1. Acid - Defined as a substance that has a sour taste and turns blue litmus paper red in an aqueous solution.

Base - Defined as a substance that has a bitter taste and turns red litmus paper blue in an aqueous solution.

concepts/theories that were proposed by different scientists to define acids and bases:

Arrhenius theory

Bronsted-Lowry theory

Lewis concept

2. Buffers:

Buffers are defined as a compound or mixture of compounds that by their presence in solution resist changes in pH upon the addition of small quantities of acid or alkali.

The resistance to a change in pH is called buffer action. When a small amount of acid or alkali is added to water or sodium chloride solution, pH is altered considerably. Therefore, such systems have no buffer action.

Example – Acetic acid & sodium acetate, ammonium hydroxide and ammonium chloride.

Types of Buffer Solutions:

(i) Acidic Buffer Solution – The solution having a mixture of weak acid (e.g. acetic acid) & its salt (e.g. sodium acetate) is known as acidic buffers.

(ii) Basic Buffer Solution – The solution having a mixture of weak base (e.g. ammonia) & its salt.

(e.g. ammonium chloride) is a basic buffer.

Properties of buffer solutions:

a. The pH does not change even after the addition of small quantities of acids or bases.

b. The pH of the solution does not change on dilution.

c. The pH of buffer solution remains constant.

d. The pH of buffer solution does not change on keeping for a long time.

e. The pH of solution remaining constant is useful in a number of chemical reactions.

Role of buffers in pharmacy:

The buffers play an important role in the pharmaceutical preparations to ensure stable pH conditions for medicinally active compounds.

(i) Solubility

(ii) Colour

(iii) Stability of certain compounds to redox conditions.

(iv) Structure unstable

(v) Patient comfort

(vi) In analytical laboratory, the buffers of known pH for use as standards are used frequently.

Application of Buffers:

The applications remain same for pH & buffers solutions, because buffers are used for maintaining a definite pH of the solution.

Some of the applications are:-

(i) Enhancing solubility

(ii) Increasing stability

(iii) Improving purity

(iv) Optimizing biological activity

(v) Comforting the body.

Buffer equation / common ion effect

Henderson - Hasselbalch equation

The buffer equation is also called as Henderson - Hasselbalch equation. Two separate equations are obtained for each type of buffer, acidic & basic. Buffer equation is developed based on the effect of salt on the ionization of a *weak acid, when the salt and acid have common ion.*

(i) pH of a buffer solution containing weak acid and its salt/ common ion effect.

$$CH_3COOH \rightarrow CH_3COO^- + H_3O^+$$

$$K_a = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]}$$

Buffer equation may be obtained by considering effect of sodium acetate on the ionization of acetic acid.

The dissociation constant for acetic acid is given as:

If sodium acetate is added to the acetic acid solution, it ionizes to produce the acetate ions.

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

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

and hence the K remains unaltered. Rearranging equation:

$$[H_3O^+] = K_a \frac{[CH_3COO^-]}{[CH_3COOH]}$$

(ii) pH of a buffer solution containing weak base and its salt:

Buffer equation may be obtained by considering the effect of ammonium chloride on the ionization of ammonium hydroxide. The salt and base have ammonium ion (NH_4^+) as a common ion.

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

The dissociation of this is given as:

$$K_b = \frac{(NH_4^+)(OH^-)}{(NH_4OH)}$$

$$(OH^-) = K_b \cdot \frac{(NH_4OH)}{(NH_4^+)}$$

Applying log:

$$\log(OH^-) = \log K_b + \log\left(\frac{(NH_4OH)}{(NH_4^+)}\right)$$

Changing the signs:

$$-\log(OH^-) = -\log K_b + \log\left(\frac{(salt)}{(base)}\right)$$

$$pOH = pKb + \log\left(\frac{(salt)}{(base)}\right)$$

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or (OH^-) = Kb (NH_4OH) = Kb (base) / (NH_4^+) (salt)

$$\frac{K_w}{[H_3O^+]} = \frac{Kb\cdot(base)}{(salt)}$$

$K_w = (H_3O^+) \times (OH^-)$

and $OH^- = Kw/H_3O$

where, Kw = Ionic product of water expressing in logarithmic terms,

$$\log K_w - \log(H_3O^+) = \log Kb + \log(base) - \log(salt)$$

Rearranging,

$$-\log(H_3O^+) = -\log K_w + \log Kb + \log(base) - \log(salt)$$

$$pH = pK_w - pKb + \log\left(\frac{base}{salt}\right)$$

but,

$$pH = pK_w - pKb = pKa$$

$$\boxed{pH = pKa + \log\left(\frac{base}{salt}\right)}$$

4. Buffer Capacity

The magnitude of resistance of a buffer solution to pH change is known as buffer capacity. The buffer capacity may be defined as the equivalent of a strong base (acid) that should be added to 1 liter of the buffer solution to bring about a unit change in the pH. The buffer capacity is also referred to as buffer coefficient, buffer index, or buffer value

$\beta = \Delta B / \Delta pH$

β. It is given by:

Here, β - Buffer capacity.

