PREDICTING MOLECULAR SHAPE

The shape of simple molecules (and parts of larger molecules) can be easily predicted using the VSEPR model.

VSEPR = Valence Shell Electron Pair Repulsion Model

- Each BOND or LONE PAIR OF ELECTRONS around an atom will try to move itself as far away from other bonds or lone pairs as possible!

For the two red circles to be farthest apart, they must be 180 degrees apart.

ANY diatomic (two-atom) molecule is linear, but only some three-atom molecules are!
For the three red circles to be farthest apart, they spread out so that each is 120 degrees from the others!
These hydrogen atoms might appear at first glance to be 90 degrees apart, but remember that molecules exist in THREE DIMENSIONS, not two!

Each hydrogen atom is actually 109.5 degrees apart, forming a TETRAHEDRON.

To see the tetrahedron in three dimensions WITHOUT buying a molecular model kit, just take four balloons, blow them up, and then tie them together. The knot will be the central atom, and the balloons will line themselves up to be 109.5 degrees apart.
### VSEPR shapes

* "Groups" can be either BONDS or LONE PAIRS!

<table>
<thead>
<tr>
<th>Groups around central atom</th>
<th>Shape</th>
<th>Bond angle(s) in degrees</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>linear</td>
<td>180</td>
</tr>
<tr>
<td>3</td>
<td>trigonal planar</td>
<td>120</td>
</tr>
<tr>
<td>4</td>
<td>tetrahedral / pyramidal / bent</td>
<td>109.5</td>
</tr>
<tr>
<td>5</td>
<td>trigonal bipyramidal (and derivatives)</td>
<td>90 and 120</td>
</tr>
<tr>
<td>6</td>
<td>octahedral (and derivatives)</td>
<td>90</td>
</tr>
</tbody>
</table>

5 and 6 violate "octet rule"
More on "4 things around a central atom":

- A compound that obeys the octet rule can have a maximum of four groups around its central atom. But we describe the molecular shape based on how ATOMS are arranged around the center. What if some of those groups aren't atoms, but pairs of UNSHARED electrons?

\[
\text{CH}_4: \quad H - C - H - H
\]

This atom is behind the paper!

\[
\text{H}_2\text{O}: \quad \begin{array}{c}
\text{H} \\
\text{H} \\
\text{O} \\
\text{H}
\end{array}
\]

With two ATOMS and two LONE PAIRS, we call the shape "BENT".

\[
\text{NH}_3: \quad \begin{array}{c}
\text{N} \\
\text{H} \\
\text{H} \\
\text{H}
\end{array}
\]

With three ATOMS and one LONE PAIR, we call the shape "PYRAMIDAL".

\[
\begin{array}{c}
\text{H} \\
\text{C} \\
\text{H} \\
\text{H}
\end{array}
\]

These atoms are in the plane of the paper!

With four ATOMS around the center, we call the shape "TETRAHEDRAL".
There are five atoms bonded to the central phosphorus atom, and they will attempt to get as far apart as possible from one another!

The top and bottom atoms are 90 degrees apart from the atoms around the center.

The atoms around the center are 120 degrees apart from each other.

There are actually two different bond angles in this structure. It's called trigonal bipyramidal.

There are several derivatives of the trigonal bipyramidal shape (like the tetrahedral shape) - depending on how many things around the central atom are atoms!
There are six atoms bonded to the central sulfur atom, and they will attempt to get as far apart as possible from one another!

All bond angles in this arrangement are 90 degrees!

This shape is called **OCTAHEDRAL**, since it has eight sides.

Like the tetrahedral and trigonal bipyramidal arrangements, there are several derivatives of the octahedron - depending on how many of the six things around the center are atoms!
Examples:

**CCl₄**

Shape? There are **FOUR** different groups attached to the central atom. This gives a tetrahedral set of angles. Since all four groups are other atoms, the molecule is TETRAHEDRAL.

**CS₂**

Shape? There are **two** different groups attached to the central carbon. Both are other atoms, so this molecule is LINEAR.

