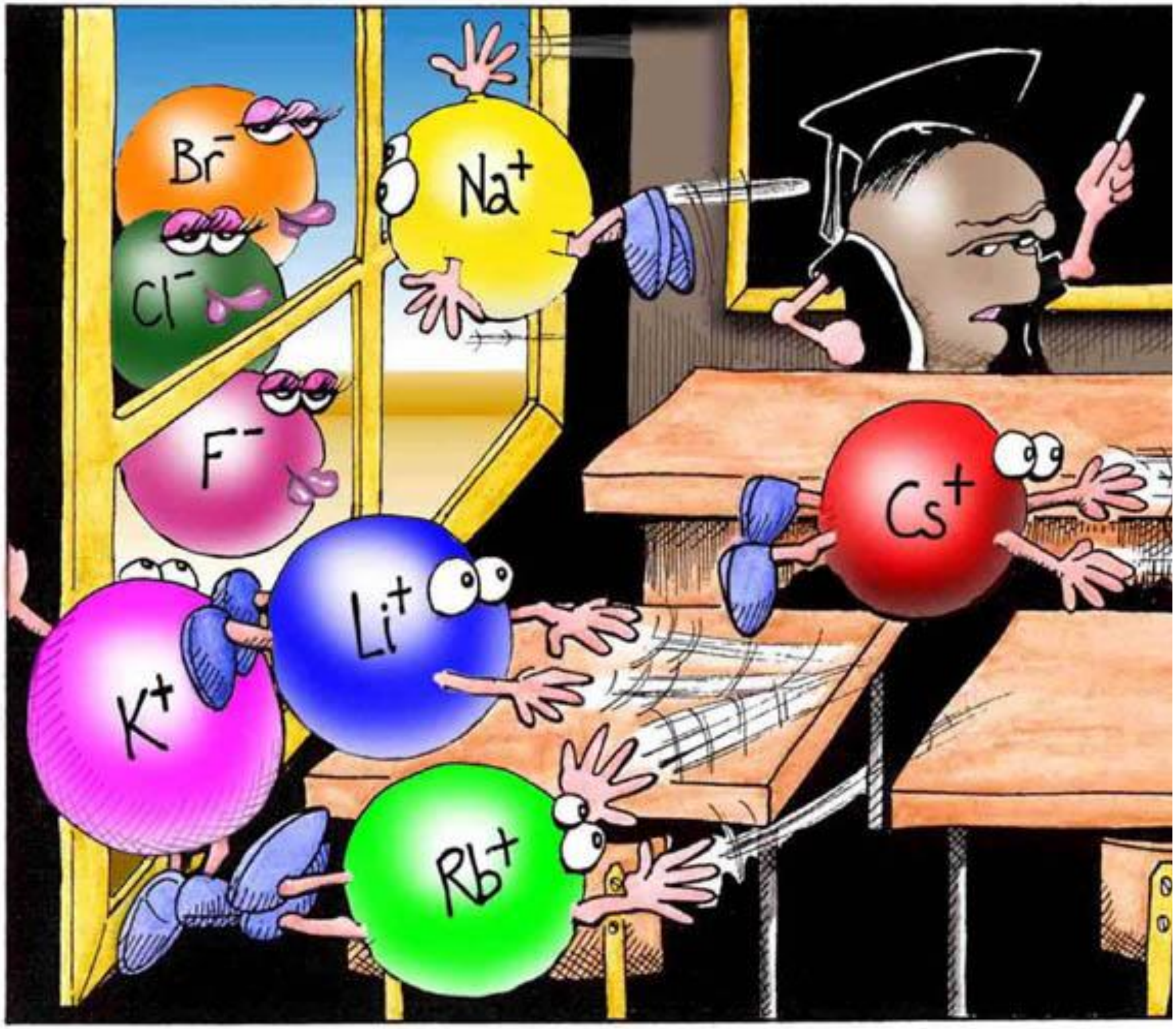


"Perhaps one of you gentlemen would mind telling me just what it is outside the window that you find so attractive..?"

Cartoon courtesy of NearingZero.net

Unit 4: Chemical Bonding & Molecules

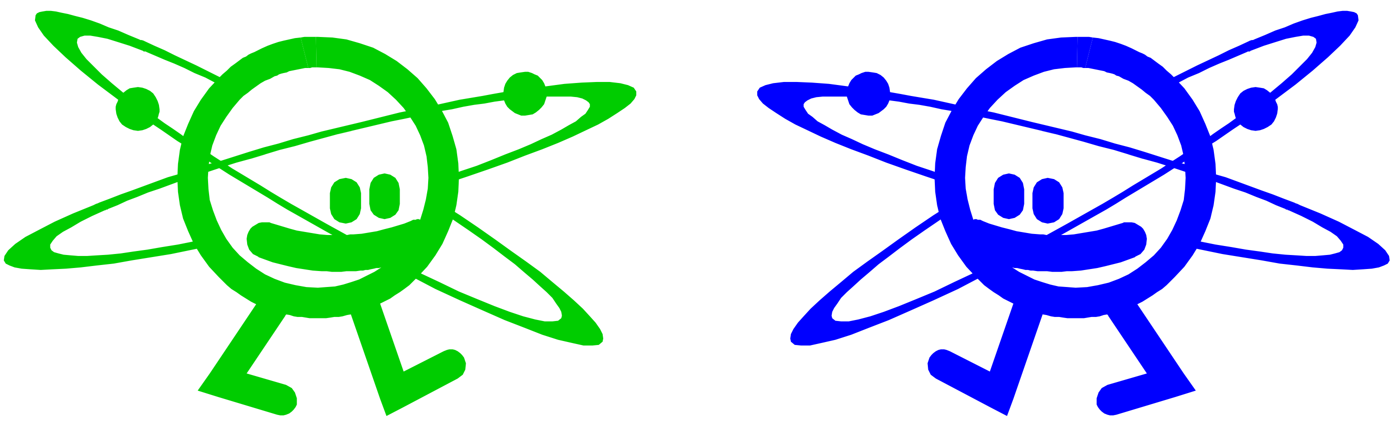
(Chapter 6 in book)

Chemical Bonding

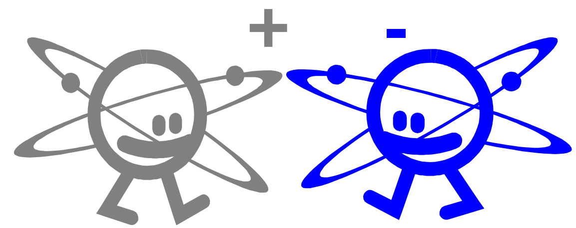
pgs.161-182

Chemical Bonds

- Attraction between the *nuclei* and *valence electrons* of different atoms that “glues” the atoms together.
  - the difference between materials as diverse as diamonds and pencils is how they're glued together.
- Why?
  - Bonded atoms are more stable than solo atoms
- How?
  - Atoms will share or exchange valence electrons to achieve a full outer shell (usually octet).

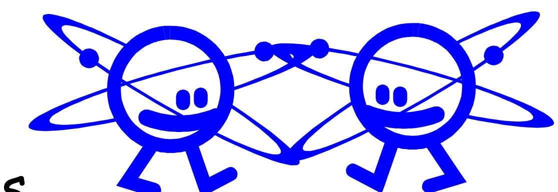


3 Main Types of Bonds



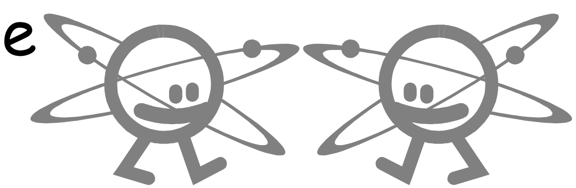
Ionic Bonds – Transfer of electrons between atoms

- electrical attraction between *cations* & *anions*
- *Formed by: metals & non-metals*



Covalent Bonds – sharing of electrons between atoms

- “co” = sharing, “valent” = outer electrons
- *Formed by: non-metals & non-metals*



Metallic Bonds – Metal atoms that share a “sea of electrons”

- *Formed by: metals & metals*

Predicting Bond Types

- Bonding is not usually purely ionic or covalent, but somewhere in between

- The difference in *electronegativity* strength of the atoms in a bond can help us estimate what percentage of the bond will be ionic

    (see example on next slide)

Using the Periodic Table to Determine Bond Types

Electronegativity

- below 1.0
- 1.0–1.4
- 1.5–1.9
- 2.0–2.4
- 2.5–2.9
- 3.0–4.0

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17
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 1.8	Ni 1.8	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 0.8	Ba 0.9	La* 1.1	Hf 1.3	Ta 1.5	W 2.4	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2
Fr 0.7	Ra 0.9	Ac† 1.1														

*Lanthanides: 1.1–1.3
†Actinides: 1.3–1.5

Ionic bond=

- metal (weak) & non-metal (strong)
- huge difference in strength (1.7 or more)

Covalent bond =

- 2 non-metals (strong)
- close to same strength

Metallic Bond =

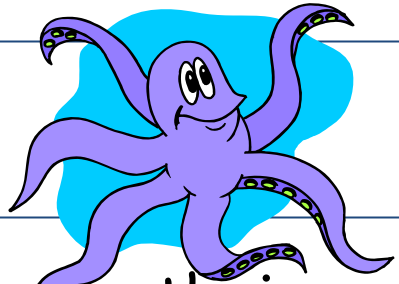
- 2 Metals (both weak)

Summary: Ionic Bonds vs. Covalent Bonds

Ionic	Covalent
Electrons are Transferred (become charged ions that are attracted)	Electrons are shared
Metal + non-metal (ex: Li + K)	2 non-metals (O + O or O + N)
One atom is a lot higher electronegativity than the other (1.7)	Close to equal electronegativities (less than 1.7)

Lewis Dot Structures

Octet Rule



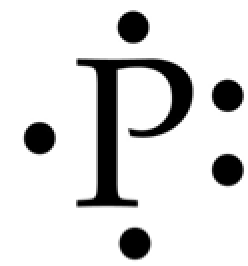
- Most* atom wants to have 8 electrons in their valence shell (outermost shell)

- Chemical compounds form so that each atom can complete their octet by gaining, losing or sharing electrons

- *Exceptions =
  - H & He (they only want 2 electrons in their valence shell)
  - B (forms bonds so it will have only 6 electrons)
  - F, O & Cl (will sometimes be surrounded by more than 8 electrons because they are so electronegative)

Lewis Dot Structure

- Picture showing how many valence electrons an atom has (dots).

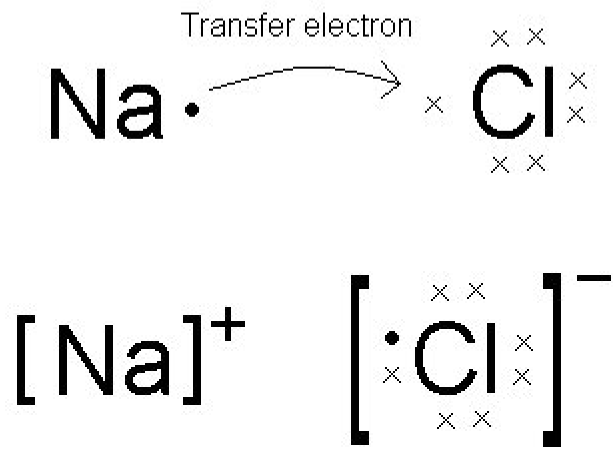


- Helps determine how atoms will bond.

Ex: Phosphorus (has 5 valence electrons)

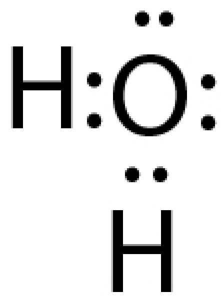
Lewis Dot Structures for Ionic Compounds

- *A* way to show how atoms achieve the octet with each other.
- Note:
  - the transfer of the electron
  - the charges ions that result

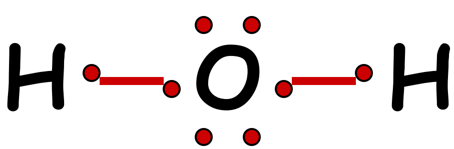


This is how we draw it

Lewis Dot Structures for Covalent Molecules



Or

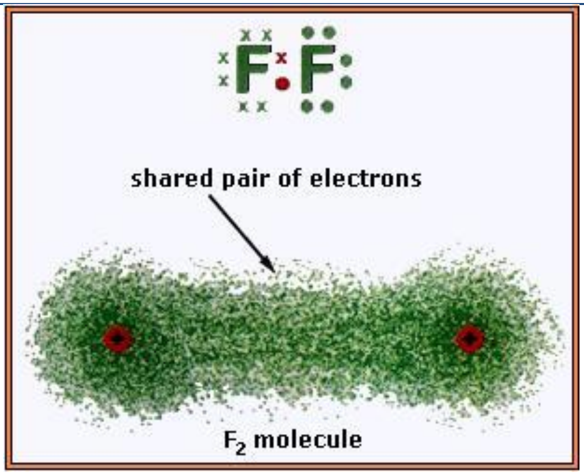


2 ways to show:
- With electrons being shared in between
- Line showing the sharing of pair of electrons

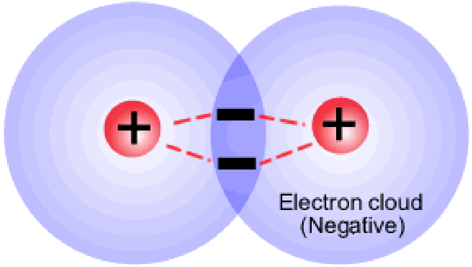
3 Bonds Types in More Depth

Covalent Bonds

- Result from the sharing of electron pairs between two atoms

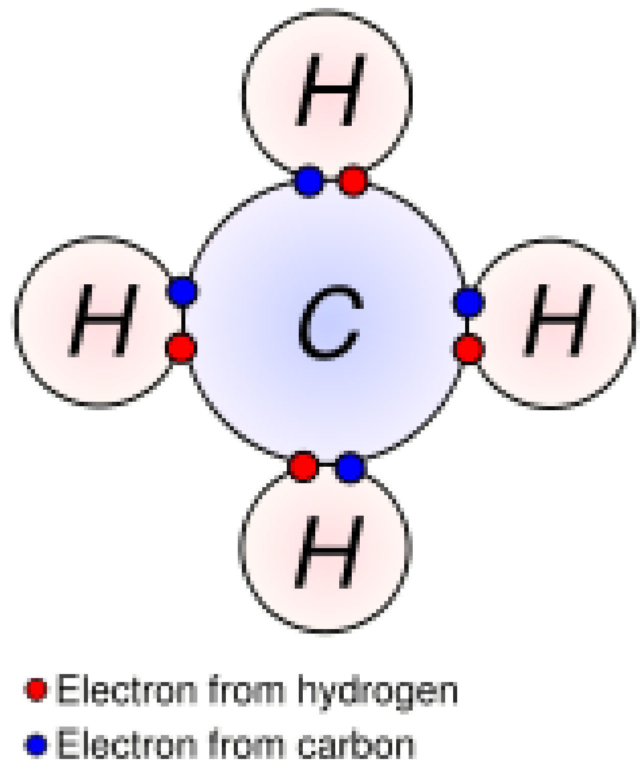


- *Molecule* = termed used to describe atoms are held together by covalent bonds



Covalent Bonds

- Occurs between 2 non-metals
- electrons are shared
- 2 types of covalent bonds: Polar and non-polar (to be discussed later)
- Ex: Water & most biological molecules (sugars, fats, proteins)
- Can form single, double, or triple bonds

