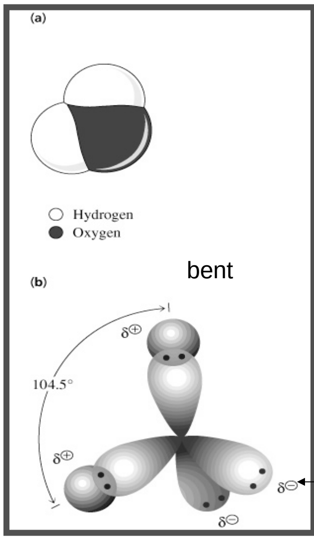
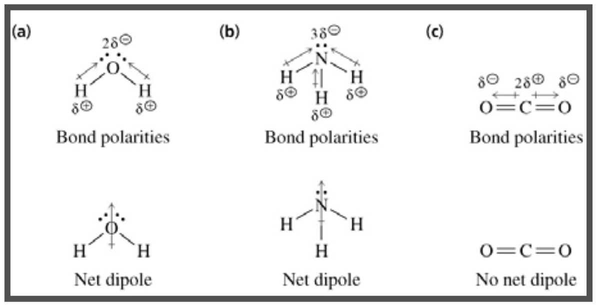
Water is the solvent for living cells.



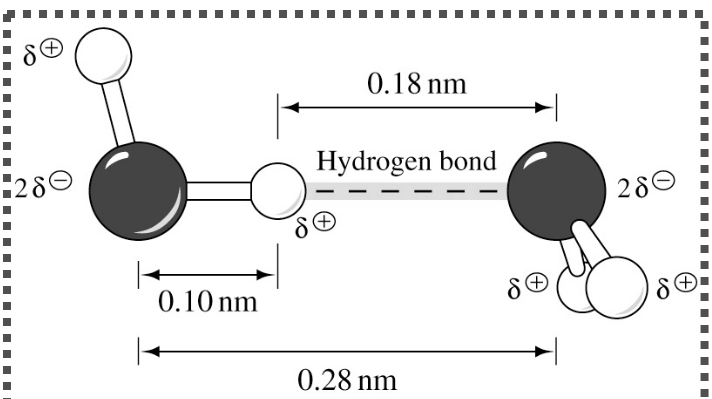
bent



polar

Hydrogen bonding between two water molecules

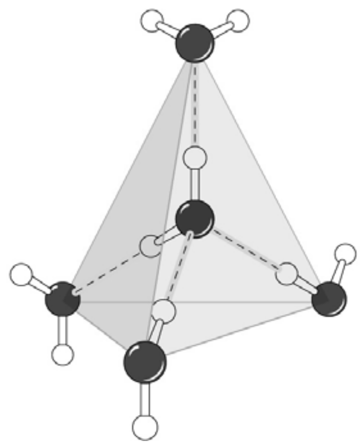
Hydrogen bond – force of attraction between O, N, F in one molecule and an H atom covalently bound to an O, N, F in another molecule.



Hydrogen bond is the strongest intermolecular force.

Hydrogen bonding by a water molecule

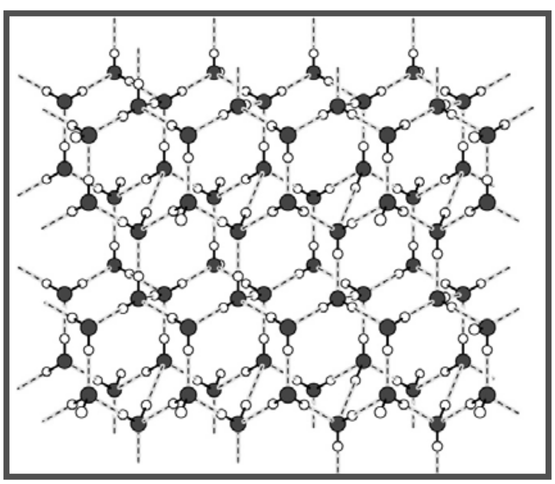
- A water molecule can form up to four hydrogen bonds
- Hydrogen bonds shown in yellow



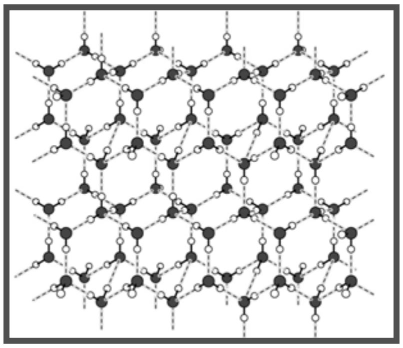
Structure of Ice.

