BP106T · B.PHARM THEORY

Pharmaceutical Inorganic and Analytical Chemistry

Unit II — Acids, Bases, Buffers, Isotonicity & Electrolytes

Review Draft

Pills & Python Notes
From-zero teaching standard: every technical term is defined before use.
Syllabus coverage: Unit II (8 hours).

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CHAPTER 5: ACIDS, BASES, BUFFERS AND pH

Unit II · ~3 hours

5.1 Introduction

Why does an injection not sting like lemon juice? Why does blood stay at exactly pH 7.4 even after you eat acidic food? The answers are **pH** and **buffers**. Almost every medicine — injections, eye drops, syrups — is formulated at a carefully chosen pH, and buffers hold it there. For a pharmacist, pH decides drug stability, drug absorption, and patient comfort.

Learning objectives — after studying this chapter, you should be able to:

- define acids, bases, and buffers (scientific + simple);
- compare the Arrhenius, Brønsted–Lowry, and Lewis theories;
- explain the pH scale and calculate pH of simple solutions;
- classify buffers and explain buffer action with equations;
- derive and apply the buffer equation (Henderson–Hasselbalch);
- explain buffer capacity and its pharmaceutical uses.

5.2 Definition

Acid — scientific: a substance that donates protons (H^+) — Brønsted–Lowry definition most used in pharmacy. **Acid — simple:** a substance that releases H^+ ions in water and tastes sour (e.g., citric acid in lemons, HCl in the stomach).

Base — scientific: a substance that accepts protons (H^+). **Base — simple:** a substance that neutralises acids, often feels soapy (e.g., NaOH, ammonia).

Buffer — scientific: a solution that resists change in pH when small amounts of acid or base are added. **Buffer — simple:** a solution that "absorbs" added acid or base so the pH hardly moves — the pH shock-absorber.

pH — scientific: the negative logarithm of hydrogen-ion concentration: $pH = -\log_{10}[H^+]$. **pH — simple:** a 0–14 scale telling how acidic or alkaline something is; 7 is neutral, below 7 acidic, above 7 alkaline.

5.3 Key Terminology

Term	Meaning
Conjugate acid–base pair	Two species differing by one H^+ (e.g., CH_3COOH / CH_3COO^-)
pKa	−log Ka; the pH at which acid and conjugate base are equal — a buffer works best near its pKa
Buffer capacity	Moles of acid/base per litre needed to change pH by 1 unit — the buffer's "strength"
Strong acid/base	Dissociates completely in water (HCl, NaOH)
Weak acid/base	Dissociates only partly (acetic acid, ammonia) — the building blocks of buffers

5.4 Detailed Explanation

5.4.1 Three theories of acids and bases

Theory	Acid	Base	Example
Arrhenius	Produces H^+ in water	Produces OH^- in water	HCl (acid), NaOH (base)
Brønsted–Lowry	Proton (H^+) **donor**	Proton (H^+) **acceptor**	$NH_4^+ \rightarrow NH_3 + H^+$ (NH_4^+ is the acid)
Lewis	Electron-pair **acceptor**	Electron-pair **donor**	BF_3 (acid) + NH_3 (base) → adduct

Exam Pearl: Brønsted–Lowry is the workhorse of pharmacy (explains NH_4^+/NH_3, HCO_3^-/CO_3^{2-}). Lewis is the broadest — it covers reactions with no H^+ at all.

Conjugate pairs: when an acid donates H^+, what remains is its **conjugate base**: CH_3COOH (acid) ⇌ CH_3COO^- (conjugate base) + H^+; NH_3 (base) + H^+ ⇌ NH_4^+ (conjugate acid).

5.4.2 The pH scale

pH = −log[H^+]. Because it is logarithmic, **each pH unit = 10× change** in [H^+]: pH 3 is ten times more acidic than pH 4.

pH	[H^+]	Example
1	10^{-1} M	Gastric juice
4.5	~3×10^{-5} M	Orange juice
7.0	10^{-7} M	Pure water (neutral)
7.4	4×10^{-8} M	**Human blood**
9	10^{-9} M	Soap solution

Worked examples:

1. *pH of 0.01 M HCl?* HCl is strong → [H^+] = 0.01 = 10^{-2} → **pH = 2**.
2. *pH of 0.01 M NaOH?* [OH^-] = 10^{-2} → pOH = 2 → **pH = 14 − 2 = 12**.
3. *Solution with [H^+] = 1×10^{-5} M?* **pH = 5** (acidic).

Significance of pH: drug **stability** (many drugs degrade outside their pH range), drug **absorption** (weak acids absorb best in acidic stomach), **patient comfort** (injections near pH 7.4 sting less), and **physiological function** (blood enzymes work only at pH ~7.4).

5.4.3 Buffers — types and action

A buffer = **weak acid + its conjugate base** (acidic buffer, e.g., CH_3COOH/CH_3COONa) or **weak base + its conjugate acid** (basic buffer, e.g., NH_4OH/NH_4Cl).

How buffer action works (acetic acid/acetate system):

- Add acid (H^+): $CH_3COO^- + H^+ \rightarrow CH_3COOH$ — the added H^+ is *swallowed* into weak acid.
- Add base (OH^-): $CH_3COOH + OH^- \rightarrow CH_3COO^- + H_2O$ — the added OH^- is *neutralised* by the weak acid.
- Net result: [H^+] barely changes → pH nearly constant.

