

**Covalent Lewis Structures**

Covalent bonding always occurs with the valence “s” and “p” electrons of the bonding atoms. Two atoms can share 2, 4, or 6 electrons (single, double, triple bonds) depending on what is needed to complete an octet (or duet for Hydrogen).

Lone pair electrons are full “s” or “p” orbitals that do not share their electrons with another atom.

Usually atoms form bonds by sharing an electron from each atom in the bond.

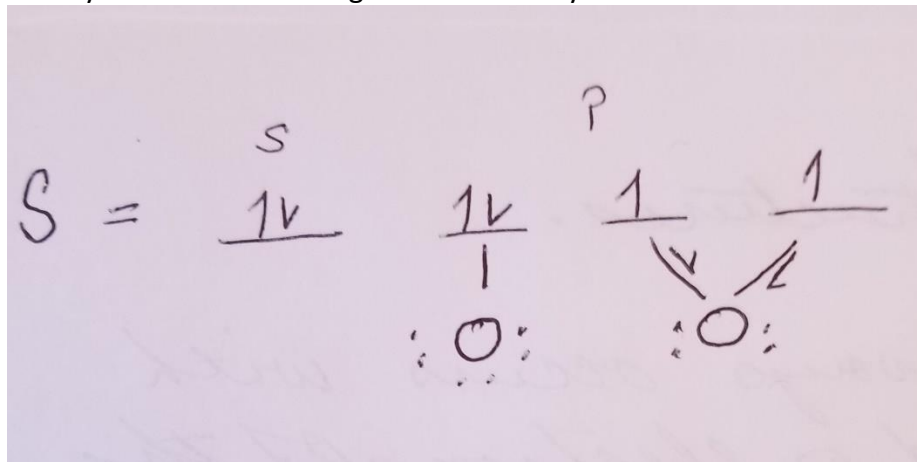
An example is: H—H (One electron from each hydrogen atom is shared in the bond between them.)

Sometimes, if an atom needs 2 electrons and another atom has 2 electrons in a full orbital, the atom that needs the 2 electrons can bond to the orbital on the other atom that has the full orbital (coordinate covalent bond).

**An example is SO<sub>2</sub>:** Sulfur has 6 electrons and needs 2 more. It needs to share 2 of its electrons with 2 electrons on other atoms. Both Oxygens have 6 electrons and need to share 2 more electrons to fill their octet.

This is what happens.

Sulfur is the “central atom” or the atom that the other electrons bond with. The central atom is usually the least electronegative and usually the one that is a lone atom of its kind.



Sulfur has 2 electrons in the “s” and 4 electrons in the “p” orbitals.

The first oxygen shares 2 electrons in two bonds (**double bond**) with sulfur completing the octet on sulfur and on the first oxygen. (The arrows represent electrons.)

The second oxygen can bond to the sulfur but this oxygen does not share any of its own electrons. It shares the full p-orbital electrons on the sulfur (**coordinate covalent bond**). Sulfur doesn't increase its number of electrons (still has an octet) but the second oxygen now completes its octet.

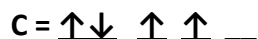
Notice, there is still a full “s” orbital on the sulfur (lone pair).

**Note:** This “lone pair” is a possible site for another coordinate bond for another oxygen to bond to. SO<sub>3</sub> could be a possible compound to form from this arrangement.

The examples I just showed you is how we are going to form our Lewis Structures. We will be using the valence electrons of the “s” and “p” orbital for the “central atom” to share and bond with “terminal atoms”. The terminal atoms are the atoms attached to the central atom.

We will always have four lines representing the one s and three p orbitals of the valence shell around the central atom. Fill in the orbitals just as you would fill in an orbital diagram.

Now, let’s look at **carbon**. Carbon has 4 valence electrons, 2 in the “s” and 2 in the “p”. Carbon needs 4 more electrons to fill its octet. This would look almost impossible using the orbital diagram for carbon:



