BIOL 300 – Foundations of Biology
Summer 2017 – Telleen
Lecture 3

Chemistry of Life: Basic Chemistry, Bonding, and Water

I. Simple Chemistry

A. Organisms are chemical machines so it is important to understand some basic chemistry in order to grasp the full scope of biology
B. Matter – **Matter** is any substance in the universe that has mass and occupies space. It is composed of extremely small particles called **atoms**.
C. Atoms are the smallest particle into which a substance can be divided and still retain its chemical properties (such a substance is called an **element**)
    1. In Greek, atom literally means ‘indivisible’
    2. Every atom has the same basic structure:
        a. A **Nucleus** is at the center of every atom
        b. **Nuclei** are made up of subatomic particles called:
            1. **Protons** – positively charged (+)
            2. **Neutrons** – no charge/electrically neutral
            3. Protons and neutrons have approximately the same mass
        c. **Electrons** orbit the nucleus is clouds
            1. Electrons are negatively charged (-)
            2. They are also much smaller than protons and neutrons
            3. There is one electron for every proton in the nucleus to balance the electrical charge of the atom
    3. The **Atomic number** refers to the number of protons in the nucleus of a type of atom. **Elements** are atoms with the same atomic number
    4. The **Atomic mass** of an atom refers to the number of protons + neutrons in the nucleus
    5. The **Periodic Table of the Elements** is a particular way to visualize the relationships between different types of. In it, elements are arranged based on their atomic number and the valence shells of electrons (which we won’t discuss in too much detail).
    6. **Electrons determine what atoms are like** because they interact with neighboring atoms. Protons also indirectly influence behavior of atoms because they dictate how many electrons are needed to balance the electrical charge of the nucleus.
    7. **Electrons carry energy**. Electricity is electrons flowing from one atom to the next.
    8. Electrons are present in electron (or valence) shells and **can be lost or gained from the outer shell of atoms**. This results in **ions**, which are electrically charged atoms (or molecules).
    9. The number of neutrons in the nucleus of an atom can vary without changing the electrical properties of an atom. **Isotopes** are atoms with the same atomic number (# of protons) but have different numbers of neutrons (i.e. they have a different atomic mass)
    10. Most elements are a mixture of isotopes. For example, Carbon (C) always has 6 protons (because its atomic # = 6), but the atomic mass of C can be 12, 13, or 14 (i.e. 6, 7, or 8 neutrons). 99% of C is ^{12}C, but other forms do exist. ^{14}C is unstable and results in radioactive decay.

D. **Molecules** are groups of atoms held together by energy through the formation of **chemical bonds**. There are three basic types of chemical bonds.
    1. **Ionic bonds** are formed by the attraction of opposite electrical charges (+ attracts -, sort of like magnets). Sodium Chlorine (NaCl), or table salt, is a good example. NaCl forms a solid crystal as the Na and Cl ionize to form Na^+ and Cl^- ions, which then attract each other due to their opposite electrical charges. Ionic bonds are fairly strong (though not as strong as covalent bonds), but are not directional (e.g. a positive charge can attract a negative charge from any direction; they are more promiscuous than covalent bonds).
    2. **Covalent bonds** are the strongest chemical bond. They are formed when atoms share electrons.
        a. Atoms seek to fill their outermost shell of electrons (which usually can contain a maximum of either 2 or 8 electrons in most biologically relevant atoms), but they are often prevented from doing this because the need to balance their electrical charge is at odds with this. Covalent bonds offer a way to share electrons to get around this.
        b. Covalent bonds come in three varieties based on how many **pairs of electrons are shared**:
            1. Single – share 2 e^- (1 pair)
            2. Double – share 4 e^- (2 pairs)
            3. Triple – share 6 e^- (3 pairs)
        c. Water (H_2O) is a good example. Oxygen (O) has 6 e^- but wants to get 2 more to fill its outer shell with 8 electrons and Hydrogen (H) has 1 e^- but wants to fill its outer shell with 2 e^-. The solution is for each to two H to share a pair of e^- with one O. This results in the formation of an O with two single covalent bonds to two different H atoms.
        d. Carbon (C) is an extremely important biological element because it can form up to 4 separate covalent bonds (or two double bonds, or a triple bond and a single bond, and so on).
        e. Covalent bonds are very strong, but they are also directional. This means that interactions with neighboring atoms through covalent bonds is very specific.
    3. **Hydrogen bonds** are weak chemical bonds that play a key role in biology.
        a. In a covalent bond, some nuclei are better at attracting the shared electrons than others. This results is **polar molecules**, in which one part of the molecule has a slightly negative charge (because it is better at attracting electrons) while another part has a slightly positive charge (because it can't hold the electrons as well).
        b. H_2O is a perfect example of a polar molecule. The O is better at holding the shared electrons, so the O side has a partial negative charge, while the H have a partial positive charge because the electrons spend more time near the O.
        c. Hydrogen bonds occur when the positive end of one polar molecule is attracted to the negative end of another.
        d. Hydrogen bonds are of extreme importance to biological systems for several reasons:
            1. They are weak and not very effective over long distances
            2. They are highly directional
        e. They are too weak to form stable molecules alone, but can act like Velcro: many weak interactions can be very stable, such as in DNA, where the two strands are held together by hydrogen bonds (but we'll talk more about that later)