5.4.4 The buffer equation (Henderson–Hasselbalch)

From the acid dissociation constant: $K_a = [H^+][A^-]/[HA] \rightarrow [H^+] = K_a \times [HA]/[A^-]$ → take −log of both sides: **pH = pKa + log([salt]/[acid])** — where [salt] = conjugate base, [acid] = weak acid.

For a basic buffer: **pOH = pKb + log([salt]/[base])**, then pH = 14 − pOH.

Worked examples:

1. *pH of a buffer with 0.1 M CH_3COOH + 0.1 M CH_3COONa (pKa = 4.76)?* pH = 4.76 + log(0.1/0.1) = 4.76 + 0 = **4.76**.
2. *0.1 M acid + 0.2 M salt?* pH = 4.76 + log(2) = 4.76 + 0.30 = **5.06**.
3. *Blood bicarbonate buffer:* pH = 6.1 + log([HCO_3^-]/[CO_2]); the body holds the ratio at 20:1 → pH = 6.1 + 1.30 = **7.40**. (Full detail in Chapter 7.)

5.4.5 Buffer capacity

The amount of strong acid/base (moles per litre) needed to change the buffer pH by 1 unit. It is **maximum when [salt] = [acid]** (pH = pKa) and increases with total buffer concentration. Effective buffer range: **pKa ± 1**.

5.4.6 Pharmaceutical buffers

Buffer system	pH range	Use
Acetate	3.7–5.7	Oral liquids
Phosphate	5.8–8.0	Eye drops, injections
Citrate	3.0–6.2	Syrups, effervescent preparations
Borate	8.0–10.0	Ophthalmic (buffered)

5.5 Easy Analogy

A buffer is the **shock absorber of a car**: the road (added acid/base) keeps throwing bumps, but the passengers (pH) barely feel them. The pH scale is like a **temperature scale for acidity** — and just as your body holds 37 °C, your blood holds pH 7.4.

5.6 Diagram — Buffer Action

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UNBUFFERED (pure water):          BUFFERED (CH3COOH/CH3COO-):
Add 1 drop HCl                    Add 1 drop HCl
pH 7 → pH 3  (plunges!)           pH 4.76 → pH 4.74  (barely moves)
Add 1 drop NaOH                   Add 1 drop NaOH
pH 3 → pH 11 (shoots up!)         pH 4.74 → pH 4.78  (barely moves)
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AI Image Prompt: "Educational textbook diagram, white background, two side-by-side beaker illustrations: left beaker labelled 'Unbuffered water' with a steep pH plunge curve, right beaker labelled 'Buffered solution' with a nearly flat pH curve, drops of acid and base shown being added, clear labels and arrows, pharmacy textbook style, minimal decoration."

5.7 Flowchart — Choosing a Buffer

> What pH is needed?
> ↓
> Pick a weak acid with pKa within ±1 of target pH
> ↓
> Set [salt]/[acid] ratio via Henderson–Hasselbalch
> ↓
> Use enough total concentration for required buffer capacity
> ↓
> Check compatibility (no precipitation with drug/excipients)

5.8 Examples

- **Blood** — bicarbonate buffer holds pH 7.4.
- **Tears** — phosphate/bicarbonate buffers (~pH 7.4); eye drops are buffered to match.
- **Aspirin** degrades faster in alkaline pH — formulated acidic.

5.9 Pharmacy Relevance

Formulation pH controls **stability** (hydrolysis of drugs like aspirin, penicillins), **solubility** (weak acids dissolve better as salts), **absorption** (pH–partition theory: unionised drug crosses membranes), and **comfort** (IV and ophthalmic products near physiological pH).

5.10 Important Points to Remember

1. **Remember:** Acid = H^+ donor; Base = H^+ acceptor (Brønsted–Lowry); Buffer = resists pH change.
2. **Exam Pearl:** pH = −log[H^+]; each unit = 10× change in [H^+].
3. **Exam Pearl:** Buffer = weak acid + conjugate base (or weak base + conjugate acid).
4. **Exam Pearl:** Henderson–Hasselbalch: pH = pKa + log([salt]/[acid]).
5. Buffer capacity is maximum at pH = pKa; effective range pKa ± 1.
6. Blood pH 7.4 is held by the bicarbonate buffer at a 20:1 ratio.
7. Strong acid ≠ concentrated acid: "strong" = fully dissociated; "concentrated" = lots of it.

5.11 Common Confusions & Common Mistakes

Do not confuse: pH ≠ pKa (pH = solution's acidity; pKa = acid's intrinsic strength) · **Strong ≠ concentrated** · **Buffer ≠ neutraliser** (a buffer resists pH change; an antacid neutralises acid outright). **Mistakes:** writing log([acid]/[salt]) upside-down; using pKa of the wrong system; forgetting pH + pOH = 14.

5.12 Memory Tricks

- "**pH = pKa + log(salt/acid)** — **S**alt on **t**op, **S**uperman flies high." (salt/acid, not acid/salt)
- "**B**rønsted **B**ase = **B**egs for protons (accepts)."

