Covalent Bonds

<u>**Covalent (Molecular) Bonding**</u>: two <u>nonmetal</u> atoms share electrons to fill the valence level of both atoms.

- Occurs between elements with Similar electronegativities, high effective nuclear charges (Z_{eff}) and small radii, so they can attract and hold each other's electrons to make shared pairs of electrons.
- The smallest group of elements held together by a covalent bond is called a molecule.
- Atoms can make <u>single</u>, <u>double</u>, or <u>triple</u> bonds depending on whether they share one, two, or three pairs of electrons respectively.
  - **Multiple bonds are most often formed by C, N, O, P and S atoms, aka** C-NOPS.

Covalent Bond Properties

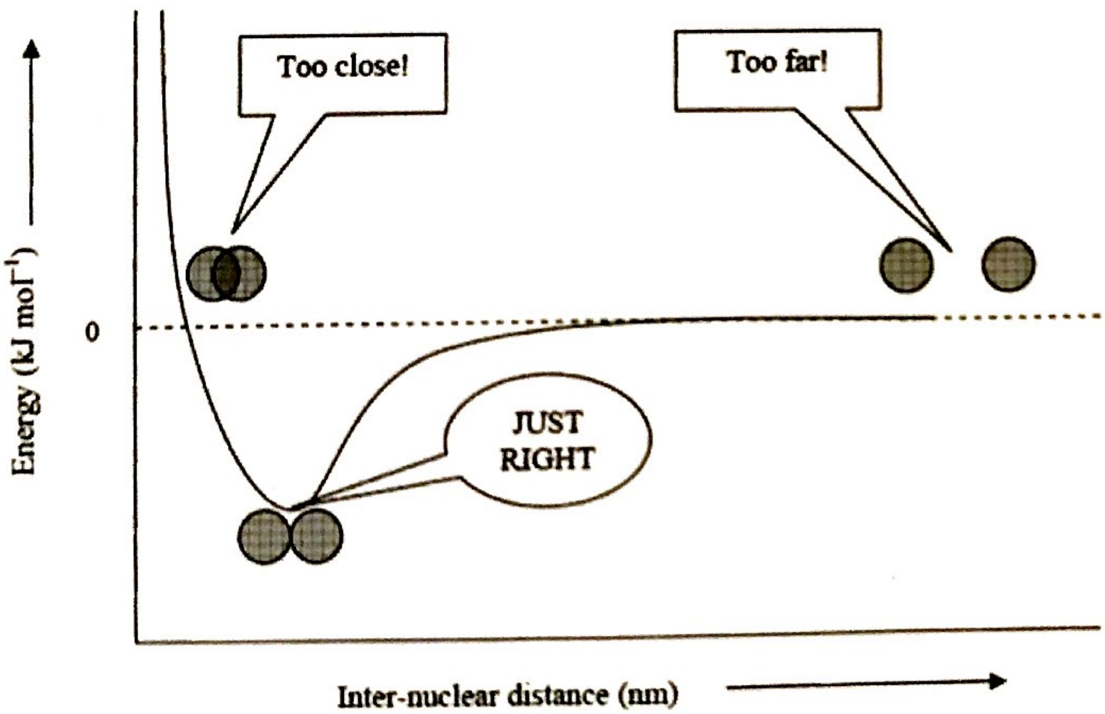
- Low melting points and boiling points. (held together by IMFs)
- Typically do NOT conduct electrical current because they lack mobile charges (EXCEPT strong acids!)

Coulomb's Law tells us:

- The negative electrons of one atom and the positive nucleus of another atom attract each other.
- If the nuclei of two atoms get too close together, their like charges repel each other.

<u>Bond length</u>: the distance two covalently bonded atoms at their lowest potential energy. It is a balance between opposing forces:

- Attractive electrostatic forces between the nucleus of one atom and the electrons of another
- Repulsive forces between the two positively charged nuclei



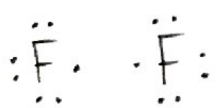
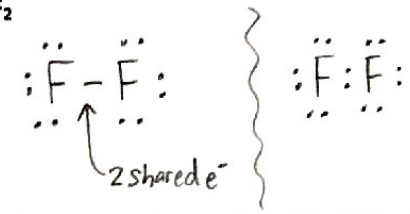
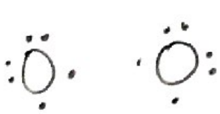
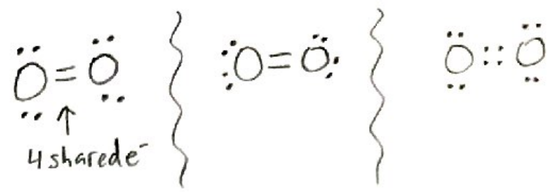
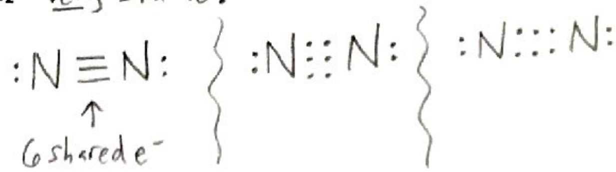
Bond Order: the number of chemical bonds between a pair of atoms; indicates the stability of a bond.

