CHAPTER

WATER AND THE FITNES OF THE ENVIRONMEN

OUTLINE

I. Water's Polarity and Its Effects
   A. The polarity of water molecules results in hydrogen bonding
   B. Organisms depend on the cohesion of water molecules
   C. Water moderates temperatures on Earth
   D. Oceans and lakes don't freeze solid because ice floats
   E. Water is the solvent of life
II. The Dissociation of Water
   A. Organisms are sensitive to changes in pH
III. Acid Precipitation Threatens the Fitness of the Environment

OBJECTIVES

After reading this chapter and attending lecture, the student should be able to:

1. Describe how water contributes to the fitness of the environment to support life.
2. Describe the structure and geometry of a water molecule, and explain what proper emerge as a result of this structure.
3. Explain the relationship between the polar nature of water and its ability to fc hydrogen bonds.
4. List five characteristics of water that are emergent properties resulting from hydro bonding.
5. Describe the biological significance of the cohesiveness of water.
6. Distinguish between heat and temperature.
7. Explain how water's high specific heat, high heat of vaporization and expansion u freezing affect both aquatic and terrestrial ecosystems.
8. Explain how the polarity of the water molecule makes it a versatile solvent.
9. Define molarity and list some advantages of measuring substances in moles.
10. Write the equation for the dissociation of water, and explain what is actu transferred from one molecule to another.
11. Explain the basis for the pH scale.
12. Explain how acids and bases directly or indirectly affect the hydrogen concentration of a solution.
13. Using the bicarbonate buffer system as an example, explain how buffers work.

14. Describe the causes of acid precipitation, and explain how it adversely affects the fitness of the environment.

KEY TERMS

polar molecule	Celsius scale	solute	hydrogen ion
cohesion	calorie	solvent	molarity
adhesion	kilocalorie	aqueous solution	hydroxide ion
surface tension	joule	hydrophilic	acid
kinetic energy	specific heat	hydrophobic	base
heat	evaporative cooling	mole	pH scale
temperature	solution	molecular weight	buffer
acid precipitation			

LECTURE NOTES

Water contributes to the fitness of the environment to support life.

- Life on earth probably evolved in water.
- Living cells are 70%-95% H_2O.
- Water covers about 3/4 of the earth.
- In nature, water naturally exists in all three physical states of matter—solid, liquid and gas.

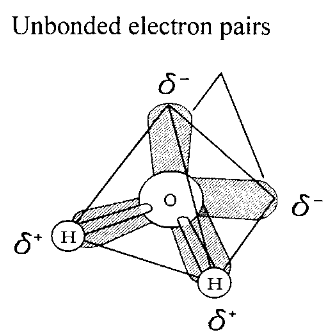
Water's extraordinary properties are emergent properties resulting from water's structure and molecular interactions.

I. Water's Polarity and Its Effects

A. The polarity of water molecules results in hydrogen bonding

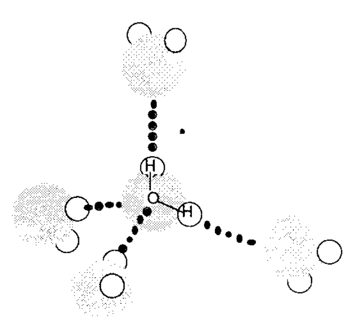
Water is a *polar* molecule. Its polar bonds and asymmetrical shape give water molecules opposite charges on opposite sides.

- Four valence orbitals of O point to corners of a tetrahedron.
- Two corners are orbitals with unshared pairs of electrons and weak negative charge.
- Two 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.



Hydrogen bonding orders water into a higher level of structural organization.

- The polar molecules of water are held together by hydrogen bonds.
- Positively charged H of one molecule is attracted to the negatively charged O of another water molecule.
- Each water molecule can form a maximum of four hydrogen bonds with neighboring water molecules.



Water has extraordinary properties that emerge as a consequence of its polarity and hydrogen-bonding. Some of these properties are that water:

- has cohesive behavior
- resists changes in temperature
- has a high heat of vaporization and cools surfaces as it evaporates
- expands when it freezes
- is a versatile solvent

B. Organisms depend on the cohesion of water molecules.

Cohesion = Phenomenon of a substance being held together by hydrogen bonds.

