

Covalent compounds

CHM 1020

Octet rule: Atoms in a compound will lose, gain or share electrons in order to achieve a stable noble gas configuration.

- Covalent bonds are between two non-metals.
- This bond involves sharing electrons.
- Electrons will exist in pairs.
- Atoms will share electrons to get 8 electrons in their outer shell. (Except Hydrogen)
- Bonding electrons: shared pairs of electrons.
- Lone pairs: unshared pairs of electrons.

Examining Lewis dot structures and Lewis structures

Hydrogen

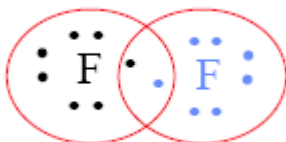
- Each hydrogen begins with one electron.
- Hydrogen needs two to get the NG configuration of He ($2e^-$)



The two atoms are bound together mainly due to the attraction of the positively charged nuclei for the negatively charged electron cloud located between them

Fluorine

- Each F begins with 7 valence electrons.
- Needs to share one more with the other F.



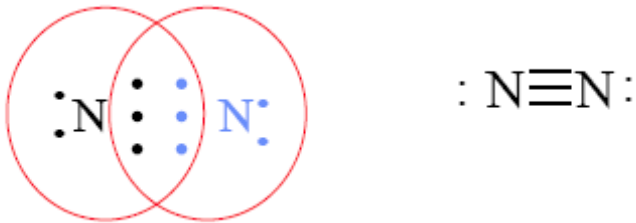
Oxygen, O₂

- Each oxygen begins with 6 electrons.
- Needs to share two more electrons.
- An example of a double bond



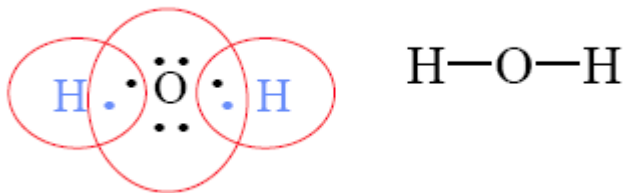
Nitrogen, N₂

- Each nitrogen begins with 5 electrons.
- Must share three more electrons.
- An example of a triple bond



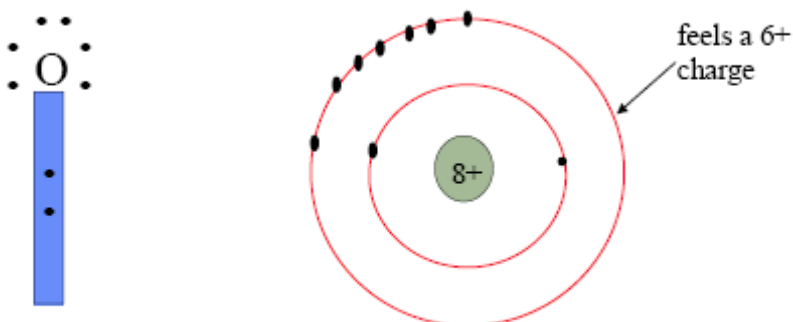
Water, H₂O

- Each H has 1 e⁻ and needs to have 1 more.
- O has 6 and needs to share 2 more



Carbonate, the Will-Nilly method (CO₃²⁻)

Let's examine one the $-O$



The outer shell feels a $6+$ charge and has 7 electrons

Formal Charge

- $FC = \text{Group\#} - \# \text{ of bonds} - e^- \text{ in lone pairs}$
- For that oxygen $FC = 6 - 1 - 6 = -1$
- Include the formal charge for every atom in a Lewis structure.
- The sum of the individual formal charges equals the overall charge.

The computer method

- Find total valence electrons.
- Create a framework using single bonds.
- Add lone pairs from the outside in to give atoms octets.
- Make multiple bonds if necessary. (Minimizing formal charge)
- Add formal charge.

