# Chapter 10: Chemical Bonding

Chemical Bonds are the attractive forces that hold atoms together in more complex units.

- An understanding of how and why atoms attach together in the manner they do is central to chemistry

## Lewis Theory

Emphasizes the importance of valence electrons

| |
|---|
| VALENCE ELECTRONS - electrons in the outermost shell - ARE THE ELECTRONS INVOLVED IN BONDING |

Uses **dots** to represent valence electrons either **ON** or **SHARED** by atoms

**Bonding between atoms occurs by either transfer or sharing of electrons to achieve outer shells with**

**8 electrons** **(exceptions: Li, Be and He)**

## Lewis Electron Dot Symbols

- Uses symbol of element to represent the **nucleus and inner (core) electrons**
- Uses dots around the symbol to represent valence electrons

(Put one electron on each side first, then pair)


Remember that elements in the **same group** have the **same number of valence electrons**; therefore, their Lewis dot symbols will **look alike**.

# Ionic Compounds: Electrons Transferred

Write the Lewis symbol for:

a) Arsenic  b) Iodine  c) Silicon

Noble gases are considered stable because they do not react with other elements. This stability is attributed to their **FULL VALENCE SHELLS** - they have a complete **OCTET of ELECTRONS** (exception: helium)

He: $1s^2$

Ne: $1s^2 2s^2 2p^6$

Ar: $1s^2 2s^2 2p^6 3s^2 3p^6$

Kr: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$

Xe: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6$

## Formation of Ionic Compounds

Atoms of many elements that LACK a complete octet of electrons in their outer shells **react in such a way to attain it.**

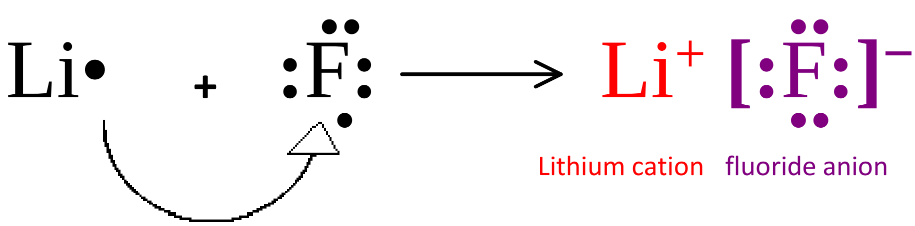
- They may lose or gain electrons depending on the type of element the atom is (metal or nonmetal)

Recall: When a neutral atom loses or gains one or more electrons an ion is formed.

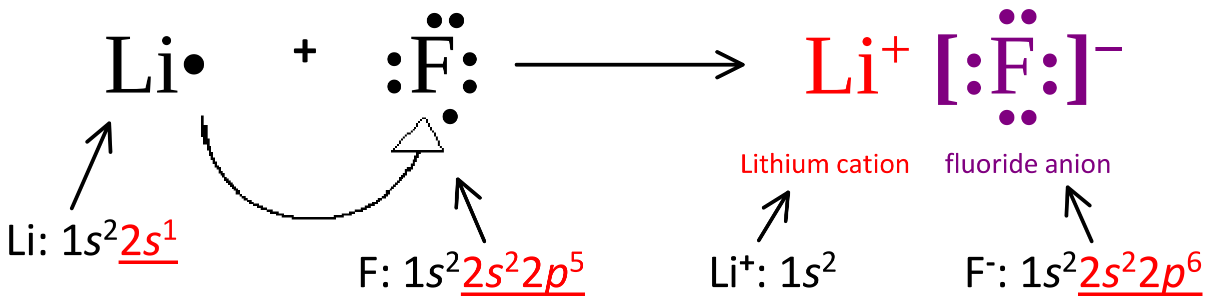
- **Ion formation occurs when atoms of two elements (a metal and nonmetal) are present.**

The metal will lose one or more electrons - forms a cation

The nonmetal will gain one or more electrons - forms an anion



## Formation of Ionic Compounds




Show how the following ionic compounds form using Lewis Electron Dot Symbols.

