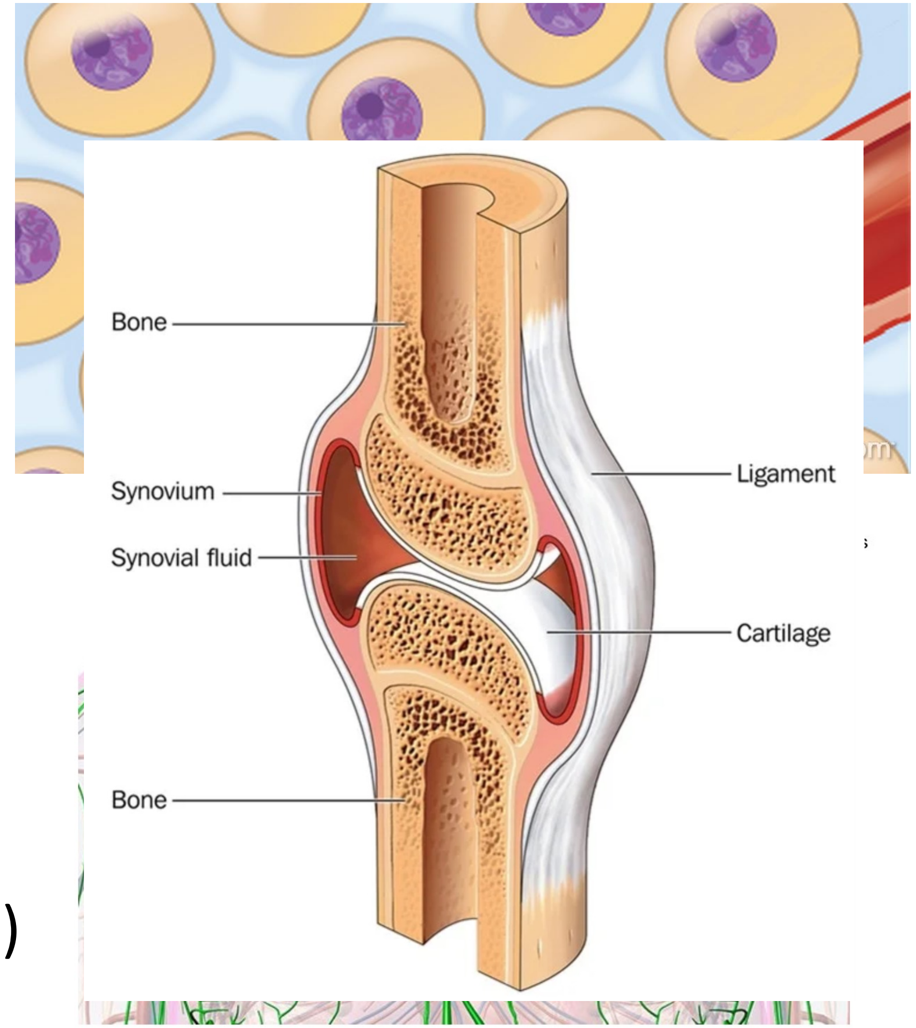
Lesson 1
Water, pH and buffers

H_2O fundamentals

- The fundamental molecule of life
- 60-95% of living human cells is H_20
  - 55% in intracellular fluids
  - 45% divided between:
    - Plasma (8%)
    - Interstitial (between cells) and lymph (22%)
    - Connective tissue, cartilage and bones (15%)



Lymph (from Latin, lympha) is the fluid that flows through the lymphatic system, a system composed of lymph vessels (channels) and intervening lymph nodes whose function, like the venous system, is to return fluid from the tissues to the central circulation (tissuse and organs drainage)