5.13 Exam-Oriented Summary

Acids donate H^+, bases accept H^+ (Brønsted–Lowry). pH = −log[H^+], 0–14. Buffers (weak acid + conjugate base) resist pH change by swallowing added H^+/OH^-. Henderson–Hasselbalch: pH = pKa + log([salt]/[acid]). Capacity max at pH = pKa. Pharmaceutical buffers: acetate, phosphate, citrate.

One-Minute Revision: Acid = H^+ donor; Base = H^+ acceptor. pH = −log[H^+]. Buffer = weak acid + conjugate base. pH = pKa + log(salt/acid). Capacity max at pKa; range pKa ± 1. Blood 7.4 via bicarbonate 20:1.

5.14 Questions

Very Short Answer: 1. Define buffer (scientific + simple). 2. State the Henderson–Hasselbalch equation. 3. What is buffer capacity? 4. Give one acidic and one basic buffer. 5. Define pH.

Short Answer: 1. Compare Arrhenius, Brønsted–Lowry, and Lewis theories with examples. 2. Explain buffer action with equations (acetic acid/acetate). 3. Calculate the pH of 0.001 M HCl. *(Ans: 3)* 4. A buffer contains 0.2 M sodium acetate and 0.1 M acetic acid (pKa 4.76). Calculate pH. *(Ans: 5.06)*

Long Answer: 1. Define acids, bases, and buffers. Derive the Henderson–Hasselbalch equation and explain buffer action and buffer capacity. 2. Discuss the significance of pH in pharmacy with examples.

MCQs

1. pH = pKa + log([salt]/[acid]) is the: (a) Nernst equation (b) Henderson–Hasselbalch equation (c) Beer–Lambert law (d) Henderson equation for solubility — **Ans: (b)**.
2. A solution of pH 3 is ___ times more acidic than pH 5: (a) 2 (b) 20 (c) 100 (d) 1000 — **Ans: (c)**. Two units = 10^2.
3. Which pair forms a buffer? (a) HCl + NaCl (b) CH_3COOH + CH_3COONa (c) NaOH + NaCl (d) H_2SO_4 + Na_2SO_4 — **Ans: (b)**. Weak acid + conjugate base; strong acid/base pairs cannot buffer.
4. Buffer capacity is maximum when: (a) [salt] = 2×[acid] (b) [salt] = [acid] (c) pH = 7 (d) concentration is minimum — **Ans: (b)**. pH = pKa at 1:1.
5. Blood pH is maintained at 7.4 mainly by: (a) phosphate buffer (b) bicarbonate buffer (c) protein buffer (d) acetate buffer — **Ans: (b)**.

Viva: Why can't a mixture of HCl and NaCl act as a buffer? What is the effective pH range of a buffer?
Application-based: An eye-drop formula needs pH 6.8. Which buffer system would you pick — acetate (pKa 4.76) or phosphate (pKa_2 7.2)? *(Ans: Phosphate — its pKa is within ±1 of 6.8; acetate is too far.)*

CHAPTER 6: ISOTONICITY

Unit II · ~1.5 hours

6.1 Introduction

Tears don't sting your eyes, and a saline drip doesn't destroy your blood cells — because both are **isotonic** with body fluids: they have the same "saltiness" as blood. Get this wrong and red blood cells burst or shrivel. Adjusting isotonicity is a core calculation skill for IV fluids and eye preparations.

Learning objectives — after studying this chapter, you should be able to:

- define isotonic, hypertonic, and hypotonic;
- explain osmosis and osmotic pressure in simple words;
- describe what happens to red blood cells in each type of solution;
- adjust isotonicity by the cryoscopic method and the sodium chloride equivalent method, with calculations;
- state the applications in IV fluids and ophthalmic solutions.

6.2 Definition

Isotonic — scientific: a solution having the same osmotic pressure as body fluids (0.9% w/v NaCl or blood plasma). **Isotonic — simple:** a solution with the same "saltiness" as blood — cells neither swell nor shrink in it.

Hypertonic: higher osmotic pressure than body fluids (cells shrink). **Hypotonic:** lower osmotic pressure (cells swell/burst).

6.3 Key Terminology

Term	Meaning
Osmosis	Movement of solvent (water) across a semipermeable membrane from dilute to concentrated side
Osmotic pressure	The pressure needed to *stop* osmosis — a measure of a solution's "pull" on water
Tonicity	A solution's effect on cell volume (iso/hyper/hypotonic) — a *biological* response
Osmolarity	Total moles of dissolved particles per litre (mOsm/L) — a *physical* property
Haemolysis	Bursting of red blood cells in hypotonic solution
Crenation	Shrinking/wrinkling of red blood cells in hypertonic solution

6.4 Detailed Explanation

6.4.1 Why isotonicity matters

- **IV fluids:** injected directly into blood. A hypotonic IV fluid makes RBCs swell and **burst (haemolysis)**; a hypertonic one makes them **shrink (crenation)**. Both can be dangerous.
- **Ophthalmic solutions:** the eye tolerates a range, but non-isotonic drops cause **stinging, tearing, and reflex blinking**, washing the drug away.

