AP Biology Notes
Water Chemistry

I. Importance of Water to Life

A. Why is water important?

1. Life probably evolved in water.
2. Living cells are 70% - 95% water.
3. Water covers about 3/4 of the earth.
4. In nature, water naturally exists in all three physical states of matter -- solid, liquid, and gas.
5. Life absolutely depends on the properties of water.
6. Water's extraordinary properties are emergent properties resulting from water's structure and molecular interactions.

II. Water and H-Bonding

A. How does water form hydrogen bonds?

1. Water is a polar molecule. Its polar bonds and asymmetrical shape give water molecules opposite charges on opposite sides.
   a. Four valence orbitals of O point to corners of a tetrahedron.
   b. 2 corners are orbitals with unshared pairs of electrons and weak negative charge.
   c. 2 corners are occupied by H atoms which are in polar covalent bonds with O. Oxygen is so electronegative, that shared electrons spend more time around the O causing a weak positive charge near H's.
2. The polar molecules of water are held together by hydrogen bonds.
   a. Positively charged H of one molecule is attracted to the negatively charged O of another water molecule.
   b. Each water molecule can form a maximum of four hydrogen bonds with neighboring water molecules.
   c. Hydrogen bonding orders water into a higher level of structural organization.

III. Properties of Water

A. Some extraordinary properties of water

1. The Cohesiveness of Liquid Water
   a. **Cohesion** = Phenomenon of a substance being held together by hydrogen bonds.
      i. Though hydrogen bonds are transient, enough water molecules are hydrogen bonded at any given time to give water more structure than other liquids.
      ii. Contributes to upward water transport in plants.
   b. **Surface tension** = Measure of how difficult it is to stretch or break the surface of a liquid.
      i. Water has a greater surface tension than most liquids.
      ii. Function of the fact that at the air/water interface, surface water molecules are hydrogen bonded to each other and to the water molecules below.
      iii. Causes water to bead (shape with smallest area to volume ratio and allows maximum hydrogen bonding).
   c. **Adhesion** = Clinging of water to hydrophilic substances (e.g. glass).
   d. **Hydrophilic** = (Hydro=water; philo=loving) Property of having a affinity for water.
   e. **Imbibition** = Process of water soaking into a porous hydrophilic substance (e.g. a sponge).
      i. Imbibition by seeds ruptures the seed coat and allows germination.
2. Water's High Specific Heat
   a. Heat and Temperature
      i. **Kinetic energy** = The energy of motion.
      ii. **Heat** = Total kinetic energy due to molecular motion in a body of matter.
      iii. **Calorie** (cal) = Amount of heat it takes to raise the temperature of one gram of

water by one degree Celsius.

iv. **Kilocalorie** (kcal or Cal) = Amount of heat required to raise the temperature of one kilogram of water by one degree Celsius (1000 cal).

v. **Temperature** = Measure of heat intensity due to the average kinetic energy of molecules in the body of matter.

vi. The Celsius Thermometer: The Celsius scale at sea level:

1. 100 degrees C (212 degrees F) = water boils
2. 37 degrees C (98.6 degrees F) = human body temperature
3. 23 degrees C (72 degrees F) = room temperature
4. 0 degrees C (32 degrees F) = water freezes

vii. To convert temperature scales:

1. degrees C = 5 (degrees F - 32) / 9
2. degrees F = 9 degrees C / 5 + 32
3. degrees K = degrees C + 273

viii. **Specific heat** = Amount of heat that must be absorbed or lost for one gram of a substance to change its temperature by one degree Celsius.

1. Specific heat of water = One calorie per gram per degree Celsius (1 cal/g/degree C).

b. How Water Stabilizes Temperature

i. Water has a high specific heat, which means that it resists temperature changes when it absorbs or releases heat.

1. As a result of hydrogen bonding among water molecules, it takes a relatively large heat loss or gain for each 1 degree C change in temperature.
2. Hydrogen bonds must absorb heat to break, and release heat when they form.
3. Much absorbed heat energy is used to disrupt hydrogen bonds before water molecules can move faster (increase temperature).

c. A large body of water can act as a heat sink -- absorbing heat from sunlight during the day and summer and releasing heat during the night and winter as the water gradually cools. As a result:

i. Water, which covers three-fourths of the planet, keeps temperature fluctuations within a range suitable for life.

ii. Coastal areas have milder climates than inland.

iii. The marine environment has a relatively stable temperature.

3. Water's High Heat of Vaporization

a. **Vaporization** (evaporation) = Transformation from liquid to gas.

i. Molecules with enough kinetic energy to overcome the mutual attraction of molecules in a liquid, can escape into the air.

b. **Heat of vaporization** = Quantity of heat a liquid must absorb for 1 g to be converted to the gaseous state.

i. For water molecules to evaporate, hydrogen bonds must be broken which requires heat energy.

ii. Water has a relatively high heat of vaporization (540 cal/g).

c. **Evaporative cooling** = Cooling of a liquid's surface when a liquid evaporates.

i. The surface molecules with the highest kinetic energy are most likely to escape into gaseous form; the average kinetic energy of the remaining surface molecules is thus lower.

d. Benefits of water's high heat of vaporization:

i. Moderates the earth's climate.

1. Solar heat absorbed by tropical seas dissipates when surface water evaporates (evaporative cooling).
2. As most tropical air moves poleward, water vapor releases heat as it condenses into rain.

ii. Stabilizes temperature in aquatic ecosystems (evaporative cooling).

iii. Helps organisms from overheating by evaporative cooling.

