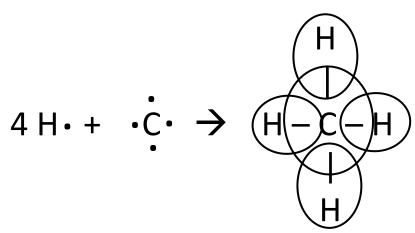
CHE 2060: Guide to covalent bonding and drawing Lewis structures

Covalent bonding
In inorganic chemistry you learned a bit about covalent bonding: bonds formed when two atoms each contribute an electron to form a shared pair of electrons that holds the two atoms together. Atoms form covalent bonds in order to fill their valence shells. Remember that a full valence shell is eight atoms in the outmost electron shell for all atoms except hydrogen and helium, satisfied with just two valence electrons.

- **Valence electron** number is an atom's column (or group) number.
- **Lewis structures** use lines to show covalent bonds between atoms in an organic compound. Free electron pairs, aka lone pairs, are represented by pairs of dots on the atoms that hold them.
- **Formal charge** is the charge carried on each atom in the structure.
  formal charge = (number of valence electrons) – (dots + sticks)

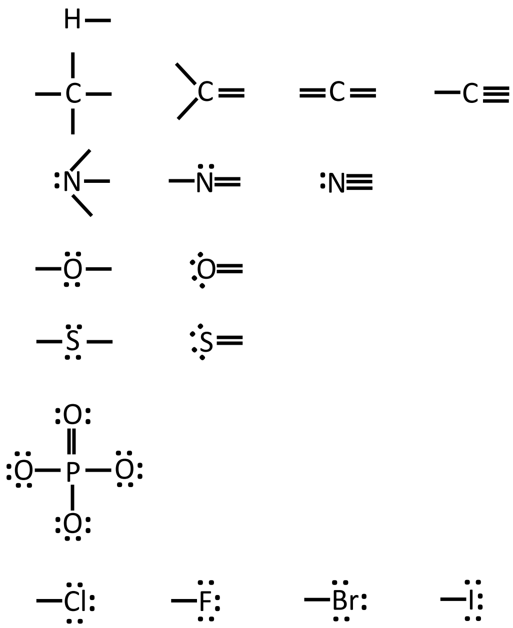
Example: Neither hydrogen (I) nor carbon (IV) have full valence shells. However, when they share their valence electrons all atoms achieve a full valence shell. The molecule shown here is methane, CH4.



Steps for drawing Lewis structures:
1. Count the number of valence electrons in the formula; add an electron for each negative charge and subtract one for each positive charge.
2. Place the symbol of the unique atom on the center and surround it symmetrically by repeated atoms. For carbon-based molecules, start with a line of carbons in the center.
3. Connect each atom in your structure with a line representing a covalent bond. Each line uses up two electrons.
4. Add electrons to the peripheral atoms to fulfill their need for an octet of eight valence electrons.
5. Place any remaining valence electrons on the central atom to fulfill its octet.
6. If the central atom does not have an octet, move pairs of electrons from a peripheral atom to form a double bond between that atom and the central atom. Repeat until the central atom has an octet.
7. Generally, symmetry is good. Check your work by checking the number of valence electrons you've used to make the structure.
8. Calculate formal charges.

Common bonding patterns
It’s easier to draw Lewis structures, and to begin to understand and visualize atomic and molecular geometry, if you can recognize common bonding patters of the atoms found in organic molecules.

Common bonding patterns for **uncharged atoms** are show here. Note these are the ‘biological’ bonding patterns of S and P.



Common bonding patterns for **charged atoms**:

- Note that S and P can have more than an octet; S, P and other large atoms, are exceptions to the octet rule. In these exceptional cases, S and P will be charged.
- Calculate the formal charge of atoms with these charged bonding patterns by subtracting the sum of ‘dots + sticks’ from the atoms valence number: FC = ve- - (dots + sticks).