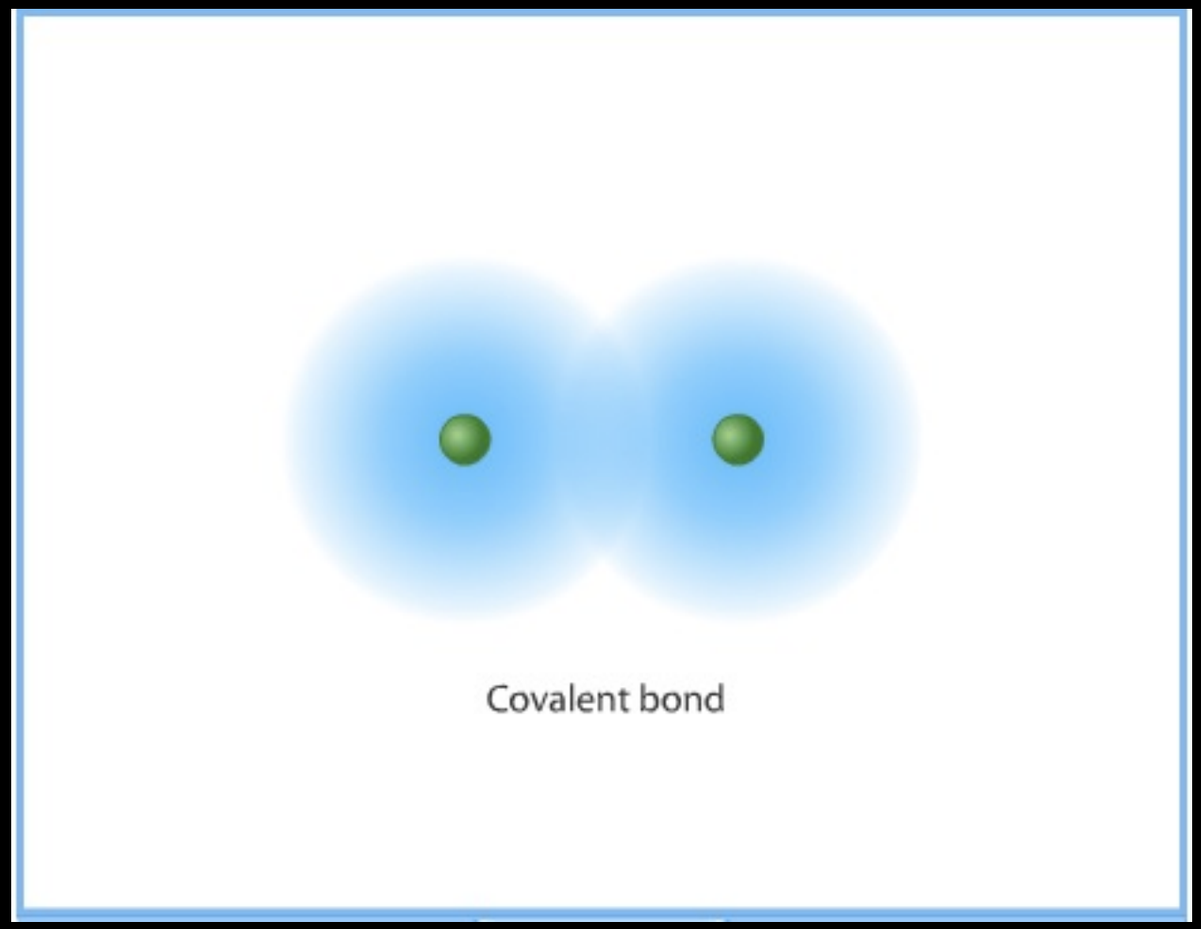
4.2 Covalent bonding

Covalent bond:

- the electrostatic attraction between a pair of electrons and positively charged nuclei.
- the sharing of electrons between atoms.



