SOLUTIONS OF ELECTROLYTES

Electrolytes and Non-electrolytes

Compounds

Electrolytes - compounds whose aqueous solutions conduct electricity

- salts
- bases
- acids

Nonelectrolytes – compounds whose aqueous solutions don't to conduct electricity

Some ionic compounds are not soluble in water, but if their molten state conducts electricity, they are also classified as electrolytes.

Why solutions of electrolytes conduct electricity?

According to theory of electrolytic dissociation, when electrolytes dissolve in water, they decompose (dissociate) into positively charged ions (cations) and negatively charged ions (anions).

Acids are electrolytes that produce cation of hydrogen in aqua solutions.

$$H_xA = xH^+ + A^{x-}$$

$$HCl = H^+ + Cl^-$$

$$HCN \Leftrightarrow H^+ + CN^-$$

$$\text{I. } H_2CO_3 \Leftrightarrow H^+ + HCO_3^-$$

$$\text{II. } HCO_3^- \Leftrightarrow H^+ + CO_3^{2-}$$

<u>Bases</u> are electrolytes that produce of hydroxide anion in aqua solutions.

$$Me(OH)_x = Me^{x+} + xOH^-$$

$$NaOH = Na^+ + OH^-$$

$$NH_4OH \Leftrightarrow NH_4^+ + OH^-$$

$$\text{I. } Ba(OH)_2 = BaOH^+ + OH^-$$

$$\text{II. } BaOH^+ = Ba^{2+} + OH^-$$

Salts are electrolytes that produce of metal cation and acid anion in aqua solutions.

$$Me_yA_x = yMe^{x+} + xA^{y-}$$

$$NaCl = Na^{+} + Cl^{-}$$

$$(NH_4)_2SO_4 = 2NH_4^{+} + SO_4^{2-}$$

$$Al_2(SO_4)_3 = 2Al^{3+} + 3SO_4^{2-}$$

DEGREE OF DISSOCIATION α

$$\alpha = \frac{n}{N}$$

n – number of molecules of electrolytes that dissociate in the given solutions.

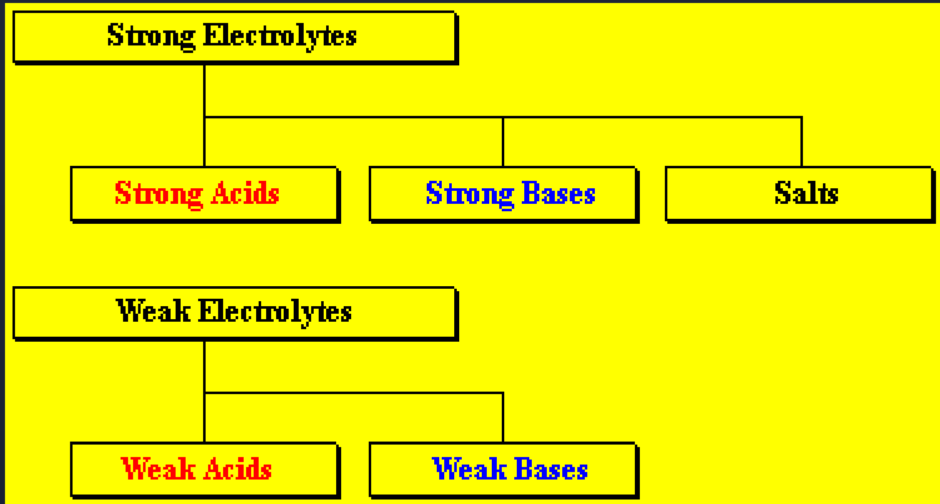
N – total number of molecules of electrolytes in the given solution.