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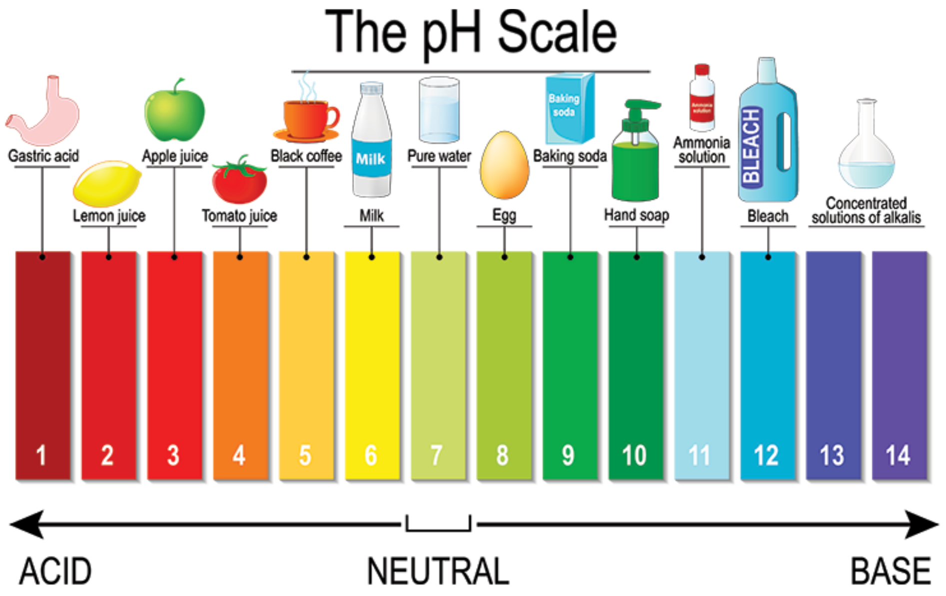
- Water balance must be preserved:
  - Loss >> intake --> dehydration
  - Intake >> loss --> edema

Water dissociation and pH

- $H_20 <-> H_30^+ + OH^-$
  - $[H_3O^+] = [OH^-]$ --> neutral
  - $[H_3O^+] > [OH^-]$ --> acidic
  - $[H_3O^+] = [OH^-]$ --> basic
- $K_w = [H_3O^+] \times [OH^-] = 10^{-14}$ (25°C)
  - In human body @ 37°C --> $K_w = 2.4 \times 10^{-14}$
- In **pure water** --> $[H_3O^+] = [OH^-] = 10^{-7}$ (25°C)
  - $pH = -\log[H_3O^+]$ --> in pure water pH = 7
  - In human blood @ 37°C --> pH =7.4 --> $[H_3O^+] = 3.98 \times 10^{-8}$
- $pK_w = -\log K_w = 14$



Weak acids and bases

- In biology we have only weak acids and bases (incomplete dissociation)

