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Program of Biophysics

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Biophysics Course for First-Year medicine
Chapter 1:
Properties of Solutions and Medical Applications

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Properties of Solutions and Their Medical Applications

I. Introduction

The biophysics of solutions is a branch of physics that studies the physical properties and behavior of solutions, particularly those containing biomolecules. It involves examining molecular interactions in solution, thermodynamic properties, molecular dynamics, and how these factors influence biological functions. Such studies are essential for understanding many biological processes and for developing new technologies in biotechnology, medicine, and pharmacology.

I.1 General Overview of the States of Matter

The states of matter are the different physical forms that matter can take. The main states are:

- **Solid:** The particles are tightly bound and vibrate around fixed positions. Solids have a definite shape and volume.
  Examples: ice, wood, metal.
- **Liquid:** The particles are more spread out than in a solid and can move relative to each other. Liquids have a definite volume but no fixed shape, taking the shape of their container.
  Examples: water, oil, alcohol.
- **Gas:** The particles are far apart and move freely. Gases have neither a definite shape nor volume and completely fill their container.

The components are the same in all three states; the difference between them depends on the nature of the **intermolecular interactions**.

I.1.1 Intermolecular Forces

Intermolecular forces are the attractive or repulsive forces that exist between molecules. They affect the physical properties of substances, such as melting point, boiling point, viscosity, and solubility. The main types of intermolecular forces are:

I.1.1.1 Van der Waals Forces

- **Dispersion Forces (London Forces):** Weak, temporary attractive forces that occur between all molecules, resulting from instantaneous dipoles created by electron movement. They are stronger in larger molecules with more electrons.
- **Dipole–Dipole Forces:** Attractive forces between polar molecules, which have permanent dipoles due to unequal electron distribution. These forces are stronger than dispersion forces.
- **Induced Dipole–Dipole Forces:** Occur when a polar molecule induces a dipole in a neighboring nonpolar molecule, leading to attraction between them.

I.1.1.2 Hydrogen Bonds

Hydrogen bonds are particularly strong intermolecular forces that occur when hydrogen is bonded to a highly electronegative atom such as oxygen, nitrogen, or fluorine. They play a crucial role in determining the properties of water and biomolecules such as proteins and nucleic acids.

I.1.1.3 Ion–Dipole Forces

These are attractive forces between an ion (cation or anion) and a polar molecule. They are important in ionic solutions, such as when salt dissolves in water.

In general, substances with strong intermolecular forces have higher melting and boiling points, while those with weaker forces are more volatile and have lower melting and boiling points.

I.1.2 Phase Transitions

Phase transitions are changes in a substance from one state of matter to another. These transformations occur when energy—usually in the form of heat—is added or removed from the substance. The main phase transitions include:

I.1. 2.1 Fusion (Melting):

The change from solid to liquid. When a solid is heated, its particles gain energy, vibrate faster, and eventually move freely to form a liquid.
Example: ice melting into water.

I.1.2.2 Solidification (Freezing):

The change from liquid to solid. When a liquid cools, its particles lose energy, move more slowly, and arrange themselves into a solid structure.
Example: water freezing into ice.

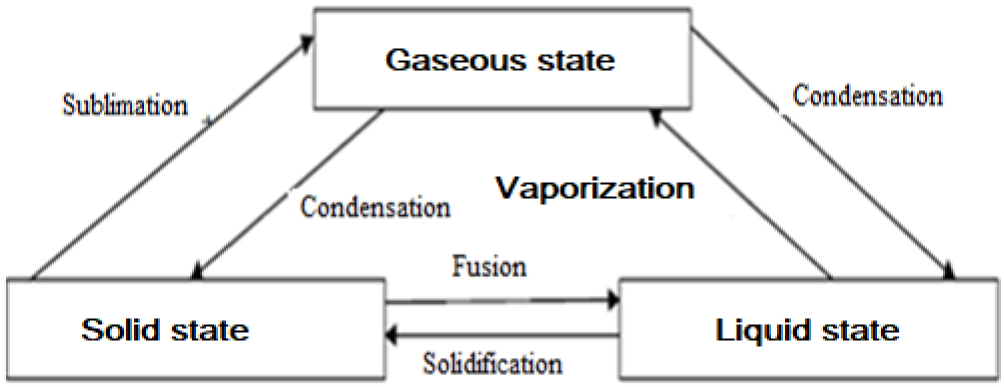
I.1.2.3 Vaporization:

The change from liquid to gas, which includes two processes:

- **Boiling:** Occurs when the liquid reaches its boiling point and vapor bubbles form throughout the liquid.
- **Evaporation:** Occurs at the surface of a liquid at temperatures below its boiling point.

- **Condensation:**
  The change from gas to liquid. When a gas cools, its particles lose energy and move closer together to form a liquid.
  Example: water vapor condensing into droplets.
- **Sublimation:**
  The direct transition from solid to gas without passing through the liquid state.
  Example: dry ice (solid carbon dioxide) subliming into gas.
- **Deposition (Solid Condensation):**
  The direct transition from gas to solid without passing through the liquid state.
  Example: frost forming on windows from water vapor.
- **Ionization:**
  The change from gas to plasma, where atoms become ionized. This occurs at very high temperatures or under high-energy conditions.
  Example: neon gas becoming plasma in a neon light tube.
- **Deionization:**
  The reverse process, where plasma returns to the gaseous state as ions recombine to form neutral atoms.

These phase transitions are often illustrated in **phase diagrams**, which show the temperature and pressure conditions under which each state of matter is stable. *Figure 1 below shows the phase transitions of water.*



I.1.3 Phase Diagrams of Solutions

Phase diagrams of solutions show how the different phases (solid, liquid, gas) of a solution change depending on temperature, pressure, and composition. They are useful for understanding how the components of a solution mix or separate.

Some common types of phase diagrams for solutions include:

I.1.3.1 Binary Phase Diagram at Constant Pressure:

- **Solubility Curve:** Indicates the maximum concentration of solute that can dissolve in the solvent at different temperatures. Above this curve, the solute begins to precipitate.
- **Melting Point Diagram:** Shows how the melting point of a mixture varies with its composition.

I.1.3 .2 Liquid–Vapor Phase Diagram

Distillation Curve:
This curve shows how the composition of a liquid solution and its vapor changes with temperature at constant pressure (or with pressure at constant temperature). Distillation curves are used to design distillation columns and to understand the separation (fractionation) of components.

I.1.3.3 Ternary Phase Diagram

This type of diagram is used for solutions containing three components. Ternary diagrams are usually represented as equilateral triangles, where each corner represents a pure component and each side represents a binary mixture line.

I.1.3 .4 Pressure–Temperature (P–T) Diagram

This diagram shows the different phases of a solution at various combinations of temperature and pressure. P–T diagrams are useful for understanding phase transitions under extreme conditions, such as those encountered in industrial processes and natural systems.

An example of a phase diagram for aqueous solutions is shown in *Figure 2* below:

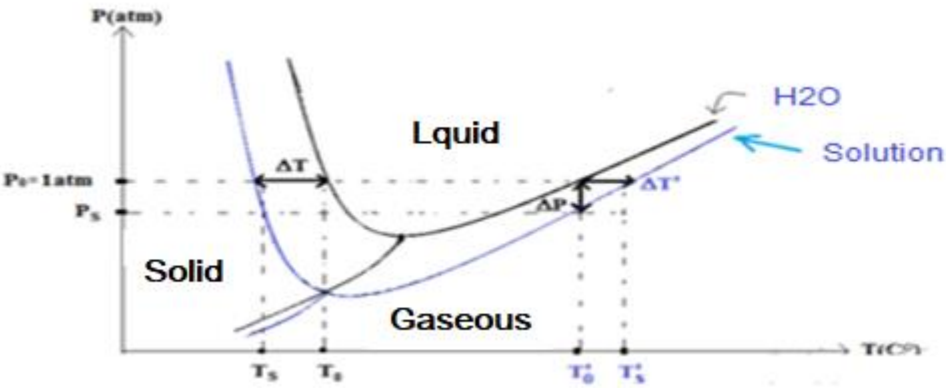


Figure 2: Phase Diagram of Aqueous Solutions

I.2 Fundamentals of Thermodynamics

Thermodynamics is the branch of physics that studies the exchange of heat and work between systems and their surroundings. It is based on four fundamental principles that help explain thermodynamic processes and their applications in various fields such as engineering, chemistry, physics, and medicine.

I.2.1 Fundamental Laws

1.2.1.1 Basic Concepts

- **System and Surroundings:**
  A *system* is the part of the universe being studied, while everything outside the system constitutes the *surroundings*, which can exchange energy or matter with the system. *Figure 3* illustrates these two concepts.

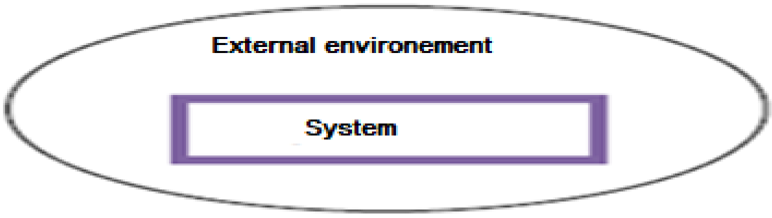


Figure 3: Concept of System and Surroundings

1.2.1.2 Types of Systems

- **Isolated System:** Exchanges neither energy nor matter with its surroundings.
- **Closed System:** Exchanges energy but not matter.
- **Open System:** Exchanges both energy and matter.