α depends on : nature electrolyte and solvent,
temperature,
concentration of electrolyte

Depending on the degree of dissociation, electrolytes are divided into

Strong electrolytes
dissociate completely

Weak electrolytes
dissociate incompletely



Strong Electrolytes	strong acids	HCl, HBr, HI, HNO_3, $HClO_3$, $HClO_4$, and H_2SO_4
	strong bases	NaOH, KOH, LiOH, $Ba(OH)_2$, and $Ca(OH)_2$
	salts	NaCl, KBr, $MgCl_2$, and many, many more
Weak Electrolytes		
	weak acids	HF, $HC_2H_3O_2$ (acetic acid), H_2CO_3 (carbonic acid), H_3PO_4 (phosphoric acid), and many more
	weak bases	NH_3 (ammonia), C_5H_5N (pyridine), and several more, all containing "N"

DISSOCIATION CONSTANT

K_d depends on : nature electrolyte and solvent, temperature,

<u>**But does not depend on concentration of electrolyte**</u>

W. Ostwald's dilution law $K_d = K_a$ or K_b

a- acid b- base

$$\alpha = \sqrt{\frac{K_d}{C}}$$

$$HF \Leftrightarrow H^+ + F^-$$

$$[H^+] = [F^-] = \alpha \cdot C$$

$$K_a = \alpha^2 \cdot C$$

$$[HF] = C - \alpha \cdot C$$

$$K_a \Leftrightarrow \frac{\alpha C \cdot \alpha C}{C - \alpha} = \frac{C \cdot \alpha^2 C}{C(1-\alpha)} = \frac{\alpha^2 C}{1-\alpha}$$

