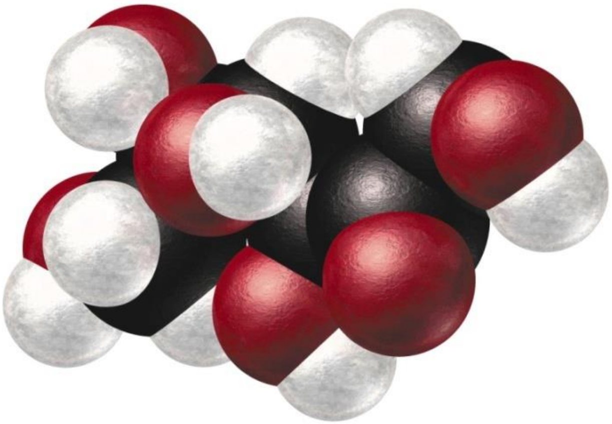
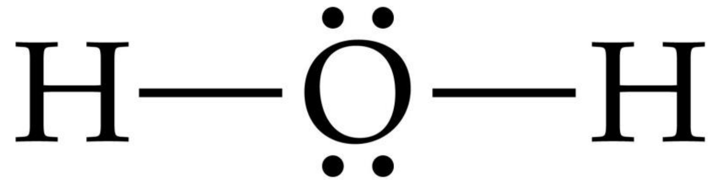