- Though hydrogen bonds are transient, enough water molecules are hydrogen bonded at any given time to give water more structure than other liquids.
- Contributes to upward water transport in plants by holding the water column together. *Adhesion* of water to vessel walls counteracts the downward pull of gravity.

Surface tension = Measure of how difficult it is to stretch or break the surface of a liquid.

- Water has a greater surface tension than most liquids; function of the fact that at the air/H_2O interface, surface water molecules are hydrogen bonded to each other and to the water molecules below.
- Causes H_2O to bead (shape with smallest area to volume ratio and allows maximum hydrogen bonding).

C. Water moderates temperatures on Earth

1. Heat and temperature

Kinetic energy = The energy of motion.

Heat = Total kinetic energy due to molecular motion in a body of matter.

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

Calorie (cal) = Amount of heat it takes to raise the temperature of one gram of water by one degree Celsius. Conversely, one calorie is the amount of heat that one gram of water releases when it cools down by one degree Celsius. NOTE: The "calories" on food packages are actually kilocalories (kcal).

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

Celsius Scale at Sea Level		Scale Conversion		
100°C (212°F)	= water boils	°C	=	$\frac{5(°F - 32)}{9}$
37°C (98.6°F)	= human body temperature	°F	=	$\frac{9°C}{5} + 32$
23°C (72°F)	= room temperature	°K	=	°C + 273
0°C (32°F)	= water freezes			

2. Water's high specific heat

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

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.

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

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

A large body of water can act as a heat sink, absorbing heat from sunlight during the day and summer (while warming only a few degrees) and releasing heat during the night and winter as the water gradually cools. As a result:

- Water, which covers three-fourths of the planet, keeps temperature fluctuations within a range suitable for life.
- Coastal areas have milder climates than inland.
- The marine environment has a relatively stable temperature.

3. Evaporative cooling

Vaporization (evaporation) = transformation from liquid to a gas.

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

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

- For water molecules to evaporate, hydrogen bonds must be broken which requires heat energy.
- Water has a relatively high heat of vaporization at the boiling point (540 cal/g or 2260 J/g; Joule = 0.239 cal).

Evaporative cooling = Cooling of a liquid's surface when a liquid evaporates (see Campbell, Figure 3.4).

- 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.

Water's high heat of vaporization:

- Moderates the Earth's climate.
  - Solar heat absorbed by tropical seas dissipates when surface water evaporates (evaporative cooling).
  - As moist tropical air moves poleward, water vapor releases heat as it condenses into rain.
- Stabilizes temperature in aquatic ecosystems (evaporative cooling).
- Helps organisms from overheating by *evaporative cooling.*

D. Oceans and lakes don't freeze solid because ice floats

Because of hydrogen bonding, water is less dense as a solid than it is as a liquid. Consequently, ice floats.

- Water is densest at 4°C.
- Water contracts as it cools to 4°C.
- As water cools from 4°C to freezing (0°C), it expands and becomes *less dense* than liquid water (ice floats).
- When water begins to freeze, the molecules do not have enough kinetic energy to break hydrogen bonds.
- As the crystalline lattice forms, each water molecule forms a maximum of four hydrogen bonds, which keeps water molecules further apart than they would be in the liquid state; see Campbell, Figure 3.5.

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

- Prevents deep bodies of water from freezing solid from the bottom up.
- Since ice is less dense, it forms on the surface first. As water freezes it releases heat to the water below and insulates it.
- Makes the transitions between seasons less abrupt. As water freezes, hydrogen bonds form releasing heat. As ice melts, hydrogen bonds break absorbing heat.

E. Water is the solvent of life

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

Solvent = Dissolving agent of a solution.

Solute = Substance dissolved in a solution.

Aqueous solution = Solution in which water is the solvent.

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

Hydrophilic

Ionic compounds dissolve in water (see Campbell, Figure 3.8).

- Charged regions of polar water molecules have an electrical attraction to charged ions.
- Water surrounds individual ions, separating and shielding them from one another.

