Nuggets of Knowledge for Chapter 1 – Atoms and Molecules Chem 2310

I. Introduction to Organic Chemistry

- Organic chemistry began in the 1700's as the study of compounds which came from living things, like sugars, urea, and quinine.
 - These compounds were found to be much more complicated than chemicals like sodium chloride or magnesium sulfate which were obtained from minerals in the earth.
 - For a time, chemists believed that they contained a special "vital energy" and could not be created in the laboratory, unlike inorganic chemicals.
 - Eventually, however, they realized that the big difference was that the organic chemicals all contained a large proportion of the element carbon, while only a few types of minerals contained carbon. Once carbon's bonding properties were understood, chemists began to make progress in analyzing and synthesizing organic chemicals.
- Organic chemistry now involves the study of all chemicals containing carbon, whether or not they originated with living things. Since those early days, organic chemists have amassed a huge body of knowledge about the structure of organic chemicals and how they behave, and have learned to make many organic compounds that are not found in nature. A two semester course can barely scratch the surface of what is now known, and there is still more to be discovered.
- Organic chemistry is typically taken in the sophomore year of college, since it is a stepping stone to more advanced chemistry courses, such as biochemistry, and is also important for the study of botany, zoology, and medical fields such as nutrition, physiology, pharmacy, medicine, and dentistry.
- Many industries also depend on organic chemistry.
 - Pharmaceutical companies create and manufacture drugs, which are nearly all organic chemicals, since they are intended to interact with the organic chemicals the body is made of.
 - Petroleum, which is thought to have come from the remains of tiny microscopic animals in the ancient oceans, is a complex organic mixture found under the earth which can be separated and modified to make many of the products we use for fuel, including gasoline, diesel, kerosene, and heating oil.
 - Plastics, which can be made in an astounding variety of properties, are also made from petroleum.
 - Pesticides and fertilizers are also made from organic compounds.

- It is unusual that an entire branch of chemistry would be based on just one element. There are several reasons for this.
 - One reason is that there is such a staggering number of different organic chemicals. 95% of all of the known compounds are organic – and that total is over 15 million different compounds.
 - The reason there are so many organic compounds is that the element carbon has a unique ability to form complex structures. No other element will form stable chains, rings, and other combinations like carbon does.
 - Organic compounds are made up with a framework of these carbon structures with hydrogen, oxygen, nitrogen, and other elements attached in a variety of stable patterns.
 - The other reason is that this makes possible the very complex molecules that are required for life, such as DNA, proteins, carbohydrates, lipids, and so on. Because of this, understanding how organic chemicals behave is necessary to studying how living things work.

II. Review of Atomic Structure

Elementary particles

- All atoms are made of three elementary particles protons, neutrons, and electrons.
- Protons, which have a positive charge, and neutrons, which have no charge, are found in the nucleus of the atom.
- Electrons, which have a negative charge, are found outside the nucleus, and are the part of the atom which participates in chemical reactions.

The Periodic Table of Elements

- There are 90 kinds of atoms that can be found in nature (more can be made using nuclear reactions). These are called the elements. They are the fundamental substances that make up all of the matter in the universe.
- The different elements are distinguished by the number of protons in their nucleus, which is called the atomic number. The number of protons is what gives the atom its identity.
 - For example, the smallest element is hydrogen, with an atomic number of one. This means that all hydrogen atoms have one proton in their nucleus, and that any atom with only one proton must be a hydrogen atom, no matter how many neutrons or electrons it has.
- The elements are organized into the Periodic Table of Elements. It starts with hydrogen, which has an atomic number of 1, and then goes across to helium, atomic number 2, then starts a new row with lithium, atomic number 3, and so on.

