CHEM 101: CHAPTER 11: CHEMICAL BONDS: THE FORMATION OF COMPOUNDS FROM ATOMS

PERIODIC TRENDS: See pages 214 - 216, 221 Table 11.3, and 227 + 228 of text.

Lewis Structures of Atoms: The Lewis Dot Diagram

Lewis Dot Diagrams (developed by chemist Gilbert Lewis) are used to indicate the number of valence e- for an atom and as a result are used in explaining chemical bonding in which one or more e- pairs are either shared (covalent bond) or transferred (ionic bond).

To draw a Lewis dot diagram, simply write the symbol for the element and use dots to represent electrons. The dots are placed at 12:00 and every 15 minutes thereafter clockwise. The dots should be left unpaired up to and including if 4 valence e- (e- = electron(s)) are present (the C family). Additional e- are paired up then.

THE NUMBER OF DOTS SHOULD EQUAL THE NUMBER OF VALENCE ELECTRONS IN THE ATOM. See Figure 11.4 pg. 225 for examples.

The Ionic Bond

Recall that when a metal and non-metal combine chemically, an ionic bond usually results. This joining of atoms is due to a direct transfer of one or more electrons from one atom (usually a metal) to another atom (usually a non-metal). Positive (cations) and negative (anions) ions result and an ionic bond forms.

IN IONIC COMPOUNDS,CATIONS, formed by metal losing e-, ARE SMALLER THAN ANIONS (formed by non-metals gaining e-) AS THE CHARGE BECOMES MORE NEGATIVE, THE LARGER THE ION IS DUE TO INCREASING REPULSIONS (like charges repel and spread apart).

One can use the Lewis Dot Diagrams and knowledge of where the electron goes during a chemical reaction to explain how a given ionic substance forms.

Noble (Inert) Gas Electron Configurations and Predicting Formulae for Ionic Compounds

Atoms generally want to have a total of 8 (octet) valence e-. There is a special stability in acquiring this. ***This is achieved by gaining (for non-metals) or losing (for metals) one or more e- and is referred to as the "Octet Rule."***

Example 1

Using the e- configuration for F, how many e- would F need in order to have an octet of valence e-?

ANSWER: F has e- configuration: $1s^22s^22p^5 \rightarrow$ 7 valence e-, so F needs one more e- from some other atom to give 8.

Example 2: Using the e- configuration for Na, how many e- would Na need to *LOSE* in order to have an octet?

Na has e- configuration: $1s^22s^22p^63s^1 \rightarrow$ 1 valence e- so Na needs to lose 1 e- to have octet below 3s ($1s^22s^22p^6$ remains).

If we then combine these ions to make a formula we would have: $Na^+ + F^- \rightarrow NaF$. *NaF is thus the formula for this particular ionic compound.*

COVALENT BONDS

Recall that covalent bonds are usually between 2 metals or non-metals and are strong bonds formed by sharing e-.

Two e- being shared make up one covalent bond; this bond is termed a <u>single</u> bond.

Two pairs of e- being shared (4) makes up a double bond; three pairs (6) make up a triple bond.

C-C represents a single bond; C=C is a double bond; C≡C a triple bond between carbon atoms.

THESE DASHED LINES INDICATE A PAIR OF ELECTRONS BEING SHARED.

If a bond results in *unequal* sharing of e-, it is termed a *polar covalent* bond. If <u>equal</u> sharing, a <u>non-polar</u> covalent bond results.

How do these bonds form? By orbitals (s, p, d, f) overlapping. The greater the overlap, the STRONGER the bond will be.

ELECTRONEGATIVITY

Electronegativity refers to the amount of attraction an atom has for its electrons. Metals tend to have low electronegativities (high electropositivities); while non-metals usually have high electronegativities (low electropositivities).

The electronegativity scale in use today was developed by Linus Pauling. The highest is 4.0 (F) and the lowest is 0.7 (Fr).

If the difference in electronegativity is zero or close to it, the bond is termed non-polar covalent.

