NOTES #26 BONDING Part A - AP Chemistry

I. Lewis Dot Diagrams: A way of representing the valence electrons as dots spread out around an atom or ion's symbol.

N N^{3-} Ca Ca^{2+}

II. Lewis Dot Diagrams can be used to show what the e- are doing when making bonds.

** Why do atoms bond? To gain/lose or share e-'s so as to____

A. **IONIC BOND** - bond created by the attractive force between two ______________________. Involves the TRANSFER of e-.

ex: **Li** + **F** →

B. **COVALENT BOND** - bond created when a pair of electrons are ____________________________ between two atoms.

ex: **F** + **F** →

** Notice that both F's have _______ electrons around them or a STABLE OCTET.

III. Focus on COVALENT BONDING.

A. Atoms that are covalently bonded are called _________________________________. ex. $C_6H_{12}O_6$, H_2O, CH_4

B. Molecules are often represented as LEWIS STRUCTURES. Lewis structures allow us to see how atoms are arranged or bonded in a molecule. For simple molecules, it's really easy to determine the Lewis structure of a molecule just by looking at the Lewis dot diagrams for each atom involved and pairing up any unpaired electrons.

1. SINGLE BONDS: H_2O NF_3

2. MULTIPLE BONDS: Sometimes, more than just one pair of electrons has to be shared between two atoms in order to fulfill all octets. In these cases, DOUBLE BONDS (when two pairs of e- are shared btwn two atoms) or TRIPLE BONDS (when three pairs of e- are shared btwn two atoms) have to be used.

O_2 N_2

** A NOTE ABOUT BOND LENGTH - For a given pair of atoms, a single bond is always _________________ than a double bond which is ___________________ than a triple bond. C–C > C=C > C≡C

C. **RULES FOR DRAWING LEWIS STRUCTURES**:

** For many molecules, the above method of determining the Lewis structure doesn't always work. The following rules work for more complex structures as well:

Ex: CO_3^{2-}

1. Write the skeletal structure of the compound.
   * The atom that wants e^-'s the least usually in the middle
   * Lowest electronegativity values/carbon is very often in the center.
2. Count the total number of valence electrons present.
   * If there is a (-) charge on the structure, you need to _________ that many e^-.
     ex. CO_3^{2-} has _____________________.
   * If there is a (+) charge on the structure, you need to subtract that many e^-.
3. Draw a single covalent bond between the central atom and each of the surrounding atoms. That's _____ e^- per bond. With the remaining e^-, start filling in the octets of the e^- starting with the surrounding atoms. When these outer atoms have full valence shells, put any remaining e^- on the central atom as *lone electron pairs*.
   * Remember, hydrogen only needs _____ e^- to have a full valence shell.
4. Is the octet rule satisfied for the central atom? If not, move a lone pair of e^- from a surrounding atom and make a double or triple bond.

*5. CHECK THAT ALL VALENCE e^- HAVE BEEN ACCOUNTED FOR AND THAT ALL ATOMS HAVE FULL VALENCE SHELLS.

6. If it's a charged structure, put it in ___________________________ indicating the overall charge. (Forgetting this is the most common mistake.)

PRACTICE: 1) NH_4^+ 2) NO^+ 3) HNO_3

IV. DETERMINING FORMAL CHARGE

A. Formal charge - is the difference between the number of valence electrons on the free atom and the number of valence electrons assigned to the atom in a molecule.
- Lone electrons belong entirely to the atom in question.
- Shared electrons are divided *equally* between the two sharing atoms.

B. Equation for formal charge for an atom: **formal charge = (# of valence e^-) - (# of unshared e^-) - 1/2 (# of bonded e^-)**

ex: What's the formal charge of the O's in the following molecule: O_3

C. The sum of the formal charges of all atoms always equal the total charge on the structure.

ex. CO_3^{2-} What are the formal charges for the atoms in this polyanion???

D. Why are formal charges important? Sometimes, different Lewis structures are possible. Formal charges help us decide which structure is most likely. Just remember:

** Lewis structures with LARGE formal charges (+2,+3,-2,-3,etc) are indicative of a BAD structure.
** Negative formal charges should always be placed on the atom that wants electrons the most (most electronegative).

ex. Based on formal charges, circle the Lewis structure that best represents the compound, formaldehyde, CH2O.

```
                                                H
H   C   O   H              or                            C          O
                                                H
```

ex. Draw the possible Lewis structures of a molecule that contains a N atom, a C atom, and a H atom. Circle the one that is most reasonable based on formal charge assignments.

V. EXCEPTIONS TO THE OCTET RULE:

A. Some atoms are "happy" with less that 8 electrons surrounding the central atom.

1. Hydrogen and helium, of course, are happy with only _____e- around them.
2. In molecules, BERYLLIUM is happy with just ______ e- around it.
   ex. Draw the Lewis structure for BeH_2
3. In molecules, Boron and _______________ are often happy with only _____ valence electrons around them.
   ex. Draw the Lewis structure for BF_3.

**Another factor that proves that BF_3 exists with less than an octet is the fact that it reacts like it's electron deficient. The Boron accepts a pair of electrons from the Nitrogen in the molecule, ammonia.

```
          H
B  +   : N ---H    ------------>
          H
```

- Do you notice something a little different about the B-N bond?