The different types are summarized in *Table 1.*

System	Mater Exchange	Energy exchange	Example
Isolated	none	none	Colorimeter
Close	none	yes	Electric battery
Open	yes	yes	To be alive

Table 1: Types of Systems — Isolated, Closed, and Open

1.2.1.3 Key Thermodynamic Quantities

- **Internal Energy (U):**
  The sum of all forms of energy present within a system (the total kinetic and potential energy of all the particles).
- **Heat (Q):**
  The energy transferred between the system and its surroundings due to a temperature difference.
  (*Thermo* ≡ heat)
- **Work (W):**
  The energy transferred by a process other than heat.
  (*Dynamics* ≡ motion ⇒ work)
- **Entropy (S):**
  A measure of disorder or the dispersion of energy within a system. In other words, it represents the irreversibility of processes.
  At **T = 0 K**, the particles that make up matter are motionless.
  T(K)=T(°C)+273T(K) = T(°C) + 273T(K)=T(°C)+273, where **0 K = -273 °C**.
- **Thermodynamic State:**
  The condition of a system described by state variables such as temperature, pressure, volume, and chemical composition.

1.2.1.4 Types of Processes

A process can be **isothermal**, **isobaric**, **isochoric**, or **adiabatic**, as summarized in *Table 2.*

Process Type	Description
Isothermal	Occurs at constant temperature.
Isobaric	Occurs at constant pressure.
Isochoric	Occurs at constant volume.
Adiabatic	Occurs with no heat exchange with the surroundings.

Table 2: Thermodynamic Processes

1.2.1.4 State and Process Functions

In addition to internal energy (U) and entropy (S), there are three other important thermodynamic functions: **enthalpy (H)**, **Helmholtz free energy (F)**, and **Gibbs free energy (G).**

- **Enthalpy (H):**

  The total energy of a system plus the product of its pressure and volume.

  H=U+PVH = U + PVH=U+PV

- **Helmholtz Free Energy (F):**

  A thermodynamic function combining internal energy and entropy, used to describe the useful work in systems at constant temperature.

  F=U−TSF = U - TSF=U−TS

- **Gibbs Free Energy (G):**

  A thermodynamic function combining enthalpy and entropy, used to predict the spontaneity of chemical reactions.

  G=H−TSG = H - TSG=H−TS

  If ΔG ≤0, the process is **spontaneous**.

1.2.2 The Four Laws of Thermodynamics

Thermodynamics is based on four fundamental laws or principles that describe the different types of exchanges between a system and its surroundings.

I.2.2.1 Zeroth Law of Thermodynamics

Principle:
If two systems are each in thermal equilibrium with a third system, then they are in thermal equilibrium with each other.

Implication:
This law defines the concept of **temperature** and allows for consistent temperature measurement.

I.2.2.2 First Law of Thermodynamics (Law of Conservation of Energy)

Principle:
Energy cannot be created or destroyed; it can only be transferred or transformed from one form to another.

Formula:

ΔU=Q−W\Delta U = Q - WΔU=Q−W

Where:

- ΔU: Change in the internal energy of the system
- Q: Heat added to the system
- W: Work done by the system

This law states that the change in a system's internal energy equals the heat added to the system minus the work performed by the system.

I.2.2.3 Second Law of Thermodynamics

Principle:
In any thermodynamic process, the total entropy of an isolated system can only increase or remain constant, it can never decrease.

Formula: $\Delta S \geq 0$

Here, ΔS: represents the change in entropy.
This law introduces the concept of entropy and states that spontaneous processes increase the total entropy of the universe, implying that some processes are irreversible.

I.2.2.4 Third Law of Thermodynamics

Principle:
As the temperature of a system approaches absolute zero, its entropy approaches a minimum constant value.

This law also states that it is impossible to reach absolute zero (0 Kelvin) in a finite number of steps. At this temperature, thermodynamic processes become extremely slow.

I.2.3 Brownian Motion

Brownian motion describes the random movement of a particle suspended in a fluid.
It results from the **collision forces** caused by **thermal agitation (T > 0)** of all fluid molecules acting on the particle.

Example:
In water at T = 20°C, the number of collisions experienced by one H_2O molecule exceeds **10^{14} collisions per second**.

The **kinetic theory** also allows calculating the **pressure (P, in Pa)** exerted by identical particles in terms of their concentration (C):

$$P = \frac{N}{V} < Ec > \frac{N}{V} = K_B\mathrm{TC}$$

Where:

- Ec_: Kinetic energy
- N: Number of molecules
- V: Volume (m^3) occupied by N molecules
- K_B: Boltzmann constant
- T: Absolute temperature

The direction of Brownian motion **cannot be predicted**, since it is random, but we can calculate the **mean square displacement**:

$$<x^2> = 2\frac{K_B .T}{f} . t$$ Where:

f: Molecular friction coefficient (solute/solvent)

T: Temperature

KB: Boltzmann constant

t: Time

x2: Square of the displacement

I.2.4 Chemical Potential (μ)

The **chemical potential (μ)** is the **partial molar free energy**, i.e., the contribution of each mole of solute to the system's Gibbs free energy.

$$\mu = \left(\frac{\partial G}{\partial n}\right)_{T,P,S}$$

In thermodynamics, the transfer of **energy and matter** is described by the **Gibbs equation**:

$$\Delta G = \mu \Delta n + V\Delta P - S\Delta T$$

Where:

- $\mu\Delta n$: Transfer of matter
- $V\Delta P$: Transfer of energy as mechanical work
- $S\Delta T$: Transfer of energy as heat

If **P and T are constant**, then:

ΔG=μΔn\Delta G = \mu \Delta nΔG=μΔn

I.3 Overview of Solutions

I.3.1 Definition

A **solution** is a **homogeneous mixture** containing one or more solutes dissolved in a solvent.

Solution=Solute + Solvent

- **Solvent:** The substance in which the solute is dissolved. It is usually present in larger quantity.
  Example: Water (H_2O) is the most common solvent.
- **Solute:** The substance that is dissolved. Usually present in smaller quantity.
  Example: In a saline solution, sodium chloride (NaCl) is the solute.

Notes:

- In an *aqueous solution*, water is the solvent.
- A key property of a solution is that the solute is **uniformly distributed**, forming a **homogeneous mixture**.

I.3.2 Classification of Aqueous Solutions and Concentrations

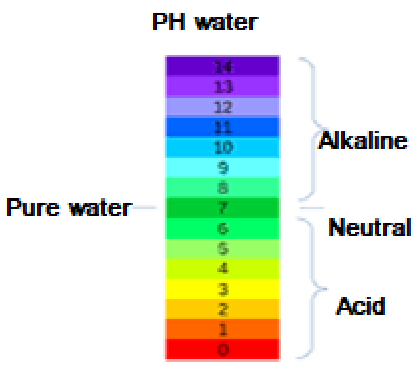
Aqueous solutions can be classified based on different criteria such as the **nature of the solute**, the **pH**, or the **concentration**.

I.3.2.1 Nature of the Solute

- **Electrolytic solutions:** Contain solutes that dissociate into ions in water, enabling the solution to conduct electricity.
  Examples: Sodium chloride ($NaCl$), hydrochloric acid (HCl)
- **Non-electrolytic solutions:** Do not dissociate into ions; thus, they do not conduct electricity.
  Examples: Glucose, sucrose

I.3.2.2 PH of the Solution

The **hydrogen potential (pH)** measures the **chemical activity of hydrogen ions (H^+)** in a solution.
In aqueous solutions, H^+ ions exist as **hydronium ions (H_3O^+)**.
pH indicates the **acidity or basicity** of a solution.We recall that the **hydrogen potential**, denoted **pH**, is a measure of the **chemical activity of protons or hydrogen ions** in a solution. In particular, in an **aqueous solution**, these ions are present in the form of **hydronium ions** (hydrated hydrogen ions, H^+ or H_3O^+).
The **pH** is used to measure the **acidity or basicity** of a solution (see Figure 4).



- **Acidic solutions:** pH < 7 (contain more H^+ than OH^-)
  Examples: Sulfuric acid (H_2SO_4), vinegar (acetic acid)
- **Neutral solutions:** pH $= 7$ (equal H^+ and OH^-)
  Example: Pure water
- **Basic (alkaline) solutions:** pH > 7 (contain more OH^- than H^+)
  Examples: Sodium hydroxide ($NaOH$), ammonia (NH_3)

I.3.2.3 Concentration

- **Dilute solutions:** Contain a small amount of solute.
  Example: Lightly salted water
- **Concentrated solutions:** Contain a large amount of solute.
  Example: Saturated sugar solution

- **Saturated solutions:** Contain the maximum solute that can dissolve at a given temperature; excess solute precipitates.
- **Supersaturated solutions:** Contain more solute than normally possible at a given temperature; unstable state.

- **Types of Solution Concentrations**

The concentration of a solution can be expressed in several ways:

1. Molarity (C_m):
Number of moles of solute per liter of solution.
Formula: Cp = $\frac{m}{V}$

Example:
If 1 mole of sodium chloride (NaCl) is dissolved in 1 liter of water, the concentration of the solution is **1 M**.

2. Molality (m):
Number of moles of solute per kilogram of solvent.
Formula: Cp = $\frac{m}{V}$

m=solventm n

Example:
If 0.5 mole of glucose is dissolved in 1 kg of water, the molality of the solution is **0.5 m**.