ΔB - Small increment in gram equivalent of strong base (or acid) added to 1 liter of the buffer solution, which brings in a change of 1 pH unit.

Buffers in pharmaceutical and biological systems:

In vivo biological buffers:

a. Blood - pH is maintained at about 7.4 by the primary buffers present in plasma as well as by secondary buffers in the erythrocytes.

b. Urine - Normal adult urine has an average pH of about 6 units and it may be as low as 4.5 or as high as 7.8.

c. Parenterals - Parenterals should be buffered only when required. Parenterals given by IV route are slowly injected.

d. Lacrimal Fluid - Lacrimal fluid has great buffering capacity having a pH of 7.4. The pH of solution meant for introduction into the eye may vary from 4.15 to 11.5.

e. Tissue or Skin - Solution applied to the skin may cause irritation if their pH is different. In such cases, the substance must be buffered to pH 7.4.

Example - Ointments, creams, gels, lotions, transdermal patches.

iii. Pharmaceutical Buffers:

a) Buffers in tablet formulations.

b) Buffers in ophthalmic preparations.

c) Buffers in parenteral preparations.

d) Buffers in creams and ointments.

6. Buffer to Isotonic Solution:

Isotonic buffered solution is defined as a solution that maintains the isotonicity and the pH as that of the body fluids. Isotonic buffer solution must be compatible with the body fluids for the following reasons:

(i) Blood and lacrimal fluids are in vivo buffer systems. Any solution that comes in contact with these fluids should be buffered to a desired pH so that these are compatible with body fluids.

(ii) In addition, some solutions are meant for the application on delicate membranes of the body. Such solutions may cause hemolysis, tissue damage, or irritation if not properly buffered.

Applications:

- Aqueous solutions used as nasal drops.
- Ophthalmic drops.

Buffered Isotonic Solutions

Tonicity

The word tonicity is simply defined as the concentration of a solution as compared to another solution.

Now, in pharmacy, the pharmaceutical buffer solutions that are meant for application inside the body must have the same osmotic pressure or the same concentration as that of the body fluids/blood.

Tonicity/concentration of blood = **0.9% w/v of NaCl**

Types of Solutions (As Per Tonicity)

There are basically three types of solutions:

- Isotonic
- Hypotonic
- Hypertonic

Isotonic Solution

A buffer solution whose concentration/osmotic pressure is equal to the 0.9% w/v of NaCl is known as a "Buffer Isotonic Solution."

Hypotonic Solution

A buffer solution whose concentration/osmotic pressure is less than 0.9% w/v of NaCl is known as a "Hypotonic Solution."

Hypertonic Solution:

Buffered Isotonic Solutions whose concentration or osmotic pressure is greater then 0.9% w/v of NaCl are known as "Hypertonic Solutions".

