

-At least one polar bond

-Asymmetrical shape (lone pairs on center… sometimes helps ID this) or asymmetrical atoms



Water



Carbon dioxide

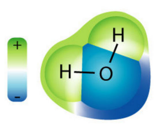
Just having polar bonds does NOT make a molecule polar!!

Polar Molecules

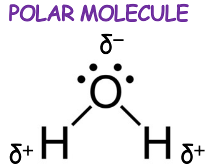
- **EXAMPLE:**

H_2O

- 2 Polar Bonds (O-H)
- Highly electronegative O pulls the e- away from the H
- Asymmetrical shape



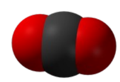
POLAR MOLECULE



Polar Molecules

- **EXAMPLE:**

CO_2

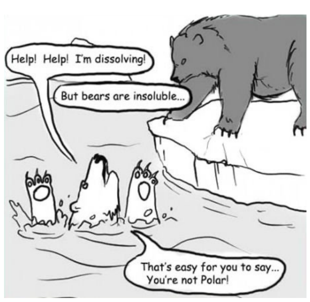


- 2 Polar Bonds (C-O)
- Highly electronegative O pulls the e- away from the C
- Symmetrical shape / cancels dipole

NONPOLAR MOLECULE

$$\overset{\longleftarrow\!\!+}{\mathrm{O}=}\mathrm{C}\overset{+\!\!\longrightarrow}{=\mathrm{O}} \qquad \overset{\delta-}{\mathrm{O}}=\overset{\delta+}{\mathrm{C}}=\overset{\delta-}{\mathrm{O}}$$

Properties

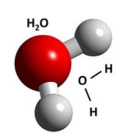


"LIKE DISSOLVES LIKE"

Bond Enthalpy

Energy required to break the bond between two covalently bonded atoms

(aka Bond Energy)



- Bonds **BREAKING** = Endothermic
- Bonds **FORMING** = Exothermic
- Change in enthalpy (energy) for a reaction can be found using the bond energies:

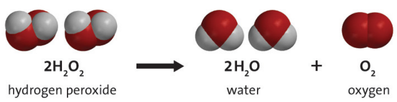
$$\Delta H_{rxn} = \Sigma(\text{energies of bonds BROKEN}) - \Sigma(\text{energies of bonds FORMED})$$

Bond Enthalpy

- **How to Calculate for a Reaction:**

1) Draw the dot structures for all the substances in order to see which bonds will break and form
2) Sum the energies of ALL of the bonds that are broken and subtract the sum of the energies of ALL of the bonds that are formed
3) Amounts of each substance (from the reaction equation) need to be considered when calculating

EX:



Average Bond Energies (kJ/mol)

H-H	432	N-H	391	Cl-Cl	239	C=C	614
H-F	565	N-N	160	Cl-Br	218	C≡C	839
H-Cl	427	N-F	272	Br-Br	193	O=O	495
H-Br	363	N-Cl	200	I-I	149	C=O	799
H-I	295	N-Br	243	I-Cl	208	C≡O	1074
C-H	413	N-O	201	I-Br	175	N=O	607
C-C	347	O-H	467	S-H	347	N=N	418
C-N	305	O-O	146	S-F	327	N≡N	941
C-O	358	O-F	190	S-Cl	253	C=N	615
C-F	485	O-Cl	203	S-Br	218	C≡N	891
C-Cl	339	O-I	234	S-S	266		
C-Br	276	F-F	154	Si-Si	340		
C-I	240	F-Cl	253	Si-H	393		
C-S	259	F-Br	237	Si-C	360		
				Si-O	452		

These are just an AVERAGE for each!

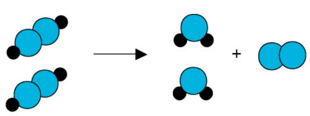
Higher energy = STRONGER bond!

Bond Enthalpy

- **EXAMPLE:**

Calculate the change in enthalpy for the reaction using bond energies.

$2\ H_2O_2 \rightarrow 2\ H_2O + O_2$



Bond Enthalpy

- **EXAMPLE:**

Calculate the change in enthalpy for the reaction using bond energies.