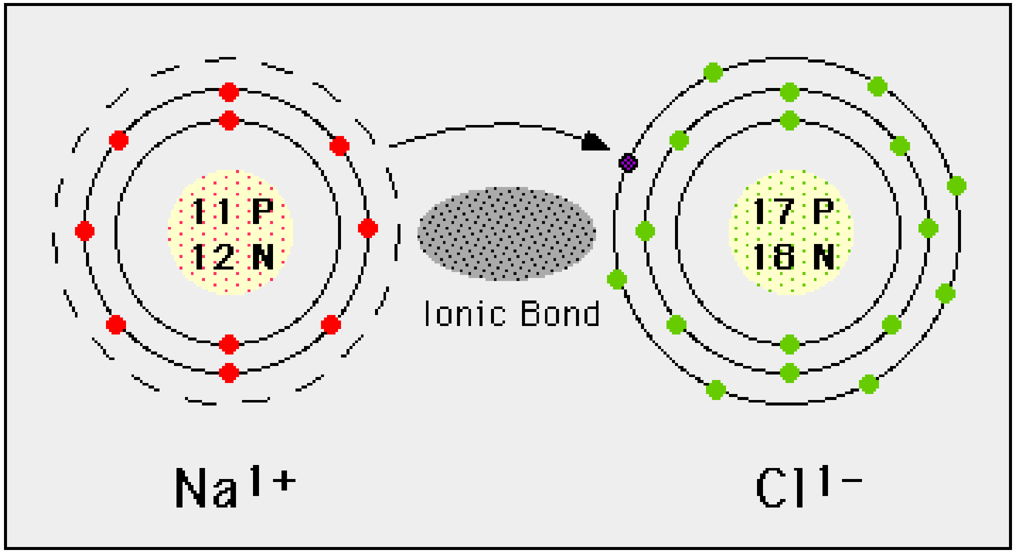


Ionic Bonds

- Forms between:

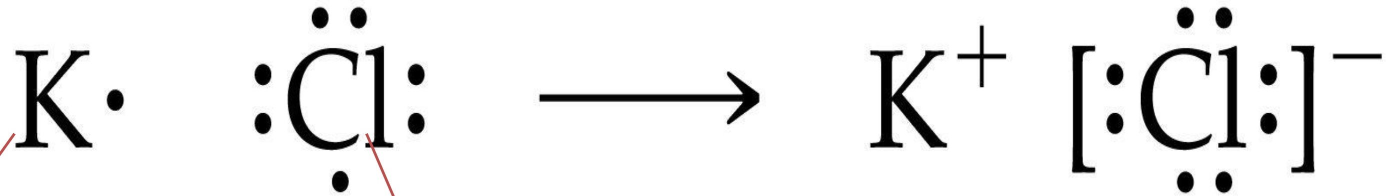
Metal + Non-Metal

- Electrons are transferred



Ionic Bonds (cont.)

Ex:



Electroneg= .8

Electroneg= 3.0

- **Cl is so much stronger that it will "take" K's electron**

- The transfer of electron causes K to be a **cation (+)** and Cl to be an **anion (-)**.

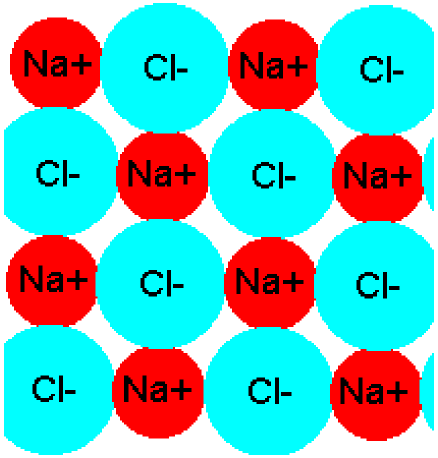
- Oppositely charged particles are highly attracted to each other… Ionic bond!

Characteristics of Ionic Compounds

- Shape- crystal lattice of alternating positive and negative ions

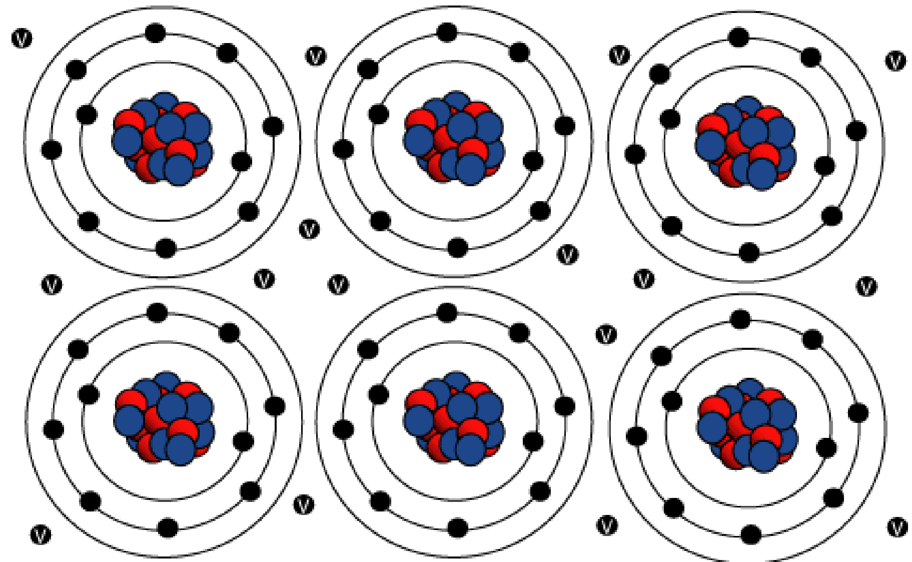
- Ex: NaCl and salts

- Ionic bonds are strong so they are:
  - hard
  - have a high melting point
  - high boiling point



Crystal Lattice

Metallic Bonds- "sea of electrons"



- Forms between 2 metals
- Metal atoms valence electrons overlap creating a "sea of electrons".

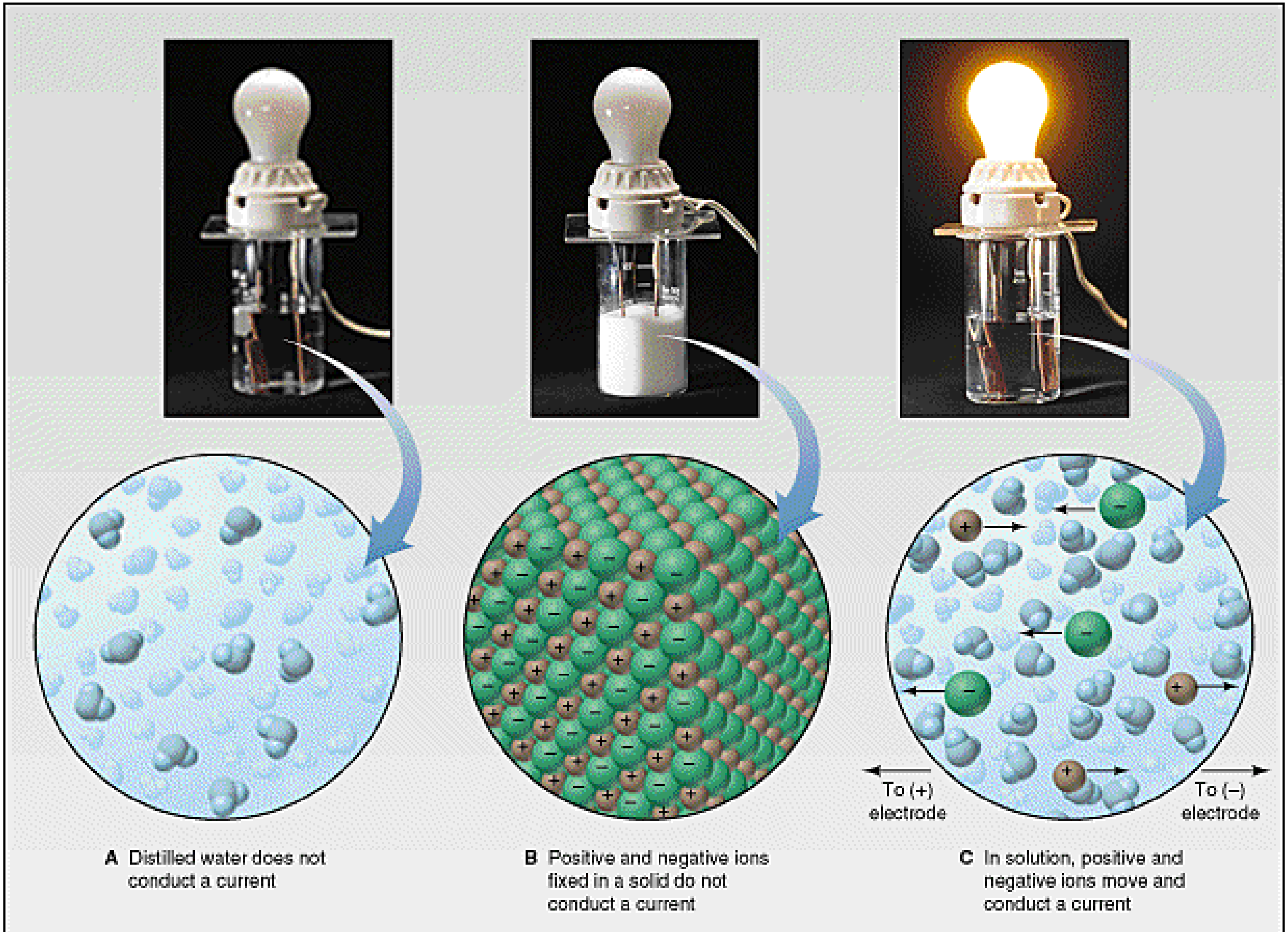
•Electrons do not belong to any one atom, but roam freely throughout the metal atoms

•**Ex: Brass (alloy of Cu + Zn)**

Metallic Characteristics

- Because of these roaming "sea" of electrons:
    - metals are great conductors of heat/ electricity
    - they are ductile (can be made into wire)
    - they are malleable (can be hammered into sheets)

Electrical Conductivity



Properties & Bonding Type

pgs.161-182

Comparison	Covalent	Ionic Bonds	Metallic Bonds
Formation			
Types of Atoms	Non-metal & non-metal	Non-metal & metal	Metal & metal
Electron Distribution	Shared	Transferred	Sea of electrons
Characteristics			
Bond Strength	Strong	Very strong	Varies
Structure	Neutral group	Crystal lattice	crystalline
Properties of Compounds			
Type of Compound	Molecular	Ionic	metallic
Melting Point	Low	Very high	n/a
Boiling Point	Low	High	Very high
Malleability	n/a	Not malleable, brittle	Very malleable
Ductility	n/a	Not ductile	Very ductile
Conductivity	Not conductive	Conductive	Highly conductive

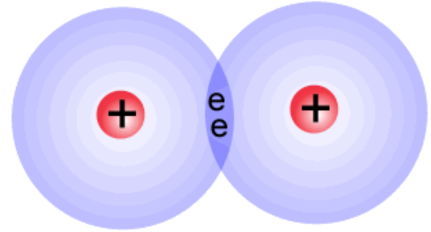
Bond Energy & Bond Length

Bond Energy-

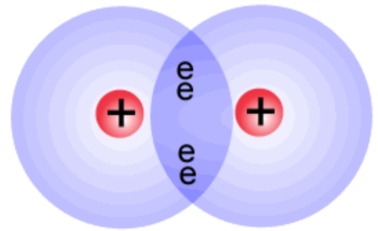
energy required to break bond

	Bond Length	Bond Energy
Single Bond	—	Low
Double Bond	=	
Triple Bond	≡	High

Only one pair of electrons holding the nuclei together



Two pair of electrons hold the nuclei tighter and closer



Bond Energy & Bond Lengths

Bond	Length (picometers)	Energy (kJ/mol)
H—Br	141	366
H—C	109	413
H—N	101	391
H—O	96	464
H—S	93	339
C—O	143	360
C=O	129	799
C—C	154	348
C=C	134	614
C C	120	839
O—O	148	145
O=O	121	498
N—N	145	170
N=N	125	418
N N	110	945

Lewis Structures in Covalently Bonded Molecules & HONC Rule

pgs. 183 - 186

Drawing Lewis Dot Structures for Molecules

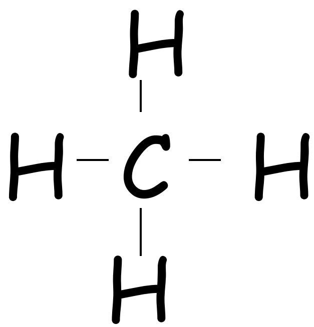
- arrange atoms to form a skeleton
  - -Carbon is center atom
  - -Hydrogen is never a central atom

- Pair up all electrons
  - unpaired electrons can pair unpaired electron from another atom to form a bond

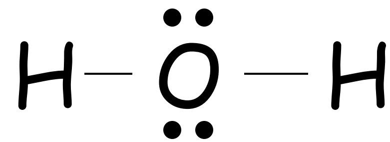
- Make sure each atom of the molecule obeys the octet rule & HONC rule
- Make sure you have correct # of valence electrons

