The Chemistry of Life Lecture Guide

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*Number in outline corresponds to the slide number in the PowerPoint presentation.

1. The Chemistry of Life. In this Lecture, we will look at the basic building blocks of a cell, starting at the smallest level, the subatomic particles. As we move through the Lecture, we will successively step up in complexity to look at how subatomic particles interact to make atoms, and then, how atoms interact to make up molecules by creating bonds.

2. The Periodic Table of Elements (PToE). You can learn a lot about the chemistry of molecules by examining how the PToE is arranged. First of all, elements are arranged based upon size. As you move across a row, the elements get progressively larger, and elements in the same column have similar chemical properties. This is a concept we will revisit a bit later when discussing how elements interact with one another, AKA create bonds. Every PToE is going to look a little different, depending on how detailed the author wants to be. Notice the color coding in this table. The red elements represent the 6 most abundant elements found in living things. The purple elements represent the 5 most common ions (elements that have a positive or negative charge) found in living systems. The dark and light blue elements represent what are known as trace elements. These are elements that are necessary for supporting life, but are only found in small amounts in an organism compared to the red elements.

3. Before we move on to the details of atomic structure, let's get some important vocabulary out of the way:

    a. There are three types of subatomic particles: **Protons** are positively charged and have a mass of 1. **Neutrons** do not have a charge and also have a mass of 1. **Electrons** are negatively charged and, for our intents and purposes, do not have a mass. Technically, they do have a mass, but is very, very small. The mass of an electron, however, is beyond where we will go in our chemistry lesson, therefore, we will just say electrons have no mass.

    b. An **element** is a substance made up of one or more **atoms** that are all chemically identical and are defined by the number of protons it has. For example, an atom of carbon, one of the most important elements to life, has 6 protons. If the proton number is different than 6, then it is not a carbon atom.

    c. A **molecule** is the joining of two or more atoms by allowing their electrons to interact with each other. A **compound** is a type of molecule that includes two or more different elements. So, for example, oxygen gas, which is two oxygen atoms bonded together, is a molecule, but not a compound. Carbon dioxide, which is 1 atom of carbon bonded to 2 oxygen atoms, is both a compound <u>and</u> a molecule.

    d. **Chemistry**, at its most basic level, is the study of how atoms interact with each other to form molecules and compounds. Technically speaking, it is the study of how electrons in different atoms interact with each other to form bonds. It is these bonds that shape the molecules and their chemical properties.

4. The structure of an atom is an important part of understanding chemistry. In the **nucleus** of an atom, you will find the neutrons and protons. Neutrons act as peacekeepers so protons can hang out together. Particularly for the larger elements that have a high number of protons, the forces that a

large number of protons exert in an attempt to repel each other would be too great without the neutrons in the mix making sure everyone plays nice. The electrons hang out in the electron shells, or outer orbitals of an atom. It is important to make the distinction here that an atom is a 3 dimensional structure, so you should think of these orbitals as having a ball shape coming out of your screen.

a. As I said before, an element is defined by the number of protons it contains. This number is also known as the **atomic number**. This number never changes. If an atom has an atomic number of 2, that means that it has 2 protons, and it is a helium atom. Regardless of other subatomic particles found in the atom, if it has more or less protons it is not helium.

b. Neutrons on the other hand, can vary. Recall that both N and P have a mass of one, so the **atomic mass** of an atom is the number of protons plus the number of neutrons. Technically speaking, the atomic mass is the total mass of all subatomic particles, but since we are under the assumption that electrons do not have a mass, they do not contribute to the overall weight of an atom.

c. The number of electrons can also vary in any given atom. Keeping in mind that electrons (E) have a negative charge and protons (P) have a positive charge, when the number of electrons is equal to the number of protons the atom has no charge. When the number of E and P are not equal, then the atom is charged and is called an ion. When there are more electrons than protons, the atom is negatively charged and called an **anion**. When there are more protons than electrons, the atom is positively charged and is called a **cation**.

5. **Isotopes** are all the different forms of an element that can exist. So, they are atoms of the same element, but with a different atomic mass, i.e. same number of protons, but differing number of neutrons. Take carbon, for example. The most common isotope of carbon is C-12, 6 protons, 6 neutrons. This is the form of carbon that you will find in living organisms and should come to mind when we talk about carbon molecules in living things. There are other isotopes, such as C-13, which has 6 protons and 7 neutrons, but, regardless of which isotope we are dealing with, they all behave the same when interacting with other atoms and molecules. This is because, as you will learn shortly, it is the electrons in the atom that determine how atoms interact with each other.

   a. Carbon-14 is considered a radioactive molecule. Unlike C-12 and C-13, C-14 is unstable, meaning that the isotope holds more energy than it can safely hold. The subatomic particles within C-14 want to decay, or split, into other particles. This process releases that excess energy, and that release is referred to as radioactive decay. When Carbon-14 decays, it turns into Nitrogen-14.

   b. With the concept of isotopes under our belt, let's revisit the periodic table of elements. Take a moment to go back to slide 2 and make note of how the atomic mass is reported in that table. Go ahead, I'll wait… Did you notice that most elements have an atomic mass reported as a decimal point value? For example, carbon's atomic mass is displayed as 12.01. That .01 comes from the different isotopes that carbon can occur in. It is the average relative occurrence of all isotopes that can be found in nature. Since 99% of all carbon in existence is C-12, the average atomic mass is very close to, but not exactly, 12. Chlorine, on the other hand, has an atomic mass of 35. 45, which suggests that Cl-35 is almost as common on the planet as Cl-36.