**NF₃**

Shape? There are **FOUR** groups bonded to nitrogen, giving it a tetrahedral set of angles. Three of the groups are other atoms, and the fourth is a lone pair, so this molecule is PYRAMIDAL.
**tetrahedral**

**linear**

**pyramidal**
This molecule has two centers, each of which is surrounded by THREE groups, giving each center a TRIGONAL PLANAR geometry.

This molecule has three groups around the central atom, and all of them are atoms - so it's TRIGONAL PLANAR.
Each carbon is TRIGONAL PLANAR

\[ \text{C}_2\text{H}_4 \] "ethene"

120°

120°

"formaldehyde"

TRIGONAL PLANAR
VSEPR and large molecules

- Large molecules have more than one "center" atom
- Describe the molecule by describing the shape around each "center".

\[ \text{C}_3\text{H}_6 : \quad \text{H} \quad \text{H} \quad \text{H} \]
\[ \text{H} - \text{C} - \text{C} - \text{C} - \text{H} \]
\[ \text{H} \quad \text{H} \quad \text{H} \]

Each of the three carbon centers is TETRAHEDRAL, since each are surrounded by four groups.

\[ \text{C}_3\text{H}_3\text{CH}_2\text{OH} : \quad \text{H} \quad \text{H} \]
\[ \text{H} - \text{C} - \text{C} - \text{O} - \text{H} \]
\[ \text{H} \quad \text{H} \]

The shape around this oxygen atom is BENT.

These carbon atoms have TETRAHEDRAL geometry.
All bond angles in the propane molecule are 109.5 degrees.

Like propane, the bond angles in ethanol are also close to 109.5 degrees.
POLARITY and shape:

- A polar molecule has an uneven distribution of electron density, making it have ends (poles) that are slightly charged.

POLARITY influences several easily observable properties.

- Melting point. (Polar substances have higher melting points than nonpolar substances of similar molecular weight.)
- Boiling point. (Polar substances have higher boiling points than nonpolar substances of similar molecular weight.)
- Solubility. (Polar substances tend to dissolve in other polar substances, while being insoluble in nonpolar substances. Nonpolar substances dissolve other nonpolar substances, and generally have poor solubility in polar solvents.)

- Polar molecules contain POLAR BONDS arranged in such a way that they do not cancel each other out.

... but how can we tell whether or not a bond will be POLAR? Use experimental data on ELECTRONEGATIVITY!

ELECTRONEGATIVITY:
- A measure of how closely to itself an atom will hold shared electrons

- A bond where there is a LARGE electronegativity difference between atoms will be either POLAR or (for very large differences) IONIC: \( \text{chart, H}_2\text{O} \)
- A bond with little or no electronegativity difference between atoms will be NONPOLAR
ELECTRONEGATIVITY TRENDS

- You may look up electronegativity data in tables, but it helps to know trends!

INCREASING ELECTRONEGATIVITY:

<table>
<thead>
<tr>
<th>2</th>
<th>3</th>
<th>4</th>
<th>5</th>
<th>6</th>
<th>7</th>
</tr>
</thead>
<tbody>
<tr>
<td>IA</td>
<td>IIA</td>
<td>IIIA</td>
<td>IVA</td>
<td>VA</td>
<td>VIA</td>
</tr>
<tr>
<td>Li</td>
<td>Be</td>
<td>B</td>
<td>C</td>
<td>N</td>
<td>O</td>
</tr>
<tr>
<td>Na</td>
<td>Mg</td>
<td>Al</td>
<td>Si</td>
<td>P</td>
<td>S</td>
</tr>
<tr>
<td>K</td>
<td>Ca</td>
<td>Sc</td>
<td>Ti</td>
<td>V</td>
<td>Cr</td>
</tr>
<tr>
<td>Rb</td>
<td>Sr</td>
<td>Y</td>
<td>Zr</td>
<td>Nb</td>
<td>Mo</td>
</tr>
<tr>
<td>Cs</td>
<td>Ba</td>
<td>La</td>
<td>Hf</td>
<td>Ta</td>
<td>W</td>
</tr>
<tr>
<td>Fr</td>
<td>Ra</td>
<td>Ac</td>
<td>Rf</td>
<td>Db</td>
<td>Sg</td>
</tr>
</tbody>
</table>