- Hexagonal lattice structure
- Every water molecule is H - bonded to 4 others

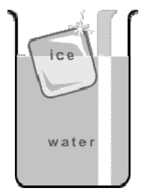
Water molecules are further apart in ice than in water.



Hydrogen bonds are shown in yellow.



When ice melts, 13% of the H-bonds are disrupted and structure collapses and molecules move closer together



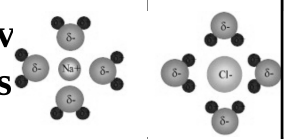
∴ Liquid water is more dense than ice.

All of unique properties of water are a consequence of its H-bonding

- High boiling and melting point
- High surface tension
- Ice floats on water

Ionic and Polar Substances Dissolve in Water

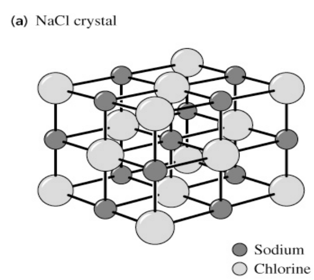
- **Hydrophilic (water-loving) substances (polar and ionic (electrolytes) readily dissolve in H_2O**
- **Polar water molecules align themselv around ions or other polar molecules**



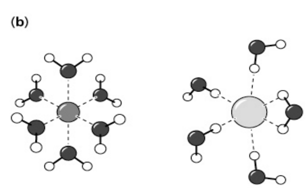
- **A molecule or ion surrounded by solvent molecules is solvated**
- **When the solvent is water the molecules or ions are hydrated**

Dissolution of NaCl in water

(a) Electrostatic forces hold ions together in crystalline sodium



(b) Water molecules form solvation spheres around Na^+ and Cl^-



Solubilities of molecules in water

- **Solubility in water depends upon the <u>ratio</u> of <u>polar</u> <u>to</u> <u>nonpolar</u> <u>groups</u> in a molecule**
- **The larger the portion of nonpolar groups the less soluble the molecule is in water**
- **The larger the portion of polar groups (e.g. hydroxyl groups (-OH)) the more soluble the molecule is in water**

Solubilities of short-chain alcohols in water

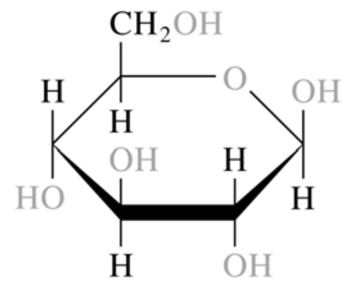
TABLE 2.1 Solubilities of short-chain alcohols in water

Alcohol	Structure	Solubility in water (mol/100 g 20°C)[a] H_2O at
Methanol	CH_3OH	∞
Ethanol	CH_3CH_2OH	∞
Propanol	$CH_3(CH_2)_2OH$	∞
Butanol	$CH_3(CH_2)_3OH$	0.11
Pentanol	$CH_3(CH_2)_4OH$	0.030
Hexanol	$CH_3(CH_2)_5OH$	0.0058
Heptanol	$CH_3(CH_2)_6OH$	0.0008

[a]Infinity (`) indicates that there is no limit to the solubility of the alcohol in water.

Structure of glucose

- Glucose has five hydroxyl groups and a ring oxygen which can hydrogen bond
- Glucose is very soluble in water

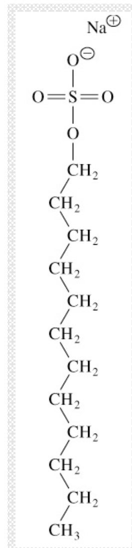


Nonpolar Substances Are Insoluble in Water

- **Hydrophobic** (water-fearing) molecules are nonpolar
- **Hydrophobic effect** - the exclusion of nonpolar substances by water (critical for protein folding and self-assembly of biological membranes)
- **Amphipathic molecules** have hydrophobic chains and ionic or polar ends. Detergents (surfactants) are examples.