6.4.2 The cell in three solutions

Solution	vs blood	Water movement	RBC effect
Hypotonic	Lower osmotic pressure	Into the cell	Swells → **haemolysis** (bursts)
Isotonic	Equal	No net movement	**Unchanged**
Hypertonic	Higher osmotic pressure	Out of the cell	Shrinks → **crenation**

> **Exam Pearl:** 0.9% w/v NaCl (normal saline) and 5% w/v dextrose are isotonic with blood.

6.4.3 Methods of adjusting isotonicity

Method 1 — Cryoscopic (freezing-point) method. Freezing point depression is a colligative property: ΔTf = Kf × m. Blood plasma freezes at **−0.52 °C**. Any solution freezing at −0.52 °C is isotonic with blood. Procedure: calculate the freezing-point depression of the drug solution, then add enough NaCl (or dextrose) to bring the total depression to 0.52 °C.

Method 2 — Sodium chloride equivalent (E value) method. The **NaCl equivalent (E)** of a drug = the weight of NaCl that has the *same osmotic effect* as 1 g of the drug (tabulated in references). Formula (per 100 ml): **NaCl to add (g) = 0.9 − (E × % w/v of drug)**.

> **Worked example:** *Make 100 ml of 1% w/v atropine sulphate eye drops isotonic (E = 0.13).* NaCl needed = 0.9 − (0.13 × 1) = 0.9 − 0.13 = **0.77 g**. So: dissolve 1 g atropine sulphate + 0.77 g NaCl, make up to 100 ml → isotonic.

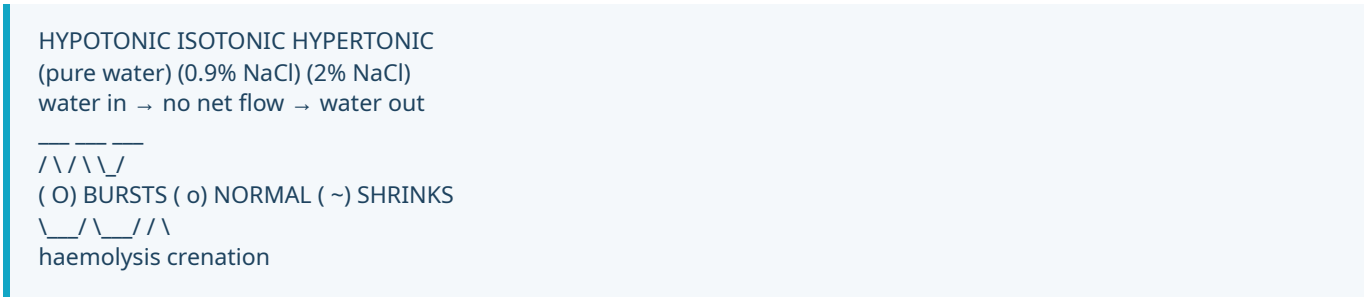
> **Worked example 2:** *2% w/v drug solution, E = 0.2, prepare 50 ml.* Per 100 ml: NaCl = 0.9 − (0.2 × 2) = 0.5 g → for 50 ml: **0.25 g NaCl**.

(The White–Vincent method is a third, volume-based method using the same E values — know its name.)

6.5 Easy Analogy

A red blood cell is a **water balloon with salty water inside**, floating in the bloodstream. Put it in pure water (hypotonic) and water rushes in — the balloon bursts. Put it in very salty water (hypertonic) and water is sucked out — the balloon wrinkles. Isotonic solution = the same saltiness inside and out — the balloon is happy.

6.6 Diagram — RBC in Three Solutions



> **AI Image Prompt:** "Educational textbook diagram, white background, three panels labelled Hypotonic, Isotonic, Hypertonic, each showing a red blood cell (swollen/bursting, normal disc, shrunken wrinkled) with blue arrows showing water movement in or out, clear black labels 'haemolysis' and 'crenation', pharmacy textbook style, minimal decoration."

6.7 Flowchart — Adjusting Isotonicity

> Drug solution prepared (hypotonic)
> ↓
> Find drug's NaCl equivalent (E) from tables
> ↓
> NaCl to add = 0.9 − (E × % w/v drug) [per 100 ml]
> ↓
> Dissolve drug + calculated NaCl → make up volume
> ↓
> Isotonic solution ready (verify: 0.9% NaCl equivalent)

6.8 Examples

- **Normal saline (0.9% NaCl)** and **5% dextrose** — isotonic IV fluids.
- **Eye drops** — adjusted with NaCl or boric acid to avoid stinging.
- **Nasal sprays** — isotonic to avoid nasal irritation.

6.9 Pharmacy Relevance

Every IV admixture a hospital pharmacist prepares must be checked for tonicity — hypotonic injections can cause haemolysis. In manufacturing, isotonicity adjustment is a routine calculation for all parenteral and ophthalmic products.

6.10 Important Points to Remember

1. **Remember:** Isotonic = same osmotic pressure as blood (0.9% NaCl).
2. **Exam Pearl:** Hypotonic → water into cell → **haemolysis**; Hypertonic → water out → **crenation**.
3. Blood plasma freezes at **−0.52 °C** (cryoscopic method).
4. **Exam Pearl:** NaCl to add (per 100 ml) = 0.9 − (E × % drug).
5. E = NaCl equivalent of 1 g of drug.
6. Applications: **IV fluids** (safety) and **ophthalmic solutions** (comfort).

