Chapter 3

Water and the Fitness of the Environment

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PowerPoint® Lecture Presentations for

Biology

Eighth Edition

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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

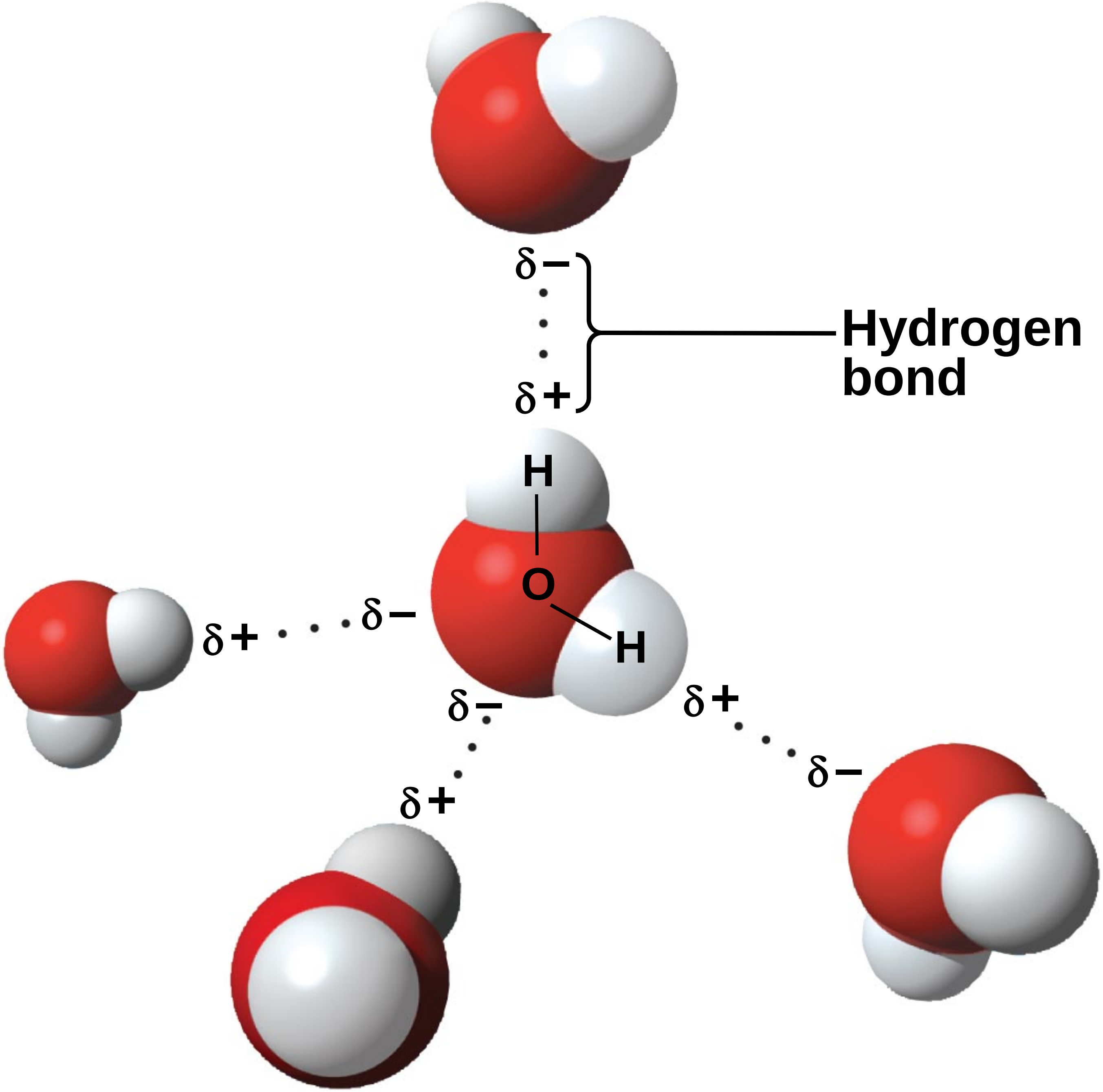
- Water is the biological medium on Earth
- All living organisms require water more than any other substance
- Most cells are surrounded by water, and cells themselves are about 70–95% water
- The abundance of water is the main reason the Earth is habitable

