LAB 2 – INTRODUCTION TO CHEMISTRY

(May 2014)

Section 1 – Atomic Structure

[2] Welcome back to the Bio 3 lab! This is Pat Farris again and I'm glad to see you've come back for more biology! But today, we're not actually doing any biology directly, because we need to start with the building blocks of organisms. And that means chemistry – you know, atoms, molecules, stuff like that. Let's get started.

[3] I suppose if we started from the VERY beginning, we could talk about how life began from just the simplest chemicals. These simple chemicals interacted for billions of years until slowly, more and more complex arrangements formed that we now recognize as the building blocks of life. These more complex molecules were things like proteins, lipids, carbohydrates and the biggie, DNA.

[4] About three and half billion years ago, these building blocks coalesced into a sustainable unit that we would recognize as a cell. These earliest cells would have had the means to collect and use energy, grow and reproduce. Luckily for us, these simple living units also started arranging into multicellular systems that we recognize today as plants and animals.

[5] We won't go through this whole evolutionary process today, but don't you think it's amazing that we can look at every organism on earth and see the same basic building blocks? Take a look at these creative structures made of Legos. The same little pieces, when arranged differently, can take on very different but meaningful shapes depending on how you put them together.

[6] It's the same with the DNA molecule. The DNA found in whales, bacteria, slime molds, chickens or crickets functions exactly the same way in all those organisms, the only difference is how the molecule is arranged.

[7] So, biology students obviously need to appreciate the roles of chemicals in sustaining life. For the next two weeks we'll examine the basics of chemistry and do some simple experiments. Now if you're already panicking about so much chemistry, relax! We're going to assume that you've not had any formal course work in chemistry. So, take a deep breath and let's start at the beginning of Section 1 in your lab book

[8] It is Democritus, a Greek who lived around 400 B.C., who is given credit for proposing the idea that everything on earth is composed of tiny, indivisible particles called atoms. The word atom comes from the Greek word for "uncuttable".

[9] So this is how we define the term atom today. Go ahead and record this definition in your lab book.

[10] Now, I'm sure that you realize that atoms are incredibly small - so small that they cannot be seen, even with the most powerful microscopes. As a matter of fact, the smallest object we can see with a light microscope contains millions of atoms.

[11] In your definition of atom, you used the term “element”, so let’s record that definition next. An element is a substance that cannot be decomposed into other substances by chemical means.

[12] This chart is called the periodic table and it lists all of the elements found on earth and even a few that are man-made. The elements are arranged according to their complexity. No, you don't have to memorize this table, but there will be several elements that you should become familiar with.

[13] Here you can see a few elements in their natural state. Each of these elements represents a pure substance - that is to say, a substance that cannot be decomposed into other substances. But we can combine elements to form new substances.

[14] If we take a couple of elements that you are breathing in right now, hydrogen and oxygen, and chemically combine them, you would make an entirely different substance called water. Water is not an element because it CAN BE chemically decomposed into other substances – back into the hydrogen and oxygen we made it from.

[15] Although there are 92 naturally occurring elements on earth, the six that you see listed here make up 99% of all living tissue. It is interesting to note that these are not the most abundant elements on earth, but for some reason the life forms that evolved on this planet found it advantageous to use these six to construct themselves. An easy way to remember them is to think of the acronym “C-H-N-O-P-S” and pronounce it "CHNOPS". Take a moment and record your answers before we go on to the table in Section 1.

[16] Chemists use an organized system to name the elements, using the first or first two letters of the Greek, Latin, French or English word for each element as a symbol. Here are some examples: "C" represents the element carbon, "S" represents the element sulfur, "P" represents the element phosphorus and "Na" represents sodium, which comes from the Latin word "Natrium", which means soda or salt.

[17] Now it’s time to introduce you to the elements we’ll be using in this class. Don’t panic, you won’t *necessarily* have to memorize every one on this table, but we do want you to have them entered in your lab book for reference. Your lecture instructor will let you know which ones are most important. But for now, please record the names of the elements and the number of protons in the table of Section 1. Return to the program when you’ve finished.

[18] All atoms, regardless of the element, are constructed from the same subatomic particles. An oxygen atom differs from a sulfur atom because each has a different number of these particles present; the particles themselves are exactly the same. Let’s look at these three particles important to biologists: electrons, protons and neutrons.