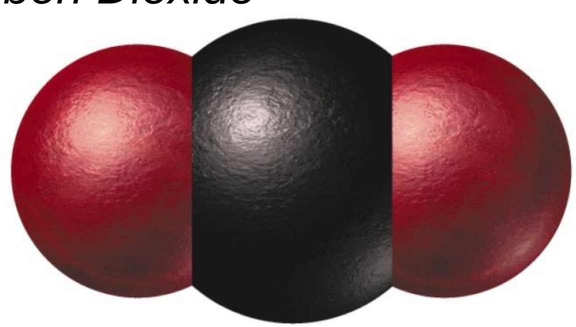
Fructose

Water

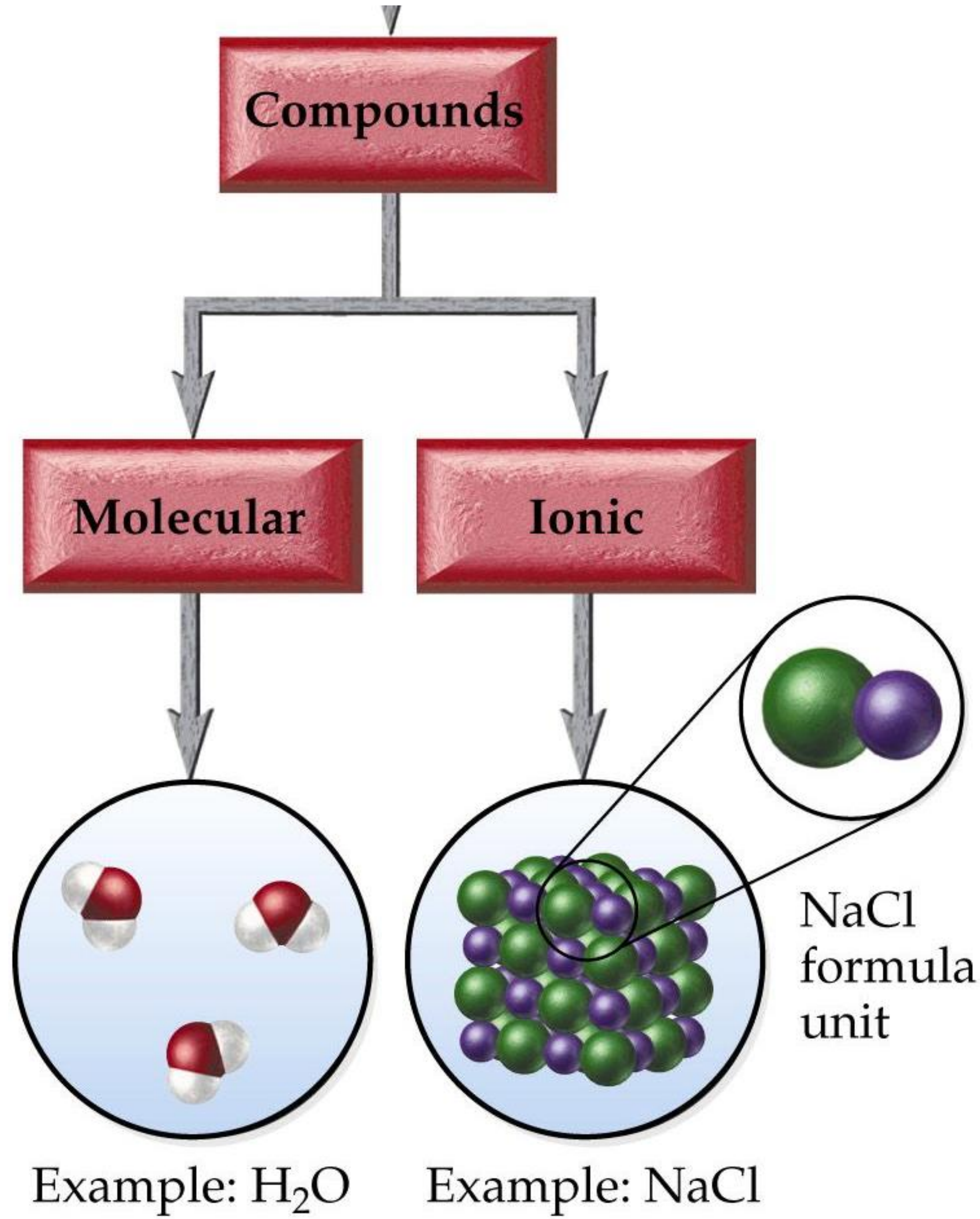


Review Notes 9.1
Covalent Bonding

Carbon Dioxide

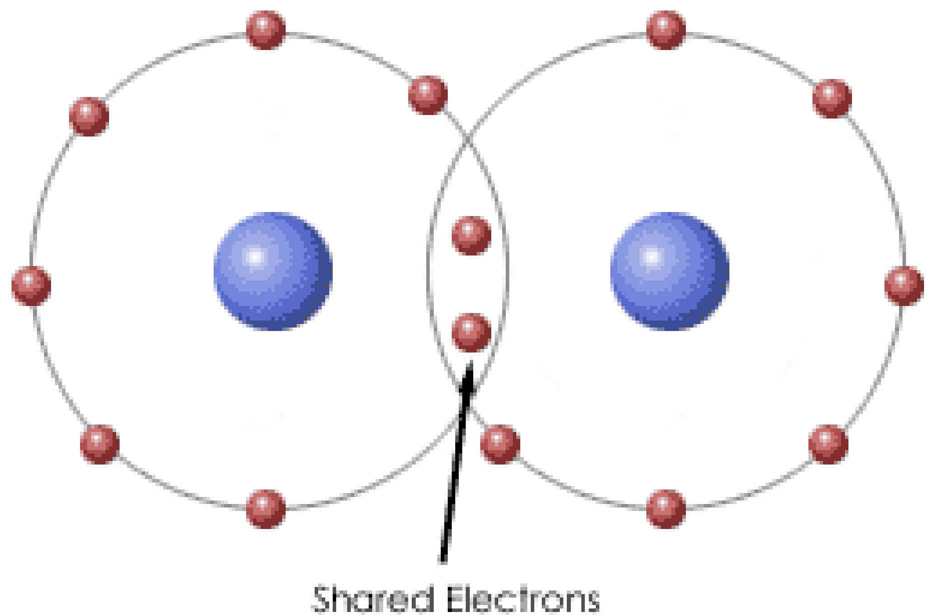


Ammonia


Compounds
Molecular
Ionic
NaCl formula unit
Example: H_2O
Example: NaCl

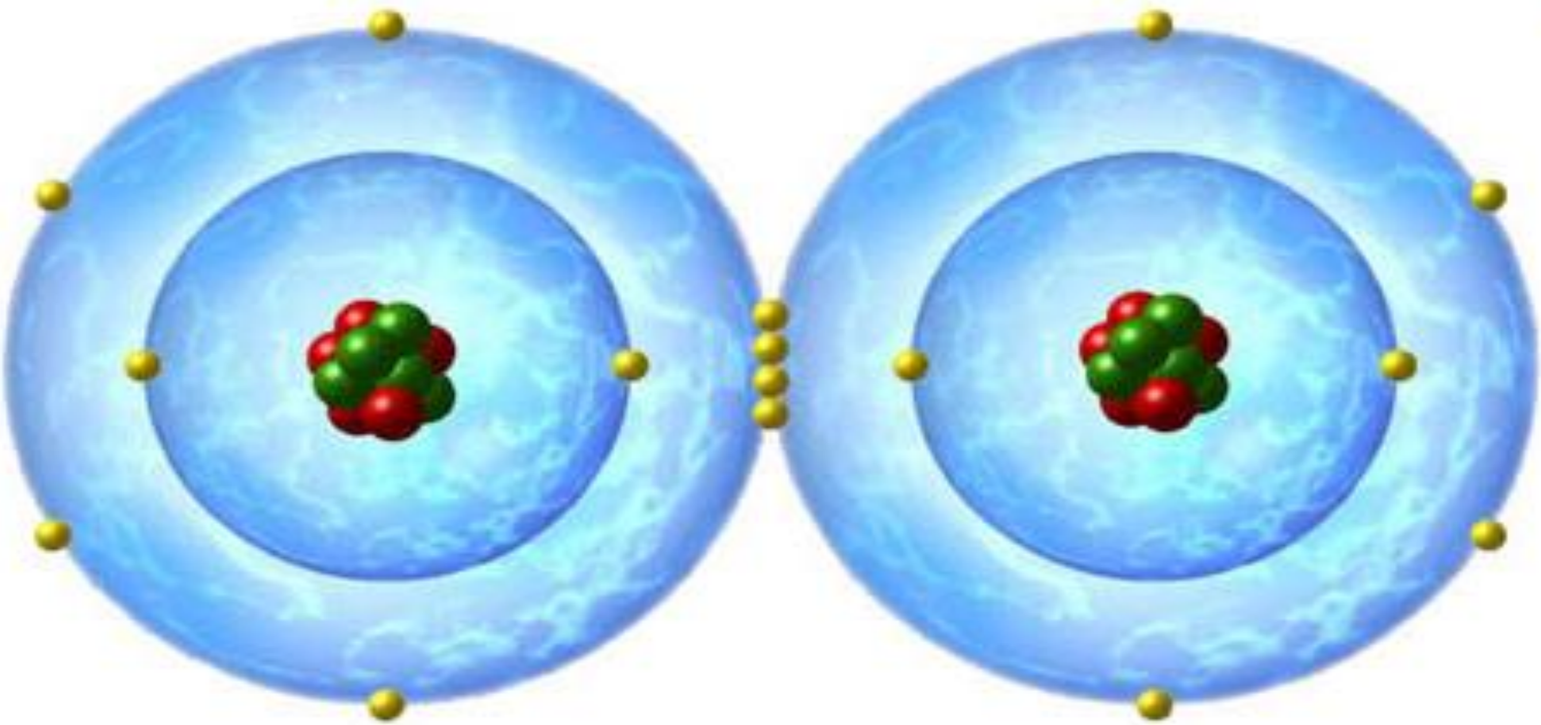
Molecules and Molecular Compounds

- Compounds bond by <u>sharing electrons</u>.
- Shared valence electrons = **<u>covalent bond</u>**.



“Molecules”

form when two or more atoms bond covalently.



O=O O_2

Covalent Bonds

- **Formed between <u>two nonmetals</u>.**
  - *Remember Ionic is between a <u>metal and nonmetal</u>*
- **Nonmetals have high electronegativity values**
  - *…means they are attracted to other electrons.*
- **Three types of sharing:**

  single bond shares 1 pair = 2 electrons —

  double bond shares 2 pairs = 4 electrons =

  triple bond shares 3 pairs = 6 electrons ≡

Learning Check

Indicate whether a bond between the following would be 1) Ionic 2) covalent

1	A.	sodium and oxygen
2	B.	nitrogen and oxygen
2	C.	phosphorus and chlorine
1	D.	calcium and sulfur
2	E.	chlorine and bromine

Lewis Structures

- Use electron-dot structures to show how electrons are arranged in molecules

- Example: Hydrogen H_2

H:H **H–H**

Drawing Lewis structures for covalent bonds:

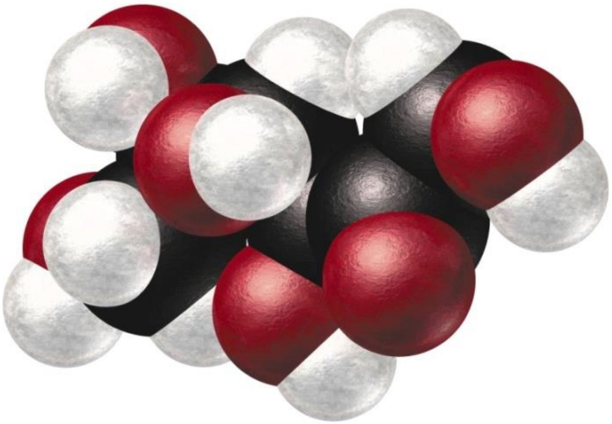
- **N** (Needed):
  - *All atoms need 8.*
  - *Hydrogen needs only 2.*
- **A** (Available):
  - *Number of valence electrons.*
- **S** (Shared):
  - *Subtract: Needed - Available*
  - *Divide by 2 = # of bonds.*
- **Draw** the molecule**.**
  - *First atom in the center.*
    - *H and Halogens are usually on outside.*
  - *Draw in bonds.*
  - *Fill in all **Available** electrons.*

Draw the Lewis-dot-structure for the following molecules

1) HF

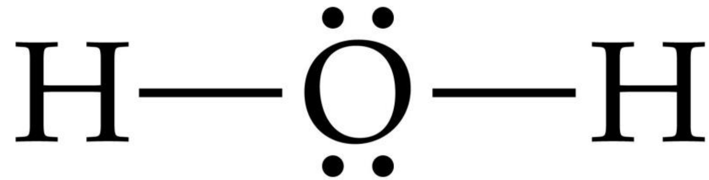
2) CCl_2H_2

3) CO



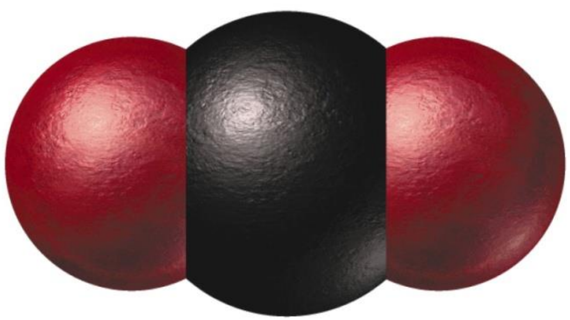
Fructose

Water



Review Notes 9.1 (pt.2)

Carbon Dioxide



Ammonia

Covalent Bonds secrets…

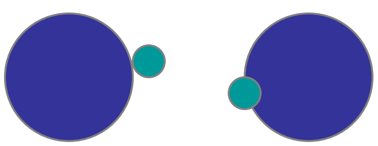
- Halogens and Hydrogen form **one** bond
  - Example: $\mathbf{H}_2$, $\mathbf{Cl}_2$
- Group 6A elements form **two** bonds
  - Example: $H_2\mathbf{O}$
- Group 5A elements form **three** bonds
  - Example: $\mathbf{N}H_3$
- Group 4A elements form **four** bonds
  - Example: $\mathbf{C}H_4$

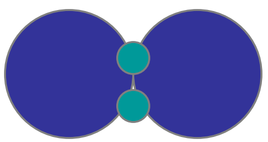
Covalent Bonds

When covalent bonds form energy is released.

- ***Exothermic reaction (energy released)***
- ***Bonds are stable = lower energy!***

H• + •H ⟶ **H:H + energy**





"hydrogen molecule"

Energy Change

- Energy is released when bonds form
- Energy is absorbed to break bonds
- Energy required is *bond dissociation energy*

- Exothermic is when more energy is released than is absorbed to break the bonds (= **hot**).
- Endothermic: more energy is absorbed to break bonds than was released (= **cold**).

