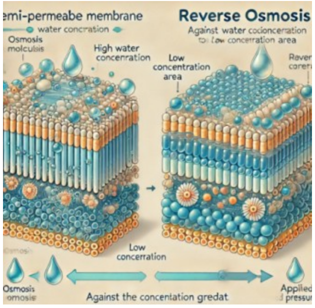
Chapter IV:

Osmosis, Reverse Osmosis, and Dialysis



Introduction:

This chapter explores the fundamental principles of osmosis, reverse osmosis, and dialysis, processes that are essential across a wide range of scientific and industrial applications. Osmosis is the natural movement of water through a semi-permeable membrane, driven by differences in solute concentration. This process is vital in biological systems, helping to regulate water balance in cells and organs.

Reverse osmosis, on the other hand, applies external pressure to force water through a membrane in the opposite direction, removing impurities and solutes. This technique is widely used in water purification and desalination, providing clean drinking water and other critical applications.

Dialysis similarly leverages membrane technology to separate smaller molecules from larger ones. It is particularly important in medical treatments, such as kidney dialysis, where it removes waste products from the blood.

The chapter also introduces the Donnan effect, which describes the distribution of ions between two solutions separated by a semi-permeable membrane and highlights the influence of charge distribution on osmotic behavior. Together, these topics provide insight into the molecular processes that govern the movement of substances in solutions, with implications in biology, medicine, and environmental sciences.

IV. 1 Osmosis:

Osmosis is the phenomenon of matter diffusion characterized by the movement of solvent molecules from one solution to another through a semi-permeable membrane that separates two solutions with different solute concentrations. The overall solvent transfer occurs from the less concentrated solution (***hypotonic medium***) to the more concentrated solution (***hypertonic medium***) until equilibrium is reached (***isotonic media***). This phenomenon pertains exclusively to exchanges between two liquid solutions with different solute concentrations, separated by a ***semi-permeable*** barrier.

This concept led to a better understanding of the behavior of aqueous solutions in chemistry in the late 19th century. It is also particularly useful in physiology and cellular biology for explaining chemical exchanges within living organisms.

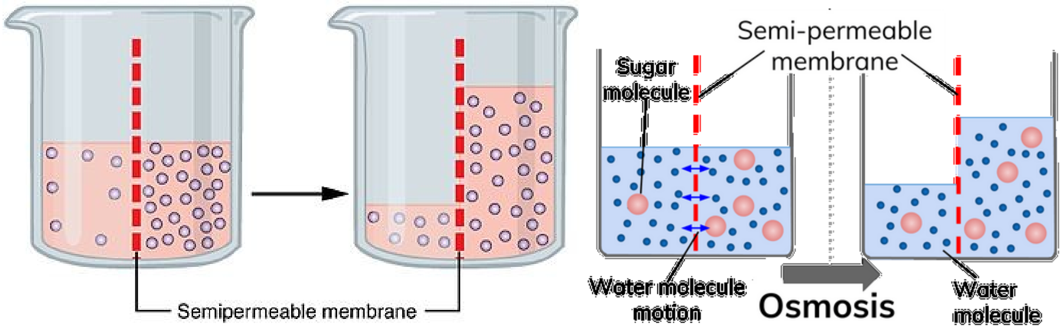


Fig IV. 1: Osmosis[1]

IV.1.1 Definition:

Osmosis is the process where water moves from a region of low solute concentration to a region of high solute concentration through a semi-permeable membrane.

IV.1.1.1 Examples:

- Absorption of water by plant roots.
- Cell swelling when placed in a hypotonic solution.
  - **Hypotonic**: Lower solute concentration outside the cell, causing the cell to swell.
  - **Hypertonic**: Higher solute concentration outside the cell, causing water loss from the cell.
  - **Isotonic**: Equal solute concentrations, resulting in no net change in cell volume.