Examples of Lewis Dot Structure

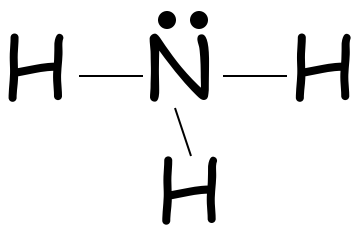
CH_4



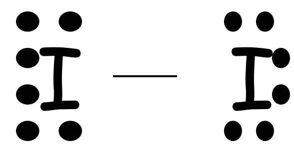
H_2O



NH_3



I_2



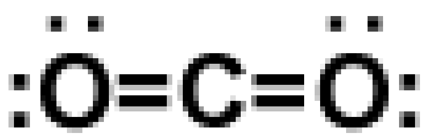
Multiple Covalent Bonds:
<u>Double bonds</u>

Two pairs of shared electrons

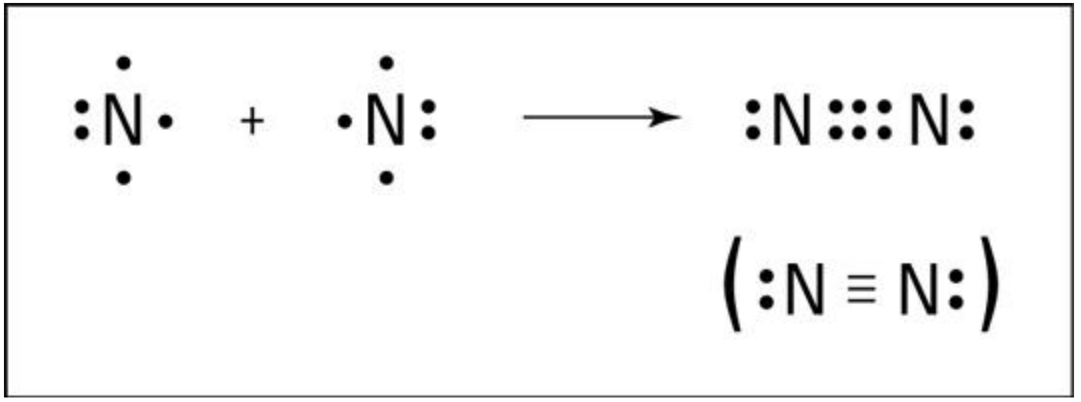
O_2 :

each oxygen
has 8 electrons
in the valence shell

CO_2 :



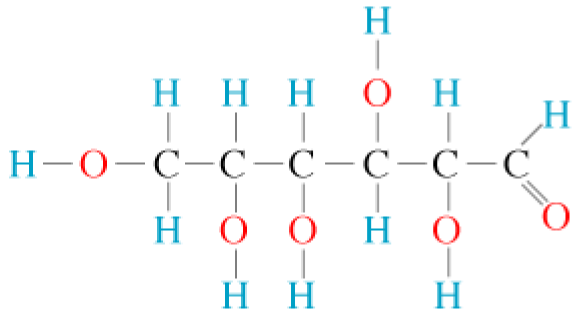
Multiple Covalent Bonds:
Triple bonds



Three pairs of shared electrons

Molecular vs. Structural Formulas

- Molecular formulas – show how many atoms of each element are in the molecules
  - Ex: $C_6H_{12}O_6$ = 6 carbons, 12 hydrogens & 6 oxygens
- Structural formulas – show the 2-dimensional shape of the molecule
  - Ex:



HONC 1-2-3-4 Rule

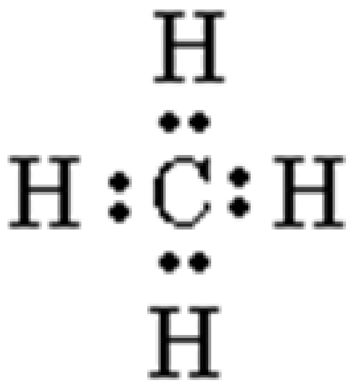
- Hydrogen, oxygen, nitrogen & carbon are common elements found in biological molecules.
  - Hydrogen needs 1 electron to fill its “octet”
  - Oxygen needs 2 electrons to fill its octet
  - Nitrogen needs 3 electrons to fill its octet
  - Carbon needs 4 electrons to fill its octet
- “1-2-3-4” can be used to predict how these atoms will form bonds with other atoms to build molecules.

Molecular Geometry
Seeing Molecules in 3-D

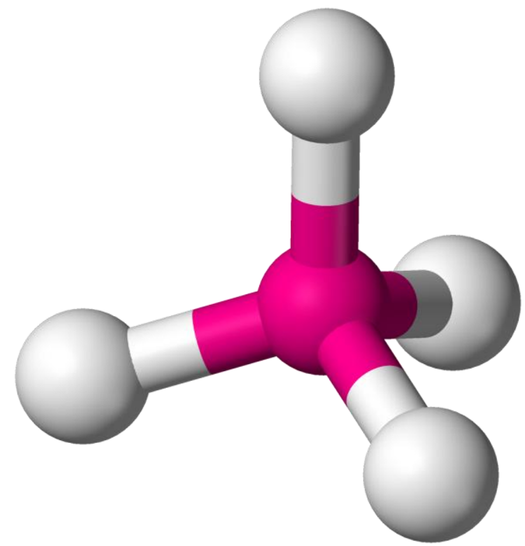
Molecular Geometry

molecules are really 3-D!

CH_4 in 2-D on a sheet of paper



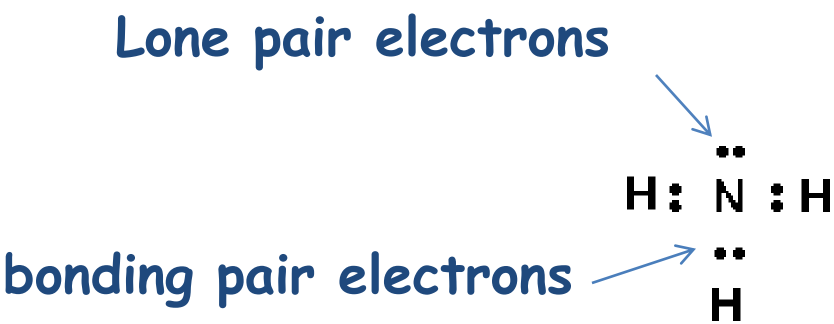
CH_4 looks like this in 3-D



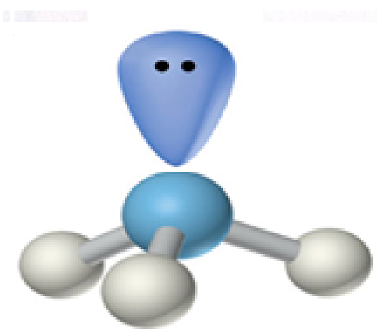
Valence Electrons determine Molecular "VSEPR" Shape

- VSEPR = "Valence-Shell Electron Pair Repulsion"

- Electron pairs (bonding or lone pairs) in a molecule repel each other and will try and get as far away from each other as possible… this determines the shape.



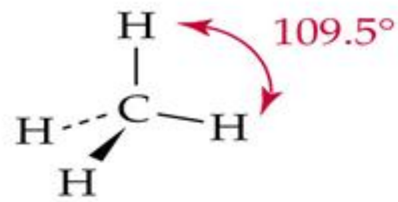
NH_3 in 2-D

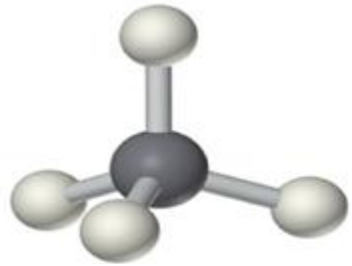


NH_3 VSEPR shape in 3-D

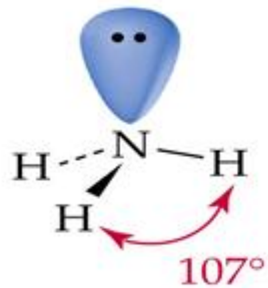
4 Shapes to Know

Tetrahedral

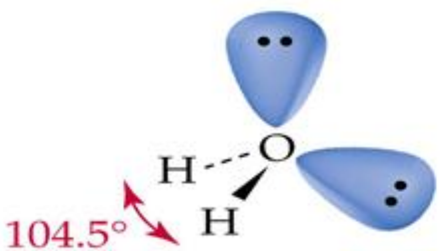




Pyramidal

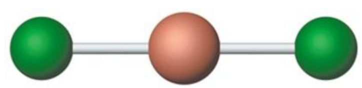


Bent

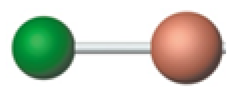


Linear

$O=C=O$

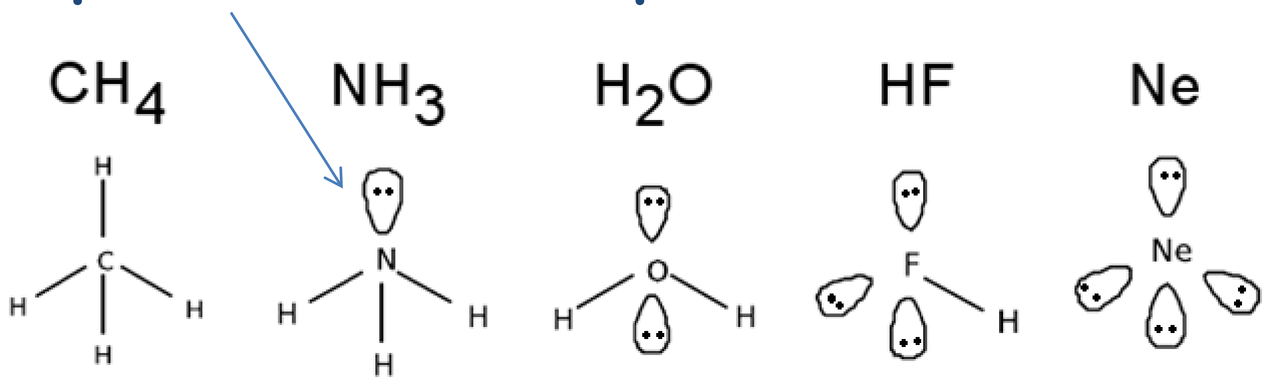


H — Br

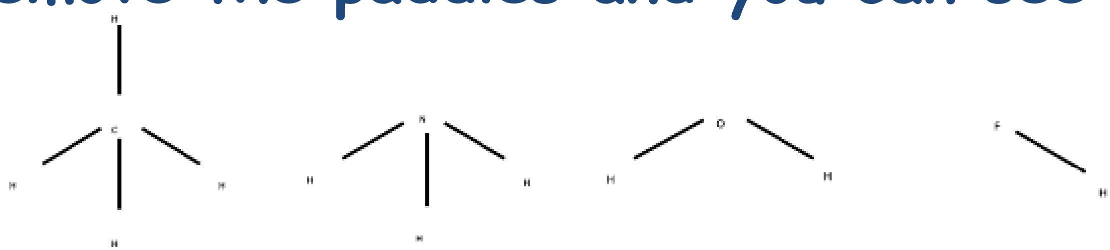


How Lone Pairs Affect Molecular Shape

"paddles" are lone pairs of electrons.



Remove the paddles and you can see the shapes.



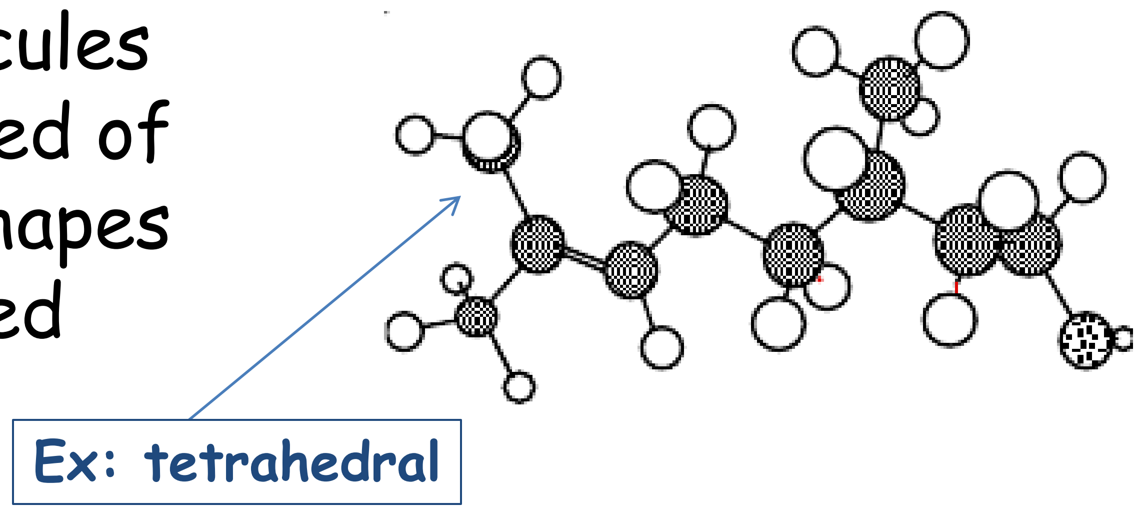
tetrahedral pyramidal bent linear

Steps for Determining Molecular Geometry

1. Draw Lewis dot structure
2. Count number atoms bonded to the central atom
3. Count number of lone-pair electrons on the central atom
4. Look up the Geometry on the chart

Shapes in Large Molecules

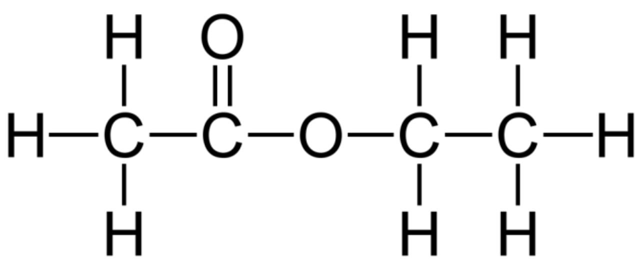
Large molecules are composed of the small shapes we've studied



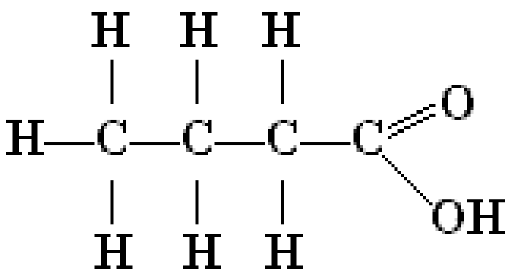
ball-and-stick model of citronellol

Why Shape Matters

Ethyl Acetate ($C_4H_8O_2$)



Butyric Acid ($C_4H_8O_2$)



