8.51.27 A student scheduled for three final examinations in one day shall report the conflict to the instructors as soon as the Common Examination schedule is posted. **On the day of the examinations, the student shall take two examinations in the order in which these classes occur in the week of the student's class schedule** (Lecture has priority over Recitation and Laboratory). It is the responsibility of the student to arrange a time with the instructor of the third class not later than 7 days prior to the last class day to take the final examination at a time that would not result in the student taking three final examinations in one day. **The instructor of the class occurring third shall make reasonable accommodation for the student to take the examination in a timely fashion even should it result in the student having two examinations on another day during finals.** #15-16-30

Chapter Thirteen

Physical Properties Of Solutions

Review: Solution Terms & Types

Solvent: Larger portion of a solution
Solute: Smaller portion of a solution
Solution: A homogeneous mixture of two or more elements or compounds
Solubility: Measure of max amount of solute in solution

TABLE 13.1 Types of Solutions

Solute	Solvent	State of Resulting Solution	Examples
Gas	Gas	Gas	Air
Gas	Liquid	Liquid	Soda water (CO_2 in water)
Gas	Solid	Solid	H_2 gas in palladium
Liquid	Liquid	Liquid	Ethanol in water
Solid	Liquid	Liquid	NaCl in water
Solid	Solid	Solid	Brass (Cu/Zn), solder (Sn/Pb)

More Solubility Terms

Saturated solution:

Maximum amount of solute that stays in solution
- Any additional solute will precipitate

Unsaturated solution:

Contains less solute than in saturated solution

Supersaturated solution:

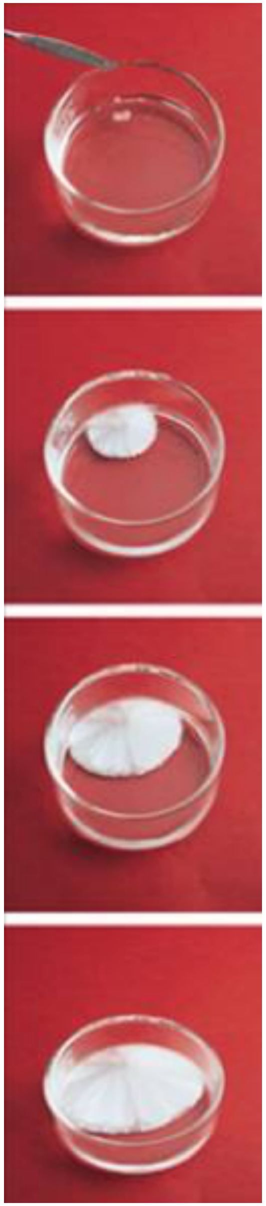
Contains more solute than in saturated solution
- Extremely unstable (easy to make extra precipitate)

Crystallization:

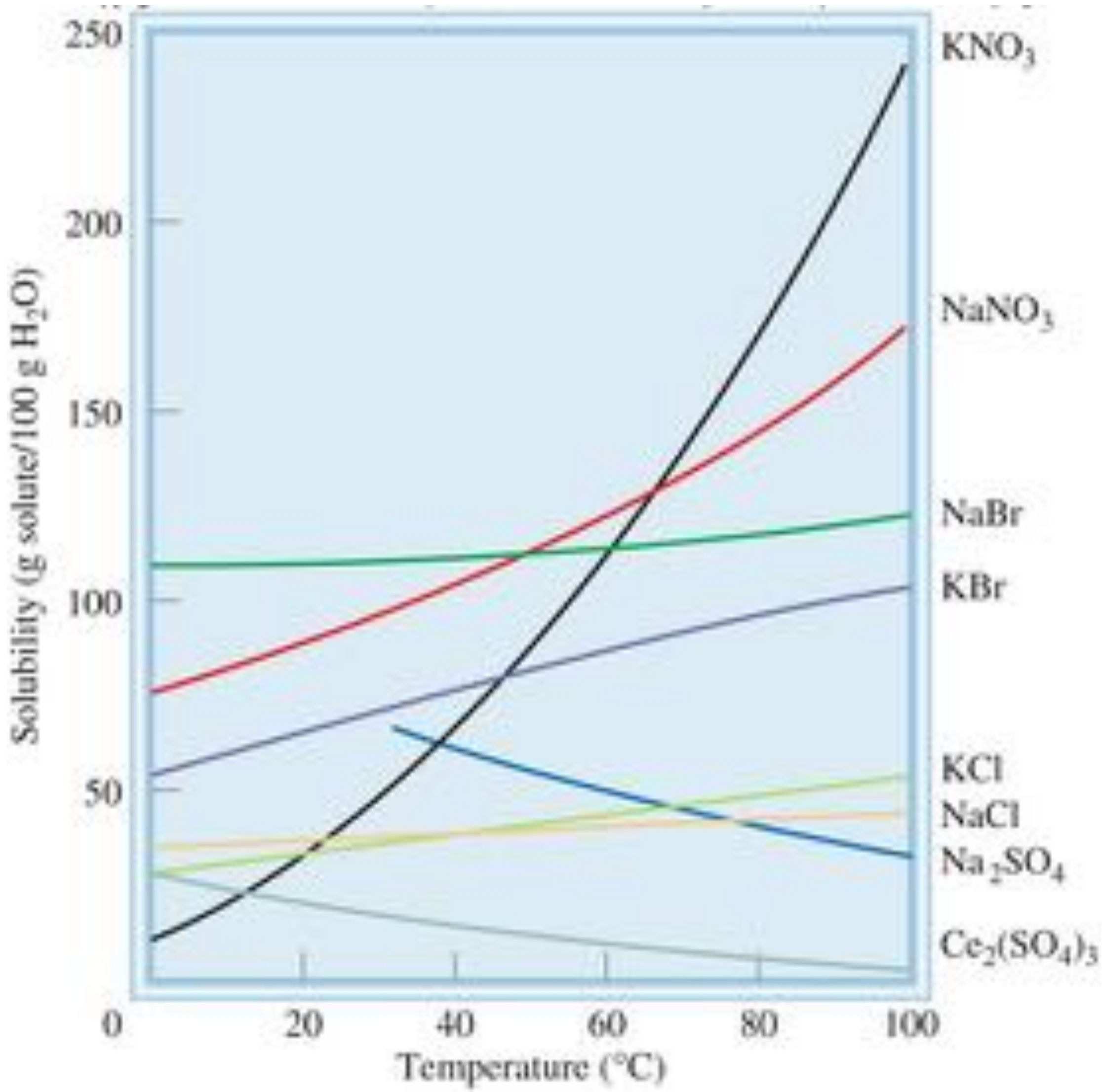
Extra solute in supersaturated solution precipitates & forms crystals