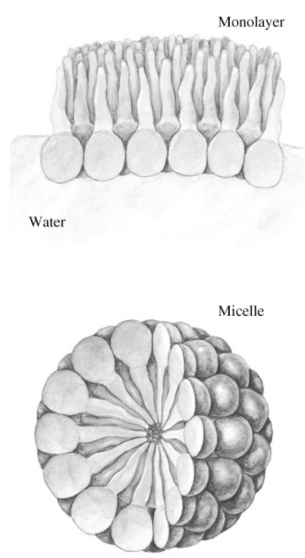
Sodium dodecyl sulfate (SDS)

- A synthetic detergent
- A 12-carbon tail
- Polar sulfate group



Detergents in water

- Monolayers can form on the surface
- At higher concentrations detergents can form **micelles**



Noncovalent Interactions in Biomolecules

Weak noncovalent interactions are important in:

- Stabilization of proteins and nucleic acids
- Recognition of one biopolymer by another
- Binding of reactants to enzymes

Noncovalent forces

There are four major types of noncovalent forces:

(1) **Charge-charge interactions**

(2) **Hydrogen bonds**

(3) **Van der Waals forces**

(4) **Hydrophobic interactions**

A. Charge-Charge Interactions (Ion Pairing)

- Electrostatic interactions between two charged particles
- Can be the strongest type of noncovalent forces
- Can extend over greater distances than other forces
- Charge repulsion occurs between similarly charged groups

Types of attractive charged interactions

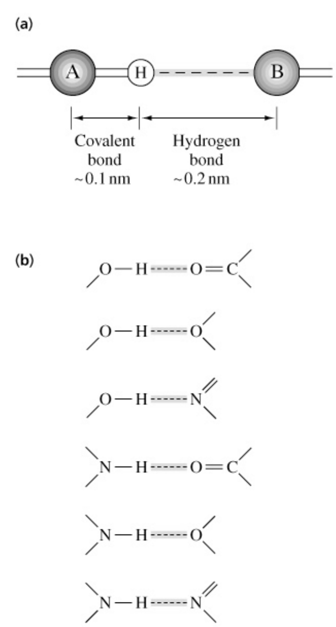
- **Salt bridges** - attractions between oppositely-charged functional groups in proteins
- **Ion pairing** - a salt bridge buried in the hydrophobic interior of a protein is stronger than one on the surface

B. Hydrogen Bonds

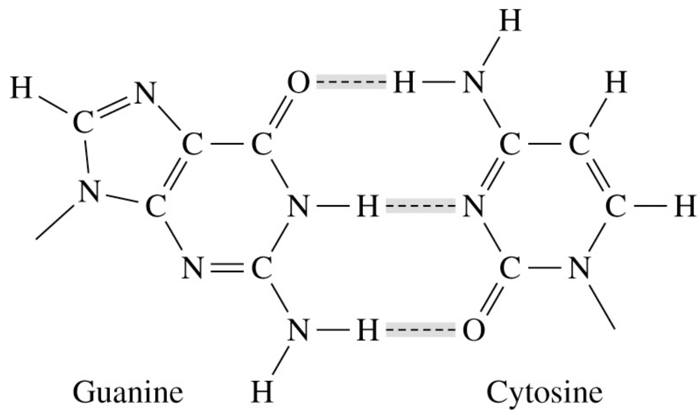
- Among the strongest of noncovalent interactions
- H atom bonded to N, O, S can hydrogen bond to another electronegative atom (~0.2 nm distance)
- Total distance between the two electronegative atoms is ~0.27 to 0.30 nm
- In aqueous solution, water can H-bond to exposed functional groups on biological molecules

(a) Hydrogen bonding between A-H and B

(b) Some biologically important H-bonds



Hydrogen bonding between complementary bases in DNA

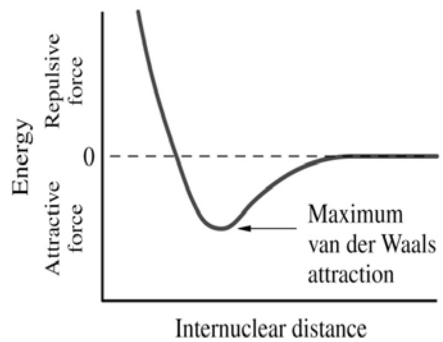


C. Van der Waals Forces

- **<u>Weak</u> <u>short</u> <u>range</u> <u>forces</u> between:**
  (a) Permanent dipoles of two uncharged molecules
  (b) Permanent dipole and an induced dipole in a neighboring molecule
- **Although individually weak, many van der Waals interactions occur in biological macromolecules and participate in stabilizing molecular structures**

