Rules for Writing Lewis Dot Structures

Please utilize this information and ***FOLLOW*** these rules step by step when doing practice problems. This is how you can best learn and understand the system of drawing Lewis structures. This is good study practice.

1. Calculate total number of valence electrons in the structure (add group numbers for each atom).

(a) If structure is a positive ion, subtract electrons equal to the charge.
(b) If structure is a negative ion, add electrons equal to the charge.

2. Draw a symmetrical skeleton structure connecting atoms bonded with a dash

(a) Generally the least electronegative atom is the central atom (but never hydrogen as it always forms only one bond).
(b) In ternary acids (ones containing polyatomic ions), hydrogen bonds to oxygen and not to the central atom.
(c) If several structures are possible, choose the one where atoms have the usual number of bonds (usual number of bonds = 8 – group number)

3. Count each dash bonding atoms as two electrons then add lone pairs (non bonding pairs) to atoms surrounding the central atom until each has eight electrons around it (except H – 2 electrons)

4. Add remaining electrons as lone pairs on the central atom

(a) If you are two electrons short on the central atom, move a lone pair from a surrounding atom to become a bonding pair and form a double bond.
(b) If you are four electrons short on the central atom, move two lone pairs from surrounding atoms to become bonding pairs and form a triple bond or two double bonds.
(c) If you have extra electrons, place them as lone pairs on the central atom (central atom may have up to twelve electrons around it).

Valence Shell Electron Pair Repulsion Theory (VSEPR)

Lewis structures will tell us where the electrons are placed (in terms of which atoms they are between or which atoms they are sitting on as lone pairs). Once we have this, we can use valence shell electron pair repulsion theory to determine the electronic and molecular geometry of the molecule.

We start by counting the total regions of electron density in the Lewis structure. Note that a bond (either single, double or triple) counts as one region. Then we use the following table to identify the electronic structure and the molecular structure.

Electronic structure – structure of the all the electron density regions (bonded and non-bonded)

Molecular structure – structure of the bonded electron density regions only

When there are no lone pairs the electronic and molecular structures will be the same.

Total Regions of Electron Density	Number of Lone Pairs	Molecular Structure	Electronic Structure	Central Atom Hybridzation
2	0 1	linear linear	linear	sp
3	0 1 2	trigonal planar bent (angular) linear	trigonal planar	sp^2
4	0 1 2 3	tetrahedral trigonal pyramid bent (angular) linear	tetrahedral	sp^3
5	0 1* 2* 3* 4	trigonal bipyramid seesaw t-shaped linear linear	trigonal bipyramid	sp^3d
6	0 1 2** 3 4** 5	octahedral square pyramid square planar t-shaped linear linear	octahedral	sp^3d^2

* - lone pairs appear in equatorial positions (ones with 120° angles) before axial positions (ones with 90° and 180° angles)

** - second lone pair is 180° from the first lone pair

Bond Angles and Lone Pairs

1. Since a lone pair occupies more space near the central atom, it repels bonding electron pairs more strongly and decreases bond angles for binding electron pairs

e.g. CH_4 (0 lone pairs) → bond angle H-C-H = 109.5°
NH_3 (1 lone pair) → bond angle H-N-H = 107°
H_2O (2 lone pairs) → bond angle H-O-H = 104°

2. Double and triple bond electron pairs occupy more space near the central atom and decreases bonding electron pair angles

Polarity

Nonpolar bond – bond where atoms are shared equally

Polar Bond – bond where the electrons are shared unequally due to difference in electronegativity (EN) of atoms bonded

$H^{\delta+}$ --- $F^{\delta-}$

$\delta+$ → means partially positive end of bond (or atom with smaller EN)
$\delta-$ → means partially negative end of bond (or atom with larger EN)

In molecules with more than one bond, dipoles are *additive* and if they cancel, the molecule is nonpolar, if not then the molecule is polar

A polar molecule usually has

(a) lone pairs on the central atom → except square planar and linear (from 5 total regions of electron density)
(b) not all atoms bonded to the central atom are the same

A nonpolar molecule usually has

(a) no lone pairs on the central atom (or lone pairs are symmetrically arranged)
(b) atoms bonded to the central atom are the same (or different atoms are symmetrically arranged)