<u>Example</u>: The C-C bond above would be non-polar covalent as the electronegativity difference is 0.

If the electronegativity difference is 0.5 or higher but less than 1.9, the bond is termed polar covalent. These bonds result in a DIPOLE forming.

$\Delta E > 1.7$ but < 1.9 = strongly polar covalent (bond tends to be more ionic)

A dipole results when 2 atoms of significant differing electronegativity share e- unequally. The result is a lop-sided e- cloud which indicates one atom controls most of the e-. ***See Figure 11.10 pg. 229 for examples of this.***

Finally, if the electronegativity difference >2.0, the bond is termed IONIC.

This indicates that electrons being shared by the atoms belong nearly completely in control by one of the atoms (usually a non-metal) in the bond.

<u>Example</u>: When K bonds with Cl (electronegativity difference = 3.0 – 0.8 = 2.2), the electron from K is transferred to Cl and Cl takes a negative charge (an anion, Cl^-); while K a positive charge (a cation, K^+).

<u>Lewis Structures for Compounds</u>

Recall that the Lewis Dot Diagram simply indicates the number of e- with dots to represent valence (outermost, those involved in chemical bonding) electrons.

By combining Lewis Dot Diagrams for the elements in a compound, the LEWIS STRUCTURE can be obtained.

The following steps can be helpful in doing this.

1. *Obtain the total number of valence e- by adding up the value for each of the atoms involved in the compound. If a negative ion is present, add an e-, if a positive ion, subtract an e-.*
2. *Write the symbols for each of the atoms in the compound and connect them with a two dots or a single dash. NOTE: H atoms can form only one bond; O atoms do not usually bond to each other except in peroxides or superoxides. O can have a maximum of two covalent bonds. The halogens (F, Cl, Br, I) usually only form one bond.*
3. *Complete octets of those atoms that surround the central atom (except H).*
4. *If not enough e- are around the central atom, form double bonds by shifting a pair of e- or if necessary triple bonds. Each bond counts for two e-.*
5. *If the compound has a charge (an ion), indicate that with a superscript.*

See pg. 232 of text for steps as well.

EXAMPLES

Resonance and Polyatomic Ion Containing Compounds Lewis Structures

Some compounds can have more than one Lewis structure. This is due to a bond (usually a double bond) shifting its location from the central atom to any of the other atoms attached to it.
<u>All structures resulting from this are equivalent (termed "resonance" structures).</u>
See examples on pgs. 235 + 236 of book.

MOLECULAR SHAPE

Molecules formed by atoms do NOT all have the same shape. This is due to the arrangement of the e- pairs around the central atom(s). Electrons being negative (-) repel each other, so the atoms want to arrange themselves in a molecule in such a way as to reduce these repulsions and give the molecule better stability. The theory that explains and predicts these shapes is termed the VALENCE SHELL ELECTRON PAIR REPULSION MODEL or VSEPR.

The following shapes are predicted by VSEPR theory based upon the number of atoms present, number of bonding e- (those forming a bond between two atoms, and the number of non-bonding pairs of e- (those around the central atom but NOT forming a bond).

Molecule	**Number of e- pairs**	**e- pair arrangement**	**Molecular shape**
CO_2	**4 bonding, *0 non-bonding***	**Tetrahedral**	**Linear**
H_2O	**2 bonding, *2 non-bonding***	**Tetrahedral**	**Bent or Angular**
CF_4	**4 bonding, *0 non-bonding***	**Tetrahedral**	**Tetrahedral**
$AlBr_3$	**3 bonding, *0 non-bonding***	**Trigonal Planar**	**Trigonal Planar**
NH_3	**3 bonding, *1 non-bonding***	**Tetrahedral**	**Trigonal Pyramidal**
PI_5	**5 bonding, *0 non-bonding***	**Trigonal Bipyramidal**	**Trigonal Bipyramidal**
SF_6	**6 bonding, *0 non-bonding***	**Octahedral**	**Octahedral**

***SEE PGS. 239 – 241 for examples and further explanation.**