Effect of internuclear separation on van der Waals forces

- Strongly repulsive at short internuclear distances, very weak at long internuclear distances
- Van der Waals attraction is maximal when two atoms are separated by their van der Waals radii



Van der Waals radii of several atoms

TABLE 2.2 Van der Waals radii of several atoms

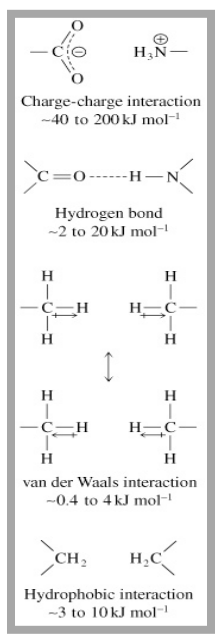
Atom	Radius (nm)
Hydrogen	0.12
Oxygen	0.14
Nitrogen	0.15
Carbon	0.17
Sulfur	0.18
Phosphorus	0.19

D. Hydrophobic Interactions

- **Association of a relatively <u>nonpolar</u> <u>molecule</u> or group with other nonpolar molecules**
- **Depends upon the <u>increased</u> <u>entropy</u> (+ΔS) which occurs when water molecules surrounding a nonpolar molecule are freed to interact with each other in solution**
- **The cumulative effects of many hydrophobic interactions can have a significant effect on the stability of a macromolecule**

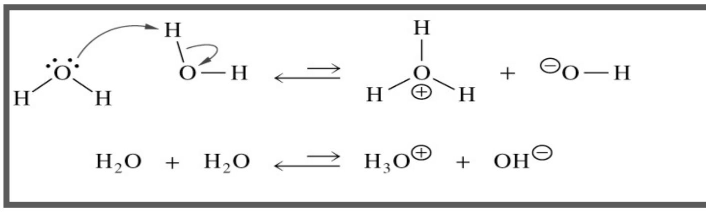
Noncovalent interactions in biomolecules

- Charge-charge interactions
- Hydrogen bonds
- Van der Waals interactions
- Hydrophobic interactions



Ionization of Water

- Pure water consists of a low concentration of **hydronium ions** (**H_3O^+**) and an equal concentration of **hydroxide ions** (**OH^-**)
- **Acids** are proton donors (e.g. H_3O^+) and **bases** are proton acceptors (e.g. OH^-)



(2.2)

The pH Scale

- **pH** is defined as the negative logarithm of the concentration of H^+

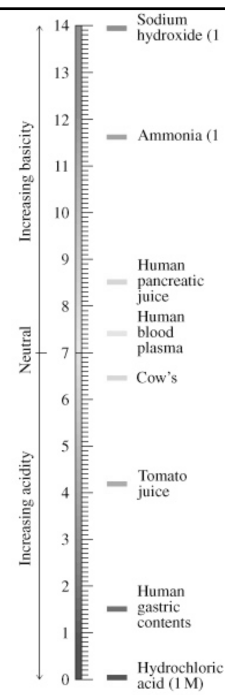
$$pH = -\log[H^{\oplus}]$$

TABLE 2.3 Relation of $[H^{\oplus}]$ and $[OH^{\ominus}]$ to pH

pH	$[H^{\oplus}]$ (M)	$[OH^{\ominus}]$ (M)
0	1	10^{-14}
1	10^{-1}	10^{-13}
2	10^{-2}	10^{-12}
3	10^{-3}	10^{-11}
4	10^{-4}	10^{-10}
5	10^{-5}	10^{-9}
6	10^{-6}	10^{-8}
7	10^{-7}	10^{-7}
8	10^{-8}	10^{-6}
9	10^{-9}	10^{-5}
10	10^{-10}	10^{-4}
11	10^{-11}	10^{-3}
12	10^{-12}	10^{-2}
13	10^{-13}	10^{-1}
14	10^{-14}	1

pH values for some fluids

- Lower values are acidic fluids
- Higher values are basic fluids



Acid Dissociation Constants of Weak Acids

- Strong acids and bases dissociate completely in water

$$HCl + H_2O \rightleftarrows Cl^- + H_3O^+$$

- Cl^- is the **conjugate base** of HCl
- H_3O^+ is the **conjugate acid** of H_2O