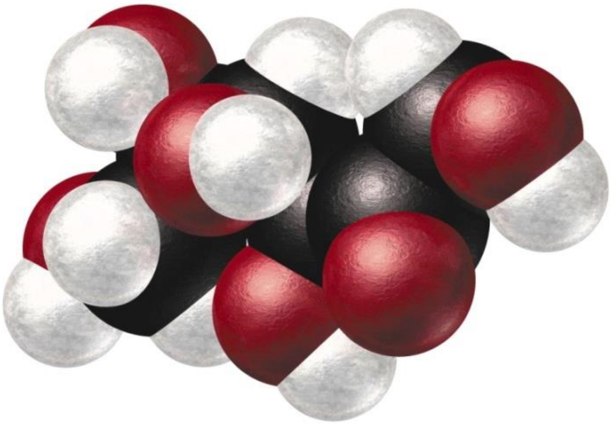
Strength of Covalent Bonds

- Distance between nuclei is *bond length*
- Bond length depends on:
  1. size of the atoms
     - *larger atom = larger distance*
  2. # of shared electron pairs (*bonds*)
     - *more shared pairs = shorter bond*
     - *Shorter = stronger*

Strength of Covalent Bonds

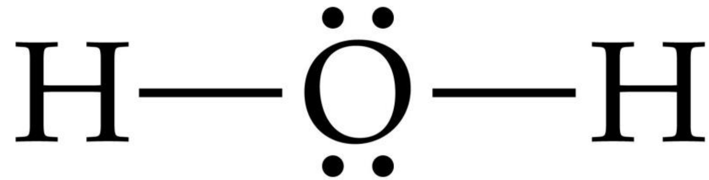
So…

- Single bond (*sigma bond*)
  - longest and weakest
  - lowest bond dissociation energy
- Double Bond (sigma+*pi*)
  - middle length and strength
- Triple bond (sigma+*pi*+*pi*)
  - shortest and strongest
  - highest bond dissociation energy



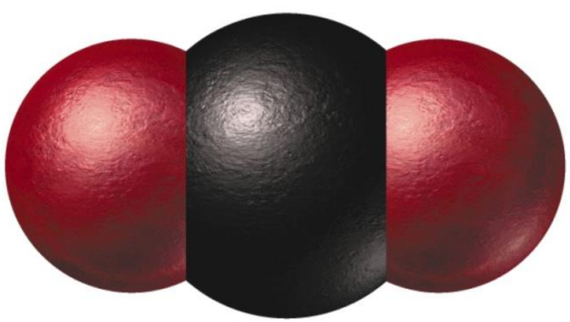
Fructose

Water



Review Notes 9.1 (pt.3)

Carbon Dioxide

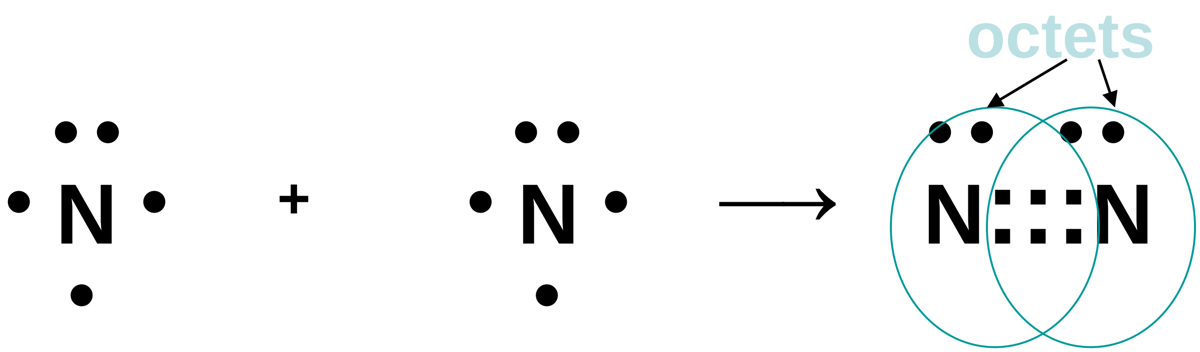


Ammonia

Diatomic Molecules

Seven elements are "diatomic"

- **more stable together than apart**
- **Halogens: F_2, Cl_2, Br_2, I_2**
- **Other: H_2, N_2, O_2**

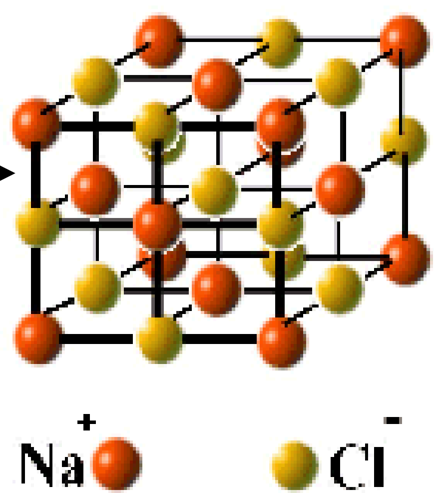


- **only when they are alone (i.e. no other elements)**

Properties of Molecular Compounds

- Usually **gases or liquids**.
- Relatively **low melting and boiling points**
  - *Reason: no bonds between molecules, so easier to separate!*

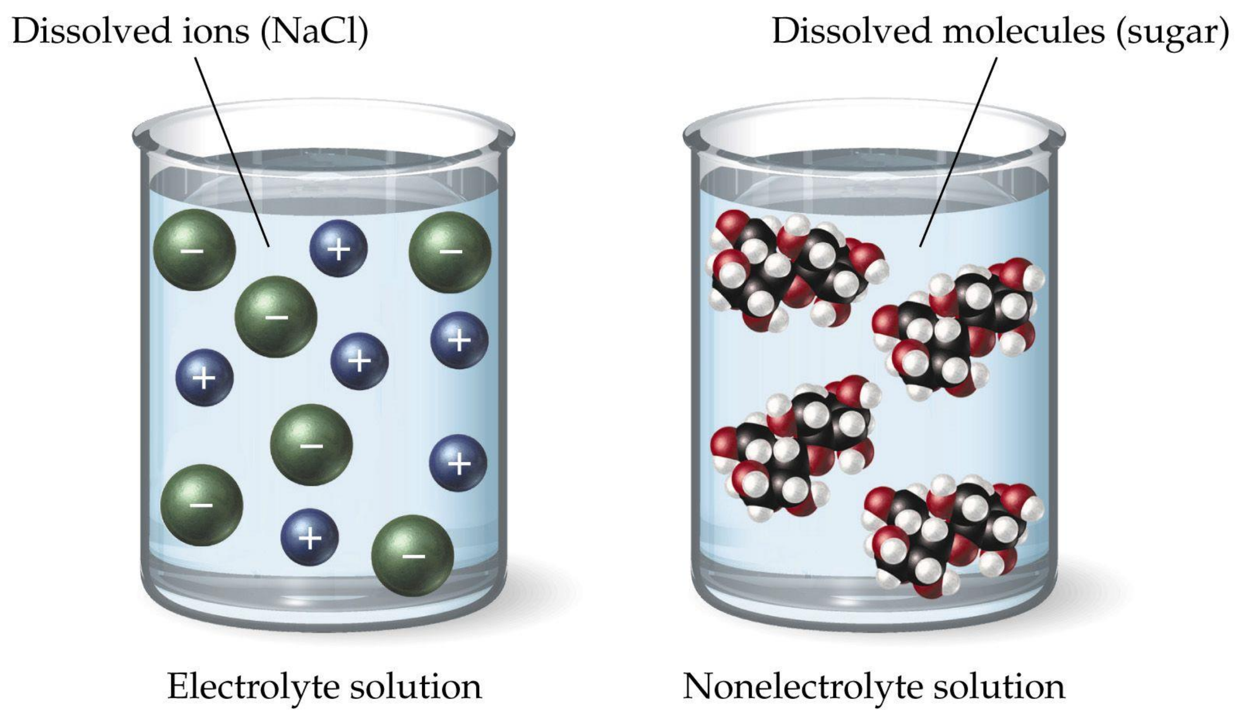
Ionic Compound



Molecular Compound

Properties of Molecular Compounds

- **Nonelectrolytes**
  - Do not conduct electricity.
  - Reason: do not break apart into ions.



Energy Change

- Energy is released when bonds form
- Energy is added to break bonds
- Energy required is *bond dissociation energy*

- Endothermic: more energy is required to break bonds than was released (cold).
- Exothermic is when more energy is released than is required to break the bonds (hot).

Review Notes 9.2:
Naming Binary Molecules

Naming Binary Molecular Compounds

- Binary = 2 nonmetals

- Scientific names reveal <u>composition</u>:
    - indicate the <u>number</u> and <u>type</u> of atoms

1. prefix+first nonmetal
    - ***mono-*** is only used with oxygen

2. prefix+second nonmetal w/***ide*****.**

Greek Prefixes

We use Greek all the time!

- What do you call a train that runs on one rail?
  - MONORAIL
- What do you call the non-motorized object that has pedals and **2** wheels?
  - **Bicycle**
- On your paper come up with a good example for the prefix we use daily for **3**?

Prefixes in Covalent Compounds			
Number of atoms	**Prefix**	**Number of atoms**	**Prefix**
1	mono-	6	hexa-
2	di-	7	hepta-
3	tri-	8	octa-
4	tetra-	9	nona-
5	penta-	10	deca-

Practice

Example: N_2O_4

Dinitrogen tetroxide

- SF_6 Sulfur hexafluoride
- Diphosphorus pentoxide P_2O_5
- SO_2 Sulfur dioxide
- nitrogen pentachloride NCl_5
- IF_7 Iodine heptafluoride

Anyone want a cold glass of

dihydrogen monoxide?

