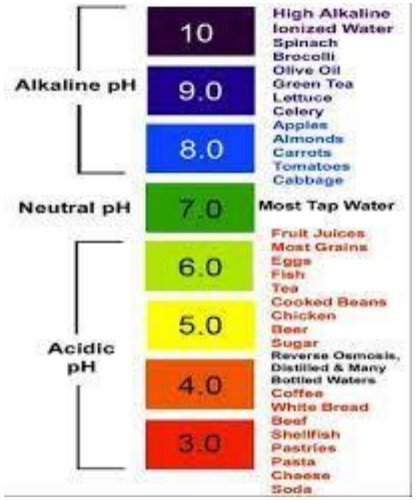
Unit 2 Topics

- Acids, Bases, & Buffers.
- Major intra & extracellular ELECTROLYTES.
- Dental Products



Acid, Base & Buffers

Acid, Base & Buffers

1. Types of Acid, Bases & Buffers
4. Uses
5. Theories of Acid, Bases & Buffers
6. Examples
7. References

Buffer or Buffer Solutions or Buffer Systems

- Solutions composed of a weak acid and one of its salts (e.g. CH3COOH + CH3COONa or CH3COONH4) or
- Weak base and one of its salts (e.g. NH4OH + NH4Cl).
- Purpose: to resist to a greater/lesser extent the changes in pH when some acid or base is added to them.

Types of buffers

- **Strong acid buffers**: concentration of oxonium ions ($H3O^+$) is high. E.g., Nitric acid
- **Strong base buffers:** checked using the equation: pH + pOH = 14
- **Weak acid buffers:** Buffer solutions with constant pH values of between 4 and 7 can be prepared from a weak acid and one of its salts.

$$CH_3COONa_{(s)} + aq \longrightarrow CH_3COO\text{-}_{(aq)} + Na^+_{(aq)}$$

- ethanoic acid is only partially ionised:

$$CH_3COOH_{(aq)} + H_2O \rightleftharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$$

$$H_3O^+_{(aq)} + OH^-_{(aq)} \rightarrow 2H_2O$$

- **Weak base buffers:** Buffer solutions with constant pH values between 7 and 10 can be prepared from a weak base and one of its salts.
- E.g..

- A solution of ammonium hydroxide and ammonium chloride is typically used.

$$NH_4Cl_{(aq)} \rightarrow NH_4^{+}{}_{(aq)} + Cl^{-}{}_{(aq)}$$

$$NH_{3(aq)} + H_2O_{(aq)} \rightleftharpoons NH_4^{+}{}_{(aq)} + OH^{-}{}_{(aq)}$$

- Uses of Buffers:

In standardizing pH meter

Buffering biological systems (e.g. blood)

Maintaining the pH of eye drops and shampoos etc

Buffer Equation

- Consider a buffer solution obtained by mixing a weak acid HA and its salt MA.
- The dissociation of HA is given as:

$HA \rightleftharpoons H^+ + A^-$ & $MA \rightleftharpoons M^+ + A^-$

- Degree of dissociation will depend on the value of the dissociation constant Ka.
- Greater the magnitude of Ka, greater is the dissociation of HA into H^+ & A^-

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

Note: Making the approximation that in dilute solution, activity is equal to concentration. This relationship exists in the mixture of HA and MA

- Therefore,

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

- Suppose the concentration of the acid is Ca and that of the salt is Cs. The concentration of the undissociated acid will be:

$$Ca - [H^+], \text{ i.e., } [HA] = Ca - [H^+]$$

- In this buffer solution, A^- ions will be produced from the dissociation of MA and also from that of HA.
- The concentration of MA [salt] is Cs, the conc. of A- ions will also be Cs.
- A molecule of HA on dissociation gives a H^+ ion and a A^- ion, Hence $[A^-] = [H^+]$. Thus
- The total conc. of A- ion in the solution will be Cs + $[H^+]$.

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

$$HA \leftrightharpoons H^+ + A^-$$

$$MA \leftrightharpoons M^+ + A^-$$

common ion

$$[H^+] = Ka \times \frac{Ca - [H^+]}{Cs + [H^+]}$$

- Due to common ion effect, the dissociation of weak acid HA will be further decreased.
- [H+] will be small and can be neglected in comparison to Cs and Ca.

$$[H^+] = Ka \times \frac{[Acid]}{[Salt]}$$

$$\log [H^+] = \log \left[Ka \times \frac{[Acid]}{[Salt]} \right]$$

$$-\log [H^+] = -\log Ka - \log \frac{[Acid]}{[Salt]}$$

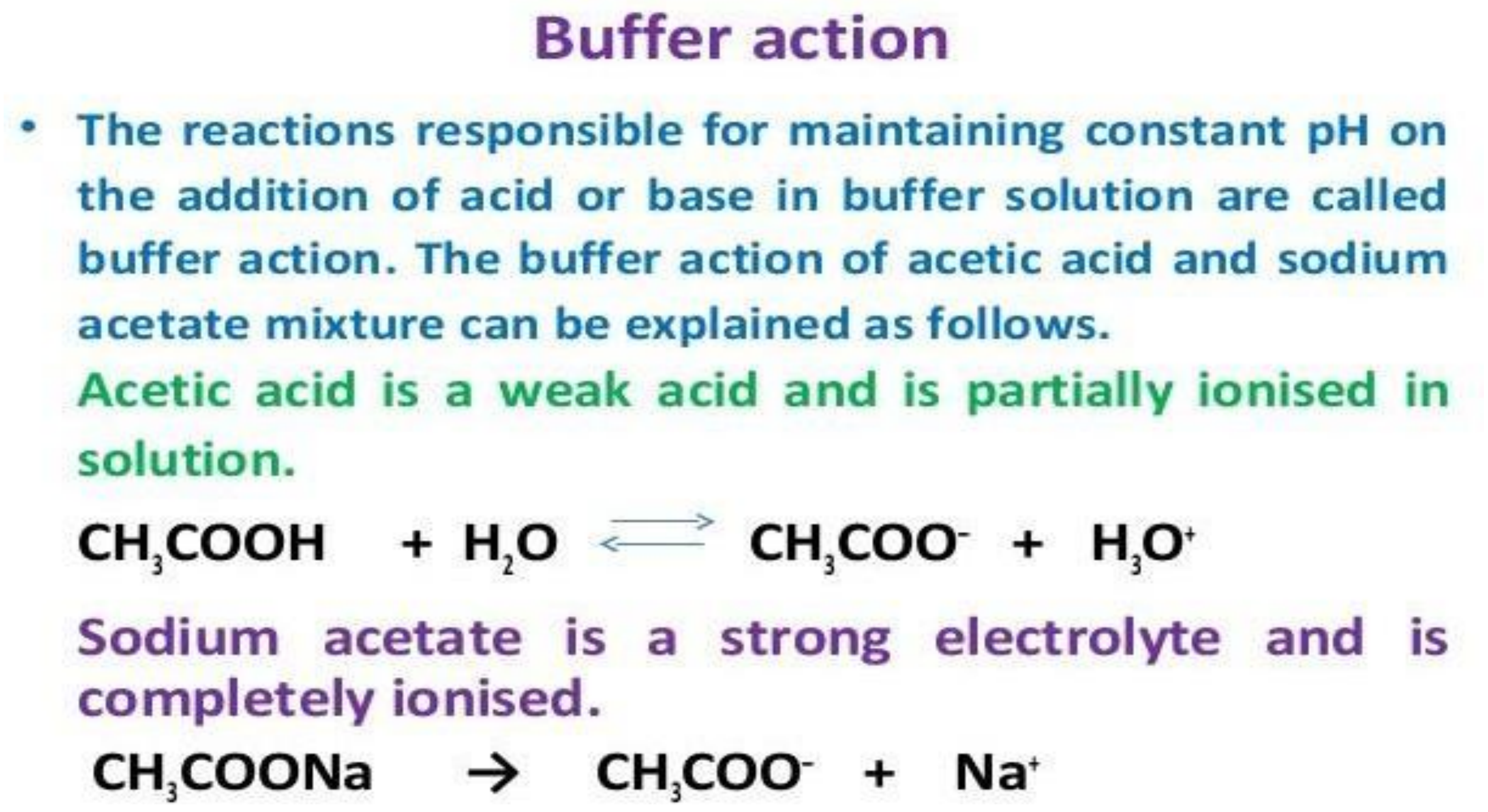
$$\log [H^+] = \log Ka + \log \frac{[Acid]}{[Salt]}$$

- - log is and /or can written as *p.*
- *Log a/b = -log b/a*

$$-\log[H^+] = -\log Ka - \log \frac{[Acid]}{[Salt]}$$

- $pH = pKa + \log \frac{[Salt]}{[Acid]}$
- *Similarly for a buffer solution containing a weak base and its salt, it can be shown that:*
- $pOH = pKb + \log [Salt] / [Base]$

Buffer Action or Buffer Capacity



Buffer action

- The reactions responsible for maintaining constant pH on the addition of acid or base in buffer solution are called buffer action. The buffer action of acetic acid and sodium acetate mixture can be explained as follows.

  Acetic acid is a weak acid and is partially ionised in solution.

  $$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$$

  Sodium acetate is a strong electrolyte and is completely ionised.

  $$CH_3COONa \rightarrow CH_3COO^- + Na^+$$

Buffer action

- When a small amount of acid is added to this mixture, H^+ ions of the acid combines with equal number of acetate ions to form acetic acid molecule. Therefore, the net pH remains constant.

$$CH_3COO^- + H^+ \rightarrow CH_3COOH + H_2O$$

When a small amount of base is added, the OH^- ions of the base neutralise acid molecules to form sodium acetate. Therefore, the net pH remains constant.

$$CH_3COOH + OH^- \rightarrow CH_3COO^- + H_2O$$

Buffer Capacity

- The resistance to changes of pH now remains to be discussed in a more quantitative way.
- The magnitude of the resistance of a buffer to pH changes is referred to as the buffer capacity, β. It is also known as buffer efficiency, buffer index, and buffer value.
- Koppel, Spiro and Van Slyke introduced the concept of buffer capacity and defined it as:
- the ratio of the increment of strong base (or acid) to the small change in pH brought about by this addition.

- For the present discussion, the approximate formula.

$$\beta = \frac{\Delta B}{\Delta \mathrm{pH}}$$

- ✓ delta, Δ: a finite/limited change
- ✓ ΔB is the small increment in gram equivalents (g Eq)/liter of strong base added to the buffer solution.
- ✓ ΔpH: change in a pH

- According to this equation, the buffer capacity of a solution has a value of 1 when the addition of 1 g Eq of strong base (or acid) to 1 liter of the buffer solution results in a change of 1 pH unit.

- The changes in concentration of the salt and the acid by the addition of a base are represented using the modified form of buffer equation

$$pH = pK_a + \log \frac{[\text{Salt}] + [\text{Base}]}{[\text{Acid}] - [\text{Base}]}$$

- the buffer capacity is not a fixed value for a given buffer system but instead depends on the amount of base added.

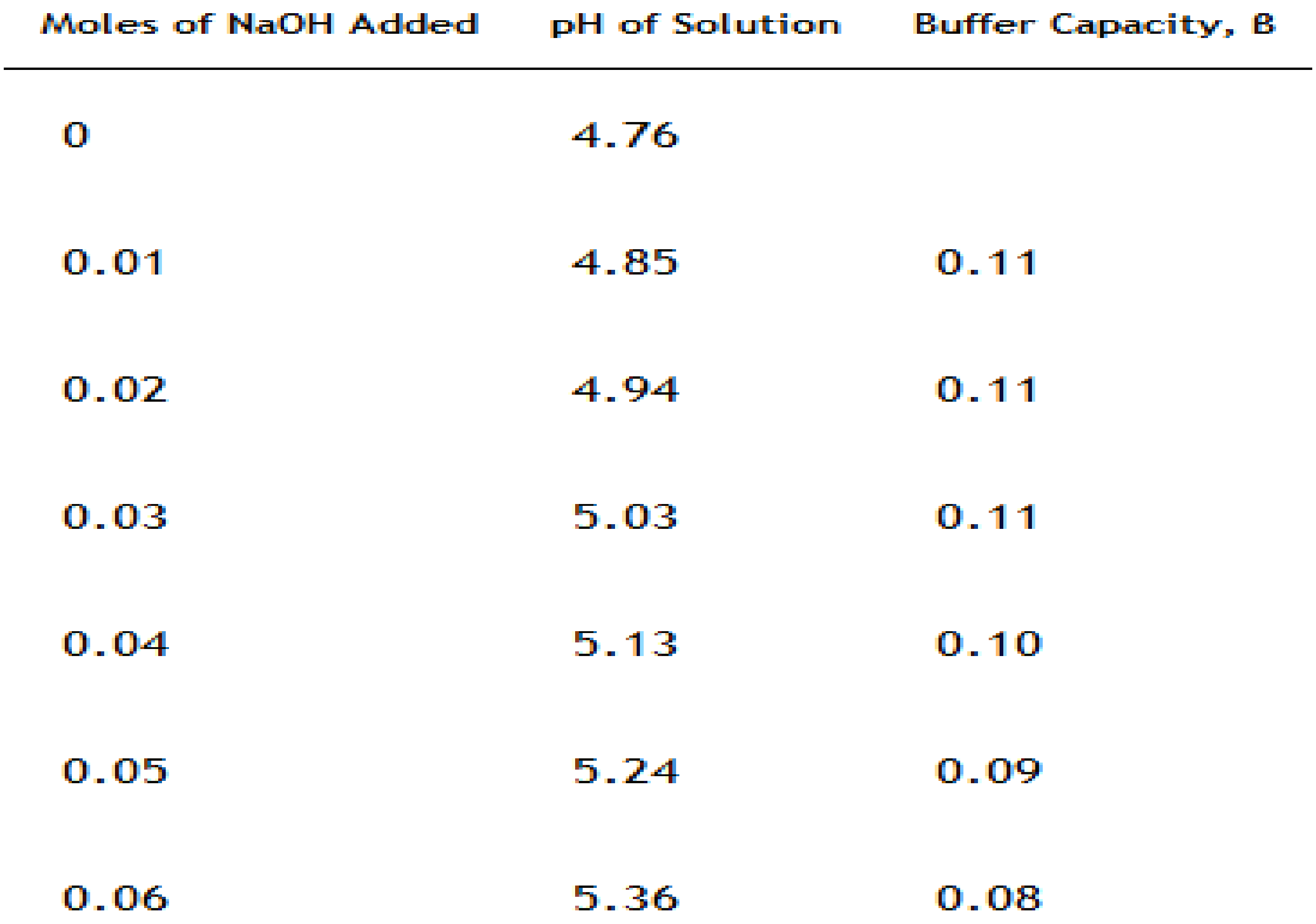
- The buffer has its greatest capacity before any base is added, where [Salt]/[Acid] = 1, and, therefore, pH = pKa.
- The buffer capacity is also influenced by an increase in the total concentration of the buffer constituents - a great concentration of salt and acid provides a greater alkaline and acid reserve.
- The buffer capacity calculated is only approximate.
- It gives the average buffer capacity over the increment of base added.

- Koppel and Spiro and Van Slyke developed a more exact equation:

$$\beta = 2.3C\frac{K_a[H_3O^+]}{(K_a + [H_3O^+])^2}$$

- C is the total buffer concentration, that is, the sum of the molar concentrations of the acid and the salt.
- This equation permits one to compute the buffer capacity at any hydrogen ion concentration—for example, at the point where no acid or base has been added to the buffer.

