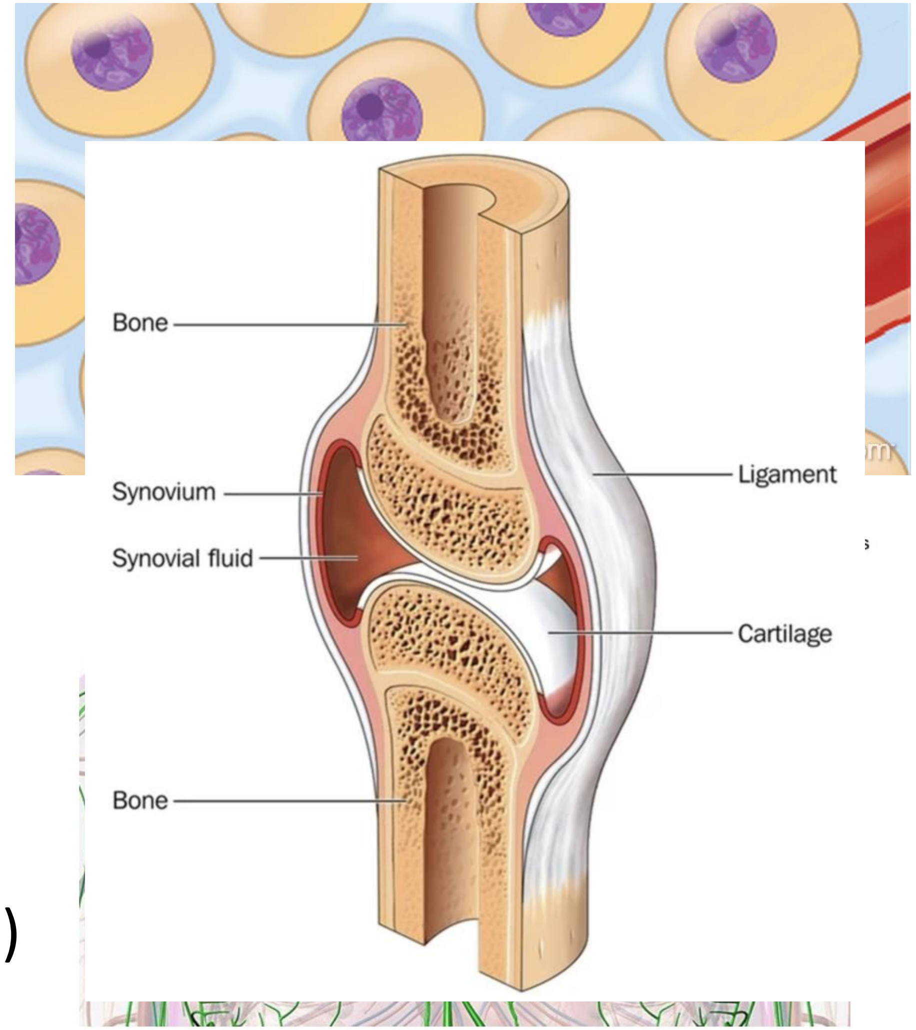
Lesson 1
Water, pH and buffers

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

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  - Body temperature maintenance
  - Solvent in the GI and excretion system