Bond Type	Bond Order	Bond Length	Bond Strength
Single bond	1	longer	weaker
Double bond	2	medium	medium
Triple bond	3	shorter	stronger

Multiple bonds increase the electron density between two nuclei

- Decreases repulsions between the two nuclei
- Added electrons ↑ attraction between nuclei and electron density → multiple bonds ↑ bond strength!
- Nuclei can move closer together → multiple bonds ↓ bond length!

Directions: On the left, show the neutral, separate atoms using Lewis valence electron dot structures. On the right, depict the bonding atoms sharing electrons.

2 atoms of F 	**F_2** 
2 atoms of O 	**O_2** 
2 atoms of N 	**N_2** very stable! 

1. Explain: why does it require more energy to break the bond between O_2 than F_2?

b/c the O_2 double bond (with 4 shared e^-) is stronger than the F_2 single bond (w/ only 2 shared e^-)

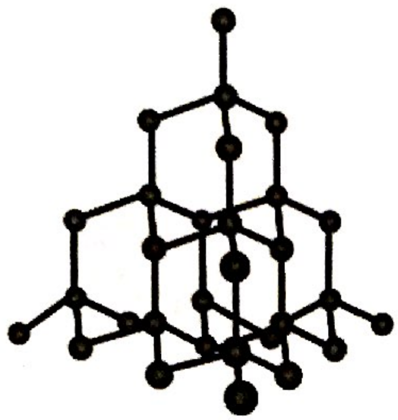
2. Rank the following in order of increasing bond length: O_2, F_2, N_2, H_2

$H_2 < N_2 < O_2 < F_2$

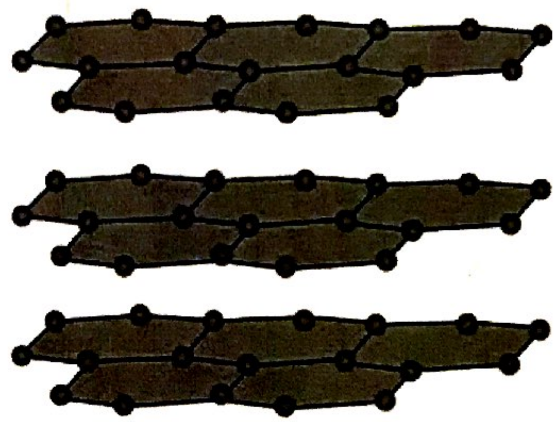
Network Covalent Solids

Network Covalent Solid: a crystalline structure is formed by non-metals covalently bonded together into 2-D (sheets) or 3-D networks.

- A perfect single crystal of a network covalent solid is a single, giant molecule!
- VERY high melting and boiling points
- Very rigid and hard
- Chemically inert (non-reactive); rarely dissolve in water



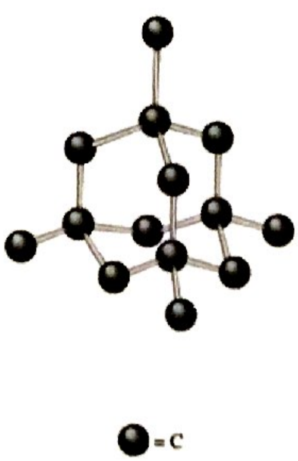
Diamond



Graphite

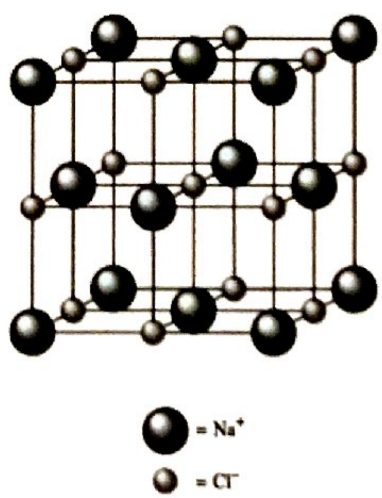
Types of Crystalline Solids

Network Covalent Solids
e.g. giant molecules
carbon (diamond, graphite)