- **3. Mass Percentage (% m/m):**
Mass of solute divided by the total mass of the solution, multiplied by 100.
Formula:
- % mass=msolutemsolution×100\% \, \text{mass} = \frac{m_{solute}}{m_{solution}} \times 100%mass=msolutionmsolute×100

- **Example:**
If 5 grams of NaCl are dissolved in 95 grams of water, the concentration of the solution is:
- 5100×100=5% by mass.\frac{5}{100} \times 100 = 5\% \text{ by mass.}1005 ×100=5% by mass.

- **4. Normality (N):**
Normality measures the number of equivalents of solute per liter of solution. It depends on the particular chemical reaction.
Formula: m = $\frac{n}{m_{solvant}}$
- **Example:**
For a sulfuric acid (H_2SO_4) solution, if 1 mole of H_2SO_4 is dissolved in 1 liter of water, the solution is **2N**, because each mole of H_2SO_4 can release **2 moles of H^+ ions**.

5. Mass Concentration (C_p):
Also called **gravimetric** or **weight concentration**, it measures the amount of solute

(in grams) per liter of solution.
Formula:

Cp = $\frac{m}{V}$ **Example:**
If 5 grams of salt (NaCl) are dissolved in 1 liter of water,

- Cp=5g/L

Relation between C_m and C_p:

C_M et Cp : $n = \frac{m}{V} \Rightarrow CP = \frac{n.M}{V} \Rightarrow CP = C.M$

6. Osmolarity (ω):
Osmolarity is a measure of the total concentration of solute particles in a solution. It is defined as the number of **osmoles of solute per liter of solution**.
Formula:

- $\boldsymbol{\omega = \frac{n}{V}}$

Ionization Coefficient (i):

It is defined as the **ratio between the number of molecules dissociated into ions** and the **total number of molecules initially present** in the solution.
It is expressed by the following formula:

- **Formula:** : $i = \frac{[A^-]}{[A_{initial}]}$

i=[A−][Ainitial]i = \frac{[A^-]}{[A_{initial}]}i=[Ainitial][A−]

Where:

- [A−] : is the concentration of dissociated ions
- [Ainitial]: is the initial concentration of the solute

Interpretation of the ionization coefficient:

- **i = 0:** No dissociation; the acid or base is completely undissociated.
- **i = 1:** Complete dissociation; the acid or base is fully dissociated into ions.
- **0 < i < 1:** Partial dissociation; the acid or base is only partially dissociated into ions.

Example: For a weak acid HA ⇌ H+ + A−

Equivalent Concentration (Ceq)

The **gram equivalent** represents the amount of ions (in grams) carrying one Faraday of charge:

1F=NA×e=6.023×1023×1.6×10−19=96500 coulombs

- Formula : $Ceq = \frac{\text{Nombre d'équivalents grammes}}{\text{Volume de la solution}}$,

  $= \frac{\text{Nombre des charges életriques}}{\text{Volume de la solution}}$
- Unit of Ceq is : Eq/L
- $Ceq = Ceq^{+} + Ceq^{-} = \Sigma Ci^{+}Zi^{+} + \Sigma Ci^{-}Zi^{-}$, and :
- Ci : Ionic Concentration ,
- Zi : valence of ion.
- Total ionic concentration : $Ci_{total} = Ci^{+} + Ci^{-}$

  $= N(\text{cations})\ \alpha C + N(\text{anions})\ \alpha C$

  $= \text{Nions}\ \alpha C = \beta \alpha C$
- Electro-neutrality implies: $\Sigma Ci^{+}\ Zi^{+} = \Sigma Ci^{-}\ Zi^{-}$

  $\Rightarrow Ceq = 2\Sigma Ci^{+}\ Zi^{+}Zi^{+} = 2\Sigma Ci^{-}Zi^{-}$

- **Relationship with the Acid Dissociation Constant (Ka)**
  The ionization coefficient is related to the acid dissociation constant **Ka** for weak acids.
  The acid dissociation constant is defined by the formula:
-

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

By substituting the equilibrium concentrations:

$$K_a = \frac{(C\alpha)(C\alpha)}{C(1-\alpha)}$$

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

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

For small values of **α**, we can approximately say that **1 − α ≈ 1**.

$$K_a = C.\alpha^2 \Rightarrow \alpha = \sqrt{\frac{K_a}{C}}$$

Remark:
Each method of measuring concentration is used depending on the type of analysis or chemical application.

1.3.3 Importance of Electrolytes

An **electrolyte** is a substance that, when dissolved in a solvent (usually water), dissociates into ions and thus allows the conduction of electricity in the solution. Electrolytes can be **ionic compounds** (such as sodium chloride) or **molecules** that dissociate partially or completely into ions when in solution.

1.3.3.1 Types of Electrolytes

Strong Electrolytes:
They dissociate completely into ions in an aqueous solution.
Examples: Hydrochloric acid (HCl), sodium hydroxide (NaOH), potassium nitrate (KNO_3).
Solutions of these electrolytes conduct electricity well due to the presence of many mobile ions.

Weak Electrolytes:
They dissociate partially into ions in an aqueous solution.
Examples: Acetic acid (CH_3COOH), ammonia (NH_3).
Solutions of these electrolytes conduct electricity, but less efficiently than strong electrolytes because they contain fewer ions.

Non-Electrolytes:
They do not dissociate into ions in solution and therefore do not conduct electricity.
Examples: Glucose, urea, ethyl alcohol.

1.3.3.2 Applications of Electrolytes

Electrolytes play a crucial role in many chemical and biological processes. For example:

- **Biology and Medicine:**
  Electrolytes are essential for nerve conduction, water balance, and proper cell function in the human body.
  An electrolyte imbalance can lead to health disorders.
- **Industry:**
  Used in electrochemical processes such as **electrolysis** to produce metals, purify compounds, or generate energy in **cells and batteries**.

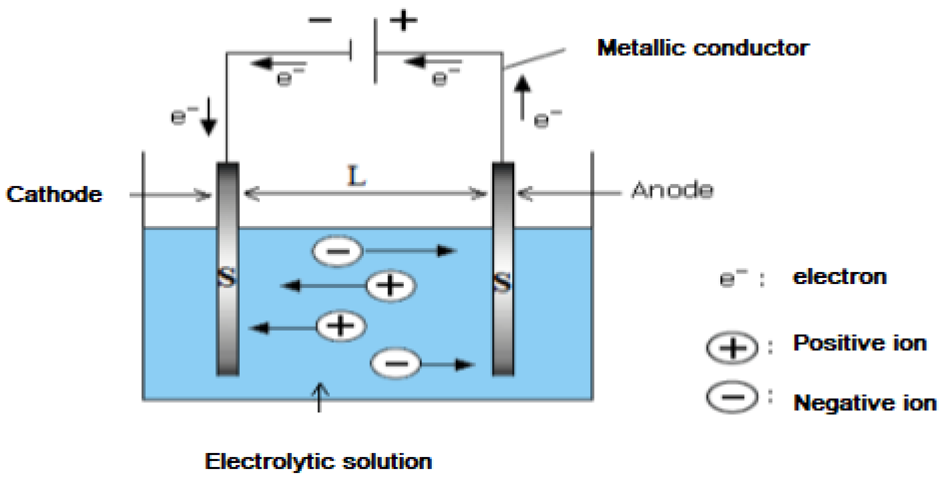
1.4 Electrical Properties of Electrolytic Solutions

1.4.1 Electrical Conductivity

Electric current results from the movement of charge carriers.
The charge carriers are:

- **Ions** in electrolytic solutions (for example: copper ions Cu^{2+} and sulfate ions $(SO_4)^{2-}$ in an aqueous solution of copper sulfate $CuSO_4$).

Let's consider **two electrodes** immersed in an electrolytic solution and apply a potential difference (U) ⇒ the ions and electrons will move as shown in Figure 5 below.



- The current intensity **I** is given by **Ohm's law:**
- Formula : U = R.I
- On the other hand: $R=\rho\ \frac{L}{S}$ et $U=E.L$, Avec : E, repesents electric current

- Electric conductivity is : $\sigma = \frac{1}{\rho} = \frac{L}{R.S} = \frac{I}{E.S}$ in unit : $\Omega^{-1}\ m^{-1}$.
- Equivalent conductivity : $\boldsymbol{\sigma eq}$
  - Formula : $\Lambda eq = \frac{\sigma}{Ceq}$; où l'unité de σeq est : $\Omega^{-1} m^{2}\ Eq^{-1}$;
  - Ceq : is equivalent concentration

- Equivalent conductivity limit : Λeq_0
  - For weak electrolyte : $AxBy \rightleftarrows xA - + yB +$

à t = 0	C	0	0
à t = t'	(1-∝)C	x∝C	y∝C

- The degree of dissociation **α** is given by::
  - Formule : $\propto = \frac{\Lambda eq}{\Lambda eq_0}$

σeq_∞: is the **limiting equivalent conductivity** extrapolated at **infinite dilution**, that is to say, the conductivity value when the **ions are so far apart** that **they no longer interact with each other**.
($Ceq \rightarrow 0$)

- For strong electrolyte (∝ = 1) : AxBy $\rightleftarrows\ xA^{-}\ +\ yB^{+}$

C	0	0
0	xC	yC

$\propto = 1 \Rightarrow \Lambda eq = \Lambda eq_0$

- **Relationship between conductivity (σ) and electrical mobility (μ_{el}):**
  The electrical conductivity of an electrolytic solution is given by the following relation:
- Formula : $\sigma = \Sigma n_+ q_+ \mu_{el^+} + \Sigma n_- |q_-||\mu_{el^-}|$

- n_+ (n_-) : Nombre de cations (anions) de charge q^+ (q^-) par unité de volume.