- This arrangement of the elements is very powerful because it allows us to deduce many things about an element based on its position of the table.
 - The elements which are metallic in their elemental state are located to the left of the stairstep line, while elements which are gases or crystals in their elemental state (except hydrogen) are located to the right of the stair-step line.
 - Metals are further divided into main group metals, transition metals, and inner transition metals. Main group metals are found in the first two columns and under the stair step line; transition metals are located in the middle of the metals, and inner transition metals are in the two lines below the table (which actually fit before the transition metals).
 - A row on the Periodic Table is called a period.
 - Periods are numbered from the top hydrogen and helium make up the first period, lithium through neon the second, and so on.
 - The elements in a period are about the same size.
 - A column on the Periodic Table is called a family or group.
 - Elements in the same family tend to have similar chemical reactivity. However, they get larger as you go down, which changes their behavior some.
 - Some of the families have names, such as the alkali metals (the first group), the alkaline earth metals (the second group), the halogens (the second to last group), and the noble gases (the last group).
- Organic chemistry only uses a few of the elements on the Periodic Table.
 - Most organic chemicals are made of carbon (C), hydrogen (H), oxygen (O), and nitrogen (N).
 - Halogens, the second to last family on the Periodic Table, are also found in many organic chemicals. They are fluorine (F), chlorine (Cl), bromine (Br), and iodine (I).
 - Occasionally phosphorus (P) and sulfur (S) are also found.
 - Two metals, lithium (Li) and magnesium (Mg), can be connected to carbon to form compounds.
 - Other metals such as sodium (Na) and potassium (K) are often found as counterions in charged organic molecules.
 - Many other elements, including boron (B), aluminum (Al), chromium (Cr), manganese (Mn), mercury (Hg), palladium (Pd), platinum (Pt), iron (Fe), silver (Ag), zinc (Zn), copper (Cu), and osmium (Os), are found in reagents that react with organic chemicals.
 - You will need to memorize the positions of all of these elements except the transition

metals.

Electronegativity

- One of the most important properties of an element is its electronegativity. This is a measure of how strongly the atom pulls electrons towards it.
- Electronegativities of the elements have been measured and can be found in tables. However, relative values are usually more important that actual values; in other words, it is more important to know whether one element is more electronegative than another than to know the actual value. Many such comparisons can be made simply by looking at the positions of the elements on the Periodic Table.
 - The most electronegative element is fluorine, and the distance an element is from fluorine can give a rough estimate of how electronegative it is.
 - Metals have much lower electronegativities than nonmetals.
 - Electronegativity increases significantly as you go across a period (except for the noble gases).
 - Electronegativity decreases somewhat as you go down a family.
- When comparing two elements, it is important to consider whether they are in the same period or the same family. (Comparisons between elements in that are not in the same period or family are not usually necessary.)
 - If two elements are in the same period, the most important difference between them is their electronegativity, whereas the small size difference is not important.
 - If two elements are in the same family, the most important different between them is their size, which usually overwhelms differences in electronegativity.

Atomic Mass and Isotopes

- The mass of an atom is made up mostly of the protons and neutrons; the electrons are so tiny in comparison that their mass is negligible.
- An isotope is an atom with a specific number of neutrons.
 - For example, there are three common isotopes of hydrogen hydrogen-1, hydrogen-2, and hydrogen-3. The number gives the mass of each isotope, which can also be written as a superscript: 1H, 2H, or 3H.
 - All three have the same atomic number, since in order to be hydrogen atoms they must all have one proton. But one has no neutrons, one has one neutron and one has two neutrons.
 - Since they have different masses they have slightly different physical properties, and can also have different nuclear properties for example, 3H, also called tritium, is radioactive,

and can be traced using the beta particles that it gives off when it undergoes radioactive decay. However, their chemical behavior is very similar – for example, all three isotopes of carbon form the same kinds of bonds.

- Radioactive isotopes can be used in organic chemistry to "label" a specific atom so we can follow where it goes. Commonly used isotopes are hydrogen-2 or deuterium (usually labeled D); carbon-14, and oxygen-18.
- The mass of hydrogen given on the Periodic Table is 1.007, and is a weighted average of all of the naturally occurring isotopes of hydrogen. The heavier ones are very rare, so the average is only slightly higher than the mass of the most common isotope, 1H.

<u>Ions</u>

- Since the electrons are on the outside of the atom, they are the part of the atom involved in chemical reactions. Unlike protons or neutrons, electrons may be gained or lost or shared through interactions with other atoms.
- The negative charge of an electron exactly matches the positive charge on a proton. The relative numbers of protons and electrons determines the charge on the atom.
 - \circ When an atom has the same number of electrons as it has protons, then it is electrically neutral.
 - If the atom gains an electron, then the number of negative charges will outweigh the number of positive charges, and the atom will be negative.
 - If the atom loses an electron, then the number of positive charges will outweigh the number of negative charges, and the atom will be positive.
- When atoms lose or gain electrons and become charged, they are called ions. Positive ions are also called cations, and negative ions are also called anions.