[19] These three subatomic particles are distinguished based on their electrical charges – the electron has a negative electrical charge, the proton has a positive electrical charge and the third particle, the neutron does not have an electrical charge -in other words, it is neutral. In addition to these charges, the subatomic particles are only found in specific regions of the atom – protons and neutrons are only ever found in the central nucleus of the atom and electrons are found in the outer energy shells. Please fill in the table at the end of Section 1.

Section 2 – Energy Shell Diagrams

[20] It may seem like a pretty simple arrangement with just three basic pieces, but think about what this means. Everything in the universe is built of those same pieces – protons, neutrons and electrons, so how do we arrive at the 92 different elements? Why do some substances float around us like xygen, and other substances are strong enough to hold up this building?

[21] It all has to do with how the protons, neutrons and electrons are arranged. Each element has a characteristic number of protons, as you saw in your table of elements. For example, hydrogen has one proton in its nucleus, carbon has six protons, oxygen has eight and chlorine has seventeen. Because each element has a characteristic number of protons, we assign what's called a unique atomic number. Make a note of the definition of atomic number at the beginning of Section 2.

[22] Now you remember where each particle is located - protons and neutrons are in the nucleus and electrons are located in the energy shells. When you measure the electrical charge of the entire atom, you find that the atom is electrically neutral. This is because an atom always contains the same number of electrons as protons. Therefore a hydrogen atom contains one electron in its energy shell to balance its single proton charge; carbon has six electrons to balance its six protons, oxygen has eight and so on.

[23] Atoms are often diagramed in a way that gives you the impression that the electrons are confined to circular tracks like race cars or satellites orbiting the earth. In reality, electrons are not confined to single tracks, but instead are confined to certain volumes of space called energy shells. Let's see what an energy shell really represents.

[24] This image shows the region where we could find the single electron of hydrogen around its positively charged single proton. In fact this image really should be in 3-D so you could see the full range of the electron's movements. So even though the electron is NOT confined to any particular path, isn't it much easier for us to show the energy shell as a single circle around the nucleus with a single dot indicating the electron? I thought you'd like that better than a big smudgy cloud.

[25] Examine the two atoms shown at the beginning of Section 2 – one of hydrogen and one of nitrogen. Hydrogen has one proton, nitrogen has seven. We know that the atoms will be electrically neutral, so the hydrogen will have just one electron, and the nitrogen has seven, but notice that there are two energy shells on the nitrogen.

[26] The distribution of electrons within these energy shells follows a very definite pattern. For the elements we're most concerned with, the first energy shell has to fill before going on to the next outer shell. When we say an energy shell "fills", we mean that the shell has reached its capacity and must use another energy shell if there are more electrons than that first energy shell can hold.

[27] This table, which is also in your lab book, shows the capacities of each energy shell. You should memorize these capacities because they are critical for understanding how different elements will bond together.

[28] Let's start with an easy one like lithium. Lithium has an atomic number of 3, so we know it has three protons. An atom of lithium will have three electrons as well, so how many energy shells will it have? Well the shells have to fill up in order, so it's a bit like grabbing the seats available in a car. Once the first two seats are filled, you have to go to the back seat.

[29] So our lithium atom has to have a second shell to accommodate that third electron, because the first shell can only hold two.

[30] The capacities of each shell are very important because this will determine how a particular element acts. Atoms react with other atoms because they are attempting to fill their outermost energy shells with electrons. For example, hydrogen has one electron in its first and only energy shell. Remembering the first energy shell can hold up to two electrons, hydrogen reacts with other atoms in an attempt to fill that shell with electrons.

[31] Before we go on, take a moment and complete the energy shell diagrams for the elements carbon, oxygen, sodium and chlorine at the end of Section 2. Return to the program when you have finished.

[32] You can check your diagrams here. As you placed your electrons on the shells, you may have wondered exactly *where* on the shell to put them. Well, it doesn't really matter how you arranged them, because electrons are free to move anywhere within that energy level. What is critical, though, is that you have the right number of electrons on each shell.

Section 3 - Molecules

[33] Now, we know that atoms react with other atoms because they are trying to fill their outermost energy shells with electrons. When two atoms react and chemically combine, they form a molecule. Please record the definition in your lab book.

[34] While we're talking about molecules in general, there are two different types. Please record the definitions of compounds and diatomic molecules, and notice that they differ in how many elements are present.

[35] Now we're ready to look at these different combinations. We list the elements found in a molecule in something called the chemical formula. Each formula also tells you exactly how many atoms are present as well, and that's shown by the little subscript below the symbol for each element.

