Resonance Structures

The more correct way to do Lewis Dot Structures (book method)

- 1. Get the sum of all valence electrons from all atoms. Ignore which electrons came from which atom.
- 2. Arrange the elements
- 3. Place the electrons anywhere in the compound to satisfy the octet and duet rule.

Counting Valence Electrons

- Lone pair electrons belong entirely to the atom in question.
-
- Shared electrons are divided equally between the two sharing atoms.
- (Valence electrons)$_{assigned}$ =
- (# lone pair electrons) +
- 1/2 (# of shared electrons).

Formal Charge

- To determine the formal charge on each atom, take the number of valence electrons assigned to the atom in the molecule and subtract if from the number of valence electrons on the free, neutral atom.

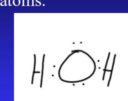
- Formal charge = valence electrons of the free atom- valence electrons assigned

Rules

- <u>The sum of the formal charges of all atoms in a given molecule or ion must equal the overall charge on that species.</u>
- If different Lewis structures exist for a species, those with formal charges on all atoms **<u>closest to zero</u>** are the best.
- All negative formal charges should be on the most electronegative atoms.

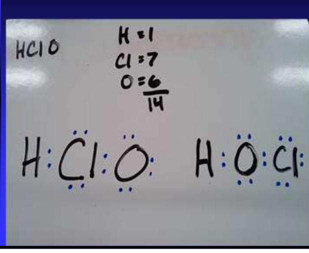
Formal charge

- Most molecules have a formal charge of 0 on all atoms.

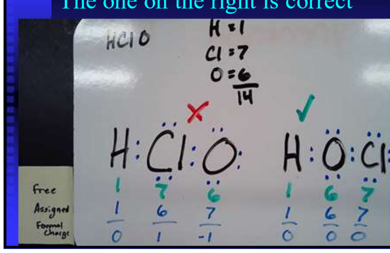


- H 1 electron on free – 1 assigned = 0
- O 6 electrons on free – 6 assigned = 0

HClO



The one on the right is correct



Ions

- **For ions**, an <u>electron is added for each negative charge</u> and an <u>electron is subtracted for each positive charge</u> from the total valence electrons.
- All ions will have a formal charge on at least one atom

- Na^+ would subtract 1 electron for the +1
- SO_4^{2-} would add two electrons for the 2-

Molecules/ions with a formal charge

- O_3
- CO
- NO_3^-
- CO_3^{2-}

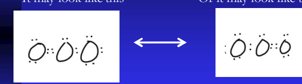
Resonance

- A very important point of this method is showing that electrons don't "belong" to any atom in a molecule.
- The electrons can "flip" places.

Lets look at ozone

It may look like this

Or it may look like this



Data suggests it looks like both at the same time. These are called resonance structures, which are all possible Lewis Dot structures for a molecule.

What this means

- Double bonds and a single bond are different lengths.
- Looking at ozone, you would expect one oxygen to be closer to the middle oxygen than the other.
- Experiments put their bonds at the same length that is somewhere in between the length of a single and double bond.

Stability

- The resonance structure makes it like there are 2 "one and a half bonds" instead of 1 single and 1 double bond.
- This makes compounds much more stable or non reactive.

Differences between Covalent Bonding and Ionic Bonding

Major difference

- Covalent bonding is a sharing of electrons, Ionic bonding is a transfer of electrons.
- Covalent bonds are between a small number of atoms.
- H-O-H one oxygen bonding to two hydrogens
- Ionic bonds are between a very large number of ions stuck together.
- In NaCl every sodium ion is bonded to every single chloride anion near.

Shortcut to determining type of bond

- When a metal and nonmetal bond you get an ionic bond
- ~ something from the left excluding H bonds with something from the right = ionic bond.
- When two nonmetals bond you get a covalent bond
- ~things from the right bond with each other =covalent bond.
- Metals don't bond with each other.

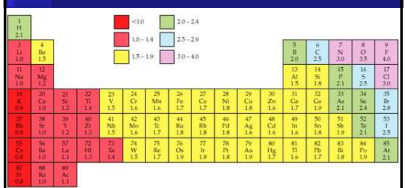
Why this works

- Electronegativity- ability of an atom to attract and hold bonding electrons.
- Elements with a large difference in electronegativity will form an ionic bond, elements with a small difference will form covalent bonds.

Using the periodic table to determine electronegativity

- electronegativity generally increases up and to the right excluding noble gases.
- Fluorine is the most electronegative element (4.0) followed by oxygen (3.5) and chlorine (3.0).
- A full chart is on page 344.

Electronegativity Chart



What about the middle ground?

- What if the difference in electronegativity isn't large or small but in the middle?
- For example H (2.1) and O (3.5)
- These elements form a polar covalent bond.
- Polar Covalent Bond- unequal sharing of electrons in the bond
- so the electrons stay around oxygen more than hydrogen

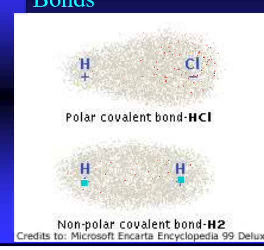
Polar covalent

- 4 electrons occupy this cloud.
- Notice how much larger the cloud is around oxygen as compared to hydrogen.

Do any bonds have an equal sharing?

- Yes, (normally the same element) when elements are equally electronegative like O_2
- In fact, anything with a very slim difference (less than 0.5) in electronegativity will pretty much equally share electrons.
- Nonpolar covalent bonding- equal sharing of electrons in a bond

Bonds

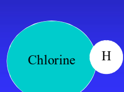


Why it is called polar

polar implies different ends have different charges similar to a magnet.

HCl as a polar bond, meaning the electrons stay around Cl more than H

That makes this side negative



and this side positive

Denoting positive and negative

Neither side is completely positive or negative, they are only partially positive and partially negative.

The symbol δ (lower case delta) means partial

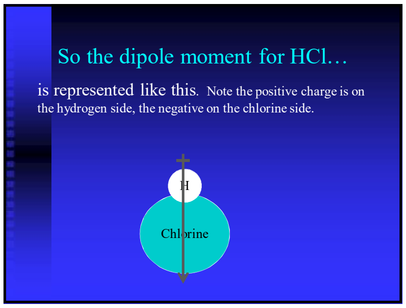
δ -

δ +

Dipole Moment

- Dipole moment- property of a molecule where the charge distribution can be represented by a center of a positive charge and a center of negative charge.
- It is represented by this symbol

- Positive center Negative center


So the dipole moment for HCl…
is represented like this. Note the positive charge is on the hydrogen side, the negative on the chlorine side.
H
Chlorine