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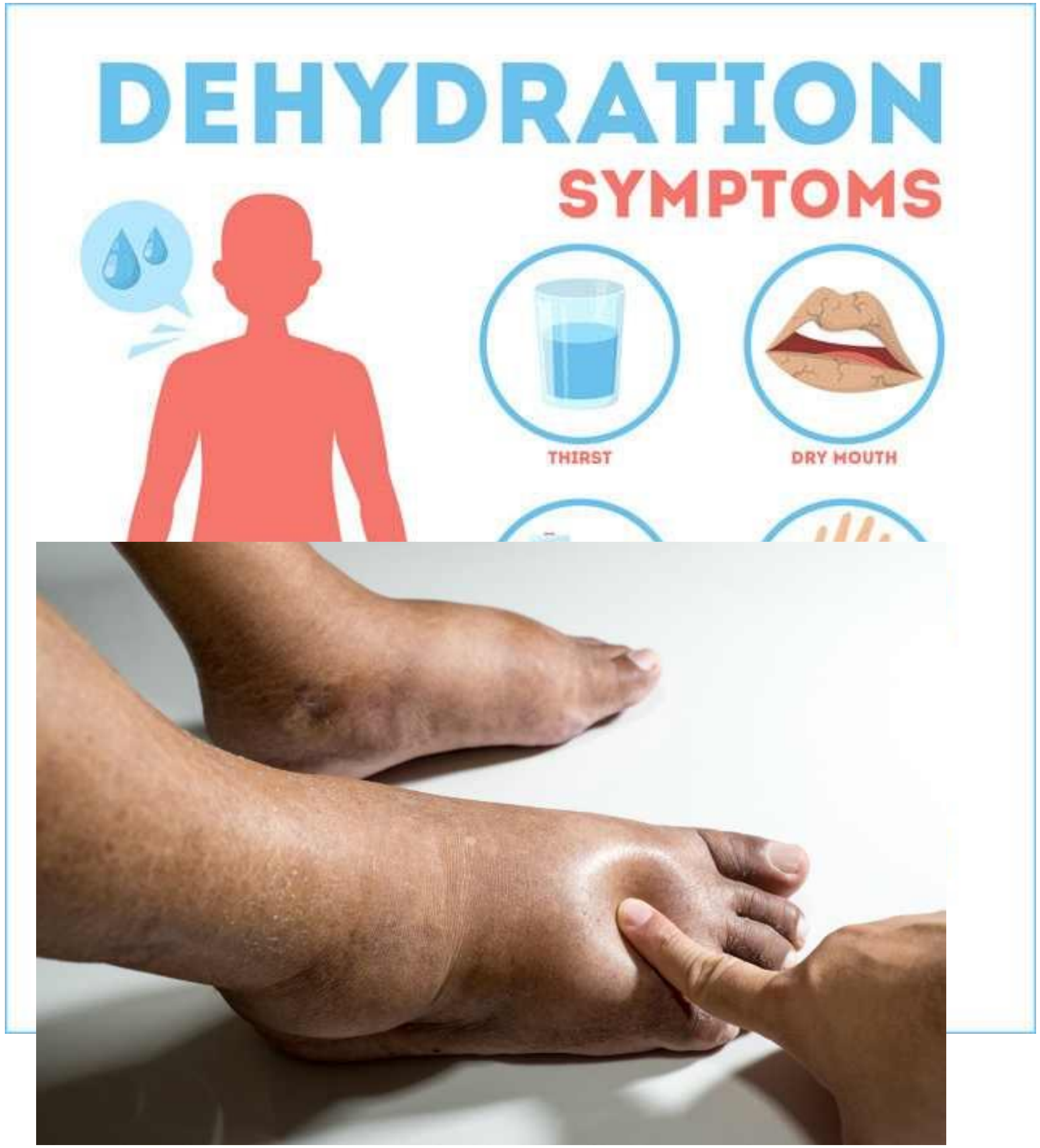
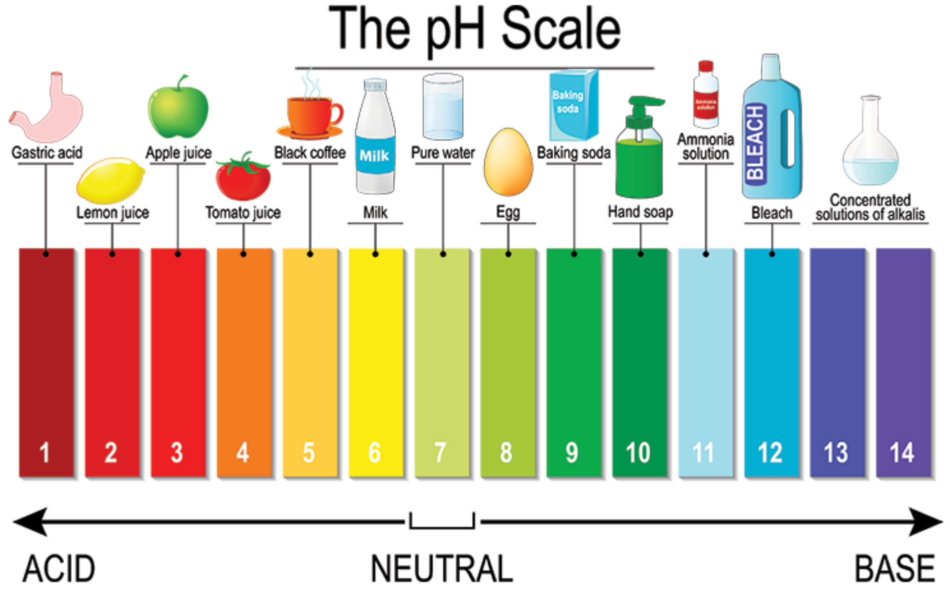
- Water balance must be preserved:
  - Loss >> intake --> dehydration
  - Intake >> loss --> edema