Buffer Capacity of Solutions Containing Equimolar Amounts (0.1 M) of Acetic Acid And Sodium Acetate



Moles of NaOH Added	pH of Solution	Buffer Capacity, β
0	4.76	
0.01	4.85	0.11
0.02	4.94	0.11
0.03	5.03	0.11
0.04	5.13	0.10
0.05	5.24	0.09
0.06	5.36	0.08

Buffer Systems

- Blood is maintained at a pH of about .
- The plasma contains carbonic acid/bicarbonate and acid/alkali sodium salts of phosphoric acid as buffers.
- Plasma proteins, which behave as acids in blood, can combine with bases and so act as buffers.
- In the erythrocytes (RBC), the two buffer systems consist of hemoglobin/oxyhemoglobin and acid/alkali potassium salts of phosphoric acid.

Pharmaceutical Buffer Systems

- Buffer solutions are used frequently in pharmaceutical practice, particularly in the formulation of ophthalmic solutions.
- One of the most common biological buffers is phosphate buffered saline (PBS).
- Contains: sodium chloride (NaCl) and dibasic sodium phosphate (Na2PO4). It may also contain potassium chloride (KCl), monobasic potassium phosphate (KH2PO4), calcium chloride (CaCl2), and magnesium sulfate (MgSO4).

General Procedures for Preparing Pharmaceutical Buffer Solutions

1. Select a weak acid having a pKa approximately equal to the pH at which the buffer is to be used.

2. From the buffer equation, calculate the ratio of salt and weak acid required to obtain the desired pH. The buffer equation is satisfactory for approximate calculations within the pH range of 4 to 10.

3. Consider the individual concentrations of the buffer salt and acid needed to obtain a suitable buffer capacity.

A concentration of 0.05 to 0.5 M is usually sufficient, and a buffer capacity of 0.01 to 0.1 is generally adequate.

4. Other factors of some importance in the choice of a pharmaceutical buffer include:

- availability of chemicals
- sterility of the final solution
- stability of the drug and buffer on aging
- cost of materials and freedom from toxicity

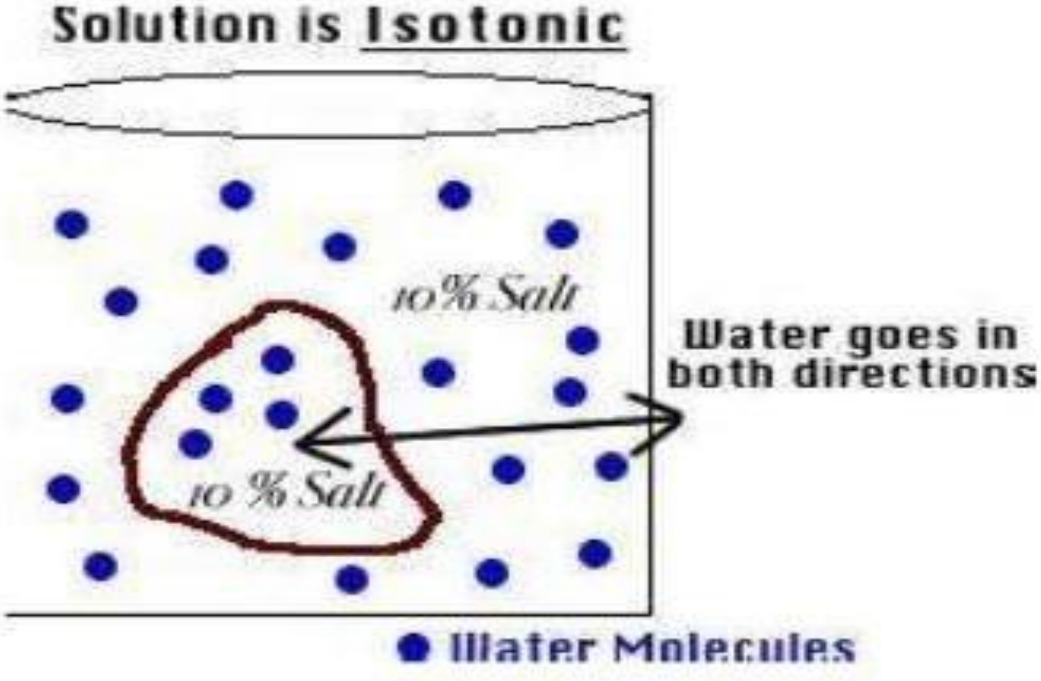
For example: a borate buffer, because of its toxic effects, certainly cannot be used to stabilize a solution to be administered orally or parenterally.

5. Finally, determine the pH and buffer capacity of the completed buffered solution using a reliable pH meter.

- In some cases, sufficient accuracy is obtained by the use of pH papers. Particularly when the electrolyte concentration is high,

Note: it may be found that the pH calculated by use of the buffer equation is somewhat different from the experimental value. This is to be expected when activity coefficients are not taken into account

ISOTONIC SOLUTIONS



ISOTONIC SOLUTIONS

- "When two solutions have same osmotic pressure and salt concentration are said to be isotonic solutions". Iso (same) and tonic (concentration).
- Physiologically, isotonic solutions are solutions having the same osmotic pressure as that of the body fluids when separated by a biological membrane. Biological fluids including blood and lachrymal fluid normally have an osmotic pressure corresponding to that of 0.9% w/v solution of sodium chloride.
- 0.9% w/v solution of NaCl has essentially the same concentration and osmotic pressure as the RBCs contents. Thus it is said to be isotonic with blood, and is also known as " ***Physiological Saline solution***", or "***Normal saline solution***"
- If RBCs are suspended in 0.9% solution of NaCl, they undergo no change in their size and shape due to isotonicity.

HYPERTONIC SOLUTION

Hypertonic

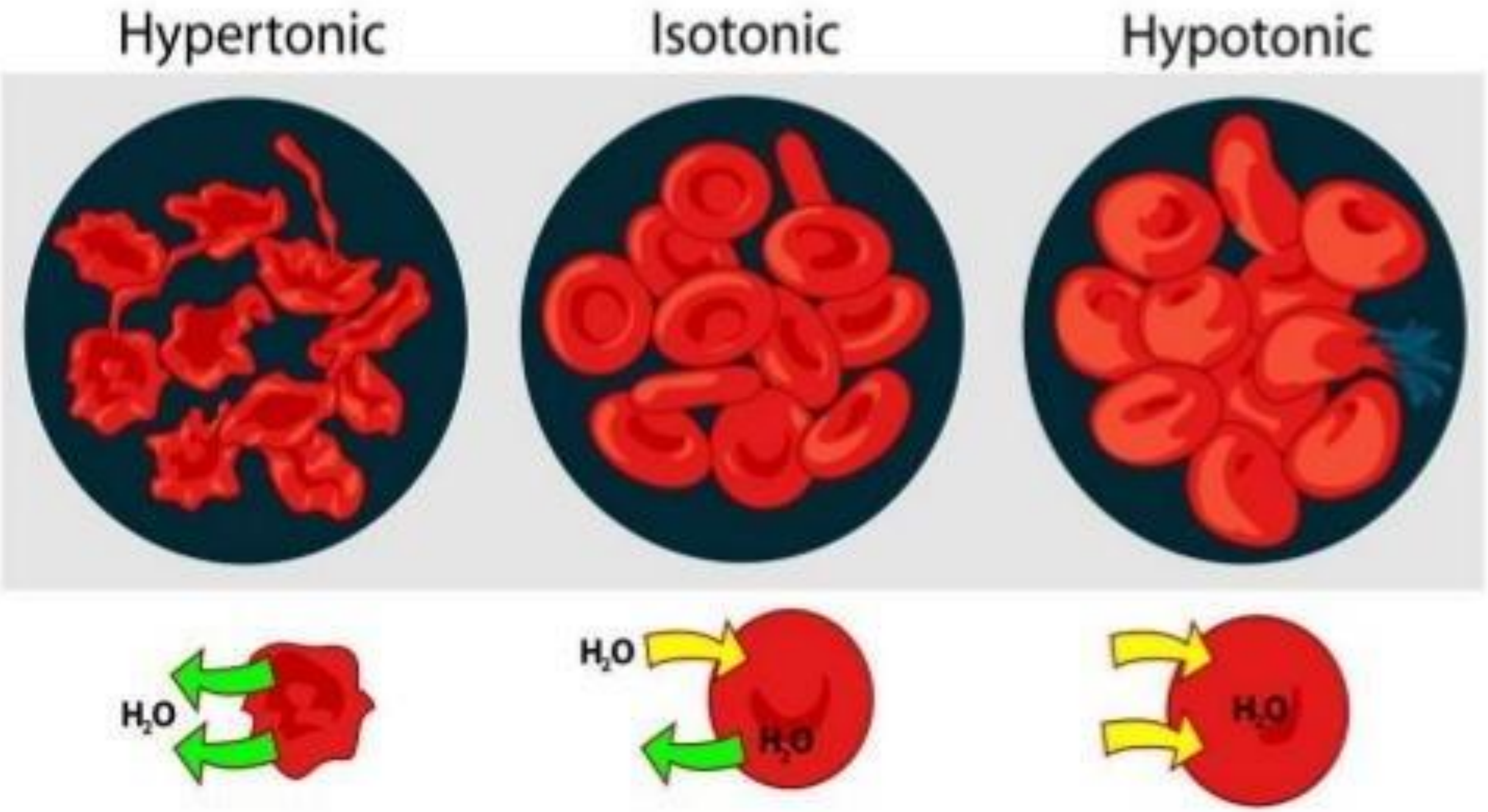
- "A hyper-tonic solution is one that has greater concentration than reference solution (i.e. RBCs Contents)."
- A hyper-tonic solution has greater osmotic pressure than that of reference solution.
- If RBCs are suspended in 2% w/v solution of NaCl (i.e. hypertonic solution), then water present within the RBCs will come out (i.e. due to osmosis, from dilute RBCs fluid to concentrated hypertonic solution) into the surroundings to dilute the NaCl solution (hypertonic solution). This exit of water from RBCs causes their shrinkage and RBCs become wrinkled in shape. This shrinkage of RBCs is known as ***"Plasmolysis"***.

HYPOTONIC SOLUTION

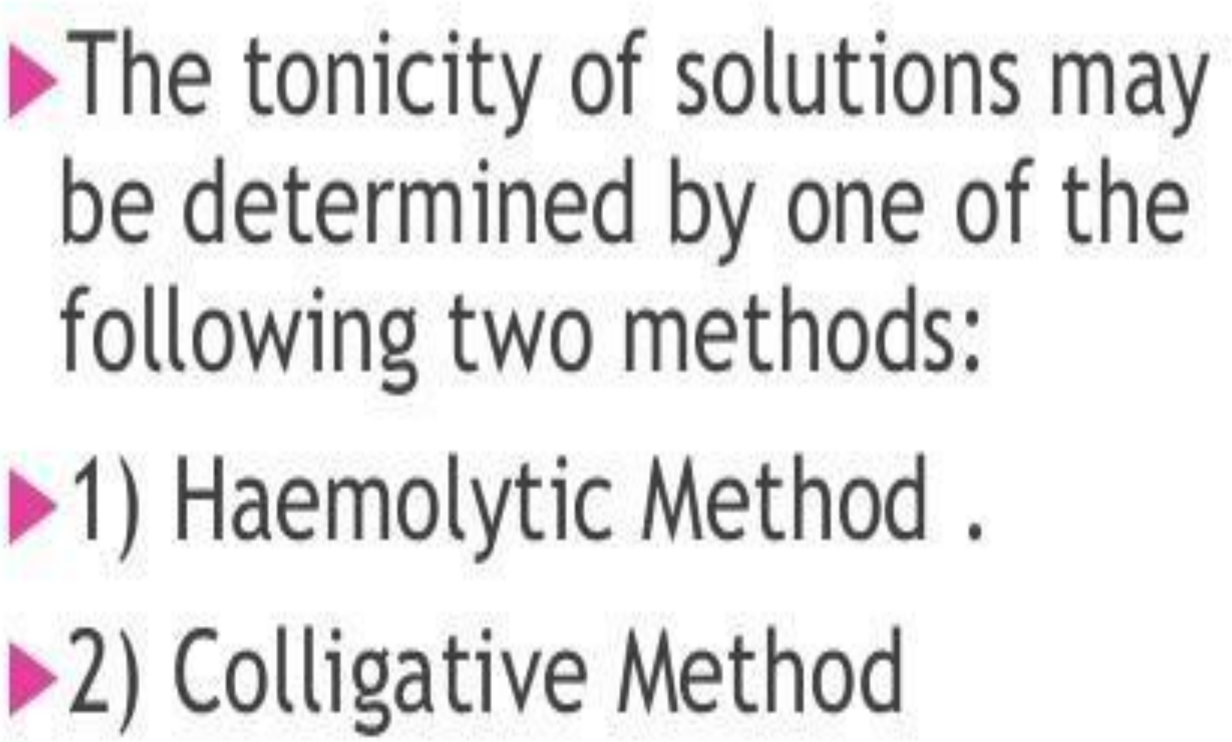
Hypotonic

- "A hypo tonic solution is one that has lower concentration than reference solution (i.e. RBCs contents).
- A hypo-tonic solution has lower osmotic pressure than that of reference solution.
- If RBCs are suspended in 0.1 % w/v solution of NaCl (i.e. hypotonic solution), then water from this solution will enter the
- RBCs (i.e. due to osmosis, from dilute hypotonic solution to RBCs fluid) to dilute the fluid within the RBCs causing their swelling, which may later result in rupturing of RBCs and release of haemoglobin. This rupturing of RBCs is known as "***Haemolysis***"