6.11 Common Confusions & Common Mistakes

Do not confuse: Isotonic ≠ iso-osmotic (iso-osmotic = same osmolarity physically; isotonic = same *effect on cells* — e.g., urea solution is iso-osmotic but hypotonic because urea enters cells). **Osmolarity ≠ tonicity** (physical property vs biological response). **Mistakes:** forgetting that E × %drug must be *subtracted* from 0.9; using the formula per 100 ml but preparing a different volume without scaling.

6.12 Memory Tricks

- "**H**ypotonic → **H**$_2$O in → **H**aemolysis (bursts). **H**ypertonic → water **H**auled out → crenation." (Both start with H — remember *in* vs *out*.)
- "**E** = **E**qual NaCl: 1 g drug = E g NaCl in osmotic effect."

6.13 Exam-Oriented Summary

Isotonic solutions match blood's osmotic pressure (0.9% NaCl). Hypotonic bursts RBCs (haemolysis); hypertonic shrinks them (crenation). Adjust by cryoscopic method (freezing point −0.52 °C) or NaCl equivalent method: NaCl = 0.9 − (E × %). Applications: IV fluids and eye preparations.

One-Minute Revision: Isotonic = like blood (0.9% NaCl). Hypo → in → haemolysis; Hyper → out → crenation. Cryoscopic: −0.52 °C. E method: 0.9 − (E × %). Uses: IV + ophthalmic.

6.14 Questions

Very Short Answer: 1. Define isotonic solution. 2. What is haemolysis? 3. State the freezing point of blood plasma. 4. What is NaCl equivalent (E)? 5. Name two isotonic IV fluids.

Short Answer: 1. Describe the effect of hypo-, iso-, and hypertonic solutions on RBCs. 2. Explain the cryoscopic method of isotonicity adjustment. 3. Calculate NaCl needed for 100 ml of 2% w/v boric acid solution (E = 0.5). *(Ans: 0.9 − 1.0 < 0 → already hypertonic; no NaCl needed — dilute instead.)* 4. Why must IV fluids be isotonic?

Long Answer: 1. Define isotonicity. Describe the methods of adjusting isotonicity with a worked example, and explain its applications in IV fluids and ophthalmic solutions.

MCQs

1. Haemolysis occurs when RBCs are placed in: (a) isotonic (b) hypertonic (c) hypotonic (d) saturated solution — **Ans: (c)**. Water rushes in and bursts the cell.
2. Blood plasma freezes at: (a) 0 °C (b) −0.52 °C (c) −1.86 °C (d) −5.2 °C — **Ans: (b)**.
3. For 100 ml of 1% drug solution (E = 0.2), NaCl needed is: (a) 0.9 g (b) 0.7 g (c) 0.2 g (d) 1.1 g — **Ans: (b)**. 0.9 − 0.2 = 0.7 g.
4. Which is isotonic with blood? (a) 0.9% NaCl (b) 5% dextrose (c) both (d) neither — **Ans: (c)**.
5. Crenation is: (a) bursting of RBCs (b) shrinking of RBCs (c) clotting (d) settling — **Ans: (b)**.

Viva: Why do non-isotonic eye drops sting? What is the White–Vincent method? **Application-based:** A student prepares 30 ml of 1% w/v pilocarpine eye drops (E = 0.24). How much NaCl is needed? *(Ans: per 100 ml: 0.9 − 0.24 = 0.66 g; for 30 ml: 0.66 × 0.3 = 0.198 g ≈ 0.2 g.)*

CHAPTER 7: ELECTROLYTES AND PHYSIOLOGICAL ACID–BASE BALANCE

Unit II · ~3.5 hours

7.1 Introduction

Your nerves fire, your heart beats, and your muscles contract — all because of tiny charged particles called **electrolytes** (Na^+, K^+, Ca^{2+}…). Lose too much through diarrhoea or vomiting and the body's chemistry collapses; that is why ORS saves millions of lives. This chapter connects inorganic chemistry directly to the living patient.

Learning objectives — after studying this chapter, you should be able to:

- define electrolyte and distinguish intracellular from extracellular fluids;
- list the major physiological ions with their locations and normal levels;
- explain the functions of Na^+, K^+, Ca^{2+}, Mg^{2+}, Cl^-, HCO_3^-, and phosphate;
- describe the electrolytes used in replacement therapy (NaCl, KCl, $CaCl_2$, ORS);
- give the WHO reduced-osmolarity ORS formulation and explain how each ingredient works;
- explain how the body maintains blood pH (buffers, lungs, kidneys);
- distinguish respiratory from metabolic acidosis and alkalosis.

7.2 Definition

Electrolyte — scientific: a substance that dissociates into ions in solution and conducts electricity; in physiology, the major body ions (Na^+, K^+, Ca^{2+}, Mg^{2+}, Cl^-, HCO_3^-, phosphates). **Electrolyte — simple:** the body's "salts" dissolved in body water — they carry electrical charges that run nerves and muscles and control water balance.

