STUDY GUIDE - *Ch.2.3 - The formation & function of molecule depend on chemical bonding between atoms.* **NAME:** ______________________

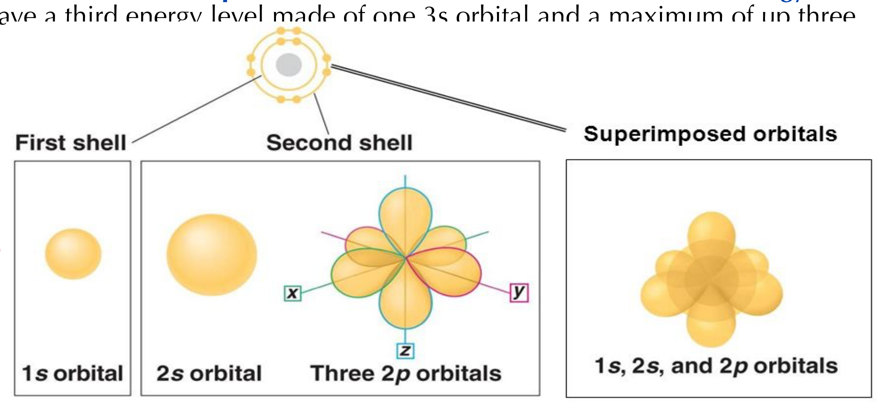
- **PHYSICALLY PRINT OUT this PDF and HANDWRITE (with a black or blue pen) your answers directly on this PDF.** *Typed or digitally-written work is **not** be accepted. Do **not** answer questions on separate paper.*
- **Importantly, study guides are NOT GROUP PROJECTS!!!** *You, and you alone, are to answer the questions as you **read** your assigned textbook. You are **not** to share answers with other students. You are **not** to copy any answers from any other source, including the internet.*
- **Get in the habit of writing LEGIBLY, neatly, and in a medium-sized font.** *AP essay readers and I will skip grading anything that cannot be easily read so start perfect your handwriting, and don't write so large you can't add all the relevant details and key elaborations in the space provided.*
- **SCAN physical documents in color and with good resolution. Then, upload your final work as PDFs to Archie.** *Avoid uploading dark, shaded, washed out, side ways, or upside down scans of homework. Keep completed physical study guides organized in your biology binder to use as future study and review tools.*
- **READ FOR UNDERSTANDING and not merely to complete an assignment.** ***First**, read a section quickly to get an overview of the topic covered. Then, read it a **second** time slowly, paraphrasing each paragraph **out loud** and analyzing every figure. Finally, read it a **third** time as you answer the study guide questions if assigned and to start building your memory. Try to write answers out in your own words when possible and to purposefully and accurately use all new terminology introduced.*

Remember, **that in a neutral atom, the TOTAL number of electrons in the electron cloud around the nucleus equals the number of protons in the nucleus**. Though negatively charged electrons are attracted to the positive charges of the protons, electrons also repel each other. There are, therefore, some areas around the nucleus where there exists a higher probability of finding an electron at any given instance. These regions are known as **orbitals**, a maximum of *two electrons* occupying any given orbital. Electrons added to an atom composed of just the nucleus fill the orbitals closest to the nucleus first. Electrons will first come to occupy a spherical area around the nucleus, this area being known as the ***1s*** orbital. A neutral atom of hydrogen, which has one proton and, thus, one electron would have only one electron in its electron cloud, this electron being located in the first energy level, which is composed of only one orbital, the 1s orbital.

Unlike that of hydrogen, most atoms contain more than one electron. After two electrons occupy this 1s region, additional electrons will occupy a spherical area further out from the nucleus, known as the ***2s*** orbital. This is the case for lithium, which has three protons and, thus, can attract three electrons, the third electron being pushed into the second s orbital (2s). After two electrons have been added to the 2s orbital, additional electrons added to an atom's electron cloud will occupy regions that look like dumbbell shapes, referred to as p orbitals. In the second energy level of an atom's electron cloud, you will find a maximum of *three p orbitals* (one referred to as $2p_x$, one $2p_y$, and one $2p_z$), each holding up to two electrons. (Note that there are no p orbitals in the first energy level discussed in the previous paragraph). If we look at atoms of the element fluorine, for example, the atom has nine protons and, thus, nine electrons, the first two being found in the region known as the 1s orbital, the next two filling the 2s orbital, and the last five occupying the areas referred to as the 2p orbitals. Because only two electrons can occupy an orbital, two out of the three 2p orbitals are filled with two electrons each, the last electron occupying the third 2p orbital, which could hold one more, but doesn't in a neutral fluorine atom. There are other atoms though that have so many electrons, all their 2s and 2p orbitals are filled, and the outermost electrons end up occupying either a spherical space with an even greater diameter knows as the 3s orbital or the 3s orbital AND up to three dumbbell-shaped regions with a larger diameter known as the 3p orbitals etc.