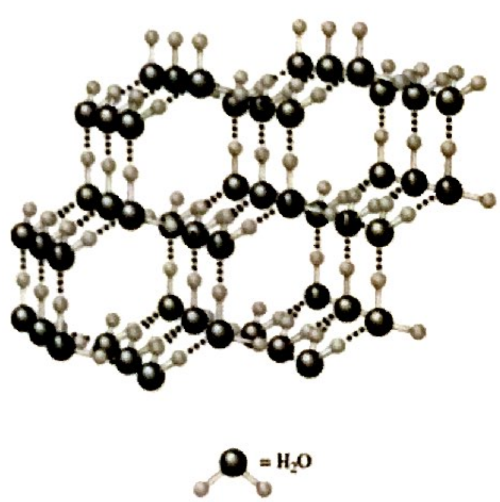
Diamond

Ionic Solids
e.g. salts like NaCl



Sodium chloride

Molecular Solids
e.g. protein crystals, sucrose



Ice

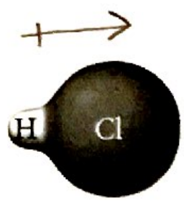
Non-Polar Covalent, Polar Covalent, or Ionic Bonding?

The attraction or "pull" on the bonded electron pair (i.e. electronegativity) determines bond polarity.

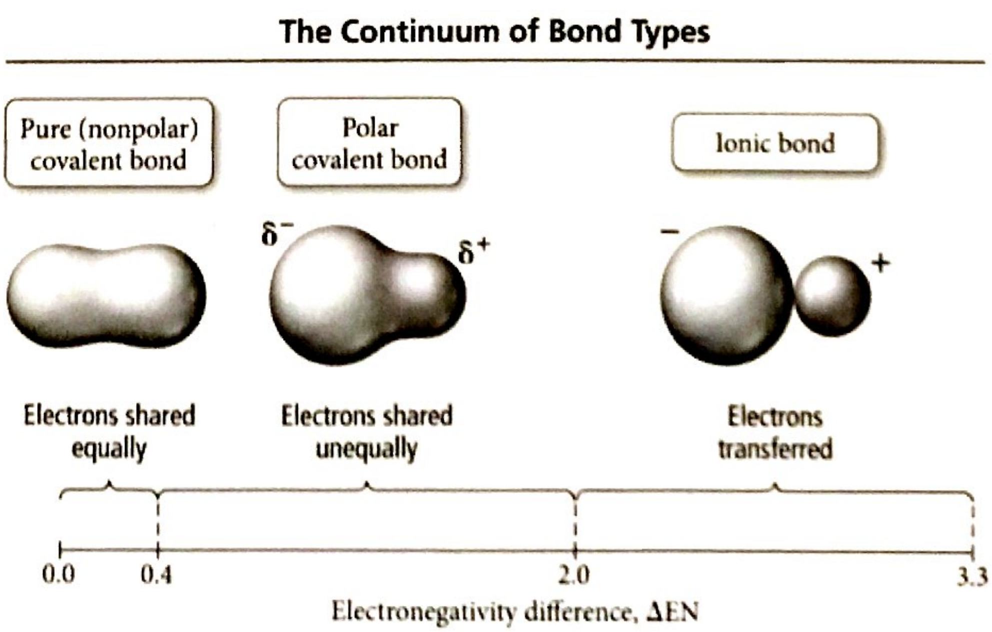
1. **Non-polar** covalent bond: bonding electrons are shared equally by the bonded atoms.
   - Electronegativity difference between atoms (ΔEN) < 0.4.
   - Examples:
     i. Diatomic molecules (Br_2, I_2, N_2, etc)
     ii. Any C-H bond.

2. **Polar** covalent bond: bonded atoms have an unequal attraction for the shared electrons.
   - 0.4 < ΔEN < 2.0
   - The atom with the greatest electronegativity has a greater attraction for the shared electrons, so they claim a greater amount of electron density.
   - The uneven electron density creates a dipole: a partial − charge on the atom with higher electron density and a partial + charge on the atom with lower electron density.
   - An arrow is used to represent the dipole (sometimes called a dipole moment): the arrow points towards the − pole (i.e. the most electronegative atom) and has a crossed tail at the + pole (least electronegative atom).

$\overset{+\!\!\longrightarrow}{H—F}$ *or* $\overset{\delta^+}{H}—\overset{\delta^-}{F}$



3. **Ionic bond**: a bond where the electronegativity difference between the two atoms is so extreme that one atom takes custody of all the contested electrons! (ΔEN > 2.0)



Most chemical bonds are somewhere between purely ionic and purely covalent!