

Rules and Steps for Lewis Dot Formulas

- Determine if bonds are ionic or covalent (covalent follow this sheet)
 - In most compounds the elements in each atom achieve the isoelectronic configurations of the closest noble gas (s and p block only, NOT transition metals)
- Figure out which element is the least electronegative, except H.
- Place the least electronegative atom in the center; this is called the central atom.
- Even space the remaining atoms around the central atom, keeping in mind the following guidelines;
 - H always only forms one bond, so it should be terminal (it is never the central atom)**
 - Atoms tend to form (8 – group number) of bonds, so....**
 - Halogens (Group 7A) usually form one bond**
 - Oxygen (Group 6A) usually forms 2 bonds (2 single or 1 double bond)**
 - Nitrogen (Group 5A) usually forms 3 bonds (3 single, 1 single and 1 double, or 1 triple bond)**
 - Carbon (Group 4A) usually forms 4 bonds (4 single, 2 double, 2 single and 1 double, or 1 single and 1 triple bond)**
 - Boron (Group 3A) usually forms 3 bonds (does not follow the octet rule)**
 - Beryllium (Group 2A) usually forms two bonds (does not follow the octet rule) All others, Ionic bonds!!**
 - Lithium (Group 1A) usually forms one bond, but they are always ionic!!**
- Add up the total number of electrons needed for each atom to achieve noble gas configuration, N. (H and He=2 and all others = 8)
- Add up the total available electrons for each atom, A (Group no.)
- Subtract A from N, this is the total number of shared electrons, S
- Divide S by 2, this is the total number of electron pairs (no. of bonds)
- Start by placing the shared electrons as single bonds around the central atom
- Place any remaining single bonds to atoms not bonded to the central atom
- How many bonds do you have? Subtract this from the total (S). If zero then go to step 14. If not go to step 12.
- Check the octet rule for each atom, starting with the central atom.
- Fill in double or triple bonds where needed. Do you have resonance? Show each structure for resonance.
- Have all the pairs been used? Check the number of available electrons. Subtract the number of shared electrons from the available electrons (A-S); these “extra” electrons are used to fill in lone pairs.
- Fill in the “extra” electrons as pairs around the appropriate atoms.
 - General guide to lone pairs**
 - 1. Group 4A (C, Si...), no lone pairs/ 4 bonds**
 - 2. Group 5A, (N, P....)1 lone pair/ 3 bonds**
 - 3. Group 6A, (O, S....)2 lone pairs/ 2 bonds**
 - 4. Group 7A (F, Cl....)3 lone pairs/ 1 bond**
- Now Check your drawing does it make sense? Remember Lewis structures can also include resonance, isomers, and expanded valence.
- Can't attach everything without formal charges? Follow these rules for formal charges;
 - Group 5A, (N, P....)1 lone pair/ 3 bonds = 0 Formal Charge**
 - Group 5A, (N, P....)0 lone pair/ 4 bonds = +1 Formal Charge**
 - Group 5A, (N, P....)2 lone pair/ 2 bonds = -1 Formal Charge**
 - Group 6A, (O, S....)2 lone pairs/ 2 bonds = 0 Formal Charge**
 - Group 6A, (O, S....)3 lone pairs/ 1 bonds = -1 Formal Charge**
 - Group 6A, (O, S....)1 lone pairs/ 3 bonds = +1 Formal Charge**