- μ_{el^+} : Ctionic mobility:$=v_+/E$; μ_{el^-} anionic mobility $=v_-/E$

- v : speed of ion (+ or -) In steady stat.

Thus, σ can be written in the form: :

- Formule : $\sigma = FC\ (\mu el^+ + \mu el^-)$. Avec : C represent concentration.

I.5 Displacement in the Liquid Phase

Displacement in the liquid phase generally refers to the way a liquid moves or behaves in various contexts — such as within a fluid, a solution, or a specific environment.
These movements can result from different **physical or chemical phenomena**, such as **diffusion**, **flow**, or **convection**.
Below is an overview of the main mechanisms of displacement in the liquid phase.

I.5.1 Diffusion

Diffusion is the process by which molecules move from a region of **high concentration** to a region of **low concentration**, until the concentration becomes uniform throughout the medium.
It is a **passive phenomenon**, occurring due to the **kinetic energy** of the molecules.

Example:
When you add a drop of food coloring to a glass of water, the dye slowly diffuses throughout the liquid until the color is evenly distributed.

Fick's First Law:

Diffusion can be described by **Fick's First Law**, which states that the **diffusion flux (J)** is proportional to the **concentration gradient (dC/dx):**

- $J = -D\frac{dC}{dX}$

Where:

- **J** = diffusion flux (amount of substance diffused per unit area per unit time),
- **D** = diffusion coefficient, a constant depending on the nature of the material and the temperature,

- $\frac{dC}{dX}$ = concentration gradient (change in concentration C with respect to distance x).

Example:
Imagine a tube filled with water where one end has a **high concentration of salt** and the other has a **low concentration**. The salt will move from the high-concentration end toward the low-concentration end.

Fick's Second Law:

It describes how the concentration changes **over time** as diffusion occurs.
It is expressed by the following formula:

$$\frac{\partial C}{\partial t} = D\frac{\partial^2 C}{\partial x^2}$$

Example:

Imagine a pond into which a dye is poured. Over time, the dye diffuses throughout the pond, and the concentration of the dye changes with time.

I.5.2 Electrophoretic Migration

Electrophoretic migration refers to the movement of charged particles under the influence of an **electric field**.
This phenomenon is commonly used in **electrophoresis** to separate biomolecules such as **proteins** and **nucleic acids**.

Electrophoretic mobility (μ):
It depends on the **charge of the particle** and the **resistance it encounters** in the medium.

- Formule : $v = \mu E$

Where:

- **v** = migration velocity,
- **μ** = electrophoretic mobility,
- **E** = electric field strength.

I.5.3 Convection

Convection is the movement of molecules in a fluid (liquid or gas) due to **differences in density** or an **external force** such as a pump or gravity.

- **Natural convection:** Caused by density differences resulting from temperature gradients.
- **Forced convection:** Induced by external forces such as pumps or fans.

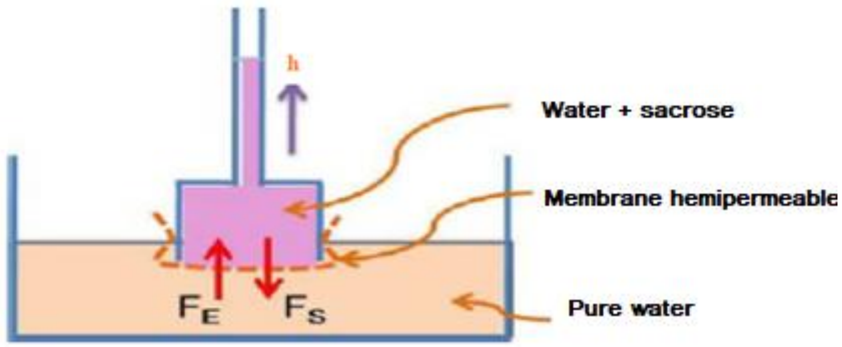
I.5.4 Osmosis

Osmosis is the movement of **solvent molecules** through a **semi-permeable membrane**, from a region of **low solute concentration** to a region of **high solute concentration**, until equilibrium is reached.

Dutrochet's Experiment:

In 1826, **Dutrochet** designed a device called the **osmometer**, consisting of a **glass reservoir** filled with a **colored solution (water + solute)** — shaped like a **vertical cylinder** whose **base is sealed by a semi-permeable membrane** (impermeable to the solute, allowing only the solvent to pass in both directions).

The **upper part** of the cylinder is connected to a **long, narrow vertical tube**, and the entire setup is **immersed in a crystallizing dish containing pure water** (see Figure 6 below).



This setup demonstrates how water enters the solution through the membrane, causing the **liquid level in the tube to rise** a clear manifestation of **osmotic pressure**.

Thus, the osmotic pressure of a solution

The **osmotic pressure** of a solution is the **hydrostatic pressure** that must be applied to the solution to **prevent the pure solvent from passing through the semi-permeable membrane**.
It has been established empirically through numerous experimental studies and is finally expressed theoretically by **Van't Hoff's law**:

- Formule : $\pi = iC_m$RT en (Pa.s) ou (N/m)

where π is expressed in **Pascals (Pa)** or **Newtons per square meter (N/m²)**.

With:

- **i** = ionization coefficient, given by:
  - i=1+α(β−1)i = 1 + \alpha(\beta - 1)i=1+α(β−1) for a **weak electrolyte**
  - i=βi = \betai=β for a **strong electrolyte**
- **C_m** = molar concentration (mol/L)
- **R** = ideal gas constant (8.314 $J \cdot mol^{-1} \cdot K^{-1}$)
- **T** = absolute temperature (K)

Osmotic Pressure:

The osmotic pressure is the **pressure required to stop the osmotic flow** through the membrane.

Example: Solvent Diffusion (Osmosis)

Consider **two compartments**, (1) and (2), at the same temperature, containing solutions of concentrations $\mathbf{C_1}$ and $\mathbf{C_2}$, respectively, with $\mathbf{C_1 > C_2}$.

The two compartments are separated by a **perfect semi-permeable membrane** (allowing only $\mathbf{H_2O}$ to pass through).

- Water diffuses from the **less concentrated medium (2)** to the **more concentrated medium (1)** — this is the **phenomenon of osmosis**.
- The compartment (2) exerts a **hydrostatic pressure** on (1), and at equilibrium, the resultant of all pressures on the membrane becomes **zero**.

We then say that compartment **(1)** exerts on compartment **(2)** an **osmotic pressure** (or counter hydrostatic pressure). **Figure 7** below illustrates this osmotic process.

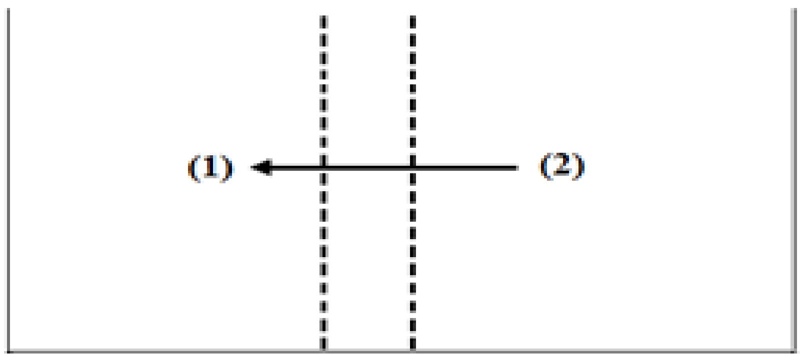


Figure 7: Illustration of Osmotic Pressure

Its formula is: $\boldsymbol{\pi = WRT}$

where:

- $\boldsymbol{\pi}$: osmotic pressure (in pascals, Pa),
- **W** : osmolarity (in osmol/L),
- **R** : ideal gas constant (8.314 $J \cdot mol^{-1} \cdot K^{-1}$),

- **T** : absolute temperature (in kelvins, K).

This equation expresses that **osmotic pressure is directly proportional** to the **osmolarity** of the solution and to its **temperature**.

I.6 Membranes

A **membrane** is a thin, semi-permeable barrier that separates two different environments and controls the passage of substances between them. Membranes can be **natural**, such as cell membranes, or **synthetic**, used in various industrial and medical applications.

I.6.1 Types of Membranes

I.6.1.1 Biological Membranes

A **biological membrane**, or **cell membrane**, is a thin structure mainly composed of **lipids** (primarily phospholipids) and **proteins**, which separates the interior of the cell from its external environment.
It plays a crucial role in **regulating the exchange** of substances between the inside and outside of the cell, by controlling the movement of ions and other molecules.

Functions of Biological Membranes

- **Selective Barrier:**
  They separate the interior of the cell (or organelles) from the exterior, controlling what enters and exits.
  Membranes are semi-permeable, allowing some substances to pass while blocking others.
- **Transport Regulation:**
  They control the movement of ions, nutrients, waste, and other molecules through mechanisms such as **diffusion**, **osmosis**, **facilitated transport**, and **active transport**.
- **Cell Communication:**
  They contain **receptors** and **signaling proteins** that enable cells to receive and transmit chemical signals.
- **Anchorage for Enzymes and Cellular Structures:**
  They provide anchoring sites for enzymes and cytoskeletal elements, facilitating **biochemical reactions** and maintaining **cell structure**.