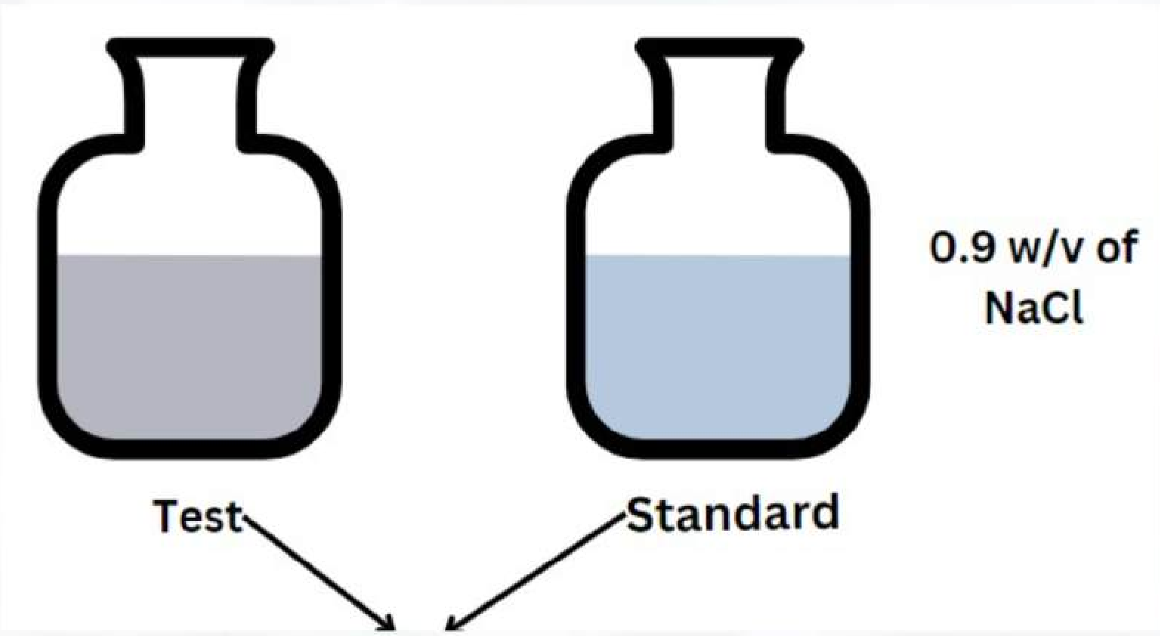
Measurement/Determination of Tonicity

- Cryoscopic/Colligative Method
- Haemolytic Method

Cryoscopic/Colligative Method

This method is based on the colligative properties of the solution such as freezing point, boiling point, vapor pressure, and temperature difference.

In this method, we basically compare the colligative properties of our test solution (whose tonicity has to be determined) with the standard isotonic solution (0.9% w/v of NaCl).



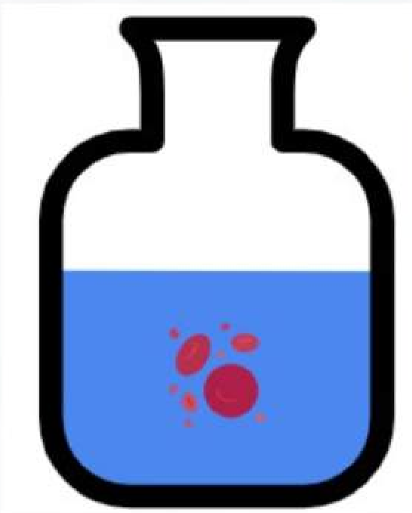
After Comparison:

- If Test = Standard → Isotonic
- If Test < Standard → Hypotonic
- If Test > Standard → Hypertonic

Haemolytic Method

In this method, we determine the tonicity of the solution based on the appearance of red blood cells suspended in the solution.

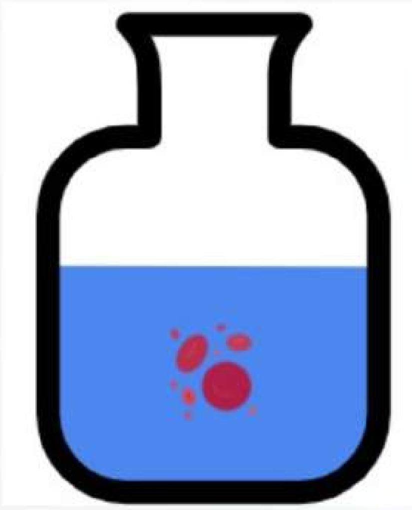
We know that according to osmosis, solvent particles move from an area of low concentration to an area of high concentration. In this method, we first dissolve the red blood cell in the given test solution, then the following 3 conditions can occur:



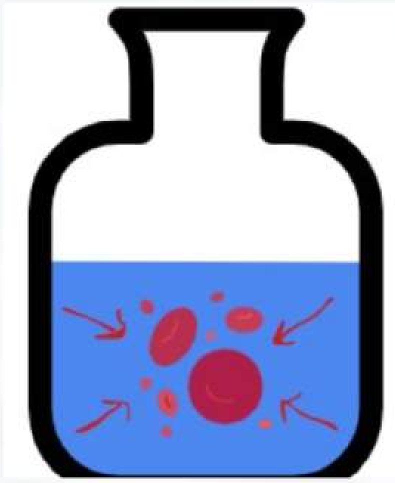
Condition I (Cell Shrinkage)

If the concentration of the solution is greater than the concentration of blood cells, the solvent moves from blood to the solution, causing cell shrinkage. This solution will termed Hypertonic.