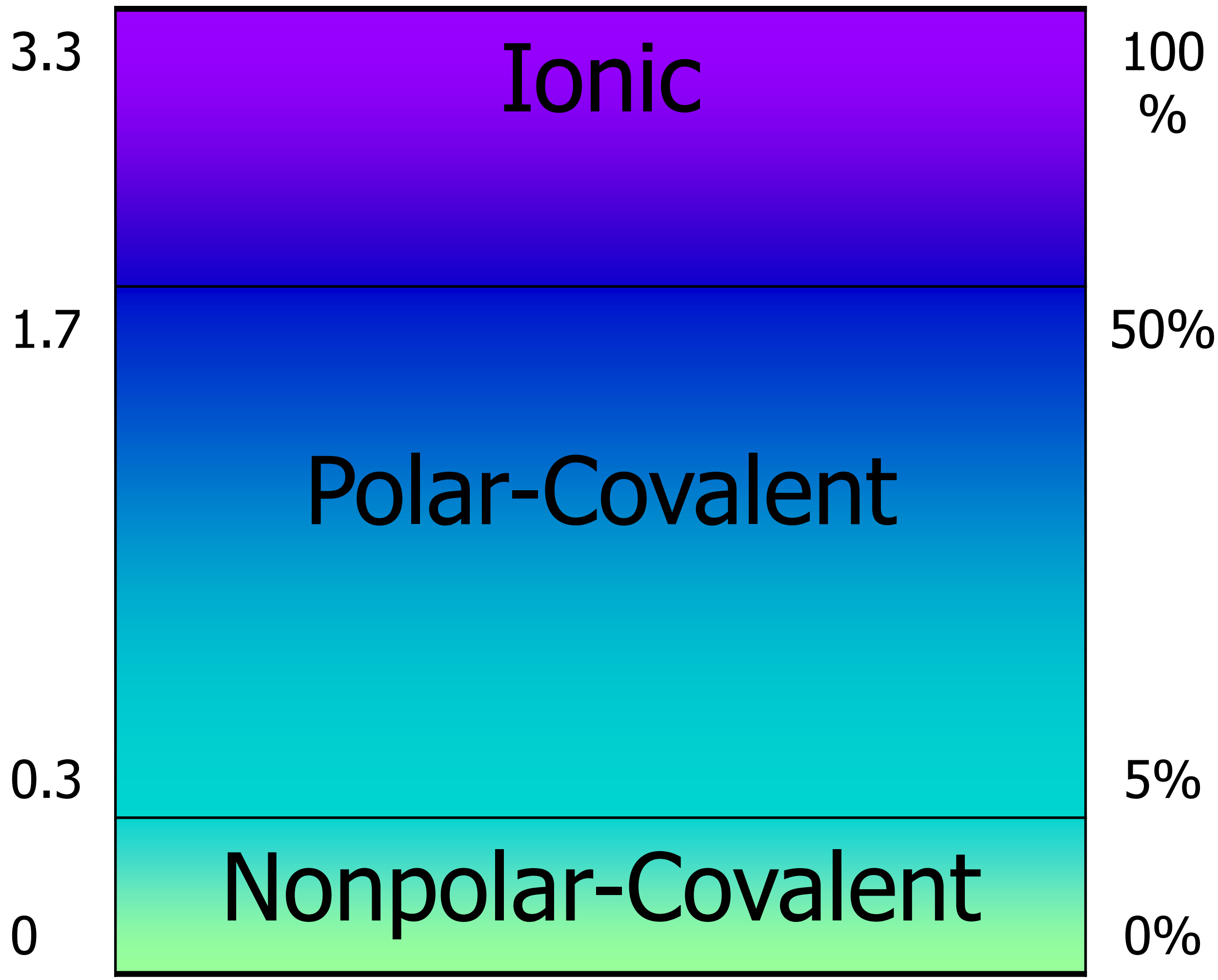
Same formula, but different shapes
= very different smells

Rum extract smell

Rancid butter smell

Polarity

Differences In Electronegativities

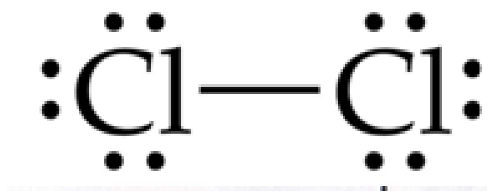


Practice Problems

Bonding Between:	Difference in Electronegativity	Bond Type
Cl & Ca	**3.0 – 1.0 = 2.0**	**Ionic**
O & H	**3.5 – 2.1 = 1.4**	**Polar-covalent**
B & H	**|2.0 – 2.1| = 0.1**	**Nonpolar-covalent**

2 types of Covalent Bonds:

Non- Polar

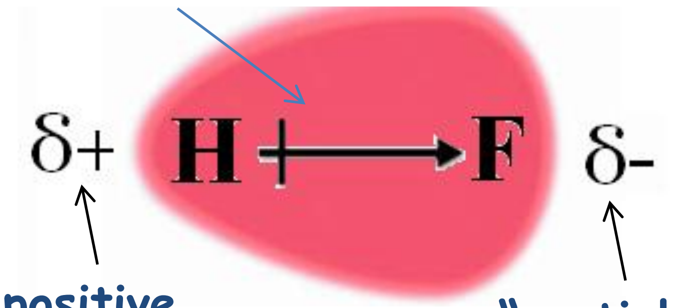


-Electrons are shared equally

- Usually the same element bonded to itself

Polar

(Arrow shows F is "pulling" electrons)

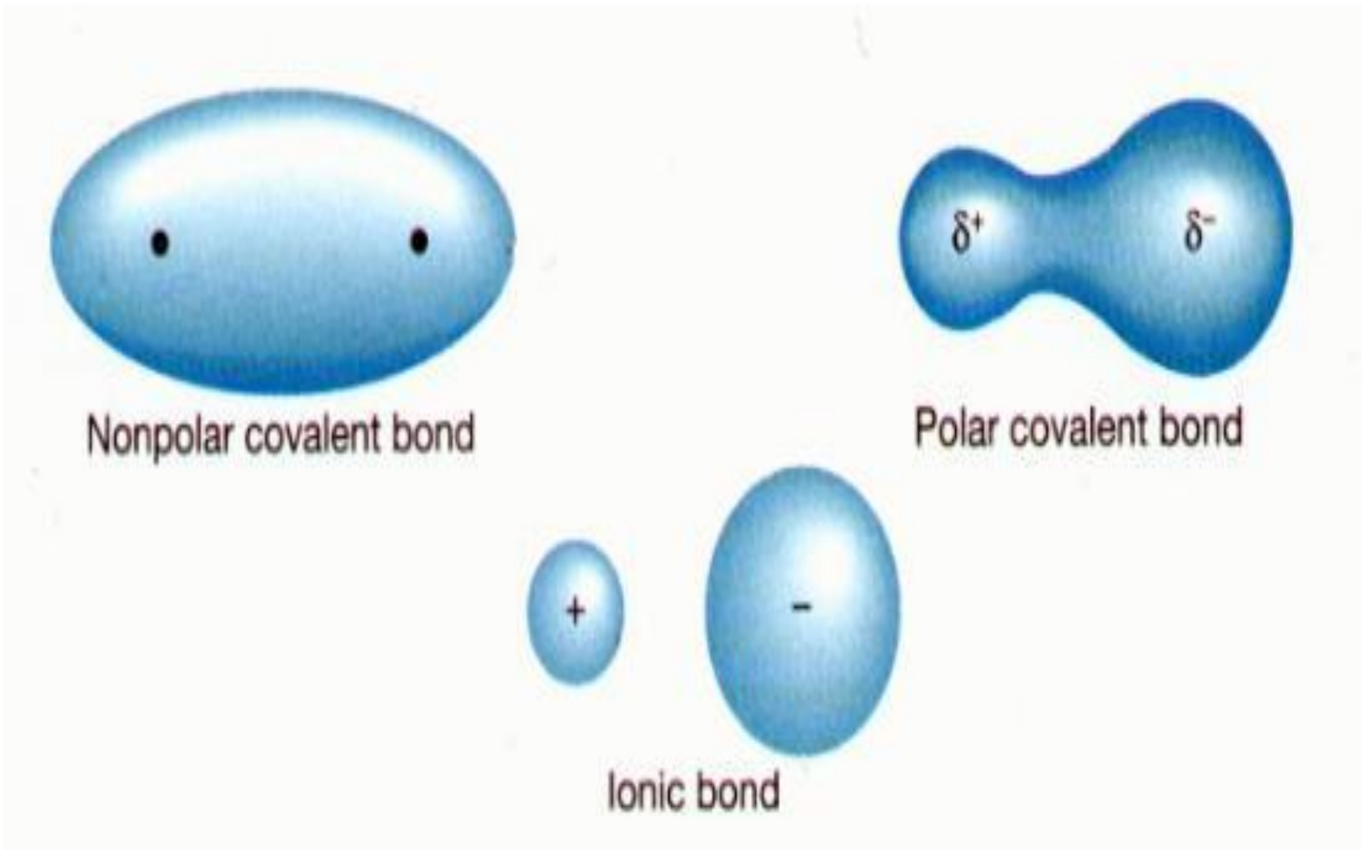


"partial positive charge"

"partial negative charge"

- Unequal sharing of electrons between atoms
- more electronegative atoms "hogs" electrons

Visual Comparison of Bond Types

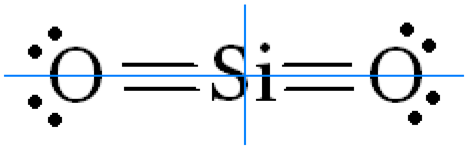


Determining Polarity

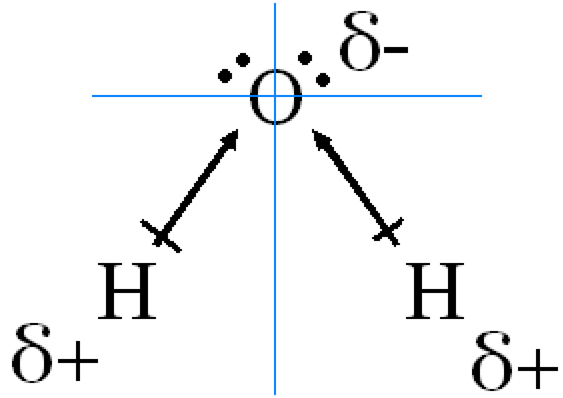
1. Draw correct VSEPR Shape
2. Determine if molecule is symmetrical.
3. If the molecule is symmetrical = **non-polar**

   - no partial charges are needed!
4. If the molecule is NOT symmetrical = **polar**

   - you must show partial charges.
   - always bent or pyramidal shapes



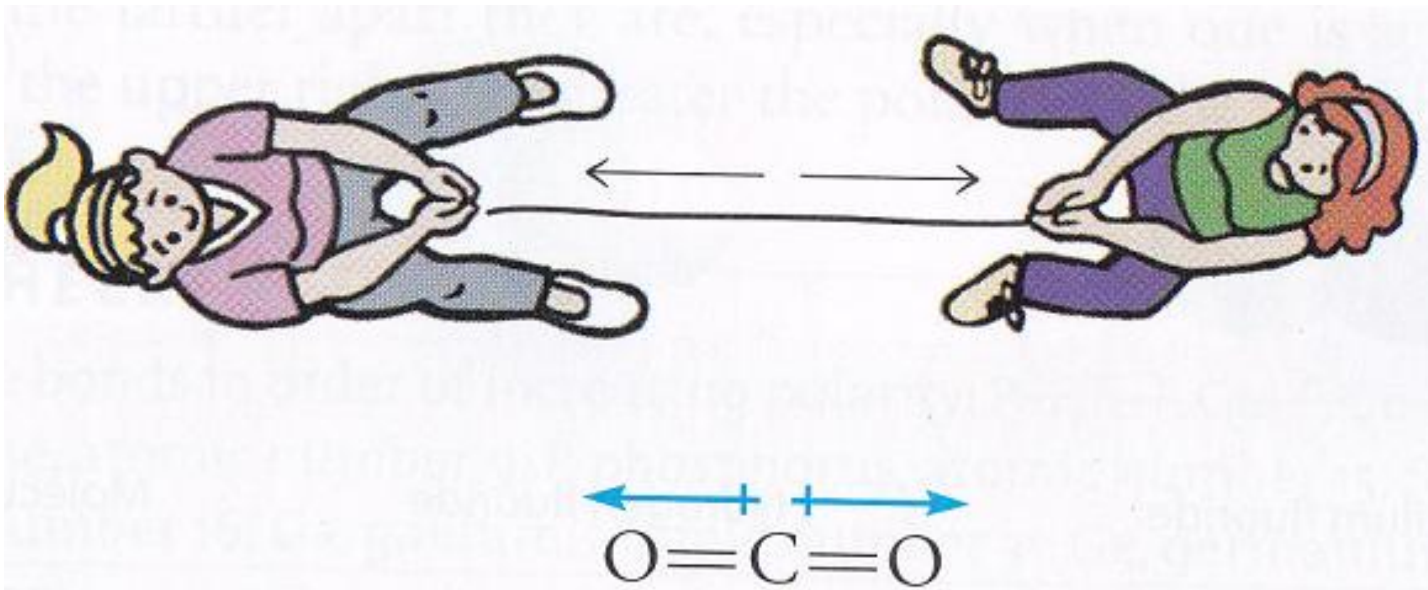
Linear shape (symetrical)



Bent shape (assymetrical)

Ex: CO_2

- Carbon dioxide= nonpolar
- has polar bonds, but they cancel each other out.