a) $Na_2S$

b) $K_3P$

c) $MgBr_2$

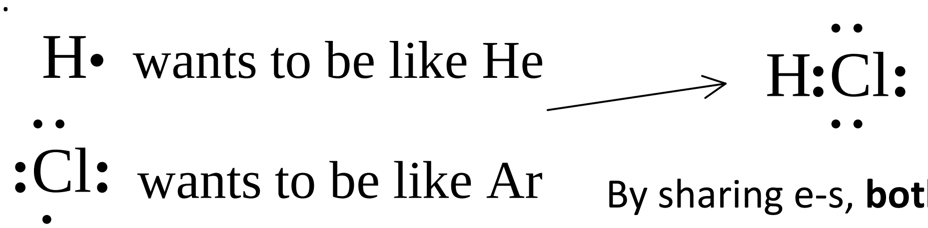
- $Li^+$ has a full valence shell - it has an electron configuration **just like He**
- $F^-$ has a full valence shell (a complete octet) - it has an electron configuration **just like Ne**

## Covalent Bonds

Covalent bonding results from the SHARING of one or more electron pairs between atoms

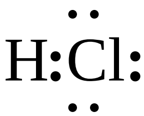
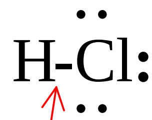
- Most nonmetal elements try to achieve a NOBLE GAS CONFIGURATION by sharing electrons with other nonmetals

Example:




By sharing e-s, **both** attain a noble gas configuration

# Covalent Lewis Structures

 can also be written as 

Nonbonding $e^-$ pair

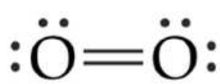
The HCl molecule has
1 bonding pair of electrons
and 3 nonbonding pairs of electrons (also called "lone pairs")

In Lewis theory:

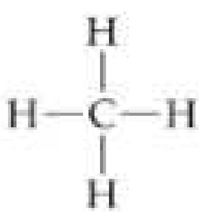
- Nonmetal atoms share electrons to complete their octet, called:

  **"OCTET RULE"** (Note: there are exceptions to the octet rule)

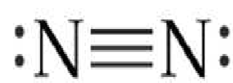
- Completing octets may involve sharing electrons with multiple atoms or sharing multiple pairs of electrons with the same atom.



Double covalent bond
Atoms share **two $e^-$ pairs**



Single covalent bonds
Atoms share **one pair** of electrons
(carbon sharing $e^-$s with **multiple** hydrogen atoms)



Triple covalent bond
Atoms share **three $e^-$ pairs**

## Rules for Writing Lewis Structures for Covalent Compounds

1. Decide on the central atom (it will <u>never</u> be H or F).
   *(The central atom is usually the one that is by itself)*

2. Determine the **<u>total number of valence electrons</u>** in the structure.
   - If the structure is an ion:
     - Add $1e^-$ for a -1 charge, add $2e^-$ for a -2 charge, etc...
     - Subtract $1e^-$ for a +1 charge, subtract $2e^-$ for +2, etc…

3. Form covalent bonds between the central atom and the surrounding atoms - called the "skeletal structure". Count how many electrons have been used to form these bonds.

4. <u>Subtract</u> electrons used to form covalent bonds <u>from total number of valence electrons</u> in the molecule or ion to determine how many electrons remain (if any).

5. Any remaining electrons become **<u>lone pairs</u>**, FIRST ON THE OUTSIDE ATOMS to complete their octets, and then on the central atom.

6. If any atoms that need an octet of electrons do not have it, form double and triple bonds as necessary by bringing <u>outer atom</u> lone pair electrons down between two atoms so they can share them.

   <u>Important: **C, N, O and F always follow the octet rule**</u>

## Write the Lewis structure for $NH_3$

1. Central atom?
2. Valence electrons?
3. Form skeletal structure
4. Remaining electrons?
5. Lewis Structure?

## Writing Lewis Structures

# C, N, O and F

ALWAYS have 8 electrons surrounding them in their Lewis structures. They always follow the octet rule.

NOTE: Boron may have less than $8e^-$s (usually 6). Phosphorus and atoms of higher atomic number have AT LEAST $8e^-$s but sometimes can have more that $8e^-$s in Lewis structures. These are elements that can be exceptions to the octet rule.

**We will not write Lewis structures of this type.**

Write the Lewis Structures for the following molecules and polyatomic ions.

a) $CH_4$

c) $O_2$

b) $CH_2O$

d) $CO_2$

## Resonance Structures

Write the Lewis Structures for the following molecules and polyatomic ions (continued).

e) HCN f) $SO_3^{2-}$ g) $NH_2Br_2^+$

## **Resonance Structures**

Sometimes we can draw more than one valid Lewis structure for a molecule or polyatomic ion.

