PROFESSOR: Here is the game plan for the next two or three lectures. I'm going to start by talking about the chemical forces that are important for the structure and function of these biomolecules. And then I'm going to relate them, as we go along, to how these properties influence the characteristics of these key macromolecules. And in particular we'll be talking about covalent bonds, hydrogen bonds, ionic bonds, a force known as van der Waals forces, and something that's not really a force but it's a characteristic that's very important, particularly when we think about proteins and lipids, called hydrophobicity-- literally "fear of water."

And then the order of the molecules. As we talk, I'll talk about carbohydrates first. I'll try and get to that today. Then we'll talk about proteins, nucleic acids, and lipids, in that order. As you'll see, these two will be sufficient to understand most of the characteristics of carbohydrates. Whereas we're going to need all five of these to be able to get an intelligent understanding of how proteins work.

Now, I'll caution you. It's going to seem-- God, he's going to talk about covalent bonds, as everybody is rolling their eyes. I heard about covalent bonds in grade one or something. But the difference here is that we're going to be looking at some of these forces, some of which you've been exposed to already, but from a biological perspective. And I hope if you kind of watch that, you'll begin to see that you're looking at something that may be sort of familiar to you. But you have to start thinking about it in a different way once you start thinking of what are the implications of the properties of these forces and the way these molecules behave for biology.

So, begin with the one that everybody undoubtedly knows, which are covalent bonds. And this is the principal force that holds atoms together. And it's based on sharing electrons. And as I'll say, these are very strong bonds.

And so in the simplest sort of example, hydrogen atom has one unpaired electron, a carbon has four. And so you can make methane, CH_4. And commonly in chemistry and biology we use a line to represent a pair of electrons.

So there's methane. As I said, apart from you know it burns, if you go out in a swamp or in a beach and you see bubbles, muddy bottom coming up, those are bubbles of methane made by methanogens that are living in the anaerobic layer underneath. A cow has a special fermentation, digestion cavity inside. It's huge, called a rumen, stuff sloshing around. And it's full of archaea, that are methanogens.

And a cow makes about 400 liters of methane a day. And Penny will tell you, it's a very bad greenhouse gas. It's much more potent than carbon dioxide.

And so the typical length of a covalent bond is about 1.5 to 0.2 nanometers. And I hope you'll try and begin to get a sense of the links of some of these things, too. But the key point about this is to break a carbon-carbon bond needs 83 kilocalories per mole. So that's a lot of energy.

At 25 degrees centigrade, if you take, say, a typical vibrational mode of a covalent bond, the energy that it has is about 0.6 kcals per mole. So what that means is that covalent bonds don't break on their own under physiological conditions. They can bend, they can rotate, and they can stretch. So they're back and forth this way, they can go this way, this way, but they don't break.

And so this sort of leads to another topic that we'll talk about, which is utterly key-- It's one of the secrets of how life works-- are these protein molecules that are known as enzymes. And we'll also talk a little bit about a similar thing made of RNA called a ribozyme. But what these are are biological catalysts that enable specific bonds-- and this is important-- specific bonds to be broken or formed under physiological conditions.

And this part is so important. If you're trying to work out a chemical reaction, the original process for taking nitrogen gas and making ammonia, the Haber process, involved some very, very tough molecule to break the bond of. So just heat it up to 500 degrees and put in a catalyst.

But if you're a living organism, you don't have that option. You have to continually make and break bonds under the conditions -- the very, very narrow conditions where life is possible. If you go a little too high, things like proteins unfold. And then they don't work as properly as machines anymore. So we'll be talking more about that as we go along.

There are different types of covalent bonds. And again, the first part of this isn't going to surprise you. There are single bonds, like this. There are double bonds, and triple. Excuse me. I'll just stay with carbon for the moment. The more electrons that are shared, the stronger the bond.

And these two are referred to, if it's a carbon compound, as being unsaturated bonds, the same term you hear when you hear about unsaturated fats. And what that means is a fat with an unsaturation, that's unsaturated, will have somewhere in it a double bond, or in some cases, many double bonds. However, there's another aspect of this which might not have been relevant to you, but you'll see it becomes relevant for thinking about proteins as soon as the next lecture.

And that is, a single bond is able to rotate this way. These guys can't rotate. And that, as you'll see, becomes important in quite a variety of situations. But we'll run into a very important example of that when we're thinking about the very backbone of all proteins, the peptide bond, which is at the heart of being a protein.

There are other molecules that have more than one bond that are important. Oxygen is one. And nitrogen, as I said, is a particularly hard nut to crack. Most organisms, as I said the other day, are unable to break this bond. The only organisms that have learned how to do it are bacteria.

The vast majority of them use one, single enzyme called nitrogenase that evolved that's a very complicated enzyme and has very, very stringent requirements and needs a huge energy input. But it is able to crack this bond and get it made into ammonia. But it's an example of another molecule that has a triple bond in it.

Let's see, how are we doing here? Okay. So another aspect of these covalent bonds that you need to think about has to do with when you're thinking about carbon. And it's a property called chirality. And it comes from the fact that carbon has four bonds but they come out as a tetrahedron.

So that doesn't matter in the case of methane. But I'm going to depict the tetrahedron in this way, so that this bond is coming -- these two are in the plane of the board, this one's coming out, that one's going back. And let's just put on four different substituents.