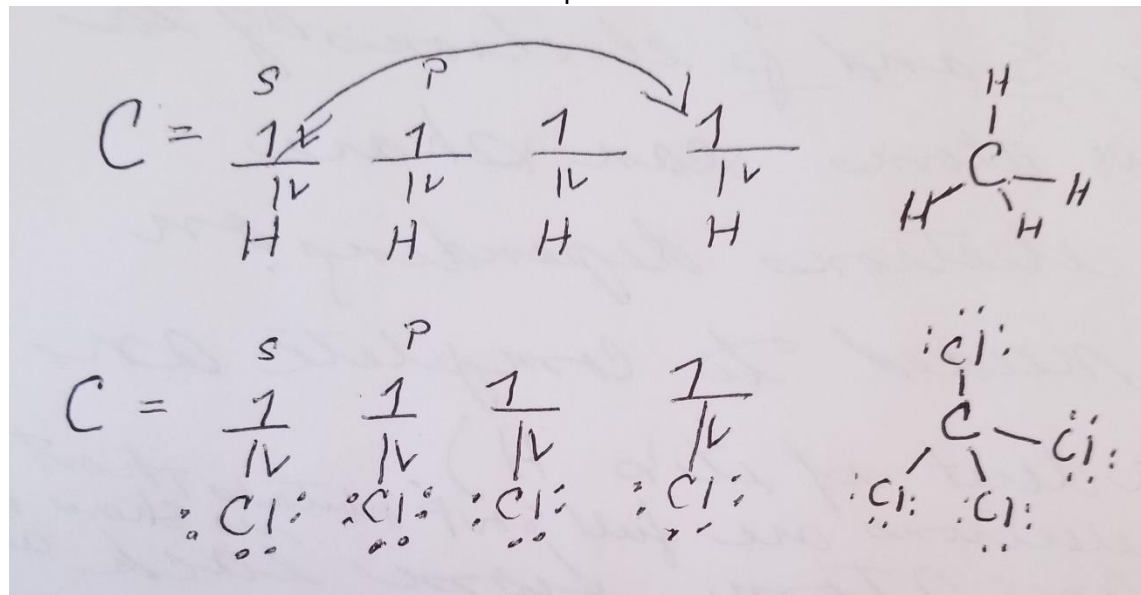
It looks as though carbon only has 2 electrons that can share and one empty orbital. (half-filled orbitals are the only orbitals that can share with another atom and produce a single bond.

A common compound of carbon is **CH<sub>4</sub>**.

Single bonds are mainly made with Hydrogen and the Halogens (F, Cl, Br, and I). Halogens **always** form **single bonds**.

It appears that carbon only has 2 electrons that can single bond. What happens is that carbon advances an electron from the “s” orbital to the empty “p” orbital producing 4 half-filled orbitals. Now carbon can accommodate 4 hydrogen atoms sharing an electron with each.

The same would hold true for **CCl<sub>4</sub>**. Only there would be 6 electrons around each Cl atom in the Lewis structure. A “bar” between atoms represents 2 electrons in a bond.



Notice I did not draw methane or carbon tetrachloride as a 2-dimensional figure. It has 3-dimensional shape. They are both tetrahedral. The Shape Chart that I linked to this lesson is very, very important. **You need to print a copy of this chart in order to do the practices.** The chart will tell you what shape you made and what type of **hybridization** your central atom has gone through.

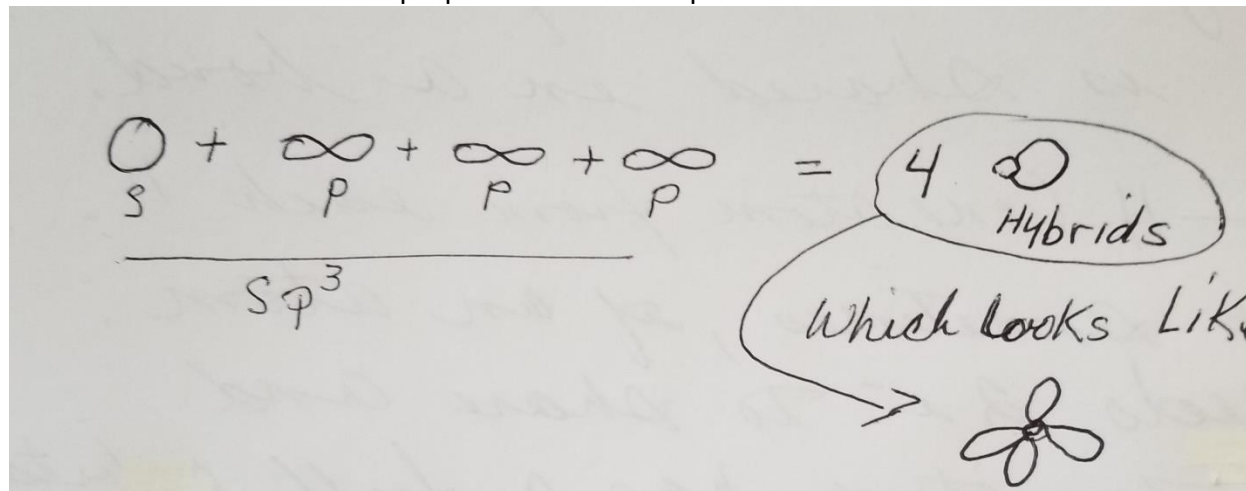
## Hybridization—Changing of an orbital shape to accommodate bonding.

Hybridization bring us to the VSEPR Theory. It stands for valence shell electron pair repulsion theory. It is highly dependent on lone electron pairs.

Before discussing VSEPR we need to discuss why molecules change shape to begin with and how they possible can bond with orbitals of different shapes and energies.

