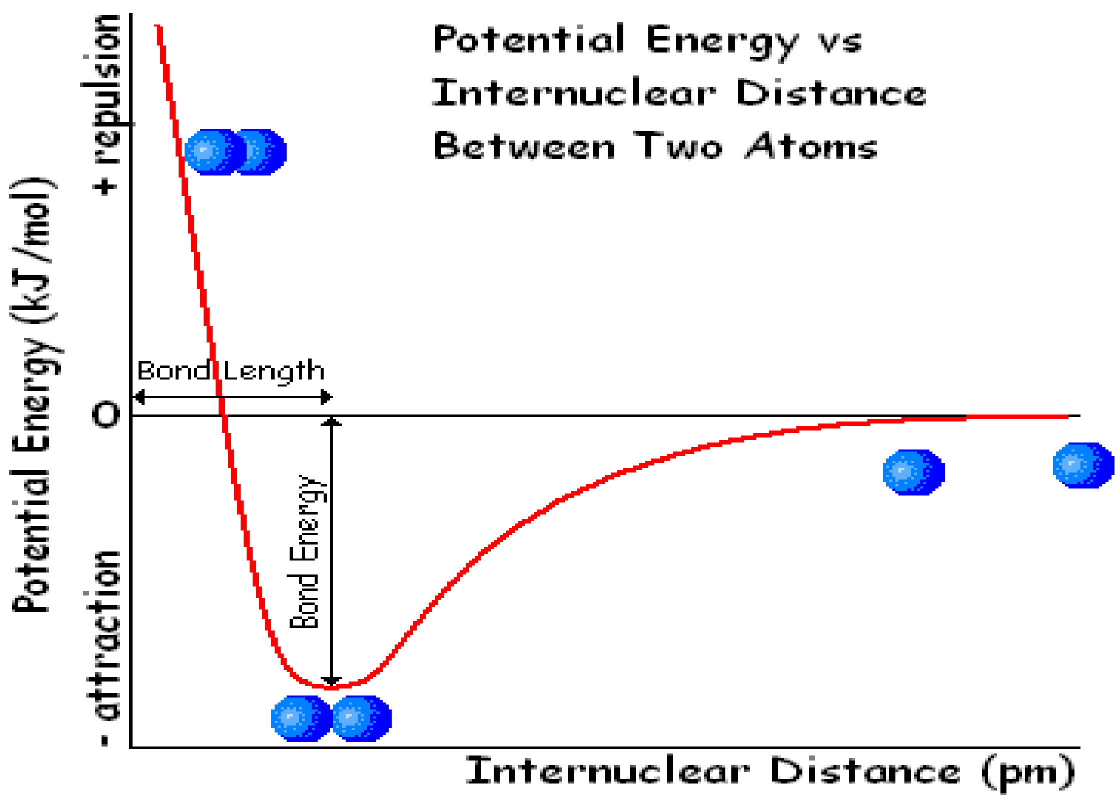
Chemical Bonding

All bonding involves the interactions of electrons
We use theories and models to explain and predict

- How many atoms are connected
- Which atoms are connected
- Molecular shape
- Strength of the bond and the bond energy

Covalent Bonding

Usually between two non metals, where electrons are **shared** between the bonded atoms.
Forces of attraction between the e^- of one atom and the protons of the other atom must be **greater** than the **repulsive forces** between like particles
Bond occurs at the **lowest potential energy**



Sharing of e^- results in a more stable configuration with electron pairing in unpaired orbitals
Covalent compounds have definite shapes
If atoms have similar electronegativities a **non-polar covalent bond** forms

Polar Covalent Bonding

When e^- are shared **unequally** between 2 atoms a **polar covalent bond** is formed
Polar covalent bonds form between atoms which have **electronegativity** differences between ~0.2 and ~2.0
Molecules containing a polar bond CAN have overall molecular polarity.
A polar bond is called a **dipole**

			$\delta+$ $\delta-$	
Ex.	**Bond:**	HCl	H –– Cl	3.0-2.1 = 0.9
	Electronegativity:		2.1 – 3.0	
			$\delta+$ $\delta-$	
Ex.	**Bond:**	CCl_4	C –– Cl	3.0-2.5 = 0.5
	Electronegativity:		2.5 – 3.0	