Formula	Common Name	Molecular compounds name
H_2O	**Water**	~~Dihydrogen monoxide~~
NH_3	**Ammonia**	~~Nitrogen trihydride~~

Review Notes 9.2 (pt.2): Naming Acids

What are Acids?

- Water solutions of some molecules are acidic.
  - *pH less than 7 = acid*
- Acids produce **H^+** in water
  - Ex: $HCl \rightarrow$ **H^+** $+ Cl^-$
  - *HINT: Any compound that starts with an H!*

Naming **Binary** Acids

Contains **H** + **X** (*an anion*)

1) Start with ***hydro-***
2) Change ending to ***–ic***
3) Add "***acid"***

Hydro- X -ic acid

Acid with no oxygen? Name as a binary acid!

– Ex: HCN= ***h***

Naming **Oxyacids**

- *Oxyacids* = H^+ + oxyanion
- Oxyanions: *Anion w/oxygen (NO_3^-, PO_4^{3-}, ClO^-)*
- Name comes from **polyatomic**!
  1) Identify the polyatomic
     - replace ***–ate*** with ***-ic***
     - replace ***–ite*** with ***–ous***
  2) Add ***acid.***

“ate→ic
ite→ous”

Naming Oxyacids

- Example:
  - HNO_3
  - $HNO_3 \rightarrow H^+ + \mathbf{NO_3^-}$
  - the polyatomic nitr**ate** → nitr**ic** + acid

  - $HClO_2$
  - $HClO_2 \rightarrow H^+ + \mathbf{ClO_2^-}$
  - the polyatomic chlor**ite** →chlor**ous** acid

hydro- is **NEVER** part of oxyacids!

Classify then name:

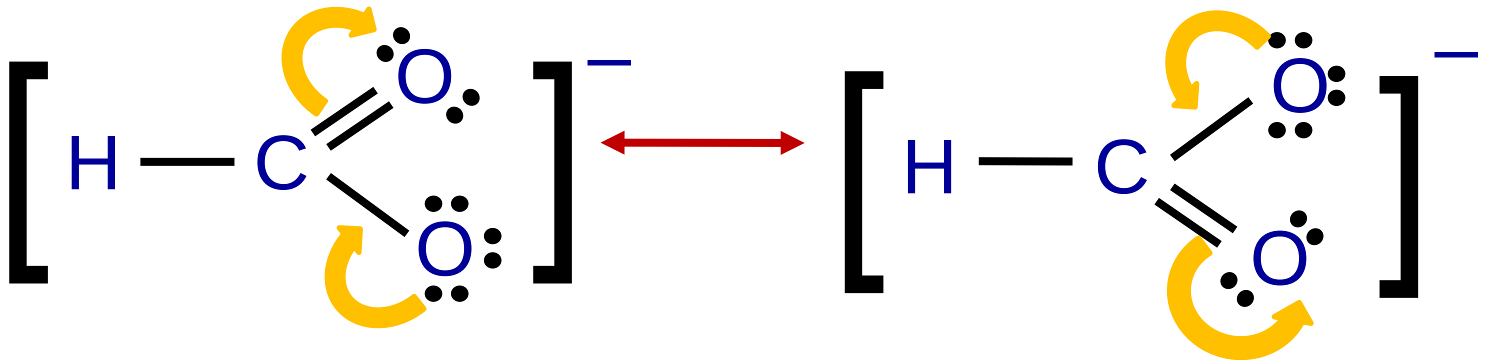
Assume they are dissolved in water!

1) HI
2) HClO
3) Phosphoric acid
4) Hydrocyanic acid
5) H_2CO_2
6) Acetic acid

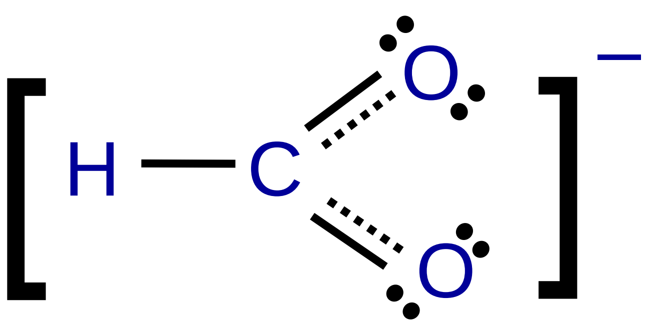
Review Notes 9.3

Resonance Structures

- Occur when there is more than one valid Lewis structure
- Differ in position of electron pairs, never atom positions

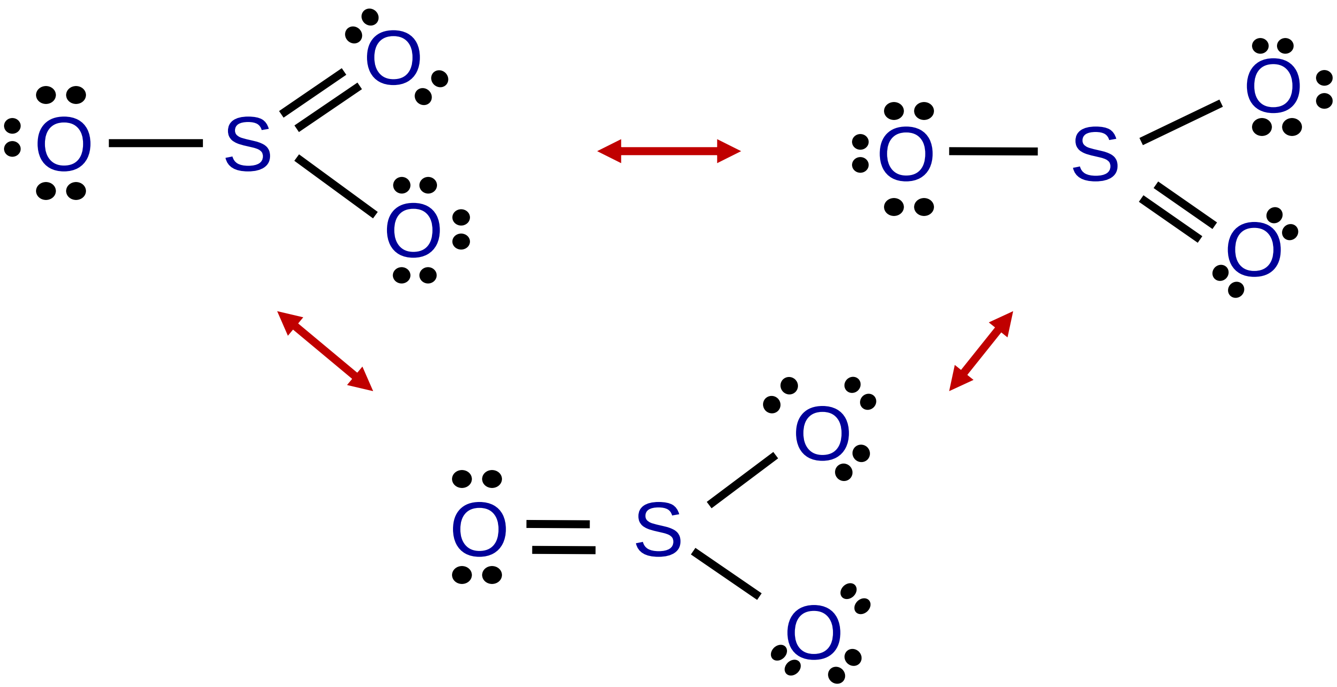


- Actual molecule behaves as one structure



Question

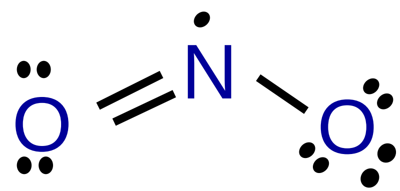
1) Draw the Lewis Dot structures for SO_3.



Exceptions to the Octet Rule

1) A small group of molecules has an odd number of valence electrons.

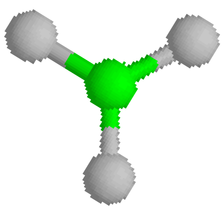
   -Ex: NO_2



2) **BORON is stable with 6 electrons!**

   - Ex: BH_3

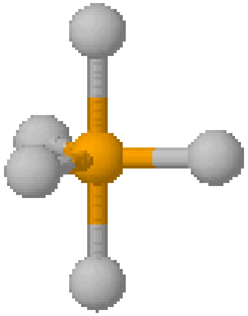
```
H - B - H
    |
    H
```



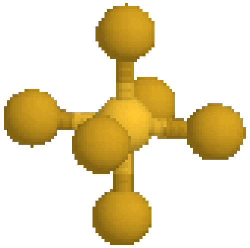
Exceptions to the Octet Rule

3) Some central atoms have more than eight valence electrons
   - Referred to as an "expanded" octet
   - Explained by including d-orbitals
   - Extra lone-pairs/bonds are added to the central atom.