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)

$CH_3COOH + H_2O \text{ <--> } H_3O^+ + CH_3COO^-$

CH_3COO^- conjugate base

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

$pK_a = -\log K_a$

$NH_3 + H_2O \text{ <--> } NH_4^+ + OH^-$

NH_4^+ conjugate acid

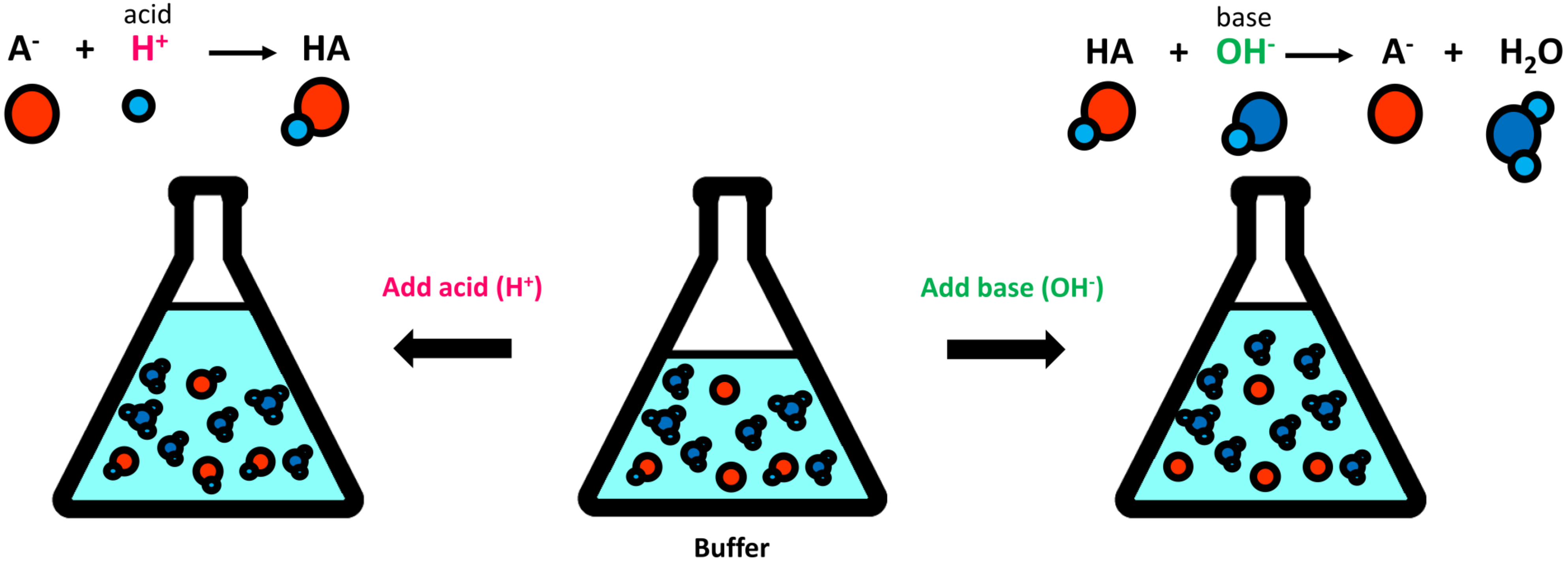
$K_b = [NH_4^+][OH^-]/[NH_3]$

$pK_b = -\log K_b$

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