Condition II (No Change): If the concentration of the solution is equal to the concentration of blood cells, then there will be no net movement of solvent, and due to this, there will be no change in the size of the blood cell, or it will remain constant. This solution will be termed Isotonic.



Condition III (Cell Swelling): If the concentration of the solution is less than the concentration of blood cells, then solvent particles move from the solution to the blood cells, causing cell swelling. This solution will be termed Hypotonic.



Methods of Adjusting Tonicity

It basically includes two classes:

- Class I (For Hypotonic)
- Class II (For Hypertonic)

Class I

Cryoscopic/Freezing Method: This method is basically used to adjust the tonicity of a hypotonic solution. In this method, we add sodium chloride to make the solution isotonic.

It includes:

- Cryoscopic/Freezing Point Depression Method
- Sodium Chloride Equivalent Method

Cryoscopic/Freezing Point Depression Method:

$w/V = 0.52 - a/b$

Where:

w = amount of adjusting substance
a = freezing point of 1% solution of un-adjusted solution
b = freezing point of 1% solution of adjusting substance

Sodium Chloride Equivalent Method:

$E = 17 \times Liso/M$

Class II

This method is basically used to adjust the tonicity of hypertonic solutions. In this method, we basically add water to make the solution isotonic. This includes:

0. White-Vincent Method

0. Sprowls Method

1. White-Vincent Method

V=W×E×111.1

Where:

V = Volume of isotonic solution prepared by mixing the drug with water
W = Weight of drug in grams
E = Sodium chloride equivalent

2. Sprowls Method

This is basically the simplification of the White-Vincent method. In this, we set the value of W = 0.3.

V=0.3×E×111.1

or

V=33.33×EV

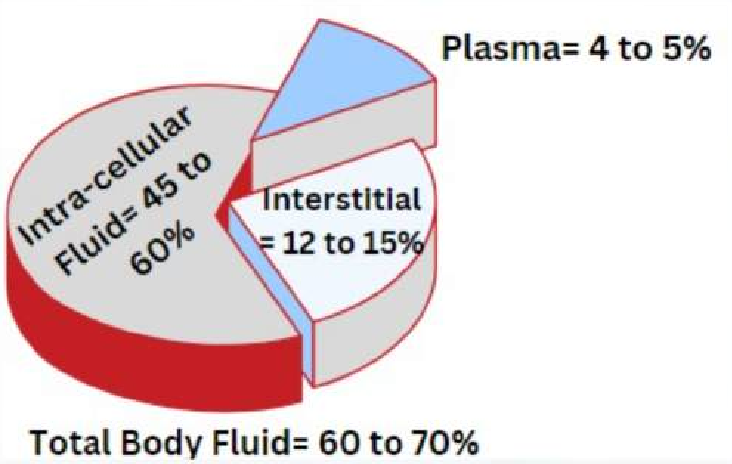
Major Intra & Extracellular Electrolytes

Body Fluids

60-70% part of the body consists of fluid (water).

The body fluid can be divided into two compartments:

- Intracellular Fluid (ICF) (45-60%)
- Extracellular Fluid (ECF) (20-25%)
- Interstitial Fluid (12-15%)
- Plasma (4-5%)



Electrolytes

Electrolytes are substances that, when dissolved in aqueous solutions or body fluids, dissociate into ions (cations and anions).

Electrolytes present inside the body can be further divided into two categories:

- Intracellular Electrolytes (Present inside the cell)

- Extracellular Electrolytes (Present outside the cell)

Major Physiological Ions

Cation

- Sodium ion (Na^+)
- Potassium ion (K^+)
- Calcium ion (Ca^{2+})
- Magnesium ion (Mg^{2+})

Anion

- Chloride (Cl^-)
- Bicarbonate (HCO_3^-)
- Phosphate (PO_4^{3-})
- Sulfate (SO_4^{2-})

Sodium

- It is present in the most abundant amount in extracellular fluid.
- It transmits nerve impulses in the nerve fibers.
- It associates with chloride and bicarbonate and regulates the acid-base balance of the body.
- It also protects the body against excessive fluid loss.
- A low level of sodium leads to Hyponatremia.
- High level of sodium leads to Hypernatremia.

Potassium

- It is present in the most abundant amount in intracellular fluid.
- It plays a major role in the contraction of muscles, especially cardiac muscles.
- It performs various biological activities inside the cell.
- It also helps in the transmission of nerve impulses.
- A low level of potassium leads to Hypokalemia.
- High level of potassium leads to Hyperkalemia.