If we look at the first energy level, made of one 1s orbital, the maximum number of electrons that can occupy this region is two **(the maximum possible number of electrons that can be held in the first energy level of an atom is two)**. If we look at the entire second energy level, made of one 2s orbital and up to three 2p orbitals, the maximum number of electrons that are able to occupy these regions or **the maximum possible number of electrons in the second energy level of atoms is eight.** This is true for atoms that have a third energy level made of one 3s orbital and a maximum of up three 3p orbitals as well, for example.

For an atom to be stable, its outer most electron shell must be filled, no matter if it only has a 1s orbital or if it could fill both one s and three three p orbitals in that outermost energy level or energy shell.



Hybrid orbitals will form later when two atoms get close and the nuclei of each start pulling on the electrons in the electron valence shell of the other.

1. Fill in the blank: Atoms are stable when they have their outermost electron level, a.k.a. their **valence shell**, filled. If they don't, **they can transfer or share valance electrons**, which results in atoms being held closely together in attractions we call ______________________________________.

2. **The two strongest types of attractions between two atoms are covalent bond and ionic bonds** *(as long as ionic bonds are **not** occurring between ions in water).* What is a **covalent bond**?

3. What is a **molecule**?

4. Distinguish between a **single bond** and **a double bond.**

   Single bond =

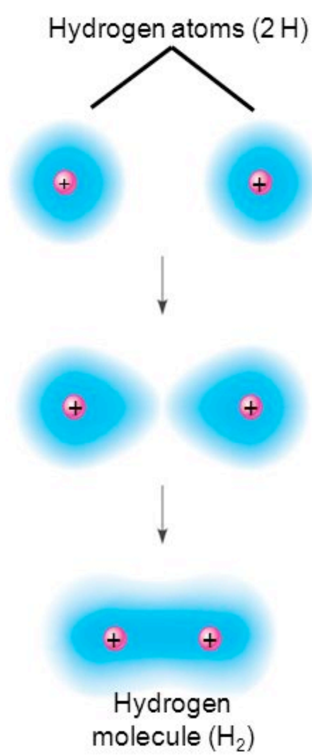
   Double bond =

5. Using Figure 2.9 as your guide, explain the three steps in the **formation of a single covalent bond between two hydrogen atoms**.

1.

2.

3.



6. a. What is meant by the **valence** of an atom?

   b. Refer back to Figure 2.7 in Section 2.2 of your textbook. What is the **valence of helium**?

c. Refer back to Figure 2.7 in Section 2.2 of your textbook. Why is the **valence of carbon 4 but oxygen 2** when carbon has 4 valence electrons yet oxygen has 6?

7. What is the difference between **a molecule that is a pure element versus one that is considered a compound**?

8. Complete the table below.

	Molecule? (y/n)	Compound? (y/n)	Molecular Formula	Structural Formula
Water				
Carbon dioxide				
Methane				
O_2			O_2	

9. a. What does the term **electronegativity** refer to?

Electronegativity

F > O > N > C ≃ H

4.0 3.5 3.0 2.5 2.1

b. What happens **as an atom's electronegativity increases**?

11. What is a **nonpolar covalent bond**?

12. What is a **polar covalent bond**?

13. a. How do the **electronegativities of an oxygen and hydrogen** atom compare?

b. Because of their electronegativity difference, **what happens when an O and an H atom covalently bond**?

c. How do the **electronegativities of a carbon and hydrogen** atom compare?

14. a. What are **ions**?

b. In terms of electronegativity, **when** do two oppositely charged ions form from two neutral atoms?

c. Distinguish between a **cation and an anion.**

Cation =

Anion =

15. What is an **ionic bond?**

16. a. If a (neutral) sodium atom has 11 protons, how many total electrons does it have in its **electron cloud**?