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

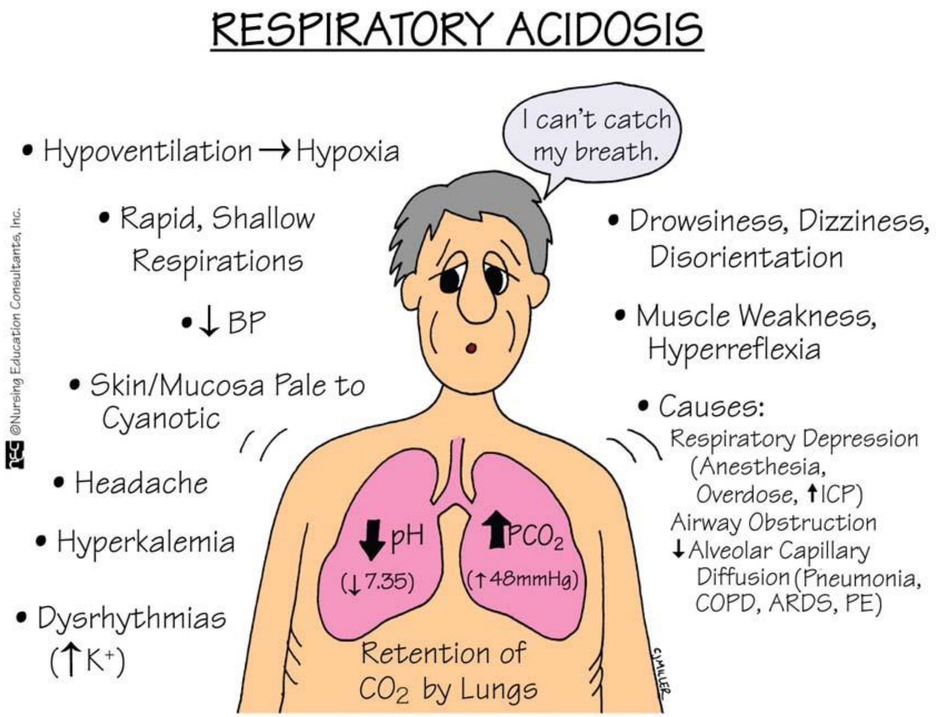


Identifying common physiological buffers

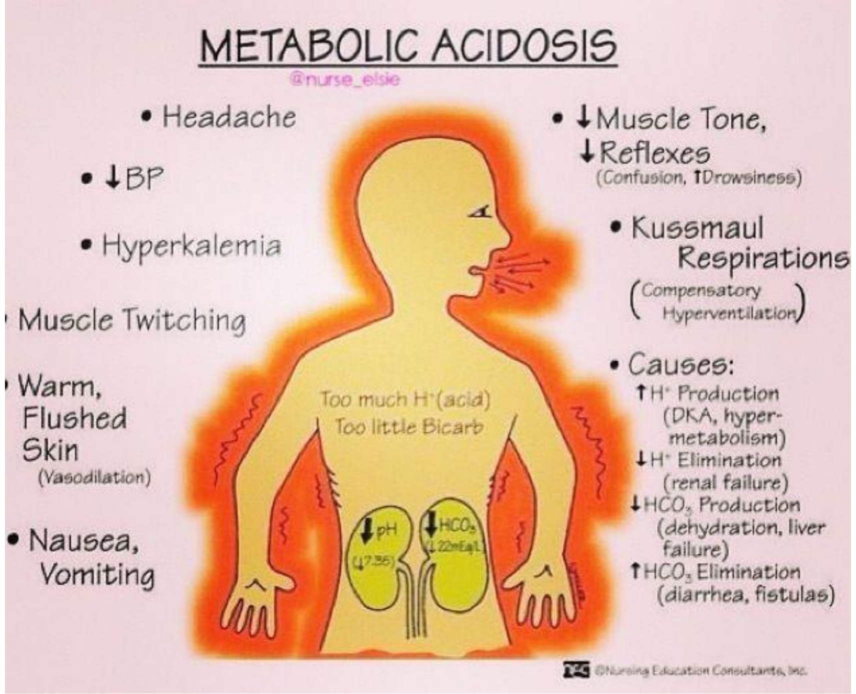
- Stomach pH = 1-2
- GI tract pH = 8-9
- Blood pH = 7.4
  - Blood pH < 7.2 --> pathological condition = acidosis
  - Blood PH < 6.8 --> death
  - Blood pH > 7.6 --> pathological condition = alkalosis
  - Blood pH > 8.6 --> death

