Chemistry Review: Inorganic Components of The Cell

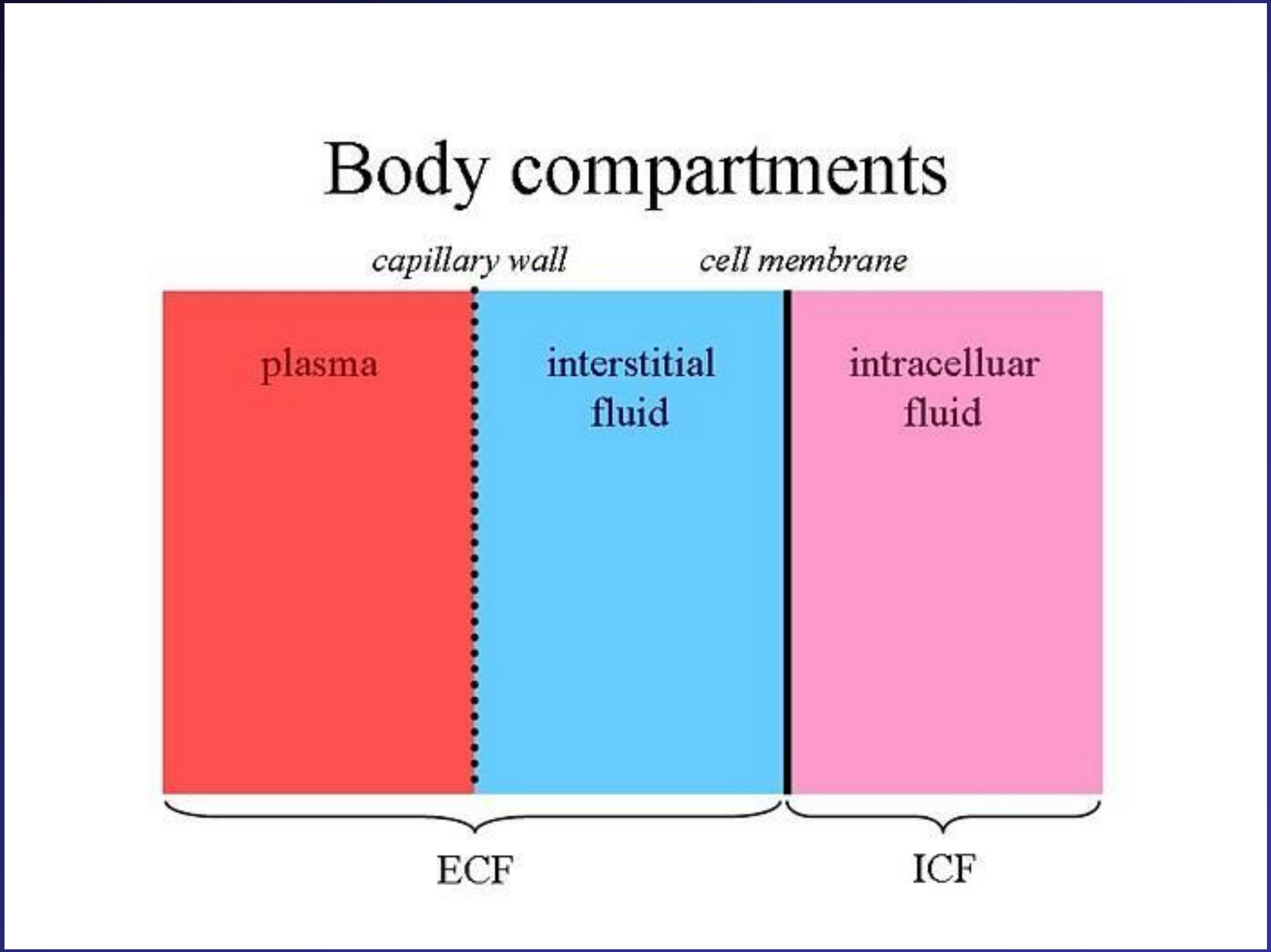
Lecture 3

Objectives

Understand/know/focus on/note

- what the inorganic components of the body are & fluid compartments
- what hydrogen ion concentration is and its relationship to pH
- what strong and weak acids and bases are, what a buffer is, and what makes a good buffer
- what extremes of pH can occur naturally in cell physiology
- what extremes of pH can indicate a disorder: alkalosis or acidosis