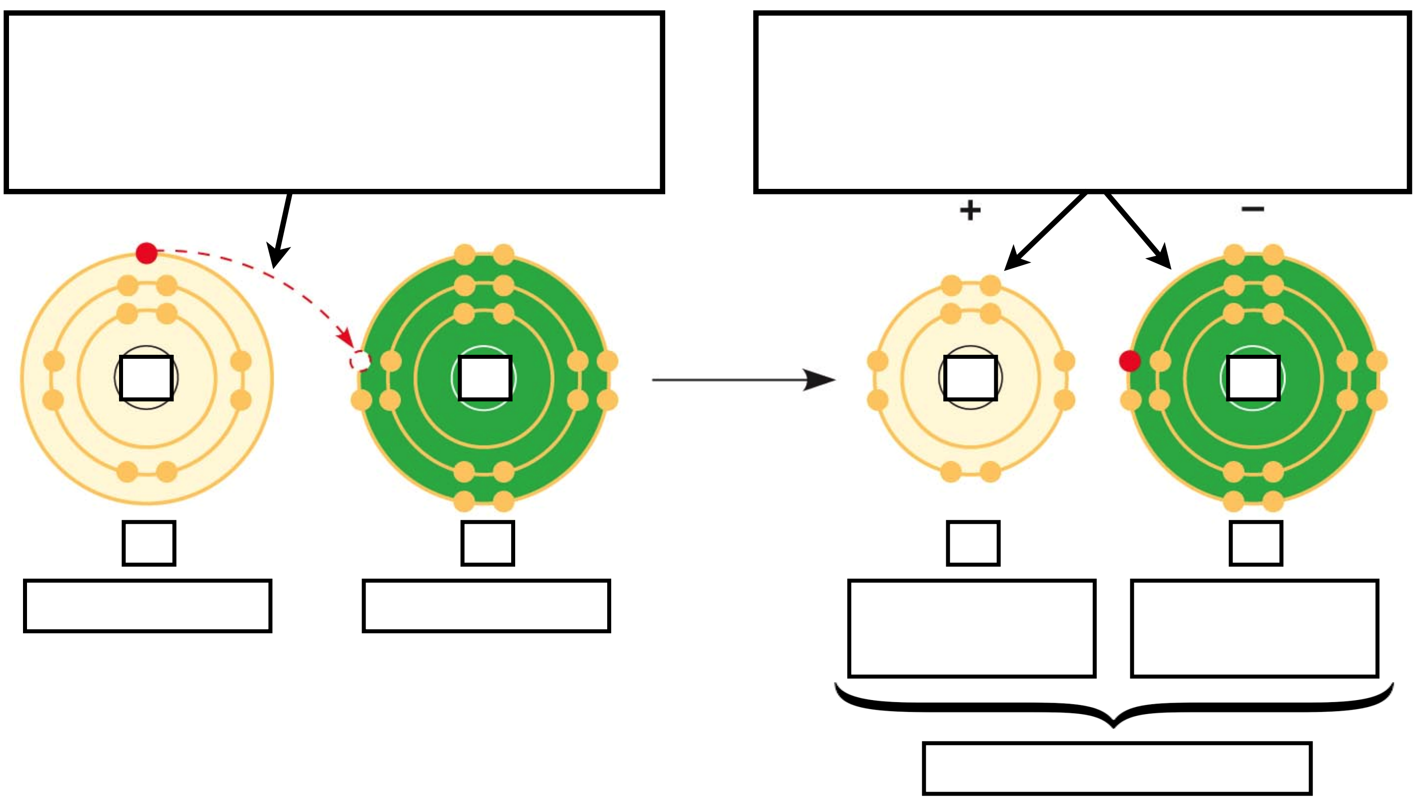
b. How many **valence electrons** does a sodium atom have? *(Valence electrons being those electrons in the valence shell or the outermost electron level (in the outermost s and p orbitals))*

c. If a (neutral) chlorine atom has 17 protons, how many total electrons does it have in its **electron cloud**?

d. How many **valence electrons** does a chlorine atom have?

e. Based on you reading of Ch.2.3, are sodium and chlorine stable atoms? Why or why not?

f. Explain the **steps in the formation of the sodium and chloride ion**, as well as the **formation of the ionic bond** that forms between an Na and a Cl atom by filling in the boxes in the figure below.



g. Why does the chloride <u>anion</u> have a complete valence shell?

h. Why does the sodium <u>cation</u> have a complete valence shell?