ISOTONIC SOLUTIONS



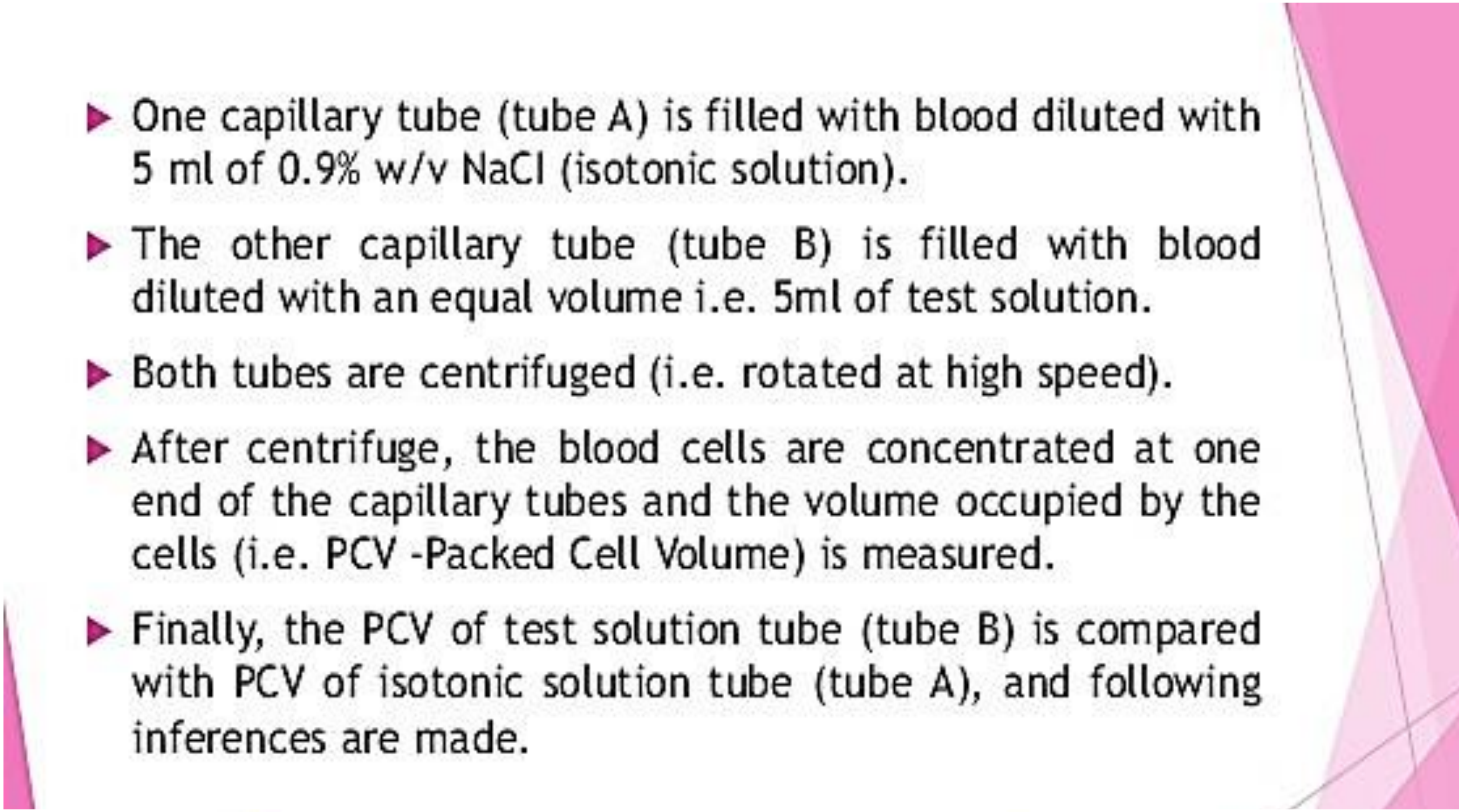
Measurement Of Tonicity



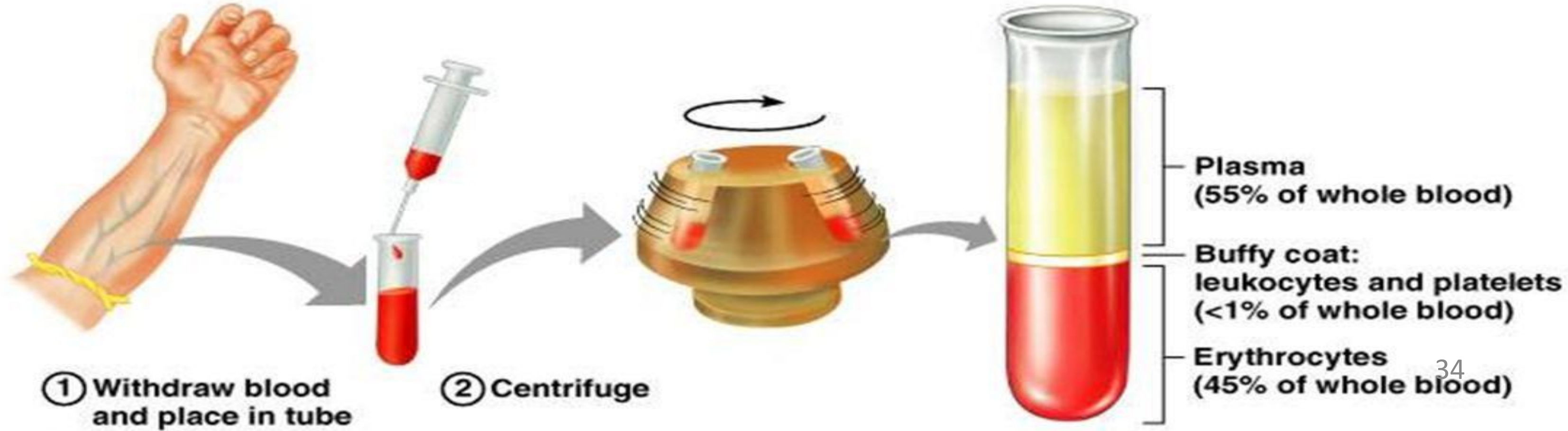
- The tonicity of solutions may be determined by one of the following two methods:
- 1) Haemolytic Method .
- 2) Colligative Method

Haemolytic method:

- The acting principle of this method is the observation of the effect of various solutions of drugs on the appearance of RBCs when suspended in those solutions. If, there is no change in size and shape of RBCs when immersed in test solution on observing with microscope, then this solution is isotonic to the blood.
- This method can be made more accurate by using a hematocrit, which is a centrifuge head in which a graduated capillary tube is held in each of the two arms.



- One capillary tube (tube A) is filled with blood diluted with 5 ml of 0.9% w/v NaCl (isotonic solution).
- The other capillary tube (tube B) is filled with blood diluted with an equal volume i.e. 5ml of test solution.
- Both tubes are centrifuged (i.e. rotated at high speed).
- After centrifuge, the blood cells are concentrated at one end of the capillary tubes and the volume occupied by the cells (i.e. PCV -Packed Cell Volume) is measured.
- Finally, the PCV of test solution tube (tube B) is compared with PCV of isotonic solution tube (tube A), and following inferences are made.



RESULTS:

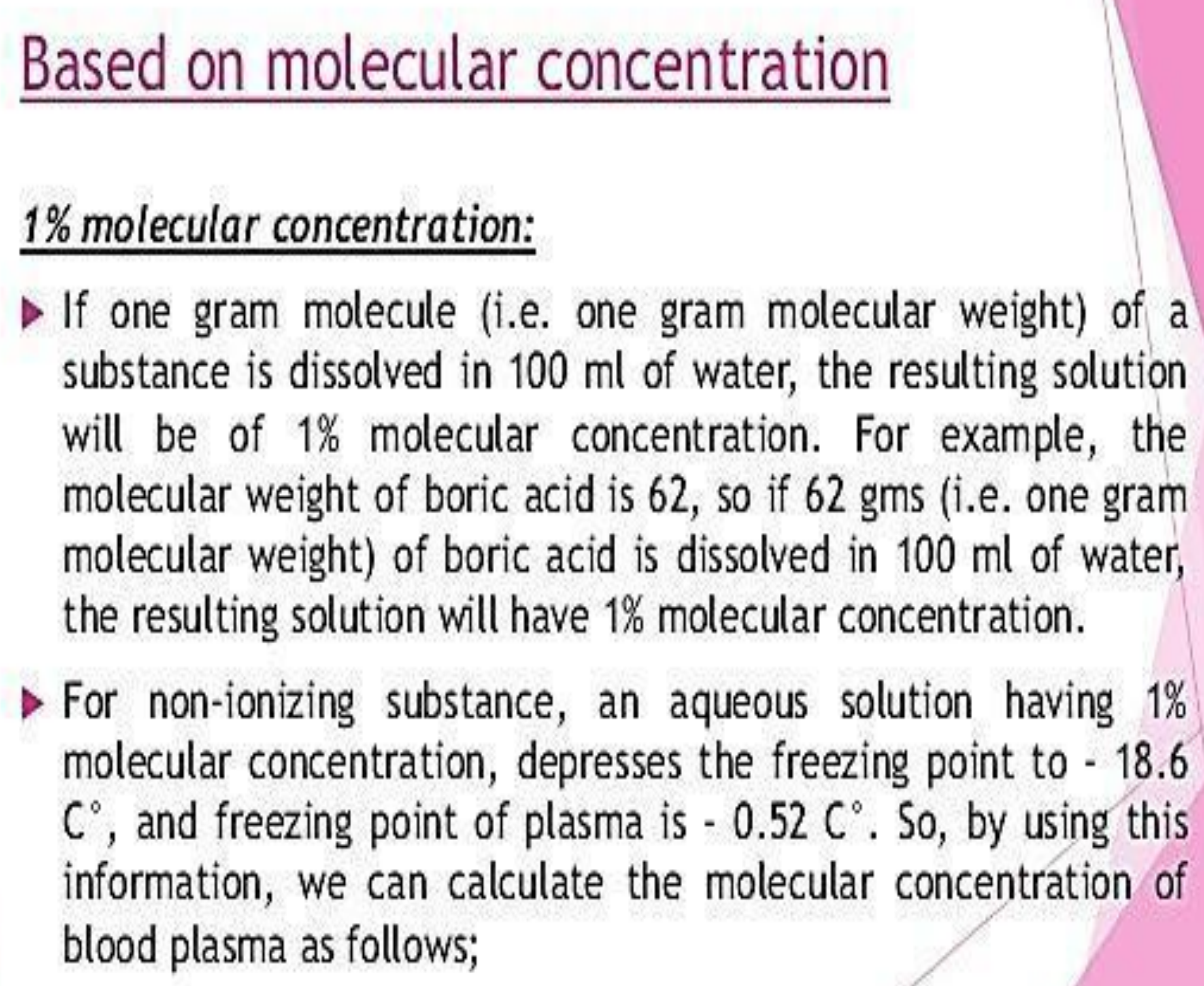
- If PCV of test solution (tube B) is same as that of tube A, then test solution is regarded as isotonic.
- If RBCs volume (i.e. PCV) of tube is more than that of tube A, then test solution is regarded as hypotonic solution (increase in PCV is due to swelling of RBCs, which occurs in case of hypotonic solution).
- If RBCs volume (i.e. PCV) of tube is less than that of tube A, then test solution is regarded as hypertonic solution (decrease in PCV is due to shrinkage of RBCs, which occurs in case of hypertonic solution).

Colligative Method:

- It has been determined that solutions having same tonicity exhibit similar behavior with respect to their colligative properties such as lowering of vapour pressure, depression in freezing point, etc. Hence, tonicity of a solution may be determined by determining its colligative properties.

For making isotonic solutions, the quantities of substances to be added may be calculated by following methods:

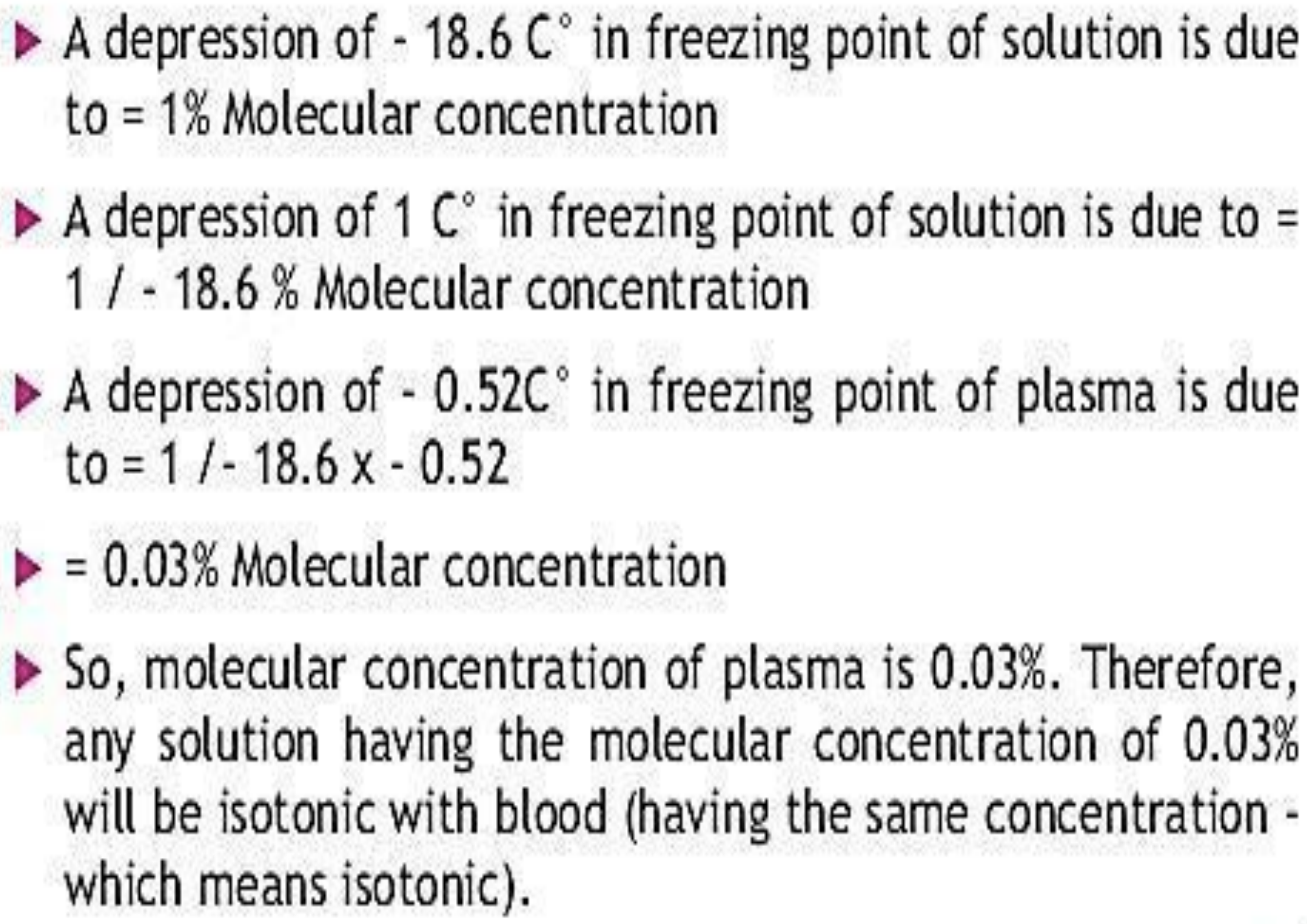
- Based on molecular concentration
- Based on freezing point data
- Based on sodium chloride equivalent (E) value
- White-Vincent method



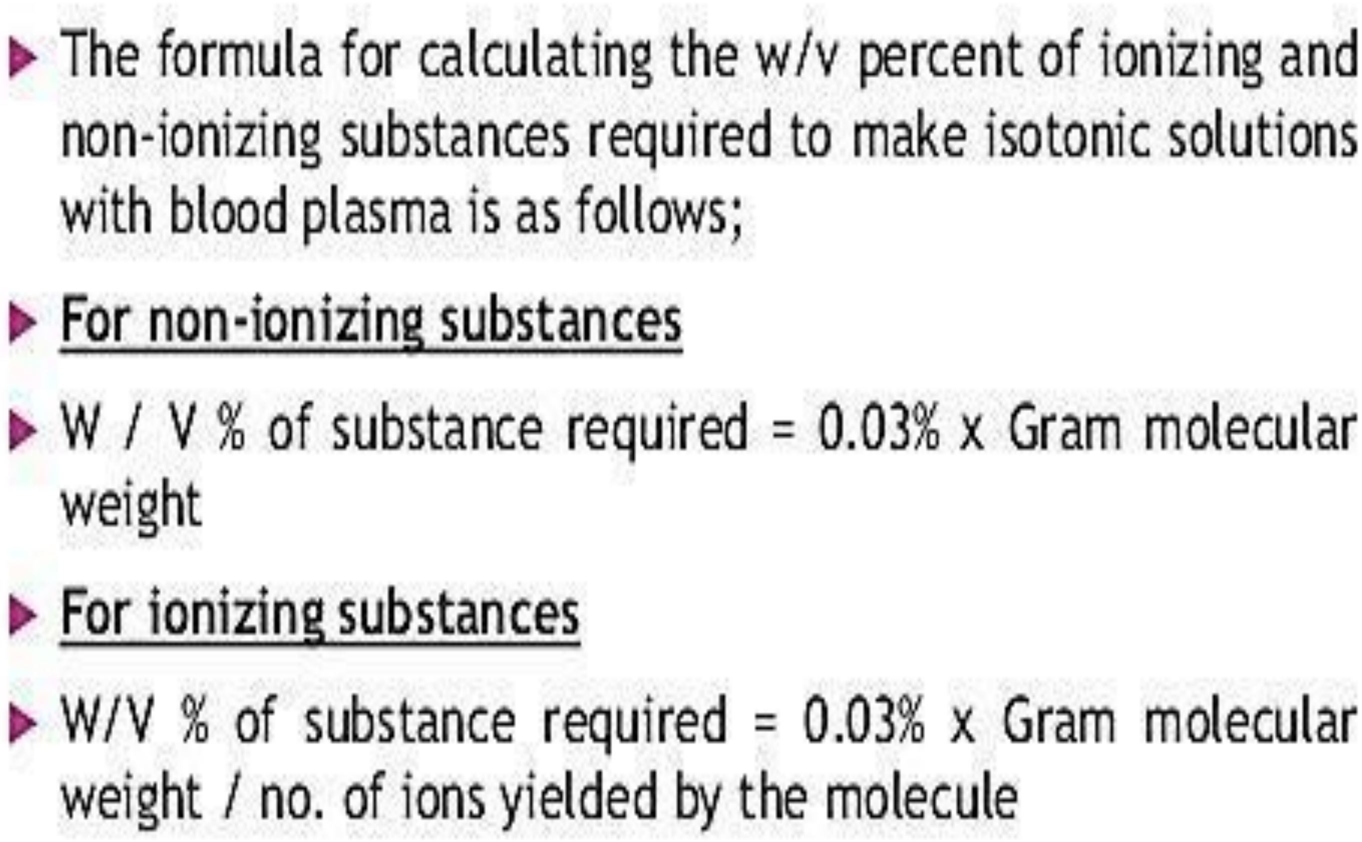
Based on molecular concentration

1% molecular concentration:

- If one gram molecule (i.e. one gram molecular weight) of a substance is dissolved in 100 ml of water, the resulting solution will be of 1% molecular concentration. For example, the molecular weight of boric acid is 62, so if 62 gms (i.e. one gram molecular weight) of boric acid is dissolved in 100 ml of water, the resulting solution will have 1% molecular concentration.
- For non-ionizing substance, an aqueous solution having 1% molecular concentration, depresses the freezing point to - 18.6 C°, and freezing point of plasma is - 0.52 C°. So, by using this information, we can calculate the molecular concentration of blood plasma as follows;



- A depression of - 18.6 C° in freezing point of solution is due to = 1% Molecular concentration
- A depression of 1 C° in freezing point of solution is due to = 1 / - 18.6 % Molecular concentration
- A depression of - 0.52C° in freezing point of plasma is due to = 1 / - 18.6 x - 0.52
- = 0.03% Molecular concentration
- So, molecular concentration of plasma is 0.03%. Therefore, any solution having the molecular concentration of 0.03% will be isotonic with blood (having the same concentration - which means isotonic).