IV.1.2 Chemical Potential of Water as a Function of Solute Concentration:

The chemical potential of water in each compartment depends on the solute concentration (x) via the following relationship

$$\mu = \mu^0 + R.T.\mu(1-x)$$

$$dG = VdP - SdT + \sum_{i=1}^{n} \mu_i dn_i$$

$$\mu_i = \frac{\partial G}{\partial n_i}$$

Where:

- μ_0 : is the chemical potential of pure water (without solute),
- R : is the gaz constant,
- T : is the temperature in Kelvin
- X : is the mole fractioin of solute in the compatement ($0 \leq x \leq 1$)

IV.1.3 Osmosis and Chemical Potential Difference:

Consider two compartments separated by a semi-permeable membrane, with different solute concentrations:

- In Compartment 1, the solute concentration is x1

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[1] https://en.wikipedia.org/wiki/Osmosis

- In Compartment 2, the solute concentration is x2, where x2 > x2

The chemical potential of water in each compartment can then be written as:

$$\mu_1 = \mu^0 + R.T.\mu(1 - x_1)$$

$$\mu_2 = \mu^0 + R.T.\mu(1 - x_2)$$

Since x2 > x1 we have m1 > m2 water thus moves Compartment 1 (lower solute concentration, higher chemical potential) to compartment 2(higher solute concentration, lower chemical potential) in order to equalize the chemical potentials**.**

IV.1.5 Osmotic concentration / osmolarity:

Osmotic concentration or osmolarity **is** a measurement that indicates the total concentration of solutes (ions, molecules) dissolved in a solution, and it is expressed in osmoles per liter of solution (osmol /L). Osmolarity allows us to understand how water will move between different solutions by a phenomenon called **osmosis**.

Osmolarity is essential in many biological and medical processes, such as regulating the body's water balance and administering intravenous solutions.

To calculate the osmotic concentration or osmolarity of a solution, one must know the molar concentration of the solutes and the number of particles that these solutes produce when they dissociate in the solution.

The general formula for osmolarity (O_{sm}) is:

$$O_{sm} = C \times i$$

- **O_{sm}** : Osmolarity in osmoles /L (osmol /L)
- **C** : Molar concentration of the solute in moles/L (mol/L)
- **i** : Dissociation coefficient, i.e. the number of particles (ions or molecules) produced by the dissociation of each solute molecule.

Calculation example:

Let's take a solution of sodium chloride (NaCl) at a concentration of 0.1 mol/L:

1. *NaCl dissociates into Na^+ and Cl- ' giving i=2 .*
2. *The molar concentration is 0.1 mol/L.*

So O_{sm} =0.1 mol/L×2=0.2 osmol /L

(a) Calculation of entities formed by dissociation:

The total number of particles (entities) present in the solution after the dissociation of a solute can be calculated with the following formula:

$$\mathbf{i = 1 + \alpha\,(n - 1)}$$

- **i : Van 't** Hoff coefficient , which corresponds to the number of particles or entities after dissociation.
- **α** : Degree of dissociation, which represents the fraction of solute molecules that dissociate (between 0 and 1).
- **n** : Total number of particles formed by a solute molecule when it is completely dissociated.

Explanations:

- **Degree of dissociation (α)**: It indicates the proportion of solute molecules that dissociate into ions. For example, if α = 0.7. it means that 70% of the solute molecules dissociate into ions, while 30% remain as intact molecules.
- **Total number of ions (n)** : This is the number of ions or particles formed when a solute molecule completely dissociates. For example, sodium chloride (NaCl) dissociates into two ions (Na^+ and Cl^-), so n = 2.

Calculation example:

Let us take the example of a NaCl solution with a degree of dissociation of α=0.8.

- *NaCl dissociates into* Na^+ *and* Cl^-*, so n = 2.*
- *α=0.8, which means that 80% of the NaCl molecules dissociate into ions.*

The formula becomes:

$i=1+0.8\,(2-1) = 1.8$

(b) Van 't Hoff 's law (for osmosis):

Hoff 's law allows us to relate the osmotic pressure (π) to the osmolarity of a solution:

Jacobus Henricus van 't Hoff Jr. (Dutch; 30 August 1852 – 1 March 1911) was a Dutch physical chemist. A highly influential theoretical chemist of his time, van 't Hoff was the first winner of the Nobel Prize in Chemistry. His pioneering work helped found the modern theory of chemical affinity, chemical equilibrium, chemical kinetics, and chemical thermodynamics