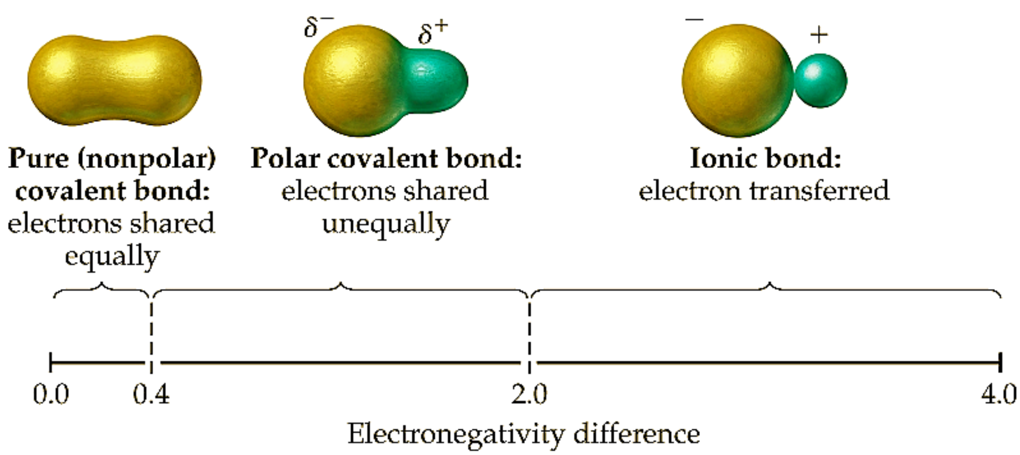
i. The transfer of the electron from the sodium to chlorine atom is **<u>NOT</u>** an ionic bond (a strong attraction between oppositely charged ions). By drawing on the figure above, **<u>where would you say the ionic bond is "located"?</u>**

17. a. What is a **<u>crystal of salt</u>**? *(Note that not all salts are the familiar table salt, NaCl)*

b. **The molecular formula for a molecule indicates the exact number of atoms of each element that make up that one molecule.** The formula for glucose for instance **$C_6H_{12}O_6$**. **The formula for an ionic salt, however, does <u>not</u> indicate the exact number of atoms of each element that make up that one compound.** The formula for calcium chloride for instance is $CaCl_2$. Explain what the **<u>ionic formula</u>** thus indicates.

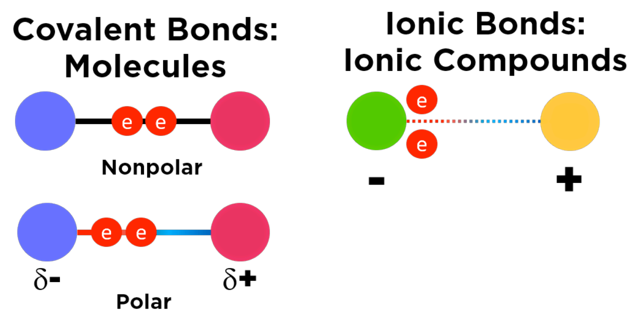
c. **When a molecule is added to an aqueous solent (water), the molecule's covalent bonds do <u>NOT</u> break.** However, why do we say that the **<u>environment surrounding a salt crystal affects the strength of the ionic bonds in the salt</u>**?

To review = **The <u>electronegativity</u> of an atom is a measure of its affinity for electrons. The relative electronegativity of two interacting atoms also plays a major part in determining what kind of strong chemical bond forms between them.** The <u>absolute value of the difference in electronegativity</u> (**ΔEN**) of two bonded atoms provides a rough measure of the polarity to be expected in the bond and, thus, the bond type. When the difference is very small or zero, the bond is non-polar covalent. When it grows somewhat larger, the bond is polar covalent. When it is large, it is ionic.

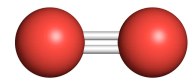


The absolute values of the electronegativity differences between the atoms in the bonds H–H, H–Cl, and Na–Cl are **0** (non-polar), **0.9** (polar covalent), and **2.1** (ionic), respectively.

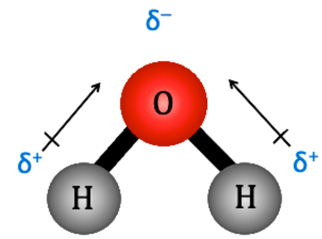
The degree to which electrons are shared between atoms varies from completely equal (non-polar covalent bonding) to not at all (ionic bonding).