- The formula for calculating the w/v percent of ionizing and non-ionizing substances required to make isotonic solutions with blood plasma is as follows;
- **For non-ionizing substances**
- W / V % of substance required = 0.03% x Gram molecular weight
- **For ionizing substances**
- W/V % of substance required = 0.03% x Gram molecular weight / no. of ions yielded by the molecule

Find the proportion of Boric Acid required to make a solution isotonic. The molecular weight of boric acid is 62, and it is a non-ionizing substance

Solution:

- By applying formulae for non-ionizing substances;
- W/V % of boric acid required to make isotonic solution = 0.03% x gram molecular weight
- = 0.03 x 62 = 1.86
- So, 1.86 gms of boric acid is required to make 100ml isotonic solution.

Find the proportion of Sodium sulphate required to make a solution isotonic. The molecular weight of sodium sulphate is 148, and it is an ionizing substance.

By applying formulae for ionizing substances;

- W/V % *of* ionizing substance required to make isotonic solution = 0.03% x gram molecular weight / no. of ions
  - Na2SO4 —> 2Na + SO4
- So, total no. of ions produced by sodium sulphate = 3
- W/V % of Na2SO4 required to make isotonic solution = 0.03% x gram molecular weight / no. of ions
- = 0.03 % x 148 / 3 = 1.48
- *SO*, 1.48 gm of sodium sulphate is required to make 100ml isotonic solution.

Find the proportion of Dextrose required to make a solution isotonic. The molecular weight of dextrose is 180, and it is a non-ionizing substance?

= 0.03 x 180 = 5.4g

Based on freezing point data/Cryoscopic method

- Blood plasma & lachrymal secretions have a freezing point of -0.52 C° due to different solutes present in them.
- 0.9% NaCl solution also has freezing point aprox. -0.52 C°
- Hence, all solutions which freeze at this temperature will be isotonic with these fluids.

Adjustment to the tonicity of solutions:

- If the freezing point of 1% solution whose tonicity is to be adjusted is known.
- And freezing point of 1% solution of adjusting substance is known.

- **Tonicity of solution based on the freezing point data calculated from the general formula**

Amount of adjusting substance required = -0.52 – a/b

- **a = freezing point of 1% solution of un-adjusted solution**
- **b = freezing point of 1% solution of adjusting solution**

Find the amount of sodium chloride required to render *1%* solution of cocaine hydrochloride isotonic with blood plasma. The freezing point of 1% w/v solution of cocaine hydrochloride is - 0.09 *C*°, and that of 1% NaCl is-0.576 *C*°.

- In this example,
- Unadjusted solution (whose tonicity is to be adjusted) is *1%* cocaine HCl
- Adjusting substance is NaCl
- Freezing point of 1% w/v solution of cocaine HCl (unadjusted solution) = a = - 0.09 *C*°
- Freezing point of *1%* w/v solution of NaCl (adjusting substance) = b = - 0.576 *C*°

Amount of adjusting substance required = - 0.52 -a /b

- *= -0.52-(-0.09) / -0.576 = 0.746 gms*
- Hence, by adding 0.746 gms of NoCl in 1% cocaine HCl solution, the final solution becomes isotonic.

Based on Sodium Chloride equivalent (E)

Sodium chloride equivalent i.e. (E) of a drug is defined as;

- "The grams of Sodium chloride that will produce the same osmotic effect as 1 gm of that drug."
- For example, potassium chloride has sodium chloride equivalent (E) value of 0.76 gm NaCl / gm of KCl-. This means 0.76 gm of NaCl produce the same osmotic effect as 1 gm of KCl.
- To make a solution of a particular drug isotonic with blood plasma, the sodium choride equivalent value (E) of that drug is noted from the reference table. This E value is multiplied with the %age of the drug solution, and result so obtained is subtracted from 0.9%. The difference in value so obtained is the amount of NaCl needed to adjust the tonicity of the solution to isotonic value.
- Amount of NaCl required = 0.9% - {%age of solution × E)

SODIUM CHLORIDE EQUIVALENTS (*E* VALUES)

SUBSTANCE	MOLECULAR WEIGHT	IONS	SODIUM CHLORIDE EQUIVALENT (E VALUE)
Boric acid	61.8	1	0.52
Chlorobutanol	177	1	0.24
Pilocarpine hydrochloride	245	2	0.24
Pilocarpine Nitrate	271	2	0.23
Phenacaine hydrochloride	353	2	0.20
Silver nitrate	170	2	0.33
Potassium nitrate	101	2	0.58

Find the amount of sodium chloride needed to make a - solution of 0.5% of KCl isotonic with blood plasma. Sodium chloride equivalent value (E) of KCl is 0.76.

- Given solution (not isotonic) = 0.5% KCl
- E value of KCl = 0.76 So, by applying formula,
- Amount of NaCl required = 0.9 - (%age of drug × E)
- = 0.9-(0.5 x 0.76)
- = 0.9 -0.38 = 0.52 gm —> Hence, 0.52 gm of NaCl must be added in 0.5% KCl solution to make it isotonic.

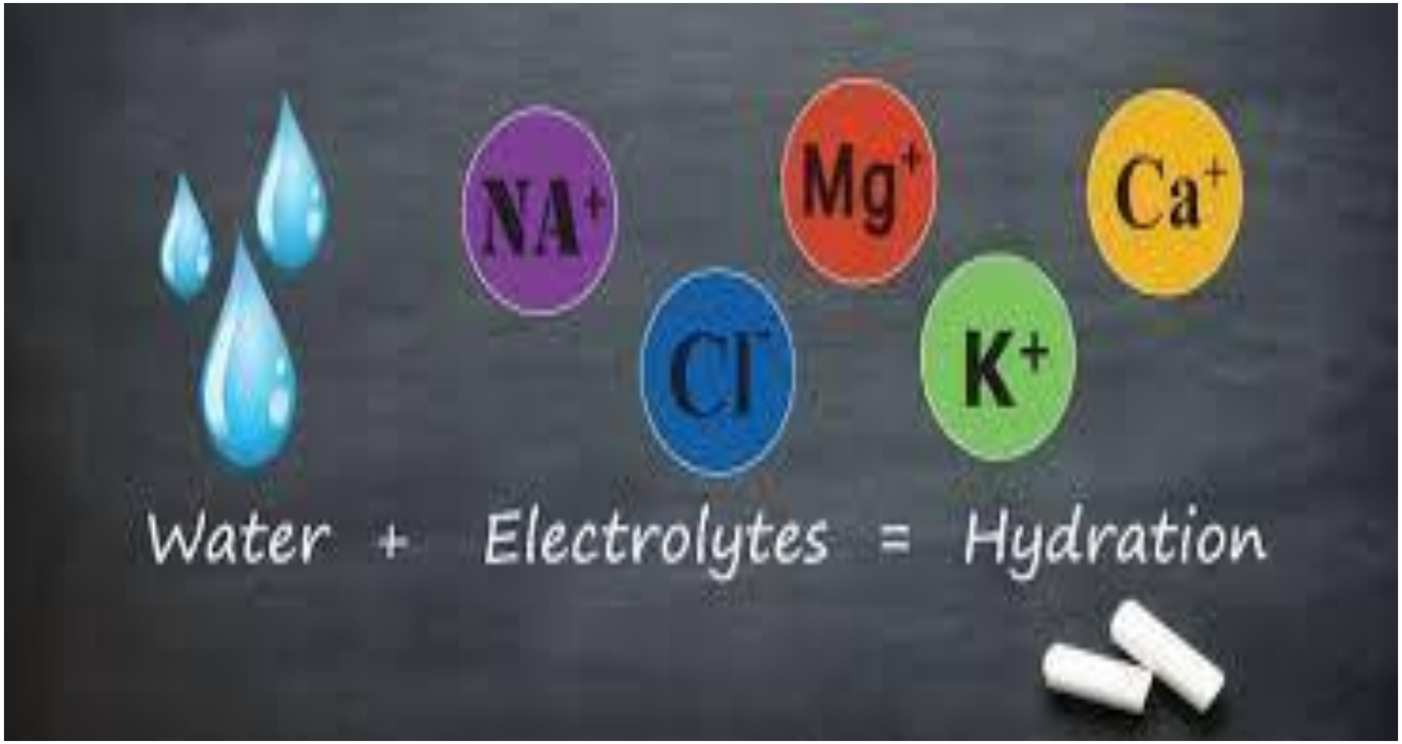
White-Vincent method

- This method involves the addition of water to the given amount of drug to make isotonic solution, followed by the addition of some other isotonic solution (e.g. 0.9% NaCl) to make the final volume.
- The volume of water that should be added in given amount of drug to make isotonic solution is calculated by using following formula;
- $V = W \times E \times 111.1$
- Where, V = volume of water needed to make isotonic solution
- W = given weight of drug in grams
- E = NaCl equivalent value of drug
- 111.1 = constant

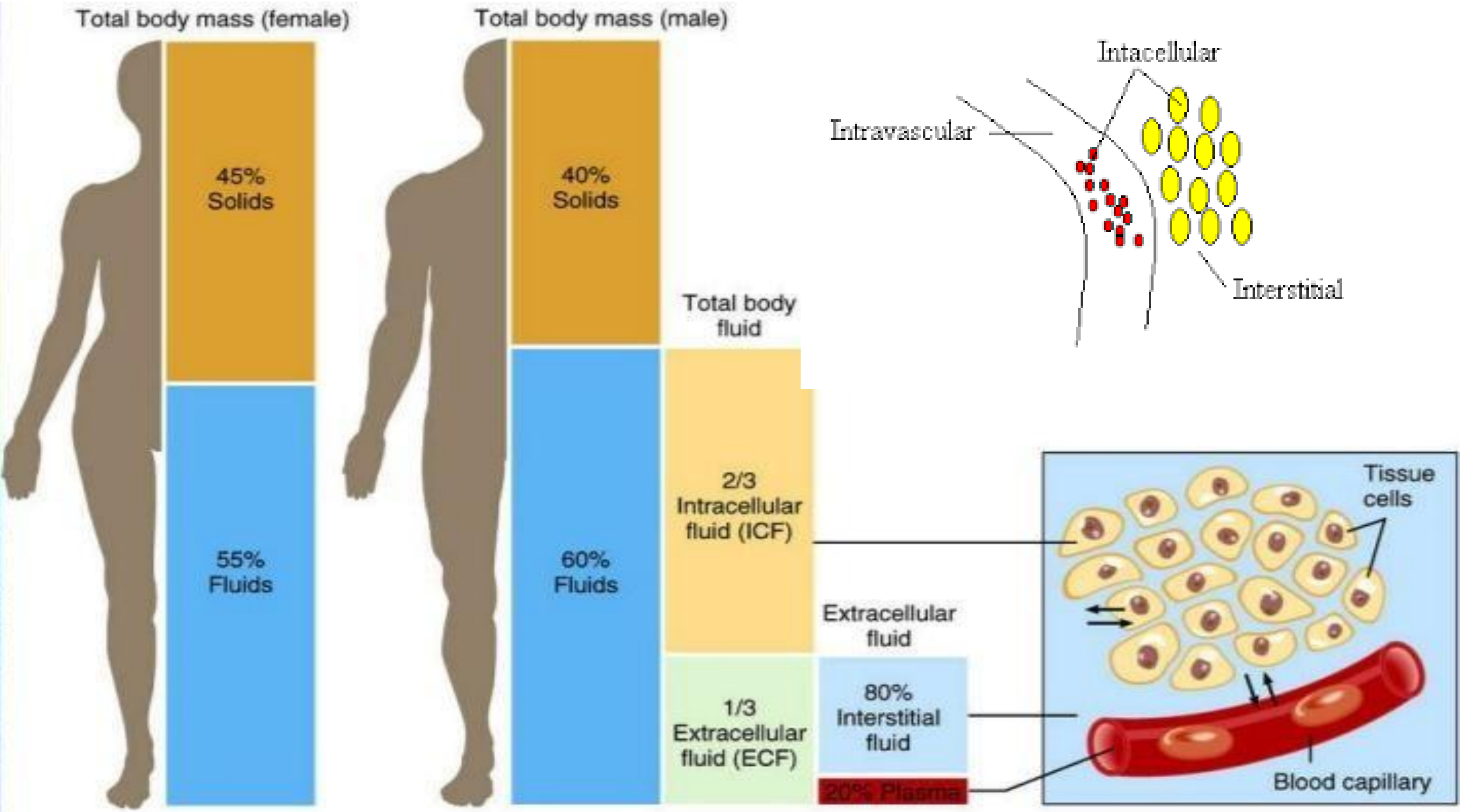
Make 50ml isotonic solution from 0.5 gm of boric acid. E value of boric acid is 0.50.

- Given amount of boric acid = 0.5 gm
- Required volume = 50 ml*
- E value of boric acid = 0.50
- Firstly, we calculate the amount of water that should be added in 0.5 gm of boric acid to make isotonic solution by using formula,
- V= W x E x 111.1
- V = 0.5 x 0.5 x 111.1 = 27.8 ml
- So, 0.5 gm of boric acid is dissolved in 27.8 ml of water to make isotonic solution.
- BUT, final volume that is required is 50 ml. so, remaining 22.2ml (50 - 27.8 = 22.2) of some other isotonic solution (e.g. 0.9% NaCl) are added to make up final 50 ml volume.