Fig. 3-1

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



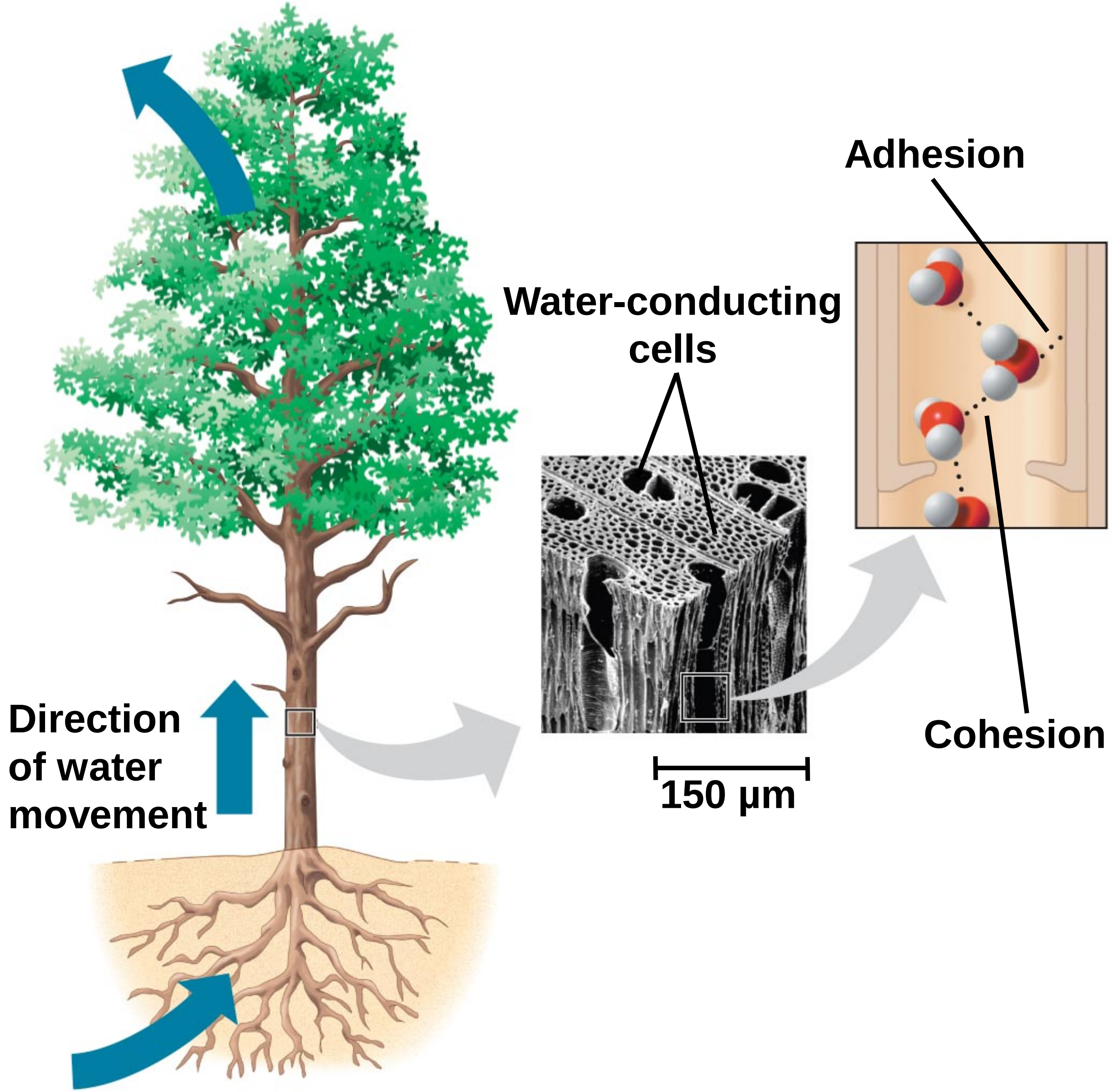
Concept 3.2: Four emergent properties of water contribute to Earth's fitness for life

