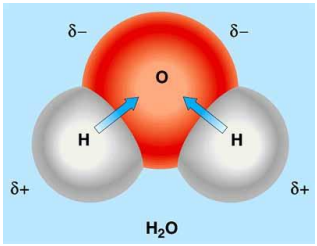
Polar & Non-polar covalent bonds

- Covalently bonded molecules are either **polar** or **non-polar**
- Unequal sharing of electrons results in a **Polar covalent bond**
  - Polar covalent bonds form when atoms pull on electrons in a molecule unequally
  - charges at the ends of the bonds called a **dipole**
  - A molecule is **POLAR** if: the central atom as a lone pair
  - Polar molecules align with an electric field
    - Example: Water
- **Non-polar molecules** are not attracted by an electric field

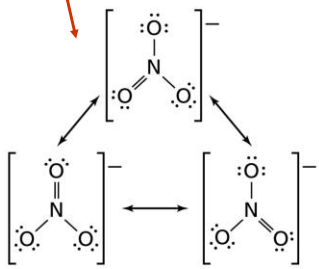


Polar Covalent Bonds

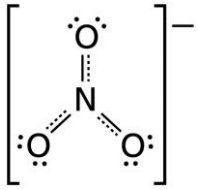
- **Solubility** is the property of a substance's ability to dissolve in another substance
- Polar molecules and ionic substances are usually soluble in polar substances
  - Water being polar allows it to dissolve ionic substances
- Non-polar molecules dissolve only in non-polar substances
- **LIKE DISSOLVES LIKE**

Resonance Structures

This figure shows three correct ways to draw the structure for $(NO_3)^{-1}$



- Resonance is a condition that occurs when more than one valid Lewis structure can be written for a molecule or ion
- The molecule behaves as though it has only one structure



Electron Affinity, Electronegativity, and Bond Character

- **Electron affinity** measure the tendency of an atom to accept an electron
  - Noble gases are not listed because they do not form compounds

Electronegativity Values for Selected Elements

1 H 2.20																	
3 Li 0.98	4 Be 1.57											5 B 2.04	6 C 2.55	7 N 3.04	8 O 3.44	9 F 3.98	
11 Na 0.93	12 Mg 1.31											13 Al 1.61	14 Si 1.90	15 P 2.19	16 S 2.58	17 Cl 3.16	
19 K 0.82	20 Ca 1.00	21 Sc 1.36	22 Ti 1.54	23 V 1.63	24 Cr 1.66	25 Mn 1.55	26 Fe 1.83	27 Co 1.88	28 Ni 1.91	29 Cu 1.90	30 Zn 1.65	31 Ga 1.81	32 Ge 2.01	33 As 2.18	34 Se 2.55	35 Br 2.96	
37 Rb 0.82	38 Sr 0.95	39 Y 1.22	40 Zr 1.33	41 Nb 1.6	42 Mo 2.16	43 Tc 2.10	44 Ru 2.2	45 Rh 2.28	46 Pd 2.20	47 Ag 1.93	48 Cd 1.69	49 In 1.78	50 Sn 1.96	51 Sb 2.05	52 Te 2.1	53 I 2.66	
55 Cs 0.79	56 Ba 0.89	57 La 1.10	72 Hf 1.3	73 Ta 1.5	74 W 1.7	75 Re 1.9	76 Os 2.2	77 Ir 2.2	78 Pt 2.2	79 Au 2.4	80 Hg 1.9	81 Tl 1.8	82 Pb 1.8	83 Bi 1.9	84 Po 2.0	85 At 2.2	
87 Fr 0.7	88 Ra 0.9	89 Ac 1.1															

Metal
Metalloid
Nonmetal

Page 265 in your book

Determining Bond Character with Electronegativity

- The character of a bond depends on how strongly each of the bonded atoms attracts electrons
  - Bonding is often not clearly ionic or covalent
- Table 8.7 lists the type of chemical bond that forms from with differences in electronegativity
- You can **subtract the electronegativities** of the elements in the compound to determine the **character and bond type**

H_20

Electronegativity O = 3.44

Electronegativity H = 2.20

Difference = **1.24**

Polar covalent !

