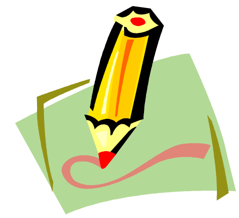
Name: ____________________________ Date:____________ **Class Notes**
Chemistry



Drawing Lewis Dot Structures

The drawing of Lewis dot symbols is a useful system to show the number of valence electrons in an atom. But an even better use for electron dot symbols is to join them together to form the Lewis structures of **molecules** or polyatomic ions.

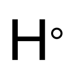
Lewis structures show the number and kinds of bonds, and the order in which the atoms or ions are connected in the molecule or polyatomic ion. However, their purpose is **not** to show the three dimensional shape of a molecule or polyatomic ion. Many of the Lewis structures are simple and can be determined by inspection. Others are a bit more complicated and require some calculating and thinking. It is the complicated structures that this handout will focus on.

SIMPLE COMPOUNDS

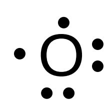
Many covalent compounds can be drawn by inspection using the valence electrons and the knowledge that covalent bonds are shared bonds.

Example: Draw the Lewis structure for H_2O

Step One: Determine the Lewis dot structures of the atoms and the number of bonds.

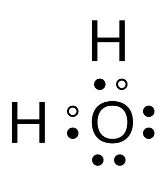


1 bond



2 bonds

Step Two: Then pair up the available electrons for the



Practice

Based on what you have learned about Lewis dot symbols and covalent bonds draw the Lewis structures for the following molecules.

A. CH_4 **B. NH_3** **C. HCl** **D. AlF_3**

Complex Compounds

Although, many molecules can be drawn by inspection others require the use of a few rules to help put them together. So, read the following rules carefully and think.

Step One Determine the "skeleton" for the molecule or polyatomic ion.

1. The least electronegative atom is the central atom, except hydrogen which is always a terminal atom.
2. Oxygen atoms do not bond with each other except in O_2 and O_3 molecules; peroxides; and super peroxides.
3. In oyxacids (ternary acids), hydrogen usually bonds to the oxygen instead of the central atom.
4. For those that have more than one central atom, the most symmetrical skeletons possible are used.

Step Two Calculate the number of electrons being shared (bonding electrons).

1. Determine the total number of electrons **needed** for each atom to complete its octet or duet (**N**).
2. Determine the total number of valence electrons already **available** (**A**). Remember to add electrons f or negative charges and subtract electrons positive charges.
3. Subtract the electrons available from the electrons needed to get the number or electrons **shared** (**S**).

$$\mathbf{S = N - A}$$

4. Divide the shared (S) by two for the number of bonds in the molecule or polyatomic ion.

S/2 = bond pairs

Step Three Place the bonding electrons in the skeleton as shared pairs.

1. Place one pair of electrons between each pair of bonded atoms.
2. If the central atom does not have a complete octet add double or triple bonds as needed.

NOT ALL ELEMENTS FORM MULTIPLE BONDS ONLY C, N, O, P, and S!!

Step Four Place the leftover electrons (A - S) in the skeleton as lone pairs.

1. Place lone pairs about each terminal atom to complete the octet rule.
2. Leftover electron pairs are placed on the central atom.
3. If the central atom is from the third or higher period, it can accommodate more than four electron pairs (expanded valance).

Example **Determine the Lewis symbol for CO_3^{2-}**

PRACTICE

Using the above rules to draw Lewis symbols for the following:

A. H_2SO_4

B. SF_4

C. ClO_2^-

D. PCl_5

RESONANCE

Empirical data shows that Ozone, O_3, has equal bond lengths, implying that there is an equal number of bond pairs on each side of the central oxygen atom. This presents a problem because there are only three bonds in the ozone molecule. A possible explanation for this could be the existence of resonance structures.

Resonance structures occur when two or more Lewis structures can be drawn for the same molecule or polyatomic ion.

Example Ozone

Note

1) Resonance structures differ only in the assignment of electron pair positions, never atom positions.

2) Resonance structures differ in the number of bond pairs between a given pair of atoms.

Oxidation Numbers

The **oxidation number** is an artificial bookkeeping system for managing electrons in covalent compounds. It is the charge that the atom would have **if** all of its bonds were considered ionic. The oxidation number on a free atom or molecule is zero.

Example Find the oxidation numbers for each atom in H_2O.

______ H ______ O

"A fool has no delight in understanding, but in expressing his own heart."
--Wisdom