Precipitation:

Solid comes out of solution, not always a crystal



Solubility of Solids As A Function Of Temperature



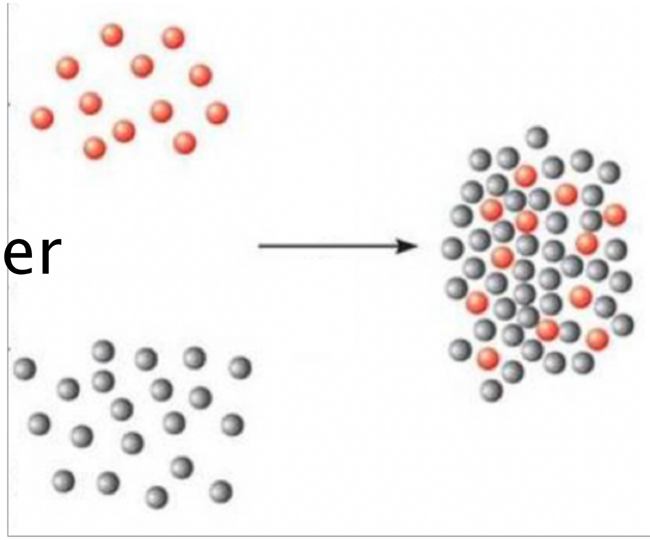
For most solids, as temp increases, solubility increases.

Solvent/Solute Intermolecular Forces & Polarity

Rule of thumb is that "like dissolves like."

Intermolecular forces are strongest for similar compounds

- Polar solvent/polar solute
- Nonpolar solvent/nonpolar solute
- Similar attractive forces lead to solvent and solute that are "miscible"
- **miscible** = fully dissolve in one another
  - resulting solution is stable



Examples:

- water/ethanol solutions
- Dissolution of ionic salts in H_2O
- CCl_4 in benzene (C_6H_6)

Concentration Units Review

$$\text{Molarity (M)} = \frac{\text{Moles Solute}}{\text{Liters of solution}} = \frac{\text{mol}}{\text{L}}$$

$$\text{Molality (m)} = \frac{\text{Moles Solute}}{\text{kg of solvent}} = \frac{\text{mol}}{\text{kg}}$$

Molarity is temperature dependent,
Molality is NOT temperature dependent

$$\text{Percent by mass} = \frac{\text{Mass solute (g)}}{\text{Mass of solution (g)}} \times 100$$

Grams cancel so no units, just % sign

ppm = part per million ($1g/1x10^6g$)

How many grams of sodium hydroxide are present in 0.500 kg of water if the solution concentration is 0.500 m?

A: 10.0 g

What mass of sodium hydroxide would be needed to make 4.00 L of a 0.300 m solution at 20°C?
(density of water is 0.998 g/mL at 20°C)

A: 47.9 g

A 14.0% by mass acetic acid (CH_3COOH) solution has a density of 1.02 g/mL. What is its molality? What is its molarity?