The osmotic pressure can be calculated using ***Van't Hoff's*** equation for a dilute solution

$$\pi = i \times C \times R \times T$$

- **π** : Osmotic pressure (in pascals, Pa)
- **i** : Dissociation coefficient (number of particles formed)
- **C** : Molar concentration of solute (mol/L)
- **R** : Ideal gas constant
- **T** : Absolute temperature (in kelvins, K)

(c) Raoult's law

Raoult's law concerns the decrease in the vapor pressure of a solvent when a solute is dissolved in it:

$$P = P_0 \times (1 - X_2)$$

- **P** : Vapor pressure of the solution
- **P_0** : Vapor pressure of the pure solvent

- **X_2** : Molar fraction of the solute

It is useful for ideal solutions and allows calculation of colligative properties such as vapor pressure depression, boiling point and freezing.

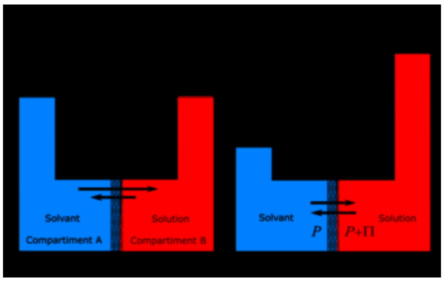

Fig.IV.2: osmotic pressure measurement[2]

(d) Measurement of osmotic pressure (Dutrochet's apparatus):

René Joachim Henri Dutrochet: (14 November 1776 - 4 February 1847) was a French physician, botanist and physiologist. He is best known for his investigation into osmosis

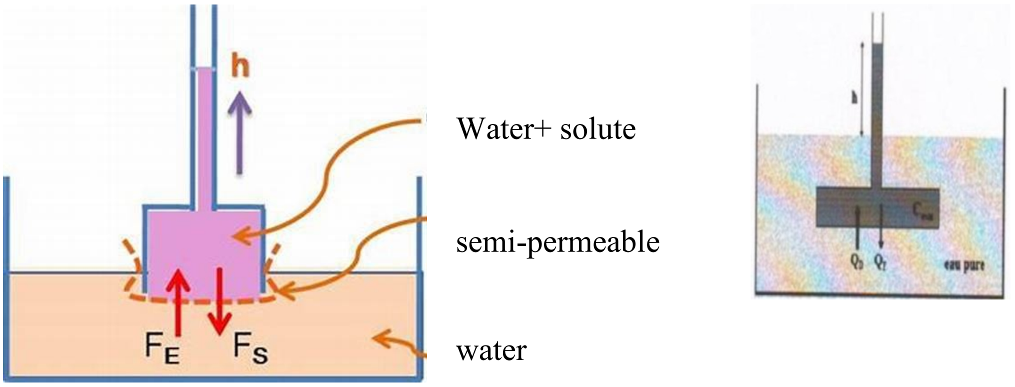

Fig. IV.3: Dutrochet Osmometer[3]

At equilibrium: osmotic pressure = hydrostatic pressure

Hydrostatic Pressure: $\Delta P = \rho g \Delta h$

Osmotic pressure: $\pi = i \times C \times R \times T$

(e) Henry's Law

Henry's law is used to determine the solubility of gases in liquids as a function of pressure:

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[3] https://en.wikipedia.org/wiki/Osmosis

$$C = k_H \times P$$

- **C** : Concentration of dissolved gas (mol/L)
- **k_H** : Henry constant (which depends on the nature of the gas and the solvent)
- **P** : Partial pressure of gas above liquid

It is particularly important for gases dissolved in liquids, for example the dissolution of CO_2 in carbonated drinks.

(f) Dalton's law (on partial pressures):

For a mixture of gases, Dalton's law allows us to determine the total pressure by adding the partial pressures of each gas:

$$P_{total} = \sum_i P_i$$

- **P_{total}** : Total pressure of the mixture
- **P_i** : Partial pressure of gas iii

This law is relevant for understanding the behavior of gases in solutions and calculating the osmolarity of dissolved gas solutions.