Ionic equations of chemical reactions between electrolytes

Molecular equation

$$2KOH + CuCl_2 = 2KCl + Cu(OH)_2 \downarrow$$

Complete ionic equation

$$2K^+ + 2OH^- + Cu^{2+} + 2Cl^- = 2K^+ + 2Cl^- + Cu(OH)_2 \downarrow$$

Net ionic equation

$$Cu^{2+} + 2OH^- = Cu(OH)_2 \downarrow$$

Ionic equations of chemical reactions between electrolytes

Molecular equation

$$KOH + HNO_3 = KNO_3 + H_2O$$

Complete ionic equation

$$K^+ + OH^- + H^+ + NO_3^- = K^+ + NO_3^- + H_2O$$

Net ionic equation

$$H^+ + OH^- = H_2O$$

Ionic equations of chemical reactions between electrolytes

Molecular equation

$$K_2CO_3 + 2HCl = 2KCl + H_2O + CO_2\uparrow$$

Complete ionic equation

$$2K^+ + CO_3^{2-} + 2H^+ + 2Cl^- = 2K^+ + 2Cl^- + H_2O + CO_2\uparrow$$

Net ionic equation

$$2H^+ + CO_3^{2--} = +H_2O + CO_2\uparrow$$

Ion product of water

$$H_2O \Leftrightarrow H^+ + OH^-$$

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