PCl_5 (10 e^-)



SF_6 (12 e^-)



Question

Xenon will form a few compounds with nonmetals that strongly attract electrons.

Draw the correct Lewis structure for **XeF_4**.

Review Notes 9.4 - Molecular Shape

– Molecules have **3-D**imensional shapes!

– Determined by number of <u>bonds</u> and <u>lone pairs</u> on **central atom**

– Use **VSEPR** theory

- **V**alence-**S**hell-**E**lectron-**P**air **R**epulsion

Using the VSEPR Model

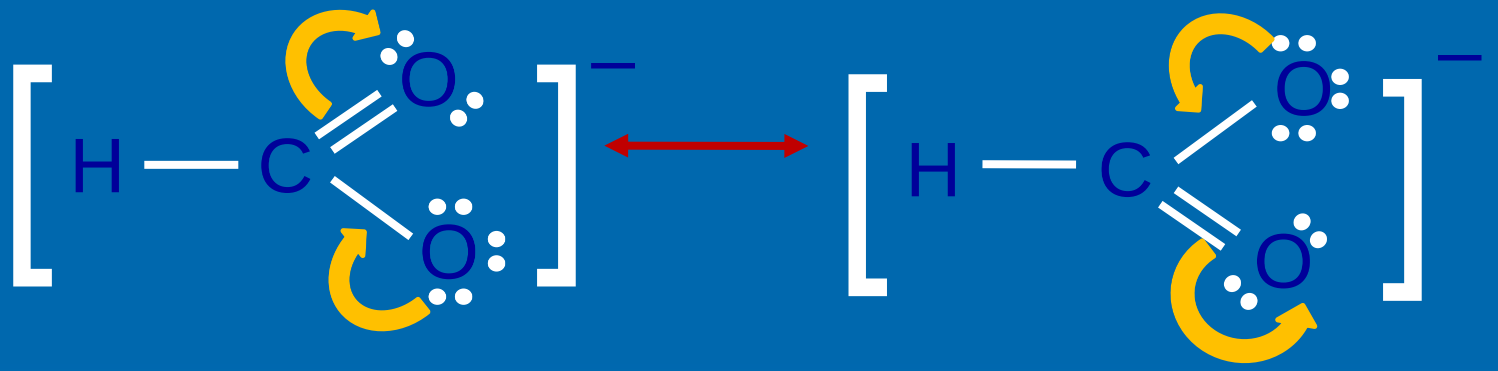
1. Draw Lewis structure *(use NAS)*
2. Identify **central** atom
3. **Count** electron pairs:
    1. *# of bonded pairs = ?*
    2. *# of lone pairs = ?*
4. Predict the shape of the molecule

Bonding and Shape of Molecules

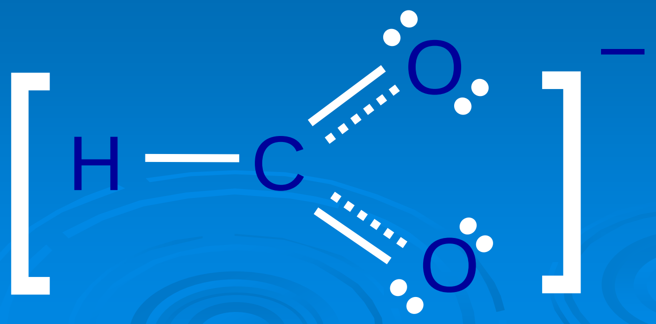
Number of Bonds	***Number of Lone Pairs***	*Covalent Structure*	***Shape***	*Examples*
2	**0**	-Be-	**Linear**	$BeCl_2$
3	**0**	—B— with a third bond below	**Trigonal planar**	BF_3
4	**0**	—C— with bonds above and below	**Tetrahedral**	CH_4, $SiCl_4$
3	**1**	—N̈— with a third bond below	**Pyramidal**	NH_3, PCl_3
2	**2**	—Ö: with a second bond below	**Bent**	H_2O, H_2S, SCl_2

Resonance Structures

- Occur when there is more than one valid Lewis structure
- Differ in position of electron pairs, never atom positions



- Actual molecule behaves as one structure



Question

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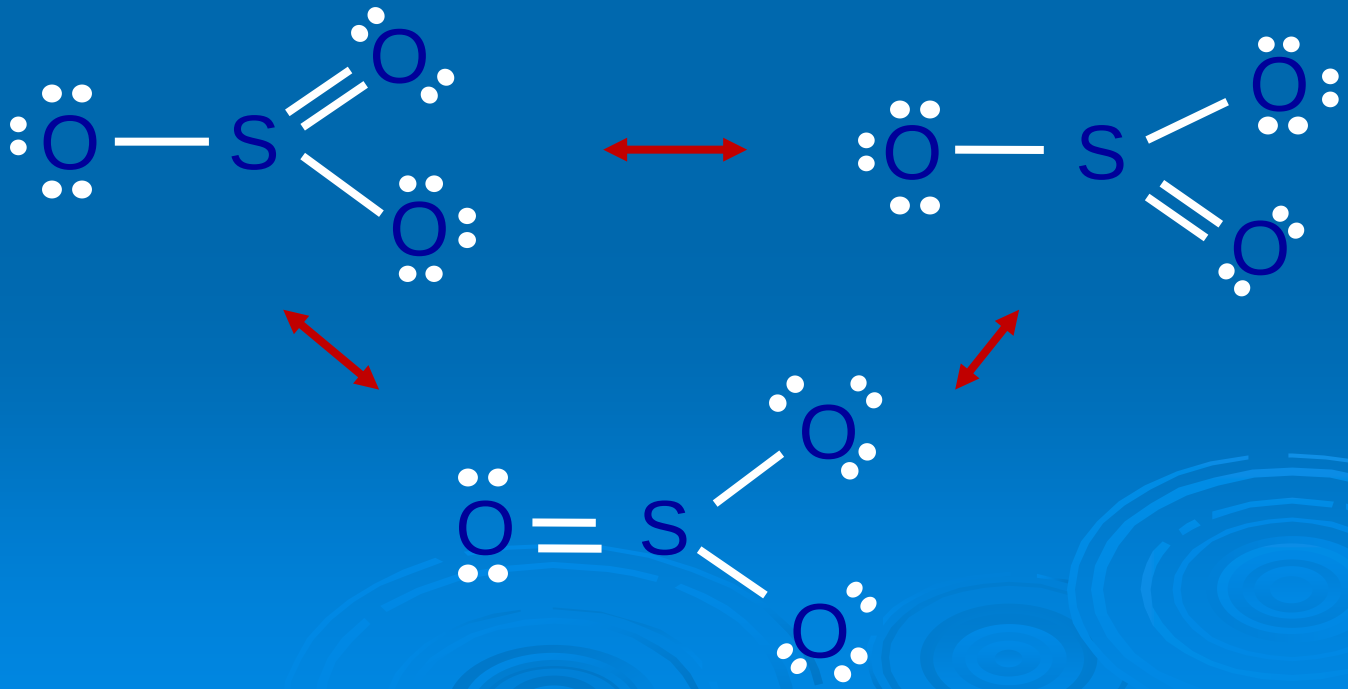
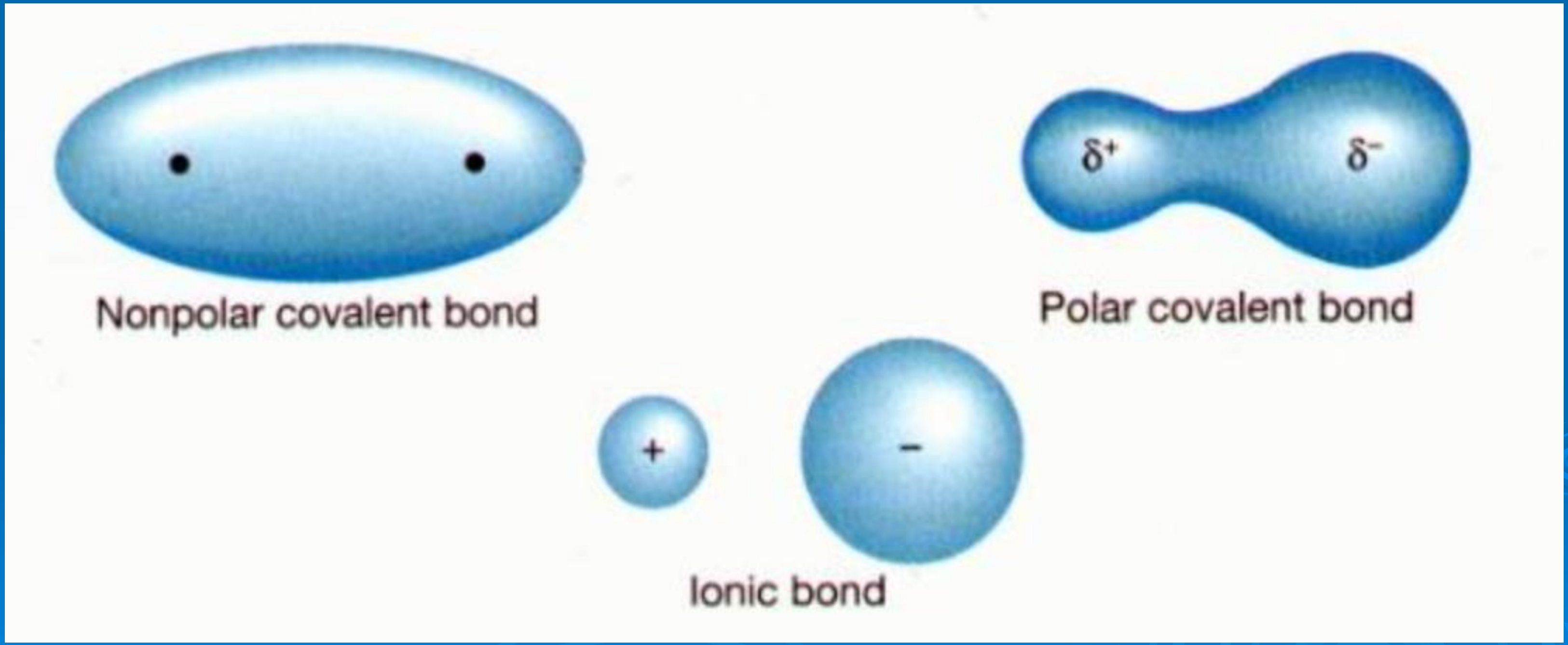


