Rules and Steps for Lewis Dot Formulas

1. 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)

1. Figure out which element is the least electronegative, except H.
2. Place the least electronegative atom in the center; this is called the central atom.
3. 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!!

4. 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)
5. Add up the total available electrons for each atom, A (Group no.)
6. Subtract A from N, this is the total number of shared electrons, S
7. Divide S by 2, this is the total number of electron pairs (no. of bonds)
8. Start by placing the shared electrons as single bonds around the central atom
9. Place any remaining single bonds to atoms not bonded to the central atom
10. 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.
11. Check the octet rule for each atom, starting with the central atom.
12. Fill in double or triple bonds where needed. Do you have resonance? Show each structure for resonance.
13. 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.
14. 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**

15. Now Check your drawing does it make sense? Remember Lewis structures can also include resonance, isomers, and expanded valence.
16. Can't attach everything without formal charges? Follow these rules for formal charges;

    1. **Group 5A, (N, P….)1 lone pair/ 3 bonds = 0 Formal Charge**
    2. **Group 5A, (N, P….)0 lone pair/ 4 bonds = +1 Formal Charge**
    3. **Group 5A, (N, P….)2 lone pair/ 2 bonds = -1 Formal Charge**

    4. **Group 6A, (O, S…)2 lone pairs/ 2 bonds = 0 Formal Charge**
    5. **Group 6A, (O, S…)3 lone pairs/ 1 bonds = -1 Formal Charge**
    6. **Group 6A, (O, S…)1 lone pairs/ 3 bonds = +1 Formal Charge**