$$K\text{д} = \frac{[H^+]\cdot[OH^-]}{[H_2O]}$$

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

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

pH

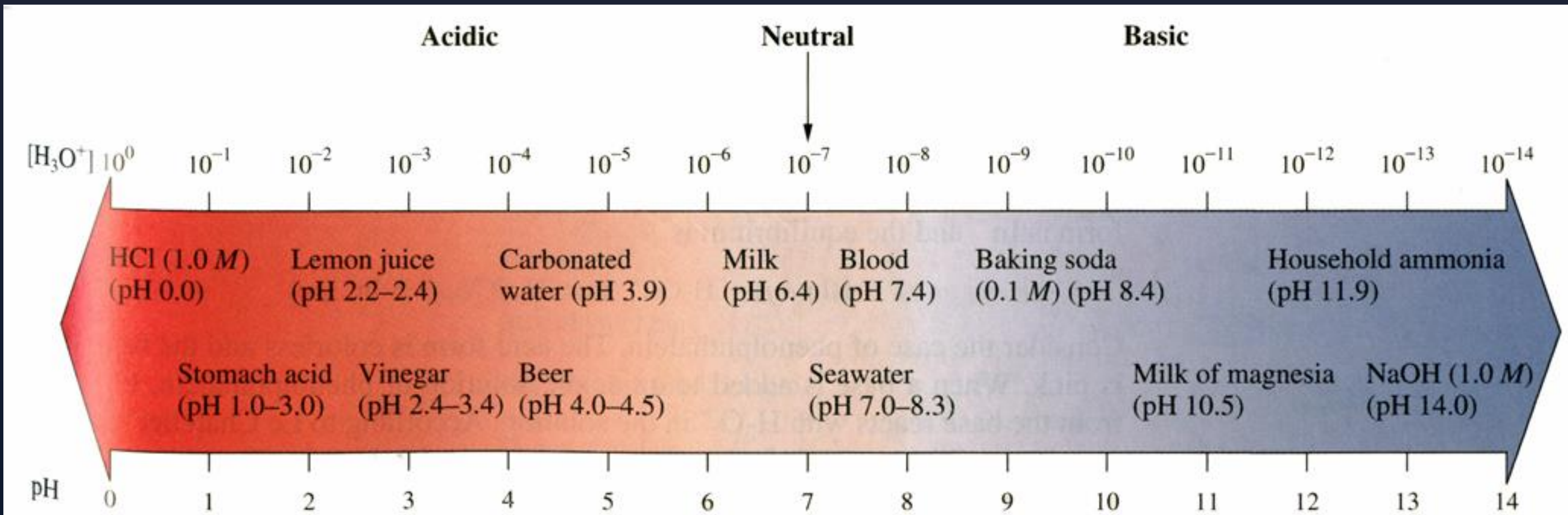
$$pH = -\log[H^+]$$

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

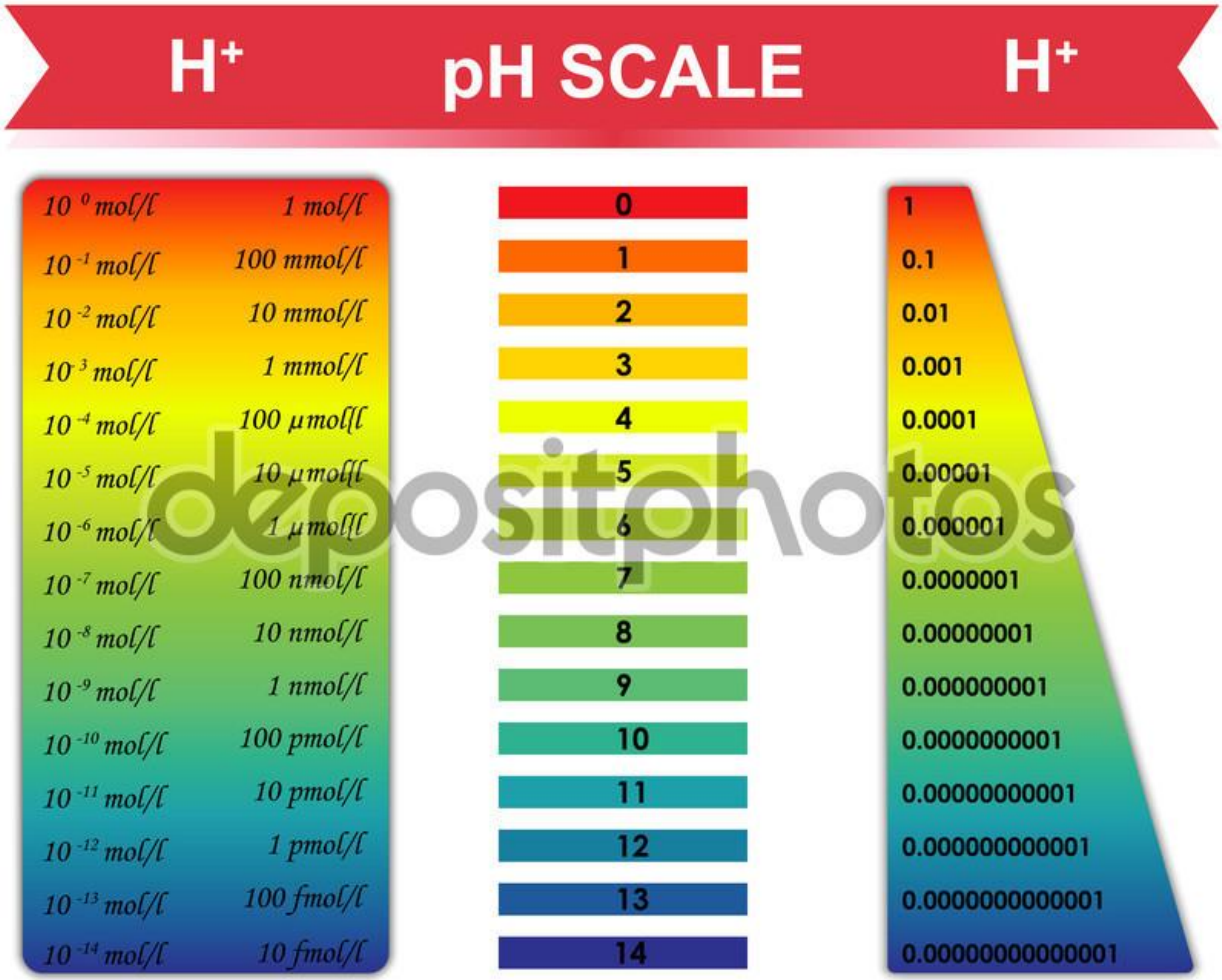
$$pH + pOH = 14$$

$$pH = 14 - pOH$$

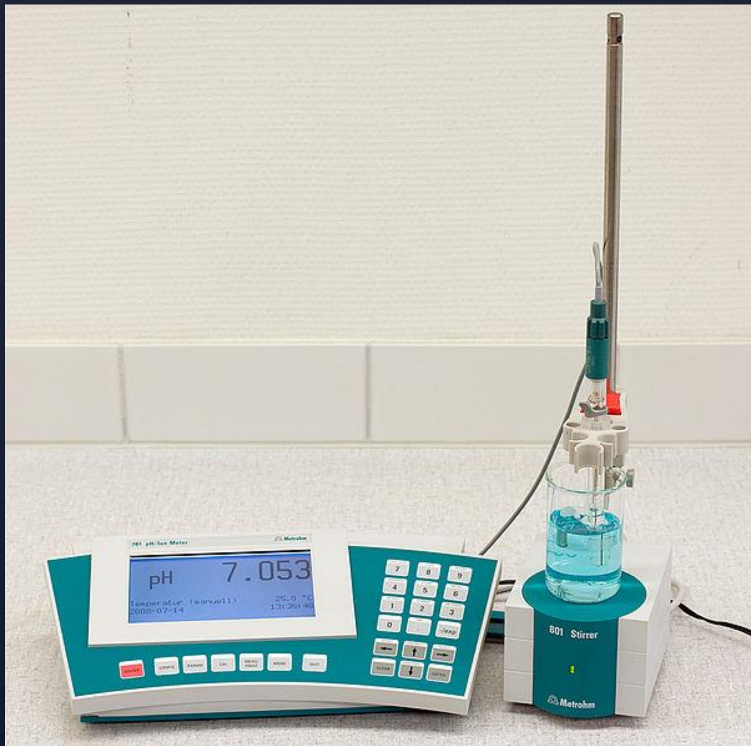
$$[H^+] = \sqrt{K_\text{д} \cdot C}$$



Acidic
Neutral
Basic
[H3O+] 10^0 10^-1 10^-2 10^-3 10^-4 10^-5 10^-6 10^-7 10^-8 10^-9 10^-10 10^-11 10^-12 10^-13 10^-14
HCl (1.0 M) (pH 0.0)
Lemon juice (pH 2.2–2.4)
Carbonated water (pH 3.9)
Milk (pH 6.4)
Blood (pH 7.4)
Baking soda (0.1 M) (pH 8.4)
Household ammonia (pH 11.9)
Stomach acid (pH 1.0–3.0)
Vinegar (pH 2.4–3.4)
Beer (pH 4.0–4.5)
Seawater (pH 7.0–8.3)
Milk of magnesia (pH 10.5)
NaOH (1.0 M) (pH 14.0)
pH 0 1 2 3 4 5 6 7 8 9 10 11 12 13 14