[36] Take a look at the chemical formula for a substance called sodium nitrite - a commonly used food preservative. This molecule is composed of three elements - sodium, nitrogen and oxygen. So is it a compound or is it diatomic?

[37] Yep, a compound. In fact, almost all the molecules we look at in a biology class will be compounds. Now how many oxygen atoms are there in the molecule?

[38] Right again! Two atoms of oxygen. But notice that the symbol for sodium, Na, is not followed by a subscript. This is how chemists indicate that there is just one atom of that element.

[39] Now is a good time to introduce you to some common biological molecules and at the same time, get you used to looking at some chemical formulas. Take a look at the table in your lab book showing some common biological molecules. Determine the type of molecule, whether it's a compound or diatomic, and complete the list of elements found in each. You will have to memorize the names of these common biological molecules, but I bet you already knew a couple of them. Return to the program when the table is complete.

[40] Another place you might see a number in relation to a chemical formula would be a large number in front of the formula, as shown here. This large number indicates how many molecules total there are, and this will come up when we start to look at chemical reactions. In this case, there are three molecules of water, and each water molecule is made up of two atoms of hydrogen and just one atom of oxygen.

[41] Just to make sure you're getting all this, try this question before we continue.

[42] Okay, smartypants molecule expert, here's the chemical formula for a molecule called glycine with a bit of a twist. We just looked at a chemical formula where the total number of atoms of each element is shown in the subscript. But notice that this formula shows groupings of atoms. This helps us figure out the arrangement of the atoms. If you look at the stick diagram of the same molecule, or what is called the "structural formula", you'll notice it shows the same groupings. Clever, huh?

[43] Now take a minute and finish up Section 3. I'll have a couple of questions for you when you get back.

[44] Let's see how well you're getting all this chemistry stuff. This is the formula for lactic acid - a molecule that makes sour milk taste sour. See if you can answer this question.

[45] Looks like you need something more challenging! Give this one a try.

[46] Congratulations! See how easy chemistry can be? Now, let's start putting some atoms together.

Section 4 – Chemical Bonds – Covalent Bonds

[47] There are several different types of chemical bonds that we will cover, and they all have their little peculiarities. Sometimes the parts being bonded together are atoms, sometimes we are bonding whole molecules together or parts of molecules. Record the definition of a chemical bond in your lab book.

[48] Before we begin our discussion of chemical bonding, it's critical you understand the basic chemistry we've gone through already. Please take a moment and answer the review questions at the beginning of Section 4 regarding subatomic particles and energy shells before we continue. If you can't answer these questions without peeking back at your lab, maybe you should review a bit more before continuing… I'm just sayin'…

[49] Now that you've reviewed a bit, we're ready to start with the most common bond, the covalent bond. A covalent bond is one that results when an atom shares one or more of its electrons with another atom. Please record this definition, and then we'll look at some examples.

[50] Here's a simple example of this property of sharing electrons. These two hydrogen atoms are sharing a pair of electrons, resulting in a molecule of hydrogen gas. Note that the atoms are drawn with overlapping energy shells. The energy shells are drawn this way because each electron is free to move around either nucleus.

[51] Here's the same molecule of hydrogen gas, shown with the beginning two states of the hydrogen atoms it was constructed from. Each hydrogen atom began with just one electron, but the energy shell ideally would contain two. If the two atoms essentially agree to share their electrons, for at least part of the time, each hydrogen nucleus will be surrounded by two electrons in the energy shell. This sharing of electrons is the force or bond that holds these atoms together and is called a covalent bond.

[52] Because the electrons are being shared, this causes the covalent bond to be a very strong bond. Once established, neither atom wants to give up its full complement of electrons. When the outer shells of an atom are filled, we say that the shell, and therefore the atom, is "satisfied".

[53] So far we've shown the covalent bond as overlapping energy shells, but here's another way to show the covalent bond – as a line between the symbols for hydrogen atoms. In other words the structural formula for the molecule. Whether we use overlapping energy shells or a line, they mean the same thing – two electrons are being shared between atoms.

[54] Here's another example of covalent bonding. This diagram depicts the formation of a water molecule. Remember that an oxygen atom has six electrons in its outer energy shell and eight are required for stability. Well, by sharing two of its electrons with two hydrogen atoms, oxygen will have eight electrons, at least part of the time, in its outer energy shell. Each hydrogen atom, in turn, will have two electrons in its outer shell--one of its own and one from oxygen. So in this case, each hydrogen atom and the oxygen atom are satisfied with filled outer shells--at least for part of the time. And I suppose it's better to be satisfied part of the time than at no time at all.