Structure of Biological Membranes

- **Lipid Bilayer:**
  Composed mainly of **phospholipids** arranged in two layers. The hydrophobic tails

face each other, while the hydrophilic heads face the inner and outer aqueous environments.

- **Membrane Proteins:**
  Include **integral proteins** (spanning the membrane) and **peripheral proteins** (attached to the surface), which play roles in transport, signaling, and structural support.
- **Cholesterol:**
  Found in animal cell membranes; it stabilizes the membrane structure and regulates its fluidity.

Types of Membrane Transport

- **Simple Diffusion:**
  Passive movement of small, nonpolar molecules across the lipid bilayer following the concentration gradient.
- **Facilitated Diffusion:**
  Passive movement of polar or charged molecules through specific transport proteins (channels or carriers).
- **Osmosis:**
  Diffusion of water through a semi-permeable membrane, following the water concentration gradient.
- **Active Transport:**
  Movement of substances against their concentration gradient, requiring energy in the form of **ATP (adenosine triphosphate)**.
- **Endocytosis and Exocytosis:**
  Processes allowing the intake (**endocytosis**) and release (**exocytosis**) of large molecules or particles via **membrane vesicles**.

I.6.1.2 Synthetic Membranes

Synthetic membranes are made from **polymers** or **inorganic materials** and are used in technologies such as **filtration**, **separation**, and **protection**.

- **Filtration Membranes:**
  Used to purify liquids and gases by removing unwanted particles and contaminants.

- **Reverse Osmosis Membranes:**
  Used in **water desalination**, allowing water to pass while retaining salts and other solutes.
- **Electrodialysis Membranes:**
  Used to separate ions in industrial and water treatment processes.
- **Biomimetic Membranes:**
  Incorporate materials and structures inspired by biological membranes, used in **medical devices** and **sensor technologies**.

I.6.1.3 Applications of Membranes

- **Medicine:**
  Renal dialysis, controlled drug release devices, biomedical implants.
- **Food Industry:**
  Filtration and purification of beverages, fruit juice concentration, protein production.
- **Water Treatment:**
  Desalination, drinking water purification, wastewater treatment.
- **Chemical Industry:**
  Separation and purification of chemicals, solvent recovery, biofuel production.

I.6.4 Importance of Membranes

- **Maintaining Homeostasis:**
  Cell membranes play a vital role in maintaining the internal balance of cells and organelles.
- **Advanced Technologies:**
  Synthetic membranes are essential for efficient **separation** and **purification** processes, providing effective solutions in industrial, medical, and environmental

I.7 Diffusion

I.7.1.Laws of Diffusion

Diffusion can be quantitatively described by **Fick's laws**.

First Law of Fick

It describes the **diffusion flux (J)** across a surface as a function of the **concentration gradient**.
It represents the transport of **uncharged macromolecules** from the **more concentrated** compartment to the **less concentrated** one.

- $J = -D\frac{dC}{dX}$

Where:

- **J** : diffusion flux (amount of substance passing through a unit area per unit time)
- **D** : diffusion coefficient (a constant depending on the nature of the particles and the medium)
- $\frac{dC}{dX}$: concentration gradient (variation of concentration **C** as a function of distance **x**)

Second Law of Fick

It describes the **evolution of concentration** of a substance as a function of **time** and **space**.

$$\frac{dC}{dt} = \frac{d^2C}{dX^2}$$

Where:

- $\frac{dC}{dt}$: ate of change of concentration over time
- $\frac{d^2C}{dX^2}$: second derivative of concentration with respect to position

I.7.2 Factors Affecting Diffusion

- **Temperature:** Increasing temperature raises the kinetic energy of particles, thus accelerating diffusion.
- **Particle size:** Smaller particles diffuse faster than larger ones.
- **Viscosity of the medium:** In a more viscous medium, diffusion is slower.
- **Concentration gradient:** A higher gradient results in faster diffusion.

Examples of Diffusion

- **In gases:** Perfumes or odors spread through the air, allowing detection far from the source.
- **In liquids:** Sugar dissolves and diffuses in water until a homogeneous solution is reached.
- **In solids:** Atoms or ions can diffuse through solids, e.g., in doping processes of semiconductors.

I.7.3 Applications of Diffusion

- **Biology:**
  - *Cellular respiration:* Oxygen diffuses through cell membranes into mitochondria for energy production.
- **Medicine:**
  - *Controlled drug release:* Diffusion regulates the release rate of drugs from medical devices (e.g., transdermal patches).
- **Chemical industry:**
  - *Mixture separation:* Diffusion is used to separate gaseous or liquid components.
- **Materials technology:**
  - *Heat treatments:* Diffusion of atoms in metal alloys modifies their mechanical properties.

I.7.4 Diffusion of Charged Particles (Ionic Diffusion)

Ionic diffusion is similar to the diffusion of neutral particles but is **influenced by electrostatic forces** in addition to normal diffusion forces.
Because ions carry electric charges, they are affected by both **concentration gradients** and **electric fields** present in the medium.

I.7.4.1 Mechanisms of Ionic Diffusion

Ionic diffusion is governed by two main mechanisms:

1. **Concentration gradient:**
   Ions move from regions of high concentration to regions of low concentration.
2. **Electric field:**
   Ions also move under the influence of an electric field **cations** (positive ions) move toward the **cathode (negative electrode)**, while
   anions (negative ions) move toward the **anode (positive electrode)**.

I.7.4.2 Laws of Ionic Diffusion

Ionic diffusion can be described by an **extension of Fick's laws** that includes the effect of the electric field.
The **Nernst–Planck equation** expresses the total ionic flux Ji of an ion *i*:

$$J_i = -\, D_i \frac{\partial C_i}{\partial x} - \frac{z_i D_i\, F C_i}{RT}\left(\frac{\partial \varphi}{\partial x}\right)$$

Where:

- $\mathbf{J_i}$: flux of ion *i*
- $\mathbf{D_i}$: diffusion coefficient of ion *i*

- $\partial C_i/\partial x$: concentration gradient of ion *i*
- z_i : valence of ion *i*
- **F** : Faraday's constant
- C_i : concentration of ion *i*
- **R** : gas constant
- **T** : absolute temperature
- $\partial\varphi/\partial x$: electric potential gradient

An alternative form of the **Nernst–Planck equation** is:

$$j_i = -\,D_i\frac{\partial C_i}{\partial x} - \frac{u_i z_i C_i\,F}{RT}\left(\frac{\partial\varphi}{\partial x}\right)$$

with

u_i : is the **ionic mobility**, t**hus**, the ionic mobility ui : is **directly proportional** to the diffusion coefficient D_i and **inversely proportional** to the temperature and the gas constant.
This relationship is known as the **Einstein–Nernst relation**. $u_i = \frac{D_i}{RT}$

I.7.4.3 Factors Influencing Ionic Diffusion

- **Ion concentration:** Higher gradients cause faster diffusion.
- **Electric field:** Can accelerate or slow down ions depending on their charge and the field direction.
- **Temperature:** Higher temperatures increase ion kinetic energy, enhancing diffusion rate.
- **Viscosity:** Greater viscosity reduces diffusion.
- **Ion–ion and ion–solvent interactions:** These can affect the effective diffusion rate.

I.7.4.4 Applications of Ionic Diffusion

- **Electrochemistry:**
  Ionic diffusion is essential in batteries, cells, and electrolytic systems where ions move between electrodes.
- **Cell biology:**
  Ionic diffusion plays a vital role in many biological processes:
  - **Nerve impulse transmission:**
    Neurons transmit electrical signals through diffusion of $\mathbf{Na^+}$ and $\mathbf{K^+}$ ions, generating and propagating **action potentials**.

- **Electrolyte balance and homeostasis:**

  Cells regulate ion diffusion via **ion pumps**, such as the **Na^+/K^+-ATPase**, to maintain internal balance.

- **Cellular respiration:**

  In mitochondria, the diffusion of **H^+ ions (protons)** across the inner membrane drives **ATP synthesis**.

- **Muscle contraction:**

  The release and diffusion of **Ca^{2+} ions** trigger interactions between actin and myosin, leading to muscle contraction.

- **Cell signaling:**

  Entry of ions such as **Ca^{2+}** initiates intracellular signaling cascades that control secretion or immune responses.

- **Intracellular pH regulation:**

  Diffusion of **H^+** and **OH^-** ions through membranes maintains intracellular pH, vital for enzyme function and metabolism.

