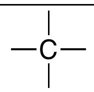
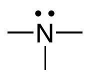
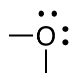
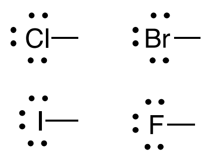
## <u>Normal Bonds (Sections 1.2-1.5)</u>

Summary of **<u>Normal, Ideal Bonding</u>** <u>(No Formal Charge)</u>

| | Valence Electrons | Valence Bonds | Lone Pairs |
|---|---|---|---|
|  | 4 | 4 | 0 |
|  | 5 | 3 | 1 |
|  | 6 | 2 | 2 |
| H— | 1 | 1 | 0 |
|  | 7 | 1 | 3 |

**<u>Rules for Drawing Lewis structures for organic molecules: (Sections 1.4-5)</u>**

1. Try to provide normal bonding for C, N, O atoms **<u>if possible</u>**. (Works > 95% of time)
2. Double or triple bonds will often be involved.
   - Double or triple bonds are often required to achieve normal bonding.
3. In any formula that has a charge, there will always be an atom with that formal charge.
4. In any formula that includes a metal, assume ionic bonding.
   - Assume positive charge for the metal,
   - Assume negative charge for the organic portion.
5. Do not draw bonds between nonmetals and metals, as if they were covalently bound.
6. <u>Be sure to specify the formal charge on any atom that has formal charge</u>.
7. Always be aware of how many lone pairs are on any atom
   - Note: We will often omit lone pairs. But you must know when they are there!

**<u>Lewis Structure Practice (Section 1-4,5)</u>**

1. Draw Lewis structures for the following formulas: (Include lone pairs or formal charges if necessary)

   a. $CH_3CH_2OH$ b. $CH_3CH_2OH$

   c. $CO_2$ d. HCN

   e. $CH_3CHO$ f. $NaOCH_3$