

Chemical Interactions

Slide 2 Don't worry. We're not going into excruciating detail about the periodic table, but it is important to remember that living organisms are governed by the laws of physics and chemistry. Out of all the elements listed on the periodic table, only a handful of them, hydrogen, carbon, nitrogen, oxygen, phosphorus, and sulfur are vitally important in biological systems. At the moment we are particularly interested in looking at how these elements interact and bond together.

Slide 3 Let's do a quick review of atoms. You probably recall that atoms consist of positively charged protons, negatively charged electrons, and neutral neutrons. Atoms are electrically neutral. That is, they have the same number of protons as electrons so that the electrical charges balance. This fact becomes important when we look at the arrangement of electrons around the nucleus. The more massive protons and neutrons make up the atom's nucleus while the electrons orbit the nucleus in different shells. The first shell can only hold up to two electrons. Each shell beyond that can contain up to eight electrons. The bottom line is that electrons like to travel in pairs and all the electron shells want to be full. When we say the electron "likes" this, or the electron shells "want" that, what we really mean is that the atom is more stable when it has what it "wants".

Slide 4. Because atoms are electrically neutral there are only going to be as many electrons as there are protons in an atom, and the electron shells will not necessarily be full. Let's take a closer look at a phosphorous atom. As you can see, the two inner shells are complete. The first shell has two electrons, and the second shell has eight electrons. But the third shell only has five electrons. This means that there are three unpaired electrons AND an empty shell. This is an unstable (and unhappy) atom. Atoms like this one whose outer shells are not full are much more likely to react than those with full outer shells. The ways in which atoms maintain electrical neutrality while filling their outer shells with electrons (~~the octet rule~~) lead to chemical interactions and bonding.

Slide 5 Ionic bonds are fairly simple to understand. They form between atoms with very different electron arrangements. The classic case is NaCl, or table salt. Sodium (Na) has only one electron in its outer shell – that's one lonely electron. Chlorine has 7 electrons in its outer shell – it "wants" to get that eighth electron to complete the shell (in chemistry this is referred to as the octet rule). Both atoms are unstable because of their electron arrangements. In ionic bonding there is a complete transfer of that lonely outer electron from sodium to complete the outer shell of chlorine. Now there is one negatively charged ion (Cl⁻) and one positively charged ion (Na⁺) with a strong attraction for each other. The electrical attraction between the two ions holds them together. Just remember, when it comes to ionic bonding, opposites attract!

Slide 6 Covalent bonding involves a sharing of electrons among atoms. Let's think about carbon since it is about the most important biological element. Carbon has 4 electrons in its outer shell available to interact with other electrons. Gaining or losing one electron as in ionic bonding will not do much to complete carbon's outer shell, so

atoms like carbon have to do things differently. As 2 atoms with unpaired electrons move closer to each other, the nuclei with their positively charged protons attract the negatively charged electrons of both atoms. The atoms end up sharing a pair of electrons, so both atoms are more stable. Carbon is most stable if it can find four atoms with unpaired electrons to share with it as in this example with methane in figure a). Figure b) simply shows some different ways of representing a covalent bond. Covalent bonds may be single – where two atoms share one pair of electrons – or multiple with more than one pair of electrons shared between 2 atoms. Double bonds between carbon and oxygen are pretty common. When you get a chance, go back and look at the periodic table to figure out why that is.

Slide 7 Atoms that participate in covalent bonding don't always share the electron pair equally. One nucleus may exert a stronger force on the electrons and so pull them a little closer. This attractive force that an atom exerts on electrons is called electronegativity. Whichever atom tends to pull the electrons closer will have a slightly negative charge to it (indicated in the figure as δ^-). Molecules with a little negative charge at one end and a little positive charge at the other are called polar molecules, and the bond is called a polar covalent bond. Water is probably the most well known (and maybe the most important) polar molecule. Its polarity is part of what makes it the “universal solvent.”

This idea of polarity gets very interesting as we start looking at larger molecules. There are many simple molecules that can be polar or nonpolar. Water is a polar molecule. Hydrocarbons are nonpolar. In the covalent bonds that form between carbon and hydrogen the electrons are shared equally. The ethane molecule in this slide is a nonpolar molecule. There are also many large molecules that may have polar or nonpolar regions.

Slide 8 The last type of bond we are going to look at is the hydrogen bond. Hydrogen bonds form between a positively charged (δ^+) hydrogen and the negatively charged (δ^-) end of a polar molecule. The most obvious example is water where the oxygen atom from one water molecule is attracted to the hydrogen atom of an adjacent molecule. Hydrogen bonding in water gives it many of its rather unusual properties like surface tension and cohesion.

Remember that hydrogen bonding isn't just found in water. Hydrogen bonds help hold proteins together and make the double helix of DNA possible.

Slide 9 We've looked at 3 types of bonds so far. This table points out the main features of ionic, covalent and hydrogen bonds and gives you an idea of which bonds are stronger. Covalent bonds are by far the strongest – about 10 times as strong as ionic or hydrogen bonds. Why are covalent bonds so much stronger? Let's recap what forms the three types of bonds. Ionic bonds result from the swapping of an electron. Hydrogen bonds pull partial positive and partial negative charges together. But in a covalent bond, an electron is being shared between two atoms. It may help to think of it this way: if you are sharing a dinner plate with someone, you have to sit very close to them so you both

have equal access to the food. This need for close proximity is what gives a covalent bond its strength.

Van der Waals forces and hydrophobic interactions are two chemical interactions that are significantly weaker than the three main chemical bonds. Van der Waals forces, which arise from the interaction of electron clouds, are relatively weak forces but can produce a big effect in large numbers. This type of interaction is what allows a gecko to climb up walls. Hydrophobic interactions have to do with the tendency for nonpolar molecules to interact among themselves. The classic example of this type of interaction is the “oil and water don’t mix” thing. When you put oil in water, you get a blob of oil in the water because polar and nonpolar molecules don’t like to interact. Hydrophobic interactions are part of what makes lava lamps work.

In this table you can also see some of the conventions for drawing the different types of bonds. The solid single or double lines indicate a single or double covalent bond. The dotted lines indicate a hydrogen bond and so on. You’ll see these representations throughout the course.

Slide 10 Now that you have seen how atoms interact with each other, let’s look at a few of the biologically important groups they form. A functional group is defined as a small number of atoms covalently bonded to each other in a particular way. What’s interesting about functional groups is that they have the same chemical properties no matter what larger molecule they are a part of. Take the hydroxyl group found in alcohols for instance. This functional group consists of an oxygen atom covalently bonded to a hydrogen atom. No matter where you find this group, it has the same chemical properties. For example, hydroxyl groups can break away from the rest of the atoms they are attached to, combine with a hydrogen atom from another molecule, and form water. In addition to hydroxyl groups, there are three other functional groups that you will encounter frequently in biology: amino groups, carboxyl groups, and phosphate groups. Amino acids (the stuff that makes up proteins) all have an amino group and a carboxyl group on either side of a central carbon atom. When your cells make proteins, these two functional groups interact to form a covalent bond that holds the amino acids together. Phosphate groups are found all over the place in your cells. This is probably because the properties of phosphate groups make them ideal for energy transfer reactions. We’ll see a lot more of each of these functional groups and their interactions later on.