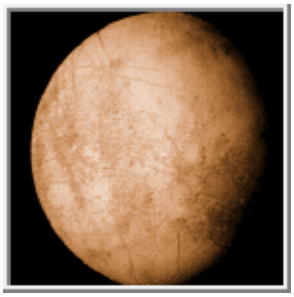
Chapter 3: Water and the Fitness of the Environment

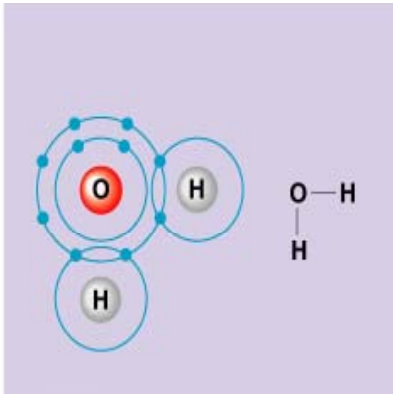




1. Structure of water molecules
2. Four properties of water
3. Acid and base

1. Structure of water molecules

- Water is a _____ molecule.
- _____ bonds between water molecules



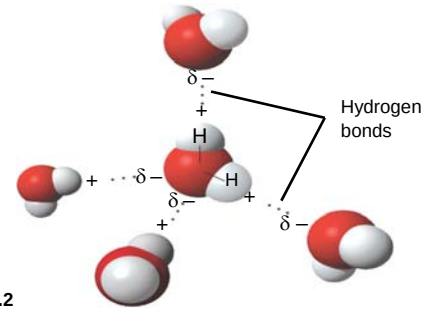


Figure 3.2

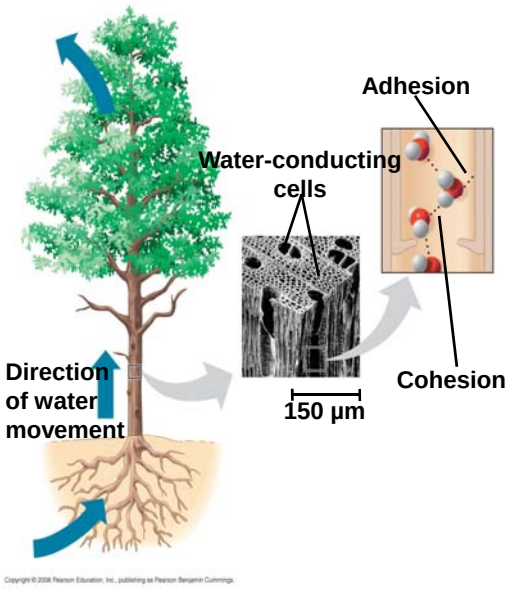
2. Four emergent properties of water

1. Cohesion
2. Temperature moderation
3. Expansion upon freezing
4. Versatile solvent

Properties result from polarity and hydrogen bonding.

2.1 cohesion

- A substantial percentage of water molecules are bonded to their neighboring molecules.
- due to _____
- Helps pull water up through the microscopic vessels of plants.



2.2 Moderation of temperature (heat bank)

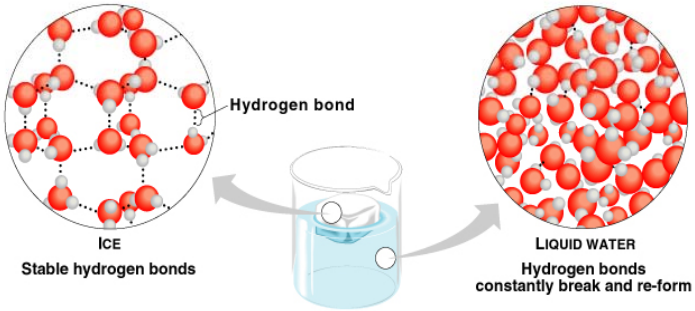
- Water absorbs heat from air that is _____.
  - ✓ Water absorb and store a large amount of heat from the sun in the daytime and during _____.
- Water releases heat to air that is _____.
  - ✓ Gradually cooling water warm the air at night and during _____.

Water has high specific heat.

- Specific heat (Cal/g/ºC): the amount of heat that must be absorbed or lost for __ g of that substance to change its temperature by __ºC.
- Water : 1 cal/g/ºC
- Ethyl alcohol : 0.6 ca./g/ºC
- Heat is absorbed when hydrogen bonds ____.
- Heat is released when hydrogen bonds ____.
- Evaporative cooling.