[55] Let's review the three different ways of showing the molecules. Here is a molecule of methane, CH_4. Record these different depictions carefully in your lab book, then come back to the program to answer a couple of questions.

[56] Take a careful look at the energy shell diagram of methane and compare it to the single atom of carbon you drew earlier in the lab. How many electrons were around the outside shell of carbon before the bonds formed? In other words, how many covalent bonds can an atom of carbon form?

[57] Is methane an example of a compound or a diatomic molecule?

[58] Atoms of a particular element consistently form the same number of chemical bonds. The reason for this, of course, is that these atoms always require the same number of electrons to fill their outermost energy shell.

[59] Now we can summarize this basic rule by taking a look at the three molecules shown here – water, ammonia and glycine, which are also shown in your lab book. I've shown you examples of hydrogen, carbon, nitrogen and oxygen forming covalent bonds in several different molecules. Now see if you can determine the number of bonds each of these elements consistently forms. Please fill in your answers in your lab book.

[60] Did you see the pattern? Nitrogen has room for three more electrons on its outside shell, so it can form three bonds. Oxygen has room for two more electrons, so two bonds are possible on oxygen atoms. The double covalent bonds in the glycine may look a little strange, but it just means that there's two covalent bonds between the same atoms. Oxygen and carbon tend to do this, so don't be surprised to see double covalent bonds when we get to the larger molecules.

[61] Now for a question - if a regular covalent bond is the sharing of two electrons, how many do you suppose would be involved in the double covalent bond?

[62] Great! I think you're really getting this stuff! Let's go on.

[63] Okay, now the fun part – we'll start building some stuff. So far, we've been representing molecules as simple two-dimensional diagrams on paper. In reality, molecules are three-dimensional structures and you'll need to appreciate the three dimensional nature of molecules to understand why they do certain things.

[64] While the shape of a certain molecule may not seem terribly important to you, I can assure you that it is extremely important to your cells. There are many molecules our cells detect and recognize by their three-dimensional shapes. We won't have you build anything as gnarly as these giant molecules, but check out the shapes – amazing what a few carbon, hydrogen, nitrogen and oxygen atoms can do.

[65] Before you begin your constructions, notice the rules for building molecules shown in the table of Section 5. In addition to the correct number of covalent bonds on each atom, the bonds will have to be certain lengths. Hydrogen is such a dinky little atom, it will be closer to the other atoms because it only has the one energy shell. Make sure you use the short bond when bonding hydrogen.

[66] Now take a crack at building the molecules shown at the end of Section 5 – water, methane, urea and glycine. Take your time and when you have completed the glycine, take the model up to the lab instructor for checking. Good luck!

[67] Now that you've completed a big molecule like glycine, this question shouldn't be any problem for you.

Section 5 – Chemical Bonds - Ionic Bonds

[68] Our next type of bonding is called the ionic bond. Please record the definition in your lab book – an ionic bond is formed when one atom donates an electron to another, resulting in oppositely charged ions.

[69] This term ion is new for us, so take a moment and record the definition of an ion – a charged particle.

[70] Certain kinds of atoms have a tendency to give up an electron or two rather than to share electrons with other atoms. Conversely, some would rather receive electrons from another atom than share. One element that would rather give up an electron is sodium. Notice that the sodium atom has one electron in its outermost energy shell. It would be difficult for sodium to pick up enough electrons to fill its third energy shell with electrons, so sodium atoms do something else to reach stability. This is one of the "beginning state" atoms shown in your lab book, so you could copy it now, if you like.

[71] Now take a look at this chlorine atom. Note that chlorine has seven electrons in its outer shell. The third shell can become stable if eight electrons are present even though it could hold 18. The chlorine atom, then, needs just one more electron to reach a stable configuration. Go ahead and copy this "beginning state" in your lab book as well.

[72] Can you predict what these atoms will do when they're brought together? Ahhh, you're pretty clever. Of course, sodium gives up its third shell electron to chlorine and its new outer shell is its second. Look carefully, you'll note that sodium's second shell contains eight electrons and is stable. Chlorine has its outermost energy shell stabilized now because it has eight electrons. This arrangement works out nicely for both. Copy this "end state" diagram carefully in your lab book.

