Equilbilbrium of Acids and Bases

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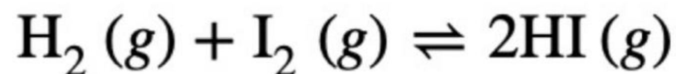
Equilibrium

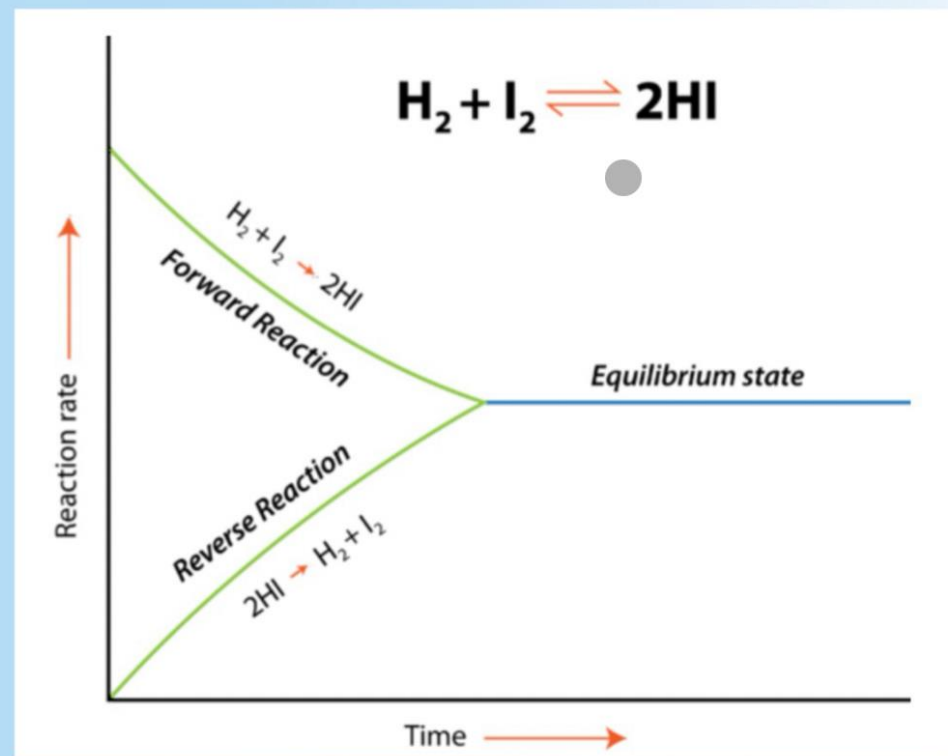
- Forward reaction rate = backward reaction rate
- No net change concentration, constant



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

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

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



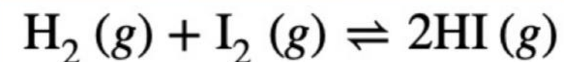
- Only solutes and gases!

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:



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

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

Brønsted-Lowry Theory

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

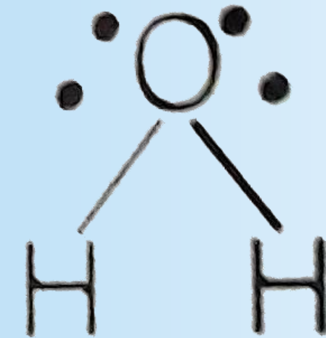
Acid —> conjugate base
Base —> conjugate acid

—> example: CH3COOH (weak acid) /CH3COO- (weak conjugate base)

By the way…is water an acid or a base? It's amphoteric (both)

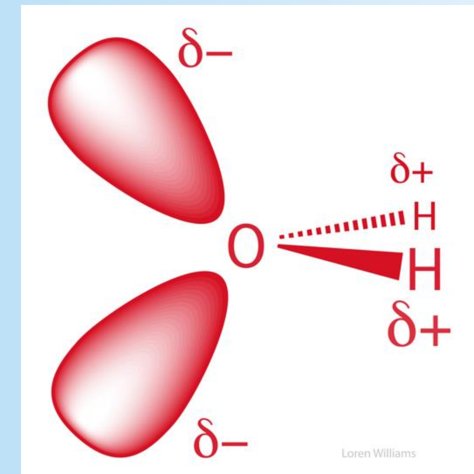
Key properties of water:

- 2 lone electron pairs
- **4 charge centers (2 bonding, 2 non-bonding)**
- Negative and positive dipole



Summary of Brønsted-Lowry and Lewis Theory on Water

Theory	Water as an Acid	Water as a Base
Brønsted-Lowry	Water donates a proton (H^+) to form OH^-. Example: $H_2O + NH_3 \rightarrow OH^- + NH_4^+$.	Water accepts a proton (H^+) to form H_3O^+. Example: $HCl + H_2O \rightarrow H_3O^+ + Cl^-$.
Lewis Theory	Water acts as a Lewis acid by accepting an electron pair. Example: $Cu^{2+} + 4H_2O \rightarrow [Cu(H_2O)_4]^{2+}$.	Water acts as a Lewis base by donating an electron pair. Example: $H^+ + H_2O \rightarrow H_3O^+$.