- In this situation, no **one** Lewis structure can adequately describe the actual structure of the molecule.
- The actual molecule or ion will have characteristics of **all the valid Lewis structures** that can be drawn. (It is a hybrid of these Lewis structures).

Write the Lewis structures of:

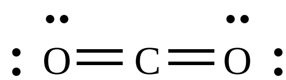
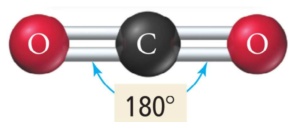
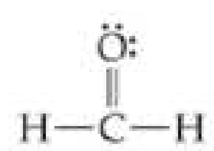
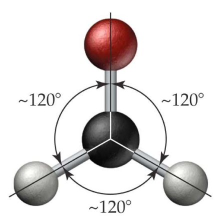
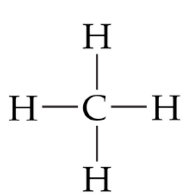
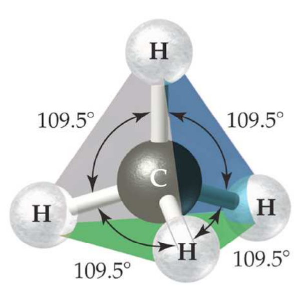
a) $SO_2$ b) $HCO_2^-$

# Shapes of Molecules

The most important factor in determining the shape of a molecule or polyatomic ion is the **relative repulsion between electron pairs.**

- A molecule or ion will be <u>most stable</u> when the electron pairs or groups are <u>as far apart as possible</u>.

(Molecular Geometry = Shape of Molecule)

| <u>Number of Electron "Groups" around the Central Atom</u> | <u>Example</u> | <u>Electron Geometry & Molecular Geometry</u> |
|---|---|---|
| 2 |    | Linear |
| 3 |    | Trigonal Planar |
| 4 |    | Tetrahedral |

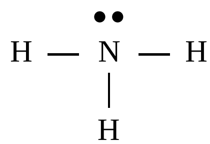
# Determining Molecular Geometry Using VSEPR Theory

Valence Shell Electron Pair Repulsion (**VSEPR**) Theory: Repulsion between the negative charges on electron groups determines the molecular geometry (shape) of a molecule or polyatomic ion

Determine the Electron and Molecular Geometry of:

a) $NH_3$

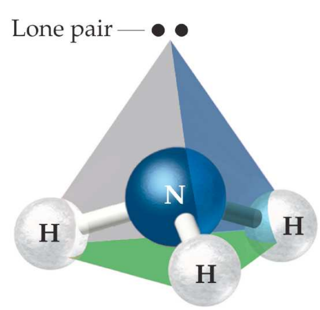
1st - Determine the Lewis structure of $NH_3$




2nd - Count electron "Groups" around the central atom 4

3rd - Determine Electron Geometry then Molecular Geometry

Four electron groups are at the corners of a tetrahedron




Electron Geometry

Tetrahedral

Molecular Geometry

Trigonal Pyramidal

b) $H_2O$

# Determining Molecular Geometry

Lewis Structure of $H_2O$

••
H — O — H
••

4 Groups
of Electrons

Electron Pair Geometry = Tetrahedral
Molecular Geometry = Bent

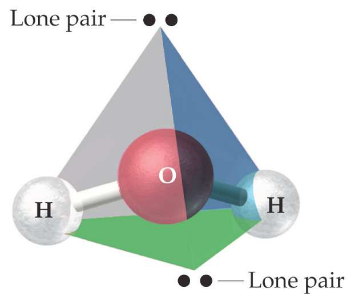
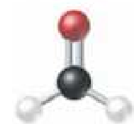



**TABLE 10.1** Electron and Molecular Geometries

| Electron Groups* | Bonding Groups | Lone Pairs | Electron Geometry | Angle between Electron Groups** | Molecular Geometry | Example | |
|---|---|---|---|---|---|---|---|
| 2 | 2 | 0 | linear | 180° | linear | :Ö=C=Ö: |  |
| 3 | 3 | 0 | trigonal planar | 120° | trigonal planar | H—C(=Ö:)—H |  |
| 3 | 2 | 1 | trigonal planar | 120° | bent | :Ö=S̈—Ö: |  |
| 4 | 4 | 0 | tetrahedral | 109.5° | tetrahedral | H—C(—H)(—H)—H |  |
| 4 | 3 | 1 | tetrahedral | 109.5° | trigonal pyramidal | H—N̈(—H)—H |  |
| 4 | 2 | 2 | tetrahedral | 109.5° | bent | H—Ö—H |  |

*Count only electron groups around the *central* atom. Each of the following is considered one electron group: a lone pair, a single bond, a double bond, and a triple bond.