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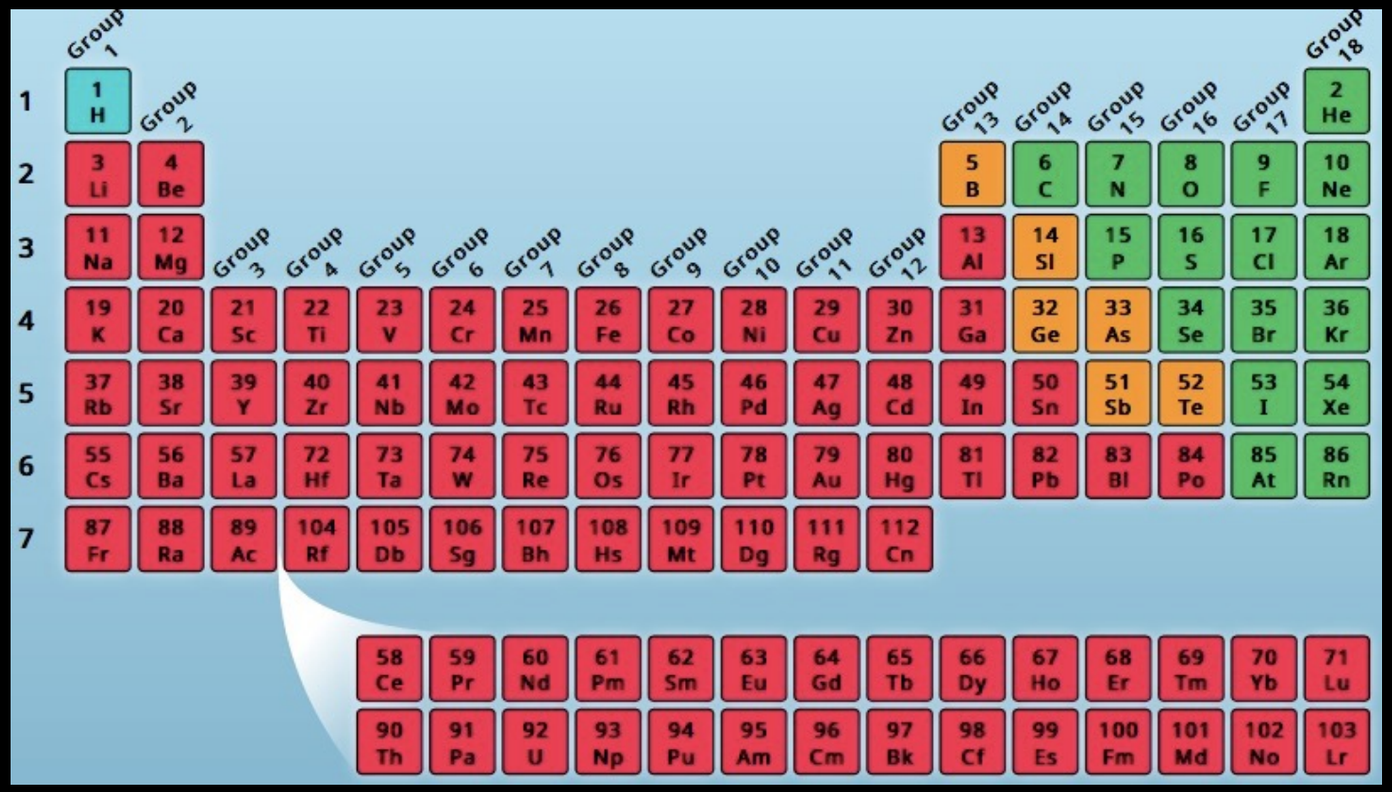
General Rule:

