Lewis Dot Diagrams

Valence electrons are electrons in the outermost level of the atom.

To show the number of valence electrons we write the symbol with a dot representing the proper number of valence electrons

Octet Rule

Atoms tend to gain, lose or share electrons in order to acquire a full set of valence electrons.

"Octet" comes from the idea that *most* atoms want to have eight electrons in its outermost level (shell).

Exceptions to this rule are H, He and the transition metals.

Drawing Lewis Dot Structures

Procedure:

1) Sum the valence electrons for all atoms
2) Write the symbols for the atoms and connect them with a single bond
3) Complete the octets of the atoms bonded to the central atom
4) Place any leftover electrons on the central atom
5) If there are not enough electrons for the central atom, try multiple bonds.

Formal Charge

In many instances, it is possible to have more than one structure that will obey the octet rule. One approach to deciding which is the correct dot structure is to use formal charge.

To calculate formal charge, we assign electrons to the atoms using these rules:

1) All unshared electrons are assigned to the atom they're found.
2) Half of the bonding electrons are assigned to each atom.

Formal Charge

The formal charge of an atom is equal to the number of valence electrons in the atom minus the number assigned to the atom by the Lewis dot structure.

The most possible structures are those with:

a) The atoms with the smallest formal charges, or
b) The atoms where negative charges reside on the more electronegative atom.

Covalent Bonds

Bonds that form due to a sharing of a pair of electrons between atoms.

Atoms can share valence electrons to achieve an octet.

In some cases, multiple covalent bonds can occur between atoms. In general, the more bonds between atoms, the closer the atoms are together (shorter bond length).

The more bonds between two atoms, the more energy (bond enthalpy) it takes to break the bond.

Bond Length

The more bonds between two atoms, the more energy (bond enthalpy) it takes to break the bond.

The more bonds between two atoms, the shorter the bond length.

The greater the electronegativity difference, the shorter the bond length.

Bonds and Electronegativity

The types of bonds within a compound is determined by use of an electronegativity chart.

Electronegativity is the ability of an atom to attract electrons within a molecule.

As a guideline, ionic bonds exist if the electro-negativity difference between two atoms is greater than 1.9.

There are two types of covalent bonds: polar and nonpolar.

Polar and Nonpolar Covalent Bonds

Polar Covalent - When atoms with different electronegativities form a covalent bond, the shared electrons are more strongly attracted to the atom that is more electronegative. The more electronegative atom acquires a partial negative charge due to stronger attraction of electrons.

Polar covalent bonds occur with an electronegativity difference between 0.5 and 1.9.

Nonpolar Covalent- Electrons are shared relatively evenly.

Nonpolar covalent bonds occur with an electronegativity difference less than 0.5

Resonance Structures

Some molecules can have equivalent Lewis structures. These structures are called resonance structures. Each one of these resonance structures exist equally in a sample of the compound.

Exceptions to Octet Rule

1) Molecules with an odd number of electrons

   Examples: NO, ClO_2

2) Less than an octet

   Compounds with boron, beryllium or transition metals can occasionally have less than an octet.

   Examples: BF_3, BeF_2

Exceptions to the Octet Rule

3) More than an octet

   Largest class of molecules

   Elements in the 3rd period or above can utilize their empty d orbital to accommodate extra electrons

   These expanded valence shells occur most often with the most electronegative atoms

Examples: PCl_5, ICl_4^{-1}