Polar compounds in general, are water-soluble.

- Charged regions of polar water molecules have an affinity for oppositely charged regions of other polar molecules.

Hydrophobic

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

1. Hydrophilic and hydrophobic substances

Ionic and polar substances are *hydrophilic*, but nonpolar compounds are hydrophobic.

Hydrophilic = (Hydro = water; philo = loving); property of having an affinity for water.

- Some large hydrophilic molecules can absorb water without dissolving.

Hydrophobic = (Hydro = water; phobos = fearing); property of not having an affinity for water, and thus, not being water-soluble.

2. Solute concentration in aqueous solutions

Most biochemical reactions involve solutes dissolved in water. There are two important quantitative properties of aqueous solutions: solute concentration and pH.

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

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

For example, to determine a mole of sucrose ($C_{12}H_{22}O_{11}$):

- Calculate molecular weight:

C = 12 dal	12 dal × 12 =	144 dal
H = 1 dal	1 dal × 22 =	22 dal
O = 16 dal	16 dal × 11 =	176 dal
		342 dal

- Express it in grams (342 g).

Molarity = Number of moles of solute per liter of solution

- To make a 1M sucrose solution, weigh out 342 g of sucrose and add water up to 1L.

Advantage of measuring in moles:

- Rescales weighing of single molecules in daltons to grams, which is more practical for laboratory use.
- 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).
- Allows one to combine substances in fixed ratios of molecules.

II. The Dissociation of Water

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.

- Only a *hydrogen ion* (proton with a +1 charge) is actually transferred.
- Transferred proton binds to an unshared orbital of the second water molecule creating a *hydronium ion* (H_3O^+).
- Water molecule that lost a proton has a net negative charge and is called a *hydroxide ion* (OH^-).

$$H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-$$

  - By convention, ionization of H_2O is expressed as the *dissociation* into H^+ and OH^-.

$$H_2O \rightleftharpoons H^+ + OH^-$$

  - Reaction is reversible.
- At equilibrium, most of the H_2O is *not* ionized.

A. Organisms are sensitive to changes in pH

1. Acids and bases

At equilibrium in pure water at 25°C:

- Number of H+ ions = number of OH- ions.
- $[H^+] = [OH^-] = \dfrac{1}{10{,}000{,}000}\ M = 10^{-7}\ M$
- Note that brackets indicate molar concentration.

This is a good place to point out how *few* water molecules are actually dissociated (only 1 out of 554,000,000 molecules).

ACID	BASE
Substance that *increases* the relative $[H^+]$ of a solution. Also removes OH^- because it tends to combine with H^+ to form H_2O. For example: (in water) $HCl \rightarrow H^+ + Cl^-$	Substance that *reduces* the relative $[H^+]$ of a solution. May alternately increase $[OH^-]$. For example: A base may reduce $[H^+]$ directly: $NH_3 + H^+ \rightleftharpoons NH_4^+$ A base may reduce $[H^+]$ indirectly: $NaOH \rightarrow Na^+ + OH^-$ $OH^- + H^+ \rightarrow H_2O$

A solution in which:

- $[H^+] = [OH^-]$ is a neutral solution.
- $[H^+] > [OH^-]$ is an acidic solution.
- $[H^+] < [OH^-]$ is a basic solution.

Strong acids and bases dissociate completely in water.

- Example: HCl and NaOH
- Single arrows indicate complete dissociation.
  $NaOH \rightarrow Na^+ + OH^-$

Weak acids and bases dissociate only partially and reversibly.

- Examples: NH_3 (ammonia) and H_2CO_3 (carbonic acid)
- Double arrows indicate a reversible reaction; at equilibrium there will be a fixed ratio of reactants and products.

$$\underset{\text{Carbonic acid}}{H_2CO_3} \rightleftharpoons \underset{\text{Bicarbonate ion}}{HCO_3^-} + \underset{\text{Hydrogen ion}}{H^+}$$

2. The pH scale

In any aqueous solution:

$[H^+][OH^-] = 1.0 \times 10^{-14}$

For example:

- In a neutral solution, $[H^+] = 10^{-7}$ M and $[OH^-] = 10^{-7}$ M.
- In an acidic solution where the $[H^+] = 10^{-5}$ M, the $[OH^-] = 10^{-9}$ M.
- In a basic solution where the $[H^+] = 10^{-9}$ M, the $[OH^-] = 10^{-5}$ M.

pH scale = Scale used to measure degree of acidity. It ranges from 0 to 14.

pH = Negative $\log_{10}$ of the $[H^+]$ expressed in moles per liter.