Valence electrons

- Electrons are organized into energy levels. The lowest energy level is closest to the nucleus, and as they get farther away from the nucleus they are higher in energy.
 - \circ The lowest energy levels are filled first, since everything tends toward the lowest energy possible.
- The electrons in the highest occupied energy level are called the valence electrons. They are important because they are the ones which participate in chemical bonds. In fact, the number of valence electrons is one of the most important factors in determining what kinds of compounds an element will form.
- You can determine the number of valence electrons by counting across the Periodic Table (skipping transition metals).

• Valence electrons can be represented by dots around the symbol for the atom, called Lewis dot structures. The dots can be on any side of the atom, but they should be separated until there are four, then put in pairs.

III. Review of Ionic and Covalent Bonds

Overview of Ionic and Covalent Bonds

- Atoms form chemical bonds in order to become more stable. They become stable by filling up their outer energy level with electrons. For all atoms except hydrogen and helium, it takes eight valence electrons to fill this outer energy level. This has often been called the octet rule atoms are most stable when the have eight valence electrons.
- There are two ways in which an atom may accomplish this.
 - Metal atoms have only a few valence electrons, so they become stable by losing those electrons. The full energy level underneath becomes the valence energy level.
 - Nonmetal atoms have half or more of the valence electrons needed, so they become stable either by gaining electrons to complete the valence shell or by sharing electrons with another atom which also needs more to complete its outer layer.
- When one atom gives an electron to another atom, the atom that gained an electron becomes negative and the atom that lost an electron becomes positive. They are then attracted to each other because of the opposite charges, forming an ionic bond.
- When atoms share electrons, they are held together by that sharing, forming a covalent bond.
- A covalent bond is represented by drawing a line between the atoms, while an ionic bond is represented by placing the charged ions close to each other.
- Both ionic and covalent bonds are important in organic chemistry.
 - Stable organic compounds are made of nonmetals, and are usually neutral, containing only covalent bonds.
 - However, during a chemical reaction the atoms within the compound may become charged and are then attracted to ions of the opposite charge, so that compounds may have both covalent and ionic bonds in the same molecule.
 - Reagents that react with organic compounds may also be covalent or ionic.

Lewis structures of covalent compounds

- Compounds containing covalent bonds may be represented by Lewis structures, which show the bonds between atoms as well as the valence electrons still on the atoms.
- When drawing Lewis structures, use the following procedure:

- First, draw out each atom with its valence electrons (the Lewis structure for the atom).
- Then look at each one to see how many bonds it will need to make in order to become stable.
- Finally, form bonds between the atoms by connecting lone electrons to make covalent bonds.
- If there aren't enough H's to satisfy all of the atoms, double or triple bonds may be formed. Remember that each bond represents two shared electrons.
- You may have been taught in the past to simply add up the valence electrons and place them around the atoms so that all obey the octet rule, but this only works with very small molecules. With larger ones, it often gives unstable structures that do not correctly represent the structure of the molecule. To obtain the correct structure, you have to look at which atom the electrons come from. Only unpaired electrons can form bonds to other atoms.
- After you have drawn a Lewis structure, make sure it obeys each of the following rules.
 - 1. All atoms (except hydrogen) must have 8 electrons around them.
 - 2. All atoms must have the correct number of bonds and electron pairs.
 - 3. The total number of electrons in the compound must equal the valence electrons from the original atoms.
- There may be more than one correct way to draw a Lewis structure for a given molecular formula. The two structures may either represent the same molecule drawn differently or two different molecules.
 - If the same atoms are connected to each other, but are arranged differently in space, then the two structures are just two ways of drawing the same molecule.
 - If the same atoms are NOT connected together, then the structures represent two different compounds.

Polarity of Covalent Bonds

- The polarity of a covalent bond is determined by the electronegativity of the two atoms in the bond.
 - When the two atoms have the same electronegativity, a nonpolar bond is formed, which means that the electrons are equally shared between the two atoms.
 - When one of the atoms is more electronegative than the other, a polar bond is formed, which means that the electrons are not shared equally, but are closer to the electronegative atom.

