

Water

the solvent for virtually all of biochemistry, ~70% of the mass within each cell is water

Carbon has **four electrons** in the **outer shell** - these hybridise into 4 sp^3 hybrid orbitals as tetrahedron. If symmetrical, the angle is **109.28 degrees**.

WATER AS A SOLVENT (substances such as household sugar **dissolve in water**, means that their molecules **separate** from each other, each becoming **surrounded** by water molecules.)

When a substance dissolves in a liquid, the mixture is termed a **solution**.

The dissolved substance is the **solute**, and the liquid that does the dissolving is the **solvent**.

Water is an excellent solvent for many substances because of its **polar bonds**.

Covalent bond	Inside a molecule
Hydrogen bond	between molecules
Polar bonds (H ₂ O)	uneven charge
Non-polar bonds (O ₂)	even charge

Henderson-Hasselbalch equation

$$pH = pK_a + \log \frac{[conjugate\ base]}{[weak\ acid]} \text{ (for weak acid)}$$

$$pOH = pK_b + \log \frac{[conjugate\ acid]}{[weak\ base]} \text{ (for weak base)}$$

pH (potential hydrogen)

The **acidity of a solution** is defined by the **concentration of H⁺ ions** it possesses.

pH scale **pH = -log10[H⁺]**

pure water **[H⁺] = 10⁻⁷ moles/liter**
pH (7)

acids substance, **proton (H⁺) donors**

bases substance, **proton acceptors (such as OH⁻)**

Water can act as both a weak acid and a weak base.

pH (potential hydrogen) (cont)

Acids in an aqueous environment proton moves from one molecule to the other

pH is a measure of **acidity (<7)** or **alkalinity (>7)**.

Higher amounts of protons in a solution results in a lower pH (acidic)

Lower amount of protons results in a higher pH (basic, or alkali)

Different Enzymes have different **optimal pH** according to their environment.

The **strength of an acid** is measured by its **dissociation constant, K_a**. The **larger the K_a** the more it **dissociates** and the **stronger the acid**.

The **pH** of a solution of a **weak acid** and its **conjugate base** is related to the **concentration of the acid and base** and the **pK_a** by the Henderson-Hasselbalch equation.

When **pH < pK_a**, the **weak acid predominates**. When **pH > pK_a**, the **conjugate base predominates**.

Buffers

A solution which pH **resists** change upon addition of either small amounts of strong acid or strong base are added.

(consist of a **weak acid and its conjugate base**)

BUFFER CAPACITY - is related to the concentrations of the weak acid and its conjugate base,

The greater the concentration of the weak acid and its conjugate base, the greater the buffer capacity.

H₂PO₄⁻ / HPO₄²⁻ is the principal buffer in cells, H₂CO₃ / HCO₃⁻ is an important buffer in blood.

Buffers work because the concentration of the weak acid and base are kept in the narrow window of the titration curve.

Biological Buffer Systems

Maintenance of intracellular pH is vital to all cells:

1. Enzyme-catalyzed reactions have optimal pH,
2. Solubility of polar molecules depends on H-bond donors and acceptors,
3. Equilibrium between CO₂ gas and dissolved HCO₃⁻ depends on pH.

Buffer systems in vivo are mainly based on:

1. Phosphate, concentration in millimolar range,
2. Bicarbonate, important for blood plasma,
3. Histidine, efficient buffer at neutral pH.

Buffer systems in vitro are often based on sulfonic acids of cyclic amines:

HEPES, PIPES, CHES.