I.7.4.5 Diffusion Coefficient (D)

The **diffusion coefficient** is given by **Einstein's relation**.
For a spherical diffusing particle with a diameter much larger than that of solvent molecules:

$$D = \frac{K_B T}{6\pi\eta r}$$

Where:

- **D** : diffusion coefficient
- **k** : Boltzmann constant
- **T** : absolute temperature
- **η** : viscosity of the medium
- **r** : radius of the particle

For macromolecules at ordinary temperature:

$$D = 3{,}21.10^{-5}\,\frac{\mathbf{1}}{\sqrt[3]{\boldsymbol{M}}}$$

If the diffusing particle has a diameter of the same order of magnitude as that of the solvent:

$$D = .\frac{\mathbf{1}}{\sqrt{\mathbf{M}}}$$

I.7.2.6 Electric mobility

The electrical mobility of an ion carrying a total charge Q = Zq is given by the equation:

$$\mu_{el} = \frac{Zq}{f}$$

Relationship between electrical mobility (μ_{el}) and diffusion coefficient (D):

$$D = \frac{K_B.T}{f} \quad \Rightarrow \quad D = \mu_{el}\,\frac{K_B T}{Zq}$$

$$K_B = \frac{R}{N_A} \quad \Rightarrow \quad D = \mu_{el}\,\frac{RT}{N_A Z_q}$$

$$F = q\,N_A \quad \Rightarrow \quad D = \mu_{el}\,\frac{RT}{ZF}$$

- Unit : R = 8.314 $JK^{-1}mol^{-1}$,

Q = 1,6.10^{-19} C,

N_A = 6,023.$10^{23}mol^{-1}$,

F = 9,652.10^{4} $C.\,mol^{-1}$,

D : m^2s^{-1}, T en °K, Z : valence of 'ion, μel : $m^2.\,V^{-1}.\,s^{-1}$.

1.8 Colligative Properties

Colligative (from *collus* = stuck): solute molecules interact with (are "stuck to") solvent molecules. A solution that obeys Raoult's law is considered an **ideal solution**.

I.8.1 Definition and Characterization

A **colligative property** of a solution is a property that depends on the **number of particles (solute)** present in the solution. It is therefore governed by a law that depends on the nature of the **solvent** and the **number** of solute particles — **not** on the chemical nature of the solute.

I.8.2 Raoult's Laws

It is observed that solutions freeze at a **lower temperature** than the pure solvent.

$$\boldsymbol{\Delta T_C = K_C\,.\,W}$$

Where:

- **Freezing point depression:**
  - KC: cryoscopic constant (depends only on the solvent). Example value: KC = 1.86 (water).
  - W: osmolarity (osmol·L^{-1})

- **Vapour-pressure lowering (Tonométrie):**
  Introducing a solute B into a solvent A (forming a solution) lowers the vapour pressure.

  The vapour-pressure lowering can be written as:

$$\Delta \mathbf{P} = \boldsymbol{K_a c_l}$$

  Where:

  - CL: molality of the solute
  - Ka: constant depending on the solvent and temperature

- **Boiling point elevation:**
  The boiling point of a solvent in solution is higher than that of the pure solvent:

$$\Delta \mathbf{T} = \mathbf{K_{eb}.C_L}$$

  Where:

  - C_L: molality of the solute
  - K_{eb}: ebullioscopic constant (depends only on the pure solvent)

I.9 Optical Properties of Solutions

Optical properties describe how solutions interact with light. They provide information on composition, concentration, and structure of solutes. Main optical properties: **absorption**, **transmission**, **reflection**, **refraction**, and **scattering**.

I.9.1 Light Absorption

I.9.1.1 Beer–Lambert Law

The Beer–Lambert law relates light absorbance to solute concentration, path length and the molar extinction coefficient:

$$A = \varepsilon\, C\, \ell$$

Where:

- A: absorbance (unitless)
- ε: molar extinction coefficient (molar absorptivity)
- c: solute concentration ($mol \cdot L^{-1}$)
- ℓ: optical path length (usually in cm)

Spectrophotometry is the technique used to measure absorbance at different wavelengths to determine concentration and identify compounds.

Transmission, Reflection, Refraction, Scattering

- **Transmission (T):** fraction of incident light passing through the solution (often expressed as %). Related to absorbance by:

$$T = 10^{-A} x100$$

- **Reflection:** light returned by the surface; depends on refractive indices.
- **Refraction:** change of light direction when entering a medium with different refractive index. **Snell's law**:

$$n_1 \text{sinus } i = n_2 \text{sinus}$$

where n1,n2 are refractive indices and i,r are incident and refracted angles.

- **Scattering:** light dispersed by particles; depends on particle size, shape and concentration.

I.9.1.2 Applications of Optical Properties

- **Chemical analysis:** spectrophotometry to quantify solutes.
- **Quality control:** monitoring purity and concentration in pharmaceutical, food and beverage industries.
- **Biomedical research:** UV–Vis spectroscopy, fluorescence, optical microscopy.
- **Environmental detection:** quantifying pollutants in water and air samples.

Example (calculation):
A solution of NADH has an optical density (absorbance) of **0.31** at **340 nm**. Measurement was made in a cuvette with path length ℓ=1 cm. The molar extinction coefficient of NADH at 340 nm is ε= 6, 2 $10^3\ l.mol^{-1}.Cm^{-1}$

Using Beer–Lambert:

Calculate the concentration
The concentration is :$Cm = 0{,}311 \times 6{,}2 \times 10^3 = 5.10^{-5} mol/l$

Nicotinamide adenine dinucleotide (NAD)
NAD is a redox coenzyme present in all living cells. It is a dinucleotide composed of two nucleotides joined by their phosphate groups.

Formule **:** $C_{21}H_{27}N_{7014}\ P_2$

Molar mass = 663,43g/mol, Point of fusion = 160 °C, solubility : Water

I.I0 Donnan Equilibrium

Consider two compartments separated by a dialysis membrane:

- **Compartment 1** contains only an ionic solution.
- **Compartment 2** contains pure water (see Figure 8).

The ions diffuse and distribute themselves on both sides of the membrane. The presence of ions in each compartment at different concentrations creates a **membrane potential difference**. These are **electrochemical potentials**.

This difference in electrochemical potentials is given by the **Nernst equation**.

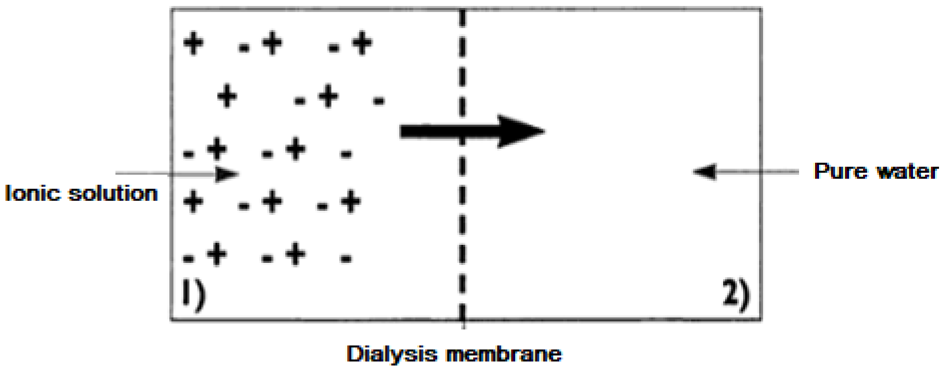


Figure 8: Dutrochet's Application

- Formula:

With:

- **Z** = algebraic value of the ion charge to consider
- Exemple : $Z(Na^{+}) = +1$, $Z(SO_4^{-2}) = -2$
- **F** = Faraday constant = 96,500 coulombs

The ions have distributed themselves **evenly on both sides of the membrane**, so C1=C2.

As a result: ΔV

$$V1 - V2 = -RTZF \ln C2C1 \quad \Delta V = V_1 - V_2 = -\frac{RT}{ZF}\ln\frac{C_2}{C_1}$$

Thus: V1=V2. There is **equality of electrochemical potentials**.

Donnan was the first to describe this phenomenon in **1911**.

I.10.1 Principle

The phenomenon is characterized at **equilibrium**, which is why it is called **Donnan equilibrium**.

As we have seen, when a solution containing only ions is separated from a solution of pure water, there is **equality of electrochemical potentials** (see Figure 9):

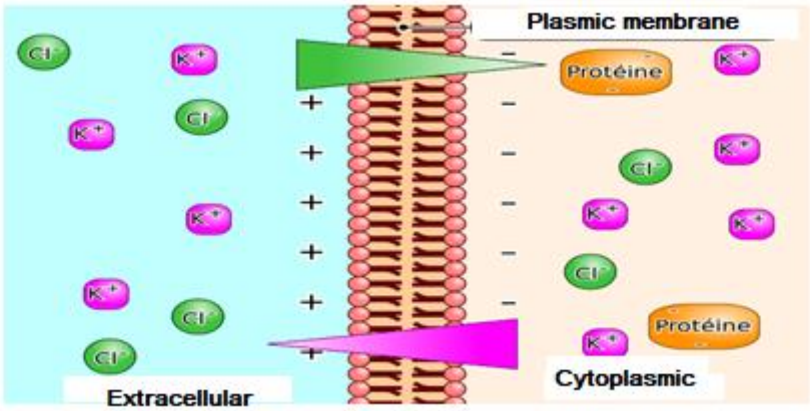


Figure 9: Plasma Membrane

The Gibbs-Donnan effect, or **Gibbs-Donnan equilibrium**,
concerns the behavior of **charged particles near a semipermeable membrane** (typically a charged protein and a dialysis membrane) that are sometimes **not evenly distributed** on each side of the membrane.

The main cause is the presence of **different charged substances that cannot cross the membrane** (proteins or non-diffusible ions), which creates an **unequal electric charge**.

At **diffusion equilibrium**, this charge difference is compensated: the **electric flux and the diffusion flux balance each other**.

The Gibbs-Donnan effect can **prevent sodium-potassium pumps in diseased cells from functioning properly**.