$CH_3COOH + H_2O$ <--> $H_3O^+ + CH_3COO^-$

CH_3COO^- conjugate base

$K_a = [H_3O^+][CH_3COO^-]/[CH_3COOH]$

$pK_a = -\log K_a$

$NH_3 + H_2O$ <--> $NH_4^+ + OH^-$

NH_4^+ conjugate acid

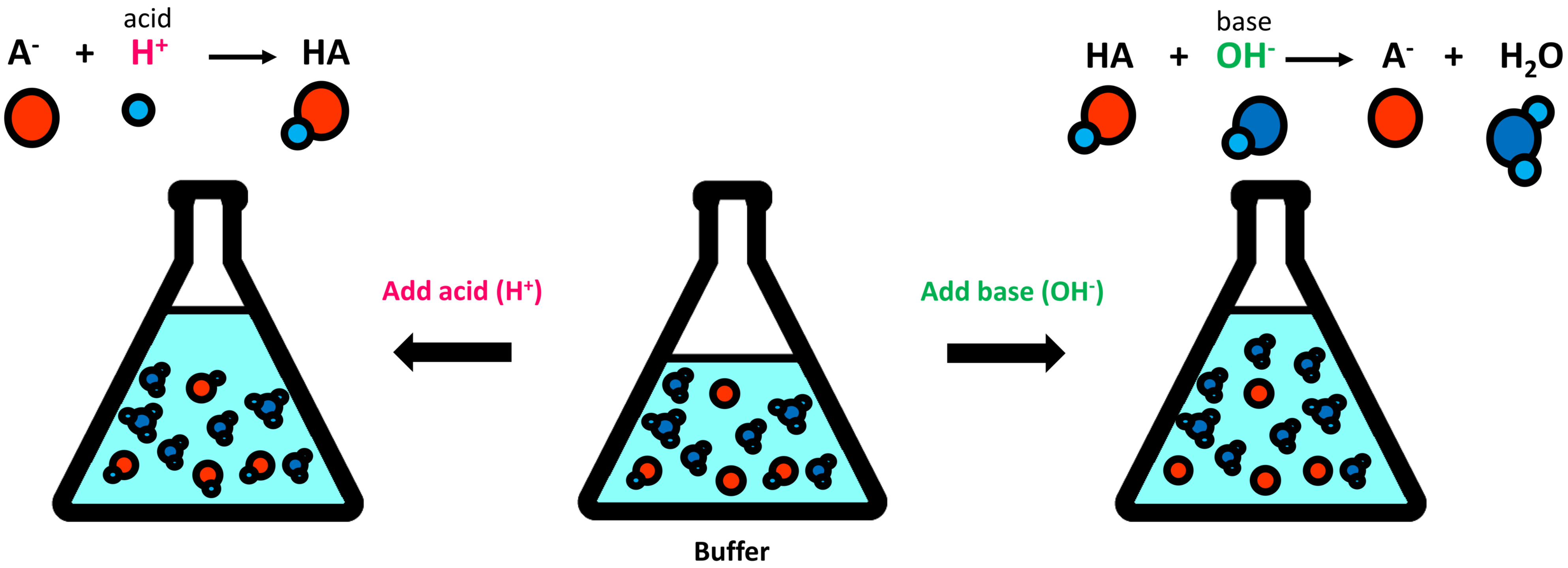
$K_b = [NH_4^+][OH^-]/[NH_3]$

$pK_b = -\log K_b$

$$K_a \times K_b = K_w = 10^{-14} \rightarrow pK_a + pK_b = 14$$

Buffers and pH control

- A solution that contains a conjugate acid-base pair of any weak acid or base in relative proportions to resist pH change when small amounts of either a (strong) acid or base are added is a buffer solution

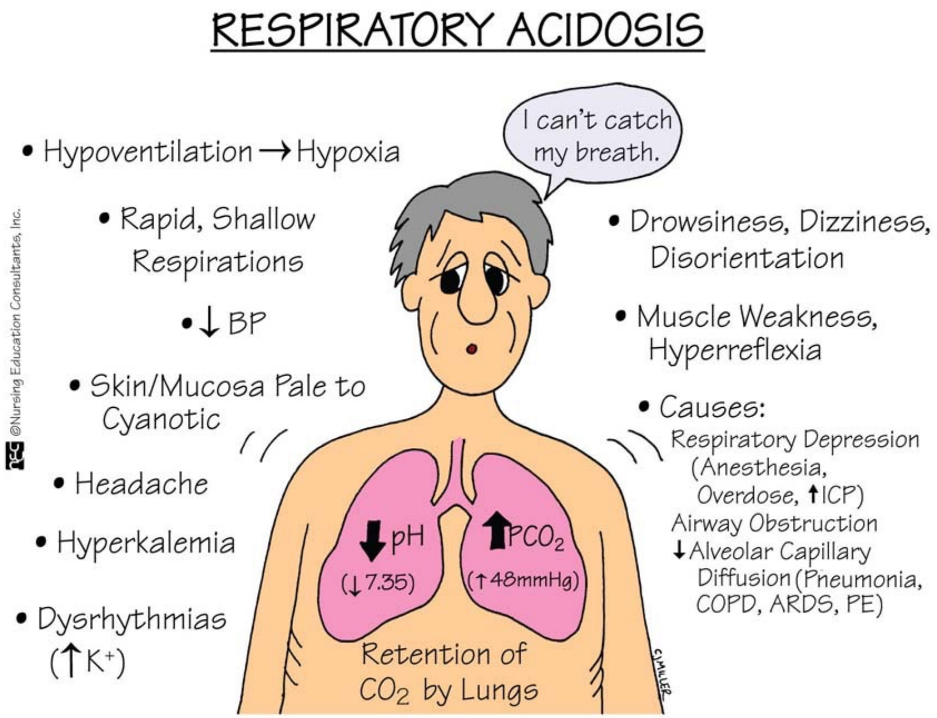


Identifying common physiological buffers

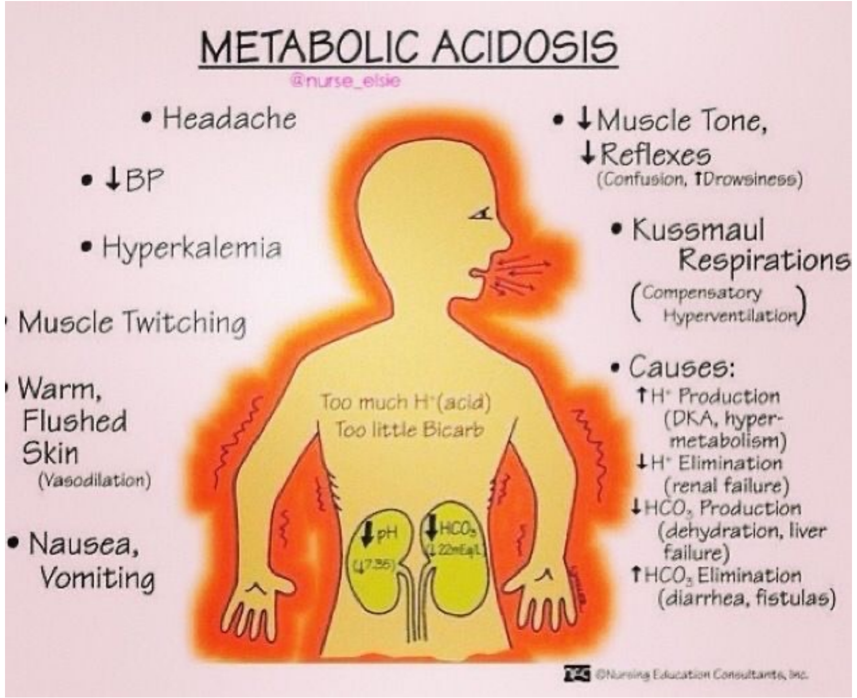
- Stomach pH = 1-2
- GI tract pH = 8-9
- Blood pH = 7.4
  - Blood pH < 7.2 --> pathological condition = acidosis
  - Blood PH < 6.8 --> death
  - Blood pH > 7.6 --> pathological condition = alkalosis
  - Blood pH > 8.6 --> death