Calcium

- It is mainly found in bones (approximately 98%) with the remaining found in extracellular fluid.
- It is essential in the clotting of blood.
- It helps in the contraction of various smooth muscles.
- A low level of calcium leads to Hypocalcemia.
- High level of calcium leads to Hypercalcemia.

Magnesium

- It is considered the second most common intracellular electrolyte.
- It helps in the formation of bones and teeth.
- It also plays an important role in myocardial function.
- A low level of magnesium leads to Hypomagnesemia.
- High level of magnesium leads to Hypermagnesemia.

Chloride

- It is mainly present in the extracellular fluid.
- It helps to maintain acid-base balance.
- It also helps to maintain the osmotic pressure of the body.
- The main source of chloride is common salt, which is used in cooking.
- A low level of chloride leads to Hypochloremia.
- High level of chloride leads to Hyperchloremia.

Phosphate

- It is present mainly in the intracellular fluid.
- It helps to maintain the acid-base balance of the body.
- The main dietary sources of phosphate are milk, nuts, etc.
- Low levels of phosphate lead to Hypophosphatemia.
- High levels of phosphate lead to Hyperphosphatemia.

Bicarbonate

- It is present in extracellular fluid.
- Along with carbonic acid, it acts as one of the most important buffer systems of the body, maintaining acid-base balance.
- It also protects tissues of the central nervous system.

Sulfate

- It is present in very small amounts in extracellular fluid.

- They play a vital role in the detoxification mechanism.
- It also helps in various biological processes.

Replacement Therapy

In different abnormal conditions like diarrhea, vomiting, dehydration, electrolytes in our body get imbalanced.

The main purpose of electrolyte replacement therapy is to overcome the electrolyte imbalance and restore the composition of body fluid and body volume.

There are three compounds used as the major sources of electrolytes:

- Sodium Chloride
- Potassium Chloride
- Calcium Gluconate

Sodium Chloride:

Molecular Weight: 58.44
Molecular Formula: NaCl
Synonyms: Rock salt, Table salt, Common salt

Method of Preparation

Sodium chloride can be obtained from natural sources as well as prepared in a laboratory.

Natural Source: Naturally, it can be obtained from rock salt and seawater, but from these sources, the obtained NaCl is in impure form. The pure form of salt can be obtained by the process of filtration and evaporation.

Laboratory Method:

Sodium chloride can be prepared in the laboratory on a small scale by the acid-base reaction. In this process, strong acid (HCl) reacts with a strong base (NaOH) and finally gives sodium chloride.

$HCl + NaOH \rightarrow NaCl + H_2O$

Properties

1. Physical Properties

State: Powder or Crystalline form
Colour: White or colorless
Taste: Saline / salty
Odour: Odourless
Solubility: Soluble in water but insoluble in alcohol.

2. Chemical Properties

It reacts with silver nitrate and forms a white precipitate of silver chloride.
$NaCl + AgNO_3 \rightarrow AgCl + NaNO_3$

It reacts with sulfuric acid and gives hydrochloric acid.
$2NaCl + H_2SO_4 \rightarrow 2HCl + Na_2SO_4$

Uses

- It is used as an electrolyte replenisher.
- Its 0.9% solution is isotonic.
- It is used as a taste enhancer and diuretic.

Assay:

Its assay is based on the Argentometric Titration.

Procedure:

- Weigh 1 gm of sample and dissolve it in 50 ml of water.
- Now add 50 ml of 0.1 M silver nitrate.
- To this, add 5 ml of 2M nitric acid and 2 ml of concentrated $KMnO_4$.
- Shake this properly and titrate with 0.1 M ammonium thiocyanate using 2 ml ferric ammonium sulfate as an indicator.
- Titration continues until a reddish-brown color appears.

Potassium Chloride

Molecular Formula: KCl

Molecular Weight: 74.55

Synonyms: Potassium muriate

Preparation

- It can be obtained by separation and purification of its minerals like carnallite ($KCl, MgCl_2 \cdot 6H_2O$)
- In the laboratory, it can be obtained by reacting potassium carbonate and hydrochloric acid. ($K_2CO_3 + 2HCl \rightarrow 2KCl + H_2O + CO_2$)

Properties

- It occurs as colorless or white crystalline powder.
- It is odourless.