II. Water (H_2O) is of central importance to all known life

A. ¾ of the Earth's surface is covered with water; 2/3 of your body is made up of water; Life as we know it cannot exist without water, so the chemistry of life is H_2O chemistry

B. H_2O has a relatively simple structure, but it can form hydrogen bonds, which are primarily responsible for its important properties:

1. Heat storage – temperature is a measure of how rapidly individual molecules are moving. Due to H-bonds, H_2O requires lots of energy to disrupt its organization. Thus, it takes lots of energy to heat, heats slowly, and holds temperature longer than most other molecules. This is why we can maintain a relatively stable body temperature
2. Ice formation – at low temperatures H-bonds don't break easily resulting in a crystal structure. Interestingly, solid water (ice) is actually less dense than liquid water (this is why a beer can will explode if you put it in the freezer for too long!)
3. High heat of vaporization – Since an increase in temperature is required to break the H-bonds in water, the liquid to vapor transition takes lots of energy. This is why sweating cools us off (i.e. there is a net loss of energy, which we perceive as a lower temperature, as the sweat evaporates into the air).
4. **Cohesion** is the attraction between individual water molecules, due to their polarity and the H-bonds. As result, one water molecule can pull another (which can pull another, and so on). This results in surface tension and capillary action, which both play important roles in biology. Cohesion is the attraction (through H-bonds) between the same type of molecule (e.g. H_2O with H_2O), while adhesion refers to H-bonding and attraction between different types of molecules (e.g. H_2O and NH_3).
5. High Polarity – Because of its partial charges, H_2O is attracted to ions and other polar molecules. These types of molecules are called **hydrophilic** (that is, they 'like' water). Water gathers closely around ions (fully charged) and polar molecules (partially charged). This is what happens when salt dissolves in water. The ionic bonds are broken as the negative ends of water interact with the Na^+ and the positive ends interact with the Cl^-. A similar situation occurs for polar molecules. This is what we mean by the term **soluble** (i.e. can be dissolved in). **Non-polar** molecules, such as oil, do not form H-bonds, so they are not water soluble and are called **hydrophobic** (they are 'afraid' of water). Hydrophobic molecules 'attract' each other because they are repelled from H_2O (and other polar molecules); this is called hydrophobic force.

C. H_2O can also ionize (i.e. covalent bonds can break).

1. $H_2O \rightarrow OH^-$ (hydroxide ion) and H^+ (hydrogen ion)
2. This is not common since covalent bonds are strong. At 25°C, one molecule out of every 550 million is ionized. This corresponds to 1/10,000,000 (or 10^{-7}) of a mole of H^+ ions. Another way to state this is $[H^+]$ (concentration of H^+) = 1/10,000,000
3. A convenient way to express $[H^+]$ is called **pH**. The relationship can be written as: **$pH = -\log[H^+]$.** Logarithm is simply the exponent of the molar concentration. Pure H_2O has a ph of 7 (remember that –7 exponent?)

4. Since each H^+ also produces an OH^-, $[H^+] = [OH^-]$, so pH = 7 is neutral
5. **Acids** have $[H^+] > [OH^-]$, so pH < 7
6. **Bases** have $[H^+] < [OH^-]$, so pH > 7
7. Acids and bases occur when things are dissolved in water that affect the $[H^+]$
8. **Buffers** are chemicals that minimize changes in pH while they are dissolved in water by taking up or releasing H^+. A good example of this is carbonic acid/bicarbonate.