Equilibrium and Acid/Base

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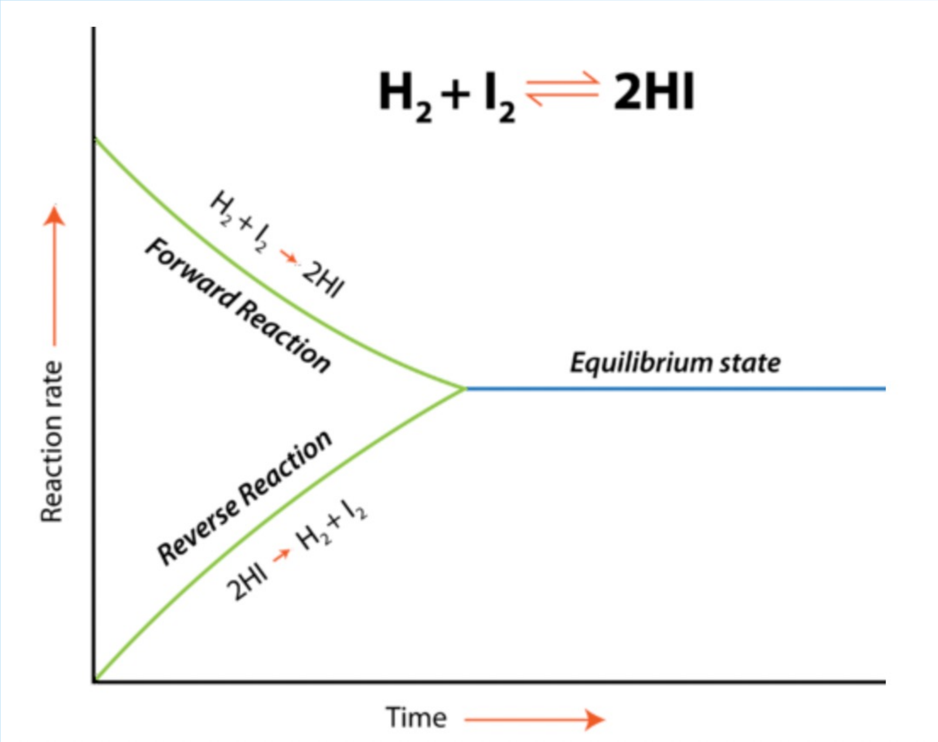
Equilibrium

- Forward reaction rate = backward reaction rate
- No net change concentration, constant

$$H_2\,(g) + I_2\,(g) \rightleftharpoons 2HI\,(g)$$

Forward reaction: $H_2\,(g) + I_2\,(g) \rightarrow 2HI\,(g)$

Reverse reaction: $2HI\,(g) \rightarrow H_2\,(g) + I_2\,(g)$



- Only solutes and gases!

Equilibrium constant

- Ratio between the concentrations of the products and the concentrations of the reactants
- Expression: $a\text{A} + b\text{B} \rightleftharpoons c\text{C} + d\text{D}$

$$K_{eq} = \frac{[\text{C}]^c[\text{D}]^d}{[\text{A}]^a[\text{B}]^b}$$

Example from last slide:

$$\text{H}_2\,(g) + \text{I}_2\,(g) \rightleftharpoons 2\text{HI}\,(g)$$

$$K_{eq} = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$$

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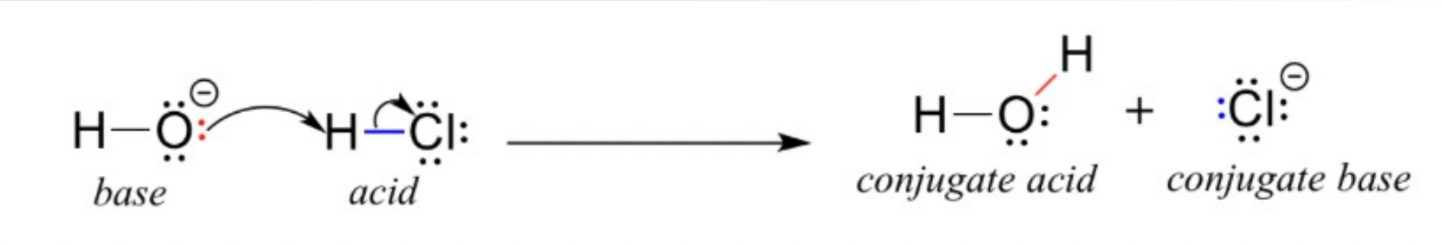
Brønsted-Lowry theory