Polar Bonds vs Polar Molecules

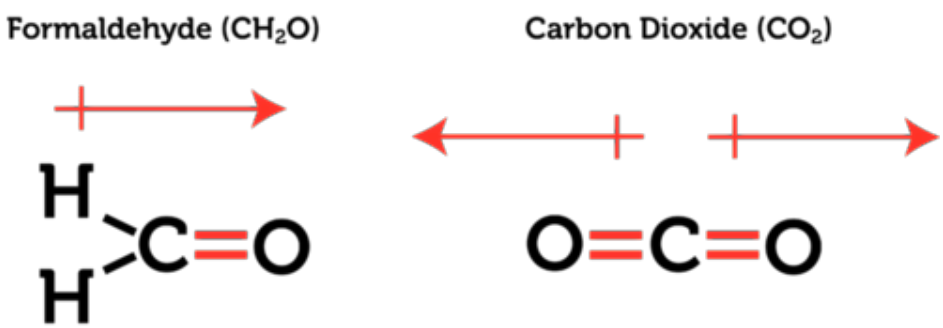


A covalent bond between two atoms is non-polar if the atoms have the same electronegativity or a difference in electronegativities that is **less than 0.4**. An example of a non-polar bond is the bond in O_2. Oxygen gas contains two oxygen atoms. **The electrons are shared equally because the electronegativity difference between the two atoms is zero. This means no net dipole moment forms: one oxygen is not, on average, more partially negative while the other is partially positive.**



A covalent bond between two or more atoms is polar if the atoms have significantly different electronegativities (**>0.4**). **Polar bonds do not share electrons equally, meaning the negative charge from the electrons is not evenly distributed in the molecule. This causes a dipole moment to form.** A dipole moment occurs when **one end of the bond is partially positive, and the other end is partially negative**. A classic example of a polar bond is the bond in water between hydrogen and oxygen. The bond is classified as a polar bond because it has a large electronegativity difference of 1.4 between the O (EN=3.5) and the H (EN=2.1).

All polar molecules contain polar bonds, but having polar bonds does not necessarily result in a polar molecule! It depends on how the atoms are arranged and so how the forces (dipole moments) are oriented. In both molecules below, the oxygen atoms attract electrons more strongly than the carbon or hydrogen atoms do, so both molecules have polar bonds. However, **only formaldehyde is a polar molecule**. **Carbon dioxide is non-polar molecule!**



In chemistry, polar molecules contain regions that are slightly or fully positive in one area and slightly or fully negative in another. They are generally asymmetrical, with an uneven distribution of the electrons. A polar molecule with two oppositely-charged poles is called a dipole.

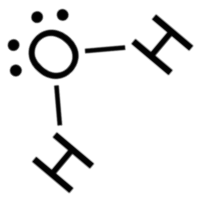
If the atoms in a molecule either share bonding electrons evenly, form non-polar bonds, or the polar bonds are symmetric, meaning that the dipole moments of the various polar covalent bonds cancel each other out, a molecule is non-polar (not a dipole). This is the case with CO_2. In this molecules, there are two polar bonds and thus two dipole moments, but they cancel each other out. **Due to the symmetrical arrangement of the atoms in the molecule, carbon dioxide being linear in shape, and the fact that the forces/dipole moments of each polar bond, therefore, try to pull electron density in equal and opposite directions, the forces cancel each other out, leaving electron density distributed more evenly along the whole molecule and one end of CO_2 NOT being more positive and the other end of CO_2 NOT being more negative.**

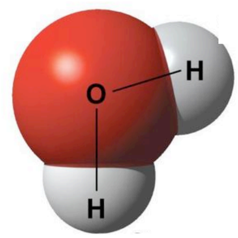
However, **if the atoms in a molecule form polar bonds which are arranged asymmetrically in the molecule, meaning that the dipole moments of the various polar covalent bonds do not cancel each other out, a molecule will be polar (a permanent dipole).** This is the case with H_2O and with formaldehyde. Polar molecules tend to attract one another more readily, which affects the properties of polar substances like water. Water molecules can actually align themselves in the presence of an electrostatic force (an electric field with a negative and a positive end).