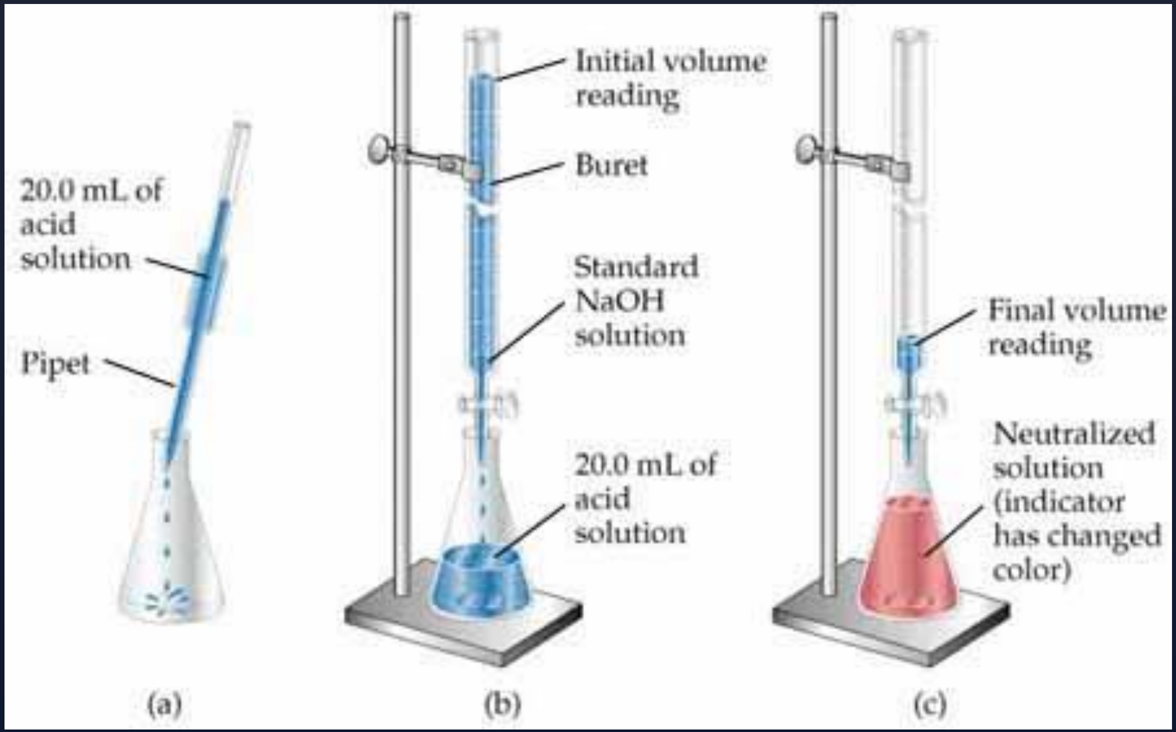


depositphotos


pH
7.053
801 Stirrer
Metrohm

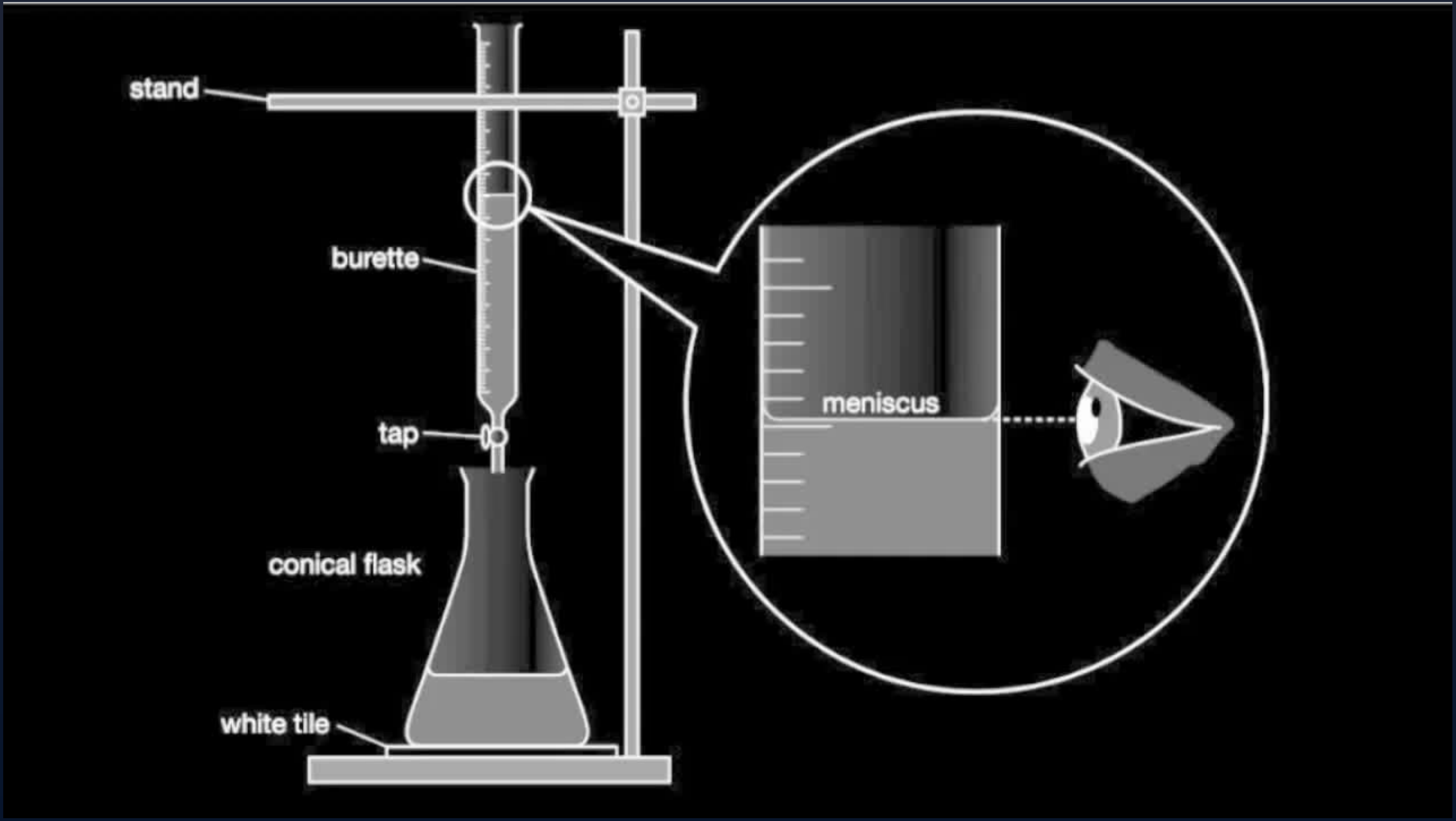

14
1
2
12
3
10
4
9
5
8
7
6
TEST PAPERS
UNIVERSAL
INDICATOR PAPER
pH1 - pH14

Titration



$$C_{N1}V_1=C_{N2}V_2$$

Titration



Hydrolysis of salts

Salt do not take a part in the hydrolysis

pH=7

NaCl

Cation hydrolysis

pH<7

$CuCl_2$

$$K_h = \frac{K_w}{K_b}$$

Anion hydrolysis

pH>7

Na_2CO_3

$$K_h = \frac{K_w}{K_a}$$

Cation and anion hydrolysis

pH≈7

CH_3COONH

$$K_h = \frac{K_w}{K_a K_b}$$

$$CuCl_2 = Cu^{2+} + 2Cl^-$$

$$Cu(OH)_2 \qquad HCl$$

Weak base **Strong acid**

Cation Hydrolysis

pH<7

$$I. Cu^{2+} + HOH \Leftrightarrow CuOH^+ + H^+$$

$$Cu^{2+} + 2Cl^- + HOH \Leftrightarrow CuOH^+ + 2Cl + H^+$$

$$CuCl_2 + H_2O \Leftrightarrow CuOHCl + HCl$$

$$Na_2CO_3 = 2Na^+ + CO_3^{2-}$$

$$NaOH \quad H_2CO_3$$

Strong base → **pH>7**

Weak acid → **Anion Hydrolysis**

$$I. CO_3^{2-} + HOH \Leftrightarrow HCO_3^- + OH^-$$