4. Freezing and Expansion of Water

a. Water is densest at 4 degrees C.

i. As water cools from 4 degrees C to freezing (0 degrees C), it expands and becomes

less dense than liquid water (ice floats).

ii. When water begins to freeze, the molecules do not have enough kinetic energy to break hydrogen bonds.

iii. As the crystalline lattice forms, each water molecule forms a maximum of 4 hydrogen bonds, which keeps water molecules further apart than they would be in the liquid state.

b. Expansion of water contributes to the fitness of the environment for life:

i. Prevents deep bodies of water from freezing solid from the bottom up.

ii. Since ice is less dense, it forms on the surface first. As water freezes it releases heat to the water below and insulates it.

iii. Makes the transitions between seasons less abrupt. As water freezes, hydrogen bonds form releasing heat. As ice melts, hydrogen bonds break absorbing heat.

5. Water as a Versatile Solvent

a. **Solution** = A liquid that is a homogenous mixture of two or more substances.

b. **Solvent** = Dissolving agent of a solution.

c. **Solute** = Substance dissolved in a solution.

d. **Aqueous solution** = Solution in which water is the solvent.

e. Water is a versatile solvent owing to the polarity of the water molecule.

i. Ionic compounds dissolve in water.

1. Charged regions of polar water molecules have an electrical attraction to charged ions.

2. Water surrounds individual ions, separating and shielding them from one another.

ii. Polar compounds, in general, are water-soluble.

1. Charged regions of polar water molecules have an affinity for opposite charged regions of other polar molecules.

iii. Nonpolar compounds (which have symmetric distribution in charge) are NOT water-soluble.

IV. Aqueous Solutions

A. Aqueous solutions

1. Most biochemical reactions involve solutes dissolved in water.

2. Two important properties of aqueous solutions:

a. Solute concentration.

b. pH.

3. Solute Concentration

a. **Molecular weight** = Sum of the weight of all atoms in a molecule (expressed in daltons).

b. **Mole** = Amount of a substance that has a mass in grams numerically equivalent to its molecular weight in daltons.

c. **Molarity** = Number of moles of solute per liter of solution.

d. Advantage to measuring in moles:

i. Rescales weighing of single molecules in daltons in grams, which is more practical for laboratory use.

ii. A mole of one substance has the same number of molecules as a mole of any other substance (6.02×10^{23}, Avogadro's number).

iii. Allows one to combine substances in fixed ratios of molecules.

4. Acids, Bases and pH

a. Dissociation of Water Molecules:

i. Occasionally, the hydrogen atom that is shared in a hydrogen bond between two water molecules shifts from the oxygen atom to which it is covalently bonded to the unshared orbitals of the oxygen atom to which it is hydrogen bonded.

1. Only a hydrogen ion (proton with a +1 charge) is actually transferred.

2. Transferred proton binds to an unshared orbital of the second water molecule creating a hydronium ion ($H3O^{+}$).

3. Water molecule that lost a proton has a net negative charge and is called a hydroxide ion (OH^{-}).

4. By convention, ionization of water is expressed as the dissociation into H^+ and OH^-.
5. Reaction is reversible.
6. At equilibrium, most of the water is not ionized.

b. Acids and Bases--At equilibrium in pure water:
   i. Number of H+ ions = number of OH- ions.
   ii. Only 1 out of 554,000,000 molecules of water actually dissociates.
      1. **Acid**
         a. Substance that increases the relative [H+] of a solution.
         b. Also removes OH- because it tends to combine with H+ to form water.
      2. **Base**
         a. Substance that reduces that relative $[H^+]$ of a solution.
         b. May alternately increase $[OH^-]$.
            For example:
            i. A base may reduce $[H^+]$ directly; $NH_3 + H+ \rightarrow NH_4$
            ii. A base may reduce $[H^+]$ indirectly; NaOH --> Na+ + OH^- then
               $OH^- + H^+ \rightarrow H2O$
      3. A solution in which:
         a. $[H^+] = [OH^-]$ = neutral solution.
         b. $[H^+]$ greater than $[OH^-]$ = acidic solution.
         c. $[H^+]$ less than $[OH^-]$ = basic solution.
   iii. The pH Scale: In any aqueous solution: $[H^+][OH^-] = 10^{-14}$ Molar
      1. In a neutral solution, $[H^+] = 10^{-7}$ M and $[OH^-] = 10^{-7}$ M.
      2. In an acid solution where the $[H^+] = 10^{-5}$ M and $[OH^-] = 10^{-9}$ M.
      3. In a basic solution where the $[H+] = 10^{-9}$ M and $[OH^-] = 10^{-5}$ M.
   iv. pH scale = Scale used to measure degree of acidity. It ranges from 0 to 14.
      1. pH = Negative log10 of the $[H^+]$ expressed in moles per liter.
      2. pH of 7 is a neutral solution.
      3. pH < 7 is an acidic solution.
      4. pH > 7 is a basic solution.
      5. Most biological fluids were within the pH range of 6 to 8. There are some exceptions such as stomach acid with pH = 1.5.
      6. Each pH unit represents a tenfold difference (scale is logarithmic), so a slight change in pH represents a large change in actual $[H^+]$.

5. Buffers
   a. Organisms must maintain pH of body fluids within a narrow range (usually pH 6-8).
   b. **Buffer** = Substance that prevents large sudden changes in pH.
      i. Are combinations of H^+-donor and H^+-acceptor forms of weak acids or bases.
   c. Buffers work by:
      i. Accepting H^+ ions from solution when they are in excess.
      ii. Donating H^+ ions to the solution when they have been depleted.
      iii. For example: Bicarbonate buffer.
         1. Bicarbonate (HCO_3^-) can take up excess H^+, forming carbonic acid (H_2CO_3). The carbonic acid is then converted to CO_2 (carbon dioxide) and water.
         2. Bicarbonate ion can release H^+, forming carbonate ion (CO_3^-) and water.