But if there is a **protein in one of the compartments**,

If it is **charged**, it tends to **retain ions of opposite charge**, thus creating **inequalities in ion concentration between the compartments**.

This results in an **equilibrium characterized by a non-zero membrane potential**, called the **Donnan potential**: this is the **Donnan effect**.

I.10.2 Protein Electrophoresis

Protein electrophoresis is a technique used to **separate proteins in a mixture based on their size, charge, or shape**. This method is widely used in **biochemistry, molecular biology, and medical diagnostics** to analyze protein samples (see Figure 10).

The charged particle is subjected to:

- An **electrostatic driving force** given by:

$$\boldsymbol{F = qE}$$

With:

q : charge in Coulombs

E : electric field in V/m

A **frictional force opposing the motion**:

$$\boldsymbol{f = 6\pi\eta vr}$$

- η\ : viscosity coefficient of the medium
- v : velocity of the particle
- rr : radius of the macromolecule

At equilibrium between these two forces, the particle **moves at a constant velocity**:

$$\boldsymbol{qE = 6\pi\eta vr}$$

The macromolecules to be separated are placed on a **support** whose ends are in contact with a **buffer solution** that conducts current from one pole to the other.

- Each buffer solution contains an **electrode**.
- The electrodes are connected to a **power supply**.

When the power supply sends current, **the charged molecules move along the support toward the electrode with the opposite charge**.

Two types of electrophoresis can be performed:

- **Free-flow electrophoresis** (in a liquid stream)
- **Zone electrophoresis** (on a support)

- **Free-flow electrophoresis**

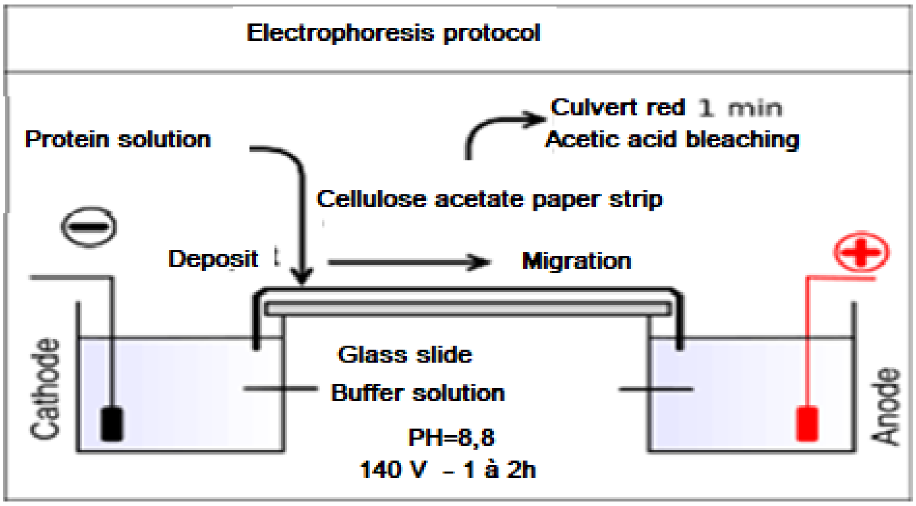
In **free-flow electrophoresis**, migration occurs **within a liquid of defined composition** under a **direct current electric field**. The illustration is shown in Figure 10.

- **Advantage:**
    - It allows **accurate determination of electrophoretic mobilities**.
- **Disadvantages:**
    - Expensive equipment.
    - Long and delicate setup.

- **Zone Electrophoresis**

This type of electrophoresis uses a **porous support** to stabilize the liquid phase.

- The support must be homogeneous, porous, and inert.
- Different types of supports can be used:
  - **Paper or acetate support:** migration of mixtures occurs on the surface.
  - Polyacrylamide or agarose support: migration of mixtures occurs inside the support.



I.11 Concepts of pH and Acid–Base Equilibrium

I.11.1 Concepts of pH

The **pH of a solution** expresses its degree of **acidity or basicity**.

It is defined as the **decimal logarithm (Log10) of the activity of H^+ ions** expressed in gram-equivalent per liter:

$$PH = - Log_{10} [H^+]$$

- The pH of a **neutral solution** is **7**.
- A pH **below 7** indicates an **acidic solution**.
- A pH **above 7** indicates a **basic (alkaline) solution**.

The pH value **depends on temperature**.

I.11.2 Acid-Base Equilibrium

Characterization
Acid-base equilibrium is an important function in the body, aimed at **regulating pH**.

- **Decrease in pH** → acidosis
- **Increase in pH** → alkalosis

For example, **blood plasma** has a pH close to **7**.

- The **pH limits compatible with life** range from **7.0 to 7.8**, with **normal value around 7.4**.

Regulation occurs through three mechanisms:

1. **Physico-chemical buffers**
2. **Respiratory system**
3. **Renal function**

I.11.3 Buffer Effect

Open buffer

- These are buffers where **some molecules can be eliminated by the body**.
- The **only open buffer in the body** is the **carbonic acid/bicarbonate buffer (H_2CO_3 / HCO_3^-)** with **pKa = 6.1**.
- It is called "open" because **CO_2 can escape from the body via respiration**.

  [$C0_3H_2 \Leftrightarrow CO_3H^- + H^+$] et [$C0_3H_2 \Leftrightarrow CO_2 + H_20$]

Closed buffer
In this case, **no molecules from the reaction are eliminated from the body**.

Example of closed buffers: the **erythrocyte buffer = hemoglobin**. These are considered to have **constant mass**, characterized as follows:

A- / AH (de pKa = 6,8).

I.11.4 Acid-Base Status of Plasma

pH of a buffer system (in the case [HCO_3^-/CO_2]):

The pH of the buffer system is written as:

- pH = $pKa + \log_{10} [CO_3H^-] / [CO_2]$ avec pka = 6,1

Henry's Law: explanation and the relationship linking **[HCO_3^-] and pH**.

The **alveolar partial pressure of CO_2 (PCO_2)** defines the concentration of **volatile acid (CO_2)** according to the equation:

- **$[CO_2]$ = a. P_{CO2}**, avec a = solubility coefficient of CO_2

where:

- **$[CO_3H^-]$ = a. P_{CO2}. $10^{(pH-6,1)}$**

CO_2 Equilibration Line – Blood Buffer Line

This function shows **variations of pH and $[HCO_3^-]$ when blood is titrated with a weak acid** (i.e., when blood is exposed to different PCO_2 levels).

This function, which is a line with a non-zero intercept **B** (representing the amount of fixed acids in plasma), is written as:

- Formule: $[CO_3H\text{-}] = - T. pH + B$

- **T** expresses the **buffering capacity of closed systems**.

The two previous equations describe the **acid-base status of plasma**.

I.11.5 Davenport Diagram

In medicine, the **Davenport diagram** is a graphical tool used to **interpret acid-base imbalances in blood**, particularly the relationship between **blood pH, bicarbonate concentration, and partial pressure of CO_2** (see Figure 11).

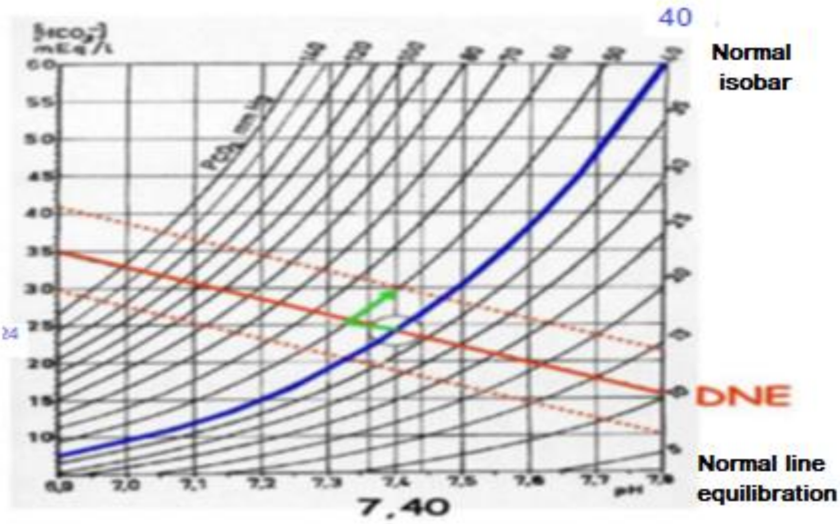


Figure 11: Davenport **Diagram**

Components of the Davenport diagram:

- **X-axis bscissa):** Represents the **bicarbonate concentration [HCO_3^-]** in plasma, usually in mmol/L.
- **Y-axis (ordinate):** Represents **blood pH**, a measure of acidity or alkalinity.
- **Isobaric pCO_2 curves:** Show the relationship between pH and bicarbonate concentration at constant PCO_2 values (e.g., 40 mmHg, 60 mmHg, etc.).
- **Normal point:** Represents the **normal acid-base balance** in the body: pH ≈ 7.4, PCO_2 ≈ 40 mmHg, [HCO_3^-] ≈ 24 mmol/L.

Interpretation of the diagram:

- **Respiratory acidosis:** Increase in PCO_2 due to hypoventilation → equilibrium shifts to lower pH.
- **Respiratory alkalosis:** Decrease in PCO_2 due to hyperventilation → equilibrium shifts to higher pH.
- **Metabolic acidosis:** Decrease in bicarbonate (HCO_3^-) due to excess acids or bicarbonate loss → equilibrium shifts to lower pH (bottom left).
- **Metabolic alkalosis:** Increase in bicarbonate → equilibrium shifts to higher pH (top right).