$$2Na^+ + CO_3^{2-} + HOH \Leftrightarrow 2Na^+ + HCO_3^- + OH^-$$

$$Na_2CO_3 + H_2O \Leftrightarrow NaHCO_3^- + NaOH$$

$$CH_3COONH_4 = CH_3COO^- + NH_4^+$$

$$CH_3COOH \quad NH_4OH$$

Weak acid **Weak base**

Cation and anion Hydrolysis

pH≈7

$$CH_3COONH_4 + HOH \Leftrightarrow CH_3COOH + NH_4OH$$

$$CH_3COONH_4 + H_2O \Leftrightarrow CH_3COOH + NH_4OH$$

Complete hydrolysis

$$Al_2S_3 + 6H_2O = 2Al(OH)3\downarrow + 3H_2S\uparrow$$

Combine hydrolysis

$$Al_2(SO_4)_3 + 3Na_2CO_3 + 3H_2O = 2Al(OH)3\downarrow + 3CO_2\uparrow + 3Na_2SO_4$$

Buffer solution is one which resists changes in pH when small quantities of an acid or an alkali are added to it.

Buffer solutions are used as a means of keeping pH at a nearly constant value in a wide variety of chemical applications.

In nature, there are many systems that use buffering for pH regulation. For example, the bicarbonate buffering system is used to regulate the pH of blood.