Notes:

1. FLUORINE is the most electronegative element, while FRANCINIUM is the least!
2. All the METALS have low electronegativity, and metal/nonmetal combinations form IONIC bonds.
3. HYDROGEN is similar in electronegativity to CARBON, so C-H bonds are considered NONPOLAR.
Examples:

\[ \text{CF}_4 \]
\[ \begin{array}{c}
\text{C} : 1 \times 4 \\
\text{F} : 4 \times 7 \\
\hline
32 \text{e}^-
\end{array} \]

1) Polar bonds? Yes, C-F bonds should be polar.  
2) Shape? This is a tetrahedral molecule. The molecule is symmetric (there's no positive side and no negative side), so the molecule is NONPOLAR.

\[ \text{CH}_3 \text{F} \]
\[ \begin{array}{c}
\text{C} : 1 \times 4 \\
\text{H} : 3 \times 1 \\
\text{F} : 1 \times 7 \\
\hline
14 \text{e}^-
\end{array} \]

1) Polar bonds? Yes, C-F bonds should be polar.  
2) Shape? This is a tetrahedral molecule. But it's not symmetric, so the fluorine side will be negative and the hydrogen side positive. This is a POLAR molecule!

\[ \text{CH}_2 \text{F}_2 \]
\[ \begin{array}{c}
\text{C} : 1 \times 4 \\
\text{H} : 2 \times 1 \\
\text{F} : 2 \times 7 \\
\hline
20 \text{e}^-
\end{array} \]

1) Polar bonds? Yes, C-F bonds should be polar.  
2) Shape? This is a tetrahedral molecule. But it's not symmetric, so the fluorine side will be negative and the hydrogen side positive. This is a POLAR molecule!

\[ \text{CO}_2 \]
\[ \begin{array}{c}
\text{C} : 1 \times 4 \\
\text{O} : 2 \times 6 \\
\hline
16 \text{e}^- \\
\end{array} \]

1) Polar bonds? Yes. C=O should be polar.  
2) Shape? Linear and symmetric, so NONPOLAR.
Fluorine is able to pull electron density through the molecule, as it is being opposed by much less electronegative hydrogen atoms.

In 2D, the fluorine atoms appear to be on the opposite sides of the molecule, but in 3D they are on the same side.
VALENCE BOND THEORY

- an attempt to explain why molecules behave in the way that the VSEPR model predicts.
- Describes the formation of bonds in terms of the OVERLAP of ORBITALS from the bonding atoms.

1. Bonds are formed when two atoms are close enough together so that their ORBITALS OVERLAP (share the same space).

2. Each SET of overlapping orbitals can contain at most a total of TWO electrons. So, two orbitals with one electron each may bond. An orbital with two electrons can only bond with an EMPTY orbital (This is called a COORDINATE COVALENT BOND.)

These 1s orbitals overlap to form what we call a "sigma bond" with overlap BETWEEN the two atomic nuclei.
Hybridization

- Look at carbon's electron configuration:

\[
\begin{array}{c}
\text{Energy} \\
\text{1s} \\
\text{2s} \\
\text{2p} \\
\end{array}
\]

Valence

You would expect that carbon would form several different kinds of bonds in a molecule like methane. But, methane's bonds are experimentally all identical. How does carbon form the four equivalent C-H bonds we see in methane?

\[
\text{CH}_4: \quad \text{H} - \text{C} - \text{H}
\]

We observe that these bonds are IDENTICAL! Same bond energy, distance, and angle.
- In valence bond theory, atomic orbitals can **combine** to make new orbitals that can then go on to bond with other molecules.

- When orbitals combine to make **hybrid orbitals**, ...

1. The overall **NUMBER OF ORBITALS** does not change.
2. The overall **NUMBER OF ELECTRONS** around the atom does not change.
3. The energy of the orbitals is between the energies of the orbitals that combine.

These sp$^3$ orbitals were formed from the combination of carbon's original 2s and 2p orbitals. These orbitals are all identical, and are spread 109.5 degrees apart from one another.