Covalent if a nonmetal + nonmetal

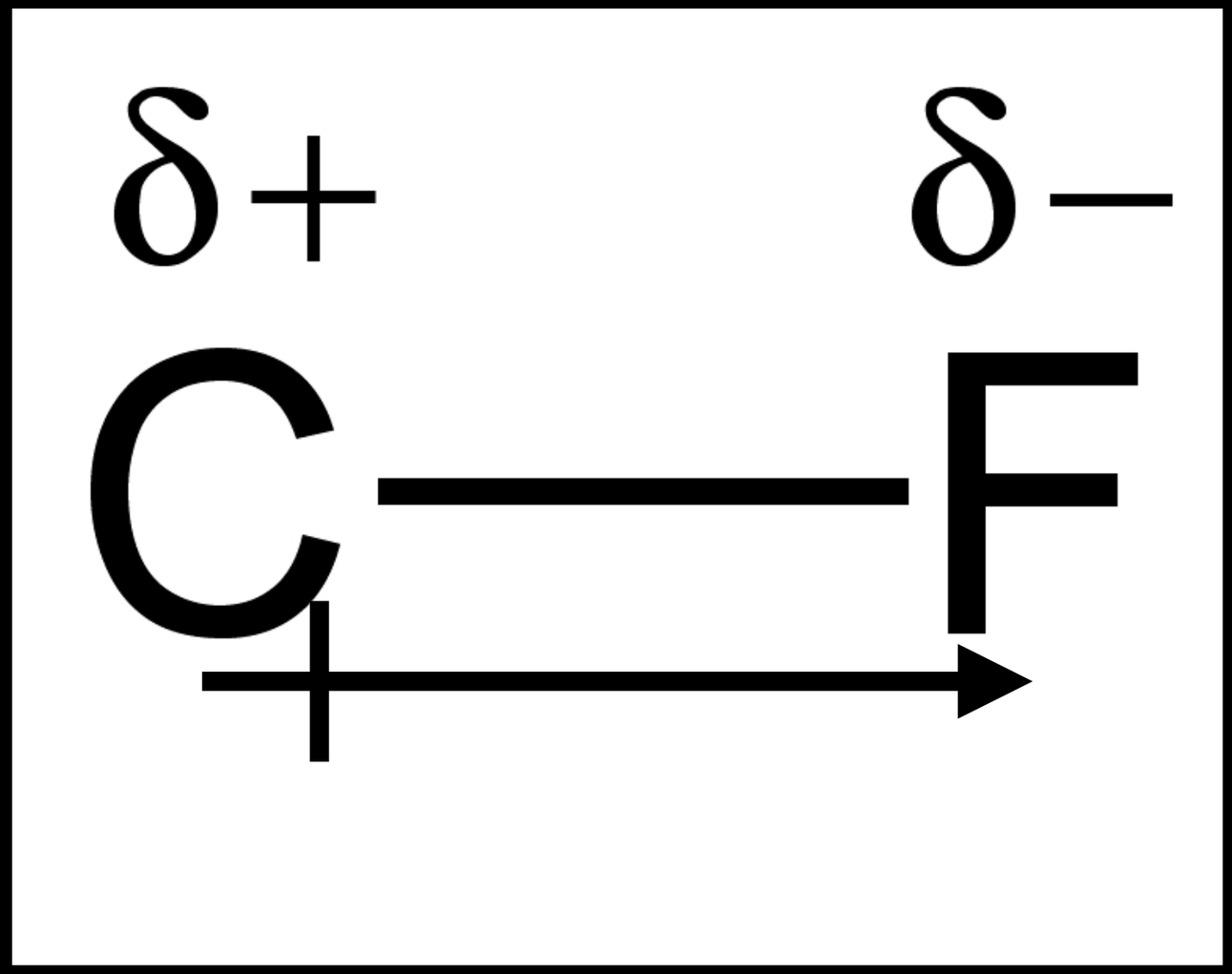
To determine covalent bond type, use electronegativity values:

Electronegativity Difference	Bond type
< 0.3	nonpolar covalent
0.3 - 1.7	polar covalent
>1.7	ionic

Online Tutorial



The Greek letter delta (δ) to represent polarity.



...or an arrow!

Naming

Use prefixes to name binary molecular compounds:

1	mono		6	hexa
2	di		7	hepta
3	tri		8	octa
4	tetra		9	nona
5	penta		10	deca

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Lewis structures:

The use of dots and dashes to represent the valence electrons of an atom, ion, or molecule.

Steps for drawing Lewis structures for molecules or polyatomic ions:

1. Total all valence electrons (and charges).
2. Connect all atoms to a central atom.
3. Add dots to complete an octet around every atom connected to a central atom (except hydrogen).
4. Add dots to complete the central atom's octet.
5. If too few electrons are available, use multiple bonds.

The HONC Rule

Exceptions to the octet rule:

1. Hydrogen will have only 1 pair of electrons surrounding it.

2. Boron will only have 3 pairs of electrons surrounding it.

3. Beryllium (a metal) can form covalent bonds with only have 2 pairs of electrons surrounding it.

4. Sulfur and Phosphorus will occasionally have more than 4 pairs of electrons by using their empty d-orbitals.

This is called an "expanded octet."

Special note regarding oxyacids:

(H bonded to an oxyanion)

Hydrogen will ALWAYS be bonded to the oxygen atom(s).

Special note regarding carbon:

<u>Carbon</u> rarely has an unshared pair of electrons.

Two exceptions:

CO
carbon monoxide

CN^-
cyanide

Is your Lewis structure correct?

Two simple questions to ask:

1.) Did you use the exact number of valence electrons available?

2.) Are all atoms (except hydrogen) surrounded by 8 electrons?