See Figure 12 for a **graphical representation of pH and [HCO_3^-] variations**.

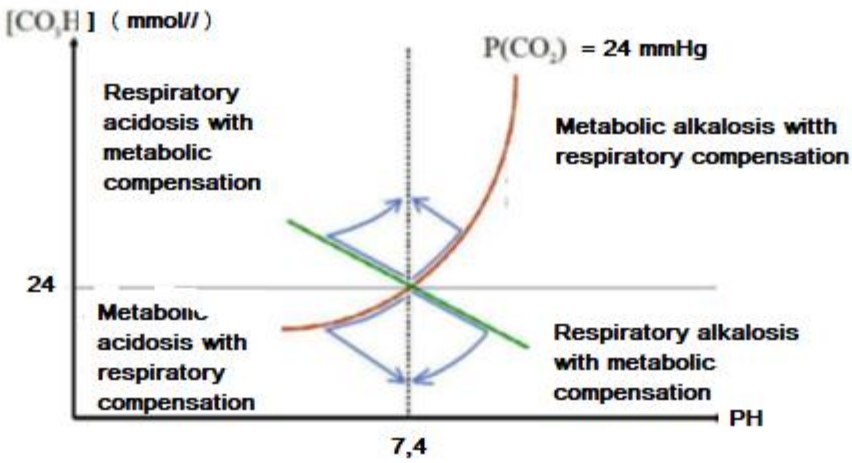


Figure12: graphical representation of pH and [HCO_3^-] variations.

I.12 Medical Applications

I.12.1 Hemodialysis

Dialysis is a medical procedure used to **replace kidney function** in renal failure. It removes **wastes, toxins, and excess fluids** from the blood.

There are mainly two types of dialysis: **hemodialysis** and **peritoneal dialysis**.

I.12.1.1 Principle of Operation

- **Extracorporeal filtration:** Blood is pumped out of the patient and passes through a **dialyzer (artificial kidney)** with a **semipermeable membrane**. Waste and excess fluid move from blood to the **dialysate**, while blood cells and proteins are retained.
- **Return of filtered blood:** Purified blood is returned to the patient.

I.12.1.2 Applications

- **Chronic renal failure (CRF):** For patients whose kidneys have lost most of their function (<15% normal capacity).
- **Acute renal failure (ARF):** Temporary use for patients whose kidneys suddenly stop working but may recover.
- **Toxins/poisoning:** Removal of certain toxins or poisons from blood.

I.12.2 Peritoneal Dialysis

Peritoneal dialysis uses the **peritoneal membrane** in the abdomen as a natural filter.

I.12.2.1 Principle of Operation

- **Dialysate exchange:** Dialysis solution is introduced into the peritoneal cavity via a catheter. Waste and excess fluid cross the peritoneal membrane into the dialysate.
- **Drainage:** After a period, the used dialysate containing waste is drained and replaced with fresh solution.

I.12.2.2 Applications

- **Chronic renal failure (CRF):** Especially for patients who prefer a **flexible method** or cannot tolerate hemodialysis well.

I.12.2.3 Complications and Management

- **Infections:** Risk at vascular access (hemodialysis) or peritoneal catheter.
- **Hypotension:** Can occur during hemodialysis due to rapid fluid removal.
- **Electrolyte imbalances:** Dialysis can alter levels of sodium, potassium, calcium, etc.
- **Fatigue:** Common after dialysis due to blood pressure and electrolyte changes.

Practical Exercises with Answers

1. A solution is:

a. A heterogeneous mixture with at least two phases.
b. A homogeneous mixture in a single phase of at least two substances.
c. None of the above.

Answer: b – A homogeneous mixture in a single phase of at least two substances, generally liquids but can also be solid or gaseous.

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2. A solution is exclusively:

a. Solid
b. Liquid
c. None of the above

Answer: c – Solutions can be solid, liquid, or gaseous.

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3. A solution is called ideal if:

a. The intermolecular forces already present in the pure solvent are not modified by the solute.
b. The intermolecular forces of the solute predominate over those of the pure solvent.
c. None of the above

Answer: a

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4. Molarity is expressed as the ratio of moles of solute present per unit of:

a. Solution volume
b. Solution mass
c. None of the above

Answer: a – Molarity = moles of solute / volume of solution

5. Molality is expressed as the ratio of moles of solute per unit of:

a. Solution volume
b. Solvent mass
c. None of the above

Answer: b – Molality = moles of solute / mass of solvent

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6. Mass concentration expresses the ratio:

a. Solution volume to solute volume
b. Solute mass to solution volume
c. None of the above

Answer: b

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7. For a glucose solution, the osmolarity is:

a. Different from molarity
b. Equal to molarity
c. None of the above

Answer: b – Glucose is non-electrolyte, $i = 1 \rightarrow W = CM$

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8. For a NaCl solution (electrolyte), osmolarity is:

a. Twice the molarity
b. Equal to molarity
c. None of the above

Answer: a – NaCl dissociates fully: $i = 2 \rightarrow W = 2 \times CM$

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9. A solution contains 20 mEq/L of Fe^{3+} (MFe = 56 g/mol). The mass concentration CP is:

a. 2.45 g/L
b. 0.37 g/L
c. None of the above

Answer: b – CP = 0.37 g/L

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10. Mixing 10 cm^3 of glucose solution ($M = 555 \times 10^{-3}$ mol/L, MM = 180 g/mol) with 60 cm^3 of glucose solution (CP = 250 g/L), the mass concentration of the mixture is:

a. 322.45 g/L
b. 228.6 g/L
c. None of the above

Answer: b – CP = 228.6 g/L

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11. Molarity of the mixture from previous question:

a. 1.30 mol/L
b. 0.78 mol/L
c. None of the above

Answer: a – CM = CP / M = 228.6 / 180 = 1.3 mol/L

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12. 0.5 L aqueous solution containing 3 g urea (MM = 60 g/mol). CP =

a. 3 g/L
b. 6 g/L
c. None of the above

Answer: b – CP = 6 g/L

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13. Molal concentration of urea in previous solution:

a. 0.05 mol/kg
b. 0.1 mol/kg
c. None of the above

Answer: b – CML = 0.1 mol/kg

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14. Mixing 2 L NaCl solution (C = 58 g/L) with 0.8 L water, resulting CP:

a. 11.6 g/L
b. 23.2 g/L
c. None of the above

Answer: c – CP = 41.4 g/L

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15. Molarity of solution from previous question:

a. 0.1 mol/L
b. 0.2 mol/L
c. None of the above

Answer: c – CM = 41.4 / 58.5 = 0.71 mol/L

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16. Mixing 500 cm^3 glucose solution (7 g/L) with 2 L water, CP:

a. 0.038 g/L
b. 1.368 g/L
c. None of the above

Answer: b – CP = 1.4 g/L

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17. Two solutions with CP = 0.1 g/L. Sugar solution vs NaCl solution (full dissociation). Osmolarity of NaCl solution:

a. 1.27×10^{-3} osmol/L
b. 3.41×10^{-3} osmol/L
c. None of the above

Answer: b – W = 2 × CM = 0.0034 osmol/L

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18. 0.5 L solution containing 1.64 g Na_2PO_4 (MM = 164 g/mol) + 4.5 g glucose (MM = 180 g/mol). Osmolarity W:

a. 0.13 osmol/L
b. 0.08 osmol/L
c. None of the above

Answer: c – W = 0.11 osmol/L

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19. Mixing 15 cm^3 10% glucose with 70 cm^3 25% glucose (diluted). CP of mixture:

a. 223 g/L
b. 145 g/L
c. None of the above

Answer: c – Needs calculation based on mass/volume

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20. Mole fraction is:

a. Ratio of moles of a component to total solution volume
b. Ratio of moles of a component to total moles of all components
c. None of the above

Answer: b

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21. NaCl solution fully dissociated. Concentration:

a. Different from molarity
b. Equal to molarity
c. None of the above

Answer: a – Osmolarity differs due to dissociation (W = 2 × CM)

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22. Mole fraction expressed as:

a. Ratio of moles to total solution volume
b. Ratio of moles to total moles of all constituents
c. None of the above

Answer: b

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23. Mixing 20 cm^3 glucose solution CM = 120×10^{-3} mol/L with 40 cm^3 CP = 400 g/L glucose, CP =

a. 416.2 g/L
b. 273.8 g/L
c. None of the above

Answer: b – CP = 273.8 g/L

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24. Two solutions CP = 3 g/L. Sugar solution vs NaCl (full dissociation). Osmolarity W of NaCl solution:

a. 1.27×10^{-3} osmol/L
b. 3.4×10^{-3} osmol/L
c. None of the above

Answer: c – W = 0.102 osmol/L

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25. 500 cm³ solution containing 4 g urea. CP =

a. 4 g/L
b. 8 g/L
c. None of the above

Answer: b – CP = 8 g/L

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26. Adding 0.5 L water to previous solution, mole fraction of urea:

a. 1.2×10^{-3}
b. 0.34
c. None of the above

Answer: a – FM = 1.2×10^{-3}

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27. Molal concentration of urea in previous solution:

a. 1.54×10^{-2} mol/kg
b. 6.67×10^{-2} mol/kg
c. None of the above

Answer: b – CML = 0.0667 mol/kg