2. 3 Insulation of water by floating ice

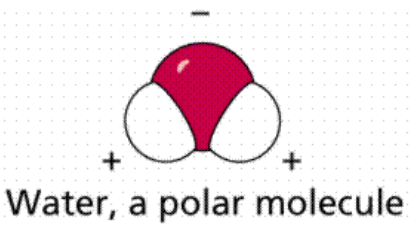
- Water is one of few substances ____ dense as solid than as liquid.
- Above 4°C, water expands as it warms and contracts as it cools.
- At _°C, water has the greatest density.
- At _°C, water becomes locked in to a crystalline lattice.



2. 4 The solvent of life

- *Nothing* worked better than water.
  - Solvent: a dissolving agent
  - Solute: a substance dissolved in solution
  - Solution: a liquid that is a homogeneous mixture of 2 or more substances
  - Aqueous solution: solution with water as solvent

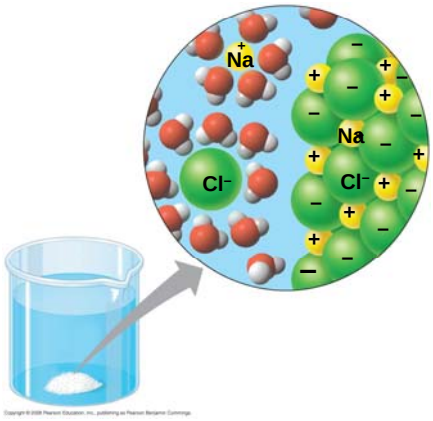
- Why is water a great solvent?



What happens when salt dissolves in water ?

$NaCl + H_2O \rightarrow$ aqueous salt solution

(solute) (solvent)

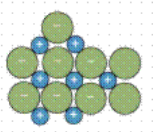


1. NaCl is hold by ionic bonds.
2. Na^+ and Cl^- ions are exposed to solvent.
3. Na^+ and Cl^- are attracted to _____ water molecules due to their electrical charges
   - Na^+ ions attracted to __
   - Cl^- ions attracted to __
   - Hydration shell

Depending on the solubility in water

Hydrophilic (hydro = water ; philo = loving)

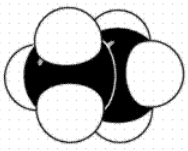
- hydrophilic substances have an affinity for water
- are therefore water soluble
- ____ and ____ substances
- example: salts



Hydrophobic (hydro = water ; phobos = fearing)

- are therefore not soluble in water
- _____ and _____ substances

  (compounds with a symmetric charge distribution)
- example: vegetable oil, O_2

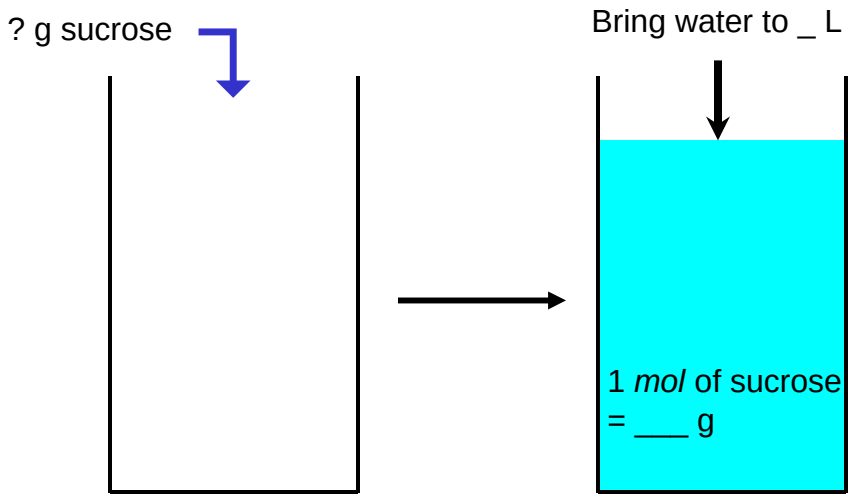


Solute concentration in aqueous solutions

- Solute concentration is calculated based on molecular mass (weight).
- Molecular mass: sum of the mass of all atoms in a molecule (expressed in _____).
- $C_{12}H_{22}O_{11}$ weighs 12(__) + 22(_) + 11(__) = ___ daltons
- Mole (*mol*): amount of a substance that has a mass in grams numerically equivalent to its molecular mass in daltons.
  - If sucrose weighs 342 daltons, then 1 mole of sucrose would equal 342 g
  - If compound 'X' weighs 67 daltons, 1 mole of 'X' = __ g ?
- Molarity (*M*): the number of ____ of solute per ___ of solution.

