Chapter 3

Water and the Fitness of the Environment

Edited by Shawn Lester

PowerPoint® Lecture Presentations for

Biology

Eighth Edition

Neil Campbell and Jane Reece

Lectures by Chris Romero, updated by Erin Barley with contributions from Joan Sharp
Learning objectives:

1. List and explain the four properties of water that emerge as a result of its ability to form hydrogen bonds

2. Distinguish between the following sets of terms: hydrophobic and hydrophilic substances; a solute, a solvent, and a solution

3. Define acid, base, and pH

4. Explain how buffers work
Overview: The Molecule That Supports All of Life

• Water is the biological medium on Earth

• All living organisms require water more than any other substance

• Most cells are surrounded by water, and cells themselves are about 70–95% water

• The abundance of water is the main reason the Earth is habitable
Concept 3.1: The polarity of water molecules results in hydrogen bonding

- The water molecule is a **polar molecule**: The opposite ends have opposite charges
- Polarity allows water molecules to form hydrogen bonds with each other
Fig. 3-2

Hydrogen bond

\[ \delta^- \ldots \delta^+ \]

\[ \delta^+ \ldots \delta^- \]

\[ \delta^+ \ldots \delta^- \]

\[ \delta^- \ldots \delta^+ \]
Concept 3.2: Four emergent properties of water contribute to Earth’s fitness for life

Four of water’s properties that facilitate an environment for life are:

- Cohesive behavior
- Ability to moderate temperature
- Expansion upon freezing
- Versatility as a solvent
Cohesion

• Collectively, hydrogen bonds hold water molecules together, a phenomenon called **cohesion**

• Cohesion helps the transport of water against gravity in plants (capillary action)

• **Adhesion** is an attraction between different substances, for example, between water and plant cell walls
Fig. 3-3

Water-conducting cells

Direction of water movement

150 µm

Adhesion

Cohesion
• **Surface tension** is a measure of how hard it is to break the surface of a liquid

• Surface tension is related to cohesion

• **Example:** Water Striders/Skippers
Moderation of Temperature

- Water absorbs heat from warmer air and releases stored heat to cooler air.
- Water can absorb or release a large amount of heat with only a slight change in its own temperature.
Heat and Temperature

- **Kinetic energy** is the energy of motion
- **Heat** is a measure of the *total* amount of kinetic energy due to molecular motion
- **Temperature** measures the intensity of heat due to the *average* kinetic energy of molecules
• The **Celsius scale** is a measure of temperature using Celsius degrees (°C)

• A **calorie (cal)** is the amount of heat required to raise the temperature of 1 g of water by 1°C

• The “calories” on food packages are actually **kilocalories (kcal)**, where 1 kcal = 1,000 cal

• The **joule (J)** is another unit of energy where 1 J = 0.239 cal, or 1 cal = 4.184 J
Water’s High Specific Heat

- The **specific heat** of a substance is the amount of heat that must be absorbed or lost for 1 g of that substance to change its temperature by 1°C

- The specific heat of water is 1 cal/g/°C

- Water resists changing its temperature because of its high specific heat
• Water’s high specific heat can be traced to hydrogen bonding
  – Heat is absorbed when hydrogen bonds break
  – Heat is released when hydrogen bonds form

• The high specific heat of water minimizes temperature fluctuations to within limits that permit life
Some common specific heats and heat capacities:

<table>
<thead>
<tr>
<th>Substance</th>
<th>S (J/g °C)</th>
<th>C (J/°C) for 100 g</th>
</tr>
</thead>
<tbody>
<tr>
<td>Air</td>
<td>1.01</td>
<td>101</td>
</tr>
<tr>
<td>Aluminum</td>
<td>0.902</td>
<td>90.2</td>
</tr>
<tr>
<td>Copper</td>
<td>0.385</td>
<td>38.5</td>
</tr>
<tr>
<td>Gold</td>
<td>0.129</td>
<td>12.9</td>
</tr>
<tr>
<td>Iron</td>
<td>0.450</td>
<td>45.0</td>
</tr>
<tr>
<td>Mercury</td>
<td>0.140</td>
<td>14.0</td>
</tr>
<tr>
<td>NaCl</td>
<td>0.864</td>
<td>86.4</td>
</tr>
<tr>
<td>Ice</td>
<td>2.03</td>
<td>203</td>
</tr>
<tr>
<td>Water</td>
<td>4.179</td>
<td>417.9</td>
</tr>
</tbody>
</table>

**Heat capacity**, ratio of heat absorbed by a material relative to the temperature change of that material.
Evaporative Cooling

- *Evaporation* is transformation of a substance from liquid to gas.

