**Lewis formulas and octet rule**

**First**, let’s see what Lewis structure is. Lewis structure, also called electron-dot structure, is a structural formula in which electrons are represented by dots; two dots between two atoms represent a covalent bond. It was named from Gilbert Newton Lewis, who introduced it in 1916. He proposed that a sharing of two electrons by two hydrogen atoms permits each one to have a stable closed-shell electron configuration analogue to helium.

**Second**, the importance of Lewis structure. In fact, Lewis structures are very important for predicting geometry, polarity and reactivity of (in)organic compounds.

**Third**, how to draw Lewis structure.
For individual atoms, the Lewis structure is drawn by placing a dot around the atom for each valence electron available. There are four positions for dots to be placed. Most chemists draw them on the top, bottom, left and right of the atom. Each of the four positions must receive one electron (one dot) before any electrons can be paired with one another. Helium is an exception; the two electrons are shown as a pair.

For molecules, before we draw their Lewis structures we have to learn octet rule. Octet rule states that in forming compounds, atoms gain, lose or share electrons to give a stable electron configuration characterized by eight valence electrons. This rule is applied to the main-group elements of the second period.

**Draw Lewis structures step by step:**
1. choose a center atom. draw the molecular skeleton.
2. count all valence electrons.
3. draw a single bond from each surrounding atom to the central atom, subtract 2 valence electrons for each bond.
4. if there are any electrons left, distribute them as lone electron pairs to maximize the number of the octets.
5. if some of the atoms lack octet structures, change as many lone electron pairs into shared electron pairs as required to complete the octet shells.

**At last**, some exceptions about octet rule.
1. As we all know, in Lewis structure all electrons are distributed as bonding or lone pairs. However, this distribution is not possible in species having an odd number of electrons, such as nitrogen oxide (e.g. NO, NO₂) and neutral methyl.
2. when there are too few electrons to form an octet electron deficiency occurs. For example, BF₃, with only 6 electrons in its bonding pairs.
3. Atoms with shells 3 and higher can expand their outer shell to form more than four bonds.