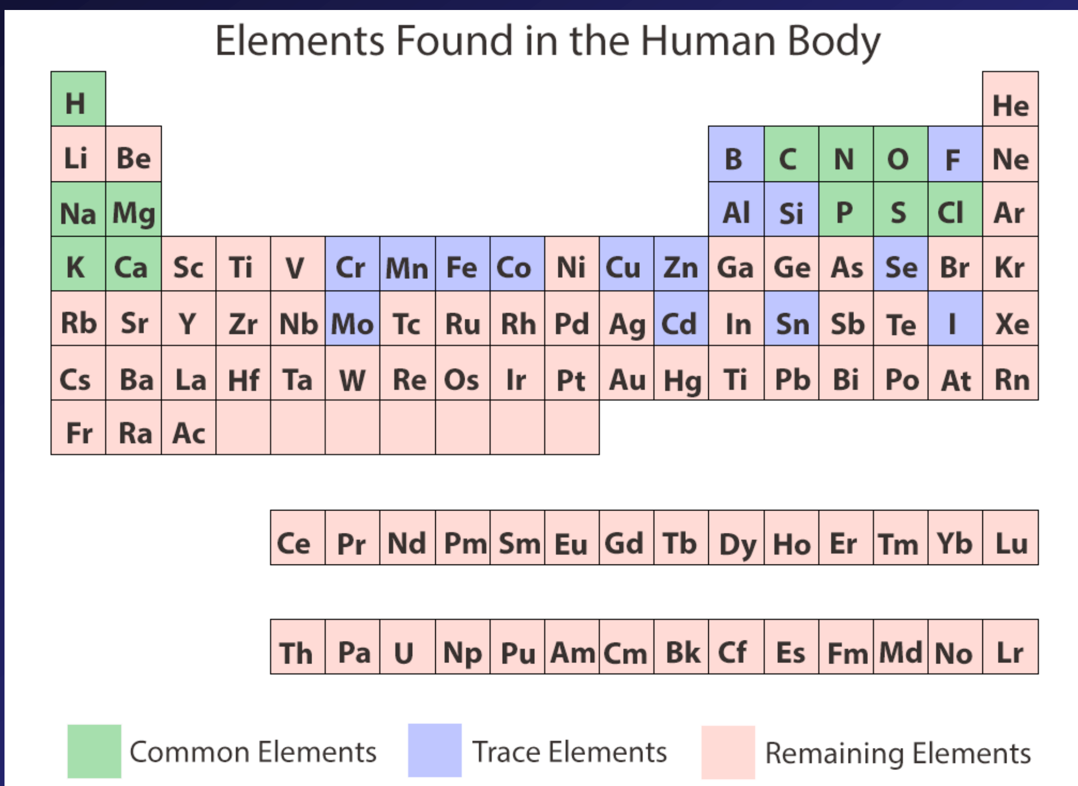
Division of Body Water

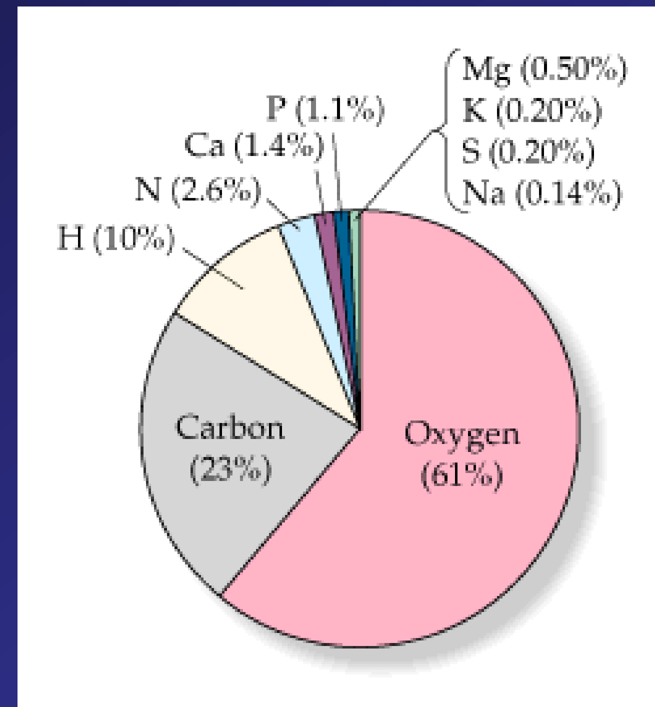


http://www.infobarrel.com/media/image/100142_max.jpg

Inorganic Components

- Water
  - Discussed as an important medium (making up 4/5th of cell mass)
- Acids & Bases & Salt forms
  - phosphates, carbonates
- Metals and Nonmetal ions
  - alkaline: Na, K
  - alkaline earth: Ca, Mg
  - transition: Fe, Cu, Zn, Mn, Mo





Percentage of elements in human body

Hydrogen Ion Concentration

- In aqueous solutions (including cytoplasm of cells), the hydrogen ion concentration $[H^+]$ is important since it can affect biochemical reactions
- Water is a weak acid (acids donate protons H^+)
- It is also a base (bases accept protons)
- Water can donate protons to itself:

$$H_2O + H_2O \leftarrow\rightarrow H_3O^+ + OH^-$$

$\leftarrow\rightarrow$ is supposed to mean that there is a forward and reverse reaction