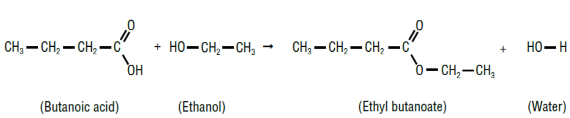
$H_2 + CO_2 \rightarrow H_2O + CO$

$CO + 2\ H_2 \rightarrow CH_3OH$

Bond Enthalpy

- **EXAMPLE:**

The formation of ethyl butanoate, one of the compounds that give pineapple its flavor, is produced according to the reaction below. Calculate the change in enthalpy for the reaction using bond energies.

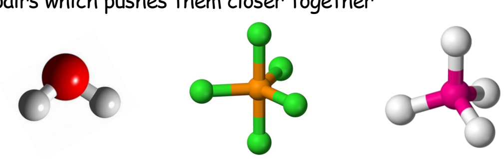


VSEPR Theory

Valence Shell Electron Pair Repulsion

Valence electron pairs attempt to get as FAR APART as possible, thus changing the 3D shape of the molecule!!!!

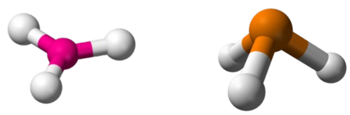
- Molecules are really 3D, not 2D
- Unshared (LONE) pairs are held closer to the atom and they repel other lone pairs and bonding pairs which pushes them closer together



VSEPR Theory

Valence Shell Electron Pair Repulsion

- Types of e- pairs:
  - **BONDING PAIRS** → electrons that form the bonds
  - **LONE PAIRS** → non-bonding electrons

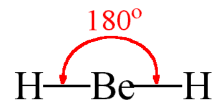


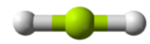
LONE PAIRS REPEL STRONGER THAN BONDING PAIRS!!!

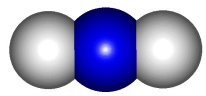
Common Molecular Shapes

- **LINEAR:**

BeH_2





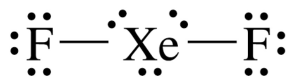


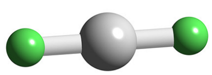
Type: $\mathbf{AB_2}$
180° bond angle
2 atoms attached
0 lone pairs

Common Molecular Shapes

- **LINEAR (another):**

XeF_2



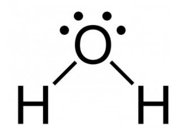


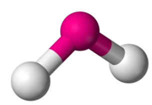
Type: $\mathbf{AB_2E_3}$
180° bond angle
2 atoms attached
3 lone pairs

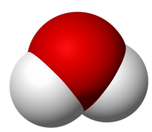
Common Molecular Shapes

- **BENT:**

H_2O





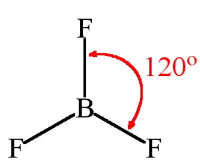


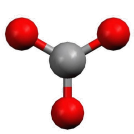
Type: $\mathbf{AB_2E_2}$
104.5° bond angle
2 atoms attached
2 lone pairs (or just 1)

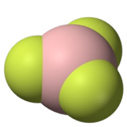
Common Molecular Shapes

- **TRIGONAL PLANAR:**

BF_3





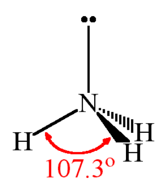


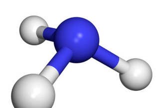
Type: $\mathbf{AB_3}$
120° bond angle
3 atoms attached
0 lone pairs

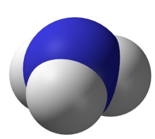
Common Molecular Shapes

- **TRIGONAL PYRAMIDAL:**

NH_3





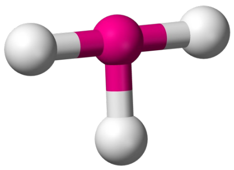


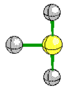
Type: $\mathbf{AB_3E}$
107.3° bond angle
3 atoms attached
1 lone pair

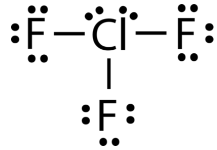
Common Molecular Shapes

- **T-SHAPED:**

ClF_3





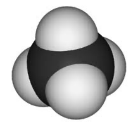


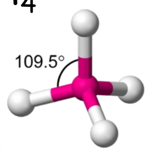
Type: AB_3E_2
90°/180° bond angle
3 atoms attached
2 lone pairs