Buffers usually consist of:

weak acid and its **salt**

CH_3COOH / CH_3COONa

weak base and its **salt**

NH_4OH / NH_4Cl

Why pH of buffer solutions do not change in result adding of some strong acid or base?

Weak electrolyte → $CH_3COOH \Leftrightarrow CH_3COO^- + H^+$

Strong electrolyte → $CH_3COONa \Rightarrow CH_3COO^- + Na^+$

$$K_a = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]} \Rightarrow$$

$$K_a = \frac{Csalt \cdot [H^+]}{Cacid} \Rightarrow$$

$[CH_3COO^-]$ - C salt

$[CH_3COOH]$ - C acid

$$[H^+] = K_a \frac{C_{acid}}{C_{salt}}$$

$[H^+]$

1. Depends on ratio of concentrations of acid and salt

2. Depends on Ka of acid

$$[H^+] = K_a \frac{C_{acid}}{C_{salt}}$$

$$pH = pK_a - \log\frac{C_{acid}}{C_{salt}}$$

$$[OH^-] = K_b \frac{C_{base}}{C_{salt}}$$

$$pOH = pK_b - \log\frac{C_{base}}{C_{salt}}$$

$$pH = 14 - (pK_b - \log\frac{C_{base}}{C_{salt}})$$

Henderson–Hasselbalch equation

What is the mechanism of the buffer action?