Major Intra & Extracellular ELECTROLYTES



Major extra and intracellular ELECTROLYTES



(b) Exchange of water among body fluid compartments

- Chemical substance dissolved in body fluid can be categorized into:

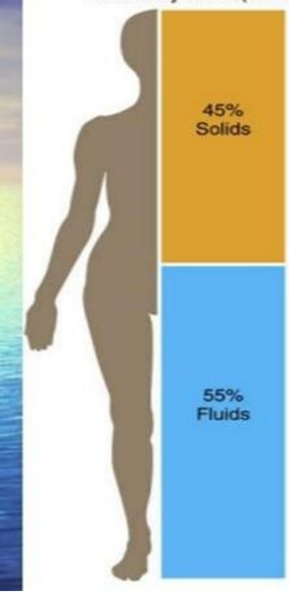
A. Non-electrolytes: Organic molecules, Do not generate ions in solution form.

e.g., Glucose, Urea, Creatine etc

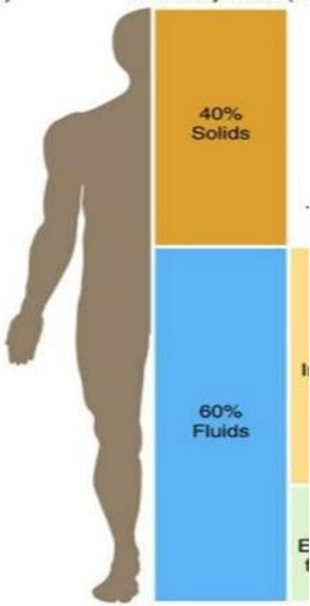
B. Electrolytes: Mostly inorganic substances, Dissociates into ions (+ve/-ve) in the body fluid.

e.g., Acids, Bases, Salts, few organic molecules like Citric acid, Lactic acid, Oxaloacetic acid etc

Body: “Both are necessary to perform physiological functions”!



Body water – *Dissolved* Necessary Elements



Substance Required

To:

Metabolize Nutrients & Drugs

Generate Energy

Maintain & Mfg. Body Components

Eliminate:

by-products & waste

Internal Homeostasis
(ionic, osmotic, pH balance)

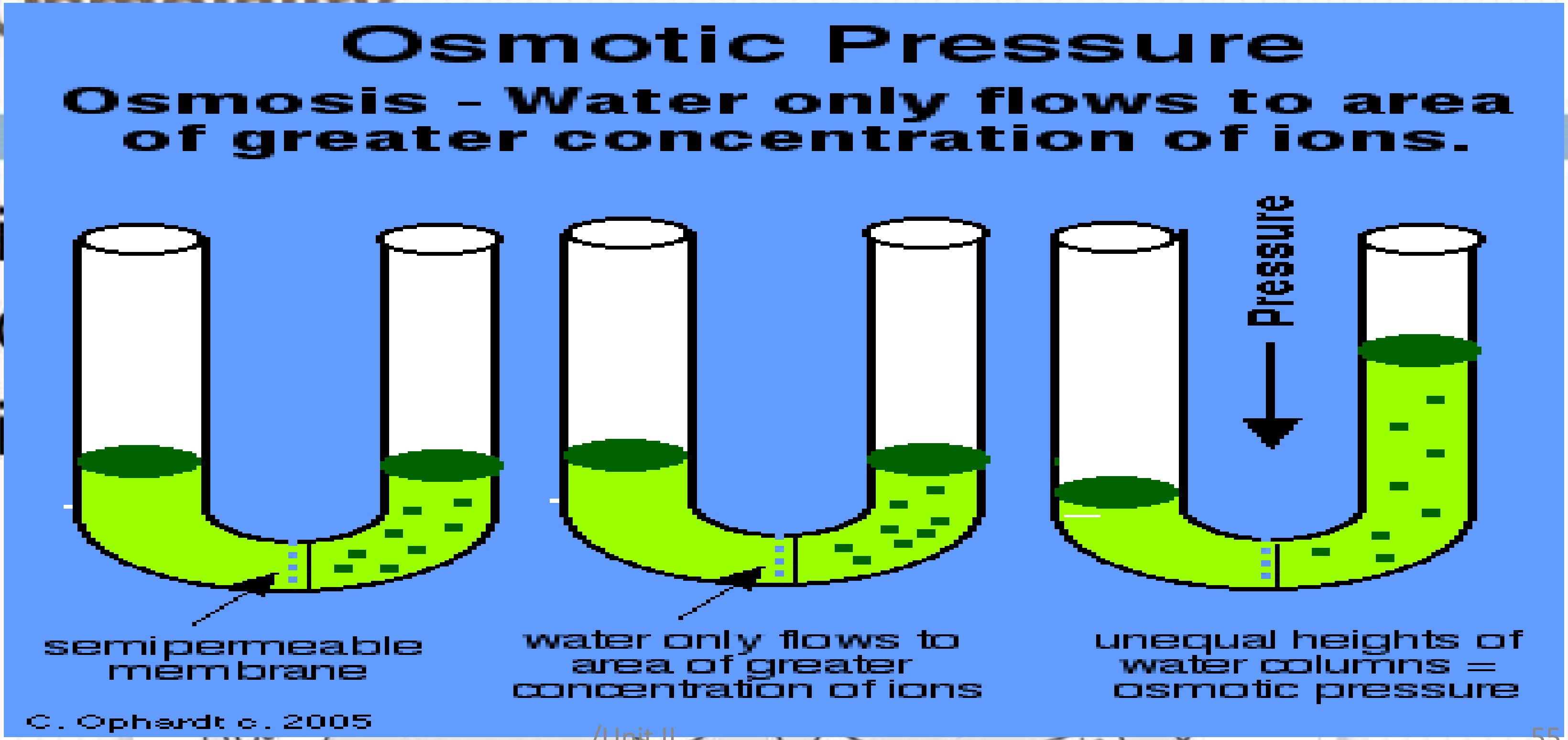
Definitions:

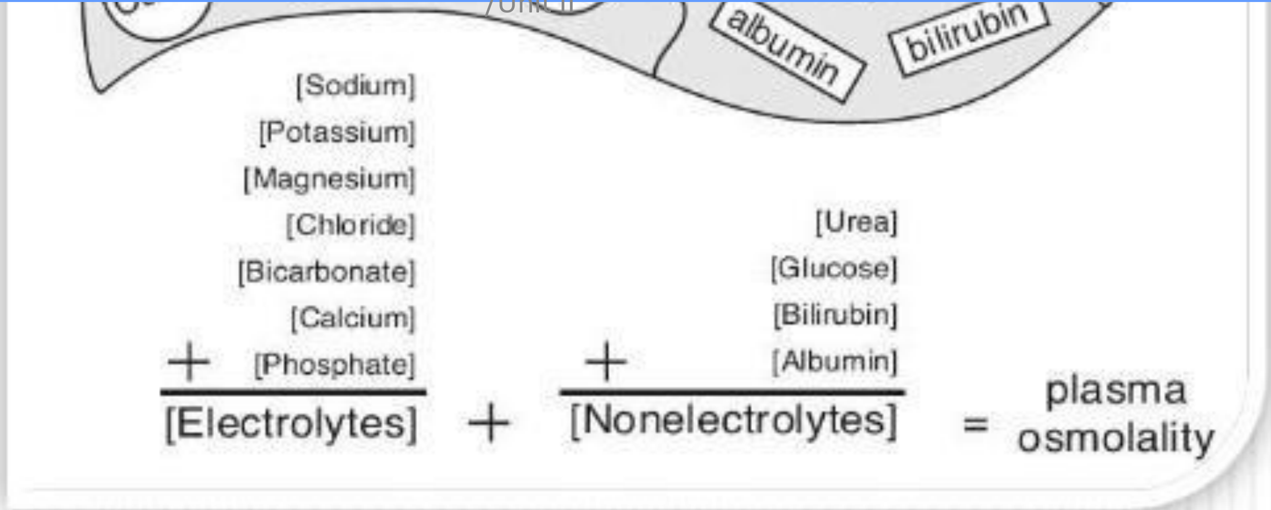
- Osmotic Pressure: concentration of electrolytes (dissolved ions) in each compartment that creates the osmotic pressure that holds water in the appropriate space.

Osmolality

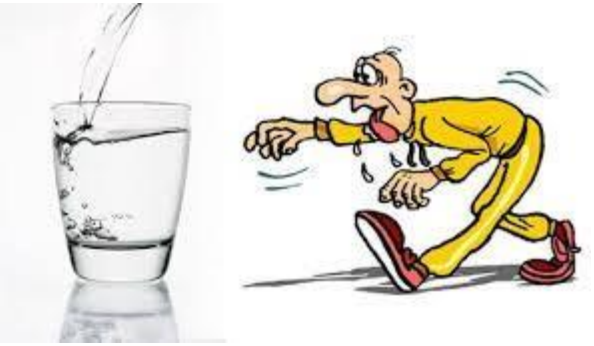
☐ i

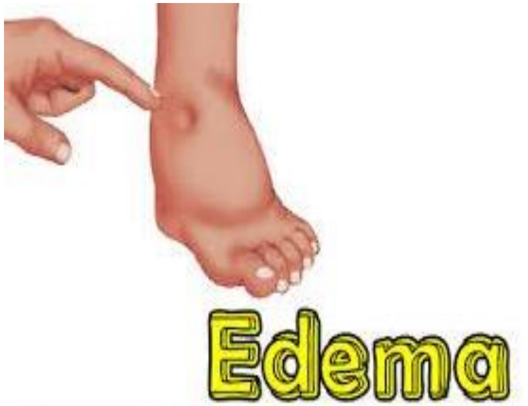
☐ i

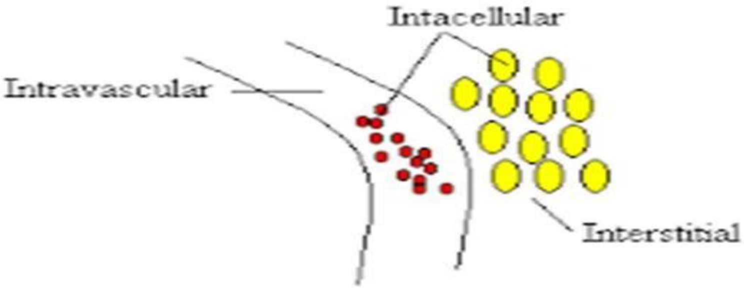




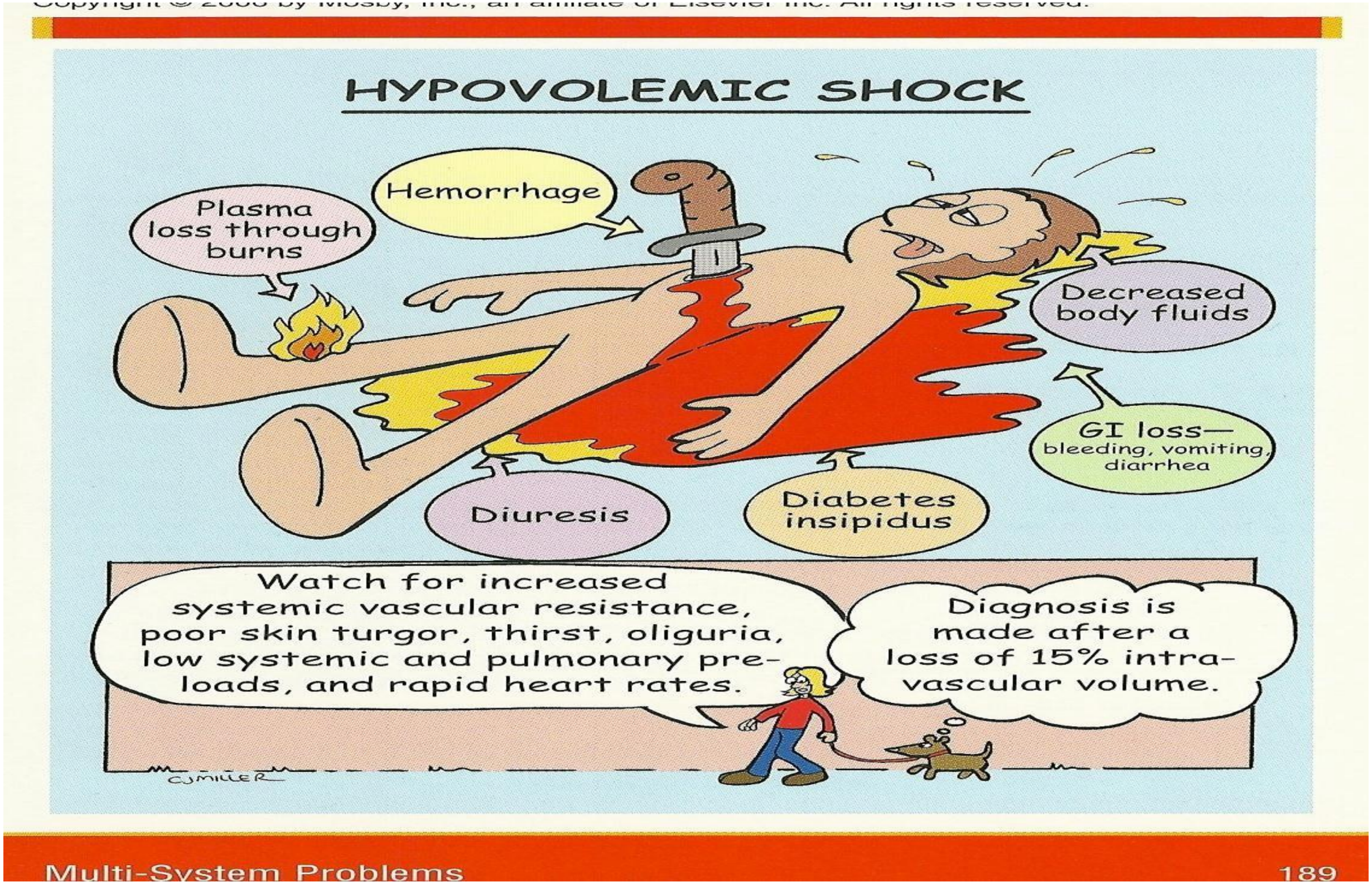
- **Dehydration:** state in which water volume is low in all 3 compartments (Intracellular, interstitial & plasma fluid).
- **Edema:** State in which fluid accumulates in the interstitial space due to low Oncotic (Protein) pressure.





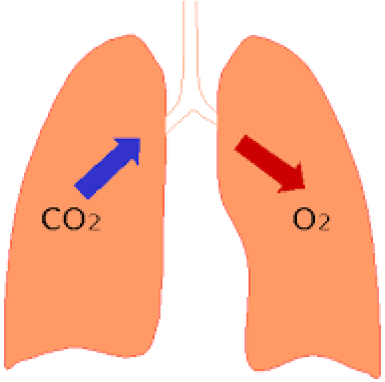


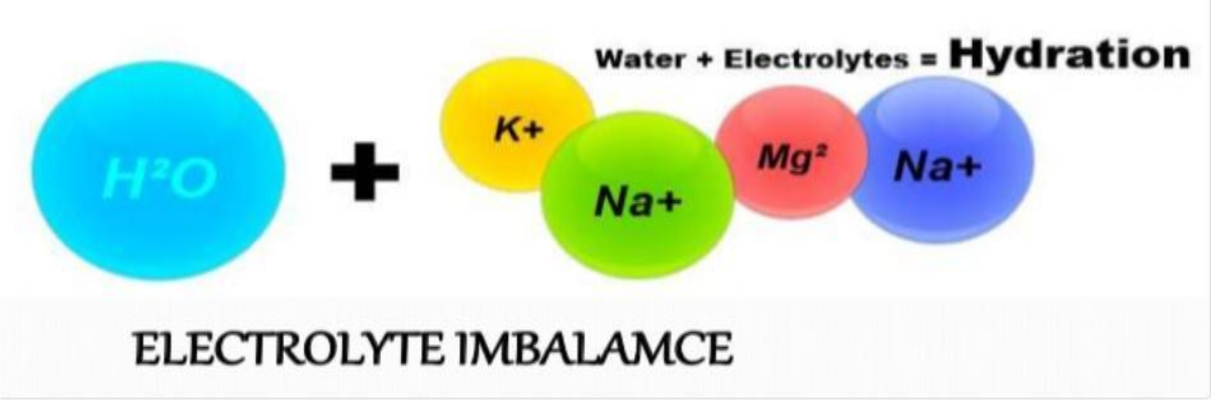
- **Hypovolemia:** State in which intravascular volume is low.