Acetic acid is a weak acid

- Weak acids and bases do not dissociate completely in H_2O

$$\underset{\text{(weak acid)}}{\underset{\text{Acetic acid}}{CH_3COOH}} \overset{K_a}{\rightleftharpoons} H^{\oplus} + \underset{\text{(conjugate base)}}{\underset{\text{Acetate anion}}{CH_3COO^{\ominus}}} \qquad (2.11)$$

The Henderson-Hasselbalch Equation

- Defines the pH of a solution in terms of:

  (1) The pK_a of the weak acid

  (2) Concentrations of the weak acid (HA) and conjugate base (A^-)

$$pH = pK_a + \log \frac{[A^{\ominus}]}{[HA]}$$

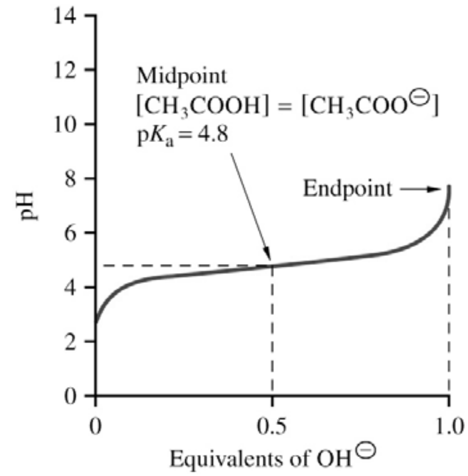
TABLE 2.4 Dissociation constants and pK_a values of weak acids in aqueous solutions at 25° C

Acid	K_a (M)	pK_a
HCOOH (Formic acid)	1.77×10^{-4}	3.8
CH_3COOH (Acetic acid)	1.76×10^{-5}	4.8
$CH_3CHOHCOOH$ (Lactic acid)	1.37×10^{-4}	3.9
H_3PO_4 (Phosphoric acid)	7.52×10^{-3}	2.2
$H_2PO_4^{\ominus}$ (Dihydrogen phosphate ion)	6.23×10^{-8}	7.2
HPO_4^{2-} (Monohydrogen phosphate ion)	2.20×10^{-13}	12.7
H_2CO_3 (Carbonic acid)	4.30×10^{-7}	6.4
$HCO_3^{\ominus}$ (Bicarbonate ion)	5.61×10^{-11}	10.2
$NH_4^{\oplus}$ (Ammonium ion)	5.62×10^{-10}	9.2
$CH_3NH_3^{\oplus}$ (Methylammonium ion)	2.70×10^{-11}	10.7

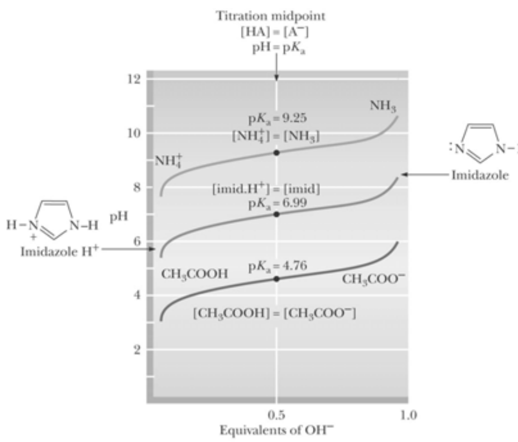
http://www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/animationsindex.htm

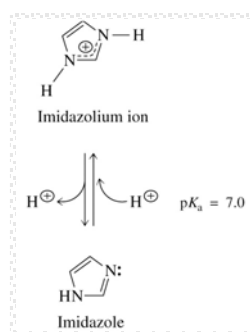
Titration curve of acetic acid (CH_3COOH)