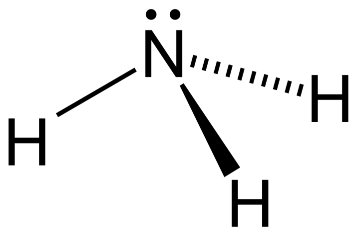
As you will learn, **polar solvents tend to dissolve polar or charged solutes** as there are temporary, but ongoing attractions occurring between the two types of particles as thermal energy causes kinetic motion of all particles. **Non-polar solvents tend to dissolve non-polar solutes,** which also move around randomly and so get distributed due to thermal energy. **Non-polar solutes do not dissolve easily in polar solvents, however, since there is no attraction between a non-polar substance and the polar molecules of the solvent while the polar solvent particles do constantly attract on and off each other, ignoring the non-polar particles.**

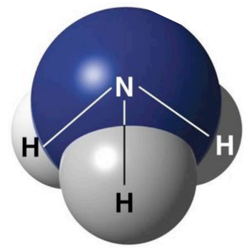
18. a. Both **water and ammonia are polar molecules** made up of partially charged atoms. Partial charges are represented by $\delta+$ and $\delta-$. Note that the **oxygen in water has <u>TWO</u> partial negative regions (where the two pairs of lone electrons are found)** that can attract positive charged regions of other particles outside the water molecule while the **nitrogen in ammonia has <u>ONE</u> partial negative region (where the one lone pair of electrons are found)** that can attract a positive charged regions of other particles outside the ammonia molecule. Of course **both water and ammonia also contain multiple partially positive hydrogen atoms** that can attract negative charged regions of other particles outside the molecules in question.

Please **<u>label all the partial charges on the molecules</u>** of water and ammonia in both images below.







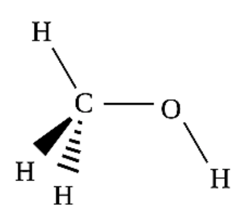


b. Unlike strong atomic attractions such as the ionic bonds in dry crystals and covalent bonds in molecules, **hydrogen bonds are considered weak intermolecular forces of attractions. <u>Draw and label the hydrogen bond</u>** in the figures above. **Note that since hydrogen bonds are weak attractive forces between distant parts of the same large molecule or between regions of different molecules** (as opposed to strong covalent bonds between atoms within a molecule), **we draw them as dotted lines and <u>not</u> solid ones.**

c. **<u>Hydrogen bonds</u> are weak attractions between partially positive hydrogen atom in a molecule and an electronegative atom that becomes partially or fully negative in a <u>neighboring</u> molecule or even somewhere further away in the same molecule. THEY'RE <u>NOT</u> THE COVALENT BONDS INVOLVING A HYDROGEN ATOM!**

<u>How (or when) does a hydrogen atom in a molecule becomes positively charged</u> *(so that a hydrogen bond can even form thereafter involving this charged hydrogen)*?

d. In the left-hand box, identify **<u>which atoms in water and methanol would be involved in hydrogen bonding</u>.** In the right-hand box, draw the structural formulas of two to three water molecules, which are hydrogen-bonding with the methanol included. **<u>Label all the partially-charged atoms in both the waters and the methanol</u>**, draw the **<u>correct orientation of the waters</u>**, and draw the hydrogen bonds (dotted lines).

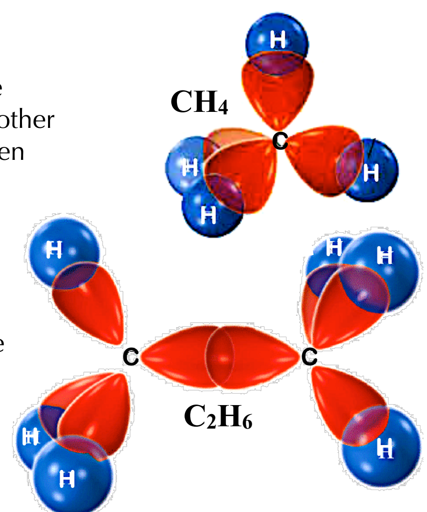


19. a. Let's read now about **<u>Van der Waals interactions</u>**. Remember, even in a molecule with nonpolar covalent bonds, in which electrons are shared equally between two atoms in question, electrons are **<u>NOT</u>** always _________________ ___________________________ within a covalent bond ; at any instant, electrons may accumulate _______ ________________ in one part of the molecule or another for a brief amount of time.