**Angles listed here are idealized. Actual angles in specific molecules may vary by several degrees.

Determine the Electron and Molecular Geometry of:

a) $CCl_4$

b) HCN

# Bond Polarity

Determine the Electron and Molecular Geometry of:

c) $CH_2S$ d) $SO_2$

Practice for next class:

e) $H_2S$ f) $PH_3$

## <u>Bond Polarity</u>

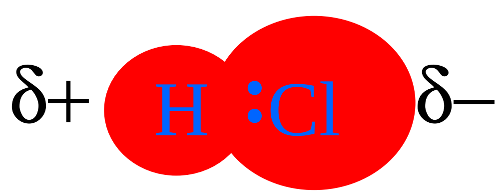
Bonding between **unlike atoms** results in unequal sharing of the electrons.

- One atom pulls the electrons in the bond closer to its side.
- One end of the bond has larger electron density than the other. The result is a <u>**POLAR BOND**</u>

The end with the larger electron density gets a **partial negative charge (δ-)** and the end that is electron deficient gets a **partial positive charge (δ+)**.

Example:

HCl



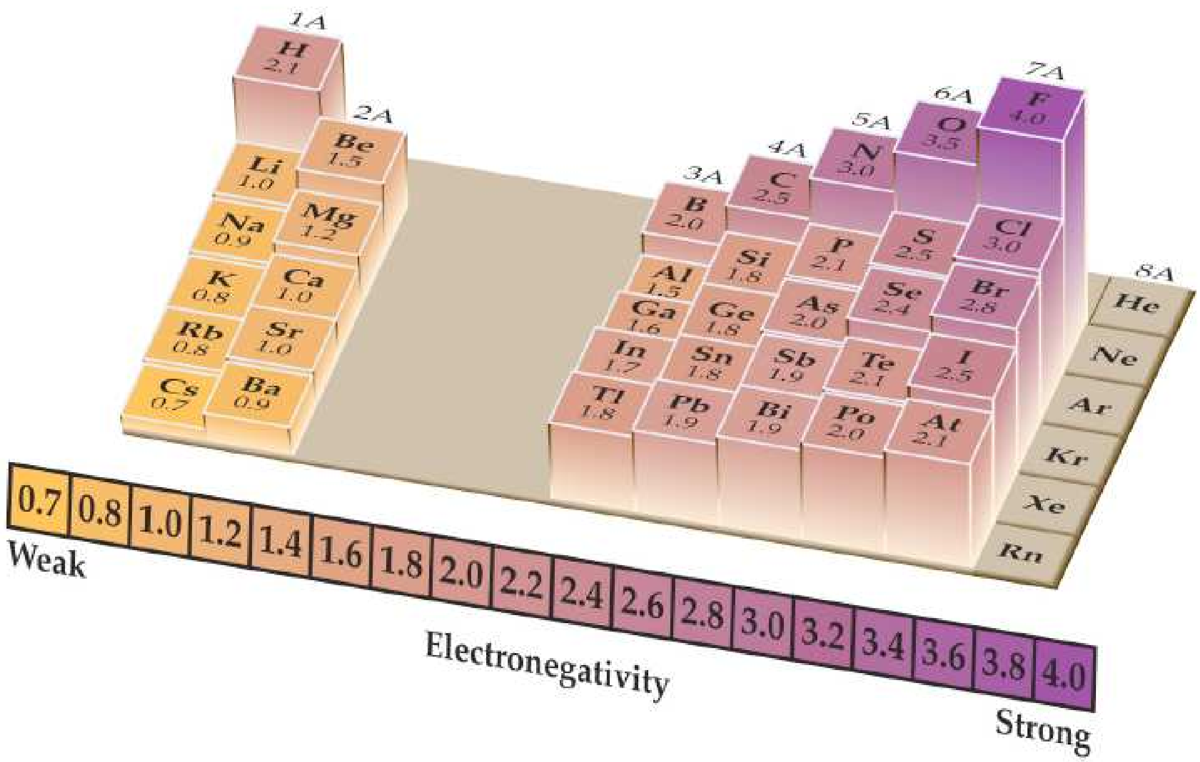
## Electronegativity

How can we determine which atom is δ+ **and which is δ- in a polar bond?**

- ➢ **Use electronegativity values of the atoms**

## <u>Electronegativity</u>

Electronegativity is a measure of relative attraction that an atom has for the shared electrons in a covalent bond