Now if I get the mirror image of that, we will have-- these two molecules are called optical isomers. And if you sit down and play with this, you will find you can't convert one to the other without actually physically breaking a bond. And this is really important, one of the central concepts that I hope you might remember from this course because it cuts across a lot of the stuff talk we'll be talking about. At a molecular level, much of biology relies on the interaction of complementary 3D surfaces.

We're actually very familiar with this at a macro level in our own lives. Imagine you've just come back from the party late on Saturday night, you're crossing the Mass. Ave. Bridge, the wind is howling, you're freezing. But no problem, you've got your gloves. And you reach in your pocket and you have two left gloves. No matter what you do, you can't get that right hand to fit properly into the left-handed glove.

One's a mirror image of the other. But we run into this problem even in our own lives. When you saw how that DNA had fit right into a groove in the protein. If we had a mirror image of the DNA or we had a mirror image of the protein, it wouldn't work. This principle goes all the way through biology.

There is another characteristic of covalent bonds that becomes important again. And that is how equally the electrons are sharing. So again, it goes back to the sharing of electrons, but with a twist.

If we have a carbon-carbon or a carbon-hydrogen bond, it's pretty much equal sharing. And this is known as a nonpolar bond. But if you have a nitrogen or an oxygen bond, it's unequal. And these are known as polar bonds.

And the term that's used to describe this unequal sharing of electrons is known as the electronegativity of the atom. It's basically a word that means the greediness of a particular atom for electrons. So if you have an oxygen and a hydrogen bond, although we write it like that on the board and you've undoubtedly seen this for many years in chemistry, in fact, the electrons spend more time down here than they spend up there. So there's a little bit of a negative charge on the oxygen and a little bit of a plus charge on the hydrogen. That's usually represented by a little delta to indicate that this has a wee bit of negative charge, that has a wee bit of positive charge.

And a molecule that's very important with respect to this is water. Because water, as you know, is H2O. But it's not symmetrical. The angle here is 104.5 degrees. And so the oxygen has a little bit of a negative charge but each of these has a little bit of a plus charge.

Actually, water is 55 molar. So it's a little dipole. You've got 55 molar, these little dipoles going on.

This property of electronegativity and nonpolar bonds then leads to the second of the forces that we're going to be talking about. That's force number two. And that's a hydrogen, or H bond.

And this is a bond that's made possible by a little bit of a negative charge that's on oxygen, or nitrogen, or a few other molecules and a little bit of a positive charge that's due to the hydrogen that's in a polar bond. This is very important, as you'll see, for proteins, nucleic acids, and for carbohydrates. And it has a huge amount to do with the way that water behaves.

Because in that 55-molar water, you'll have one water molecule that's going to be like this. And there will be another water molecule down here with a little bit of a negative charge. And this a little bit of a plus charge on this hydrogen and a little bit of a negative charge can form what's known as a hydrogen bond between them.

And what's especially important about these hydrogen bonds is they're about 1/20 the strength of a covalent bond. And that means that in a distribution of molecules at physiological temperatures, there will be some guys up in the -- the most energetic molecules within the bunch will have enough energy to break hydrogen bonds. But they're much easier to do.

And just to peer ahead, when we talk about replicating DNA, those two strands are held together by hydrogen bonds. So the backbones are really solid, just like two strips of Velcro or something. But the hydrogen bonds hold the two strands together, but 1/20 the strength. So it's basically like molecular Velcro between the two strands of DNA. And we'll see some more examples of this.

Let's see if I can go back to this and get this thing to play. This is static representation just illustrating this. But in fact, what happens, water molecules are continually changing partners. So they're constantly making shells, and cages, and so on.

And the next little movie is a picosecond simulation of water just at zero degrees. And you can see how the molecules are changing partners, making little shells and things.

And here's a picosecond simulation of water at the boiling temperature. And what you can see from this is every now and then, a molecule like this one will get enough energy to break out of this constant sharing of little hydrogen bonds and escape. And another thing, when we talk about getting something dissolved in water, this is something we'll have to think about. Because if you try and dissolve something in water, like stir a lot of oil into it, you know what happens. You can stir like mad and doesn't go in.

Part of the problem is if you put something in the water, it's going to have to break these existing hydrogen bonds. And that's an energy cost. So in order to get something to dissolve, you're going to have to get the energy back. And we'll be talking about that. But it's one of the fundamental parts of water.

You're familiar with the characteristics of water. There's surface tension. It's why trees can grow 300 feet tall, because they've got basically little nanotubes and little capillaries. And with this surface tension, water, due to hydrogen bonds, can go 300 feet up. The water can go right up.

You've seen water bugs walk around on water. There's a particularly interesting lizard in South America, Central America called the basilisk lizard that's about 2 and 1/2 feet long. It's able to run across the top of the water. It's actually called the Jesus Christ lizard. And it's able to do that because of this surface tension in the water.

In fact, when I finished my Ph.D. Thesis, I went in a competition for the theses. And mine was something like, a chemical enzymatic synthesis of oligoribonucleotides. And I was competing against a guy who said why do lizards run on water, and his entire talk consisted of movies of this thing running across the water. I thought I was toast, but I actually won that prize. But anyway, every time I see this I remember it.

For example, when they go and explore Mars or think about planets, they're always looking for water because it has this very, very special set of properties that are so important for life.