Salt & water balance:

- Oral intake of fluid & electrolytes
- Evaporation of solute – free water across the skin and lungs.
- Excretion of water & electrolytes through the kidneys : ⇅ output – antidiuretic hormone (ADH) & aldosterone.





- The fluid in each compartment is ionically balanced.
- Body has the capacity to adjust slight variations in electrolytic concentration of the fluid compartments.
- If concentration of electrolytes changes – water will migrate across the cell membrane to reestablish <u>Osmotic equilibrium</u>.

Replacement Therapy

- When body itself fails to correct an electrolyte imbalance.

Products:

- Electrolytes
- Acids & Bases
- Blood Products
- Carbohydrates
- Amino acids
- Proteins

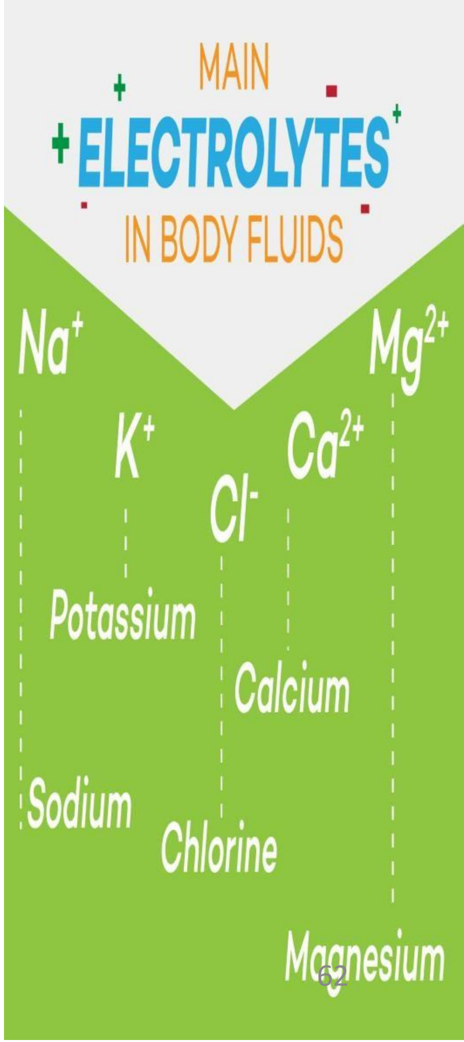
Electrolytes

- Mineral salts (inorganic compounds) are necessary within the body for all body process.
- They are usually required in small quantities.
- Main elements:

Calcium & Phosphorus: bone & teeth

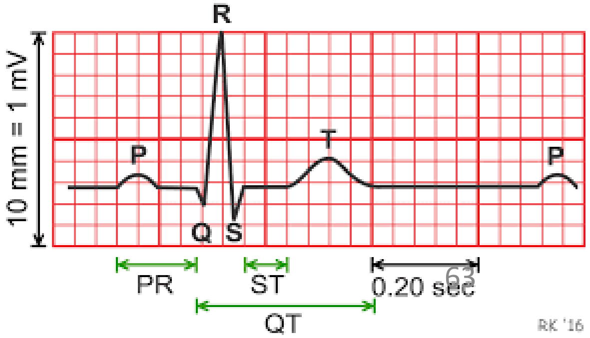
Iron: haemoglobin - convey oxygen & CO2.

Na & K: Transmission of nerve impulses & contraction of muscles



Important Functions:

- Control of osmosis of water between body compartments.
- Maintain the acid-base balance required for normal cellular activates.
- Help to generate <u>action potentials</u> & graded potentials.
- Help to control secretion of some hormones (e.g., Aldosterone, Thyroid hormones) and neurotransmitters.



Major Physiological Ions

- **Nature/Properties**
- **Important Role/Major Physiological role**
- ↑↓s

Sodium (Na+),
Chloride (Cl-),
Potassium (K+),

Calcium (Ca2+),
Magnesium (Mg2+),
Phosphate ($H2PO4^{-}$,$HPO4^{2-}$,$PO4^{3-}$),
Bicarbonate ($HCO3^{-}$)

Prepare a
Simple Report

Prepare a
Simple Report

Electrolytes used in the Replacement Therapy

- In a healthy person, at least 70 liters of fluids are exchanged (secreted and reabsorbed) across the walls of the intestines per day.
- The brain, heart, kidney, and virtually every other vital organ depend on these fluids to function.
- As the body takes in the water and salts it needs, it loses or excretes those it does not need through urine, stools, and sweat.
- Thus, the secretion and absorption rates are kept in balance.

- In various condition like prolonged fever, sever vomiting or diarrhea creates a tremendous outpouring of water (heavy loss of water) & electrolytes (body salts) state of dehydration and impairs the capacity to reabsorb the fluid & electrolytes in our system.
- To compensate this loss, Electrolyte Replacement Therapy / Oral Rehydration Therapy is required.

“Replace what it Lost” Dr. Perla D. Santos

- 2 types of solutions used

1. A solution for rapid initial replacement:

Name	Concentration Range
Sodium	130 – 150 mEq/L
Chlorine	98 – 110 mEq/L
Potassium	4 – 12 mEq/L
Bicarbonate	28 – 55 mEq/L
Calcium	3 -5 mEq/L
Magnesium	3 mEq/L

These electrolyte concentrations thus closely resemble with the electrolyte concentrations found in extracellular fluids!

Name	mOsm/Litre
Sodium	75
Potssium	20
Dextrose	75
Chloride	65
Citrate	10
Toal osmolarity in approx. 200ml water	245

2. A solution for subsequent replacement:

Name	Concentration Range
Sodium	40 – 120 mEq/L
Chlorine	30 – 105 mEq/L
Potassium	16 – 35mEq/L
Bicarbonate	16 – 53 mEq/L
Calcium	10 - 15 mEq/L
Magnesium	03 - 06 mEq/L
Phosphorus	0 – 13 mEq/L

PHYSIOLOGICAL ACID-BASE BALANCE

- Electrolytes also play an important role in regulating body's acid-base balance
- Body fluids contain balanced quantities of acids & bases.

Acidity of the solution: No of [H^+] present in fluid/solution - ECF

Sources: [H^+]

- Food
- Cellular metabolism of Glucose, Fatty acids, & Amino acids etc
- Reabsorption

- Biochemical reactions: Very sensitive to change in pH (acidity/alkalinity)

e.g., enzyme Pepsin in the stomach– helps in digestion of dietary proteins at low pH.

enzyme Ptyalin in saliva – helps in digests carbohydrates at pH between 5.4 - 7.5.

Body Fluid	pH value
Urine	4.5 – 08
Blood	7.4 – 7.5
Gastric juice	1.5 – 3.5
Saliva	5.4 – 7.5
Bile	6.0 8.5

Kidney – removes excess acid – make urine acidic

Buffer Systems

Acids-bases are continually taken into & formed by the body, the pH of fluids inside & outside cells remain fairly constant because of the presence of ‘BUFFER SYSTEMS’.

- Consists of a weak acid & the salt of that acid

Functions:

- to convert strong acids or bases into weak acids or bases.
- to prevent drastic change in pH of the blood.

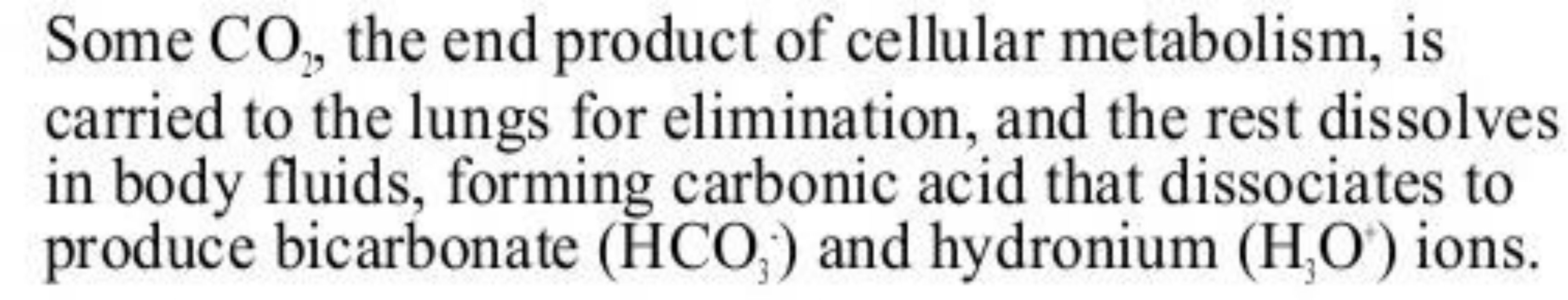
Note: However, it will be effective only if excess acid/alkali excreted out by lungs and/or kidneys.

Types of Buffer systems:

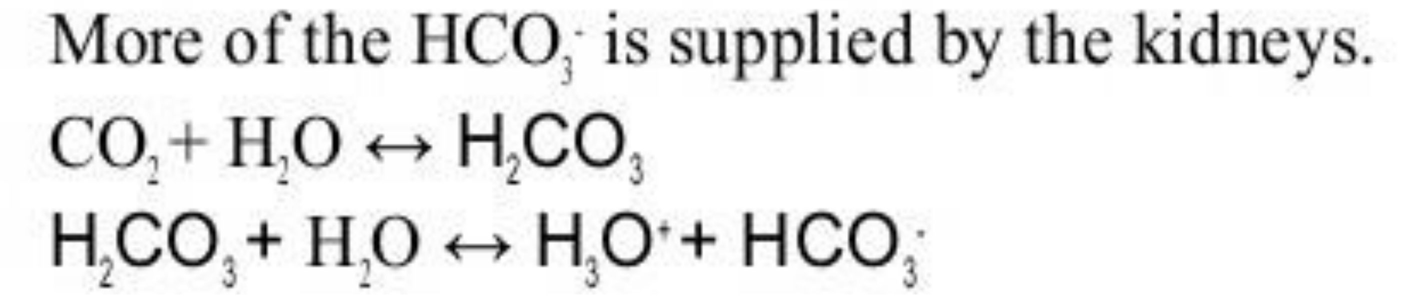
- Carbonic Acid ($H2CO3$) – Bicarbonate ($HCO3^{-}$) Buffer System
- Phosphate ($H2PO4^{-}$,$HPO4^{2-}$,$PO4^{3-}$) Buffer System
- Protein (Hemoglobin/HbH) Buffer System

Carbonic Acid (H2CO3) – Bicarbonate (HCO3$^-$) Buffer System

- Major buffer of metabolic acid/base present in Plasma & Kidneys.
- Regulates blood pH



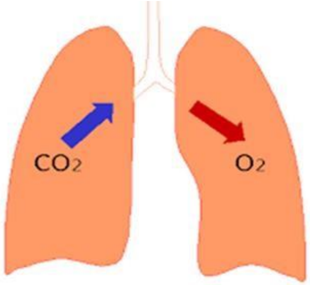
Some CO_2, the end product of cellular metabolism, is carried to the lungs for elimination, and the rest dissolves in body fluids, forming carbonic acid that dissociates to produce bicarbonate (HCO_3^-) and hydronium (H_3O^+) ions.



More of the HCO_3^- is supplied by the kidneys.

$$CO_2 + H_2O \leftrightarrow H_2CO_3$$

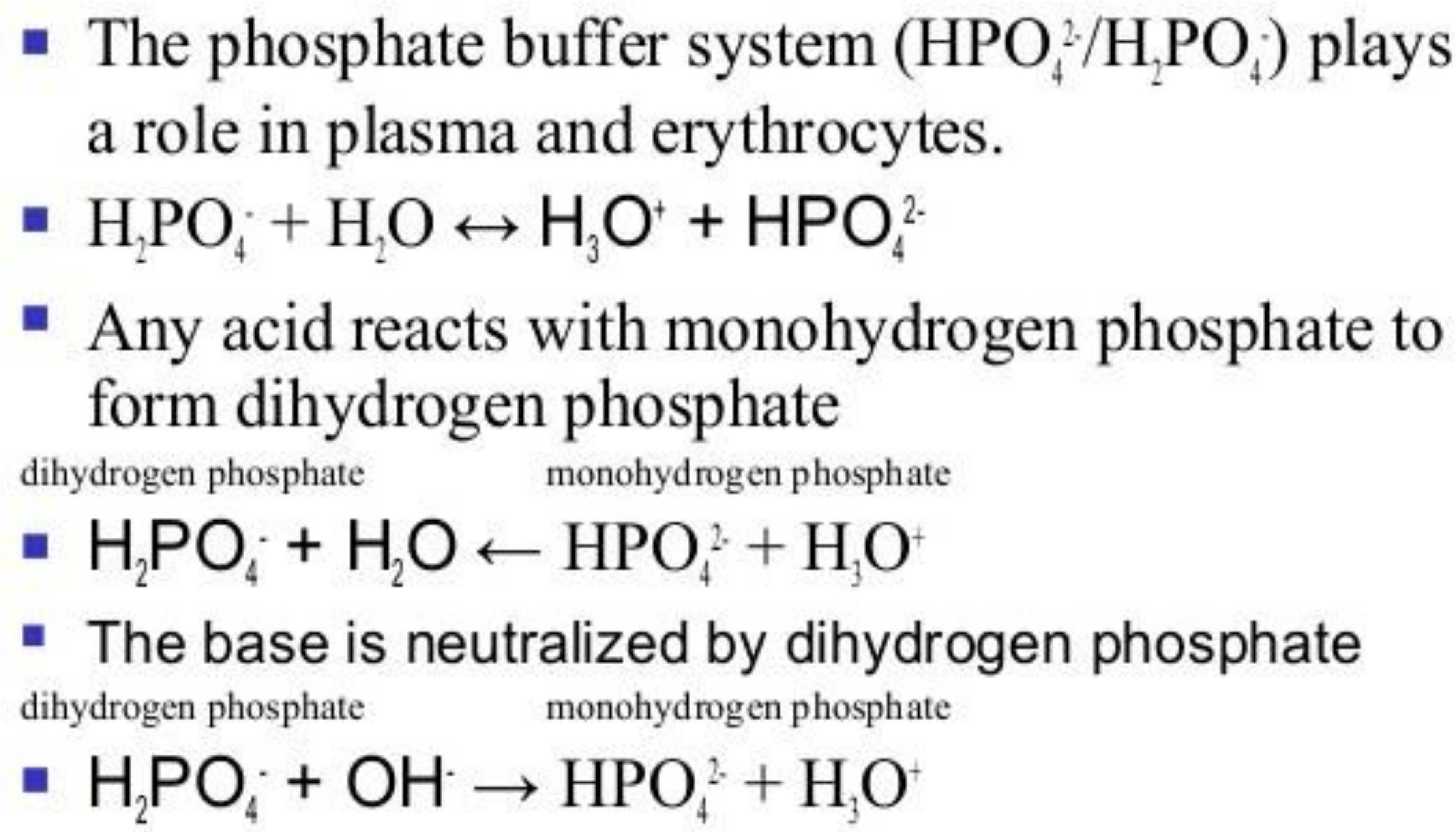
$$H_2CO_3 + H_2O \leftrightarrow H_3O^+ + HCO_3^-$$



Regulation of blood pH

- The lungs and kidneys play important role in regulating blood pH.
- The lungs regulate pH through retention or elimination of CO_2 by changing the rate and volume of ventilation.
- The kidneys regulate pH by excreting acid, primarily in the ammonium ion (NH_4^+), and by reclaiming HCO_3^- from the glomerular filtrate (and adding it back to the blood).