Add small amount of strong acid

$+ H^+$

acid

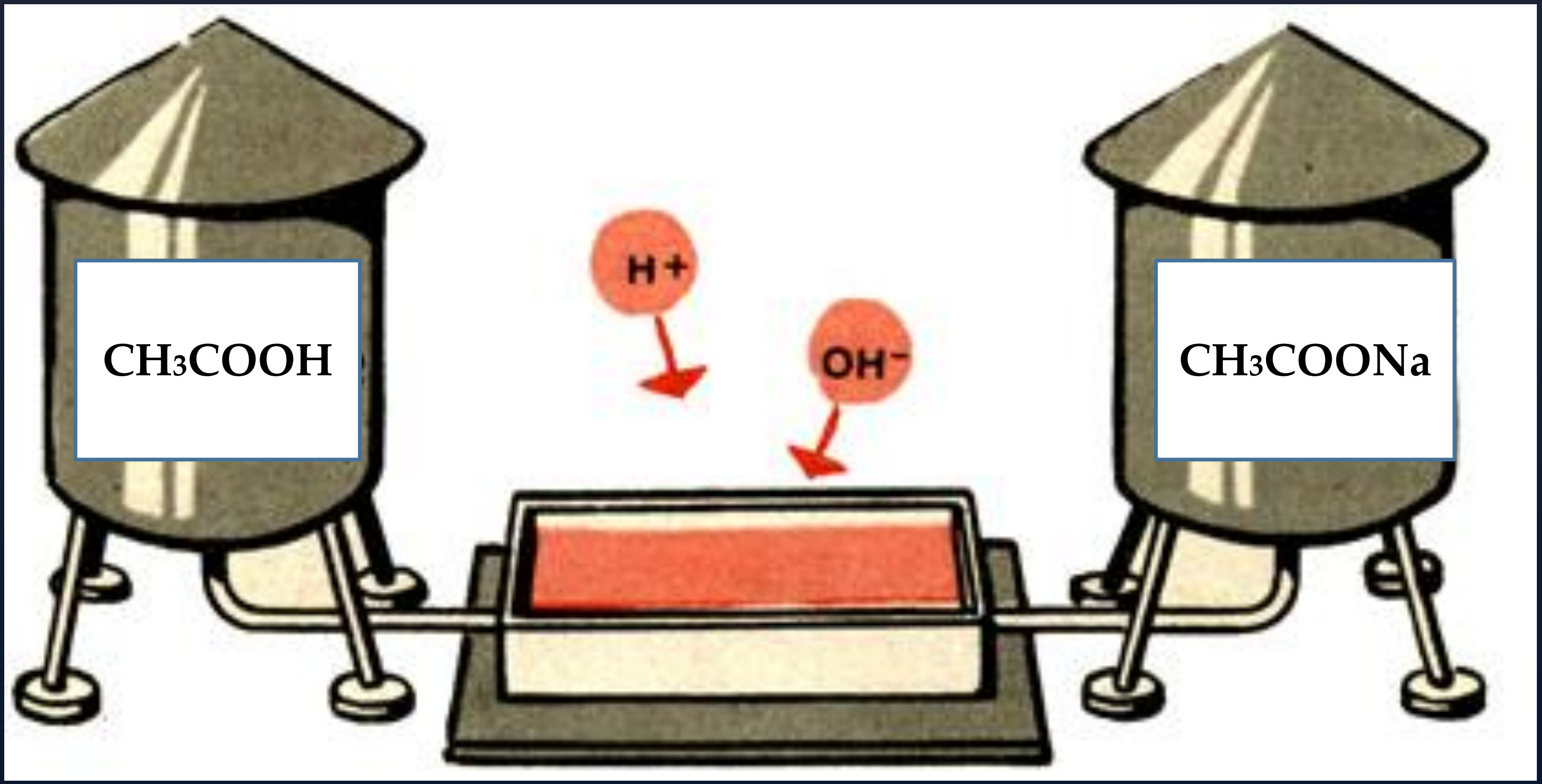
$CH_3COO^- + H^+ = CH_3COOH$

Add small amount of strong base

$+ OH^-$

base

$OH^- + H^+ = H_2O$



CH3COOH
H+
OH-
CH3COONa

Buffer solution is able to retain almost constant pH when small amount of acid/base is added.
Quantitative measure of this resistance to pH changes is called buffer capacity.

$$B = \frac{C \cdot V}{\Delta pH \cdot Vbuf.}$$

B – buffer capacity

C – concentration of acid or base, mol-eq/l

V buf. – volume of buffer solution, l

V – volume of adding electrolyte solution, l

ΔpH – changing of pH

Buffer capacity depends on concentrations of components of and its ratio.
Buffer capacity is maximal, when the ratio of its components = 1:1

acid-base balance of blood

7,36

acidosis ←

7,25-7,44

alcalosis →

a condition characterized by an abnormal increase in the acidity of the blood and extracellular fluids

+ $NaHCO_3$ 4% solution

an abnormal increase in the alkalinity of the blood and extracellular fluids

+ ascorbinic acid 5% solution