Hybrid orbitals are named from the orbitals that go into making the hybrid. 2s + 3 2p orbitals = "sp$^3"! ℙ402: picture of hybrids (£p401, 1045 e⁻)
Types of hybrid orbitals:

<table>
<thead>
<tr>
<th>Hybrid type</th>
<th>Number of orbitals</th>
<th>Molecular shape</th>
</tr>
</thead>
<tbody>
<tr>
<td>sp</td>
<td>2</td>
<td>linear</td>
</tr>
<tr>
<td>sp2</td>
<td>3</td>
<td>trigonal planar</td>
</tr>
<tr>
<td>sp3</td>
<td>4</td>
<td>tetrahedral (or derivatives)</td>
</tr>
<tr>
<td>sp3d</td>
<td>5</td>
<td>trigonal bipyramidal (or derivatives)</td>
</tr>
<tr>
<td>sp3d2</td>
<td>6</td>
<td>octahedral (or derivatives)</td>
</tr>
</tbody>
</table>

$\Phi_{02} = \text{picture of hybrids} (\Phi_{01} \mid 0_{15})$
MULTIPLE BONDS and VALENCE BOND THEORY

Valence bond theory provides an explanation of multiple (double and triple) bonding that explains some interesting observations about these kinds of bonds.

Each carbon has a TRIGONAL PLANAR geometry. This suggests that the carbons are "sp2 hybridized".

One unchanged 2p orbital

Three sp2 hybrids that are 120 degrees apart

"Original" unbonded carbon atom

Carbon atom with sp2 hybrid orbitals
The 2p orbitals overlap above and below the axis between the two carbon atoms. This OFF-AXIS overlap is called a PI BOND.

The sp2 hybrid orbitals overlap ON THE AXIS between the two carbon atoms. This bond is called a SIGMA BOND.

As you can see, the carbon-carbon double bond in ethylene is made up of TWO DIFFERENT KINDS OF BONDS!
Some notes on sigma and pi bonds:

1. SIGMA bonds are formed when orbitals overlap along the axis between two atoms. These bonds have good overlap between the bonding orbitals, meaning that they are strong. Single bonds are always sigma bonds. Double and triple bonds contain one sigma bond each.

2. PI bonds are formed when off-axis orbitals (usually p orbitals) overlap. Since the overlapping orbitals do not face each other as in the sigma bond, the overlap in pi bonds tends to be poorer than in sigma bonds. As a result, pi bonds tend to be weaker than sigma bonds. Double bonds contain a single pi bond, and triple bonds contain two pi bonds.

* Experimentally, we observe that the bond energy of the C=C bond is less than the bond energy of two C-C bonds. This suggests that the second bond in a double bond is different from the first!

3. Molecules may rotate around SIGMA bonds, since rotation around the axis between two atoms will not affect the overlap and break the bond. Off-axis PI BONDS prevent rotation because rotation would break the pi bond.
- Consider this molecule: \((\text{CH}_2\text{Cl})_2\) 

"1,2-dichloroethane"

... are these two structures different?

No! The molecule is free to rotate around the C-C single (sigma) bond, and we do not observe two different versions of 1,2-dichloroethane. Both of the forms drawn above are equivalent.
The molecule is free to rotate about the carbon-carbon bond!
... now consider "1,2-dichloroethene": \((\text{CHCl}_2)\)

... are these two structures different?

YES! The two carbon atoms in these structures are held together by a DOUBLE BOND, which contains a pi bond. The molecule cannot rotate around the C=C double bond without breaking the pi bond, so the form with the two chlorine atoms on opposite sides cannot freely flip over to the form with the chlorine atoms on the same side.

These two Lewis structures actually represent DIFFERENT MOLECULES. They are called ISOMERS, since they have the same chemical formula but different arrangements of atoms.
For this rotation to take place, the PI BOND must break and then re-form!
trans 1,2-dichloroethene
BOILING POINT: 47.5 °C
POLARITY: NONPOLAR (0 D dipole moment)
DENSITY: 1.26 g/mL

cis 1,2-dichloroethene
BOILING POINT: 60.3 °C
POLARITY: POLAR (1.9 D dipole moment)
DENSITY: 1.28 g/mL

* As you can see, some of the properties of these two molecules are very different! The presence of the pi bond (part of the double bond) in each of these molecules means conversion from one form to the other requires a chemical reaction instead of a simple rotation.