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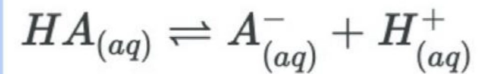
Disassociation

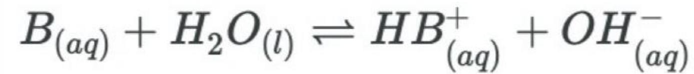
- K_a/K_b – measurement of the ability of the acid/base have to donate/accept protons

Acid expression:



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

Base expression:



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

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

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

pOH of strong base

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

pH of weak acid

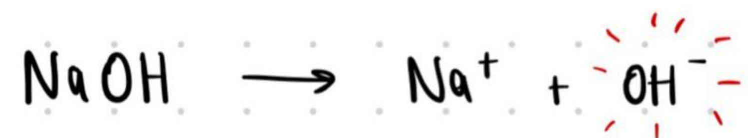
$$pH = -\log\sqrt{Ka \times Ca}$$

pOH of weak base

$$pOH = -\log\sqrt{Kb \times Cb}$$

The pH of a 0.001M solution of NaOH is equal to?

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



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

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

$= -\log(0{,}001)$

$pOH = 3$

$pH + pOH = 14$

$pH + 3 = 14$

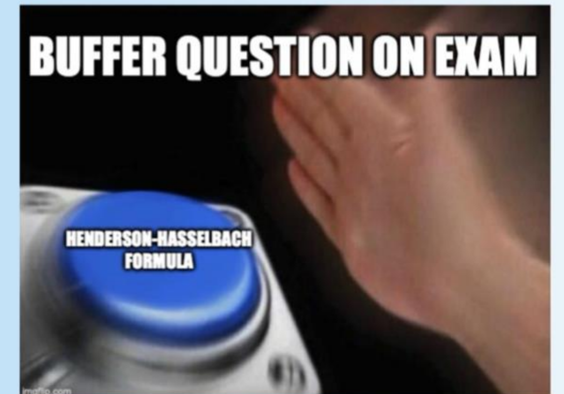
$pH = 11$

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:

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



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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

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?

(4) Calculate pH of the buffer

1) HH equation

$$pH = pKa + \log \frac{[A^-]}{[HA]}$$

$K_a = 1.75 \times 10^{-5}$

$pKa = -\log Ka = -\log(1.75 \times 10^{-5})$

≈ 4.76

2) Calculate concentrations of w. acid (HA) and w. conjugate base (A^-)

HA

$M = \frac{mol}{vol\ (L)}$ (mol/L)

$M = \frac{0.05\ mol}{2L}$

$M = 0.025\ mol/L$

A^-

$M = \frac{mol}{vol\ (L)}$ (mol/L)

$= \frac{0.09\ mol}{2L}$

$= 0.045M$

3) Plug missing values back into HH to solve equation

$HH = 4.76 + \log \frac{(0.045)}{(0.025)}$

$= 4.76 + 0.26$

$= 5.02$

Buffer Concentration

Buffer concentration = total concentration of buffer components

already calculated conc. of w. acid + w. conj. base (see prev. step)

Buffer concentration = 0.025 M + 0.045 M

= 0.07 M

Change in pH after adding 1ml of NaOH

NaOH (strong base) will react w/ the w. base and increase the the moles of conj. base. On the other hand NaOH will react w/ the w. acid and decrease its moles.

New total moles of base :

moles of NaOH + moles of conj. base (already calculated in 1st step)

moles of NaOH

$$8M \text{ NaOH} = \frac{8 \text{ mol NaOH}}{L} \times \frac{L}{1000 ml} = 0.008 \text{ mol NaOH}$$

Total moles of base (A^-)

$$A^- = 0.09 \text{ mol} + 0.008 \text{ mol}$$
$$= 0.098 \text{ mol}$$

Total moles of acid (HA)

$$HA = 0.05 \text{ mol} - 0.008$$
$$= 0.042 \text{ mol}$$

With new values calculate the change in pH when 0.008 mol NaOH is added

$$HH = pH = pKa + \log \frac{[A^-]}{[HA]}$$

$$= 4.76 + \log \frac{\left(\frac{0.098 mol}{2 L}\right)}{\left(\frac{0.042 mol}{2L}\right)}$$

$$= 4.76 + \log(2.333)$$

$$= 4.76 + 0.37$$

New pH = 5.13

Calculate change in pH before and after NaOH was added (ΔpH)

$$\Delta pH = 5.13 - 5.02$$

$$= 0.11$$

Buffer Capacity towards bases

$$\beta = \frac{\Delta n \text{ (moles of strong base added)}}{\Delta pH \times Vol(L)}$$

$$= \frac{0.08}{0.11 \times 2L}$$

$$= 0.036 \text{ mol/L per pH unit}$$

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%		

Quiz ☺

How to participate?

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