7.3 Key Terminology

Term	Meaning
Intracellular fluid (ICF)	Fluid *inside* cells (~40% of body weight); rich in K^+, Mg^{2+}, phosphate
Extracellular fluid (ECF)	Fluid *outside* cells (~20% of body weight): plasma + interstitial fluid; rich in Na^+, Cl^-, HCO_3^-
mEq/L	Milliequivalents per litre — the unit for electrolyte concentration in blood
Dehydration	Loss of body water (and usually electrolytes)
Acidosis / Alkalosis	Blood pH below 7.35 / above 7.45

7.4 Detailed Explanation

7.4.1 Body fluid compartments

Total body water ≈ **60% of body weight** → ICF ~40% (inside cells) + ECF ~20% (outside: plasma ~4–5%, interstitial fluid ~15%).

> **Exam Pearl: K^+ is the king inside** (ICF), **Na^+ is the king outside** (ECF). The cell membrane's Na^+/K^+ pump maintains this divide.

7.4.2 Major physiological ions — functions

Ion	Main location	Normal serum level	Key functions
Na^+	ECF	135–145 mEq/L	Water balance, osmotic pressure, nerve impulses
K^+	ICF	3.5–5.0 mEq/L	Muscle contraction, **cardiac rhythm**, nerve function
Ca^{2+}	ECF, bone	8.5–10.5 mg/dL	Bones/teeth, blood clotting, muscle contraction, nerve transmission
Mg^{2+}	ICF	1.5–2.5 mEq/L	Cofactor for 300+ enzymes, muscle/nerve function
Cl^-	ECF	95–105 mEq/L	Acid–base balance, gastric HCl, osmotic pressure
HCO_3^-	ECF	22–26 mEq/L	**Main blood buffer**
Phosphate	ICF, bone	—	Bones, ATP (energy), intracellular buffer

Deficiency/excess glimpses: low Na^+ (hyponatraemia) → confusion; low K^+ (hypokalaemia) → muscle weakness, cardiac arrhythmia; low Ca^{2+} → tetany (muscle spasms).

7.4.3 Electrolytes in replacement therapy

Sodium chloride* — 0.9% w/v (normal saline): isotonic fluid/electrolyte replacement, the standard IV fluid. (3% hypertonic saline is used in hospitals for severe hyponatraemia.)

Potassium chloride — treats **hypokalaemia** (from diuretics, vomiting, diarrhoea): oral syrup/tablets or IV infusion. **Safety pearl:** IV KCl must *never* be given as a bolus — rapid K^+ rise stops the heart; it is always diluted and infused slowly.

Calcium chloride — treats **hypocalcaemia** and is used in cardiac resuscitation and hyperkalaemia: IV only; extravasation causes tissue necrosis.

7.4.4 Oral Rehydration Salt (ORS)

Why ORS works: the small intestine has a **glucose–sodium co-transporter** — glucose *drags* sodium (and water follows) into the blood on a 1:1 basis, even during diarrhoea. That single mechanism is why sugar + salt + water saves lives.

WHO reduced-osmolarity ORS (2003; guidelines updated 2006) — recommended by WHO and UNICEF, replacing the old 311 mOsm/L formula:

Ingredient	g/L	Ions/molecules	mmol/L
Sodium chloride	2.6	Sodium	75
Potassium chloride	1.5	Potassium	20
Trisodium citrate dihydrate	2.9	Citrate	10
Glucose, anhydrous	13.5	Glucose	75
(Chloride)	—	Chloride	65
Total osmolarity			**245 mOsm/L**

Role of each ingredient: NaCl/KCl — replace lost electrolytes; glucose — drives Na^+/water absorption (1:1 co-transport); citrate — corrects the **acidosis** of diarrhoea (replacing the older bicarbonate, which was unstable in hot climates).

Advantages of the reduced-osmolarity formula: ~25% less stool volume, ~30% less need for IV therapy, less vomiting, no hypernatraemia risk — effective for diarrhoea of all causes, all ages.

Preparation & counselling (pharmacist's job): dissolve the *entire* sachet in **1 litre** of clean drinking water (too little water worsens diarrhoea); use within **24 hours**; continue feeding/breastfeeding. **WHO also recommends zinc** (20 mg/day above 6 months, 10 mg/day below 6 months, for 10–14 days) alongside ORS for childhood diarrhoea.

7.4.5 Physiological acid–base balance

Normal arterial blood pH: **7.35–7.45** (7.40). Three lines of defence:

1. Buffer systems (seconds): bicarbonate (H_2CO_3/HCO_3^-, pKa 6.1 — the main ECF buffer): pH = 6.1 + log([HCO_3^-]/[CO_2]); the body holds the ratio at **20:1** → pH 7.4. **Phosphate** buffer (ICF, urine); **protein/haemoglobin** buffers (blood).

2. Lungs — respiratory regulation (minutes): blow off CO_2 by hyperventilation → pH rises; hypoventilation retains CO_2 → pH falls. $CO_2 + H_2O \rightleftharpoons H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$.

3. Kidneys — renal regulation (hours to days): excrete H^+, reabsorb and generate fresh HCO_3^- — slow but the most powerful long-term control.

