Water chemistry (chapter 2)

Water (H_2O)

A “V” shaped molecule

- Living things are mostly water

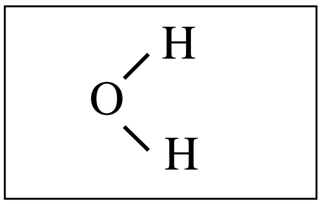


Fig 3.2

Substances are classified as either <u>hydrophilic</u> or <u>hydrophobic</u>, depending on whether they attract water or not

Hydrophilic substance = Any substance that attracts water

- Some typical traits of hydrophilic substances:

    a) Water spreads out on their surface (example: Paper)

    or…

    b) They dissolve in water (example: Table salt)

Hydrophobic substances = Any substance that does not attract water

- Some typical traits of hydrophobic substances:

    a) Water beads up on their surface (example: Wax)

    or…

    b) They do not dissolve in water (example: Oil)

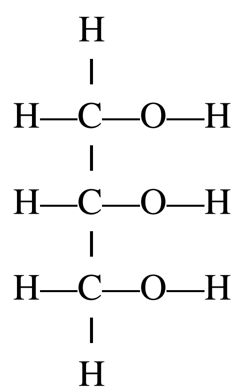
How to predict whether a substance is hydrophobic or hydrophilic by looking at its molecules:

1) Ions are always hydrophilic

√ Example: Table salt, made of Na^+ and Cl^- ions

2) Molecules where all (or most) of the carbons have partial charges are hydrophilic ("polar molecules")

√ Example: Glycerol



3) Molecules where all (or most) of the carbons have don't have partial charges are hydrophobic ("non-polar molecules")

√ Example: Octane

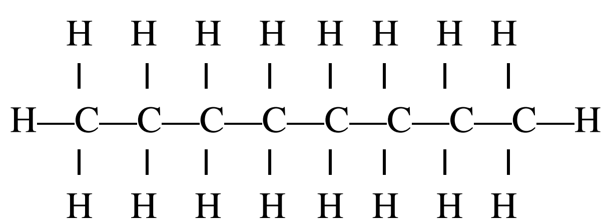
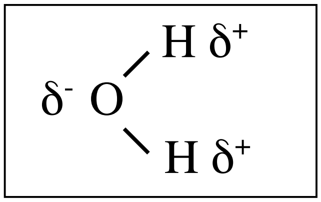


Fig 3.2 and 3.6

Each water molecule has partial charges

- The oxygen is δ^- and the hydrogens are δ^+



- Hydrophilic molecules are hydrophilic because they attract water electrostatically

  √ Water's δ^- oxygen is attracted to any positive charge (such as + ions and δ^+ hydrogen)

  √ Water's δ^+ hydrogen is attracted to any negative charge (such as – ions and δ^- oxygen or nitrogen)

- Hydrophobic molecules are hydrophobic because they have little or no charges to attract water molecules

Fig 3.2

Hydrogen bond

The electrostatic attraction between a δ^+ H and any δ^- atom (such as oxygen or nitrogen)

- Example: Water molecules bond to each other by hydrogen bonds
- Hydrogen bonds are shown as a dashed line

Figs 3.2 and 3.4

Dissolve

When a substance is evenly spread out in a liquid

- Three important terms related to dissolving a substance in a liquid:
  - √ Solvent = The liquid
  - √ Solute = The substance that is dissolved
  - √ Solution = The liquid and the substance dissolve in it
- When a substance dissolves in water, each molecule of the substance becomes coated with water molecules

Figs 2.14, 3.6, and 3.7

Concentration

The amount of solute per volume of solution

• Concentrations are usually stated as “grams per liter” (grams of solute per one liter of solution)

• A solute written in brackets means the concentration of that solute

√ Example: [Glucose] = The concentration of glucose

Diffusion

The spontaneous movement of a solute from an area of its high concentration to an area of its low concentration.

• Cell membranes can block diffusion of most solutes

Osmosis

The spontaneous movement of water across a selectively permeable membrane towards the side with the higher total solute concentration

- “Selectively permeable” = Water can cross membrane but solutes cannot
- “Water moves towards solutes” is a way to remember osmosis
- Hypertonic = Higher [solute] than a cell’s cytoplasm
    - √ Cells lose water when put in hypertonic solutions
- Hypotonic = Lower [solute] than a cell’s cytoplasm
    - √ Cells gain water when put in hypotonic solutions
- Isotonic = Equal [solute] to a cell’s cytoplasm
    - √ Cells don’t gain or lose water when put in isotonic solutions

Acid

Any molecule that adds H^+ ions to water

- Examples:

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

Hydrochloric acid

$$H_2CO_3 \rightarrow H^+ + HCO_3^-$$

Carbonic acid

Figs on pages 53 and 54

Base

Any molecule that removes H^+ ions from water

- Examples:

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

Hydroxide ion

$$HCO_3^- + H^+ \rightarrow H_2CO_3$$

Bicarbonate ion

Figs on pages 54

Water molecules can dissociate:

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

- Only a small number of water molecules are dissociated at any one moment

- In pure water, $[H^+] = [OH^-]$

Fig on page 53

pH scale

A number (from 0 to 14) that indicates how acidic or basic a solution is

- $pH = -\log[H^+]$
  - √ Example: $[H^+] = 0.0001 = 10^{-4}$ = pH 4
  - √ Example: $[H^+] = 0.000000001 = 10^{-9}$ = pH 9
- Pure water has a pH of 7 and is called “neutral” (not acidic or basic)
  - √ $[H^+] = [OH^-]$ at pH 7
- Solutions that are acidic have a higher $[H^+]$ than pure water
  - √ Acidic solutions have pHs **lower** than 7
  - √ The higher the $[H^+]$, the lower the pH
- Solutions that are basic have a lower $[H^+]$ than pure water
  - √ Basic solutions have pHs **higher** than 7
  - √ The lower the $[H^+]$, the higher the pH

Fig 3.8

Buffer

Substances that (when added to a solution) minimize changes in the solution's pH

- Buffers make a solution resistant to acids and bases

- A common buffer in living things is the carbonic acid/bicarbonate ion buffer

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

- The carbonic acid replaces any lost H^+ by:

$$H_2CO_3 \rightarrow H^+ + HCO_3^-$$

- The bicarbonate ion absorbs excess H^+ by:

$$HCO_3^- + H^+ \rightarrow H_2CO_3$$

Fig on page 55