- 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


Adhesion
Water-conducting cells
Direction of water movement
150 µm
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

Fig. 3-4

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 1ºC
- The specific heat of water is 1 cal/g/ºC
- 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:		
Substance	S (J/g 0C)	C (J/0C) for 100 g
Air	1.01	101
Aluminum	0.902	90.2
Copper	0.385	38.5
Gold	0.129	12.9
Iron	0.450	45.0
Mercury	0.140	14.0
NaCl	0.864	86.4
Ice	2..03	203
Water	4.179	417.9

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

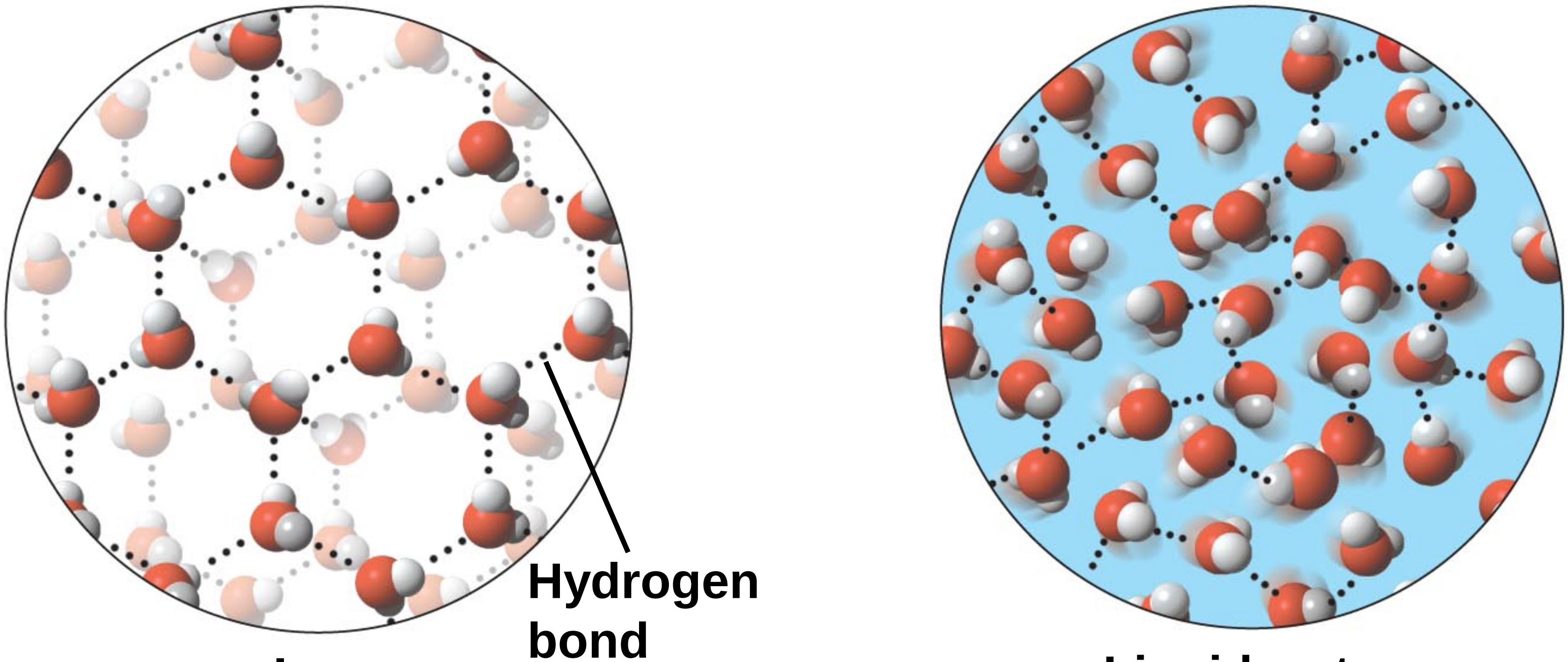
Evaporative Cooling

- *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

Fig. 3-6a



Ice

Hydrogen bonds are stable

Liquid water

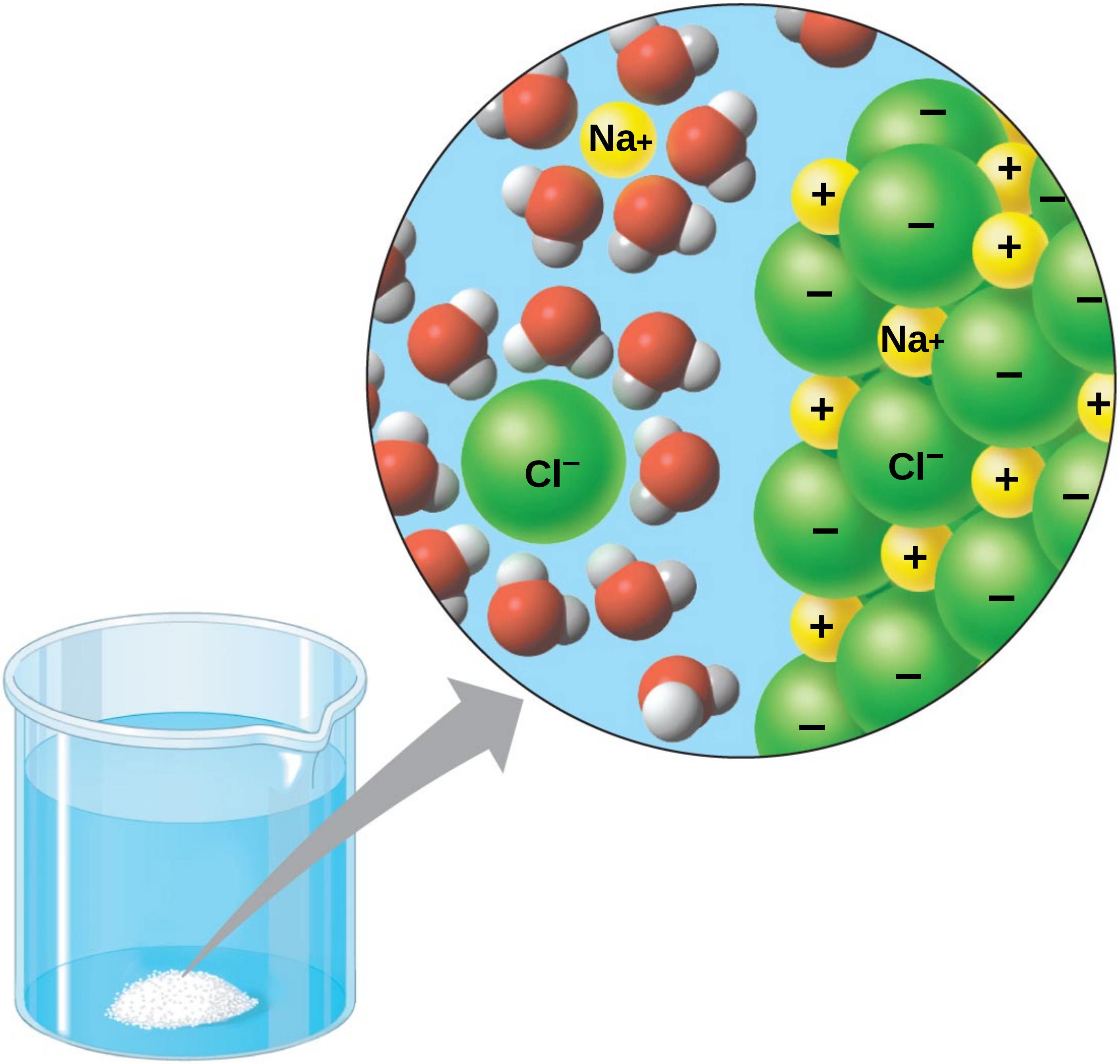
Hydrogen bonds break and re-form

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

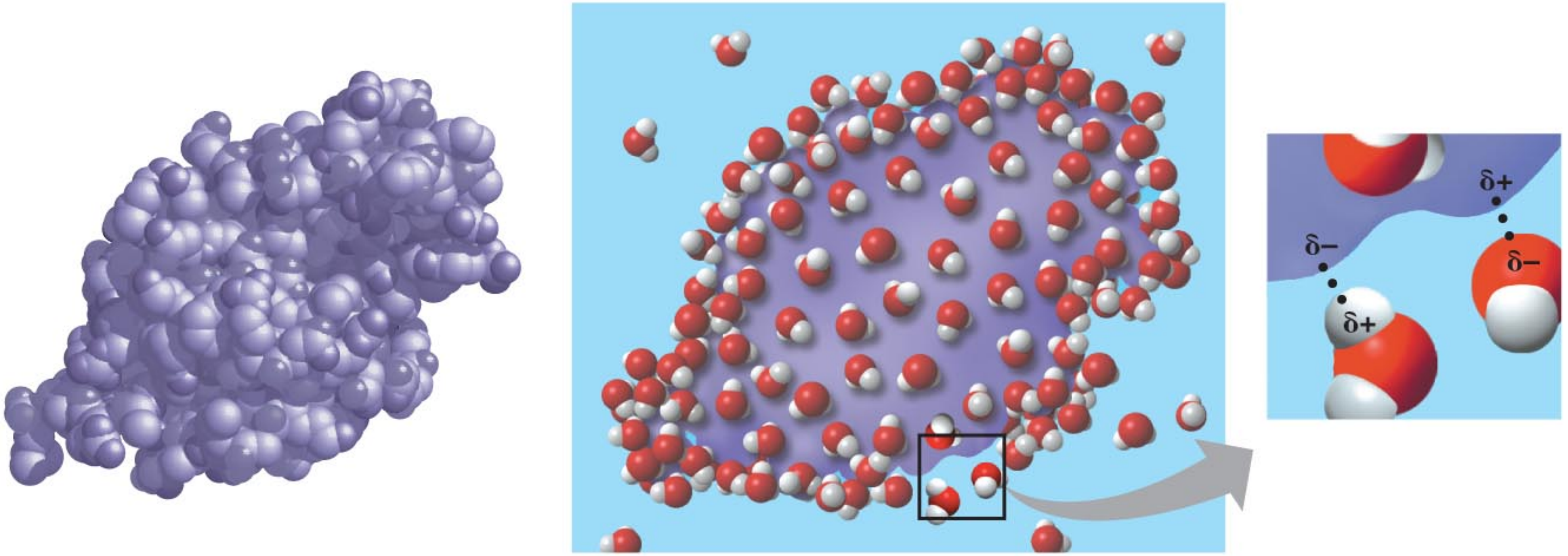
- Water is a versatile solvent due to its polarity, which allows it to form hydrogen bonds easily
- When an ionic compound is dissolved in water, each ion is surrounded by a sphere of water molecules called a **hydration shell**

Fig. 3-7



- 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

Fig. 3-8



(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)** = 6.02×10^{23} molecules
- Avogadro's number and the unit *dalton* were defined such that 6.02×10^{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 ($CaCl_2$), you would place how many gm of $CaCl_2$ into a container and then add how much pure water?