Identifying common physiological buffers

- Respiratory acidosis
  - Inefficient expulsion of CO_2, increased concentration of H_2CO_3, impaired-respiration pathologies (pneumonia, emphysema, asthma)

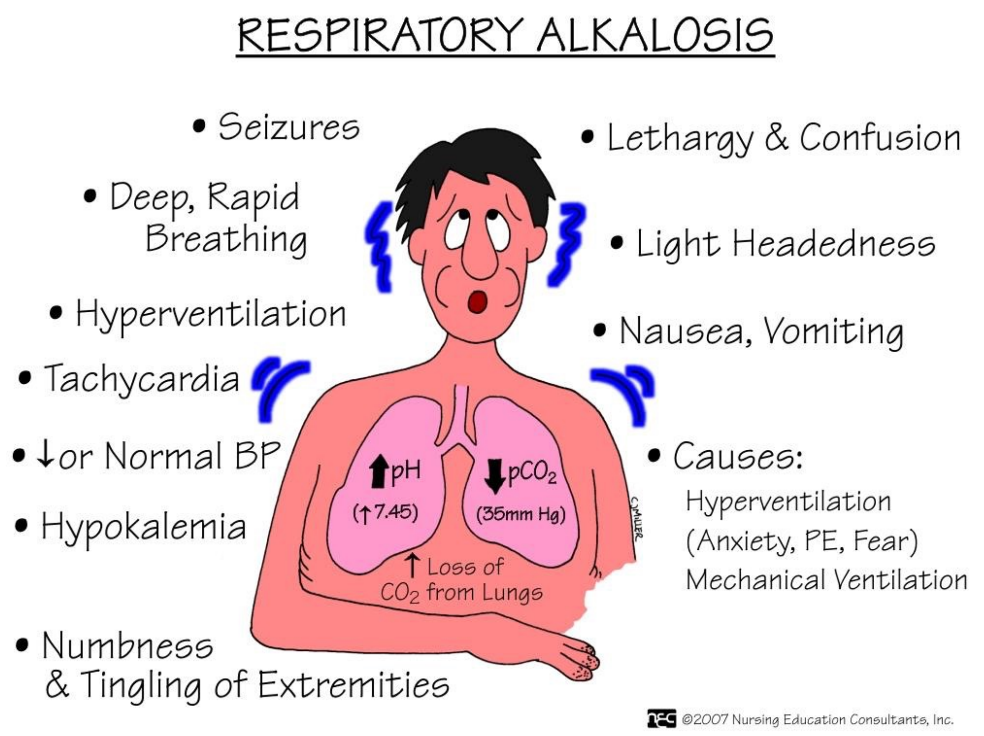


- Metabolic acidosis
  - Decreased concentration of HCO_3^-, results from various kidney diseases, uncontrolled diabetes, or vomiting of non-acid fluids