6. When looking at the periodic table of elements, these pieces of information are often displayed, and from this simple lay out you can always determine the number of subatomic particles an element has. The position of each may vary, but just remember, the atomic mass will always be the larger of the two numbers shown. Most elements will not have a charge associated with it, only the elements in

the first two columns and the second and third from the right side. We will explore why this is so in a few slides. In elements that do not have a charge displayed, it is assumed that it is a neutrally charged atom, meaning that there are just as many electrons as there are protons.

7. Practice! Identify the number of N, P, and E for each of these three elements and the atomic number and atomic mass of the hydrogen isotopes. Record your answers in the associated study guide.

8. **Electron shells**, also called **orbitals**, are where you will find the electrons. It is these particles that determine the chemical properties of an atom. For our intents and purposes, we will only consider the first 4 electron shells. Most biologically relevant elements will not be much larger than 4 shells, and the rules change a bit once we have atoms with more than 4 shells. It is important to understand that, although each atom has its own unique number of electrons, the rules that govern how electrons situate themselves within an atom are universal. The first shell, which is the smallest, can only hold 2 electrons. This means that hydrogen and helium are the only elements that will only contain 1 electron shell. All other elements will have at least 2 shells. Shells 2 and 3 hold up to eight electrons. It is also important to understand that all vacancies in the lower shells must be occupied before adding a new shell. You will not start putting electrons in larger shells if the smaller ones are not filled yet. Regardless of how large the atom is, or how many orbitals it has, the outermost shell in any given atom is called the **valence shell**. An atom is the most stable when its valence shell is full, and this is the basis of forming chemical bonds.

9. **Chemical bonds** are formed when electrons in the valence shells of atoms interact to fill vacancies in the outermost shell. An atom always wants to fill its valence shell. A full valence shell means a happy atom, and it will always seek out other atoms to get happy. (Clap along if you know what happiness is to you.) For some atoms (not all, but some), **electronegativity** is the driving force in getting happy. Electronegativity is a measurement of how hard an atom can pull electrons to itself. Those atoms that only have 1 or 2 vacancies have a high electronegativity. They can pull electrons into their valence shell, stealing electrons from other atoms. You can think of these atoms with high electronegativity as the bullies of the chemical playground, stealing lunch money from the weaker, low electronegative atoms. These weaker atoms that have only 1 or 2 electrons in its outermost shell can't pull very hard and, therefore, are more likely to lose their valence electrons to become stable.

10. Remember how columns in the PToE have similar chemical properties? The columns labeled 1 and 2 contain elements that have a weak electronegativity (column 1 more so than column 2). They are more likely to donate their electrons and lose their valence shell altogether to become stable. Conversely, columns 14 and 15 (column 15 more so than column 14), have high electronegativity and can easily steal electrons, most often from elements in the first 2 columns, to stabilize. The last column (16), are known as the noble gases and are inert. These elements have a full valence shell. They don't form bonds because they are already naturally happy atoms. No vacancies, no need to acquire or get rid of extra electrons. In your study guide or notes, prove to yourself that these rules of electronegativity are true by drawing the valence models for all the elements in the third <u>row</u> of the table.

11. There are 3 types of chemical bonds that we will explore: covalent, ionic, and hydrogen. We will dissect each of these in detail in the coming slides, but it is best to have a general idea of each before continuing.

    a. **Covalent bonds** are the strongest of all the bonds, meaning that, once they are formed, it takes a lot of energy to break that bond. Typically speaking, this type of bond occurs between atoms with similar electronegativity. Neither participant can pull on electrons hard enough to steal them away from the other, so instead they share electrons. Nonpolar covalent bonds are created when each atom has the same level of electronegativity, whereas polar covalent bonds are formed when one atom pulls just a little harder than the other. The slightly stronger atom, in a sense, hogs the electrons, and they are shared unevenly. We will see the relevance of that in a few slides.

   b. **Ionic bonds** are those that are formed between atoms that have a large difference in their electronegativity. Instead of sharing electrons, the highly electronegative atom steals electrons from the weaker atom.

   c. **Hydrogen bonds** are the weakest of the three and do not form new molecules. Instead, these bonds are formed between polar covalent molecules and allow molecules to “stick” together. More on that to come...

12. Let’s start with **ionic bonds**. Recall that for elements that are weakly electronegative it is easier to give their valence electron away instead of trying to find 7 more electrons to fill their shell. In these atoms, such as a neutrally charged sodium atom, when they donate their valence electron to another molecule, the outermost orbital disappears, making the lower orbital the valence shell with a full complement of electrons. This atom is now stable, but it also has one less electron than proton now and becomes a positively charged ion, AKA a cation. For atoms with strong electronegativity, they steal electrons from others to fill the single vacancy in their valence shell. Now, they have a negative charge because of the acquired extra electron, AKA an anion. In the example here, when sodium donates its extra electron to the chlorine, both ions are now “happy”, but with opposite charges. The opposing charges cause the ions to become attracted to each other, thus forming an ionic bond.

