There are three general classes of bonds that form between atoms. You can predict which will form by classifying the atoms as metals or nonmetals:

- Metal + metal: metallic bond
  - Au-Ag alloy
- Metal + nonmetal: ionic bond
  - MgCl$_2$
- Nonmetal + nonmetal: covalent bond
  - SO$_2$ or CH$_4$

Some compounds can contain both ionic and covalent bonds such as K$_2$SO$_4$... the sulfate ion is held together with covalent bonds... the potassium ions are ionically bonded to the sulfate ions.

Acids are exceptions... they are ionic only when dissolved.

Many ions can be explained because they have gained or lost electrons and attain a noble gas configuration. For example: P$^3-$ S$^2-$ Cl$^-$ Ar K$^+$ Cu$^{2+}$ all have the same electron arrangement: 1s$^2$ 2s$^2$ 2p$^6$ 3s$^2$ 3p$^6$

The importance of this configuration is that this is one reason why ions form. After these ions form, they stick together in a crystal lattice because opposites attract:

++ + - + - + - + - + - + -
- - + + + - + - + - + - + +

There are other reasons why some ions (ex: Cu$^+$ or Zn$^{2+}$) form.

The covalent bond between two atoms depends on the balance of attractions between one atom’s + nucleus and the other atom’s – electrons and the proton-proton repulsions as well as electron-electron repulsions.

If two atoms have half-filled orbitals, the interactions balance at a small enough distance so the e$^-$’s can be close to both nuclei at the same time... this is a covalent bond.

Lewis symbols consist of the atomic symbol surrounded by valence electrons. The four sides represent the four valence orbitals. Atoms are usually shown in their excited states. (Families II, III, & IV can also be in their “ground state.”)

$$\text{Li} \cdot \text{Be} \cdot \text{B} \cdot \text{C} \cdot \text{N} \cdot \text{O} \cdot \text{F} \cdot \text{Ne}$$

Ions include brackets and charges. Positive ions show no valence electrons while negative ions show an octet.

$$[\text{Li}]^+ \quad [\text{Mg}]^{2+} \quad [\text{O}]^{2-}$$
Example: CO$_2$

Draw the Lewis symbols for each atom.

\[ \begin{array}{ccc}
\text{O} & \text{C} & \text{O} \\
\text{O} & \text{C} & \text{O} \\
\text{O} & \text{C} & \text{O}
\end{array} \]

Connect the unpaired electrons.

Clean up your drawing.

<table>
<thead>
<tr>
<th>Compound</th>
<th>Molecular Conducts as Solid</th>
<th>Ionic Conducts as Solid</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>NO</td>
<td>NO</td>
</tr>
<tr>
<td>Conducts as Liquid</td>
<td>NO</td>
<td>YES</td>
</tr>
<tr>
<td>Conducts in Solution</td>
<td>NO</td>
<td>YES</td>
</tr>
<tr>
<td>Conducts as Gas</td>
<td>NO</td>
<td>YES</td>
</tr>
<tr>
<td>Hardness</td>
<td>soft</td>
<td>hard</td>
</tr>
<tr>
<td>MP / BP</td>
<td>low</td>
<td>high</td>
</tr>
<tr>
<td>Bonding</td>
<td>covalent</td>
<td>ionic</td>
</tr>
<tr>
<td>Examples</td>
<td>He, CH$_4$, CO$_2$, C$<em>6$H$</em>{12}$O$_6$, NaCl, KI, AgNO$_3$</td>
<td></td>
</tr>
</tbody>
</table>

You will be given a chart of electronegativity values. Memorize (F = 4.0) (O = 3.5) and (Cl = 3.0).
The noble gases have no values… no bonds.
Large electronegativity in the upper right of the per. table and small in the lower left portion of the table.
Classify the bond between any two atoms by subtracting their electronegativity values ($\Delta e$)
- Non-polar covalent $0 < \Delta e < 0.5$
- Polar covalent $0.5 \leq \Delta e < 1.7$
- Ionic $\Delta e \geq 1.7$
The more electronegative atom is more negative.
Polar covalent bonds have partial charges $\delta^+$ and $\delta^-$

Use VSEPR theory to predict the shape of molecules. The Steric Number (the # of lone pairs + bonded atoms) relates the shape of the electron pairs around a central atom.
- 1=linear, 2=linear, 3=trigonal planar, 4=tetrahedral
If each shape is symmetrical, the bond dipoles will cancel resulting in a nonpolar molecule.
If a shape has lone pairs of electrons on the central atom, the shape is often unsymmetrical, the molecule is polar.
Polar molecules and ions dissolve well in polar solvents while nonpolar molecules dissolve in nonpolar solvents.
“Like Dissolves Like”