- It has a saline / salty taste.
- It is freely soluble in water and insoluble in alcohol.
- KCl is used to produce metallic potassium, by reducing KCl with metallic sodium at 850°C. ($KCl+Na \rightarrow NaCl+K$)

Uses

- It is used as an electrolyte replenisher.
- It is used in the case of potassium deficiency.
- It is used as a substitute for sodium chloride salt.
- It is also used in digitalis poisoning.

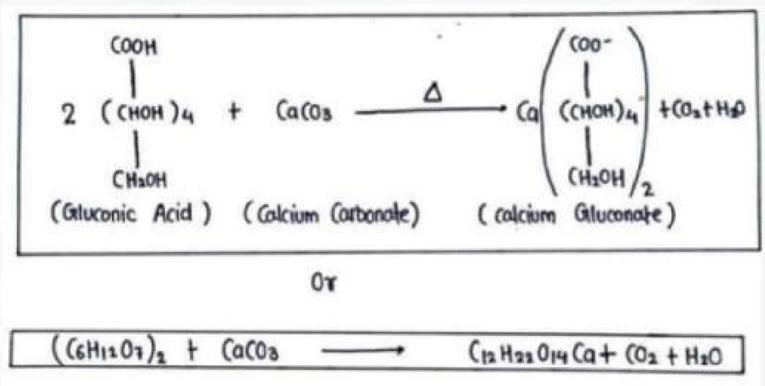
Calcium Gluconate:

Molecular Formula: $C_{12}H_{22}O_{14}Ca \cdot H_2O$

Molecular Weight: 430.373

Synonyms: Calcium salt, D-gluconic acid

Preparation: Calcium gluconate is prepared by boiling the solution of gluconic acid with excess calcium carbonate.



Properties:

- It appears in the form of white crystalline granules or powder.
- It is odourless and tasteless.
- It is soluble in water and insoluble in alcohol.

Assay:

The assay of calcium gluconate is based on complexometric titration.

Procedure:

- Weigh 0.5 g of the sample and dissolve it in 50 mL of warm water.
- Now add 5.0 mL of 0.05 M magnesium sulfate and 10 mL of strong ammonium solution.
- The resulting solution is titrated against 0.05 M disodium EDTA until a deep blue color develops.

Uses:

- It is used as an electrolyte replenisher.
- It is used in the case of calcium deficiency.
- It plays a vital role in bone building and development.

-----------------x---------------------------------x--------------------------

ORS (Oral Rehydration Salt)

The full form of ORS is Oral Rehydration Salt.
It is also known as Oral Rehydration Therapy (ORT).
It is a type of fluid replacement used mainly in the treatment of dehydration, particularly due to diarrhea.
ORS is a cheap, simple, and effective way to treat dehydration caused by diarrhea.
ORS drinks contain the main elements that are lost from the body during diarrhea.

Principle of ORS

Glucose, when given orally, enhances the intestinal absorption of salt and water, and thus maintains the electrolyte and water balance.

Formula of ORS

The formula of ORS recommended by WHO and UNICEF:

2.6 g/L sodium chloride
2.9 g/L trisodium citrate dehydrate
1.5 g/L potassium chloride
13.5 g/L glucose

Total weight = 20.5 g

Equipment Needed

- Take one liter of boiled and cooled drinking water.
- A clean glass of 200 mL capacity.
- A clean vessel to mix the solution.
- A clean spoon to mix the solution and feed the child.

Procedure

- Wash hands.
- Take one liter of clean water into a vessel.
- Open the ORS packet and pour all the content into the vessel.
- Mix the ORS into the water.
- Take some solution in a clean glass.
- Feed the child frequently with small doses of the solution.

Physiological Acid-Base Balance

Acid-base balance is a part of the homeostasis process that deals with the maintenance of pH.

Most of the reactions in our body occur only in a specific pH range; any change in this pH can cause major disturbances.

The normal pH value of blood is approximately 7.4, and the survival range of pH in the blood is between 6.8 to 8.0. If the pH limit crosses this value, it may lead to death, so it becomes very important to maintain the pH balance of our body.

System that Regulates pH Balance

- Buffer System
- Respiratory System
- Renal System

Buffer System

The buffer system converts strong acid and base into weak acid and weak base so that they do not allow rapid and drastic changes in pH.

There are three major buffer systems in our body that regulate the acid-base balance:

- Bicarbonate buffer system
- Phosphate buffer system
- Protein buffer system

1. Bicarbonate Buffer System

Function: It is an important regulator of blood pH.