Electronegativities form a continuum which shows the degree of how much an element wants electrons (with 0 being low, and 4 being high)

Ionic Bonding

When the difference in electronegativities is greater than ~2.0 we say the bond is **ionic** (this usually forms between metals and non-metals)
The electrons are **transferred** from one atom to another
The **electrostatic attractions between the ions occur** and crystals form, resulting in a decrease in potential energy and generation of heat.
Form between metals and non-metals
Very **stable** compounds with high melting points

Lewis Electron Dots (recall from sci 10)

Dots are used to represent electrons, only electrons in outermost **s** and/or **p** orbitals. Valence electrons
Dots are written in pairs on each side of the atomic symbol (symbol represents the nucleus and inner shells of electrons)

Ex.	F	$1s^2\mathbf{2s^2 2p^5}$	7 valence electrons	
	Cl	$1s^2 2s^2 2p^6 \mathbf{3s^2 3p^5}$	7 valence electrons	
	N	$1s^2\mathbf{2s^2 2p^3}$	5 valence electrons	

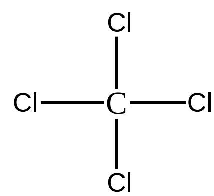
Octet Rule

Most **bonded non-metallic** elements have atoms with **8 electrons** in their outermost energy levels
Exceptions exist: **H** only has **2 e$^-$**, can have > 8 e$^-$ if accessible ***d*** orbitals)

To determine the bonding of a molecule:
    Determine the bonding experimentally
    Make an assumption

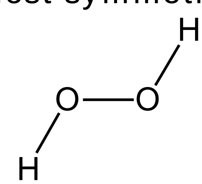
Either   a)   if formula has the form AX_n (one atom of one element and several of another) assume that the single atom is in the middle

Ex. CCl_4



or   b)   molecules will form the most symmetrical shape possible

Ex. H_2O_2

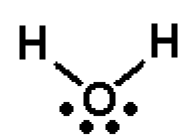


Lewis structures with Molecules

When dealing with molecules, you need to use the NAS technique (next topic) to figure out how many bonds you will have, then finish off the atoms with lone pairs (non-bonded sets of electrons, usually common in highly electronegative elements). Have to make sure every atom gets their filled octet (with exceptions, like hydrogen, discussed below)

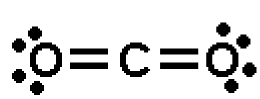
Ex. H_2O (has 2 bonds)

Total number of electrons: 8



Ex. CO_2 (has 4 bonds)

Total number of electrons: 16



NAS technique:

Needed - how many electrons in total must the atoms share to satisfy the octet rule?

– **A**vailable - What is the total number of valence electrons in all the atoms of the molecule?

Shared - how many electrons must be shared?

Ex. Single bonds

CCl_4

NH_3

C_2H_6

Ex. double bonds

O_2

C_2H_4

Ex. triple bonds

N_2

Ex. Polyatomic Ions

OH^-

SO_3^{-2}

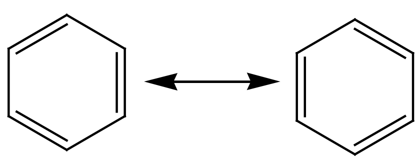
NH_4^+

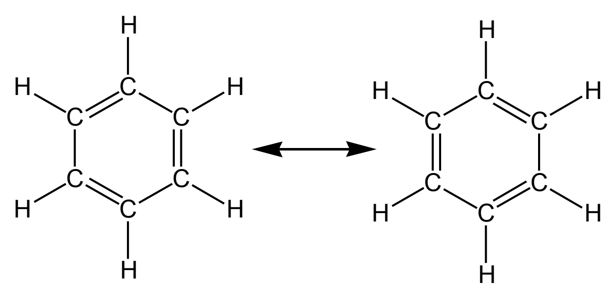
Resonance

One limitation of the electron dot model is that it doesn't explain bonds which are in between single, double and triple bonds (ex. 1.5, or 2.5). Resonance structures are used to represent a molecule when no single electron dot structure adequately depicts the molecule.
The structures *don't actually exist* but if taken together they provide a better representation than any one diagram, the actual molecule is a composite of the different structures.