13. **Covalent bonds** occur between atoms that are not quite so eager to give away their electrons or can’t pull hard enough on them to steal it away completely. So, these electrons are shared. Keep in mind that a covalent bond involves the sharing of 2 electrons, one from each participating atom. There are three things that you need to remember when evaluating and predicting covalent bonds: #1, a covalent bond is defined by a <u>pair of electrons</u>, 1 from each participant. You can’t have only one atom sharing an electron and the other just reaping the benefits. Share and share alike, as we learned in kindergarten. #2, both participants <u>must have 8 electrons each</u>, including the shared ones, in their valence (outer) electron shell to be “happy”. If one has 8 and the other has 7 after sharing electrons, one isn’t happy, and, therefore, the bond cannot occur. And #3, when considering how atoms will fit together, and how many electrons must be shared, <u>consider how many vacancies</u> each atom has <u>and how many different atoms occur in the molecule</u> you are trying to form. In these examples, the number of vacancies equal the number of bonds that need to be formed. But what about a molecule like carbon dioxide, CO_2 (1 carbon atom bonded to 2 oxygen atoms)? Carbon has 4 vacancies, but bonds to 2 oxygen atoms. Can you guess how many electrons must be shared in this scenario? After reviewing the details of H_2 and O_2, give it a try in the study guide.

   a. In the example here of hydrogen gas (H_2), the valence shell is the first shell, since hydrogen only has 1 electron in its neutral form. In this case, each hydrogen shares its electron with the other so that both have 2 electrons to fill their valence shell. The electron pair represents a single covalent bond formed between the two atoms of hydrogen.

   b. In the oxygen gas (O_2) example, let’s first think about each oxygen atom on its own. Oxygen has an atomic number of 8. Therefore, since 2 electrons fill the first orbital, there are 6 electrons in each valence shell and 2 vacancies. (Draw out two oxygen atoms in your study guide to visualize this.) What this means is that each atom needs to find 2 electrons to “get happy”. So, in the case of O_2, the two atoms share 2 electrons each with the other. Since each pair of electrons shared represents a covalent bond, this means that O_2 is formed by creating a “double bond”. The electrons being shared are represented in the model here as the electrons contained within the overlapping valence shells of each atom. Count the number of electrons in each valence shell of each atom. In this model, electrons 1 – 4 contribute to the valence shell of both atoms. Now, go back to the study guide and draw your O_2 using colored coded electrons.

14. **Polar covalent bonds** follow all the same rules as already discussed, but the difference here is that one participant pulls harder on the electrons than the other. Since water (H_2O) is such an important polar molecule to life sustaining processes we will use it as our example to demonstrate polar covalent bonds. Examine this model of water. There are a couple of things you should notice. First of all, oxygen has a lot more protons in its nucleus than hydrogen does: 8 in oxygen vs. only 1 in hydrogen. This means that the larger positive charge on the oxygen is pulling harder on the electrons that the hydrogen atoms can. So, the electrons spend more time hanging out near the oxygen than they do with the hydrogen. This, in essence, is the basis of a polar covalent bond. Although the molecule has an overall neutral charge (go ahead and count all the protons and electrons in the molecule to convince yourself), the tendency for the electrons to hang out on the oxygen end of the molecule means that that end of the molecule has a slightly negative charge, whereas the hydrogen portions have a slightly positive charge. It is this that makes it a polar molecule.

15. **Hydrogen bonds** are a little different than covalent or ionic bonds. Instead of occurring between atoms to form a molecule, hydrogen bonds form between two different polar covalent molecules, causing them to loosely "stick" together. Water, as we learned, has a partially positive end and a partially negative end. These partial charges attract each other and form loose associations with its neighbors. Notice in this image that the partially negative oxygen can form 2 hydrogen bonds with its neighbors, but the partially positive hydrogens can only form one hydrogen bond with one other molecule. This is because of the huge difference in proton numbers in the oxygen vs. the hydrogen. With the 8 electrons the oxygen owns straight out, and the two additional ones that contribute to the partial polarity, the charge is large enough to be able to hold onto the 2 very weak partial negative charges in the hydrogen. But because the hydrogen has such a weak positive charge with only the one proton, the charge isn't strong enough to hold onto more than 1 negative molecule at a time. Of all three types of bonds, hydrogen bonds are the weakest of all.

    a. The polarity of water is one of the reasons water makes such a great solvent and such an important molecule to life sustaining processes. These are 4 important aspects of water that we will explore in following Lectures. It is of the utmost importance that you have a strong understanding of the topics presented in this Lecture. Concepts such as bonding, organization of atoms, molecules and compounds, etc., will follow us through the entire course. It is for this reason that you should plan on spending a significant amount of time absorbing and understanding this Lecture before moving on to following Lectures.

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