Location: Occurs in plasma and kidneys.

Case I: When there is an excess of H^+ ions (which increases acidity), bicarbonate ions (HCO_3^-) combine with H^+ to convert into weak carbonic acid (H_2CO_3), which is a weak acid.

$$H^+ + HCO_3^- \leftrightarrow H_2CO_3$$

Case II: When there is shortage of (H^+) then carbonic acid ionized to release H^+ ions to maintains the pH.

$$H_2CO_3 \leftrightarrow H^+ + HCO_3^-$$

2. Phosphate Buffer System

Function: Phosphate buffer system is found in intestinal fluid because phosphate concentration is highest in intestinal fluid, intracellular fluid.

Components: This system consists of Monohydrogen Phosphate ions (HPO_4^{2-}) and Dihydrogen Phosphate ions ($H_2PO_4^-$).

Case I: When there is an excess of H^+ ions (which increases acidity), Na_2HPO_4 combines with H^+ and converts into NaH_2PO_4.

$H^+ + Na_2HPO_4 \leftrightarrow NaH_2PO_4$

Case II: When there is a shortage of H^+ ions (which increases basicity), NaH_2PO_4 ionizes to release H^+ ions to maintain the pH balance.

$NaH_2PO_4 \leftrightarrow Na_2HPO_4 + H^+$

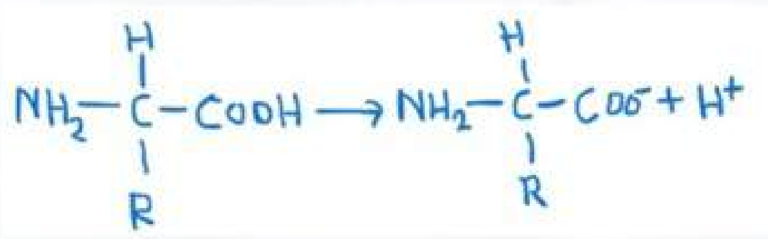
3. Protein Buffer System

Function: It is an important buffer system in blood and plasma.

Composition: Proteins are made up of amino acids, and amino acids contain one carboxyl group (-COOH) and one amino group ($-NH_2$).

Case I: In the case of excess H^+ ions, the amino group acts as a base and accepts or combines with H^+ ions.

Case II: In the case of a shortage of H^+ ions, the carboxyl group releases an H^+ ion to maintain the pH balance.



Respiratory System

Function: The increase or decrease of CO_2 is responsible for disturbances in the pH of the body's internal environment. The respiratory system works by increasing or decreasing the breathing rate in our body.

Case I:

CO_2 Increases → Acidity Increases → Breathing Rate ↑ → Excess Removal of CO_2 → pH Balanced

Case II:

Basicity Increases → Breathing Rate ↓ → Retention of CO_2 → pH Balanced

Renal System

Function: It is the most effective regulator of pH.

pH of Urine: The pH of urine is normally acidic (nearly 6.0).

Process: When the amount of H^+ increases in our body, it is eliminated from our body through urine, while bicarbonate ions (HCO_3^-) are reabsorbed in our body. This maintains the acid-base balance.

Dental Products:

Purpose: Dental products are used to maintain dental hygiene and to prevent the decay of teeth, giving freshness and cleanliness to the teeth and mouth.

Availability: There is a wide range of dental products available in the market.

Classification of Dental Products:

- Anticaries Agent
- Dentifrices
- Desensitizing Agent
- Cement and Fillers
- Anticaries Agent

1. Anticaries Agent:

Definition: Dental Caries is the medical term for tooth decay or cavity.

Cause: Dental caries or tooth decay is caused by acids produced by the action of microorganisms on carbohydrates.

Process: The disease is characterized by the decalcification of teeth.

Process Breakdown:

Some Bacteria/Microorganism>Turns Carbohydrates into Acids>Decalcification>Dental Caries/Cavity

Dental Caries Prevention:

Prevention: Dental caries can be prevented by maintaining oral and dental hygiene.

Anti-caries Agents: These are chemical compounds used to prevent dental caries produced by the action of microorganisms.

Current Use: Fluoride is the main anti-caries agent used in the treatment of dental caries.

Role of Fluoride in the Treatment of Dental Caries:

Common Agent: Fluoride is the most commonly used anti-caries agent.

Natural Occurrence: Fluoride occurs naturally in our body and is also found in small amounts in a variety of foods.