Ex. Ozone O_3

Ex. Benzene C_6H_6





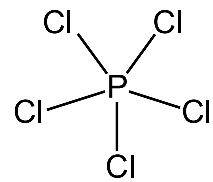
Exceptions

Some groups are exceptions to the octet rule. Instead of requiring 8 electrons in their bonded orbitals, they require:

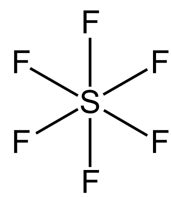
H	2	ex: H–H	H_2
Be	4	ex: H–Be–H	BeH_2
B	6	ex: H–B–H H	BH_3

Some compounds involve bonding with **d** orbitals in addition to the **s** and **p** orbital electrons (can have more than 8 bonded electrons)

Ex. PCl_5



SF_6



Some compounds can't be explained using this method and require more sophisticated bond theories (as discussed in AP chem, or university)

Ex. NO

Shapes of Molecules

Valence **S**hell **E**lectron **P**air **R**epulsion Theory (**VSEPR**)

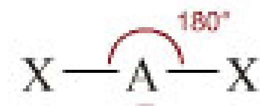
Used to predict shapes of molecules.

Bonded atoms *and lone pairs* of electrons are arranged around a central atom as far apart as possible to minimize repulsion.

Number of total electron pairs is called the Steric Number (**SN**)