- **Heat of vaporization** is the heat a liquid must absorb for 1 g to be converted to gas.

- As a liquid evaporates, its remaining surface cools, a process called **evaporative cooling**.

- Evaporative cooling of water helps stabilize temperatures in organisms and bodies of water.
Insulation of Bodies of Water by Floating Ice

- Ice floats in liquid water because hydrogen bonds in ice are more “ordered,” making ice less dense.

- Water reaches its greatest density at 4°C.

- If ice sank, all bodies of water would eventually freeze solid, making life impossible on Earth. 
  - Less dense floating ice insulates water underneath keeping it liquid.
Hydrogen bonds are stable

Liquid water
Hydrogen bonds break and re-form
The Solvent of Life

- A **solution** is a liquid that is a homogeneous mixture of substances
- A **solvent** is the dissolving agent of a solution
- The **solute** is the substance that is dissolved
- An **aqueous solution** is one in which water is the solvent
• Water is a versatile solvent due to its polarity, which allows it to form hydrogen bonds easily.

• When an ionic compound is dissolved in water, each ion is surrounded by a sphere of water molecules called a hydration shell.
• Water can also dissolve compounds made of nonionic polar molecules

• Even large polar molecules such as proteins can dissolve in water if they have ionic and polar regions
(a) Lysozyme molecule in a nonaqueous environment

(b) Lysozyme molecule (purple) in an aqueous environment

(c) Ionic and polar regions on the protein's surface attract water molecules.
Hydrophilic and Hydrophobic Substances

• A **hydrophilic** substance is one that has an affinity for water

• A **hydrophobic** substance is one that does not have an affinity for water

• Oil molecules are hydrophobic because they have relatively nonpolar bonds

• A **colloid** is a stable suspension of fine particles in a liquid
Solute Concentration in Aqueous Solutions

- Most biochemical reactions occur in water
- Chemical reactions depend on collisions of molecules and therefore on the concentration of solutes in an aqueous solution
• **Molecular mass** is the sum of all masses of all atoms in a molecule

• Numbers of molecules are usually measured in moles, where 1 mole (mol) = $6.02 \times 10^{23}$ molecules

• Avogadro’s number and the unit *dalton* were defined such that $6.02 \times 10^{23}$ daltons = 1 g

• (a dalton is equivalent to atomic mass unit, aka molecular weight, or g/mol)

• **Molarity** ($M$) is the number of moles of solute per liter of solution
To make a 1 \( M \) solution of calcium chloride (\( \text{CaCl}_2 \)), you would place how many gm of \( \text{CaCl}_2 \) into a container and then add how much pure water?

\[
\text{mass of a Ca atom} = 40; \quad \text{mass of a Cl atom} = 35; \quad \text{mass of an O atom} = 16; \quad \text{mass of an H atom} = 1
\]

A. 75 gm of \( \text{CaCl}_2 \) then add 1 liter of water
B. 110 gm of \( \text{CaCl}_2 \) then add 1 liter of water
C. 128 gm of \( \text{CaCl}_2 \) then add 1 liter of water
D. 75 gm of \( \text{CaCl}_2 \) then add water to make a total volume of 1 liter
E. 110 gm of \( \text{CaCl}_2 \) then add water to make a total volume of 1 liter
Concept 3.3: Acidic and basic conditions affect living organisms