Phosphate Buffer System



- The phosphate buffer system ($HPO_4^{2-}/H_2PO_4^-$) plays a role in plasma and erythrocytes.
- $H_2PO_4^- + H_2O \leftrightarrow H_3O^+ + HPO_4^{2-}$
- Any acid reacts with monohydrogen phosphate to form dihydrogen phosphate

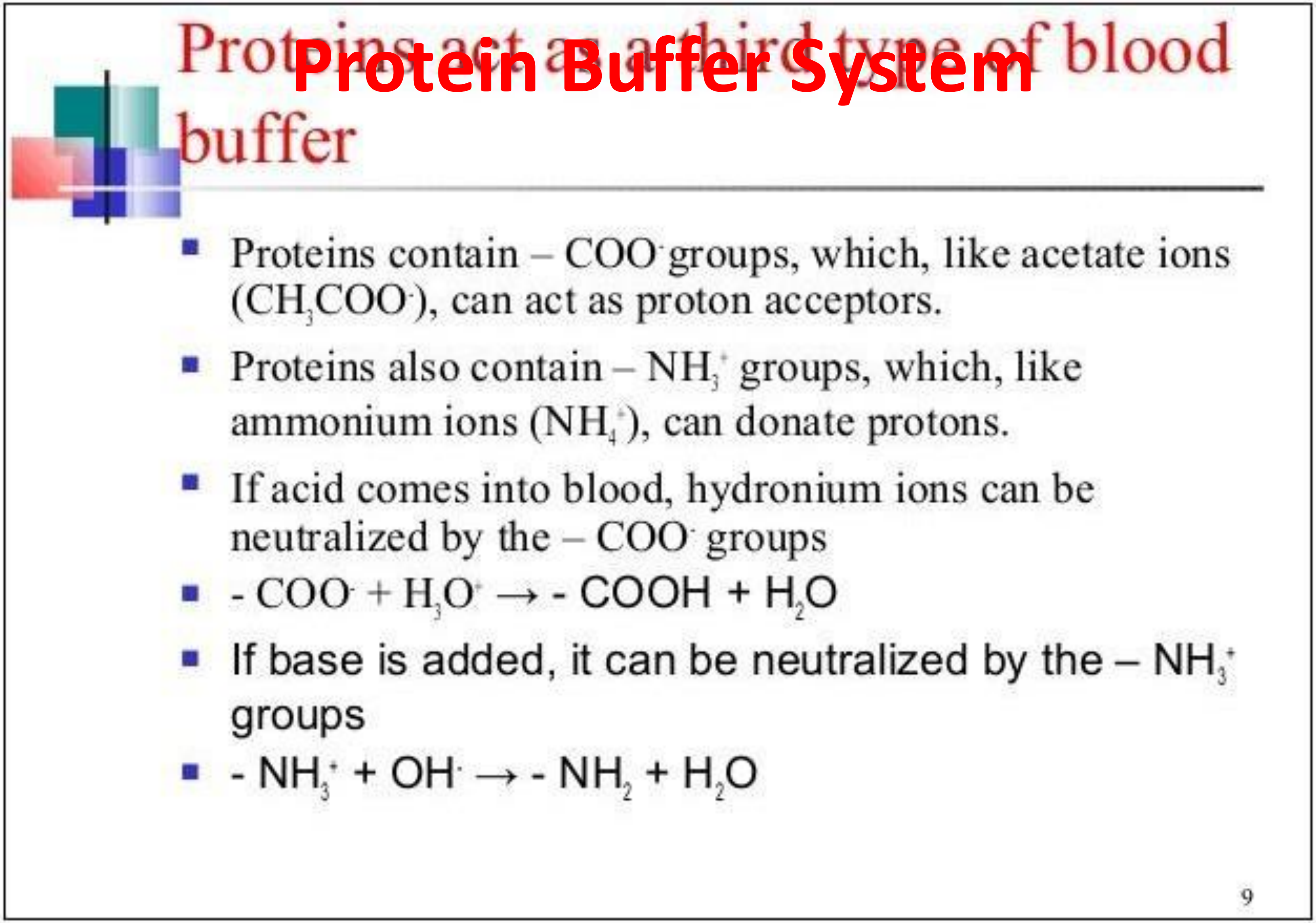
dihydrogen phosphate monohydrogen phosphate

- $H_2PO_4^- + H_2O \leftarrow HPO_4^{2-} + H_3O^+$
- The base is neutralized by dihydrogen phosphate

dihydrogen phosphate monohydrogen phosphate

- $H_2PO_4^- + OH^- \rightarrow HPO_4^{2-} + H_3O^+$

Protein Buffer System



Proteins act as a third type of blood buffer

- Proteins contain – COO^- groups, which, like acetate ions (CH_3COO^-), can act as proton acceptors.
- Proteins also contain – NH_3^+ groups, which, like ammonium ions (NH_4^+), can donate protons.
- If acid comes into blood, hydronium ions can be neutralized by the – COO^- groups
- $-COO^- + H_3O^+ \rightarrow -COOH + H_2O$
- If base is added, it can be neutralized by the – NH_3^+ groups
- $-NH_3^+ + OH^- \rightarrow -NH_2 + H_2O$

Properties, Preparation, Assay & Uses of

- Sodium Chloride
- Potassium chloride
- Calcium gluconate

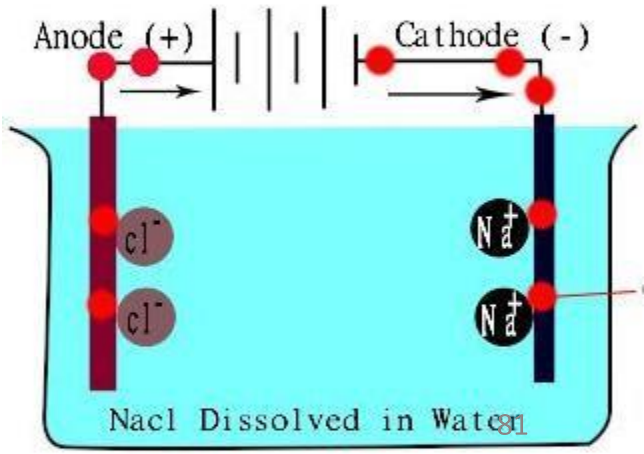
Electrolyte Replenisher

Sodium Chloride (NaCl)

- Sodium chloride is an ionic compound
- It is commonly called as table salt, halite or common salt (99.5% NaCl).
- It is the salt which is mainly responsible for the salinity of the seawater and for the extracellular fluid which is present in many multi-cellular organisms.
- It finds its application from household, medicines to industrial processes.
- Sea water is a major source of this salt.

Properties: NaCl

- It is easily soluble in water and partially in glycerine & alcohol.
- They are white crystals which does not have an odour but possess a taste.
- In its aqueous state NaCl acts as a good conductor of electricity due to the free movement of the ions.
- M.P. 801°C



Preparation of Sodium Chloride:

- 1 mol of sodium bicarbonate reacts with 1 mol of hydrochloric acid to generate 1 mol of salt, 1 mol of water, and/or 1 mol of carbon dioxide.

$$NaHCO_3 + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(aq)} + CO_2$$

Procedure:

- Accurately weigh 5 g of NaHCO3 into evaporating dish.
- Add 5 to 6 mL of distilled water to the dish to wet the bicarbonate. Cover the dish with a watch glass.
- Move the watch glass aside slightly and add, in small portions, about 6 mL of concentrated hydrochloric acid from a 10 mL graduated cylinder.
- After the addition of 6 mL of acid, continue adding acid only as long as CO2 (gas) continues to be evolved.
- Remove the watch glass and evaporate to dryness over a water bath.
- Allow the dish to cool, weigh & collect it out the crystals of NaCl.

- **Assay:** It is analysed by Precipitation Titration (Mohr's method)

$$NaCl + AgNO3 \rightarrow AgCl + NaNO3$$

Sodium chloride reacts with silver nitrate solution using potassium chromate as an indicator

$$2\ AgNO3 + K2CrO4 \rightarrow Ag2CrO4 + 2KNO3$$

Reddish brown coloured

silver chromate

Uses:

- Normal saline (0.9%) that has the same osmotic pressure (isotonic) as body fluids.
- Wet dressings
- Hypotonic solution – when patient unable to take fluid & nutrients orally.
- Hypertonic solution/injections: patients suffers from excessive loss of sodium (1.6% w/v of NaCl).

Potassium Chloride (KCl)

- Colourless, odourless white granular powder or crystals.
- It has a saline taste and is stable in air.
- Soluble in water and insoluble in alcohol.

Assay: It is analysed by Precipitation Titration (Mohr's method)

$$KCl + AgNO_3 \rightarrow AgCl + KNO_3$$

- KCl reacts with silver nitrate solution using potassium chromate as an indicator

$$2\ AgNO_3 + K_2CrO_4 \rightarrow Ag_2CrO_4 + 2KNO_3$$

Reddish brown coloured

silver chromate

Preparation:

Method 1: Potassium chloride can be prepared by treating potassium hydroxide (KOH) or other potassium bases (potassium carbonate, potassium sulphat) with hydrochloric acid:

$$KOH + HCl \rightarrow KCl + H_2O$$

- This conversion is an acid-base neutralization reaction.
- The resulting salt can then be purified by recrystallization.

Method 2:

- By allowing potassium to burn in the presence of chlorine gas (exothermic reaction)

$$2\,K + Cl_2 \rightarrow 2\,KCl$$

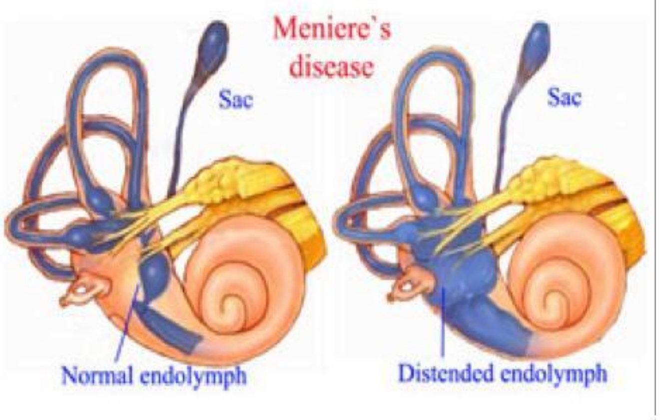
Uses of KCl

- Potassium replacement (hypokalemia or hypochloremic alkalosis condition).
- As an isotonic solution – alone

Or Mixed with NaCl or 5% dextrose solution

- Paralysis
- Menier’s syndrome
- Digitalis intoxication

Note: cautiously given in heart & renal diseases.



Calcium Gluconate

- It appears odourless, tasteless, white crystalline granules or powder.
- Soluble in water and insoluble in alcohol & other organic solvents.
- Its solution remains neutral to litmus.
- Decomposed by dilute mineral acids (HCl) into Gluconic acid and Calcium chloride of the mineral acid used.

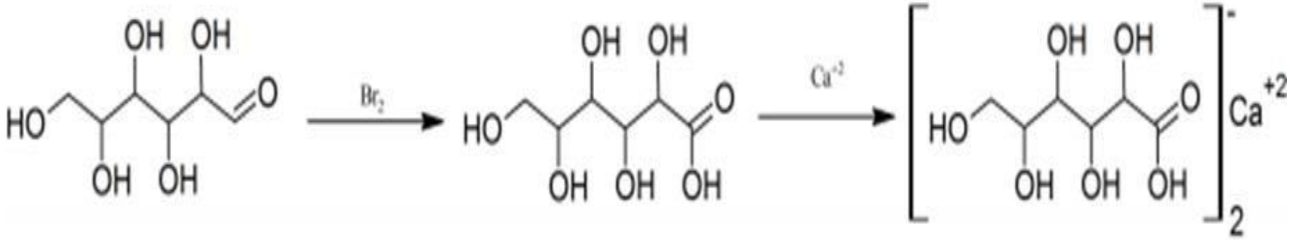
Assay: By Complexometric Titration

- An accurate weighed sample is dissolved in small quantity of water, acidified with dil. HCL.
- To the above solution add 1.0 N NaOH solution, murexide indicator and a solution of naphthol green and titrate against disodium EDTA (Ethylenediamintetraacetic acid) until deep blue colour develops.

Uses: Orally, I.V. or I.M. in the treatment of Hypocalcaemia or in calcium deficiency.

Note:

- Calcium gluconate injection represents 92 – 103% of calcium gluconate .



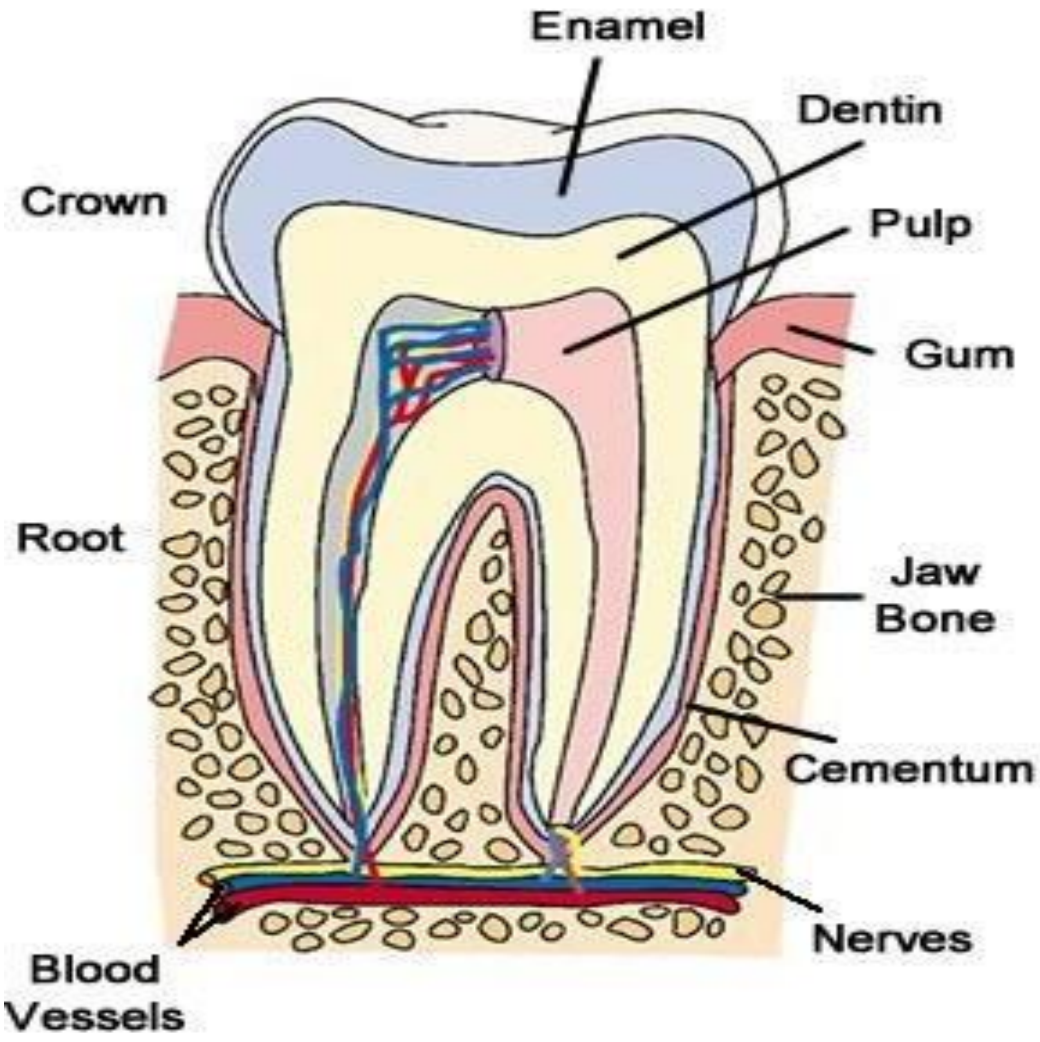
Preparation of calcium gluconate

- To a 200 g of anhydrous glucose in 1000 ml of water, 200 g of bromine are gradually added.
- After the reaction is over the excess of bromine is boiled off and the golden-yellow solution is cooled and the volume measured.
- Add lead carbonate to the above solution - lead gluconate is then formed and this prevents the lead bromide from crystallizing out.
- The resulting mixture is concentrated and allowed to stand in the ice box for 24 hours, after which the lead bromide is filtered off and washed with a little ice-cold water.
- In the presence of silver oxide or silver carbonate, and hydrogen sulfide is passed in to remove minute amounts of lead and silver ions in solution.
- Gluconic acid, is boiled with an excess of calcium carbonate. After cooling, and filtering off the excess of carbonate.
- Filter & concentrate the solution of calcium gluconate.