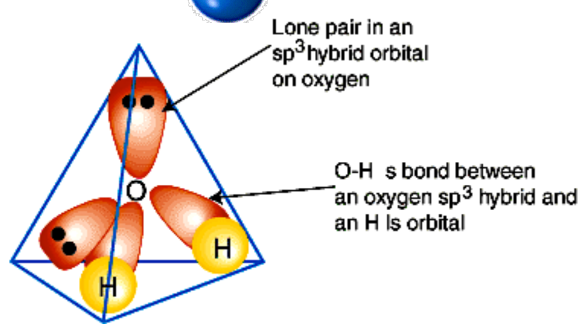
b. What is the **<u>results of these short-lived, temporary charges on molecules</u>**?

c. **<u>When do weak, momentary Van Der Waals Interactions tend to occur</u>**?

20. Why are weak **<u>Van Der Waals Interactions</u>** *(which includes London Dispersion Forces, Dipole-Induced Dipole Forces, Dipole-Dipole Forces like hydrogen bonding, weaker ionic bonds between fully charged ions in water, and even intermolecular repulsions between opposite partial charges)* **<u>so important in biology</u>**?



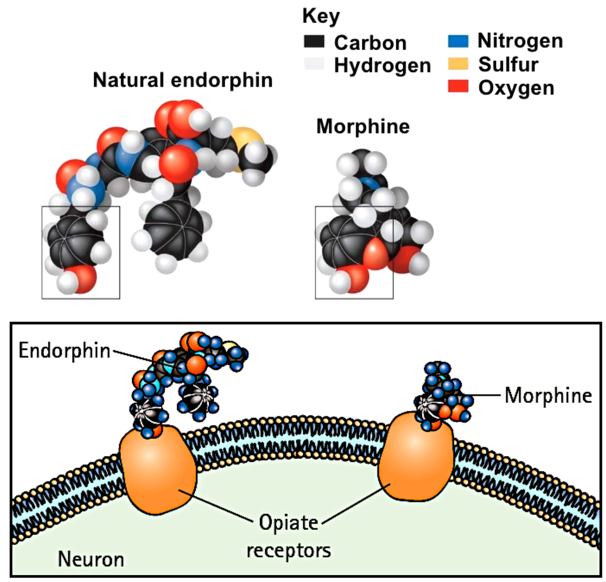
Atoms with valence electron in both s and p orbitals (like carbon with its four valence electrons, initially in its outer s and certain p orbitals) will form covalent bonds with other atoms that also have incomplete valence shells (like other carbons or hydrogens). When bonding occurs, the single s and up to three p orbitals of the valence shell of an atom combine to form up to four teardrop-shaped hybrid orbitals (like when a carbon forms covalent bonds with four hydrogens and/or other carbons). These hybrid orbitals of the central carbon's outermost valence shell are now filled with bonding electrons from other carbon and hydrogen its bonding with, the main carbon being more stable now on account of it now having filled or often-filled valence shell. These teardrop-shaped hybrid orbitals extend from the nucleus of the carbon to four corners of an imaginary tetrahedron where the hydrogen or other carbons its bonding with are located. **This tetrahedral shape of a carbon atom bonded to four other atoms (bond angles 109.5° apart) is often a repeating motif within biological, carbon-based molecules.**



In water, not all four hybrid orbitals form between atoms since two of the four hybrid orbitals are filled with the lone pair, nonbonding electrons of the O atom. The other two teardrop-shaped hybrid orbitals though do fill with bonding electrons from the oxygen and each of the two hydrogen atoms that covalently bond with the oxygen. **The result is a water molecule shaped roughly like a "V", its two covalent bonds located an angle of 104.5° degrees.**

21. Why is **<u>molecular shape</u>** crucial?

22. **Molecular shape is critical** and exemplifies the theme of **form fitting function** *(molecules have the 3-D shape they need to perform the function they need to do)*. In what way is morphine a **molecular mimic** of natural endorphins?



23. Explain why the structure **H - C = C - H** fails to make sense chemically? *(Check your answers to question 23 by going to the Ch.2.3 **Concept Check Question #1** in Appendix A of your textbook)*

24. What holds the atoms together in a crystal of magnesium chloride ($MgCl_2$): ionic bonds, covalent bonds, or Van Der Waals interactions? Explain. *(Check your answers to question 24 by going to the Ch.2.3 **Concept Check Question #2** in Appendix A of your textbook)*

25. If you were a pharmaceutical researcher, why would you want to learn the three-dimensional shapes of naturally occurring signaling molecules in the cell or body? *(Check your answers to question 25 by going to the Ch.2.3 **Concept Check Question #3** in Appendix A of your textbook)*