- A hydrogen atom in a hydrogen bond between two water molecules can shift from one to the other:
  - The hydrogen atom leaves its electron behind and is transferred as a proton, or **hydrogen ion** ($\text{H}^+$)
  - The molecule with the extra proton is now a **hydronium ion** ($\text{H}_3\text{O}^+$), though it is often represented as $\text{H}^+$
  - The molecule that lost the proton is now a **hydroxide ion** ($\text{OH}^-$)
Water is in a state of dynamic equilibrium in which water molecules dissociate at the same rate at which they are being reformed.
Hydronium ion (H$_3$O$^+$)  
Hydroxide ion (OH$^-$)
• Though statistically rare, the dissociation of water molecules has a great effect on organisms

• Changes in concentrations of H\(^+\) and OH\(^-\) can drastically affect the chemistry of a cell
Effects of Changes in pH

- Concentrations of $\text{H}^+$ and $\text{OH}^-$ are equal in pure water.
- Adding certain solutes, called acids and bases, modifies the concentrations of $\text{H}^+$ and $\text{OH}^-$.
- Biologists use something called the pH scale to describe whether a solution is acidic or basic (the opposite of acidic).
Acids and Bases

- An **acid** is any substance that increases the $\text{H}^+$ concentration of a solution.

- A **base** is any substance that reduces the $\text{H}^+$ concentration of a solution.
The pH Scale

- In any aqueous solution at 25°C the product of $H^+$ and $OH^-$ is constant and can be written as $[H^+][OH^-] = 10^{-14}$

- The **pH** of a solution is defined by the negative logarithm of $H^+$ concentration, written as $pH = -\log [H^+]$

- For a neutral aqueous solution $[H^+]$ is $10^{-7} = -(-7) = 7$
• Acidic solutions have pH values less than 7
• Basic solutions have pH values greater than 7
• Most biological fluids have pH values in the range of 6 to 8
Fig. 3-9

Acidic solution

Neutral solution

Basic solution

pH Scale

0

1
Battery acid

2
Gastric juice, lemon juice

3
Vinegar, beer, wine, cola

4
Tomato juice

5
Black coffee

6
Rainwater

7
Urine

8
Pure water

9
Human blood, tears

10
Seawater

11
Milk of magnesia

12
Household ammonia

13
Household bleach

14
Oven cleaner

Neutral 
\[ [H^+] = [OH^-] \]

Increasingly Acidic 
\[ [H^+] > [OH^-] \]

Increasingly Basic 
\[ [H^+] < [OH^-] \]
Bases donate OH\(^{-}\) or accept H\(^{+}\) in aqueous solutions

Acids donate H\(^{+}\) in aqueous solutions

- Acidic: \([H^{+}] > [OH^{-}]\)
- Neutral: \([H^{+}] = [OH^{-}]\)
- Basic: \([H^{+}] < [OH^{-}]\)
Buffers

• The internal pH of most living cells must remain close to pH 7

• **Buffers** are substances that minimize changes in concentrations of $\text{H}^+$ and $\text{OH}^-$ in a solution

• Most buffers consist of an acid-base pair that reversibly combines with $\text{H}^+$

• Our cells and tissues use bicarbonate and phosphate to maintain proper pH
In humans, blood pH is around 7.4, and a decrease in blood pH to 6.4 would be fatal. A drop by 1 pH unit represents which of these?

A. 1/10 as many H\(^+\) ions in the solution
B. 1/7 as many H\(^+\) ions in the solution
C. 1/2 as many H\(^+\) ions in the solution
D. twice as many H\(^+\) ions in the solution
E. ten times as many H\(^+\) ions in the solution
The chemical equilibrium between carbonic acid and bicarbonate acts as a pH regulator in our blood. As the blood pH begins to rise, what will happen?

\[
\begin{align*}
\text{H}_2\text{CO}_3 & \quad \leftrightarrow \quad \text{HCO}_3^- \quad + \quad \text{H}^+ \\
\text{Carbonic acid} & \quad \text{Bicarbonate ion} \quad \text{Hydrogen ion}
\end{align*}
\]

A. reaction proceeds to the right; more carbonic acid dissociates

B. reaction proceeds to the right; more carbonic acid forms

C. reaction proceeds to the left; more carbonic acid dissociates

D. reaction proceeds to the left; more carbonic acid forms