[73] Now, what holds the sodium and chlorine together? Compare the number of protons in sodium's nucleus with the number of electrons it has after donating one electron to chlorine. Sodium is no longer neutral, but has a net positive charge because it is missing an electron. If you count the protons and electrons in chlorine, you'll note that it now essentially has an extra electron compared to the number of protons, so it carries a net negative charge.

[74] And now we no longer technically have "atoms" – we have electrically charged ions. And since they have opposite charges, they will stick together like opposite ends of a magnet. So the sodium chloride, or table salt, that we just made is an example of a substance held together by ionic bonds.

[75] Notice that the final state of sodium chloride must show the electrical charges to be an accurate depiction of an ionic bond. Finish Section 5 by looking over your "beginning state" and "end state" diagrams, and make sure you have recorded the charges on the sodium and chlorine ions. Return to the program when you are ready to begin Section 6.

Section 6 – Chemical Bonds - Hydrogen Bonds

[76] The last type of bonding we'll cover today is the hydrogen bond. Please record the definition of hydrogen bonding – a weak attraction between molecules or between parts of very large molecules.

[77] The hydrogen bond may be weak, but it's extremely important to biology. Two examples of molecules that have lots of hydrogen bonds are water and DNA. We'll get to the DNA later in the semester, but for now let's take a closer look at water.

[78] Take a look at that hydrogen bond definition again – this bonding is between molecules, not atoms or ions. You've already built the three dimensional model of water, and you know it has sort of a tweak to it – in other words, the hydrogens stick off to one side because of how the electrons relate to the oxygen.

[79] When you think about where the electrons are spending most of their time, you know that there are more electrons piled up on the oxygen side of the molecule. This results in what we refer to as a partial negative charge. It's not the same as being a negative ion, but it does mean that there is a slight negative charge on that side of the molecule.

[80] Now if we look at the other half, the hydrogen half, we see there are single protons of the hydrogen atom's nucleus with a positive charge. The electron of the hydrogen atoms are being shared with the oxygen atom because of the covalent bond, so now we've got a partial positive charge at each hydrogen nucleus.

[81] This results in what we call a polar molecule – a molecule which has a positively charged side and a negatively charged side, just like a magnet has two oppositely charged ends. Record these partial charges on the water molecule shown in your lab book.

[82] So when two or more water molecules meet, they will arrange themselves to have the opposite charges next to each other. The oxygen of one molecule will be next to the hydrogen of the other molecule.

[83] Now see if you can figure out the proper arrangement of four water molecules in the space provided in your lab book, making sure you indicate the partial charges that are holding them all together. Continue when you're ready for Section 7.

Section 7 – Chemical Reactions

[84] Okay, now we're ready to tackle some chemical reactions. If you've never had chemistry before, it might look a little intimidating, but let's start with the basics.

[85] Take a look at this outline of a chemical reaction - substances A and B are called the "reactants". These are simply the substances that react with one another during the reaction and are written to the left of the arrow. Substances G and H are called the "products" because they are produced when chemicals A and B react. The arrow always points to the products. The letter "E" written above the arrow represents an enzyme and is always written above the arrow.

[86] So let's look at an actual chemical reaction. Examine the reaction shown in your lab book at the beginning of Section 7. It shows a single reactant, two products and an enzyme. Fill in the blanks below the reaction to identify each component of the reaction.

[87] The chemical decomposition of hydrogen peroxide is an example of a chemical reaction that occurs in our cells. It is an important reaction since hydrogen peroxide is a toxic substance produced by our cells, could destroy the cell if it were not broken down into some harmless products like water and oxygen.

[88] Enzymes are very powerful chemicals that will enable the reaction to happen much faster than it normally would. Almost every reaction happening in your body right now is using an enzyme, so you can see why we need to appreciate them before we go on with a biology class.

[89] Enzymes work by lowering the amount of energy required by the reaction. We'll get to the details of enzyme function next week when we introduce organic molecules, but for now just believe me, we would NOT want to wait around for reactions that don't have enzymes to speed them up. We've got things to do, experiments to run!

[90] For our first chemical reaction, we'll look at a reaction involving an enzyme called tyrosinase. Check out the reaction in your lab book and notice the tyrosinase written above the arrow of the chemical reaction.

[91] This reaction is one you should memorize, because it just may pop up again next week, if ya know what I mean. So here's what to look for: this reaction has two molecules of a chemical called pyrocatechol that react with one molecule of oxygen to form two molecules of quinone and two molecules of water. Tyrosinase is the enzyme required for the reaction. Don't worry about memorizing the chemical formulas for each molecule, but you should know their names.