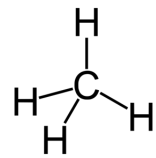
Common Molecular Shapes

- **TETRAHEDRAL:**

CH_4





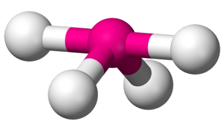


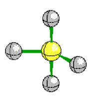
Type: AB_4
109.5° bond angle
4 atoms attached
0 lone pairs

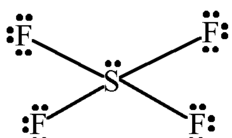
Common Molecular Shapes

- **SEESAW:**

SF_4





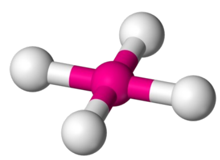


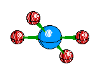
Type: AB_4E_1
90°/120°/180° angles
4 atoms attached
1 lone pair

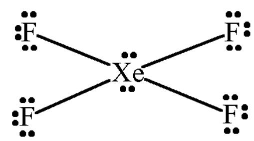
Common Molecular Shapes

- **SQUARE PLANAR:**

XeF_4





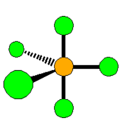


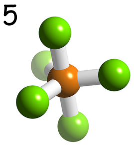
Type: AB_4E_2
90°/ 180° bond angle
4 atoms attached
2 lone pairs

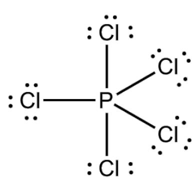
Common Molecular Shapes

- **TRIGONAL BIPYRAMIDAL:**

PCl_5





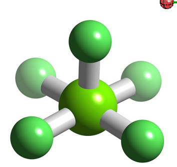


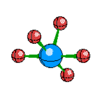
Type: AB_5
90°/120° bond angle
5 atoms attached
0 lone pairs

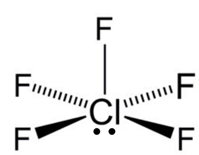
Common Molecular Shapes

- **SQUARE PYRAMIDAL:**

ClF_5





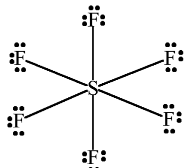


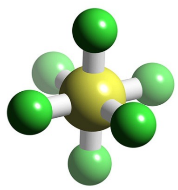
Type: AB_5E
90°/180° bond angle
5 atoms attached
1 lone pair

Common Molecular Shapes

- **OCTAHEDRAL:**

SF_6





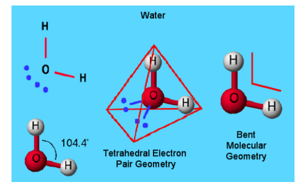
Type: $\mathbf{AB_6}$
90° bond angle
6 atoms attached
0 lone pairs

Electronic vs. Molecular

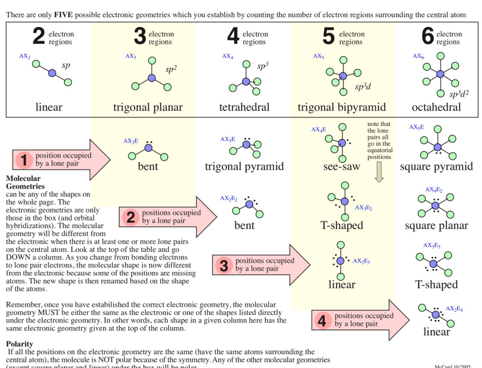
- ELECTRONIC DOMAINS: based on the electron groups on the central atom (ignores shared vs. lone pairs) / helps with bond angle
- MOLECULAR GEOMETRY: used to describe the shape of the actual molecule
- EXAMPLE: H_2O

4 Electron Domains… Tetrahedral Electron Geometry! (4 attached)

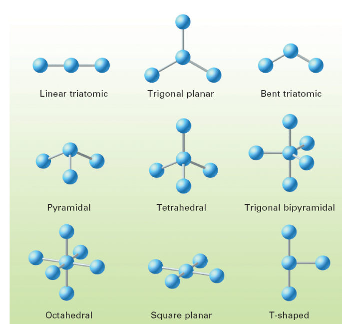
But, a Bent Molecular Geometry! (2 attached and 2 lone pairs)



Electronic vs. Molecular



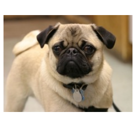
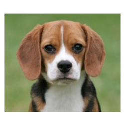
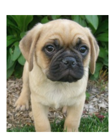
Common Molecular Shapes