Electronegativity

- Increases across the period (left to right)
- Decreases down the group (top to bottom)

**The larger the difference in electronegativity, the more polar the bond.**

# Main Classes of Chemical Bonds

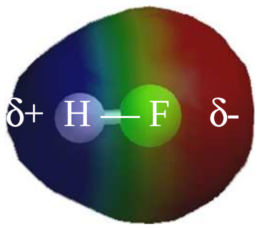
## Nonpolar Covalent Bond

- Electronegativity difference between zero and 0.4
- Many times between two identical atoms

Example: H—H

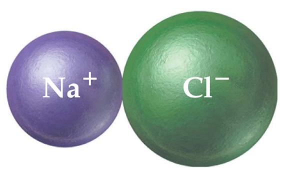
## Polar Covalent Bond

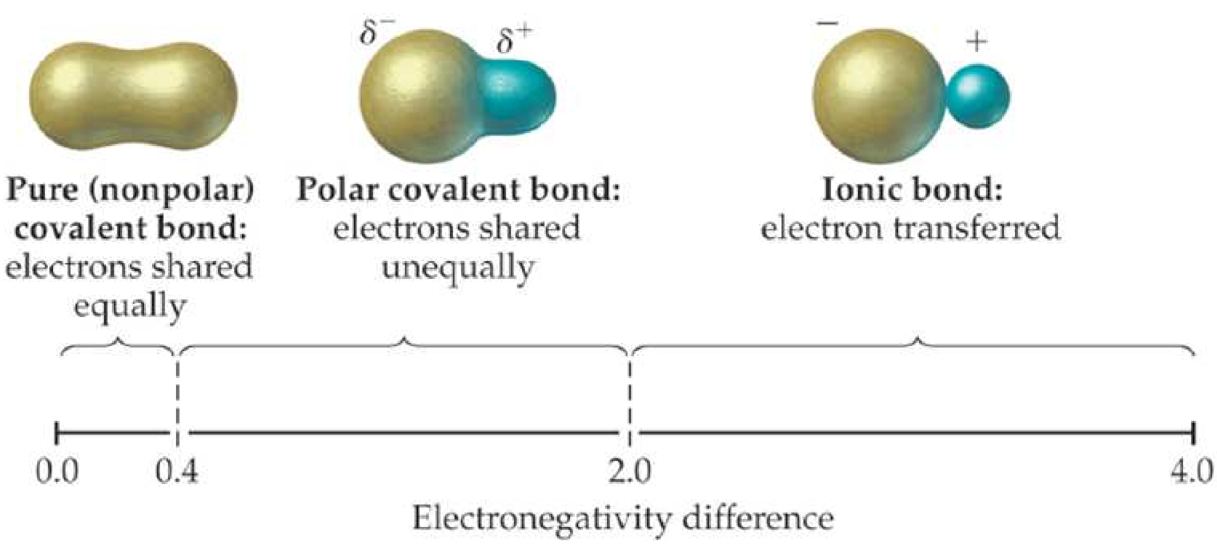
- Electronegativity difference between 0.4 and 2
- Between two different NONMETAL atoms




## Ionic Bond

- Electronegativity difference is greater than 2
- Primarily exists between METALS and NONMETALS



# Polar Molecules

Can show the direction of <u>bond</u> polarity with $\delta$+ and $\delta$- and/or a special arrow:


Show the direction of bond polarity for the bond in HCl.