Disorders:

Disorder	pH	Primary cause	Example
Respiratory acidosis	↓	CO_2 retention	COPD, airway obstruction
Respiratory alkalosis	↑	CO_2 blown off	Hyperventilation, anxiety, high altitude
Metabolic acidosis	↓	HCO_3^- loss / acid gain	Diarrhoea, diabetic ketoacidosis, renal failure
Metabolic alkalosis	↑	HCO_3^- gain / acid loss	Prolonged vomiting, excess antacids

7.5 Easy Analogy

Electrolytes are the body's **electrical wiring and plumbing crew**: sodium manages the water levels, potassium keeps the heartbeat's rhythm, calcium is the muscle's "on" switch. Blood pH is guarded like a **thermostat at 7.4** — buffers react in seconds (like opening a window), lungs in minutes (like a fan), kidneys in days (like rebuilding the insulation).

7.6 Diagram — Body Fluid Compartments & pH Defence

```
TOTAL BODY WATER (~60% body weight)
|
+-----+-----+
| |
ICF (~40%) ECF (~20%)
K+, Mg2+, Na+, Cl-, HCO3-
phosphate (plasma + interstitial)

pH DEFENCE (blood 7.35–7.45):
Buffers (seconds) → Lungs (minutes) → Kidneys (hours–days)
```

AI Image Prompt: "Educational textbook diagram, white background, two-part illustration: (1) human silhouette with pie-style body fluid compartments labelled ICF 40% (K+, Mg2+, phosphate) and ECF 20% (Na+, Cl-, HCO3-); (2) three shields labelled Buffers/seconds, Lungs/minutes, Kidneys/days guarding a pH 7.4 gauge; clean black labels, pharmacy textbook style, minimal decoration."

7.7 Flowchart — Classifying Acid–Base Disorders

```
Blood pH abnormal?
├─ pH < 7.35 → ACIDOSIS
│  ├─ CO2 high → Respiratory acidosis
│  └─ HCO3- low → Metabolic acidosis
└─ pH > 7.45 → ALKALOSIS
   ├─ CO2 low → Respiratory alkalosis
   └─ HCO3- high → Metabolic alkalosis
```

7.8 Examples

- **Diarrhoea in a child** → ORS + zinc (the pharmacist's commonest counselling).
- **Vomiting patient on diuretics** → KCl supplementation (hypokalaemia risk).
- **Diabetic ketoacidosis** → metabolic acidosis (hospital emergency).

7.9 Pharmacy Relevance

Counselling a mother on ORS (*exact* 1 litre, 24-hour use, continue feeding) is core community-pharmacy work. Hospital pharmacists monitor IV fluid choice (normal saline vs dextrose) and flag IV KCl bolus errors — a classic fatal medication error.

7.10 Current Update — 2026

The **WHO reduced-osmolarity ORS (2003/2006)** remains the global standard (245 mOsm/L; Na^+ 75, glucose 75 mmol/L) — it is on the WHO Model List of Essential Medicines. WHO pairs it with **zinc supplementation** (10–14 days) for childhood diarrhoea, which shortens duration and reduces recurrence.

7.11 India-Specific Context

ORS is supplied through India's public health system (sub-centres, ASHA workers) for diarrhoea management — one of the most cost-effective interventions in the National Health Mission. IP monographs cover Sodium Chloride Infusion, Potassium Chloride Oral Solution, and ORS powder.

7.12 Important Points to Remember

1. **Remember:** Electrolytes = body ions running nerves, muscles, water balance.
2. **Exam Pearl: K^+ inside (ICF), Na^+ outside (ECF)** — 60% body water = 40% ICF + 20% ECF.
3. HCO_3^- (22–26 mEq/L) is the main blood buffer; blood pH 7.35–7.45.
4. **Exam Pearl:** ORS works by **glucose–sodium co-transport (1:1)** in the small intestine.
5. **Exam Pearl:** Reduced-osmolarity ORS: NaCl 2.6, KCl 1.5, trisodium citrate 2.9, glucose 13.5 g/L → 245 mOsm/L (Na^+ 75, K^+ 20, Cl^- 65, citrate 10, glucose 75 mmol/L).
6. **Exam Pearl:** IV KCl never as bolus — cardiac arrest risk; always diluted, slow infusion.
7. pH defence: buffers (seconds) → lungs (minutes) → kidneys (days).
8. Blood buffer equation: $pH = 6.1 + \log([HCO_3^-]/[CO_2])$, ratio 20:1 → 7.4.

7.13 Common Confusions & Common Mistakes

Do not confuse: ICF ≠ ECF ions (K^+ inside, Na^+ outside) · **Acidosis ≠ alkalosis** (pH < 7.35 vs > 7.45) · **Respiratory ≠ metabolic** (CO_2 problem vs HCO_3^- problem) · Old ORS (311 mOsm/L, Na 90) ≠ new ORS (245, Na 75). **Mistakes:** dissolving ORS in less than 1 litre ("stronger is better" — wrong, worsens diarrhoea); keeping ORS solution beyond 24 hours.

7.14 Memory Tricks

- "**K**$^+$ stays **K**aged **in**side (ICF); **Na**$^+$ stays **Na**turally **out**side (ECF)."
- "**B**icarbonate **B**uffers **B**lood (the 3 B's); ratio **20:1** → **7.4**."
- "ORS: **2.6 – 1.5 – 2.9 – 13.5** g/L (NaCl–KCl–citrate–glucose)."