Electronegativity and Polarity
Review Notes 9.5 pt.1



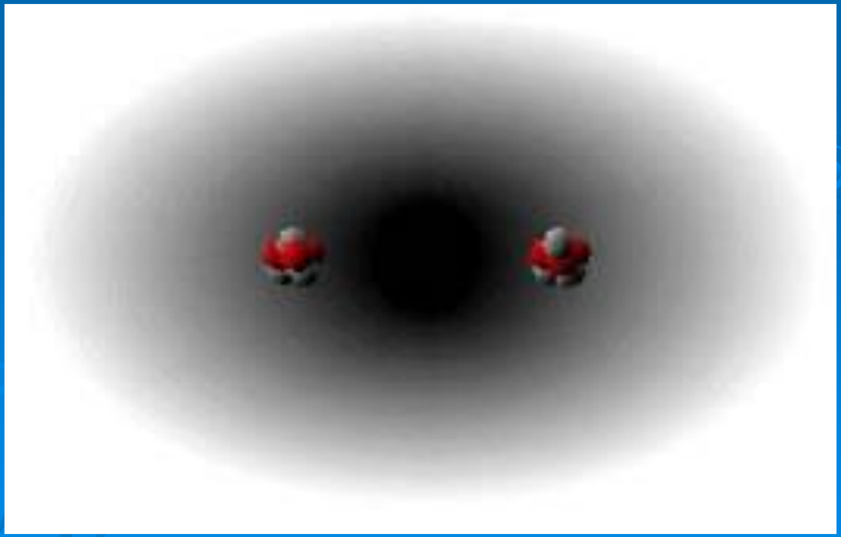
Recall Electronegativity…

^{9}F

- Electronegativity = ability to attract electrons in a bond
  - Increases across a period
  - Decreases down a group
- Fluorine is the most electronegative
  - *N, O, F are electron HOGS!!!*
- Metals have low electronegativities.

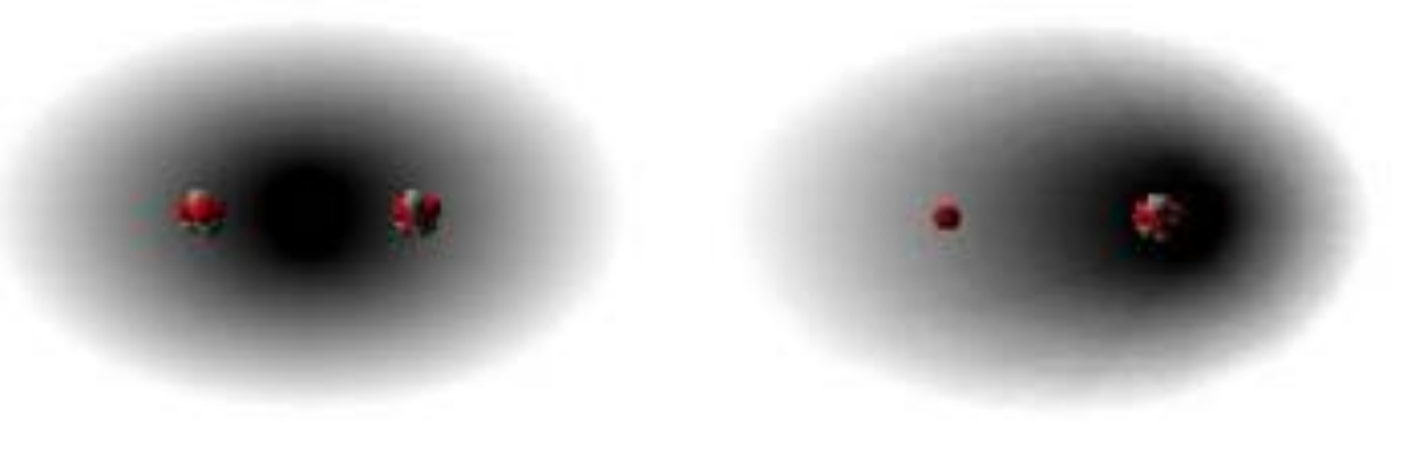
Electronegativity & Bond Character

- Character and type predicted by electronegativity differences.
- Identical atoms: *electronegativity difference = zero*
  - Electrons are equally shared
  - **Pure** covalent bond = **NON-POLAR**
  - Ex: N_2 Br_2