- Acid = proton (H^+) donor
- Base = proton (H^+) acceptor

- Acid → conjugate base
- Base → conjugate acid

→ example: CH_3COOH/CH_3COO^-



Disassociation

- K_a/K_b – measurement of the ability of the acid/base have to donate/accept protons

Acid expression:

$$HA_{(aq)} \rightleftharpoons A^-_{(aq)} + H^+_{(aq)}$$

$$K_a = \frac{[A^-][H^+]}{[HA]}$$

Base expression:

$$B_{(aq)} + H_2O_{(l)} \rightleftharpoons HB^+_{(aq)} + OH^-_{(aq)}$$

$$K_b = \frac{[HB^+][OH^-]}{[B]}$$

Strong acid/base: high K_a/K_b
Weak acid/base: low K_a/K_b

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Formulas to remember

$$pH + pOH = 14$$

pH of strong acid

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

pH of weak acid

$$pH = -\log\sqrt{K_a \times C_a}$$

pOH of strong base

$$pOH = -\log[OH^-]$$

pOH of weak base

$$pOH = -\log\sqrt{K_b \times C_b}$$

1. The pH of 0.001 M NaOH solution equals to: **A.** 13 **B.** 11 **C.** 7 **D.** 4 **E.** 3

$NaOH \rightarrow Na^+ + OH^-$

$[OH^-] = 0,001\ M$

$pOH = -\log [OH^-]$

$= -\log (0,001)$

$pOH = 3$

$pH + pOH = 14$

$pH + 3 = 14$

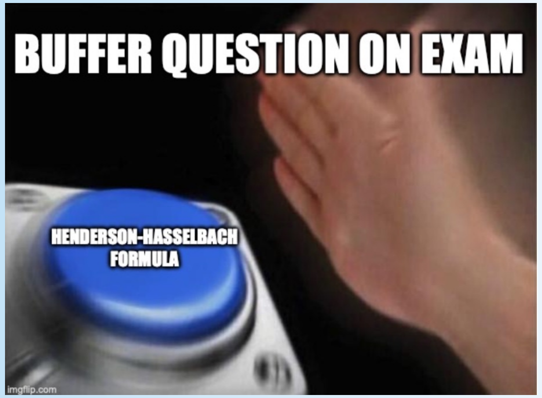
$pH = 11$

Buffers - concept

- Solution with a weak acid/conjugate base or weak base/conjugate acid
- Ability to resist changes in pH when acid or base is added
- UWAGA: physiological buffers in our body!

- Henderson-Hasselbach equation:

$$pH = pK_a + \log \frac{[base]}{[acid]}$$



Buffers

- Buffer concentration = $C_a + C_b$
- Buffer capacity = quantity of strong acid/base that must be added to change the pH of 1L of the solution by one pH unit
- → more concentrated = larger capacity to resist change

$$\beta = \frac{\Delta n}{\Delta pH}$$

n – moles added of H+ / OH- to 1L buffer

change caused by the addition

- Buffer range = pKa +/- 1

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6. What is the pH of acetate buffer containing 0.05 moles of acetic acid (CH_3COOH) and 0.09 moles of sodium acetate (CH_3COONa) in 2L of buffer? What is the buffer concentration? What would be the change in pH if 1 mL of 8M NaOH was added to above buffer? For acetic acid $K_a = 1.75\times 10^{-5}$. What is the buffer capacity towards bases?

Physiological buffers - bicarbonate system

- Maintain pH in blood → metabolic function

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

- Excess H^+ → CO_2 exhales
- Dec. H^+ → Shifts to right

Important Normal Values on ABG

pH	7.35	-	7.45
pCO_2	35 mmHg	-	45 mmHg
pO_2	75 mmHg	-	100 mmHg
HCO_3^-	22 mEq/L	-	26 mEq/L
O_2 Sat	Greater than 95%		

Acid/Base disorders

- **Alkalosis - high pH**
- **Acidosis - low pH**

- Respiratory – PaCO2 - acidic
- Metabolic – HCO_3^- - basic

ABG	pH	PaCO2	HCO3
Respiratory Acidosis	↓	↑	normal
Respiratory Alkalosis	↑	↓	normal
Metabolic Acidosis	↓	normal	↓
Metabolic Alkalosis	↑	normal	↑

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