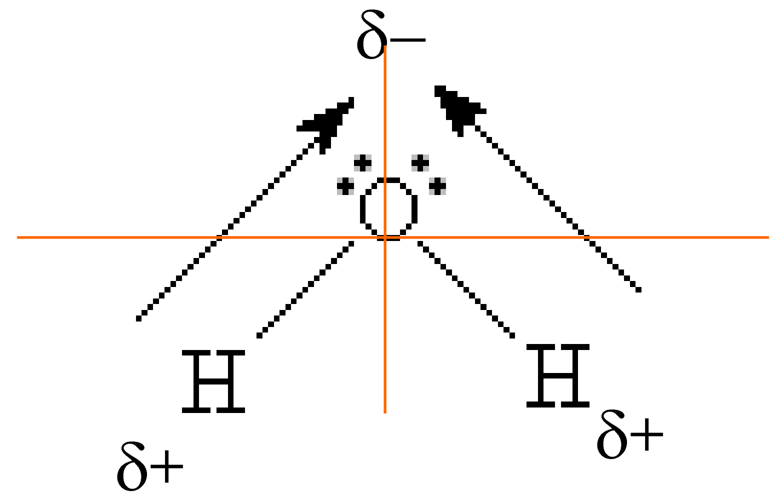
EX: Water= Polar Molecule

How we know:

1) Cut the molecule on 2 planes

   - see how it's different above the horizontal line = non symmetrical

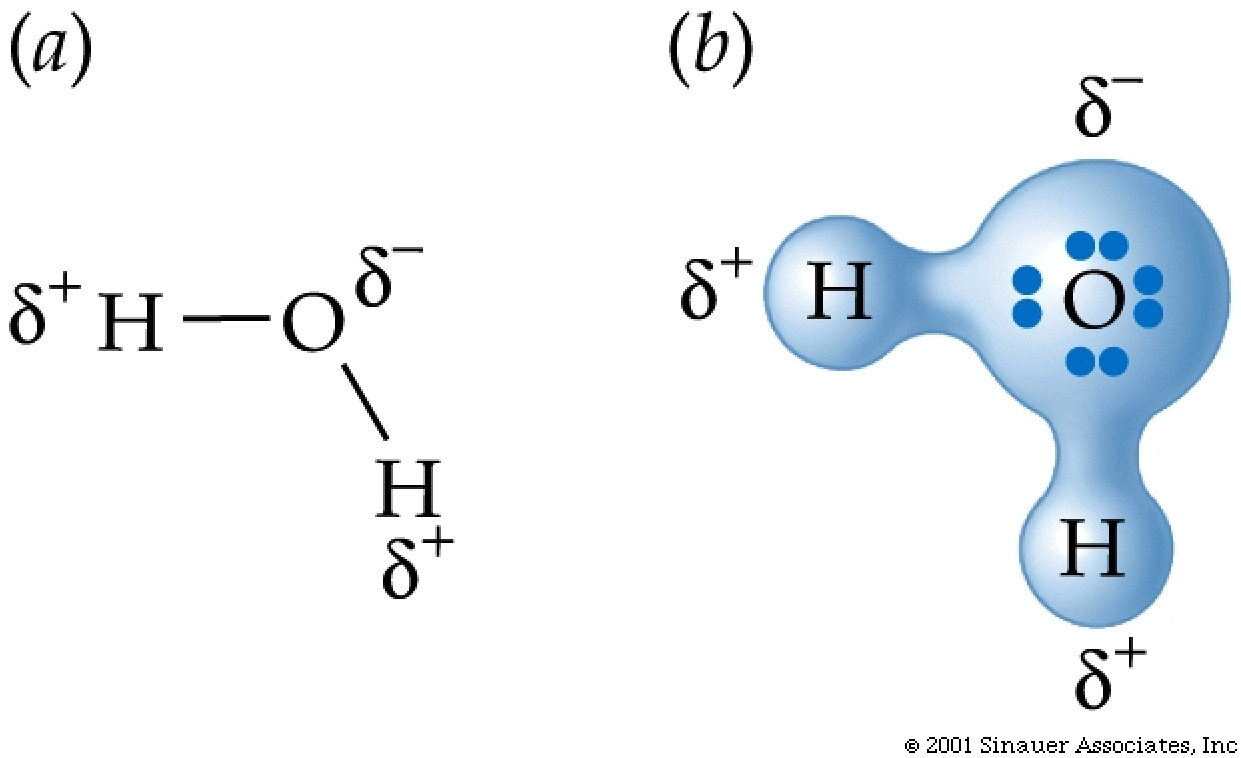
1) One atom is "pulling", look at periodic table to determine which one.



→ Indicates which atom "pulls" the electrons

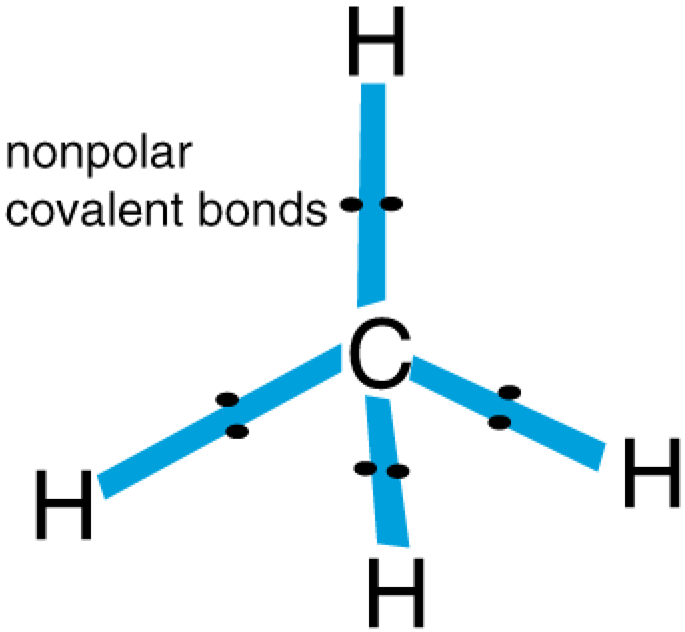
δ− Means oxygen is slightly negative because it "hogs" electrons

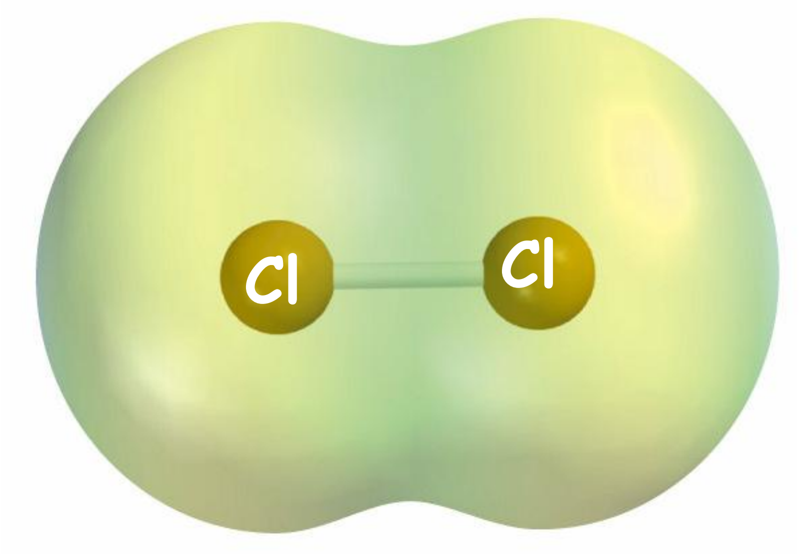
2 views of Polar water



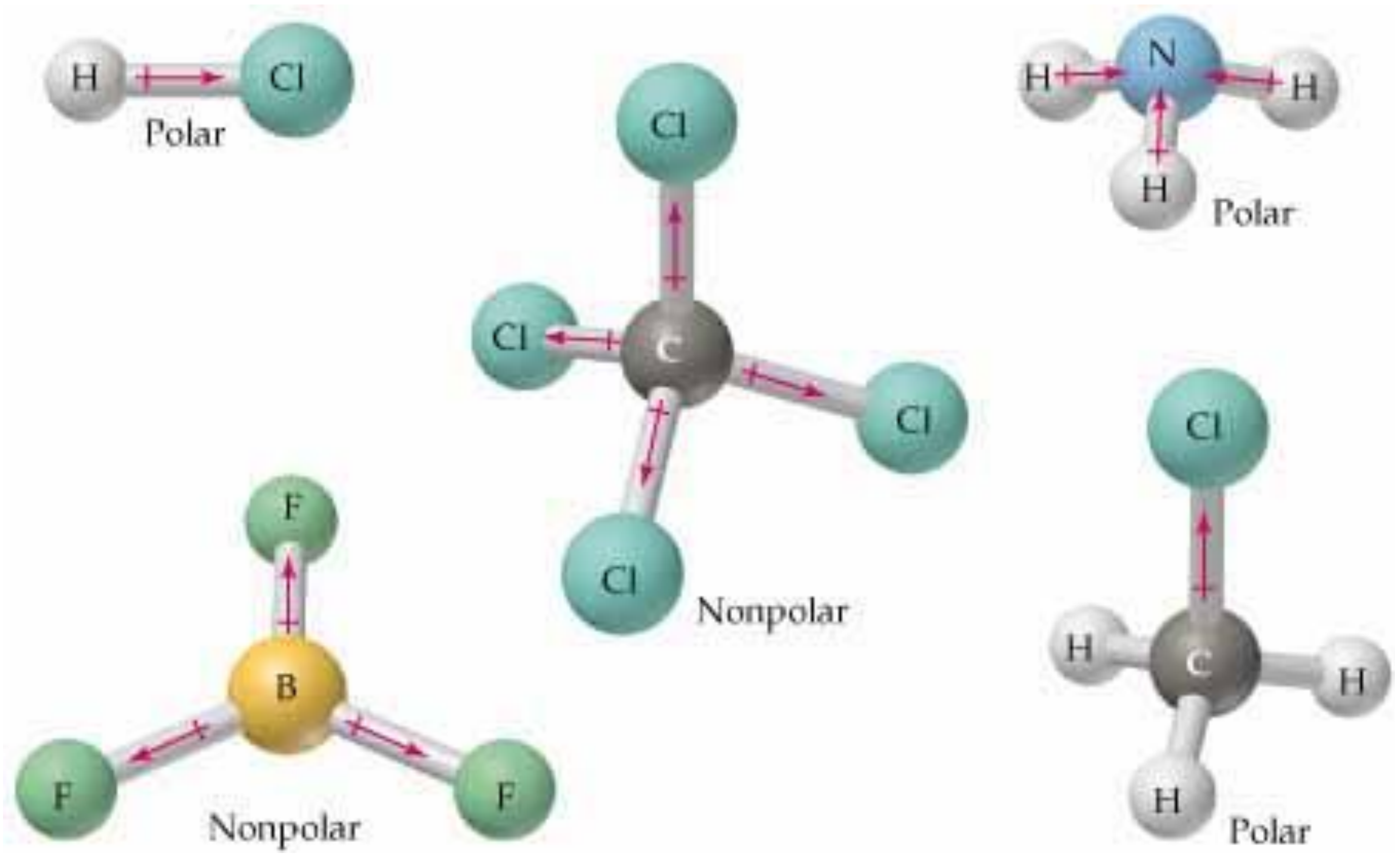
Non Polar Molecules

- Non Polar molecule= "no pull"
  - equal sharing of electrons
  - No difference in electronegativity
  - symmetrical in shape





Examples of Polar & Nonpolar Molecules



Electronegativity

Colour key	Range
blue	0.5 - 0.9
light cyan	1.0 - 1.4
light green	1.5 - 1.9
light yellow	2.0 - 2.4
orange	2.5 - 2.9
dotted orange	3.0 - 3.4
light red	3.5 - 3.9
red	4.0 +

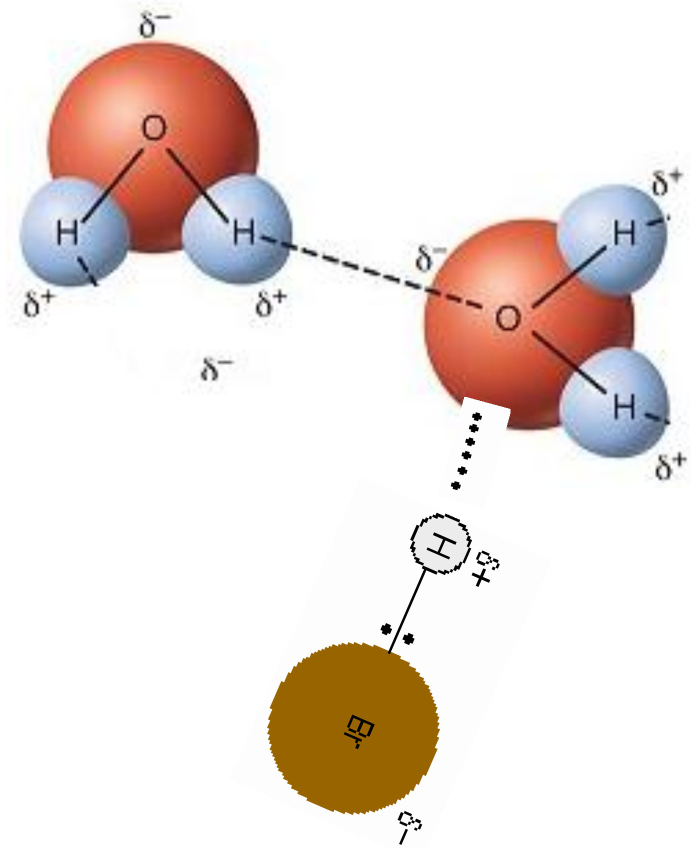
1	2	(3)	(4)	(5)	(6)	(7)	(8)	(9)	(10)	(11)	(12)	3 (13)	4 (14)	5 (15)	6 (16)	7 (17)	8 (18)
H 2.1																	He --
Li 1.0	Be 1.6											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne --
Na 0.9	Mg 1.3											Al 1.6	Si 1.9	P 2.2	S 2.5	Cl 3.0	Ar --
K 0.8	Ca 1.3	Sc 1.4	Ti 1.5	V 1.6	Cr 1.7	Mn 1.6	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.7	Ga 1.6	Ge 2.0	As 2.2	Se 2.6	Br 2.8	Kr --
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.3	Nb 1.6	Mo 2.2	Tc 2.1	Ru 2.2	Rh 2.3	Pd 2.2	Ag 1.9	Cd 1.7	In 1.8	Sn 2.0	Sb 2.1	Te 2.1	I 2.7	Xe 2.6
Cs 0.8	Ba 0.9	La 1.1	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 2.0	Pb 2.3	Bi 2.0	Po 2.0	At 2.2	Rn --
Fr 0.7	Ra 0.9	Ac 1.1	Rf --	Db --	Sg --	Bh --	Hs --	Mt --	Uun --	Uuu --	Uub --		Uuq				

Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

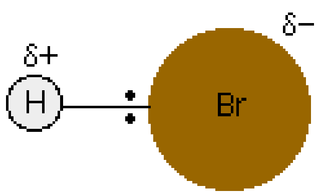
Inter vs. Intra molecular Forces

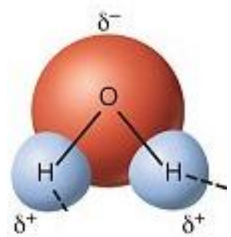
Why Polarity Matters: Molecular Attractions

