# Module 6: Drawing and Interpreting Lewis Structures

## Formal Charge

Fundamentals of Chemistry Open Course

# Learning Objectives | Module 6

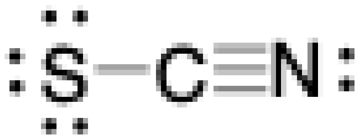
1. Infer the number of valence electrons in a neutral atom from the position of its element on the periodic table.
2. Draw Lewis symbols for atoms with the appropriate number of electrons.
3. Infer the most likely skeletal structure of a simple molecule from the molecular formula.
4. Distinguish between bonding and nonbonding electrons in Lewis structures.
5. Define and apply the octet rule when drawing Lewis structures.
6. Draw Lewis structures for simple molecules containing only single bonds and formally neutral atoms.
7. Draw Lewis structures for simple molecules containing multiple bonds.
8. Deduce the formal charge of an atom in a Lewis structure.

- **Formal charge:** the charge on an atom in a Lewis structure deduced by splitting each pair of bonding electrons evenly between the two bonded atoms. Formal charge is the difference between the formal charge electron count and the valence electron count of the neutral atom.
    - **Formal charge electron count (FCEC):** the number of nonbonding electrons plus *half* the number of bonding electrons associated with an atom.
    - **Valence electron count of the neutral atom (VEC):** the number of valence electrons in a neutral free atom.

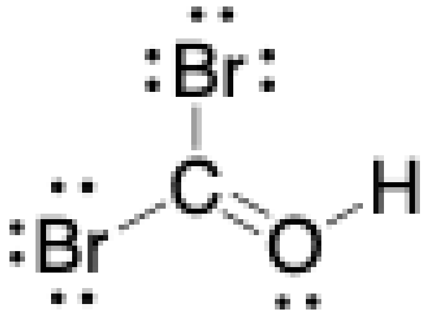
- The sum of all formal charges in a structure must equal the overall charge of the molecule.

- Most atoms are neutral…but not all!

**Example.** Determine the formal charge of each atom in the structure of thiocyanate ion ($SCN^-$) below.



**Example.** Determine the formal charge of each atom in the structure below.



- Formal charge is never greater than ±1 in physically realistic (valid) Lewis structures.
  Keep this in mind when drawing structures—make bonds to keep formal charges low.

# Drawing Lewis Structures Revisited

Drawing a Lewis structure given a molecular formula…

1. Determine the total number of valence electrons. For cations, subtract one electron for each positive charge. For anions, add one electron for each negative charge.

2. Draw a **skeletal structure** of the molecule or ion, arranging the atoms around a central atom. Connect each atom to the central atom with a single bond (one electron pair).

3. Distribute the remaining electrons as lone pairs on the peripheral atoms (except hydrogen), completing an octet around each atom.

4. Place all remaining electrons on the central atom.

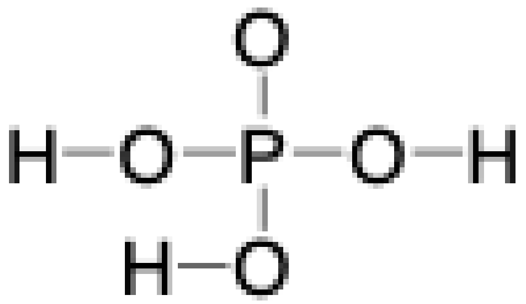
5. Rearrange the electrons of the outer atoms to make multiple bonds with the central atom to create octets wherever possible.

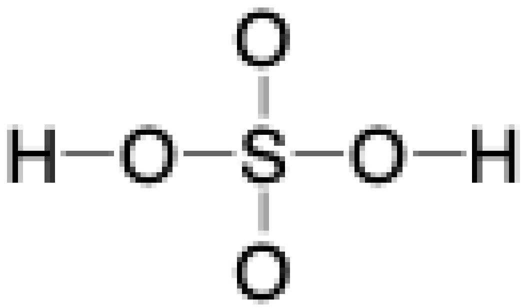
6. Assign formal charges, making sure no atom has charge greater than ±1 and the sum of all charges equals the total charge of the molecule or ion. Revisit step 5 if necessary!

# The Octet Rule Revisited

- Atoms in periods 3 and below will frequently require more than an octet of electrons to have neutral formal charge. Such “exceptions” of the octet rule are generally fine.

**Example.** Complete the Lewis structures of phosphoric acid ($H_3PO_4$) and sulfuric acid ($H_2SO_4$) below.
What exceptions to the octet rule do you observe?





# Lingering Questions

- It is sometimes possible to draw more than one Lewis structure for a molecule, even when keeping the connectivity of the atoms and the total number of valence electrons the same. Why does this happen? How do we decide which Lewis structure is "best" in this case?

- Are formal charges physically realistic? Do they actually represent regions of positive and negative charge density in molecules? Why or why not?