- Preparation of calcium gluconate
- To a 200 g of anhydrous glucose in 1000 ml of water, 200 g of bromine are gradually added.
- After the reaction is over the excess of bromine is boiled off and the golden-yellow solution is cooled and the volume measured.
- Add lead carbonate to the above solution - lead gluconate is then formed and this prevents the lead bromide from crystallizing out.
- The resulting mixture is concentrated and allowed to stand in the ice box for 24 hours, after which the lead bromide is filtered off and washed with a little ice-cold water.
- In the presence of silver oxide or silver carbonate, and hydrogen sulfide is passed in to remove minute amounts of lead and silver ions in solution.
- Gluconic acid, is boiled with an excess of calcium carbonate. After cooling, and filtering off the excess of carbonate.
- Filter & concentrate the solution of calcium gluconate.

Dental Products

Dental Anatomy……In Brief



Common Dental Problems

- Dental Caries: formed by growth and implantation of cariogenic microorganisms.

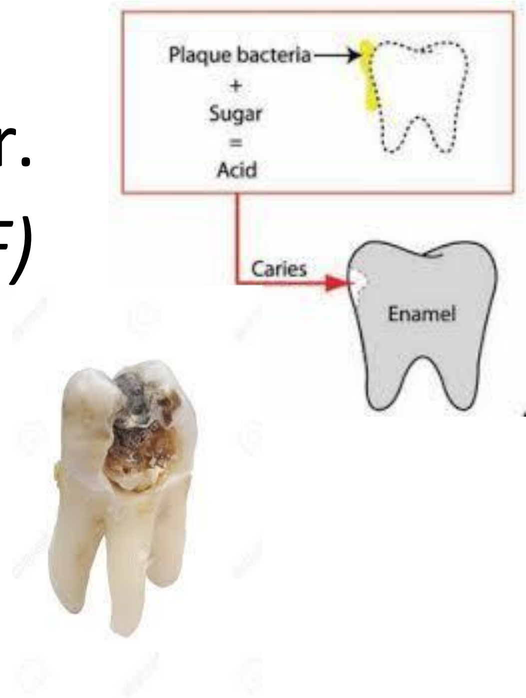
Causes:

- Bacteria – produce acids (lactic acid) that demineralize enamel &
- Food with high concentraiton of sugar.

Treatment: *Products with fluorides (NaF)*

& to alleviate the pain – Oral analgesics

e.g., *Acetaminophen*



- Plaque and Calculus:

-Plaque: sticky substance formed by the attachment of bacteria to the pellicle which is this, cellular glycoprotein.

-Calculus: Calcium salts ppt from saliva & it formes when plaque is not treated within 24 hr.

Treatment:

- Soft, rounded, nylon bristles toothbrushes.
- *Dentifrices*: Toothpastes (Na bicarbonate, Ca carbonate, Dicalcium phosphate & Fluoride)
- *Desensitizing agents*: with 5% potassium nitrate compounds used to reduce the sensitivity of teeth to hot/heat & cold

Gingivitis:

- Inflammation of gingiva caused by microoganisims.

Treatment: eugenol or benzocaine (anesthetics), Mouthwahes, Acetaminophen (Paracetamol)

Lesions:

- Occur on nonkeratinized mucosal surface in mouth
- Treatment: 7 – 14 days – heal,

Protectants, local anesthetics,

wound-cleansing agents

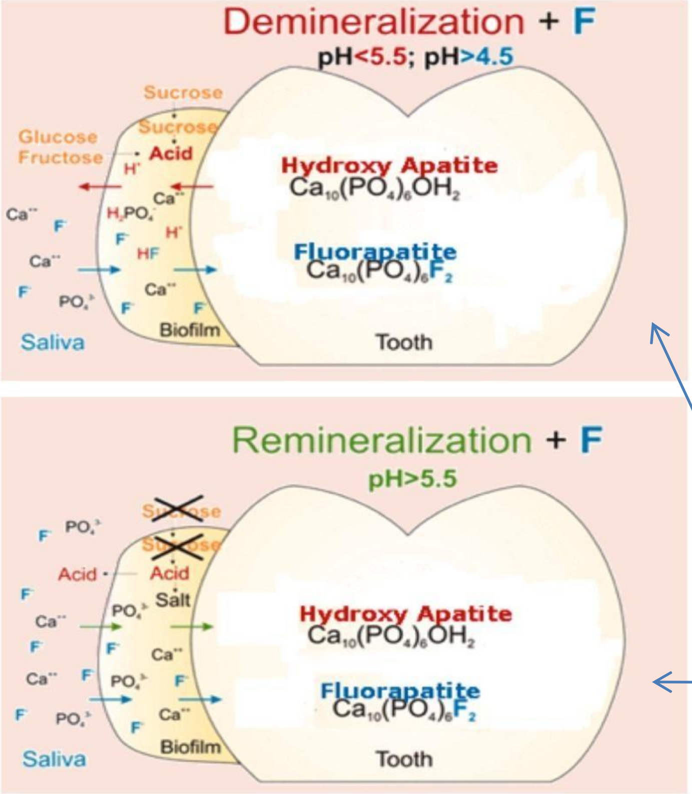
DENTAL PRODUCTS

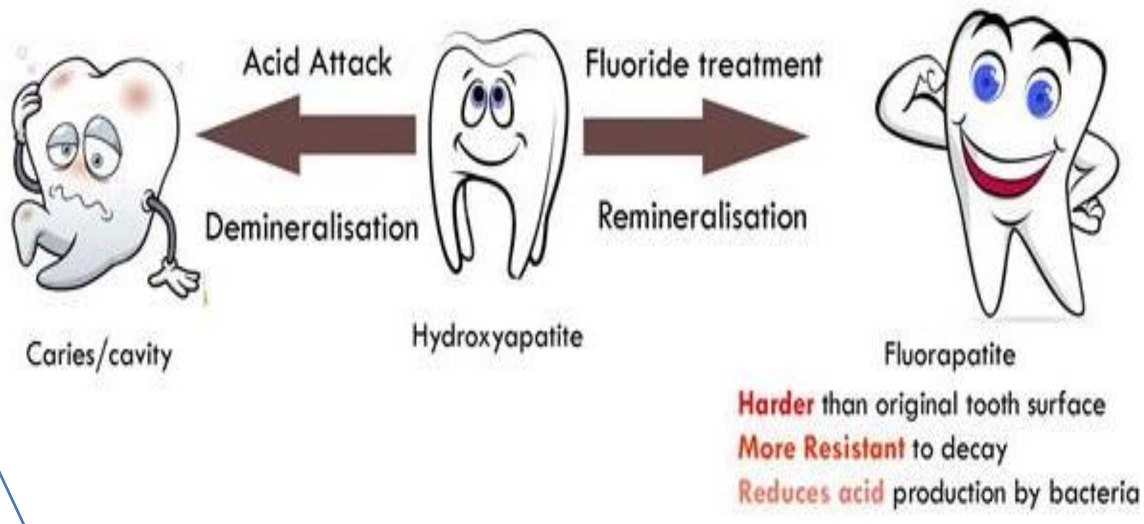
- Products used to treat/prevent various common dental problems.
- Classification:

1. Anticaries Agents: Fluorides, Sodium Flouride (NaF)
2. Dentifrices: Fluorides, Desensitizing agents, Calcium carbonate

Fluorides

- Agent used to prevent the tooth decay (caries) – it get deposited on teeth.
- It is toxic in 5 – 10mg/kg. Acute toxicity causes nausea, vomiting and diarrhea. (Max. Conc. in toothpaste is 250mg/container.





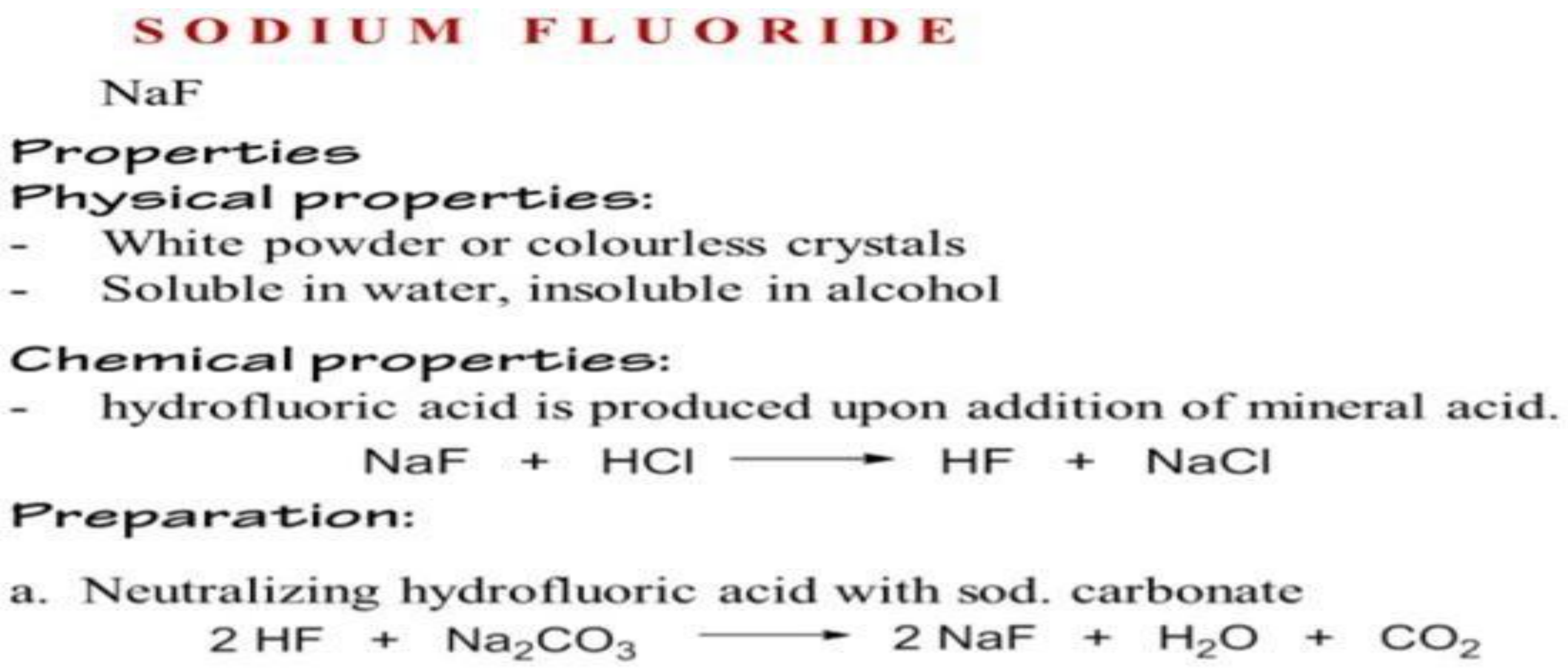
Harder than original tooth surface

More Resistant to decay

Reduces acid production by bacteria

Top – When sugar is present, acid gets produced, pH drops, and minerals leave the tooth surface Bottom – When sugar is not present, acid doesn't get produced, pH rises, and minerals enter the tooth surface

- Administration: Oral (Fluoridation of public water (0.5 to 1 ppm)) & Topical. E.g., Sodium Fluoride (NaF)



SODIUM FLUORIDE

NaF

Properties

Physical properties:

- White powder or colourless crystals
- Soluble in water, insoluble in alcohol

Chemical properties:

- hydrofluoric acid is produced upon addition of mineral acid.

$$NaF + HCl \longrightarrow HF + NaCl$$

Preparation:

a. Neutralizing hydrofluoric acid with sod. carbonate

$$2\,HF + Na_2CO_3 \longrightarrow 2\,NaF + H_2O + CO_2$$

DENTIFRICES

It is a material used to clean the accessible surface of the teeth using preferably a tooth brush.

- Formulation
  - pastes, powders, slurries etc
- Content
  - Phosphate salts
  - Calcium carbonate
  - Mg carbonate
  - Aluminium oxide
  - Silicates
  - Foaming agents
  - Flavoring agents

Desensitizing Agents

- desensitizing agents

  Agents those used in dental preparation to reduce sensitivity of teeth to heat and cold.

- Teeth are sensitive to heat and cold
- On the site of tooth decay there is strong sensitivity of toothache of heat and cold.
- desensitizing agents reduce sensitivity may be by acting as local anesthetic.

Examples

Pumice

It is complex silicates of aluminium, K and Na.

Zinc chloride $ZnCl_2$

CALCIUM CARBONATE $CaCO_3$

CALCIUM CARBONATE

$CaCO_3$

Synonym: Calc carb, Precipitated chalk

Physical properties:

- white, odourless, tasteless, very light powder
- Very slightly soluble in water

Physical properties:

- Calcium carbonate becomes soluble in presence of CO_2 and ammonium salts

$$CaCO_3 + \underbrace{H_2CO_3}_{H_2O + CO_2} \longrightarrow Ca(HCO_3)_2$$

$$CaCO_3 + 2NH_4 \longrightarrow Ca^{++} + 2NH_4 + H_2O + CO_2$$

CALCIUM CARBONATE $CaCO_3$

- It neutralizes acids,

$$CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$$

Preparation:

a. From calcium chloride

- Adding boiling soln. of Sod. Carbonate to calcium chloride
- Ppt of calcium carbonate are washed and purified

$$Na_2CO_3 + CaCl_2 \longrightarrow NaCl + CaCO_{3\ ppt}$$

b. From lime water

- Passing CO_2 through lime water

$$CO_2 + Ca(OH)_2 \longrightarrow H_2O + CaCO_{3\ ppt}$$

CALCIUM CARBONATE $CaCO_3$

Assay:

- Principle: Complexometric titration
- Sample is dissolved in aq. dil. HCl solution
- Boil to remove formed CO_2
- Titrated with std. disodium edetate
- Indicator: Calcon mixture

Use:

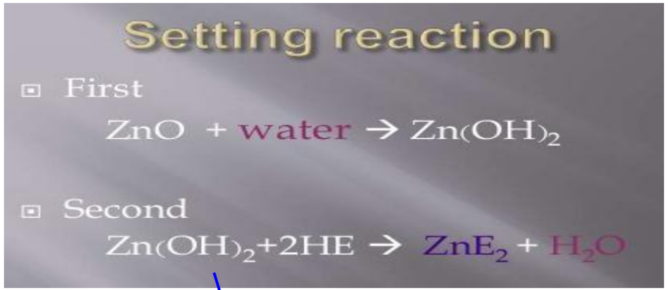
- non-systemic antacid
- Dentrifrices
- Pharmaceutical aid

Storage:

- Well closed containers

ZINC EUGENOL CEMENT

- ZINC OXIDE EUGENOL CEMENT (Zinc eugenolate)
- Introduced in 1858, is used as a Protective, sedative lining in deep carious cavities for
- Temporary Filling
- Temporary Cementing
- Pulp capping
- Root canal filling
- Surgical packing
- Available in powder & liquid form



Setting reaction

- First
  $ZnO + water \rightarrow Zn(OH)_2$
- Second
  $Zn(OH)_2 + 2HE \rightarrow ZnE_2 + H_2O$

Top 5 Tips to Keeping Your Teeth Healthy

Visit your dentist every 6 months without fail

Brush your teeth at least twice a day

Remember to Floss!

Limit the amount of artificial sugar in your diet

Cut down on your drinking and stop smoking

End of Unit II