- 1 molecule can be attracted to another molecule
  - "inter"molecular force



- You can predict how one molecule might react with another: Ex: HBr + H_2O

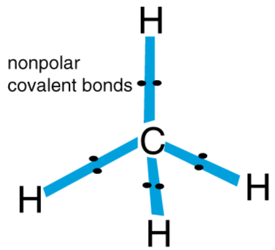




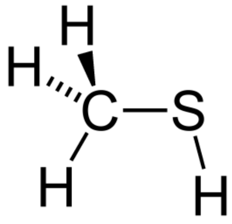
Intermolecular Attractions & Smell

- Besides shape, polarity also plays a role in your ability to smell.
  - Polar molecules = smell
  - Non-polar = don't smell
- Your smell receptors are polar and surrounded by mucous (a watery substance)

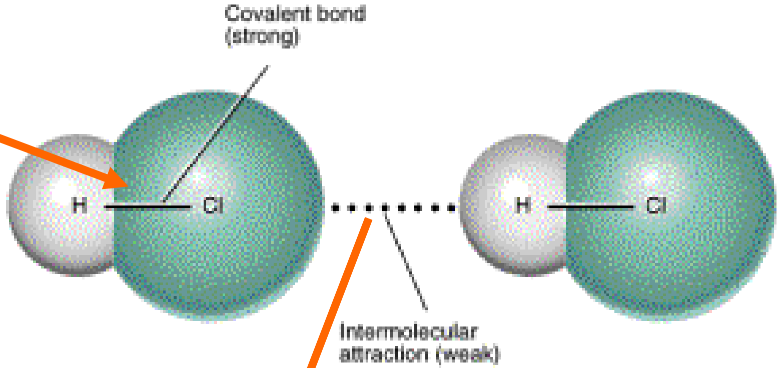
Ex: Methane gas is odorless



-They add a this stinky chemical to it so that you can smell it it:



Intermolecular Forces vs. Intramolecular Forces



Intramolecular Forces:
(within in a molecule)

Ex: -Covalent bond

- Ionic bond

-Metallic bond

Intermolecular Forces:
(between molecules)

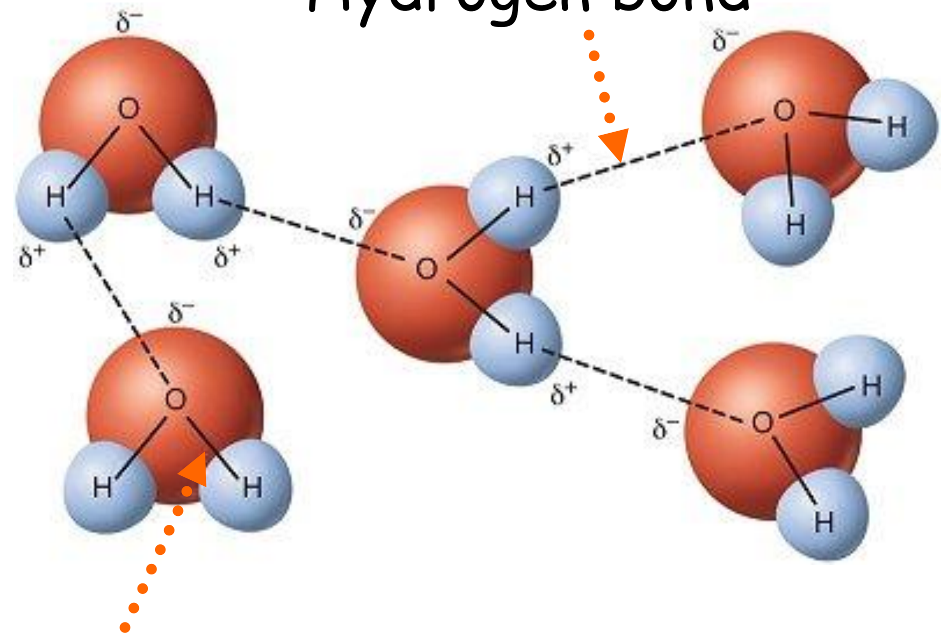
Ex: Hydrogen bonds

-Weaker than covalent, ionic, & metallic bonds

Hydrogen Bonding- (an intermolecular force) in Water

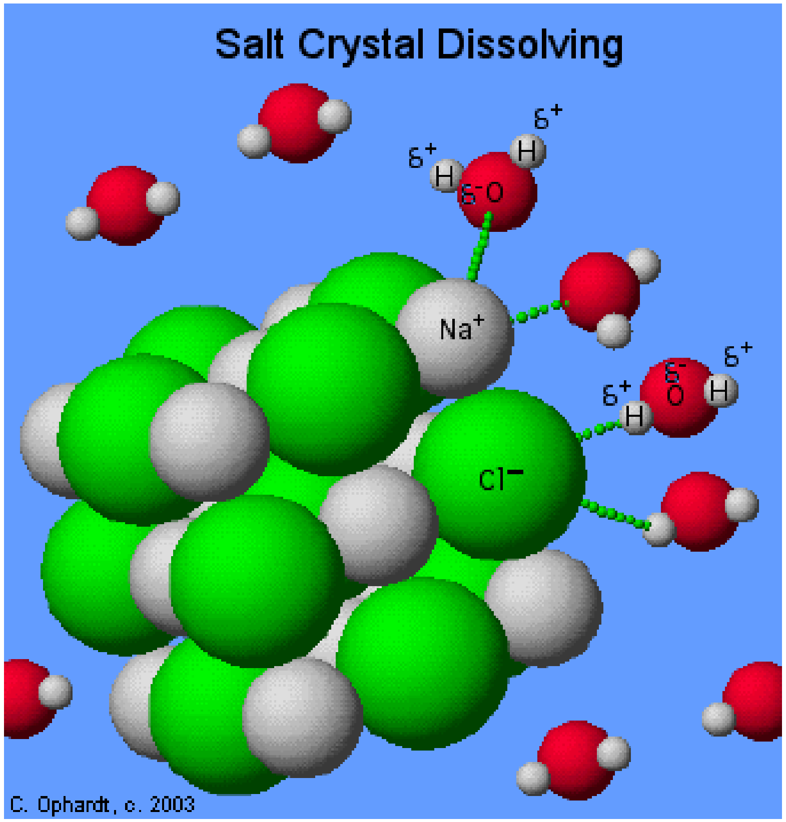
- Water is polar
  - (has a + and – end)
  - It's "sticky"
- Will stick to any other thing that is:
  - polar (ex: other water molecules)
  - charged ionic substances (NaCl)

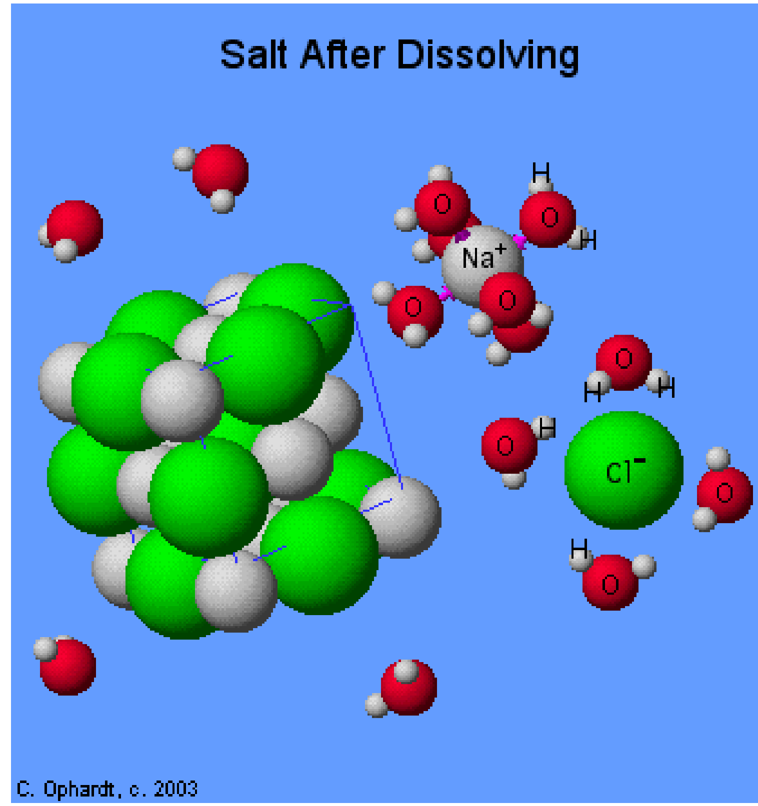
Hydrogen bond



Covalent bond

Water's Polarity leads to its ability to dissolve things so well





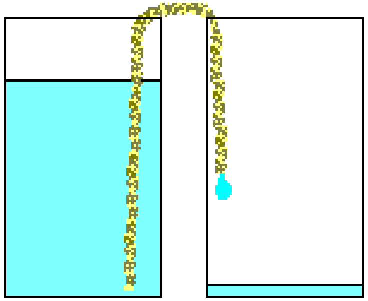
The slight charges on water attract the NaCl's ions and cause them to separate from each other

Unique Properties of Water due to polarity & hydrogen bonding

1) Surface tension (hydrogen bonds create surface on water)

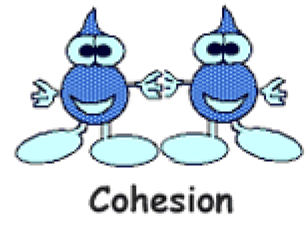
3) Adhesion/ Cohesion (water is attracted to other water molecules)

4) Capillary action water is attracted to other water molecules and will "rise"

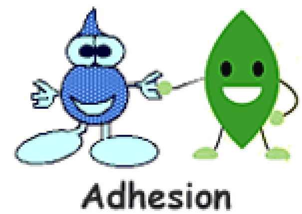


Properties of Water due to Hydrogen Bonding & Polarity

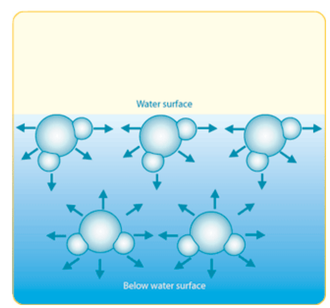
- **Cohesion** – water molecules are attracted to one another
  - Causes water to be "Sticky"
  - This is why water forms droplets

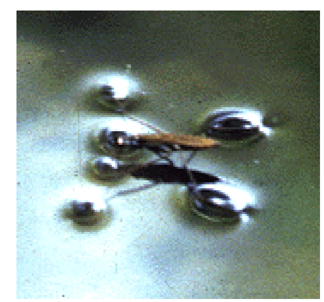


- **Adhesion** – water is attracted to other substances
  - Water will "stick" to containers & objects



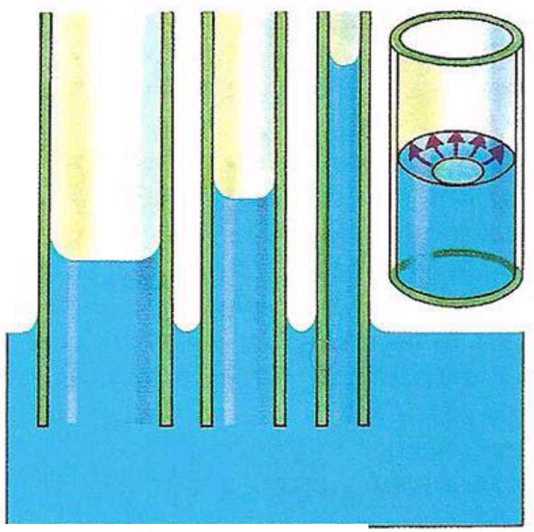
- **Surface tension** – strong forces between molecules cause the surface of a liquid to contract

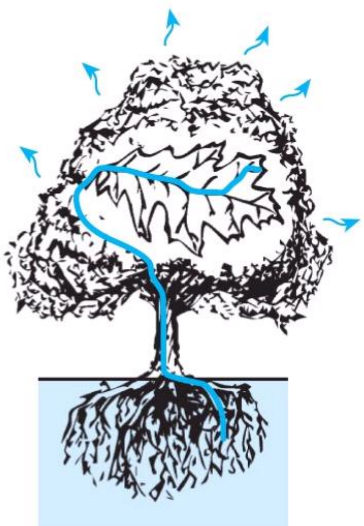




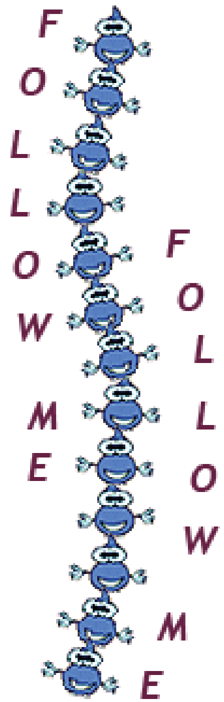
"USGS Water Science for Schools: All about water!" US Geological Survey. 9 December 2011. http://ga.water.usgs.gov/edu/index.html

More properties…

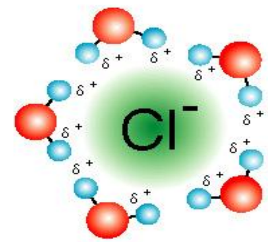




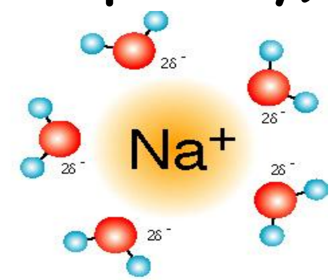
- **<u>Capillary Action</u>** – the movement of water within the spaces of a porous material due to the forces of adhesion, cohesion, and surface tension.