- Bonds which have only a slight polarity are often referred to as nonpolar bonds because the polarity is too small to affect their properties or reactivity.
- When a polar bond is formed, the atoms become slightly charged because of this unequal sharing of the electrons.
 - The more electronegative element, which is pulling the electrons towards it, becomes slightly negative.
 - The less electronegative element, which is having the electrons pulled away from it, becomes slightly positive.
 - The bigger the difference in electronegativity between the two atoms, the more polar the bond will be and the larger these partial charges will be.
 - These partial charges are different from full positive or negative charges that come from the loss or gain of electrons because the electrons aren't actually lost or gained they are just closer or farther away from the nucleus of the atom.
- Multiple bonds (double and triple bonds) are more polar than single bonds.
- The partial charges in a polar bond are represented by either an arrow with a cross on the positive end, or by a lower case Greek delta with a plus or minus charge.
- Polarity of bonds is important for predicting both the physical and chemical behavior of a compound.

Strength of Covalent Bonds

- Knowing how strong a bond is can help determine whether that bond will break in a reaction; weaker bonds are always more reactive than stronger bonds.
- There are several factors by which you can predict the strength of a bond.
 - Multiple bonds are stronger than single bonds because there are more electrons shared between them.
 - Larger atoms typically have longer and weaker bonds than smaller atoms because the electrons involved in the bond are farther away from the nucleus. This can happen with either a small atom bonded to a large atom, or two large atoms bonded together.
 - Third, when two electronegative atoms are bonded together, the bond is typically weak and reactive because both atoms are trying to pull the electrons toward them.
- When a bond is broken, it can happen in two ways homolytically or heterolytically.
 - Homolytic cleavage means that both atoms get one of the two electrons from the bond. It occurs in nonpolar bonds (when the atoms have about the same electronegativity).

- Heterolytic cleavage means that both of the electrons in the bond go to one of the two atoms. It occurs in polar bonds (when the atoms have different electronegativities) the most electronegative atom gets both of the electrons.
- Heterolytic cleavage is much more common than homolytic cleavage.

Geometry of molecules

- The atoms in a molecule exist in three-dimensional space. To describe the shape of a molecule, we focus on one atom at a time and describe the shape made by the atoms connected to it.
- These shapes can be predicted by the fact that the electron pairs around the atom repel each other, and atoms will arrange themselves so that all of the electrons are as far apart as possible. This is sometimes called VSEPR theory (valence shell electron pair repulsion what a mouthful).
 - Nonbonding electron pairs influence the shape of the molecule, but don't count as part of the shape.
- The most common shapes of atoms in organic molecules are:
 - tetrahedral four atoms, no nonbonding electron pairs (109.50 angles)
 - trigonal pyramid three atoms, one nonbonding electron pair (109.50 angles)
 - bent two atoms, two nonbonding electron pairs (109.50 angles)
 - trigonal planar three atoms, no nonbonding electron pairs (1200 angles)
 - linear two atoms, no nonbonding electron pairs (1800 angles)
 - bent two atoms, one nonbonding electron pair (1200 angles)
- The easiest way to envision these shapes is to use molecular models.
 - In the set most often used in this class, black is carbon or nitrogen, red is oxygen, white is hydrogen, and green is a halogen. (I wish they had provided blue nitrogen atoms, but no such luck.)
 - Single bonds are formed with a green tube, while double bonds are formed with two white tubes, and triple bonds with three white tubes.
 - When used correctly, these models give the correct shape of organic molecules. They can also be useful in seeing whether two Lewis structures are the same or different.

Molecules with a charge

- When a molecule contains an ionic bond, the atoms participating in the ionic bond must have a charge. In order for this to happen, one atom has to gain an electron while the other has to lose one.
- Carbon, nitrogen, oxygen, and halogen atoms can all lose or gain an electron to become positive or negative. When these atoms become charged, the number of bonds and lone pairs that they have will be different than for neutral compounds. You must learn to recognize when one of these atoms will have a charge.
- These charges are different from the partial charges that are present in a polar covalent bond. If a charged atom is also part of a polar bond, the charge will be increased or decreased by the polarity of the bond, which will have the effect of spreading the charge out.
- In general, neutral molecules are more stable than molecules containing charged atoms.
- Some charged atoms are more stable than others (and therefore less reactive).
 - More electronegative atoms are more stable with a negative charge than less electronegative atoms.
 - Less electronegative atoms are more stable with a positive charge than more electronegative atoms.
 - If the charged atom has an incomplete octet, it will be very reactive regardless of electronegativity.

