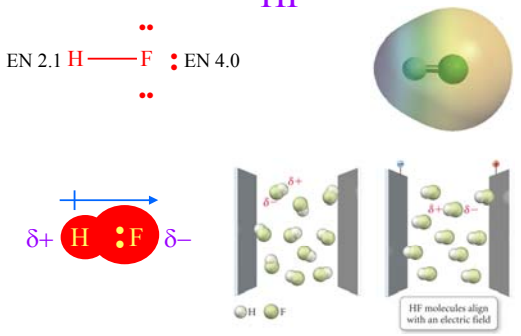
Bond Polarity

- covalent 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 **polar covalent bond**
  - ✓bond polarity
  - ✓the end with the larger electron density gets a partial negative charge
  - ✓the end that is electron deficient gets a partial positive charge

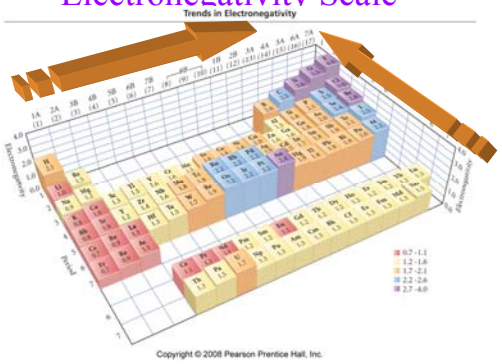
HF



Electronegativity

- measure of the pull an atom has on bonding electrons
- increases across period (left to right) and
- decreases down group (top to bottom)
  - ✓fluorine is the most electronegative element
  - ✓francium is the least electronegative element
- the larger the difference in electronegativity, the more polar the bond
  - ✓negative end toward more electronegative atom

Electronegativity Scale



Electronegativity and Bond Polarity

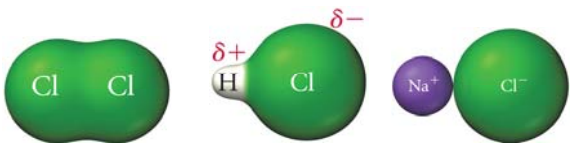
- If difference in electronegativity between bonded atoms is 0, the bond is **pure covalent**
  - ✓ equal sharing
- If difference in electronegativity between bonded atoms is 0.1 to 0.4, the bond is **nonpolar covalent**
- If difference in electronegativity between bonded atoms 0.5 to 1.9, the bond is **polar covalent**
- If difference in electronegativity between bonded atoms larger than or equal to 2.0, the bond is **ionic**

Percent Ionic Character

4%	51%	"100%"
0 0.4	2.0	4.0

Electronegativity Difference

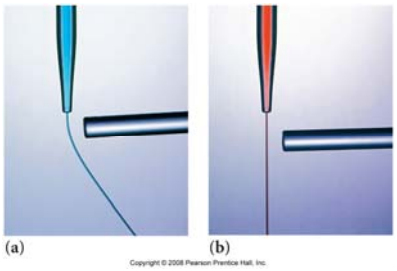
Bond Polarity



$EN_{Cl} = 3.0$	$EN_{Cl} = 3.0$	$EN_{Cl} = 3.0$
3.0 - 3.0 = 0	$EN_H = 2.1$	$EN_{Na} = 1.0$
Pure Covalent	3.0 – 2.1 = 0.9	3.0 – 0.9 = 2.1
	Polar Covalent	Ionic

Water – a Polar Molecule

stream of water attracted to a charged glass rod



stream of hexane not attracted to a charged glass rod

Example 9.3(c) - Determine whether an N-O bond is ionic, covalent, or polar covalent.

Lewis Structures of Molecules

- shows pattern of valence electron distribution in the molecule
- useful for understanding the bonding in many compounds
- allows us to predict shapes of molecules
- allows us to predict properties of molecules and how they will interact together

Lewis Structures

- use common bonding patterns
  - ✓ C = 4 bonds & 0 lone pairs, N = 3 bonds & 1 lone pair, O= 2 bonds & 2 lone pairs, H and halogen = 1 bond, Be = 2 bonds & 0 lone pairs, B = 3 bonds & 0 lone pairs
  - ✓ often Lewis structures with line bonds have the lone pairs left off
    - ➢ their presence is assumed from common bonding patterns
- structures which result in bonding patterns different from common have formal charges

Writing Lewis Structures of Molecules HNO_3

1) Write skeletal structure
   - ✓ H always terminal
     - ➢ in oxyacid, H outside attached to O's
   - ✓ make least electronegative atom central
     - ➢ N is central

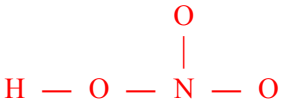
H O N O
O

2) Count valence electrons
   - ✓ sum the valence electrons for each atom
   - ✓ add 1 electron for each − charge
   - ✓ subtract 1 electron for each + charge