- Titration curves are used to determine pK_a values

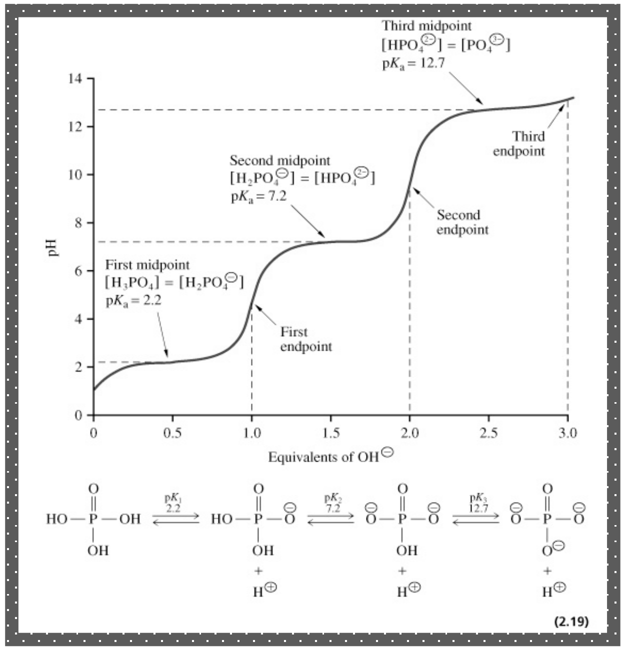


Titration of the imidazolium ion





Titration curve for phosphoric acid (H_3PO_4)



Buffered Solutions Resist Changes in pH

- Buffer capacity is the ability of a solution to resist changes in pH
- Most effective buffering occurs where:

  solution pH = buffer pK_a
- At this point: [weak acid] = [conjugate base]
- Effective buffering range is usually at pH values equal to the pKa ± 1 pH unit

Regulation of pH in the blood of animals

- **Blood plasma of mammals has a constant pH which is regulated by a buffer system of: carbon dioxide /carbonic acid /bicarbonate**
- **Buffer capacity depends upon equilibria between:**
  - **(1) Gaseous CO_2 (air spaces of the lungs)**
  - **(2) Aqueous CO_2 (dissolved in the blood)**
  - **(3) Carbonic acid**
  - **(4) Bicarbonate**

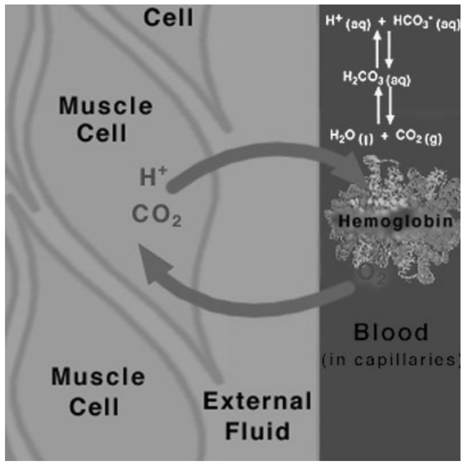
Carbonate buffering equilibria

$$H_2CO_3 \rightleftharpoons H^{\oplus} + HCO_3^{\ominus}$$

$$CO_2(\text{aqueous}) + H_2O \rightleftharpoons H_2CO_3$$

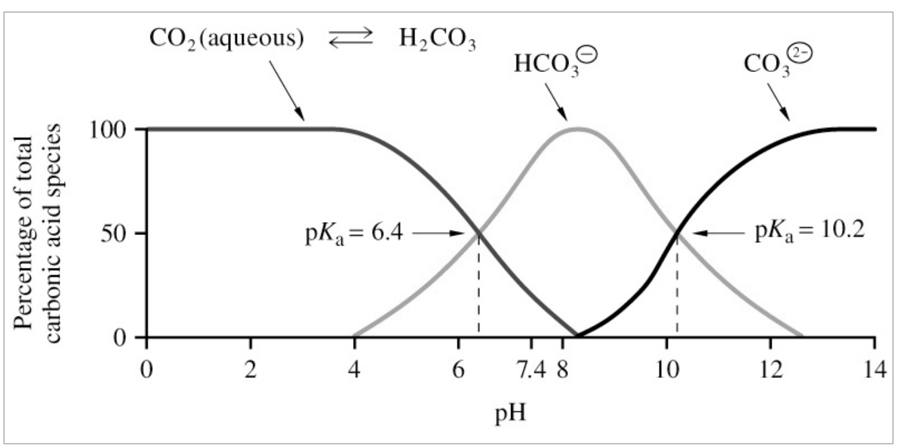
$$CO_2(\text{gaseous}) \rightleftharpoons CO_2(\text{aqueous})$$

$$H^+ (aq) + HCO_3^- (aq) \rightleftharpoons H_2CO_3 (aq) \rightleftharpoons H_2O (l) + CO_2 (g)$$



What happens to the blood during exercise?

Percentages of carbonic acid and its conjugate bases as a function of pH



Regulation of the pH of blood in mammals