9.6 Electronegativity and Polar Bonds

Consider a covalent bond between two atoms: **X–Y**:

- The electrons in a covalent bond are not necessarily shared evenly between the two atoms.
- The **Electronegativity** of an element is a measure its ability to attract the shared electrons in a covalent bond.
- Pauling scale of EN: Relative to F = 4.0
- Electronegativity increases with increasing $Z_{\text{effective}}$, so electronegativity increases as you go up a group or left to right across a period (Figure 9.8).
- The greater the electronegativity difference (ΔEN) in a bond, the **more uneven the electron sharing**.
- If a bond is between elements with different electronegativity (ΔEN), the electrons are shared unevenly. This is called a **polar covalent bond**.
- Polar bonds have both *covalent* and *ionic* character.

Pauling Scale: #1 F #2 O #N, Cl

δ^+ δ^- ← partial charge figure 9.6

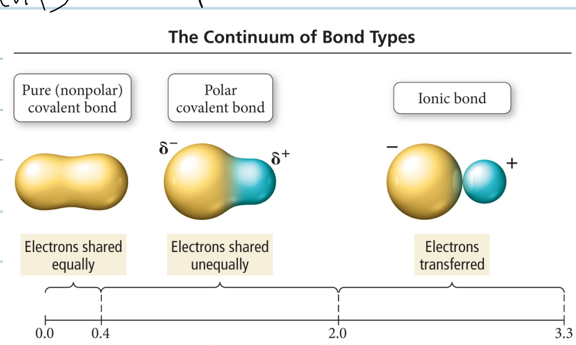
H – F

less EN more EN

δ^+ δ^- δ^+

H–H H – O – H $\overset{+}{\text{Na}} - \overset{-}{\text{Cl}}$

nonpolar bond (pure covalent) polar covalent ionic



9.7 Lewis Structures of Molecules and Polyatomic Ions

Guidelines for drawing Lewis structures:

1) Arrange the atoms with the correct connectivity:
   - The least electronegative element is usually **central**
   - H and halogens are generally **terminal**
   - Symmetrical structures preferred
2) Calculate the total number of valence electrons in the molecule (for polyatomic ions, add or subtract electrons due to charge). A Lewis structure **must** contain this number of VE, in bonds or lone pairs.
3) Join atoms with single bonds (uses 2 e^- /bond).
4) Add remaining VE as lone pairs to complete the octets of all the atoms (except H).
5) If there are not enough VE to give each atom an octet, convert lone pairs to double or triple bonds.

O=C=O

N_2O

N–N–O

- Try to follow the "common bonding patterns":
  - C = 4 bonds & 0 lone pairs
  - N = 3 bonds & 1 lone pair,
  - O = 2 bonds & 2 lone pairs,
  - H = 1 bond, no lone pairs (always)
  - halogens = 1 bond, 3 lone pairs
  - B = 3 bonds & 0 lone pairs

Lewis structure examples:

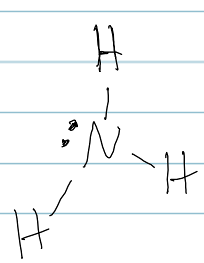
1. NH_3 Valence electrons: N - 5

H - 1 x 3

8 VE

-6 (single bonds)

2 VE

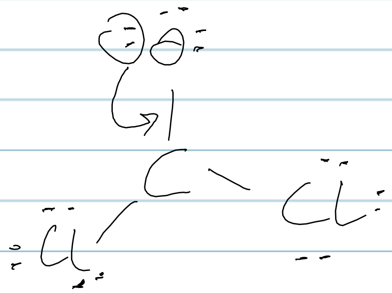


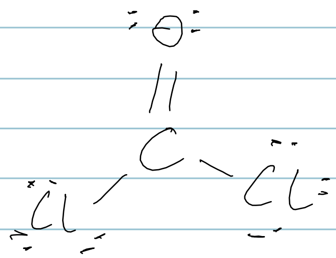
2. $COCl_2$ (phosgene)

V.E:

C – 4
O – 6
Cl – 7 ×2

24 VE

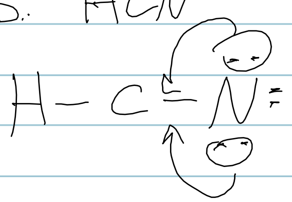




3. HCN

C – 4
N – 5
H – 1

10 VE



$H-C\equiv N:$

4. NO_3^- nitrate ion

N – 5
O – 6 × 3
(–) charge 1

24 VE