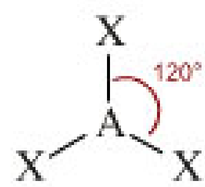
1) Linear

- SN = **2**
- Bond angle: 180°
- Ex: $BeCl_2$



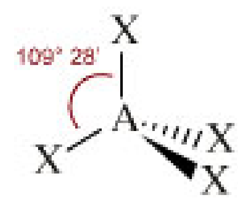
2) Trigonal Planar

- SN = **3**
- Bond angle: 120°
- Ex: BCl_3, SO_3



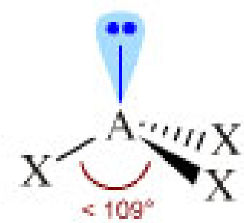
3) Tetrahedral

- SN = **4**
- Bond angle: 109.5°
- Ex: CH_4, $SiCl_4$



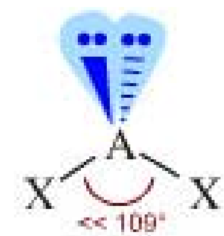
i. Trigonal Pyramidal

- SN = **4**
- 1 lone pair
- Bond angle: <109.5°
- Ex: NH_3



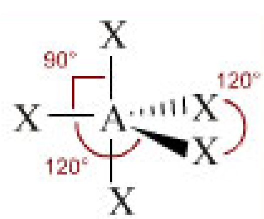
ii. Angular / Bent

- SN = **4**
- 2 lone pairs
- Bond angle: <109.5°
- Ex: H_2O



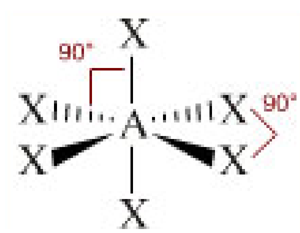
4) Trigonal Bipyramidal

- SN = **5**
- Bond angle: 90 and 120°
- Ex: PCl_5

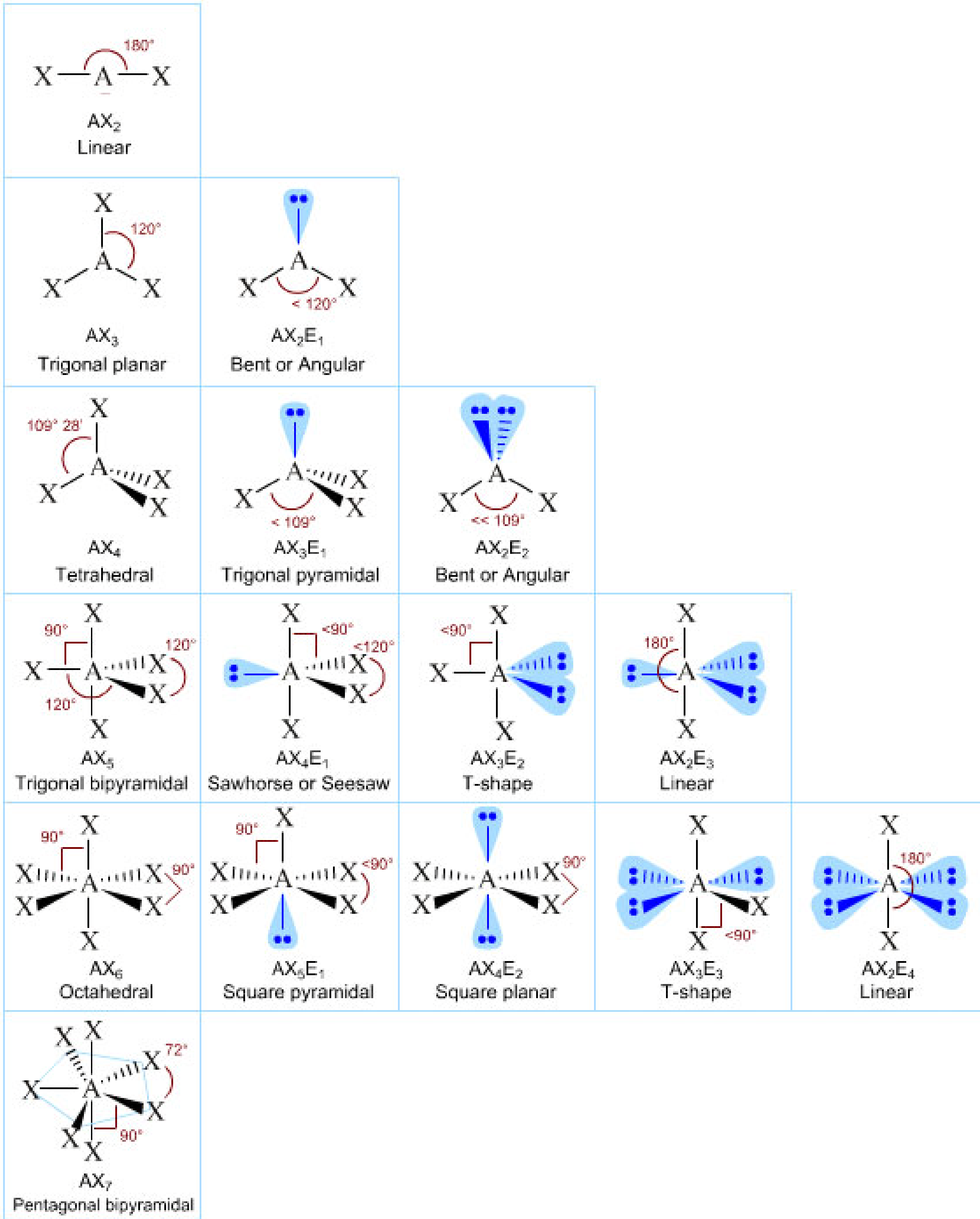


5) Octahedral

- SN = **6**
- Bond angle: 90 and 180°
- Ex: SF_6



All the rest!!!


180°
AX_2
Linear
120°
AX_3
Trigonal planar
< 120°
AX_2E_1
Bent or Angular
109° 28'
AX_4
Tetrahedral
< 109°
AX_3E_1
Trigonal pyramidal
<< 109°
AX_2E_2
Bent or Angular
90°
120°
120°
AX_5
Trigonal bipyramidal
<90°
<120°
AX_4E_1
Sawhorse or Seesaw
<90°
AX_3E_2
T-shape
180°
AX_2E_3
Linear
90°
90°
AX_6
Octahedral
90°
<90°
AX_5E_1
Square pyramidal
90°
AX_4E_2
Square planar
<90°
AX_3E_3
T-shape
180°
AX_2E_4
Linear
72°
90°
AX_7
Pentagonal bipyramidal