(g) Law of tonicity:

Tonicity describes how a solution influences the volume of a cell through osmosis. The concepts of **hypotonic** , **isotonic** , and **hypertonic solutions** are used to explain how a cell will expand or contract when placed in a solution, depending on the relative osmolarity of the solution and the intracellular environment.

(h) Fick 's Law (for diffusion):

Fick 's law describes the diffusion of particles across a membrane:

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

Adolf Eugen Fick (3 September 1829 – 21 August 1901) was a German-born physician and physiologist.

- **J** : Diffusion flux (amount of substance per unit area and time)
- **D** : Diffusion coefficient
- $\frac{dC}{dX}$: Concentration gradient

It is useful for understanding the diffusion of solutes in a solution, a process related to osmosis.

These laws provide a deeper understanding of the phenomena related to osmolarity and diffusion in solutions. Each applies to specific contexts and can be used as needed for precise calculations.

IV.2 Reverse osmosis:

Reverse osmosis is a process in which solvent molecules pass through a semi-permeable membrane in the opposite direction of natural osmosis. In this case, external pressure is applied to the more concentrated solution (hypertonic), forcing solvent molecules to move toward the less concentrated solution (hypotonic). This process reverses the natural flow of the solvent, allowing for the purification or desalination of water by separating solutes from the solvent.

Reverse osmosis is widely used in water treatment, particularly for producing drinking water from seawater, water purification in chemistry, and in industrial and medical applications where pure solutions are required.

IV.3 Dialysis:

c. In medicine, dialysis specifically refers to a method of blood purification through a membrane. There are different types of membranes: semi-permeable membranes, which allow only the solvent to pass, and dialysis membranes, which have pores on the scale of nanometers (nm) that are uniform and well-defined, allowing the solvent and smaller solutes to pass through.

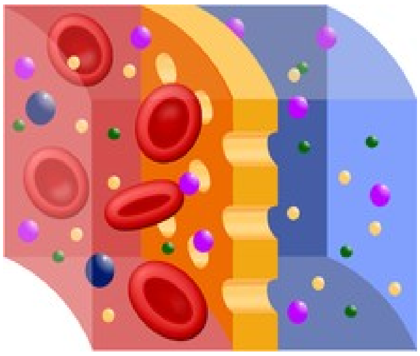

Fig. IV.4: Dialysis[4]

The main medical application of dialysis is for individuals whose kidneys have stopped functioning, whether temporarily (acute kidney failure) or permanently (end-stage chronic kidney disease).

- **Definition**: Dialysis is a process where a semi-permeable membrane is used to separate small molecules and ions from large molecules in a solution.
- **Principle**: The solvent and small molecules diffuse through the membrane, while large molecules remain trapped.

IV.4 Donnan Effect

Introduction

The Donnan Effect, also known as Donnan equilibrium, is a phenomenon observed when charged ions distribute unevenly across a semi-permeable membrane. This imbalance is caused by the presence of charged particles that cannot cross the membrane, leading to an unequal distribution of diffusible ions on each side of the membrane

The Gibbs–Donnan effect (also known as the Donnan's effect, Donnan law, Donnan equilibrium, or Gibbs–Donnan equilibrium) is a name for the behaviour of charged particles near a semi-permeable membrane that sometimes fail to distribute evenly across the two sides of the membrane. The usual cause is the presence of a different charged substance that

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[4] https://en.wikipedia.org/wiki/Dialysis_(chemistry)

is unable to pass through the membrane and thus creates an uneven electrical charge.

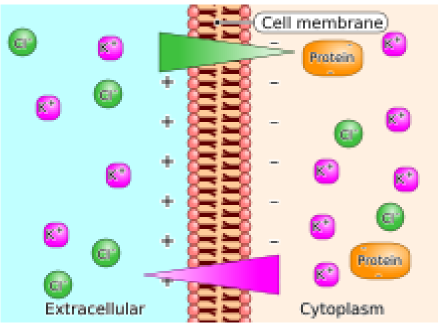

Fig. IV.5 Schematic: Donnan equilibrium across a cell membrane[5]

For example, the large anionic proteins in blood plasma are not permeable to capillary walls. Because small cations are attracted, but are not bound to the proteins, small anions will cross capillary walls away from the anionic proteins more readily than small cations.