Identifying common physiological buffers

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

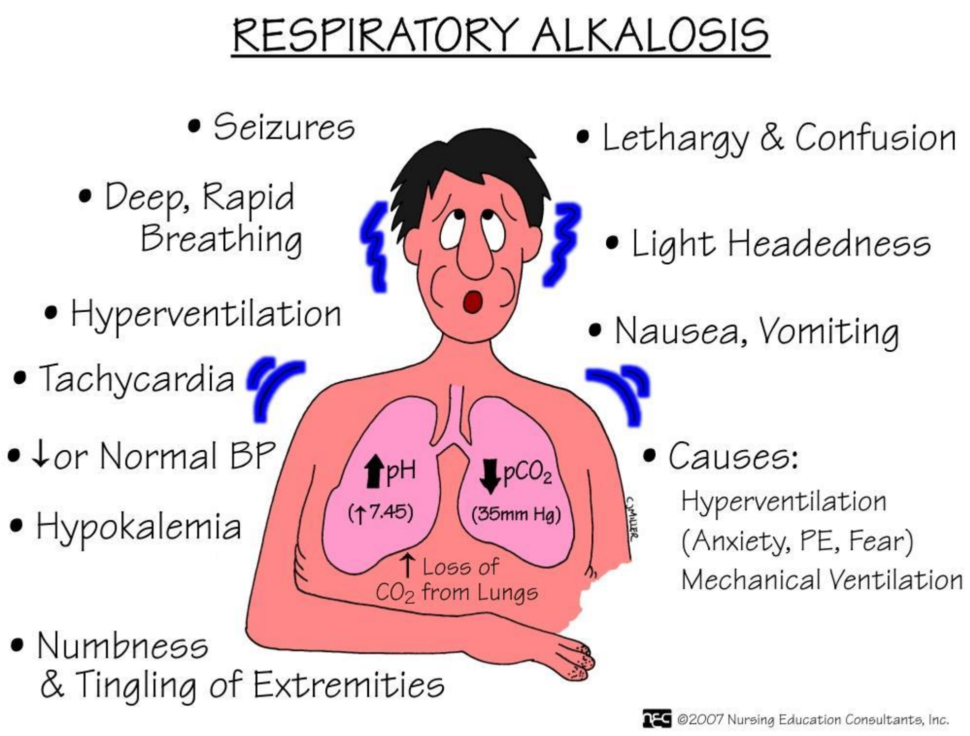


- Metabolic acidosis
  - Decreased concentration of HCO_3^-, results from various kidney diseases, uncontrolled diabetes, or vomiting of non-acid fluids

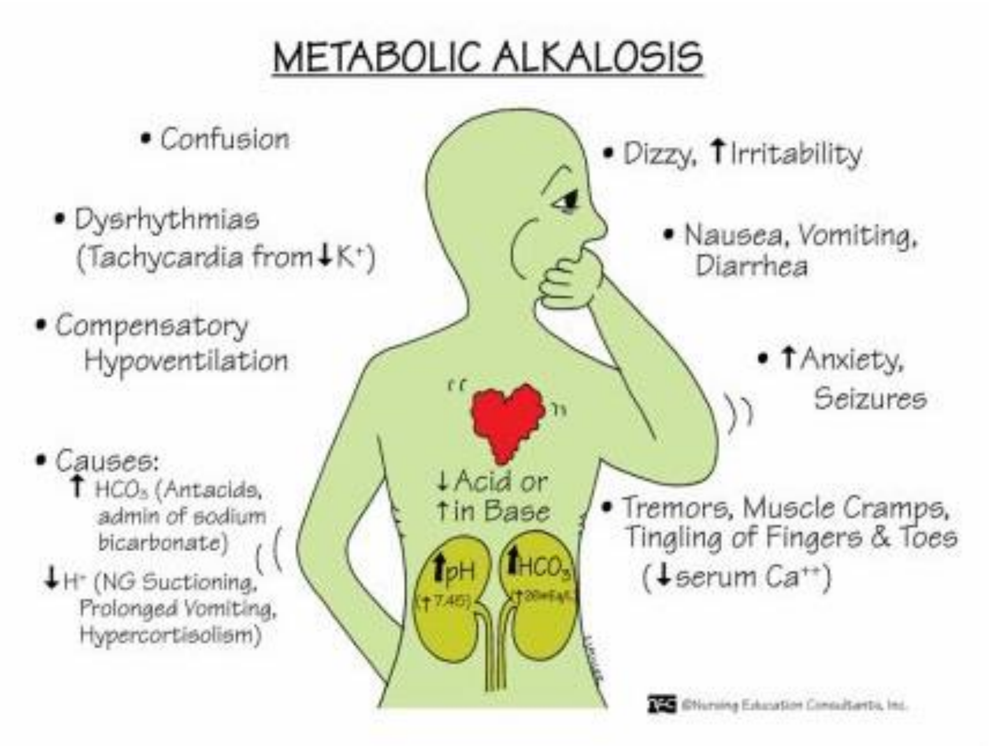


Identifying common physiological buffers

- 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:
    - $CO_2 + H_2O$ <--> H_2CO_3
    - H_2CO_3 <--> $H^+ + HCO_3^-$
- $K_{eq} = [H^+][HCO_3^-]/[CO_2]^* = 7.95 \times 10^{-7}$ --> $pK_{eq} = -\log K_{eq} = 6.1$
- Rearranging:

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

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

*$[H_2O]$ = 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