Molecular Polarity

Polar covalent bonds *sometimes* make polar molecules, for molecules with polar covalent bonds:

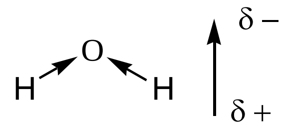
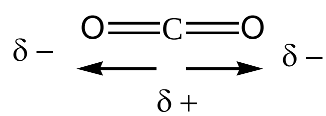
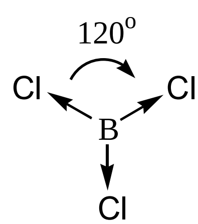
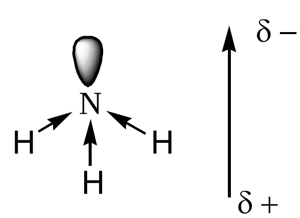
Polar molecules cannot be symmetrical – **ALWAYS POLAR**

- Linear with 2 different atoms
- Angular
- Trigonal pyramidal

Non-polar molecules where dipoles "cancel out" – **SOMETIMES POLAR**

- Linear with 3 atoms
- Trigonal planar
- Tetrahedral
- Trigonal bipyramidal
- Octahedral

Ex.

	Bond Polarity	**Molecular Polarity**	
H_2O	Polar covalent 3.5 – 2.1 = 1.4		Polar molecule (imbalanced)
CO_2	Polar covalent 3.5 – 2.5 = 1.0		Non-polar molecule (dipoles cancel)
BCl_3	Polar covalent 3.0 – 2.0 = 1.0		Non-polar molecule (dipoles cancel)
NH_3	Polar covalent 3.0 – 2.1 = 0.9		Polar molecule (imbalanced)

Intermolecular Forces

Van Der Waals Forces: Forces of attraction between positive and negative dipoles

Non – dipole forces

London (dispersion) forces arise when momentary dipoles occur due to uneven distribution of electrons.
Very weak

Dipole – Dipole Forces

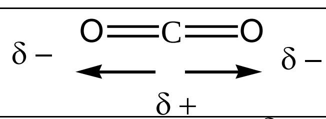
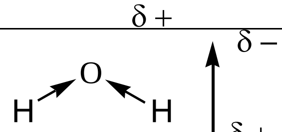
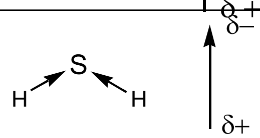
Simple *dipole-dipole*: attraction between molecules that have permanent dipoles. Positive end attracted to negative end (and vice versa)
Weak attraction
Ex. HCl

Hydrogen Bonds: special type of dipole-dipole force that occurs when H is bonded to a very electronegative element with a small radius (**F / O / N**)
Stronger than most dipole-dipole interactions
Ex. H_2O

Consequences of IM forces

Intermolecular forces account for the differences in properties of non-polar and polar compounds. Explain the differences in boiling point for the following compounds: (what affects BP in molecules?)

CO_2	-78.4^{o}C	all small molecules having 3 atoms &
H_2O	100.0^{o}C	relatively small molar mass, so why are
H_2S	-60.3^{o}C	the BP's so different?

	Structure	Molecular Polarity	Largest IM Force	Relative BP
CO_2		Non-polar	London Forces	Lowest
H_2O		Polar	Hydrogen Forces	Highest
H_2S		Polar	Dipole-Dipole Forces	Middle

Bond length and Strength

Bond energy is the energy required to separate 2 atoms that are bonded; indicates the strength of the bond.

$$H_{2(g)} + 436\ kJ \leftrightarrow H_{(g)} + H_{(g)}$$

Type of Bond	Bond Strength	Bond Length	
Single	347 kJ/mol	Longest 0.154 nm	c-----------------c
Double	619 kJ/mol	Medium 0.134 nm	c=======c
triple	812 kJ/mol	Shortest 0.120 nm	c≡≡c

Metallic Bonding

Outer electrons in metals are free to move among atoms.

Sometimes called a “sea of electrons” or “electron soup” because electrons are free to move around (aka delocalized).

Delocalization of e^- accounts for the strength, malleability, ductility and conductivity of metals.