Resonance Structures

- Resonance structures are used for molecules that cannot be accurately represented with a single Lewis structure.
 - The reason for this is that nonbonding electron pairs or single electrons can in some circumstances spread over more than one atom. Bonds can also be spread across more than two atoms.
 - $\circ~$ The kinds of molecules in which this occurs usually have multiple bonds next to charged atoms.
- One way to represent these kinds of molecules is called resonance structures.
 - These structures show all of the places where the bonds, nonbonding electron pairs, or single electrons are spread.
 - Each structure must follow the familiar conventions and rule of Lewis structures.
 - When we write resonance structures, we understand that the actual molecule is a mixture of them.

- Resonance structures can be used to accurately predict the behavior of these molecules.
- Resonance structures are not free to rotate relative to each other; the atoms in a certain position in one resonance structure correspond to those same atoms in another resonance structure.
- Resonance structures are always shown with brackets around them, and with a special doubleheaded arrow between them. This shows that all structures are actually part of the same molecule.
- Having resonance structures makes a compound more stable. This is usually because charges make a compound unstable, and if you can spread those charges out across different atoms, the molecule will be more stable. (Half a negative charge on two atoms is better than a whole negative charge on one atom.)
- Resonance structures of a molecule may have the same stability or different stabilities.
 - If the resonance structures are equally stable, then the actual molecule will be exactly halfway between them. These are said to be equal resonance contributors.
 - This occurs when bonds or nonbonding pairs are spread between atoms that are the same.
 - One or more resonance structures is more stable, then the actual molecule will be more like the stable structure. The more stable structures are said to be greater resonance contributors, while the less stable structures are said to be lesser resonance contributors.
 - This occurs when bonds or nonbonding pairs are spread between different atoms.
- Resonance structures of a single molecule are very different from a reaction in equilibrium.
 - In an equilibrium, a chemical reaction is occurring in both the forward and backward direction, so that a molecule is going back and forth from structure to another. It is one way, then the other, not both at the same time.
 - In a molecule with resonance structures, the molecule is not changing; it is always somewhere in the middle of the two structures shown.
 - The easiest way to determine whether two structures represent the resonance structures of a single compound, or two different compounds in equilibrium is to look for what has changed between them.
 - In resonance structures, only electrons have changed positions. Electrons, because of their wave-like nature, can spread out across two or more atoms.
 - In a chemical equilibrium, atoms will have changed positions. A chemical reaction will have occurred in which bonds are broken and formed again. Atoms are not waves, and cannot spread themselves out in space like electrons can. They must be in one place or another.

IV. Orbitals and Hybridization

Atomic orbitals

- Electrons in atoms are organized into energy levels; within those levels, the electrons are also organized into orbitals. Before an atom forms bonds, these orbitals are known as atomic orbitals.
- You can think about an orbital as a place with a certain shape where electrons reside. A more exact definition is that it is a mathematical equation describing an electron wave it defines the probability of finding an electron in a certain space.
- There are four kinds of atomic orbitals: s, p, d, and f.
 - There is one s orbital in each energy level, and they are spherical in shape.
 - There are three p orbitals in each energy level (starting with the 2nd), and they are shaped like a dumbbell, with two lobes on opposite sides of the nucleus. One p orbital is pointed along the x axis, one is pointed along the y axis, and one is pointed along the z axis.
 - There are five d orbitals in each energy level (starting with the 3rd), and seven f orbitals (starting with the 4th). Don't worry about the shape of d or f orbitals.
 - Each individual orbital can hold 2 electrons.

energy level	orbitals			
1st energy level	S			
2nd energy level	S	ррр		
3rd energy level	S	ррр	d d d d d	
4th energy level	S	ррр	d d d d d	ffffff
5th energy level	S	ррр	d	ffffff
6th energy level	S	ррр	d	ffffff

- The electron configuration of an atom gives the number of electrons in each orbital, starting with the lowest energy level. For example, the electron configuration of carbon is $1s^2 2s^2 2p^2$.
 - The energy levels correspond to the periods on the Periodic Table. You can read across the table to see what the electron configuration should be.
- Organic molecules are made mostly of C, H, N, and O. Therefore, we will nearly always be concerned with s and p orbitals.
 - H only has electrons in the s orbital of the first energy level.
 - The valence electrons of C, N, and O are in the s and p orbitals of the second energy level.