- **<u>Universal Solvent</u> things dissolve in water- polarity)**



Slightly positive hydrogen are attracted to chlorine anions

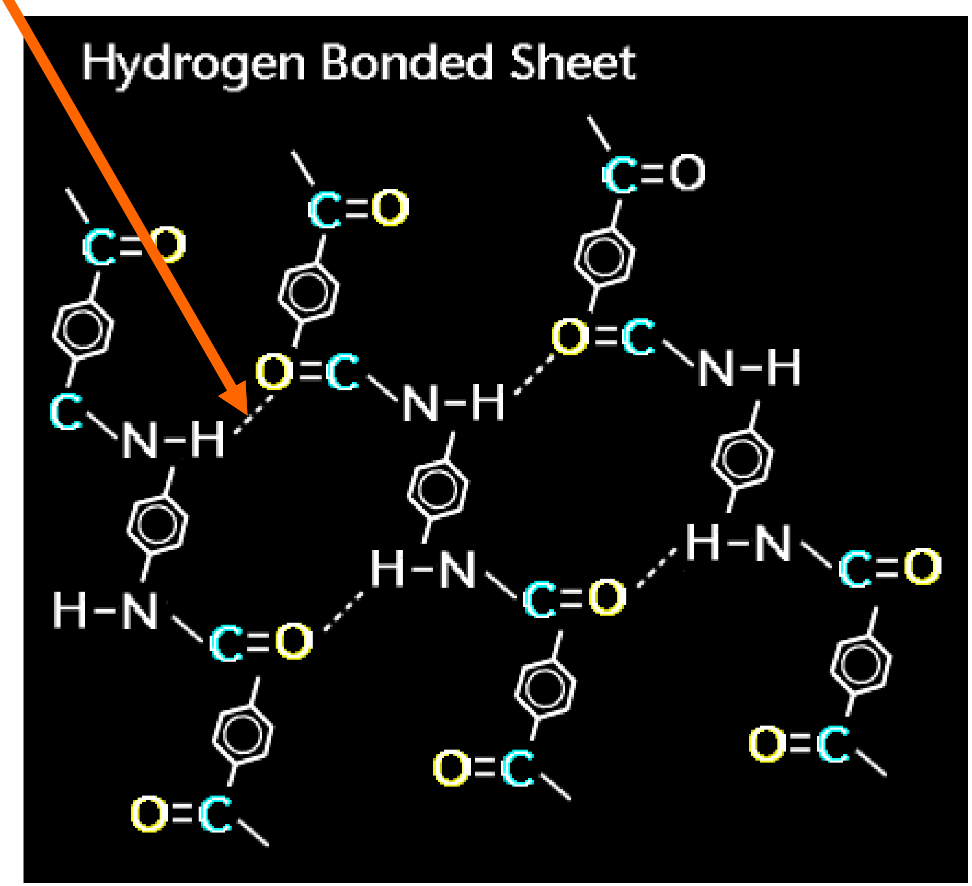


Slightly negative oxygen are attracted to sodium cations

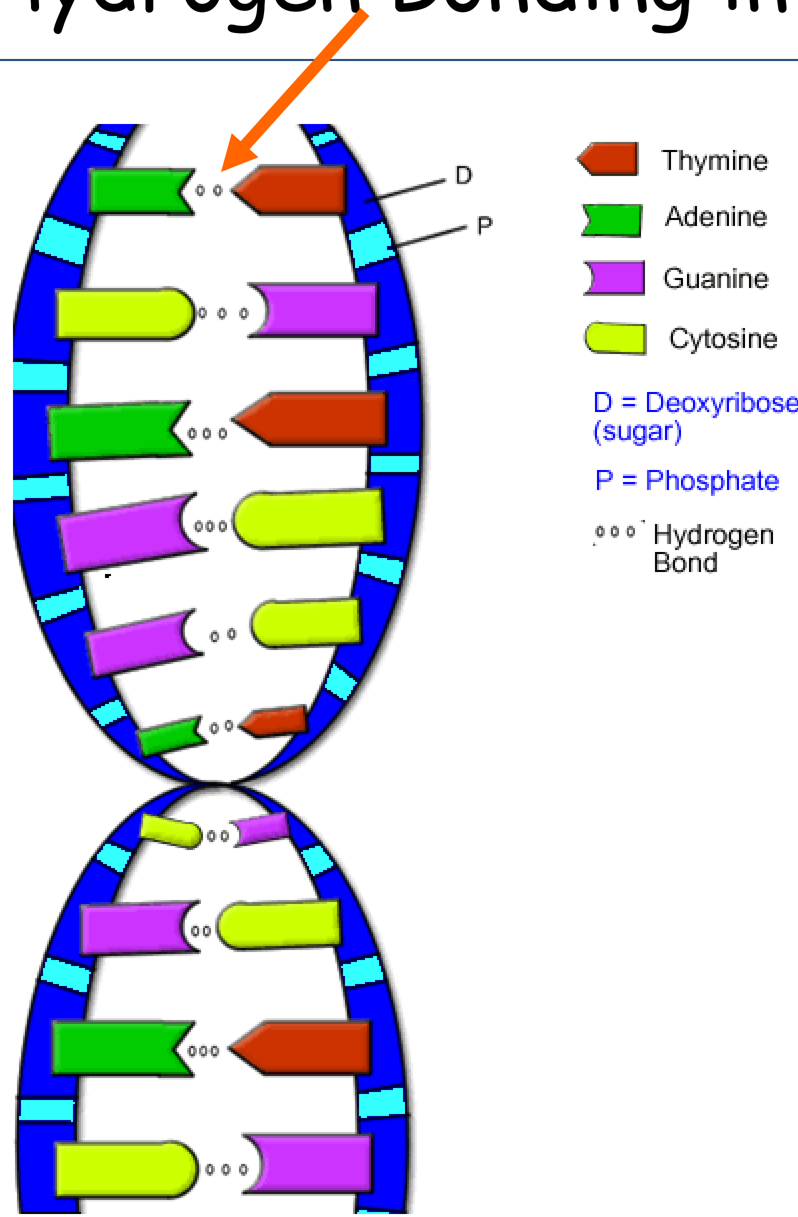
"USGS Water Science for Schools: All about water!" US Geological Survey. 9 December 2011. http://ga.water.usgs.gov/edu/index.html

Hydrogen Bonding in Kevlar

Hydrogen bonding in Kevlar, a strong polymer used in bullet-proof vests.

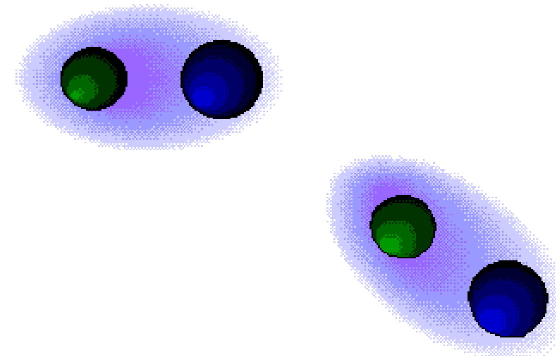


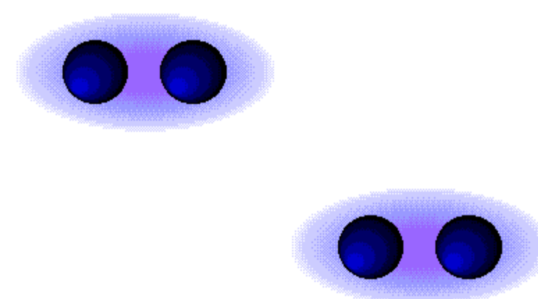
Hydrogen Bonding in DNA



Other Intermolecular Forces
(FYI… not part of this class)

- Van der Waals Forces include:
  - **Dipole-Dipole forces** – results from the tendency of polar molecules to align themselves so that the positive end of one molecule is near the negative end of another molecule.
  - **London (Dispersion) forces** –results from the small, instantaneous dipoles that occur because of the varying positions of the electrons during their motion about nuclei

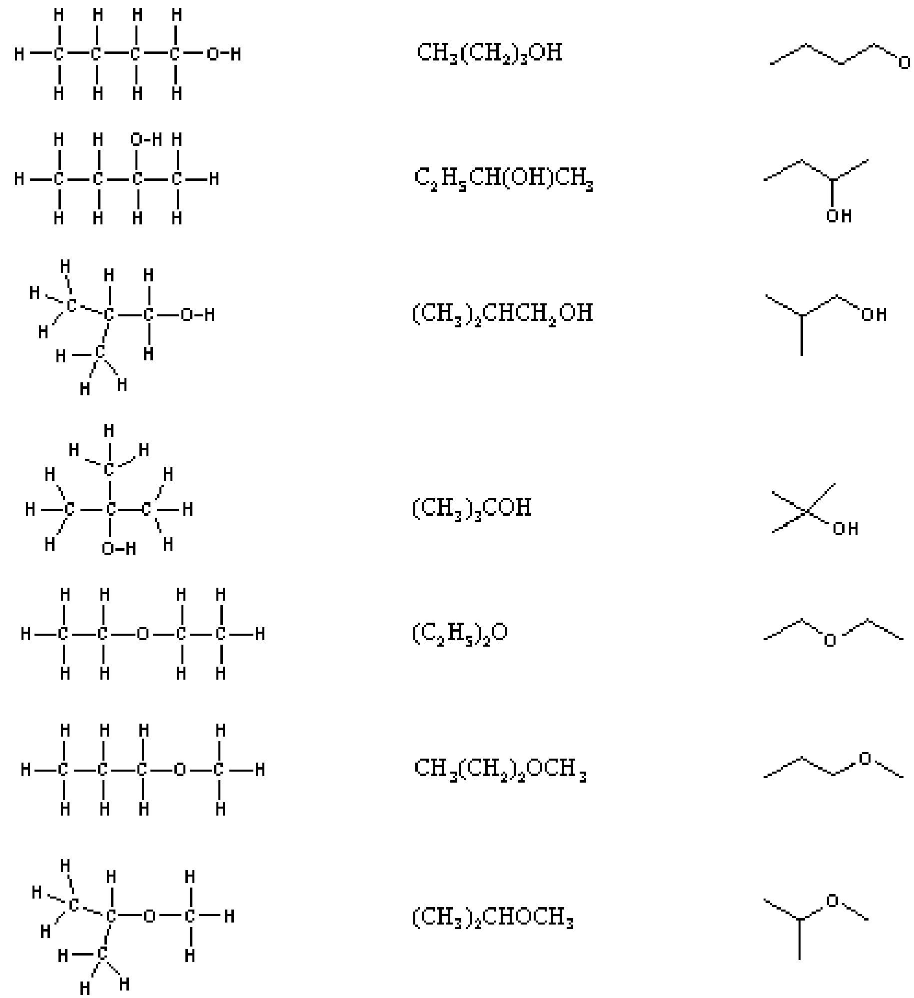




Organic Chemistry

Organic Chemistry-
shows the versatility of carbon

- has 4 valence electrons = 4 bonding spaces available.
- Backbone to many large, complex biological molecules (Carbs, Lipids, Proteins, Nucleic Acids)
- Over 16 million carbon-containing compounds are known.

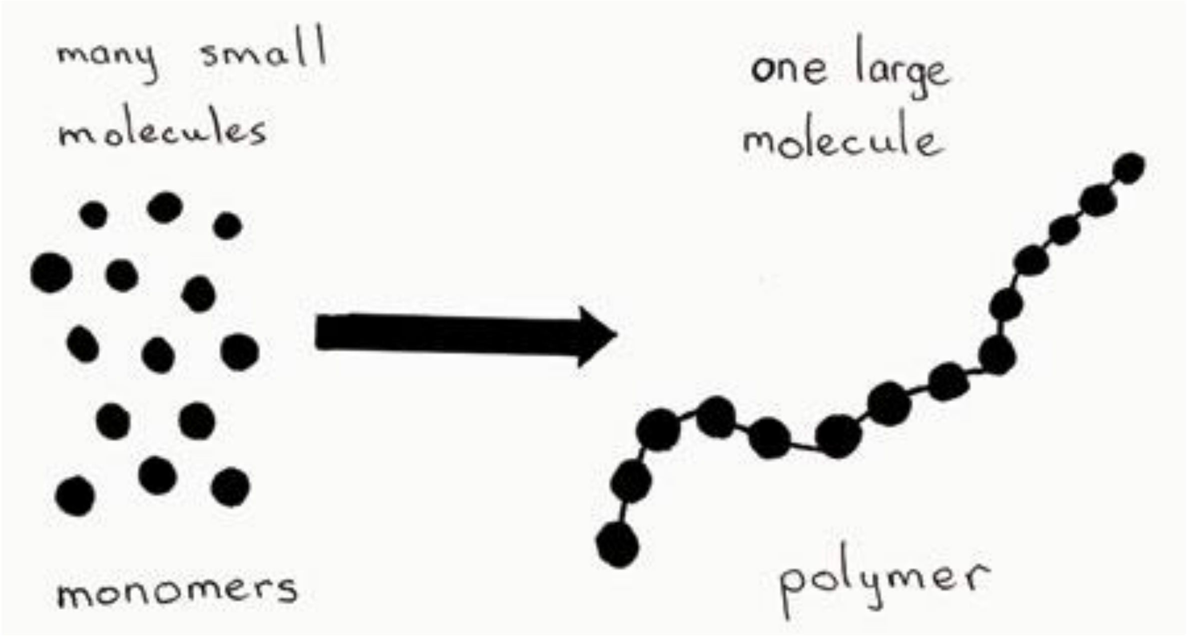


Monomers combine to make Polymers
(small unit) (large)

Common Examples of Polymers:

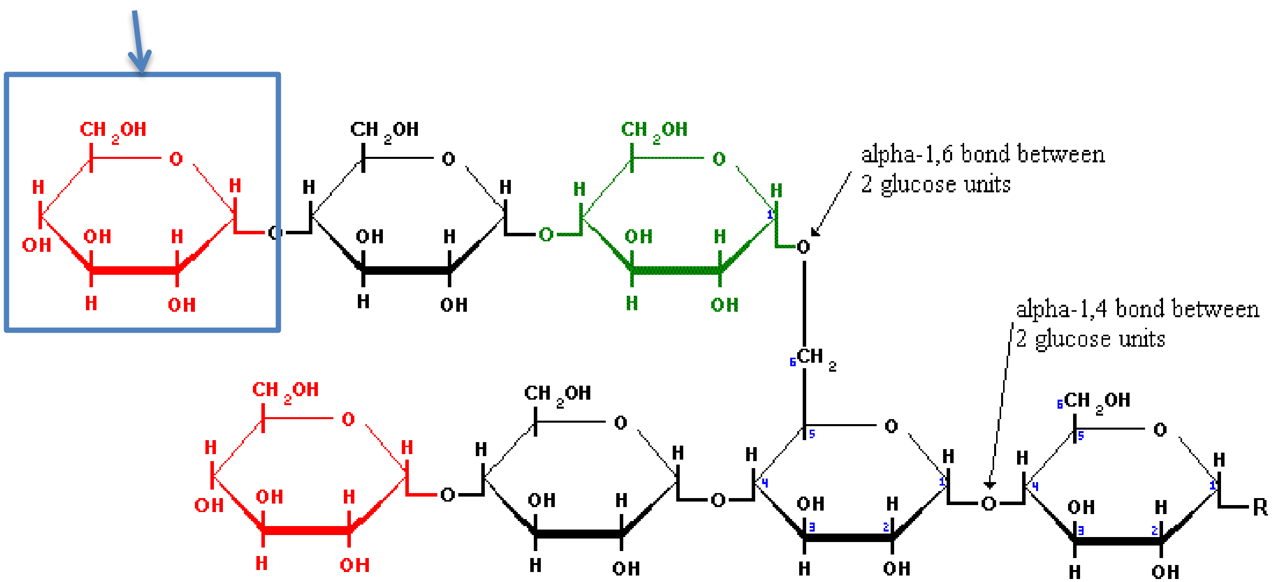
- **Carbohydrates**
- **Lipids**
- **Proteins**
- **Nucleic Acids**