The central atom

- The first non-hydrogen atom is usually the central atom.
- Hydrogen is always terminal and so never can be the central atom.
- Carbon is always a central atom.
- There is often more than one central atom.
- If hydrogen is the first atom in the formula, it is usually acidic and attached to oxygen.

Carbonate, the computer method

HCN, the computer method

Acetate, CH_3CO_2^-

Shapes of molecules, VSEPR

Based on 2 simple assumptions.

- electrons repel
- areas of electron density repel equally
 - Areas of electron density
 - single bond
 - double bond
 - triple bond
 - lone pair

In this class we are only concerned with the geometry around the central atom.

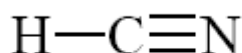
Types of shapes

Area of electron Density	Shape	Bond Angles
2	linear	180°
3	Trigonal planar	120°
4	tetrahedral	109.5°

Linear

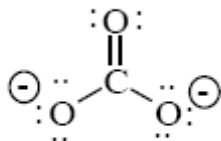


Lewis

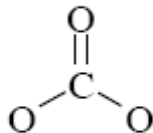


VSEPR

Trigonal planar

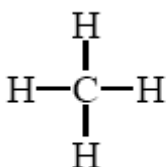


Lewis structure

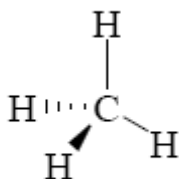


VSEPR

Tetrahedral

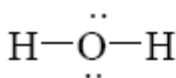


Lewis structure

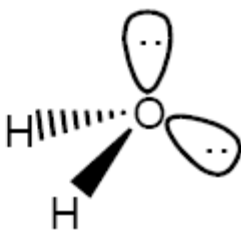


VSEPR diagram

Water is also tetrahedral



Lewis



VSEPR

Naming simple covalent compounds

1. Name the first non-metal by its elemental name.
2. Name the second nonmetal by its elemental name and an -ide ending.
3. Use the prefixes mono, di, tri, tetra, penta and hexa to denote number of atoms of that element in the molecule.
4. If mono is the first prefix it is understood and not written.

Examples

- CO

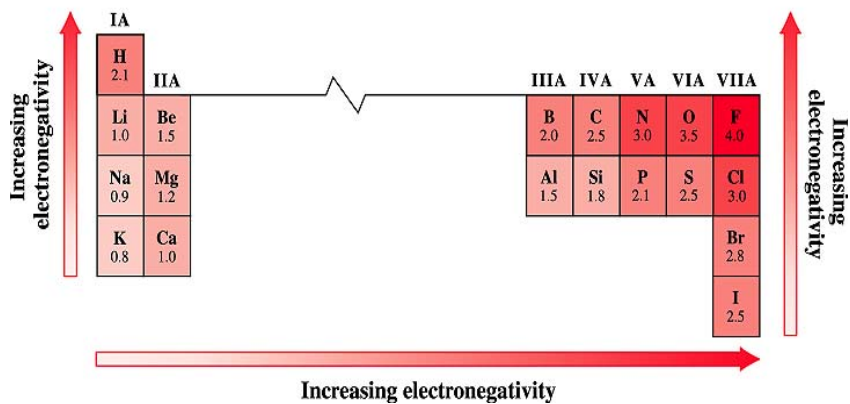
- CO_2
- P_2O_5
- NO
- N_2O
- NO_2

Bond Polarity

Electrons in bonds are not always shared equally

- Electronegativity is the atom's tendency to attract the *bonding* electrons to itself.
- Bonds in which the electrons are not shared equally have a partial charge distribution and are called polar bonds.
- A polar bond has a dipole.
- Concept proposed by Linus Pauling (1901-1994)

Pauling Scale of Electronegativities



Electronegativity Difference	Bond Type
0-0.4	Non-polar covalent
0.5-1.9	Polar Covalent
Metal + nonmetal (most greater than 1.9)	Ionic

C_6H_{14} has no polar bonds.