Molecular orbitals

• When two atoms join together to make a covalent bond, two atomic orbitals, one from each

atom, are joined together to make a molecular orbital.

- When the two electrons occupy the new molecular orbital, they are spread out between the two atoms so that the atoms are held together, making a covalent bond.
- There are two kinds of molecular orbitals sigma (s) orbitals and pi (p) orbitals.
 - Sigma orbitals may be formed by joining s orbitals together, and pi orbitals are made by p orbitals together.
 - Single bonds and the 1st bond of a double or triple bond are sigma orbitals. They are shaped like a fat sausage, with the nucleus of one atom inside each end.
 - The 2nd bond of a double bond, and the 2nd and 3rd bond of a triple bond are pi orbitals. They are shaped like both sides of a hot dog bun, half above and half below the sigma orbital.
- The reason for these two different orbitals with different shapes is that orbitals can't occupy the same space.
 - In order to have a double bond between two atoms, the sigma orbital goes in the middle (the sausage), and the pi orbital goes above and below it, (the hot dog bun).
 - In order to have a triple bond between two atoms, the sigma orbital goes in the middle (the sausage again), and one pi bond goes above and below, and the other goes to either side (imagine a fat sausage with two buns around it).
- Whenever a sigma or pi bonding orbital is formed, an antibonding orbital is also created.
 - Bonding orbitals occupy the space between the atoms (the sausage and buns), while antibonding orbitals occupy the space away from the two atoms. If the antibonding orbitals had electrons in them, they would pull the atoms apart rather than join them together.
 - Bonding orbitals are lower in energy than the two atomic orbitals that they came from, while antibonding orbitals are higher in energy than the two atomic orbitals they came from.
 - Electrons in the covalent bond will therefore occupy the bonding orbitals, rather than the antibonding orbitals.
 - Antibonding orbitals are designated with a star: s* or p*, and are often referred to as "sigma-star" or "pi-star" orbitals.

Hybridized atomic orbitals

- In order to create the kind of orbitals that a molecule has, we will need more s-type orbitals.
 - An atom with four single bonds needs four s-type orbitals to make 4 sigma orbitals.

- An atom with two single bonds and a double bond needs three s-type orbitals and one p orbital to make 3 sigma bonds and one pi bond.
- An atom with one single bond and a triple bond needs two s-type orbitals and two p orbitals to make 2 sigma bonds and two pi bonds.
- In order to get more s-type orbitals, the s orbital is mixed with one, two, or three p orbitals (depending on how many you need for making pi orbitals). The blending or mixing process is called hybridization, and the blended orbitals are called hybrid orbitals.
 - \circ An atom with four single bonds, all three p orbitals are mixed with the s orbital. The result is 4 identical orbitals that are called sp³ hybridized orbitals.

 $s + p + p + p = sp^3 + sp^3 + sp^3 + sp^3$

 \circ For an atom with two single bonds and a double bond, two p orbitals are mixed with the s orbital. The result is 3 identical sp² hybridized orbitals plus the leftover p orbital.

 $s + p + p + p = sp^2 + sp^2 + sp^2 + p$

• For an atom with one single bond and one double bond, one p orbital is mixed with the s orbital. The result is 2 identical sp hybridized orbitals plus two leftover p orbitals.

$$s + p + p + p = sp + sp + p + p$$

- Having an s orbital in the mix makes the hybrid orbital behave like an s orbital (no matter how many p's were added). Therefore, the hybridized orbitals can be joined together just like s orbitals to make sigma bonds.
- Hybridized orbitals spread out around the atom so that they are as far apart as they can get. (The p orbitals have to join to make a pi bond in the same direction as one of the sigma orbitals, so they can't spread out.)
 - Four sp³ orbitals spread out to make a tetrahedral shape, trigonal pyramid shape, or bent shape (depending on how many orbitals make bonds vs how many are nonbonding electron pairs.
 - \circ Three sp² orbitals spread out to make a trigonal planar shape, bent shape, or no shape.
 - Two sp orbitals spread out to make a linear shape, or no shape.
 - Nonbonding electron pairs usually occupy hybrid orbitals, as this also allows them to spread out better. (When they participate in resonance structures, this can change.)
 - Empty orbitals are always unhybridized p orbitals.