The reaction is often shown as this $H_2O \leftarrow\rightarrow H^+ + OH^-$

- In pure water, the H^+ concentration reaches 10^{-7} moles per liter = 10^{-7} molar = 10^{-7} M
- This also implies that the $[OH^-] = 10^{-7}$ M

K_w & pH

- An equilibrium constant K is a quotient: it is the arithmetic product of the concentrations of all products divided the arithmetic product of the concentrations of all reactants

$$K = \frac{[\text{product1}][\text{product2}] \ldots [\text{product}n]}{[\text{reactant1}][\text{reactant2}] \ldots [\text{reactant}n]}$$

- The equilibrium constant for water dissociation is K_w,:

$$K_w = [H^+][OH^-] = [10^{-7}][10^{-7}] = 10^{-14}$$

those concentrations came from the previous slide

- Definition: $pH = -\log[H^+]$

for pure water: $pH = -\log 10^{-7} = 7$

Mineral Acid

- Hydrochloric acid (HCl) is an example of a strong acid

  whenever any uses term "mineral" acid, they imply strong acid

- Strong acids when added to water completely react with water molecules (which act as bases) to form hydronium ions (H_3O^+)

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

- Most chemists just leave out the fact that a proton (H^+) is transferred to an accepting base (H_2O) and write:

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

Mineral Acid

- If we add 1 mole HCl to 1 liter H_2O, all of the protons in HCl contribute to the $[H^+]$ concentration (100% ionization)
- So we have $[H^+] = 1$ M, and so the pH = –log 1 = 0
- If we added 0.001 mole to 1 L water, then

  [H+] = 0.001 M, so pH = –log 0.001 = 3

Mineral Bases

- Sodium hydroxide (NaOH) is an example of a strong (mineral) base
- It produces OH^- ions immediately, lowering the $[H^+]$
- As OH^- and H^+ are in an equilibrium with H_2O and affected by its equilibrium constant K_w
- Suppose we add 1 mol NaOH to 1 L H_2O, which produces $[OH^-] = 1$ M. What is the $[H^+]$ and pH?

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

the math says $[H^+] = 10^{-14}$ M, so pH = 14

- Now suppose we prepare 0.001 M NaOH. The math says $[H+] = 10^{-11}$ M, so pH = 11

Weak Acids & Bases

- Unlike HCl, weak acids do not completely give up their protons
- Acetic acid (CH_3COOH) is an example of a weak acid: when added to water, not all of it ionizes (gives H^+ to the solution)

- Unlike NaOH, weak bases do not completely take up protons in solution
- Ammonia (NH_3) is an example of a weak base

Buffers

- HA ←→ $H^+ + A^-$

  if HA = HCl, then [HA] = 0 (no HA left)

  if HA = CH3COOH, there is much HA remaining
- B + H^+ ←→ BH^+

  if B = NaOH, then [B] = 0 (no NaOH left)

  if B = NH3, there is much B remaining
- If you add HCl to just water, you get a change in pH that is dependent just on the H^+ produced by HCl
- But if you add HCl to water containing a high concentration of CH_3COOH, the H^+ resists changing because of the weak acid present

This makes weak acids and bases good buffers.

A buffer is a chemical that resists a change in pH when lots of H+ or OH– ions are produced

Buffering in the Body

- The pH of all body fluids, inside and outside of the cells, must be maintained in a narrow range or else enzymes stop doing metabolism
- Cytosolic pH values should be between 7.1-7.4 Phosphate (H_2PO4^-, HPO_4^{2-}) optimally buffers at pH 7.2, and along with proteins, keeps intracellular pH stable
- Blood is buffered using the equilibrium between bicarbonate and carbon dioxide

$$HCO_3^- + H^+ \leftrightarrow H_2CO_3 \leftrightarrow H_2O + CO_2$$

The left side ions HCO_3^- and H^+ are dissolved forms in the blood, while the CO_2 is a gas that forms in the lungs (to be exhaled)

Acidosis & Alkalosis

- Normal cell metabolism produces acids that can lower pH. Proper kidney and lung function ensures there is no acid increase (metabolic acidosis)
- Hyperventilation causes exhalation of CO_2 in excessive quantities and respiratory alkalosis

  this pulls the equilibrium reaction shown on the previous slide to the right, causing a depletion of the H^+ ion in the blood, thus increasing the pH and leading to alkalosis

Physiological Extremes of pH

- Extremes of pH can be generated quite naturally
- Within cells: membrane-bound lysosomes particularly within phagocytic immune cells (neutrophils) will produce an internal pH of about 2: shown by enzymes that work in the lysosome having a maximal activity at that low pH
- At organ/system level: it is well known that parietal cells in the stomach are engines that pump H^+ into the gastric fluid for digestion, and that enzymes like pepsin have maximal activity at these low pH values

Reading (Sources)

- Becker's WotC: Chapter 2
- Raven: Chap 2
- Marieb: pp 40-42