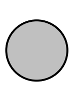
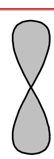
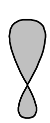
***NOTE: Double or triple bonds in a molecule often give linear or planar shapes!!**

Orbital Hybridization Theory

Combines information about molecular bonding and molecular shape

 +  = 

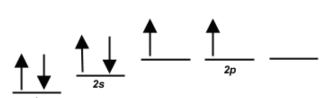
Pug + Beagle = Puggle

 +  = 

s orbital + *p* orbital = *sp* orbital

Orbital Hybridization Theory

- Lets look at CARBON:



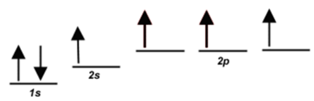
Can you see a problem with this?

Carbon only has TWO electrons available for bonding… That is not enough!

So what's going on here??

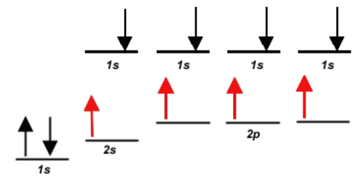
Orbital Hybridization Theory

- First thought was that carbon **promoted** one of its 2s electrons to the **empty** 2p orbital



- Now try to bond to four hydrogen atoms to form methane (CH_4)…

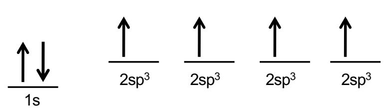
Orbital Hybridization Theory



- Three of the C-H bonds would involve a 2p matched with a 1s, but a fourth bond would be between a 2s and a 1s
- This one bond would have slightly LESS energy than the other bonds in methane… this is NOT what chemists observe though!

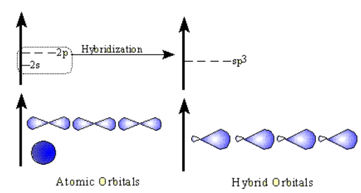
Orbital Hybridization Theory

- Measurements show that all four bonds are **EQUAL** so they must be **HYBRIDIZED**
- For methane, the s orbital combines with the three p orbitals to create four equal **sp^3** hybridized orbitals



New orbitals have slightly MORE energy than the 2s and slightly LESS energy than the 2p!

Orbital Hybridization Theory

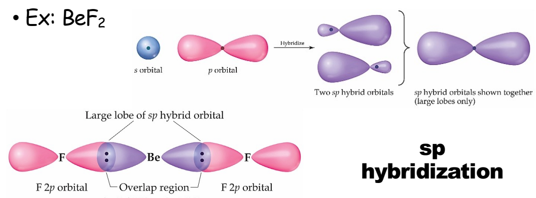


Notice the energy is in between the 2s and 2p!

Orbital Hybridization Theory

Combines information about molecular bonding and molecular shape

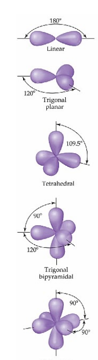
- Covalent bonds form when atomic orbitals overlap
- Mixing of s, p, and sometimes d orbitals that allows bonds to form… HYBRID ORBITALS
- Ex: BeF_2



sp hybridization

Shape Hybridizations

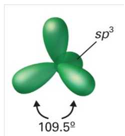
- LINEAR (2 attached): sp
- TRIGONAL PLANAR (3 attached): sp^2
- TETRAHEDRAL (4 attached): sp^3
- TRIGONAL BIPYRIMIDAL (5 attached): sp^3d
- OCTAHEDRAL (6 attached): sp^3d^2



*EACH lone pair of e- counts as an "attached" group too!!

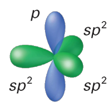
Depends on which types of orbitals are being mixed!

Multiple Bond Hybridizations



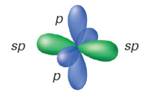
Single Bonds

(Just COUNT number of bonds to determine)



Double Bonds

sp^2

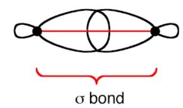


Triple Bonds

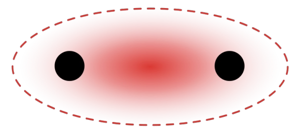
sp

Sigma (σ) Bonds

Electrons are located around the central axis of a bond



- Single bonds
- Ex: H_2



Pi (π) Bonds

Electrons are located above and below the axis of a bond

- Weaker than sigma bonds
- Double bonds = 1 sigma / 1 pi
- Triple bonds = 1 sigma / 2 pi
- Ex: O_2 / N_2