* Double (and triple) bonds prevent rotation and "fix" the structure of a molecule. This is easily explained by valence bond theory!
SOLIDS AND LIQUIDS

- Here's a brief review of the atomic picture of gases, liquids, and solids.

**GASES**

* Kinetic theory says...
  * Gas molecules are small compared to the space between them.
  * Gas molecules move in straight lines until they hit another gas molecule or the walls of the container.
  * There are no attractive or repulsive forces between gas molecules except during a collision.
  * When gas molecules collide, energy may be transferred, but no energy is lost as heat.
  * The temperature of a gas is proportional to the average kinetic energy of the gas molecules.

Gases are FLUID, COMPRESSIBLE, and DIFFUSE (NOT DENSE)!

- The properties of different gases are very similar to one another. At moderate conditions, different gases obey the simple IDEAL GAS EQUATION.

\[ PV = nRT \]
LIQUIDS

* Molecules are much closer together than in the gas phase.

* Molecules are free to move around each other, but there is much less freedom of motion than in the gas phase.

* Molecules in the liquid state are held together by attractive forces that we will call INTERMOLECULAR FORCES.

Liquids are FLUID, DENSE, and INCOMPRESSIBLE!

- The properties of different liquids are often very different from one another. Compare liquids like water and motor oil, which are different enough so that they won’t readily mix with one another!

\[
\text{no simple } \frac{PV}{nRT} \text{ for liquids!}
\]
SOLIDS

* Molecules are usually packed closer together in the solid phase than in the gas or liquid phases.

* Molecules are not free to move around each other as in the liquid phase. Molecular/atomic motion in the solid phase is limited to vibration.

* Most solids have a regular structure - unlike liquids or gases. This structure is called a CRYSTAL LATTICE.

* Molecules are held together by INTERMOLECULAR FORCES. These are usually stronger than in the liquid phase.

Solids are RIGID, DENSE, and INCOMPRESSIBLE!

- As for the liquids, the properties of different solids often differ considerably. Compare a sample of candle wax to a sample of quartz.
PHASE CHANGES

- To understand solids and liquids at the molecular level, it will help to examine PHASE CHANGES in a little more detail.

A quick review of the phase changes...

<table>
<thead>
<tr>
<th>Phase change</th>
<th>Description</th>
<th>Energy change</th>
</tr>
</thead>
<tbody>
<tr>
<td>Melting</td>
<td>Solid to liquid</td>
<td>Endothermic</td>
</tr>
<tr>
<td>Sublimation</td>
<td>Solid to gas</td>
<td>Endothermic</td>
</tr>
<tr>
<td>Vaporization</td>
<td>Liquid to gas</td>
<td>Endothermic</td>
</tr>
<tr>
<td>Deposition</td>
<td>Gas to solid</td>
<td>Exothermic</td>
</tr>
<tr>
<td>Freezing</td>
<td>Liquid to solid</td>
<td>Exothermic</td>
</tr>
<tr>
<td>Condensation</td>
<td>Gas to liquid</td>
<td>Exothermic</td>
</tr>
</tbody>
</table>

\[ \Delta H_{\text{fus}} : \text{enthalpy change to melt 1 mol of solid} \]

\[ \Delta H_{\text{vap}} : \text{enthalpy change to vaporize 1 mol of liquid} \]
Phase diagrams are a convenient way to show experimental data on when bulk phase changes occur.

The curves on the phase diagram represent bulk phase changes.

The triple point is a set of conditions where all six phase changes occur at the same time; in other words, all three phases coexist.

The critical point occurs where there stops being a distinction between the gas and liquid states. This occurs at high pressure and high temperature, where the substance has the density of a liquid but the fluidity of a gas. This is called a supercritical fluid. Supercritical fluids - like supercritical carbon dioxide - are often used as environmentally friendly solvents.

The normal boiling point and freezing points are on the curves. The normal melting point is the point on the solid/liquid curve at 1 atm, while the normal boiling point is on the liquid/gas curve at 1 atm!