7.15 Exam-Oriented Summary

Body water 60% (ICF 40%, ECF 20%); K^+ inside, Na^+ outside. Major ions and functions; replacement therapy: NaCl (normal saline), KCl (hypokalaemia — never IV bolus), $CaCl_2$ (hypocalcaemia), ORS (glucose–Na co-transport; 2003 formula 245 mOsm/L). Blood pH 7.35–7.45 defended by buffers, lungs, kidneys; four acid–base disorders.

One-Minute Revision: K^+ in, Na^+ out. HCO_3^- main buffer. ORS = glucose drags Na^+ + water (1:1); new formula 245 mOsm/L (2.6/1.5/2.9/13.5 g/L). 1 litre, 24 hours. pH 7.4 = buffers → lungs → kidneys. Acidosis pH<7.35; alkalosis >7.45.

7.16 Questions

Very Short Answer: 1. Define electrolyte. 2. Which ion dominates ICF? ECF? 3. Normal serum Na^+/K^+ levels. 4. Why does ORS contain glucose? 5. Normal blood pH range. 6. Name the four acid–base disorders.

Short Answer: 1. Tabulate major physiological ions with location and one function each. 2. Give the WHO reduced-osmolarity ORS formulation and the role of each ingredient. 3. Explain the bicarbonate buffer system of blood with the Henderson–Hasselbalch equation. 4. Why must IV potassium chloride never be given as a bolus? 5. Distinguish respiratory from metabolic acidosis with examples.

Long Answer: 1. Describe the electrolytes used in replacement therapy, with the ORS formulation, mechanism, and preparation. 2. Explain the physiological acid–base balance: buffer systems, respiratory and renal regulation, and the four disorders.

MCQs

1. The dominant intracellular ion is: (a) Na^+ (b) K^+ (c) Cl^- (d) HCO_3^- — **Ans: (b)**.
2. Reduced-osmolarity ORS has total osmolarity: (a) 311 (b) 245 (c) 200 (d) 150 mOsm/L — **Ans: (b)**.
3. In 1 litre of WHO ORS, anhydrous glucose is: (a) 20 g (b) 13.5 g (c) 2.6 g (d) 1.5 g — **Ans: (b)**. (20 g was the old formula.)
4. Blood pH is maintained at 7.4 with a $[HCO_3^-]/[CO_2]$ ratio of: (a) 1:1 (b) 10:1 (c) 20:1 (d) 40:1 — **Ans: (c)**. 6.1 + log 20 = 7.4.
5. Prolonged vomiting typically causes: (a) metabolic acidosis (b) metabolic alkalosis (c) respiratory acidosis (d) respiratory alkalosis — **Ans: (b)**. Loss of gastric HCl → HCO_3^- excess.
6. IV potassium chloride must not be given as a bolus because: (a) it precipitates (b) rapid K^+ rise can stop the heart (c) it is hypotonic (d) it inactivates the drug — **Ans: (b)**.
7. Citrate in ORS: (a) adds flavour (b) corrects acidosis (c) preserves the powder (d) increases osmolarity — **Ans: (b)**.

Viva: Why was bicarbonate replaced by citrate in ORS? What is the Na^+/K^+ pump's role in fluid compartments? **Application-based:** A mother dissolves one ORS sachet in half a litre "to make it stronger" for her dehydrated child. What is wrong, and what do you counsel? *(Ans: Concentrated ORS worsens diarrhoea (hyperosmolar) and risks hypernatraemia — always the full 1 litre, discard after 24 hours, continue feeding, add zinc per WHO.)*

UNIT II — INFOGRAPHIC CONCEPT: "pH, BUFFERS & BODY SALTS"

Title: The Body Runs on pH 7.4 and Salt Water **Sections:** pH scale strip (0–14 with examples) → buffer shock-absorber visual → Henderson–Hasselbalch mini formula → ICF/ECF ion map → ORS recipe card (2.6/1.5/2.9/13.5 g/L) → pH defence trio (buffers/lungs/kidneys). **Layout:** vertical flow, colour-coded (acid = red, alkali = blue, ions = teal).

AI Infographic Prompt: "Professional educational infographic, white background, title 'The Body Runs on pH 7.4 and Salt Water', sections: pH scale strip 0-14 with everyday examples, buffer shock-absorber illustration, Henderson-Hasselbalch formula card, body fluid compartment ion map, ORS recipe card with gram amounts, three shields for pH defence (buffers/lungs/kidneys); clean sans-serif typography, pharmacy textbook style, minimal decoration."

UNIT II — REFERENCES

1. Beckett, A.H. and Stenlake, J.B., *Practical Pharmaceutical Chemistry*, Part I, The Athlone Press — buffers, isotonicity, electrolytes.
2. Bentley, R. and Driver, J., *Textbook of Pharmaceutical Chemistry*, Oxford University Press — acid–base chemistry, electrolytes.
3. Indian Pharmacopoeia Commission, *Indian Pharmacopoeia*, Ghaziabad (current edition) — ORS, electrolyte infusions, buffer solutions.
4. WHO/UNICEF, *Oral Rehydration Salts: Production of the new ORS* (2006); reduced-osmolarity ORS recommendation (2003).
5. Vogel, A.I., *Vogel's Textbook of Quantitative Chemical Analysis*, Pearson — pH, buffer calculations.