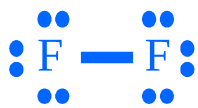
In order for a **<u>MOLECULE to be polar</u>**:

It <u>must</u> have polar bonds

<u>AND</u>

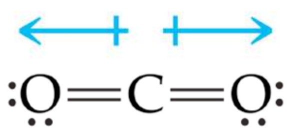
It <u>must</u> have an **unsymmetrical** shape

If there are no polar bonds, then molecule is NONPOLAR



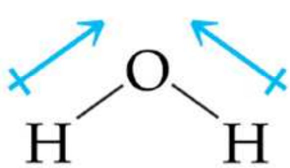
If there are polar bonds **<u>and</u>** the bond dipoles **cancel out**

- ➢ Molecule is NONPOLAR



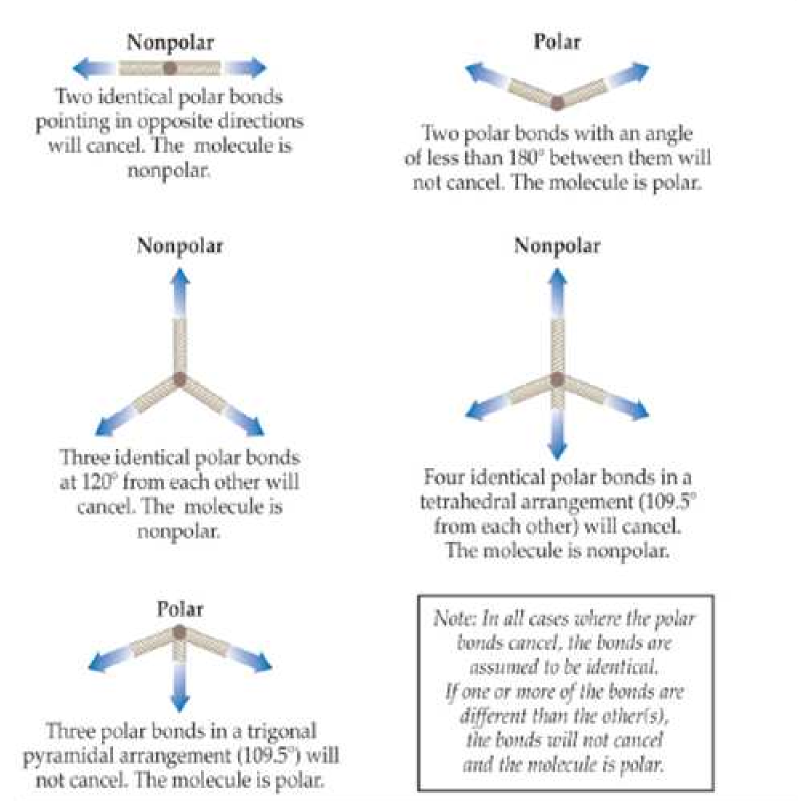
If there are polar bonds **<u>and</u>** the bond dipoles <u>DO NOT</u> **cancel out**,

- ➢ Molecule is POLAR



# Molecular Polarity

**TABLE 10.3** Common Cases of Adding Dipole Moments to Determine Whether a Molecule Is Polar




Determine whether the following molecules are POLAR or NONPOLAR.

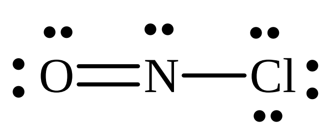
a) $CF_4$

b) $CH_2F_2$

c) $NH_3$

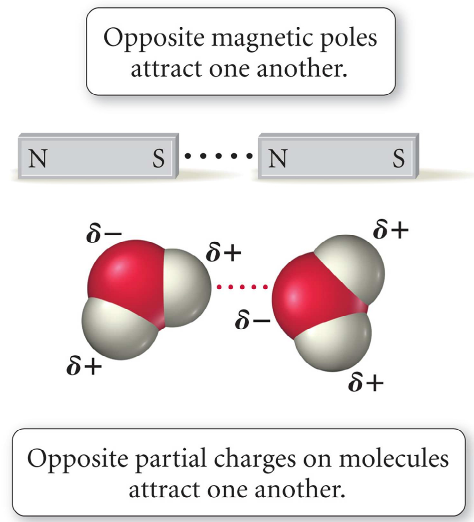
d) NOCl

Lewis Structure



# Molecular Polarity Affects Solubility

- Polar molecules are attracted to other polar molecules
- Since water is a polar molecule, other polar molecules dissolve well in water
  - (Many ionic compounds dissolve in water as well).
- Nonpolar molecules are attracted to other nonpolar molecules and dissolve in each other






"LIKE DISSOLVES LIKE"