[mass of a Ca atom = 40; mass of a Cl atom = 35; mass of an O atom = 16; mass of an H atom = 1]

g
mw | moles

A. 75 gm of $CaCl_2$ then add 1 liter of water

B. 110 gm of $CaCl_2$ then add 1 liter of water

C. 128 gm of $CaCl_2$ then add 1 liter of water

D. 75 gm of $CaCl_2$ then add water to make a total volume of 1 liter

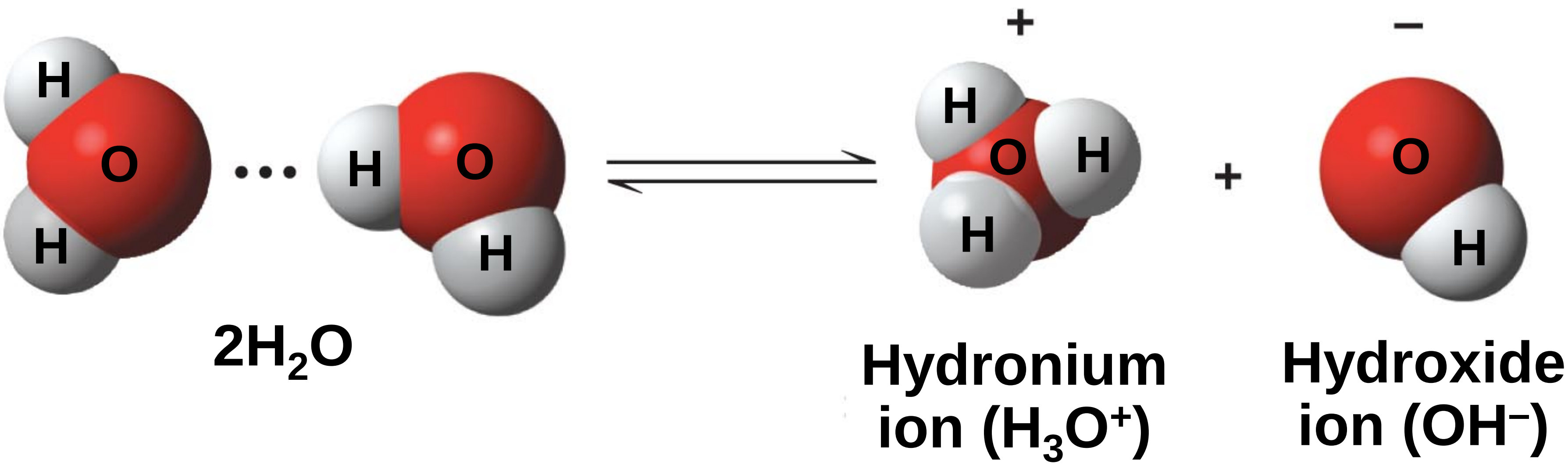
E. 110 gm of $CaCl_2$ then add water to make a total volume of 1 liter

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** (H^+)
  - The molecule with the extra proton is now a **hydronium ion** (H_3O^+), though it is often represented as H^+
  - The molecule that lost the proton is now a **hydroxide ion** (OH^-)

- Water is in a state of dynamic equilibrium in which water molecules dissociate at the same rate at which they are being reformed

Fig. 3-UN2



- 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 H^+ and OH^- are equal in pure water
- Adding certain solutes, called acids and bases, modifies the concentrations of H^+ and 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 H^+ concentration of a solution
- A **base** is any substance that reduces the 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