N = 5
H = 1
$O_3 = 3 \cdot 6 = 18$
Total = 24 e^-

Writing Lewis Structures of Molecules HNO_3

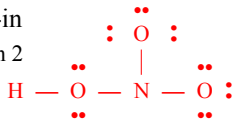
3) Attach central atom to the surrounding atoms with pairs of electrons and subtract from the total



	Electrons
Start	24
Used	8
Left	16

Writing Lewis Structures of Molecules
HNO_3

4) Complete octets, outside-in
   - ✓ H is already complete with 2
     - ➢ 1 bond

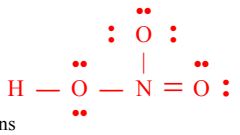


and re-count electrons

		Electrons		Electrons
N = 5				
H = 1	Start	24	Start	16
$O_3 = 3 \cdot 6 = 18$	Used	8	Used	16
Total = 24 e^-	Left	16	Left	0

Writing Lewis Structures of Molecules
HNO_3

5) If all octets complete, give extra electrons to central atom.
   - ✓ elements with *d* orbitals can have more than 8 electrons
     - ➢ Period 3 and below
6) If central atom does not have octet, bring in electrons from outside atoms to share
   - ✓ follow common bonding patterns if possible



Practice - Lewis Structures

- CO_2
- H_3PO_4
- $SeOF_2$
- SO_3^{-2}
- NO_2^{-1}
- P_2H_4

Formal Charge

- during bonding, atoms may wind up with more or less electrons in order to fulfill octets - this results in atoms having a **formal charge**

FC = valence e^- - nonbonding e^- - ½ bonding e^-

left O FC = 6 - 4 - ½ (4) = 0

S FC = 6 - 2 - ½ (6) = +1

right O FC = 6 - 6 - ½ (2) = -1

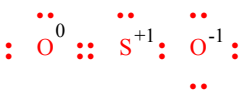
- sum of all the formal charges in a molecule = 0
  - ✓ in an ion, total equals the charge

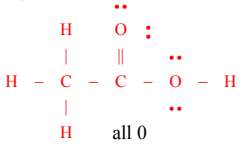
Writing Lewis Formulas of Molecules (cont'd)

7) Assign formal charges to the atoms

a) formal charge = valence e^- - lone pair e^- - ½ bonding e^-

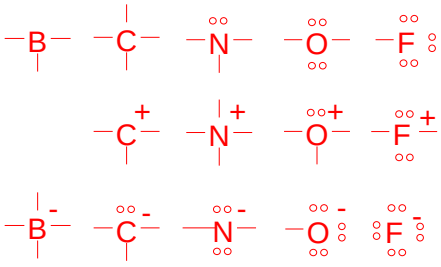
b) follow the common bonding patterns





all 0

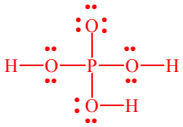
Common Bonding Patterns



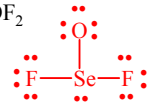
Practice - Assign Formal Charges

- CO_2

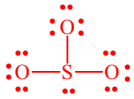
- H_3PO_4



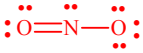
- $SeOF_2$



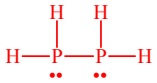
- SO_3^{-2}



- NO_2^{-1}



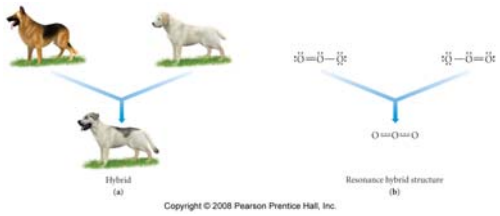
- P_2H_4



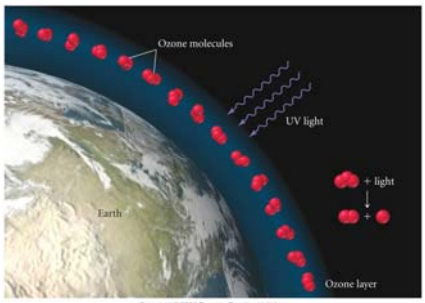
Resonance

- when there is more than one Lewis structure for a molecule that differ **only** in the position of the electrons, they are called **resonance structures**
- the actual molecule is a combination of the resonance forms – a **resonance hybrid**
  - ✓it does **not** resonate between the two forms, though we often draw it that way
- look for multiple bonds or lone pairs

Resonance



Ozone Layer



Exceptions to the Octet Rule

- expanded octets
  - ✓elements with empty *d* orbitals can have more than 8 electrons
- odd number electron species e.g., NO
  - ✓will have 1 unpaired electron
  - ✓free-radical
  - ✓very reactive
- incomplete octets
  - ✓B, Al