- pH of 7 is a neutral solution.
- pH < 7 is an acidic solution.
- pH > 7 is a basic solution.

- Most biological fluids are within the pH range of 6 to 8. There are some exceptions such as stomach acid with pH = 1.5. (See Campbell, Figure 3.9)
- Each pH unit represents a *tenfold* difference (scale is logarithmic), so a slight change in pH represents a large change in actual $[H^+]$.

To illustrate this point, project the following questions on a transparency and cover the answer. The students will frequently give the wrong response (3×), and they are surprised when you unveil the solution.
How much greater is the $[H^+]$ in a solution with pH 2 than in a solution with pH 6?

ANS:

$$\text{pH } 2 = [H^+] \text{ of } 1.0 \times 10^{-2} = \frac{1}{100}\ \text{M}$$

$$\text{pH } 6 = [H^+] \text{ of } 1.0 \times 10^{-6} = \frac{1}{1{,}000{,}000}\ \text{M}$$

<u>10,000</u> times greater.

3. Buffers

By minimizing wide fluctuations in pH, buffers help organisms maintain the pH of body fluids within the narrow range necessary for life (usually pH 6-8).

Buffer = Substance that minimizes large sudden changes in pH.

- Are combinations of H^+-donor and H^+-acceptor forms in a solution of weak acids or bases
- Work by accepting H^+ ions from solution when they are in excess and by donating H^+ ions to the solution when they have been depleted

  Example: Bicarbonate buffer

$$\underset{\substack{H^+ \text{ donor} \\ \text{(weak acid)}}}{H_2CO_3} \quad \underset{\text{response to a drop in pH}}{\overset{\text{response to a rise in pH}}{\rightleftharpoons}} \quad \underset{\substack{H^+ \text{ acceptor} \\ \text{(weak base)}}}{HCO_3^-} \quad + \quad \underset{\substack{\text{Hydrogen} \\ \text{ion}}}{H^+}$$

$$\underset{\substack{\text{strong} \\ \text{acid}}}{HCl} + NaHCO_3 \longrightarrow \underset{\substack{\text{weak} \\ \text{acid}}}{H_2CO_3} + NaCl$$

$$\underset{\substack{\text{strong} \\ \text{base}}}{NaOH} + H_2CO_3 \longrightarrow \underset{\substack{\text{weak} \\ \text{base}}}{NaHCO_3} + H_2O$$

III. Acid Precipitation Threatens the Fitness of the Environment

Acid precipitation = Rain, snow, or fog more strongly acidic than pH 5.6.

- Has been recorded as low as pH 1.5 in West Virginia
- Occurs when sulfur oxides and nitrogen oxides in the atmosphere react with water in the air to form acids which fall to Earth in precipitation
- Major oxide source is the combustion of fossil fuels by industry and cars
- Acid rain affects the fitness of the environment to support life.
  - Lowers soil pH which affects mineral solubility. May leach out necessary mineral nutrients and increase the concentration of minerals that are potentially toxic to vegetation in higher concentration (e.g., aluminum). This is contributing to the decline of some European and North American forests.

- Lowers the pH of lakes and ponds, and runoff carries leached out soil minerals into aquatic ecosystems. This adversely affects aquatic life. Example: In the Western Adirondack Mountains, there are lakes with a pH < 5 that have no fish.

What can be done to reduce the problem?

- Add industrial pollution controls.
- Develop and use antipollution devices.
- Increase involvement of voters, consumers, politicians, and business leaders.

> The political issues surrounding acid rain can be used to enhance student awareness and make this entire topic more relevant and interesting to the students.

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