$$[H^+] \text{ is } 10^{-7} = -(-7) = 7$$

- 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

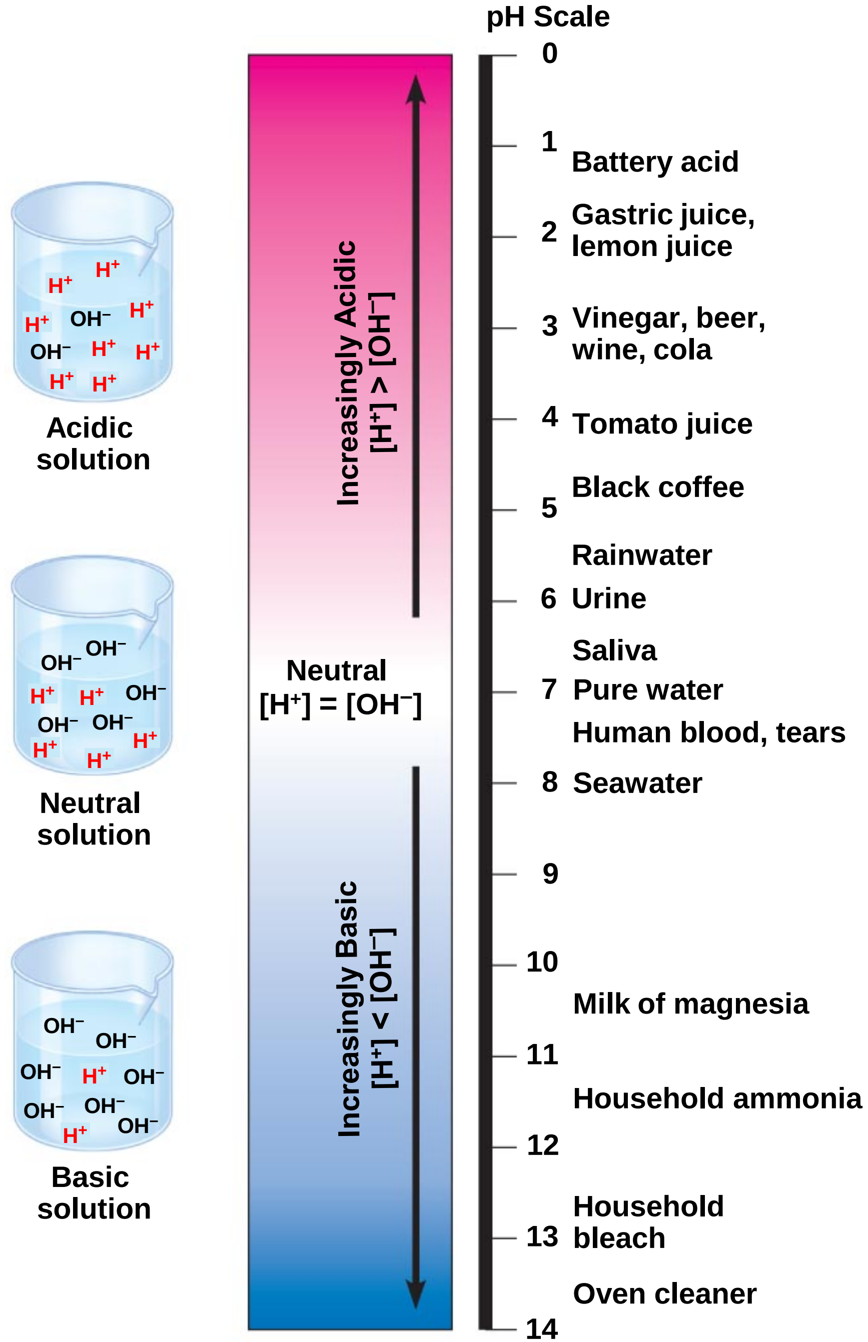
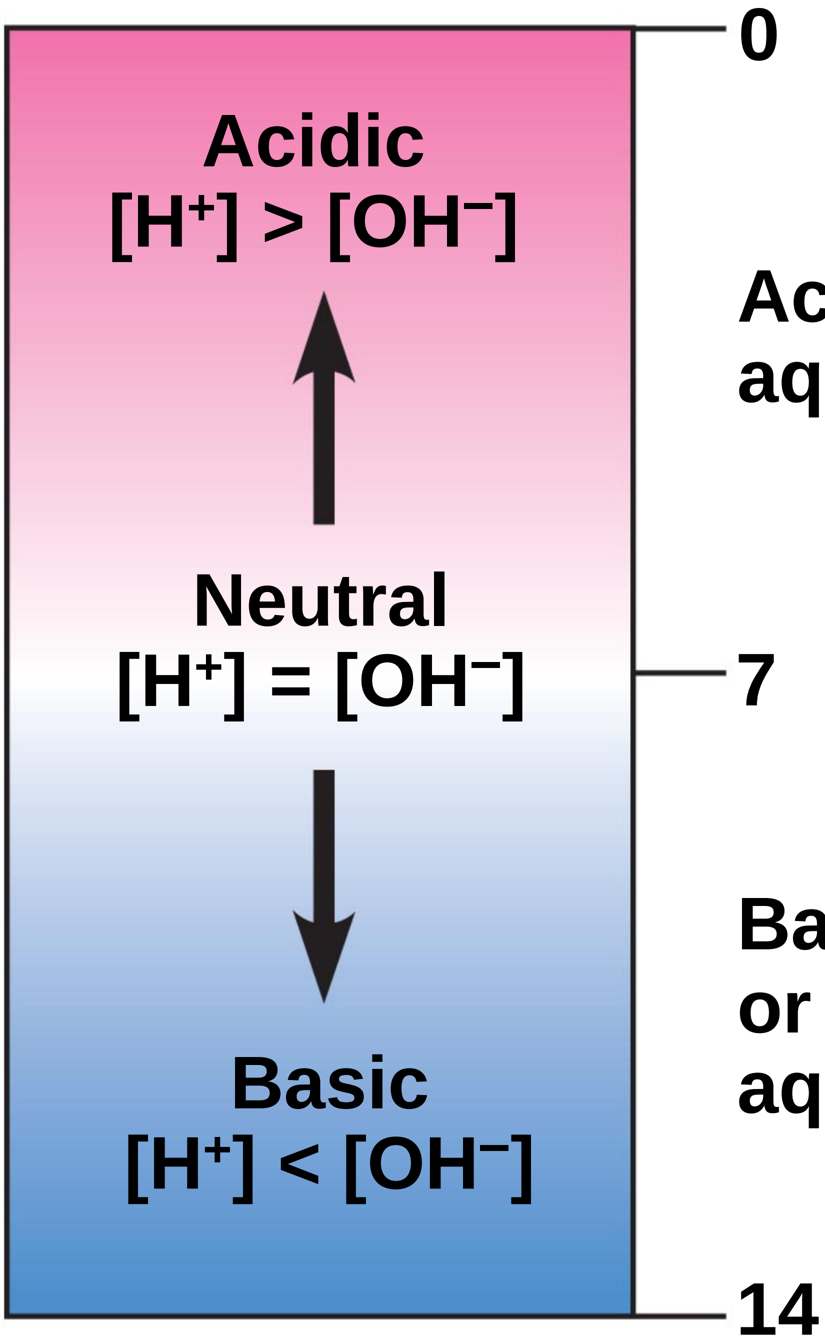

pH Scale
0
1
Battery acid
2
Gastric juice, lemon juice
3
Vinegar, beer, wine, cola
4
Tomato juice
Black coffee
5
Rainwater
6
Urine
Saliva
7
Pure water
Human blood, tears
8
Seawater
9
10
Milk of magnesia
11
Household ammonia
12
Household bleach
13
Oven cleaner
14
Increasingly Acidic
[H+] > [OH–]
Neutral
[H+] = [OH–]
Increasingly Basic
[H+] < [OH–]
H+
OH–
Acidic solution
Neutral solution
Basic solution

Fig. 3-UN5



Acidic
[H+] > [OH–]
Neutral
[H+] = [OH–]
Basic
[H+] < [OH–]
0
7
14
Acids donate H+ in aqueous solutions
Bases donate OH– or accept H+ in aqueous solutions

Buffers

- The internal pH of most living cells must remain close to pH 7
- **Buffers** are substances that minimize changes in concentrations of H^+ and OH^- in a solution
- Most buffers consist of an acid-base pair that reversibly combines with 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?

H_2CO_3	$\leftrightarrow$	HCO_3^-	+	H^+
Carbonic acid		Bicarbonate ion		Hydrogen ion

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