[92] The reason we chose this reaction to examine is that it results in a noticeable color change. The reactants, pyrocatechol and oxygen, are colorless but the product, quinone, is yellow and quite easy to find in a test tube. Remember- quinone is yellow!

[93] You're now going to get a chance to play the "mad scientist". What I'd like you to do is to read the directions very carefully and complete Section 7 and all the little questions along the way. You'll find the materials that you need either at the lab sink or at the demonstration table. Come back to the program when you are ready for some questions regarding the tyrosinase reaction.

[94] Welcome back! I hope you rinsed out your test tubes and returned everything to the sink area, if not.... please do so before going on. Let's try a question:

[95] Identify the role of each of these chemicals.

[96] Which of these would be a yellow substance?

[97] Well done! To finish up Section 7, let me mention why the experiment was designed this way. We did a couple of test tube set-ups where we didn't really expect the reaction to happen – after all, tubes A and B didn't have all the substances required for the reaction. But they certainly confirmed that the color change was the result of a reaction involving *both* pyrocatechol and tyrosinase. By doing these control tubes, we know that tyrosinase doesn't turn yellow all by itself when left alone in a test tube for five minutes. If this were true, then test tube B would have been yellow as well. These types of confirmation tests are called "controls".

Section 8 – pH, Acids and Bases

[98] In addition to enzymes, there are many other things that can affect chemical reactions. For example, the acidity or alkalinity of the solution can have a major impact on the reaction. For this reason, the last topic I'd like to discuss today involves acids, bases and pH.

[99] Take a look at the pH scale at the beginning of Section 8. I'm sure you've seen one of these before, and you may already know that substances with a pH of less than seven are called acids, pHs of more than seven are called bases, and a pH of exactly seven is considered neutral.

[100] The easiest place to start is with neutral. Go ahead and record this definition before we go on. A substance which is neutral is not acidic or basic – it's at exactly seven on our scale.

[101] Our next definition will be for acids. Acids are substances that release hydrogen ions when placed in water. Please record this definition and I'll give you some examples.

[102] Consider the chemical reaction shown here. It shows that when H_2SO_4 is placed in water, a hydrogen ion is released. When a chemical comes apart like this, we say that it dissociates, and because it's a hydrogen ion that comes off, we know that H_2SO_4 is an acid. As a matter of fact, this is the same acid that is put into car batteries - it's called sulfuric acid.

[103] Now think back to what a hydrogen ion would be – it's a hydrogen atom that has lost its only electron, so it has to have a positive charge because all that's left - a proton. Answer the questions about acids before going on.

[104] Now the other end of the pH scale shows the relative strength of bases. You may also have heard the term "alkaline" – they mean the same thing. A base is a substance that releases hydroxide ions when placed in water. A hydroxide ion consists of a hydrogen atom combined with an oxygen atom. This aggregation is called an ion because it contains one more electron than the number of protons. Please record the definition of base and answer the question about the charge of hydroxide ions before I show you a reaction involving a base.

[105] Here you see what happens when potassium hydroxide is placed in water. Remember that "K" is the symbol for potassium. Note that it dissociates into potassium ions and hydroxide ions, or O-H ions. Because it releases hydroxide ions, it's a base. The potassium ion is positively charged because it loses one of its electrons to the hydroxide group when it is placed in the water. This accounts for the extra electron that makes the hydroxide group negatively charged.

[106] Now I think you can figure out how sodium hydroxide would dissociate on your own. Don't forget to put the proper electrical charges on each of the resulting ions.

[107] We'll finish up Section 8 by testing the pHs of a variety of household substances. Read the instructions carefully and return to the program when you are ready to begin Section 9.

Section 9 – Effect of pH on Enzyme Activity

[108] Okay, we've got one last experiment to put it all together. You'll be doing the same chemical reaction you did in the first experiment, but now you'll be testing the effect of pH on your results. Be sure to answer the short review questions at the beginning of Section 9 before you run off to get your test tubes. After checking your results against the "standards" on the side of the lab, get your graph checked by the lab instructor before you come back.

[109] Well you've finished your second Bio 3 laboratory, and you survived chemistry. It wasn't really that scary after all, was it? Did you rinse out your test tubes and clean up your booth? Excellent! Let's take a look at your lab and make sure you've answered all the little questions at the end. If so, you're free to go and we'll see you next week.