Absorption: When a fluoride-containing salt or solution is ingested, it is readily absorbed, transported, and deposited in the bones or developing teeth, and remains excreted by the kidney.

Protection: The deposited fluoride on the surface of teeth prevents the action of acids or enzymes in producing cavities.

Necessary Quantity: A small quantity (1 PPM) of fluoride is necessary to prevent dental caries.

Excessive Intake: However, if more than 2-3 ppm is ingested, it can be carried to bones and teeth and cause dental fluorosis.

Administration of Fluoride:

Forms of Administration: Fluoride can be administered both internally and topically for the prevention of dental caries.

Oral Administration: It can be given in drinking water or juice at about 1 ppm/day. Sodium fluoride tablets at a dose of 2 mg per day are also used.

Topical Application: A 2% solution is generally used on teeth.

Sodium Fluoride:

Chemical Formula: NaF

Molecular Weight: 41.99 g/mol

Method of Preparation:

It can be prepared by neutralizing HF with Na_2CO_3. $2HF+Na_2CO_3 \rightarrow 2NaF+H_2O+CO_2$

It can also be prepared by double decomposition of calcium fluoride with sodium carbonate, where insoluble calcium carbonate can be removed by filtration.

$CaF_2+Na_2CO_3 \rightarrow 2NaF+CaCO_3$

Properties

It occurs as colorless, odorless crystals or as white powder.

It is soluble in water but insoluble in alcohol.

Its aqueous solutions corrode ordinary glass bottles, so it is prepared in distilled water and stored in dark pyrex bottles.

Uses

Dental Use: It is used in the prevention of dental caries.

Insecticide: It is also used as an insecticide.

Toothpaste: It is used in the preparation of toothpaste.

Storage:

Conditions: Store in well-closed, tight containers in a dark place.

2. Dentifrices:

- Dentifrices are products used for cleaning teeth and adjacent gums.
- They can be used with fingers or toothbrushes.

- Available as paste as well as powders.
- The effectiveness of dentifrices depends on the abrasive property and the rubbing force used.

Properties of Dentifrices:

- Responsible for the physical removal of plaque.
- A good dentifrice must remove stains from teeth and provide freshness to the mouth.

Drawbacks: Dentifrices are not able to clean surfaces inside cavities.

Calcium Carbonate:

Chemical Formula: $CaCO_3$

Molecular Weight: 100.09 g/mol

Synonym: Precipitated Chalk

Method of Preparation:

Commercial Preparation: Calcium carbonate is obtained by mixing the boiling solution of calcium chloride and sodium carbonate and allowing the resulting precipitate to settle down.

$CaCl_2 + Na_2CO_3 \rightarrow CaCO_3 + 2NaCl$

Properties:

Appearance: Occurs as white crystalline powder.

Odor: Odorless.

Taste: Tasteless.

Solubility: Soluble in dilute HCl and HNO_3, but insoluble in water and alcohols.

Uses:

- Used as a dentifrice and polishing agent.
- Other Uses: Also used as an insecticide and as an antacid.

3. Desensitizers:

The teeth are usually sensitive to heat and cold. During tooth decay, the perception of heat and cold has been felt strongly. Desensitizing agents reduce the pain in sensitive teeth caused by heat or cold; they reduce the sensitivity of teeth. They act as local anesthetics.

Zinc Chloride:

Chemical Formula: $ZnCl_2$

Molecular Weight: 136.28 g/mol

Method of Preparation:

It is prepared by heating granulated zinc with HCl. ($Zn+2HCl \rightarrow ZnCl_2+H_2$)

Properties:

- It occurs as a white crystalline powder.
- It is odourless.
- It is soluble in both water and alcohol.

Uses:

- It is used as desensitizer.
- It is also used as an antiseptic.
- It is also used in dental fillinos.

4. Cement and Fillers:

- Dental cements are used to temporarily cover and protect areas that have undergone operations in dental surgery.
- The cementing material is applied as a paste which gets hardened and forms a protective layer.
- After healing of the area, the cement can be removed by the dentist.

Zinc Eugenol Cement:

- Zinc oxide eugenol cements have been used in dentistry since 1830.
- It is considered as the best cementing material in dental practice.
- They are the cement of low strength.
- They are the least irritating of all dental cements.

Composition:

It is mainly composed of:

- Eugenol
- Olive Oil / Clove Oil
- Zinc Oxide

Properties:

- It is the cement of low strength, used for temporary fillings.
- It contains eugenol that provides mild antiseptic and anesthetic effects.