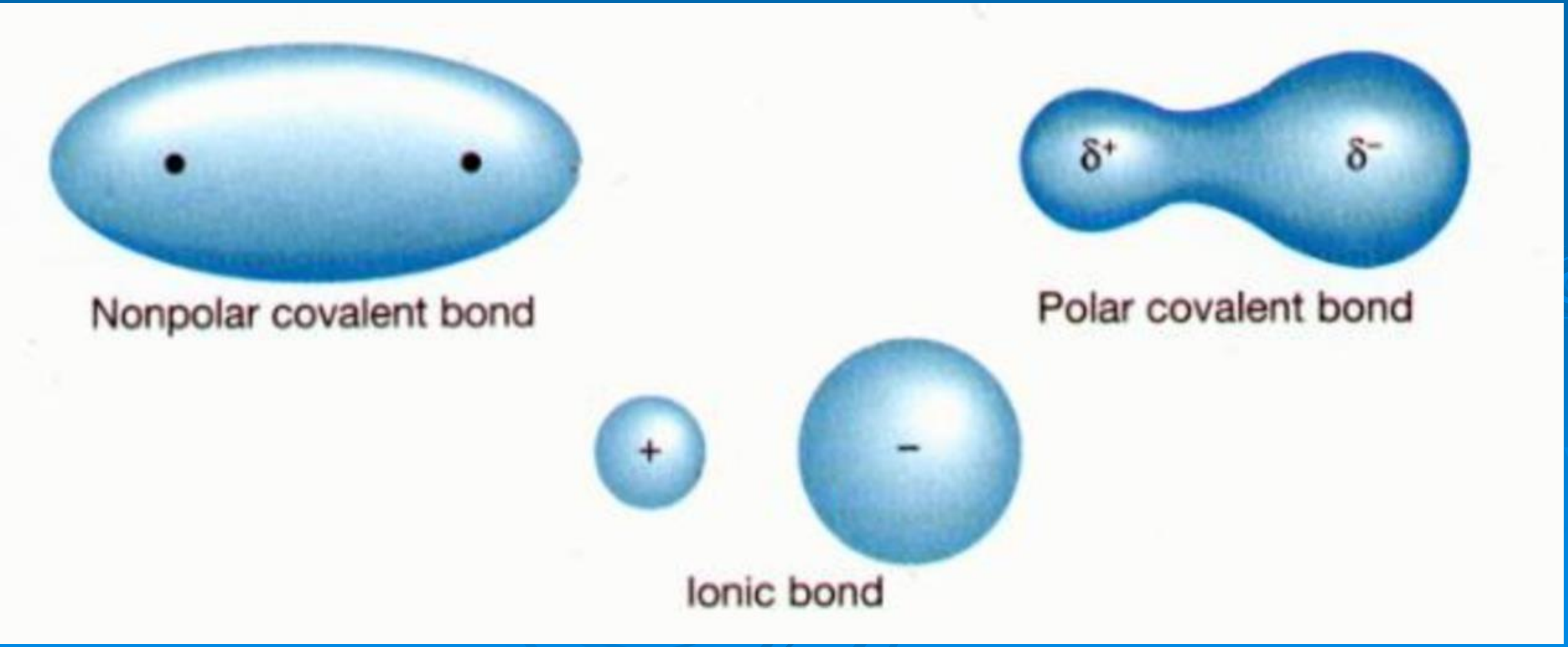
Electronegativity & Bond Character

- Different atoms: difference electronegativity $\geq$ zero
  - ~ 0-0.40 **NON-POLAR COVALENT**
  - ~ 0.4-2.0 **POLAR COVALENT**
    - Can range from *SLIGHTY* polar to *VERY* polar
    - Examples: O-Cl O-S N-Cl



Electronegativity & Bond Character

- Large difference in electronegativity result in an ionic bond (greater than 2.0).
  - Examples: NaCl KF



Learning Check

Identify the type of bond between the following atoms

A. K-N — Ionic (metal + non-metal)

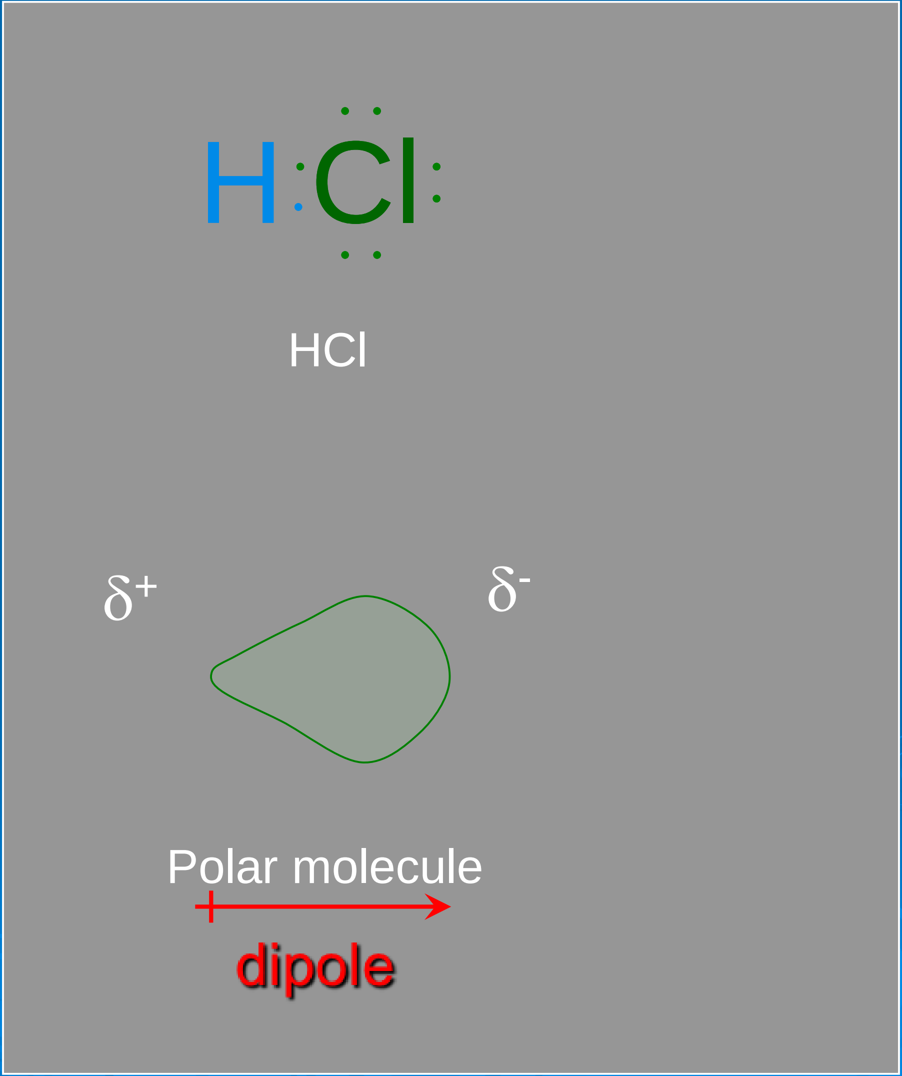
B. N-O — Slightly Polar covalent
(O to the right of N)

C. Cl-Cl — Pure covalent (equal sharing)

Electronegativity and Polarity
Review Notes 9.5 pt.2

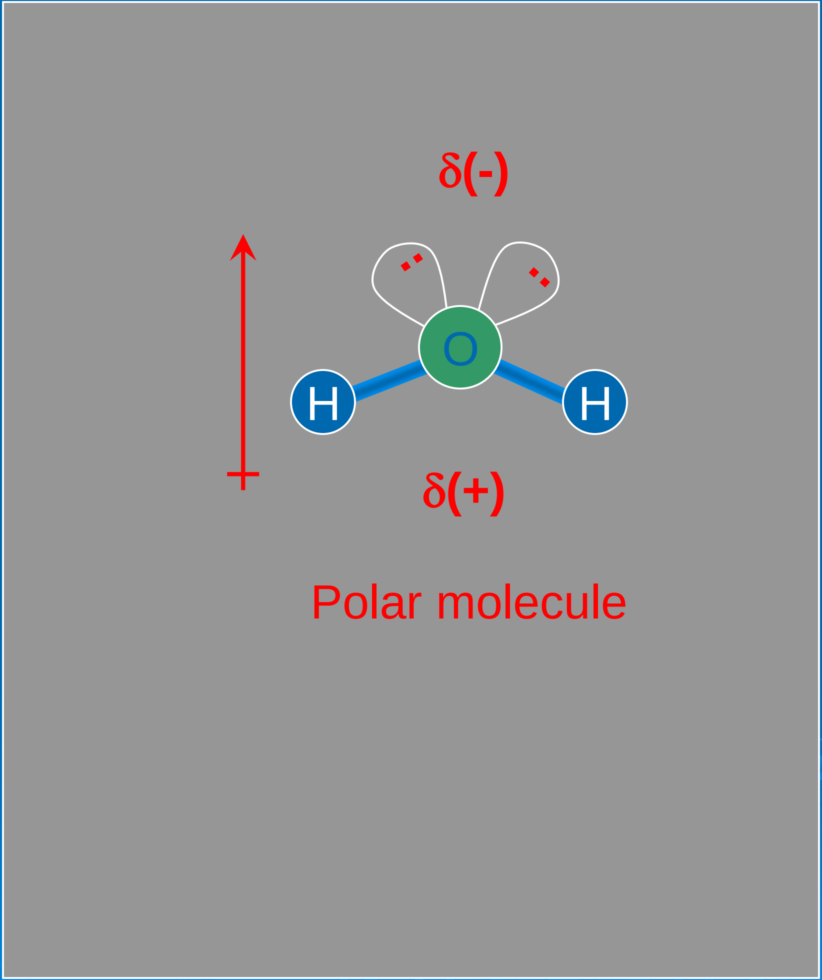
Bond Polarity

- It's a tug-of-war…
- The shared pair of electrons is pulled toward one of the atoms



Molecular Polarity

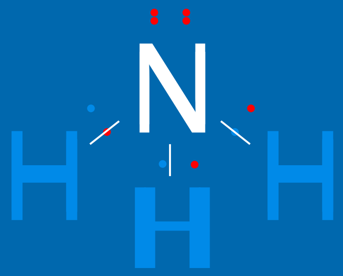
- Looks at ALL the bonds.
- A greater density of electrons on one side of the molecule =
  - partial negative δ^-
  - partial positive δ^+
- Creates a **dipole**

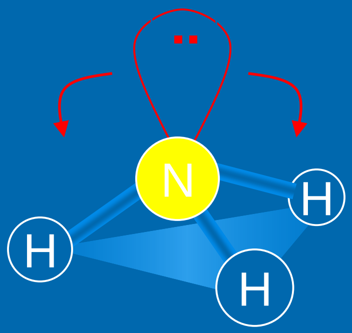


Polar Molecule or not?