When Carbon advanced and electron from the “s” to the “p” orbital (carbon will do this almost every time it bonds to anything.) it is taking an electron that has lower energy and putting it into an orbital that has higher energy. “Oops”, that isn’t a good thing because electrons always go to lowest energy. Something has to happen so that these orbitals are “degenerate” or have the same energy. Also, the “s” orbital is small and inside the three “p” orbitals so it would be very difficult to find room to bond to another atom.

What happens to **all** central atoms is that they **hybridize** forming four identical orbitals from the one “s” and three “p” orbitals. They form four “hybrid” orbitals called **sp<sup>3</sup> hybridization**. The orbitals have a little bit of the properties of both shaped orbitals.



This is a tetrahedral shape of the 4 hybrid orbitals. This is why CH<sub>4</sub> and CCl<sub>4</sub> are tetrahedral.

We can have sp<sup>3</sup>, sp, sp<sup>2</sup>, sp<sup>3</sup>d, and sp<sup>3</sup>d<sup>2</sup> hybridization.

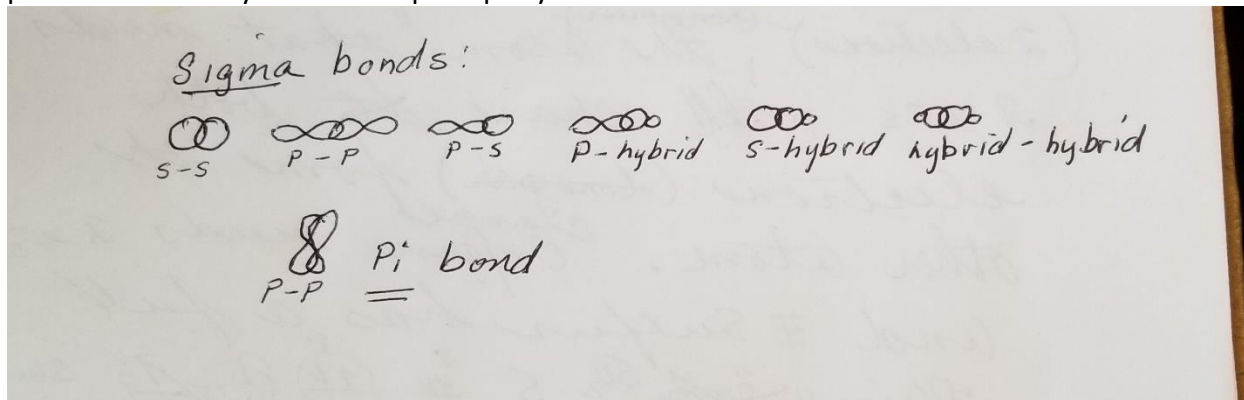
sp<sup>3</sup> is most common but with exceptions to the octet rule, we can have “empty d” orbitals on the same energy level as the valence s and p orbitals hybridize giving 5 and 6 atoms bonding to a central atom. We will be doing this in practice tomorrow and Tuesday.

There are two major types of bonding.

**Sigma** bonding (most common) and **Pi** bonding.

**Sigma** bonding is **end-on-end** bonding. **Pi** bonding is **side-by-side p-orbital** bonding.

All double bonds have one pi bond and all triple bonds have one pi bond. That is the only time there are pi bonds. But, a pi bond means the central atom cannot hybridize one of its p-orbitals. We may then have sp or sp<sup>2</sup> hybridization.



See p. 419-421 for hybrid orbitals and the way they bond.

Both the hybridization and the Lone pairs (unshared pairs) determine the shape of the molecule.

Molecules change shape and orientation in order to accommodate their areas of **high electron density** which I refer to as **HED's**.

Areas of high electron density are **bonds**, and **lone pairs** of electrons.

Single, double, and triple bonds all count as **one HED**.

We count HED's to determine the basic shape of our molecule and lone pairs to determine the specific shape of the molecule (See shape chart).

The **Worksheet for Lewis Structures** which I posted and linked to your lesson will be used to construct Lewis Structure for the molecule. **Please print off about 10 copies of this sheet**, or construct your own from the posted page but you will need the pages to work on these structures.

The **orbital structure** will give you the number of bonds and the number of lone pairs. They will add together to give you the number of **HED's**. From that information, you will determine the **shape** of the molecule. You will then **draw the 3-D shape** of your molecule and determine the **bond polarity** (electronegativity difference) and the **molecular polarity** (from shape chart). I will cover **Intermolecular Forces** next Tuesday. This will take only a few minutes (approx. 15 min).

I will tie up loose ends on Tuesday of next week and go over what will be on the test.

**Please watch all 3 videos. They are very important.**

I did not cover **formal charge** but I will talk about it in class. It will not be on the test.