(CLPN)



Ex: Carbohydrates

Monomer	**Polymer**	**Examples**
Monosaccharide	Polysaccharide	- Starch - Fiber - sucrose



Ex: Lipids

Monomers

Glycerol

& Fatty Acid tails

Polymer

Tri-glyceride

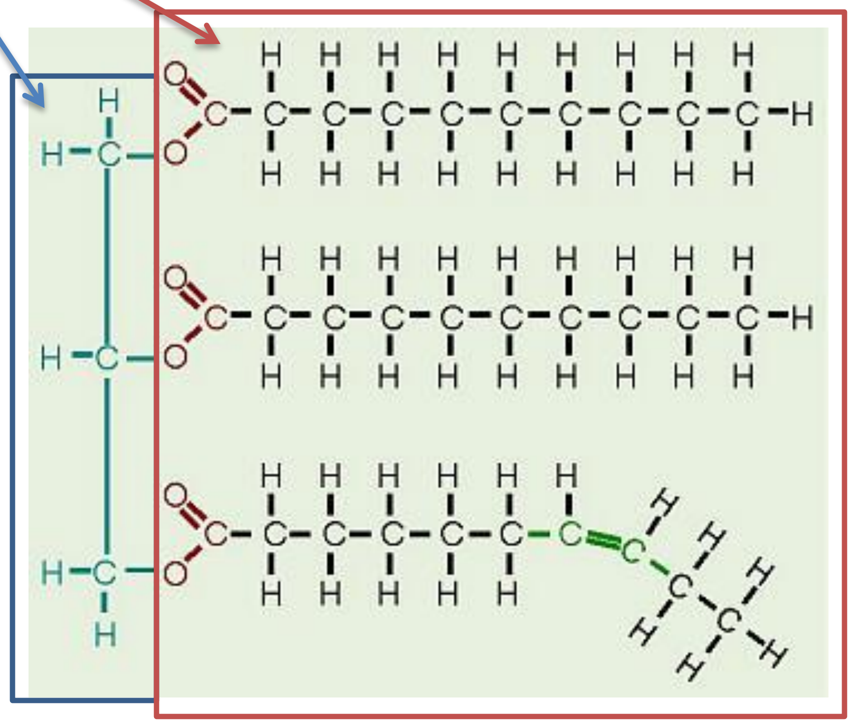
Examples

-Saturated Fats

-Unsaturated fats

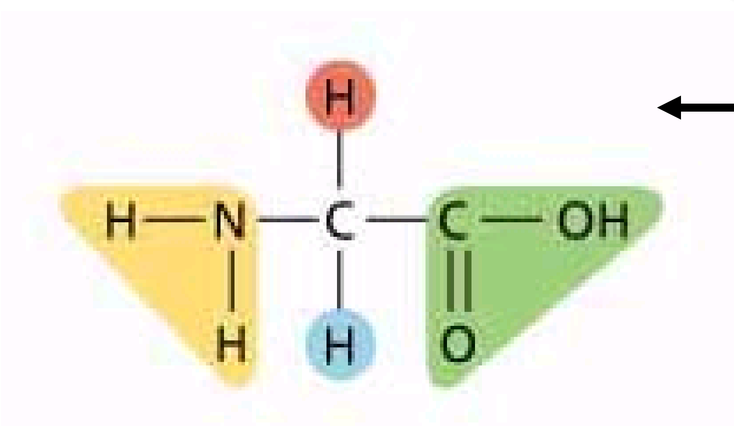
-Steroids

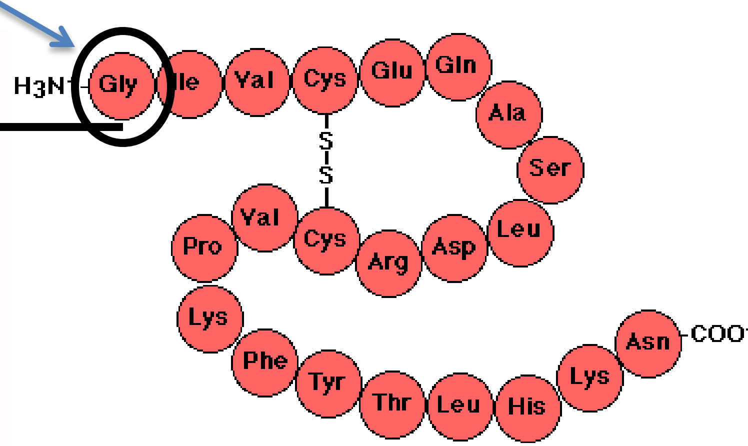
-Cholesterol



Ex: Proteins

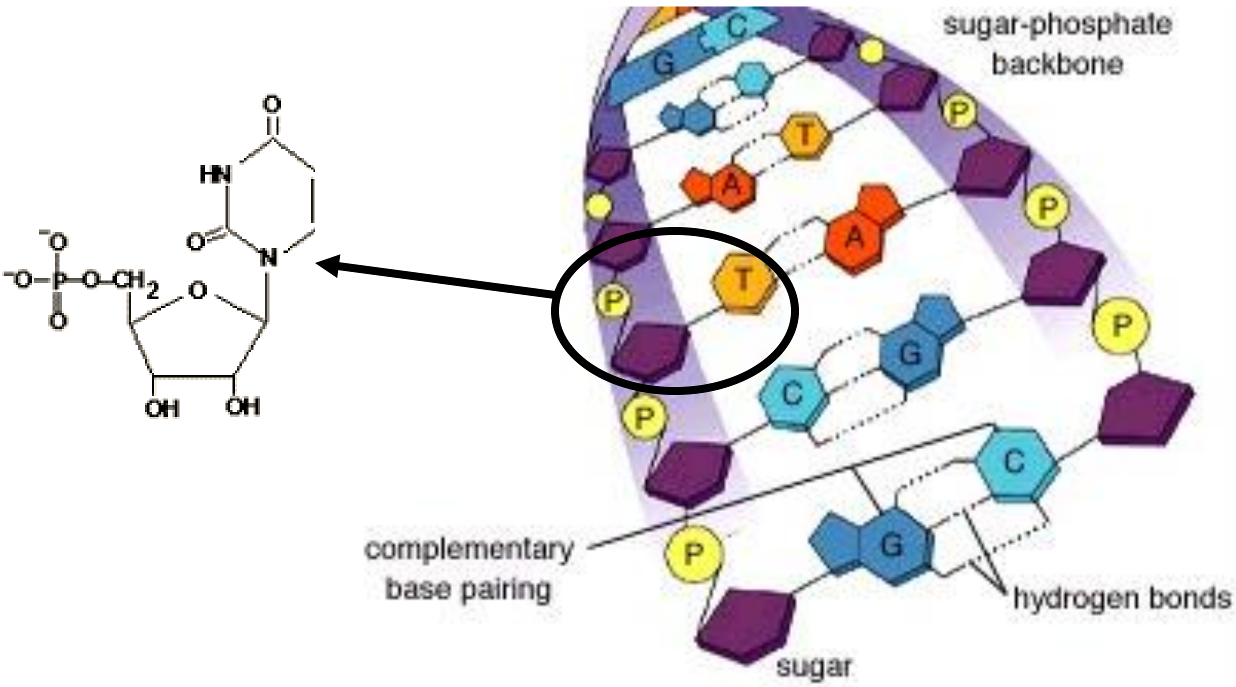
Monomer	Polymer	Examples
Amino Acids	Polypeptide	-enzymes -pigments -Meat/dairy



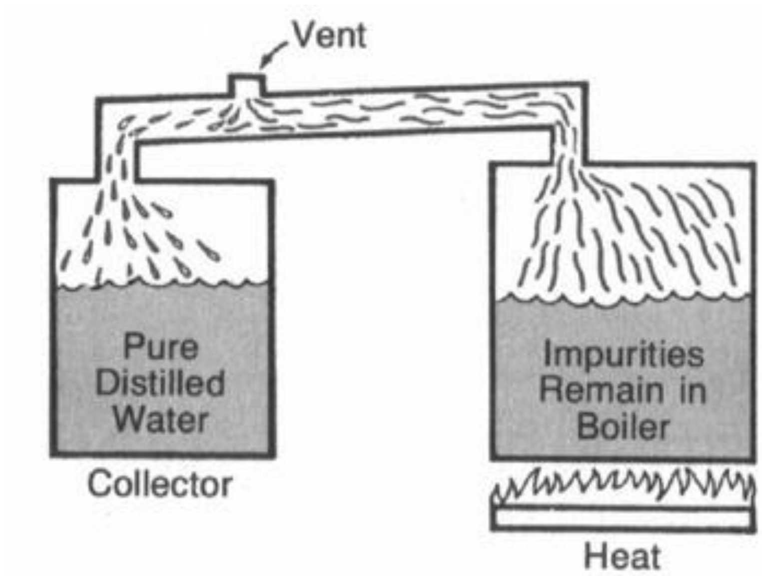


Ex: Nucleic Acids

<u>Monomer</u>	<u>Polymer</u>	<u>Examples</u>
Nucleotide	<u>Poly</u>nucleotide	-DNA -RNA



Distilled Water vs. Tap Water



Water Poisoning/ water Intoxication

Cause: excessive consumption of water during a short period of time.

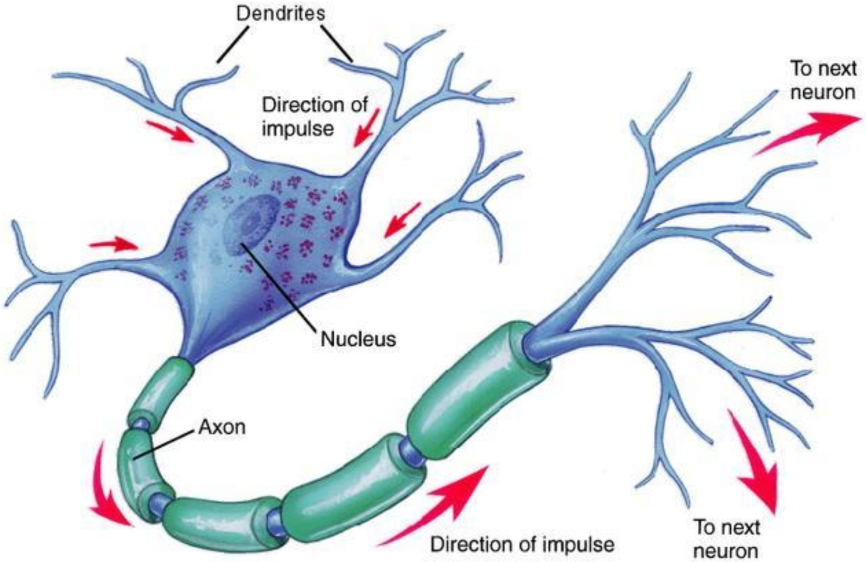
Why: leads to a disruption in normal brain function due to the imbalance of electrolytes in the body's fluids.

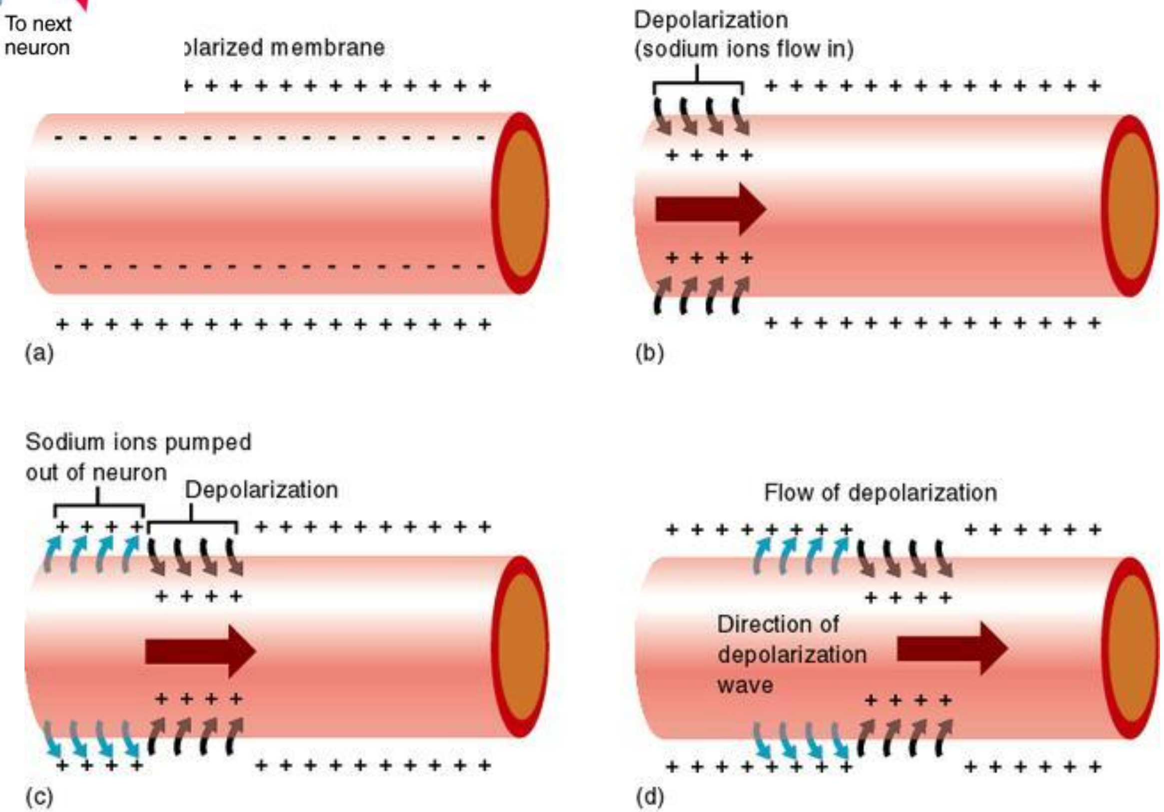
- can dilute the careful balance of sodium compounds in the body fluids

Who: individuals in water drinking contests…consume more than 10 liters (10.5 quarts) of water over the course of just a few minutes

- People doing endurance sports which electrolytes are not properly replenished, yet massive amounts of fluid are still consumed

Neural transmission





Electric Stimulation Machine-
stimulates muscles for you

See video clips on web links