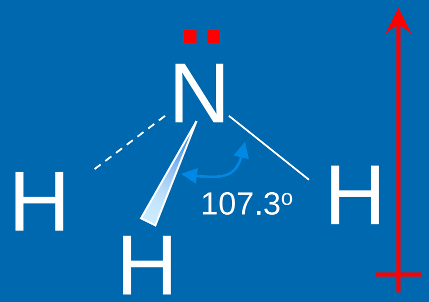
- The geometry is what is important
  - Asymmetrical = polar
  - Symmetrical = non-polar
    - Partial charges are balanced.
- Is ammonia polar or non-polar?
  1. Check symmetry
  2. Draw VSEPR
     1. *NAS*
     2. *Lewis*
     3. *Count bonded & lone pairs*
     4. *Check table*

Ammonia- NH_3



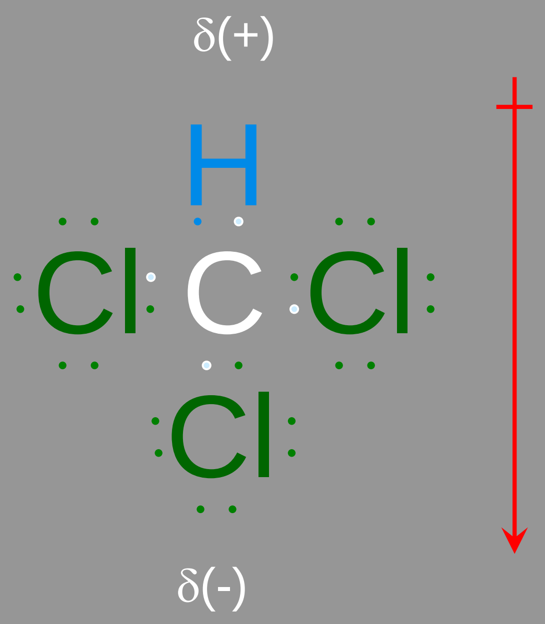


Pyramidal geometry



Molecular Polarity

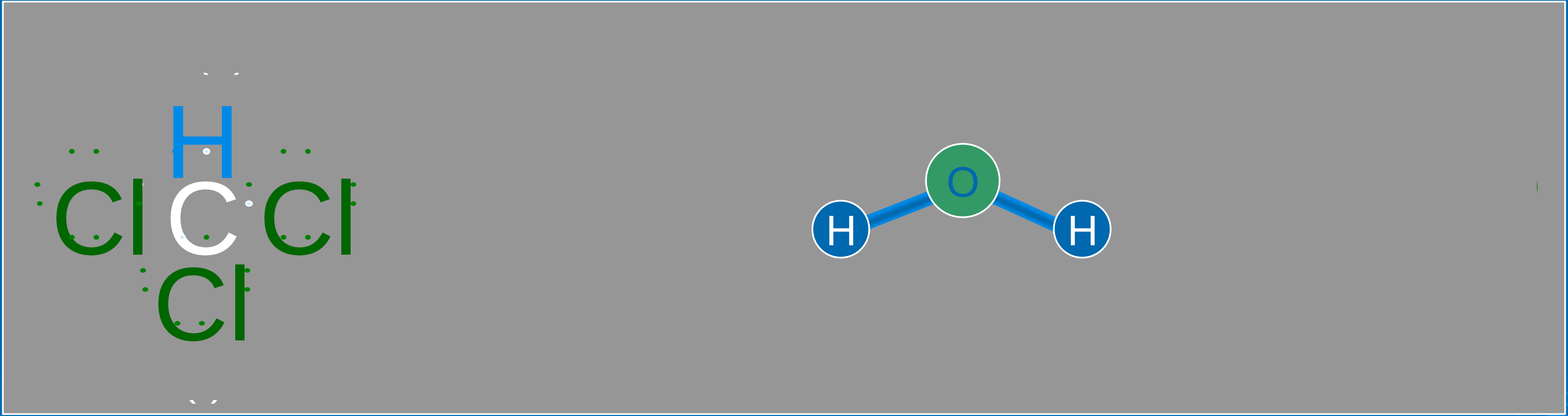
- Trichloromethane
- C-Cl are considered polar



Polar molecule

Polar Molecule or not?

- Some molecules contain polar bonds but are not polar molecules

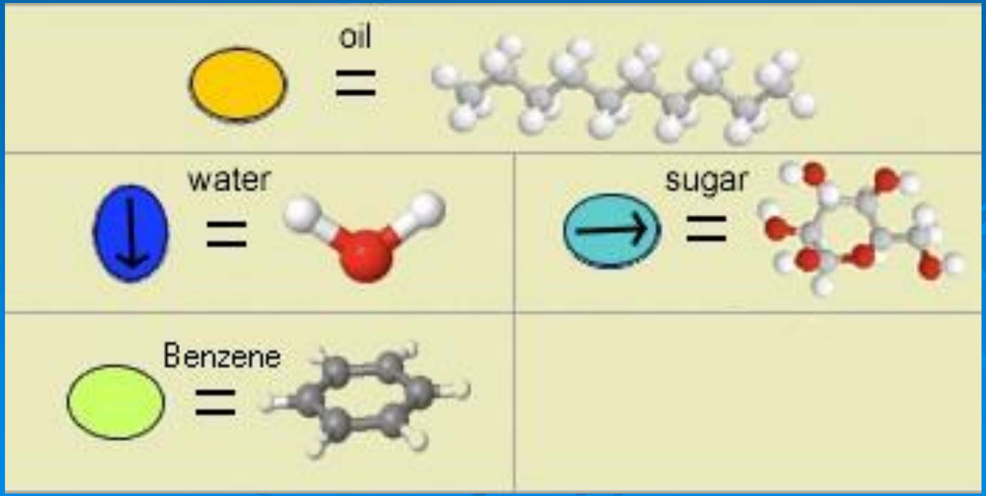


  - Carbon tetrachloride molecules are nonpolar
  - Water molecules are polar
  - Carbon dioxide molecules are nonpolar

Electronegativity and Polarity
Review Notes 9.5 pt.3

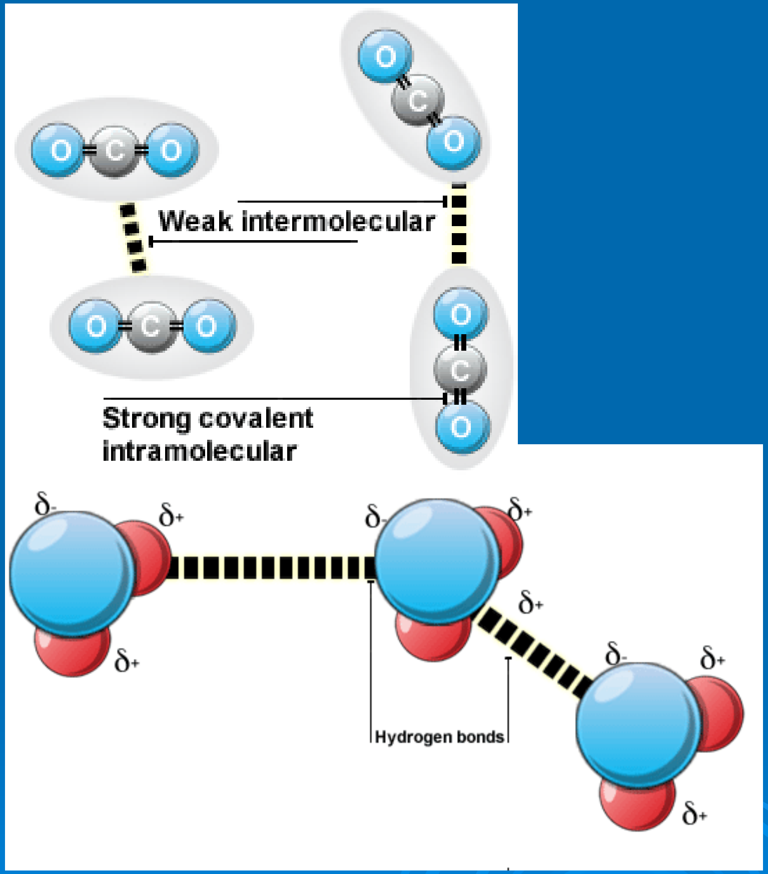
Solubility of Polar Molecules

- "LIKE dissolves LIKE"
  - Polar molecules and ionic compounds are soluble in polar substances
  - Non-polar molecules dissolve in non-polar molecules



Properties of Covalent Compounds

- *Intermolecular forces* are the attraction between molecules
  - "van der Waal forces"
- Weak for non-polar molecules
  - Dispersion forces
  - Movement of electrons
- Stronger for polar molecules
  - dipole-dipole interactions



- **Many properties determined by these forces**
  - Boiling & melting point, state of matter, etc.