Table 8.7 EN Difference and Bond Character	
Electronegativity Difference	Bond Character
> 1.7	mostly ionic
0.4 – 1.7	polar covalent
< 0.4	mostly covalent
0	nonpolar covalent

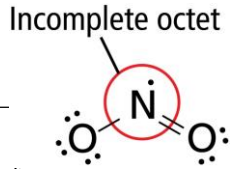
Since Oxygen electronegativity is higher than that of hydrogen, the shared pair of electrons are closer to Oxygen. Making the Oxygen end of the compound more negative

Review: Main Idea's

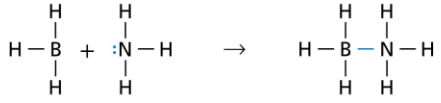
- Specific rules are used when naming binary molecular compounds and binary acids
- Atoms gain stability when they share electrons and form covalent bonds
- Structural formulas show the relative positions of atoms within a molecule
- The VSEPR model is used to determine molecular shape
- Electronegativity: the relative ability of an atom to attract electrons in a chemical bond
- A chemical bond's character is related to each atom's attraction for the electrons in the bond

Extra Information follows

Exceptions to Octet Rules



- Some molecules do not obey the octet rule
- A small group of molecules might have an odd number of valence electrons.
- NO_2 has five valence electrons from nitrogen and 12 from oxygen and cannot form an exact number of electron pairs
- A few compounds form stable configurations with less than 8 electrons around the atom—a suboctet.
- A **coordinate covalent bond** forms when one atom donates both of the electrons to be shared with an atom or ion that needs two electrons.

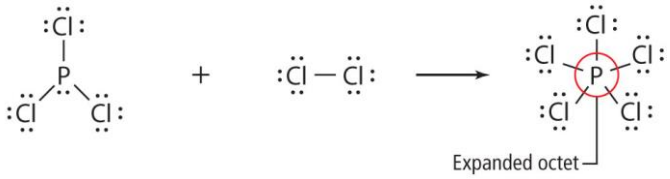


The boron atom has no electrons to share, whereas the nitrogen atom has two electrons to share.

The nitrogen atom shares both electrons to form the coordinate covalent bond.

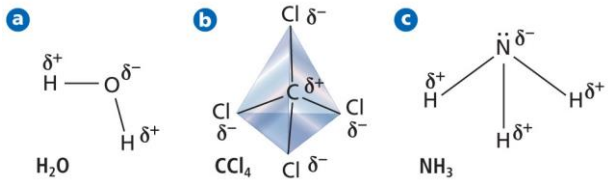
Exceptions to the Octet Rule (cont)

- A third group of compounds has central atoms with more than eight valence electrons, called an expanded octet
- Elements in period 3 or higher have a d-orbital and can form more than four covalent bonds



Polar Covalent bonds

- Compare Water and CCl_4
  - Both are polar by subtracting electronegativities
  - But only water is a polar molecule because of the shape of the molecule



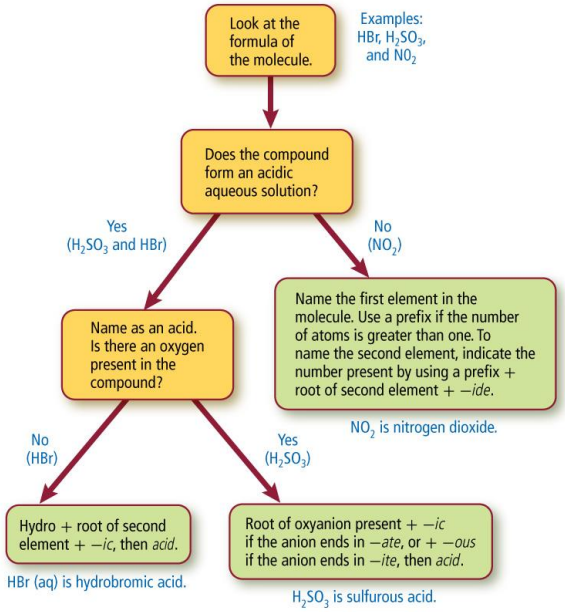
The bent shape of a water molecule makes it polar.

The symmetry of a CCl_4 molecule results in an equal distribution of charge, and the molecule is nonpolar.

The asymmetric shape of an ammonia molecule results in an unequal charge distribution and the molecule is polar.

Properties of Covalent Compounds

- Covalent bonds between atoms are strong, but attraction forces between molecules are weak.
- The weak attraction forces are known as van der Waals forces.
- The forces vary in strength but are weaker than the bonds in a molecule or ions in an ionic compound
- Non-polar molecules exhibit a weak dispersion force, or induced dipole. The force between two oppositely charged ends of two polar molecules is a dipole-dipole force.
- A hydrogen bond is an especially strong dipole-dipole force between a hydrogen end of one dipole and a fluorine, oxygen, or nitrogen atom on another dipole.
- Weak forces result in the relatively low melting and boiling points of molecular substances.
- Many covalent molecules are relatively soft solids.



Hybridization

- **Hybridization** is a process in which atomic orbitals mix and form new, identical hybrid orbitals
- Carbon often undergoes hybridization, which forms an **sp3 orbital** formed from one s orbital and three p orbitals
  - Carbon $1s^2\ 2s^2\ 2p^2$
  - Hybridization Carbon $1s^2\ 2s^1p^3$
  - Lone electrons will occupy hybrid orbitals
- Allows Single, double, and triple bonds to occur in one hybrid orbital