The Donnan Effect describes the unequal distribution of ions across a semi-permeable membrane when there are non-diffusible ions on one side of the membrane. For instance, if a membrane separates two solutions, one of which contains a negatively charged protein (non-diffusible ion), the diffusible ions (like Na^+ and Cl^-) will distribute unevenly between the two sides of the membrane to maintain electrical neutrality. However, this balance will not be perfect due to the presence of non-diffusible ions, leading to a difference in ion concentration and an electrical potential difference between the two compartments, known as the **Donnan potential**.

This effect is essential in cellular biology because it influences the distribution of ions across cell membranes, affecting osmotic balance and cell volume. For example, in cells, proteins and other negatively charged macromolecules cannot cross the plasma membrane, creating an ionic gradient and causing water movement through osmosis, which helps regulate cell volume.

IV.4.1 Principle of the Donnan Effect:

(a) Semi-permeable Membrane and Non-diffusible Ions:

A semi-permeable membrane allows the passage of certain ions (such as Na^+, Cl^-) while retaining larger molecules or ions (like charged proteins). These non-diffusible ions or molecules present on one side of the membrane create a charge difference that influences the movement of diffusible ions.

(b)-Donnan Potential: When ions are distributed unevenly on each side of the membrane, an electrical potential develops to balance this charge difference. This potential is known as the Donnan potential.

(c) Ionic Distribution: The diffusible ions (e.g., Na^+, Cl^-) will move to balance the charge on both sides of the membrane but cannot completely neutralize the non-diffusible ion, leading to an asymmetrical distribution.

- ✓ **Donnan Equilibrium Equation:**

The Donnan effect is often quantified by the Donnan equilibrium equation, which relates the ion concentrations on both sides of the membrane:

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[5] https://en.wikipedia.org/wiki/Gibbs%E2%80%93Donnan_effect

$$\frac{\left[C^{+}_{\text{inside}}\right]}{\left[C^{+}_{\text{outside}}\right]}=\frac{\left[A^{-}_{\text{outside}}\right]}{\left[A^{-}_{\text{inside}}\right]}$$

where:

- $[C^{+}]$ and $[A^{-}]$ represent the concentrations of positive and negative ions, respectively.
- "inside" and "outside" refer to each side of the membrane.

(b) Nernst Equation and Ion Equilibrium:

The **Nernst Equation** describes the relationship between an ion's concentration gradient across a membrane and the electrical potential it generates at equilibrium. This potential, known as the **equilibrium potential** (or Nernst potential), is essential for maintaining the electrochemical gradient across cell membranes.

For an ion X with a charge z, the Nernst Equation is:

$$E=\frac{RT}{zF}\ln\frac{[X]_{\text{outside}}}{[X]_{\text{inside}}}$$

Where:

- E is the equilibrium potential,
- R is the ideal gas constant,
- T is the temperature in Kelvin,
- F is the Faraday constant,
- $[X]_{\text{outside}}$ and $[X]_{\text{inside}}$ are the ion concentrations outside and inside the cell.

The Nernst Equation shows that the equilibrium potential depends on the concentration gradient and charge of the ion. In biological systems, this is used to calculate the potential across a membrane for ions like K^{+}, Na^{+}, and Cl^{-}.

(c) Donnan's ratio r is expressed by the relation:

$$r=\frac{[X]_{\text{outside}}}{[X]_{\text{inside}}}$$

(d) Membrane Potential and Relationship to Donnan Effect:

The **membrane potential** is the overall electrical potential difference across a cell membrane, largely determined by the Nernst potentials of various ions and their relative permeabilities. The Donnan Effect contributes to membrane potential by causing an uneven distribution of diffusible ions in response to non-diffusible ions. The Nernst potentials for each ion help calculate the final membrane potential.

In cells, the Donnan Effect results in an accumulation of positive ions inside the cell to counterbalance negatively charged proteins, while the Nernst Equation helps us calculate the potential for each individual ion across the membrane.