Covalent Bonding Notes

Sharing is Caring!
Don't be IONIC, be COVALENT!
Ionic vs Covalent Bonding

- **Ionic**: electron(s) leave one atom & gained by another atom to satisfy both atoms’ octets, this results in the formation of ions. The resulting opposite charges attract each other.
- **Covalent**: electrons are shared by two or more atoms to satisfy their octets.
How can you tell if a bond is IONIC or COVALENT?

• Subtract the two electronegativity values (look at an electronegativity chart p. 263).

  >1.7 to 4.0: Ionic
  >0.4 to 1.7: Polar Covalent
  0.0 to 0.4: Non-Polar Covalent

• Electronegativity: atom’s ability to attract electrons in a chemical bond. (higher electronegativity means the atom wants electrons more)
How can you tell if a bond is IONIC or COVALENT?

• Easy way:

All metals = metallic bond
Nonmetals and Metals = ionic bond
All nonmetals = covalent bond
Yes, there are 2 kinds of covalent bonds!

• **Polar covalent:** the electrons are shared, but one atom is pulling on the electrons a lot more. The electrons spend more time around that atom.

• **Nonpolar covalent:** the electrons are evenly shared between the two atoms.
O(3.5) – H (2.1) = 1.4

Polar covalent bond

The electrons are unevenly shared between the O and H atoms, forming a polar covalent bond.
• The closer the elements are on the periodic table, their electronegativities are more similar... more likely to form covalent bonds
• Farther away... greater difference in electronegativity... more likely to form ionic bonds.
  • Metal + nonmetal = usually ionic
  • Nonmetal + nonmetal = usually covalent.
Properties of Covalent Compounds

1) Covalent compounds generally have much lower melting and boiling points than ionic compounds.

2) Covalent compounds are soft and squishy (compared to ionic compounds, anyway).
Properties, Cont’d

• 3) Covalent compounds tend to be more flammable than ionic compounds.
  – There are exceptions to this rule!

• 4) Covalent compounds don't conduct electricity in water.
Covalent Bonding and Molecular Compounds

• Molecule: is a neutral group of atoms that are held together by covalent bonds
• Molecular compound: a chemical compound whose simplest units are molecules
• Chemical formula indicate the relative numbers of atoms of each kind in a chemical compound by using atomic symbols and subscripts
• Molecular formula shows the types and numbers of atoms combined in a sing molecule of molecular compound
Diatomics

- **Diatomic elements** - containing only 2 atoms
- are elements that do not exist singularly in nature because they are highly reactive.

“Which elements are the diatomics?”

“HON, it’s the halogens!”

- \( \text{H}_2, \text{O}_2, \text{N}_2, \text{F}_2, \text{Cl}_2, \text{Br}_2, \text{I}_2 \)
Bond Dissociation Energy

- **Bond dissociation energy** = energy required to break a covalent bond.

- **Triple Bond**
  - Shortest bond length
  - Strongest
  - Hardest to break

- **Double Bond**
  - Longest bond length

- **Single Bond**
  - Easiest to break
LEWIS DOT STRUCTURES

1. Arrange the symbols such that the least electronegative element is in the center and the other elements are surrounding the central atom.

   O C O

2. Give each of the elements their appropriate number of valence electrons (dots). Remember the number of valence electrons for a representative element is the same as the group number.

   :O:   :C:   :O:
3. Keep track of the total number of valence electrons for the compound by adding the valence electrons from each atom. If the compound is an ion then add electrons (dots) for each negative charge or subtract electrons (dots) for each positive charge.

- 4 for C and 6 for O (twice) = 16 electrons
LEWIS DOT STRUCTURES

4. Now move the dots around so that you have 8 dots (the octet rule) around each element (do not forget the exceptions) while at the same time keeping the dots in pairs. Electrons, at this point, exist as pairs (the buddy system).

5. If there are too few pairs to give each atom eight electrons, change the single bonds between two atoms to either double or triple bonds by moving the unbonded pairs of electrons next to a bonding pair.

\[ \text{ :O : : C : : O : } \]
6. Once the octet rule has been satisfied for each atom in the molecule then you may replace each pair of dots between two atoms with a dash.

\[ \overset{\cdot}{\overset{\cdot}{O}} = \overset{\cdot}{\overset{\cdot}{C}} = \overset{\cdot}{\overset{\cdot}{O}} \]
• 7. Now check your structure by
   a) count the total number of electrons to make sure you did not lose or gain electrons during the process.
The Basics: Drawing Lewis Structures

Step 1: Calculate the total number of valence electrons in the molecule or ion
Step 2: Determine the central atom(s) of the molecule or ion – usually it’s the least electronegative atom.
Step 3: Draw a tentative diagram for the molecule or ion.

Rules

a) A hydrogen atom always forms one bond. Hydrogen is always a terminal atom in a Lewis diagram – an atom that is bonded to only one other atom.

b) A carbon atom normally forms four bonds

c) When several carbon atoms appear in the same molecule, they are often bonded to each other. (No cyclic compounds in Chem 60/68)
Draw the Lewis Structure for the following molecules.

1. \( \text{H}_2\text{O} \)

Oxygen has 6 valence electrons & Hydrogen has 1 valence electron for a total of 8 electrons.

\[
\text{H} - \text{O} - \text{H}
\]

2. \( \text{CO} \)

Oxygen has 6 valence electrons & Carbon has 4 valence electrons for a total of 10 electrons.

\[
:\text{C} : : : \text{O} :
\]
Draw the Lewis Structure for the following molecules.

3. BH$_3$
Boron has 3 valence electrons & Hydrogen has 1 valence electron for a total of 6 electrons

\[ \text{H} - \text{B} - \text{H} \]

4. NH$_3$
Nitrogen has 5 valence electrons & Hydrogen has 1 valence electron for a total of 8 electrons

\[ \text{H} - \text{N} - \text{H} \]
Cations

\[ \text{ammonia} \quad \rightarrow \quad \text{ammonium ion} \]
Anions

\[
\begin{array}{c}
\text{O} \quad \text{S} \quad \text{O} \\
\text{O} \\
\end{array}
\]

\[\text{2}^--\]

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Resonance

\[
\begin{align*}
.\cdot & \quad .\cdot \quad S = O \\
.\cdot & \quad .\cdot & \quad \text{or} & \quad .\cdot & \quad .\cdot \\
.\cdot & \quad .\cdot & \quad O = S \quad - \quad O \\
\end{align*}
\]