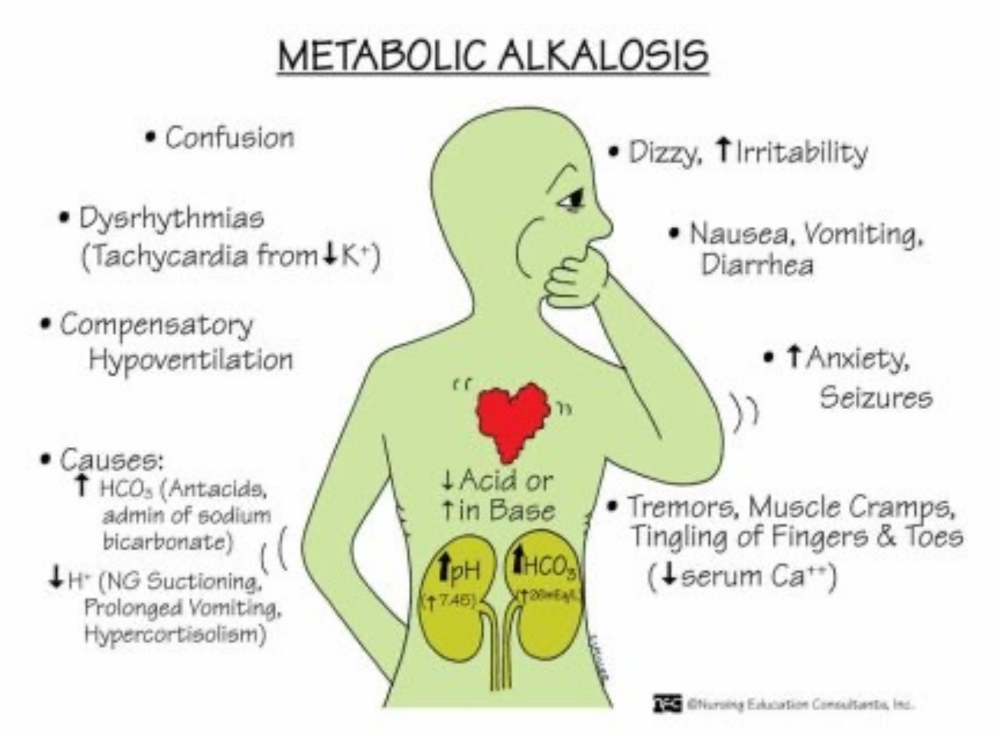


Identifying common physiological buffers

- Respiratory alkalosis
  - Excessive CO_2 removal, decreased concentration of H_2CO_3, hyperventilation



- Metabolic alkalosis
  - Increased concentration of HCO_3^-, results excessive vomiting of stomach acid



Identifying common physiological buffers

- PROTEIN BUFFER SYSTEM
  - Protein buffer system helps to maintain acidity in and around cells
- PHOSPHATE BUFFER SYSTEM
  - Phosphate buffer helps to maintain intracellular and urine pH
- BICARBONATE BUFFER
  - main extracellular buffer, main blood buffer
    - Helps in controlling CO_2 levels --> $CO_2 + H_2O$ <--> H_2CO_3 <--> $H^+ + HCO_3^-$
    - Coupled with CO_2 (blood) <--> CO_2 (lungs)

Buffer action and the pH of blood

- Normal blood pH = 7.4
  - Kept at this value by the buffering action of HCO_3^-, resulting from these two parallel physiological equilibria:
    - $CO_2 + H_2O \text{ <--> } H_2CO_3$
    - $H_2CO_3 \text{ <--> } H^+ + HCO_3^-$
- $K_{eq} = [H^+][HCO_3^-]/[CO_2]^* = 7.95 \times 10^{-7} \text{ --> } pK_{eq} = -\log K_{eq} = 6.1$
- Rearranging:

$$[H^+] = K_{eq} \times [CO_2]/[HCO_3^-] \text{ --> } \log[H^+] = \log K_{eq} + \log([CO_2]/[HCO_3^-] \text{ -->}$$

$$-\log[H^+] = -\log K_{eq} - \log[CO_2]/[HCO_3^-] \text{ --> } \color{red}{pH = pK_{eq} + \log[HCO_3^-]/[CO_2]}$$

*$[H_2O]$ = 55.6 M = const = included in the K_{eq} value

Quiz time

- A patient suffering from acidosis had a blood pH of 7.15 and a CO_2 concentration of 1.15 mM.
- If the reference range for pH = 7.4 are:

$[HCO_3^-]$ = 22.0 – 26.0 mM (average = 24 mM)

$[CO_2]$ = 1.20 mM

and pK_{eq} for the bicarbonate buffer = 6.1

- Q1. What was the patient's bicarbonate ($HCO3^-$) concentration?
- Q2. What are the implications of this value to the buffer capacity of the blood?

Quiz time

R1.

$pH = pK_{eq} + \log[HCO_3^-]/[CO_2]$ --> $7.15 = 6.1 + \log[HCO_3^-]/(1.15 \times 10^{-3})$

$10^{1.05} = [HCO_3^-]/(1.15 \times 10^{-3})$ --> $[HCO_3^-] = 12.9 \times 10^{-3}$ --> $[HCO_3^-] = 12.9$ mM

R2.

Normal $[HCO_3^-]$ average value = 24 mM --> $[HCO_3^-]$ in patient lowered by 11.1 mM --> severely impaired buffer capacity --> any further, small acid production will have serious consequences for the patient