How would we make a 1 *M* of sucrose solution ?

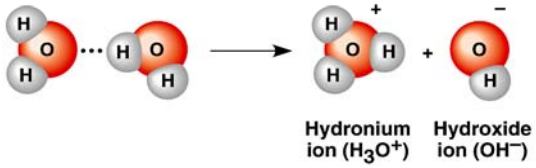
(Molarity = # of moles of solute per liter of solution)



3. Acid and base

- Occasionally, H atom shifts from one molecule to another

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



- Simply, dissociation of a water molecule into a H^+ and OH^-

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

- The concentration of each ion in pure water is 10^{-7} *M*.

What causes an aqueous solution to have an imbalance in its H^+ and OH^- concentration?

Acid: substance that _____ $[H^+]$ of a solution.

$HCl \rightarrow H^+ + Cl^-$

Base: substance that _____ $[H^+]$ of a solution.

$NH_3 + H^+ \rightarrow NH_4^+$

Types of solutions:

- Acidic solution: $[H^+]$ _ $[OH^-]$
- Basic solution: $[H^+]$ _ $[OH^-]$
- Neutral solution: $[H^+]$ _ $[OH^-]$

The pH scale

In any aqueous solution, the product of H^+ and OH^- concentrations is constant at 10^{-14} M:

$$[H+]\,[OH-] = 10^{-14}\ M$$

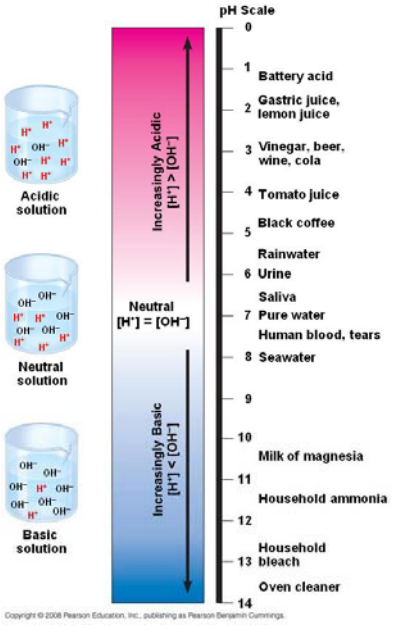
Consider neutral solution:

$[H+] = 10^{-7}$ and $[OH-] = 10^{-7}$ so $[H+]\,[OH-] = 10^{-14}$ M

If we add enough acid to increase [H+] to 10^{-5}, then [OH-] will decline to ___ such that $[H+]\,[OH-] = 10^{-14}$

Thus, if we know the [ ] of either H+ or OH-, we can calculate the [ ] of the other ion.

The pH scale



- pH: the negative logarithm of the $[H^+]$ concentration.
- $pH = -\log_{10}[H^+]$
- If $[H^+] = 10^{-7}$ M, then pH = 7
- pH scale ranges from 0 to 14.
- Acidic solutions have pH _ 7.
- Basic solutions have pH _ 7.
- Neutral solutions have pH _ 7.

1. Imagine a solution with one out of every 10 molecules being H^+

$[H^+]$ would be $1/10 = 0.1 = 10^{-1}$

$pH = -\log_{10} [H^+] = -\log_{10} [\_\_] = \_$

2. Now imagine a solution with one out of every 100 molecules being H^+

$[H^+]$ would be $1/100 = 0.01 = 10^{-2}$

$pH = -\log_{10} [H^+] = -\log_{10} [\_\_] = \_$

Buffers

- Substances _____ changes in __ in a solution.
- Internal pH of most cells is ~7.
- Wide pH fluctuations disrupt chemical processes in the cell.
- Buffers prevent pH fluctuations by
  - _____ H^+ ions when these are in excess
  - _____ H^+ ions when these are depleted
  - example: carbonic acid buffers pH of human blood

$$H_2CO_3 \underset{\text{If pH \_\_\_\_\_}}{\overset{\text{If pH \_\_\_\_\_}}{\rightleftharpoons}} HCO_3^- + H^+$$

pH and the environment

Acid precipitation: rain, snow or fog with pH < 5.6

- combustion of fossil fuels releases _____ and _____
- these compounds react with water to form strong acids which fall to earth