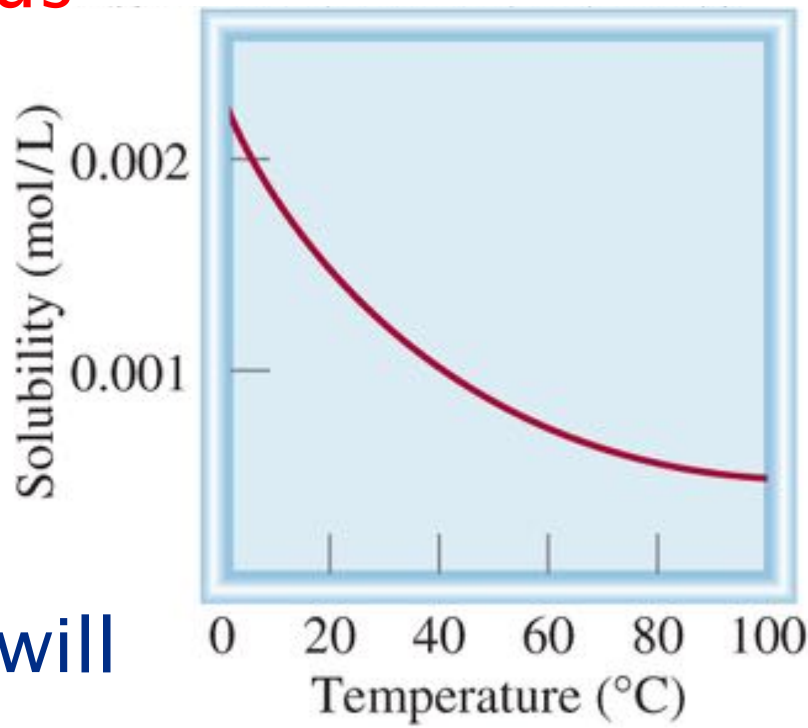
Molality = 2.71 m
Molarity = 2.38 M

The Solubilities Of Gases: Effect of Temperature

Gases are *less* soluble in liquids as temperature increases

- Molecules move faster
- Disrupt intermolecular forces
- Gases escape the liquid
- Fewer gas molecules/atoms in liquid = Lower solubility

Solubility of O_2 in water



In sealed container, pressure will increase with increasing temp.

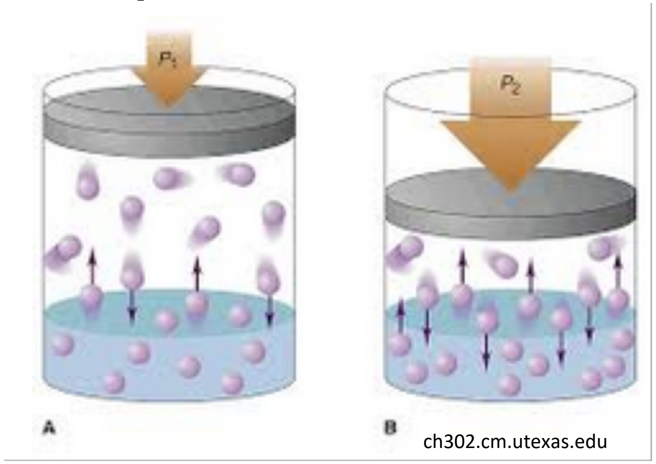
- More gas out of solution, moving faster, colliding more with container

Gas laws (Chapter 5!) govern gas phase and thus solubility

The Solubilities Of Gases: Effect of Pressure

Gas solubility in liquids increases as pressure increases

- Molecules pushed too close together so they are forced back into liquid



If pressure is decreased.....

Henry's Law c=kP

c= solubility or concentration
P= pressure
k= proportionality constant

Common units for k: c= moles/L and P= atm

If you know the concentration at one pressure, k allows you to determine concentrations at different pressures

Vapor Pressures of Solutions

Raoult's Law:

$$P_a = X_a P^\circ_a$$

P_a: Vapor pressure of solvent "a" above a solution

P^o_a: Vapor pressure of pure solvent "a"

x_a: Mole fraction of "a" in the solution

$$x_i = \frac{moles_i}{Moles_{total\ solution}} = \frac{n_i}{n_{total\ moles}}$$

Works perfectly for ideal solutions

Works ok for the solvent in dilute solutions

- Solvent molecules in environment similar to pure solvent.

What is the vapor pressure of water in a 100.0g solution that contains 10.0g sucrose, $C_{12}H_{22}O_{11}$ at 40°C? The vapor pressure of pure water is 55.5 torr at 40°C.

MM: sucrose = 342.30 g/mol; water = 18.015 g/mol)

Step 1:Determine mole fraction of water

Step 2:Determine vapor pressure given P°= 55.5 torr at 40°C

A: 55.2 torr

What mass (g) of ethanol (46.069 g/mol) must be added to 550.0 g water (18.015 g/mol) to give a solution with a vapor pressure 1.5 mmHg less than that of pure water at 30°C? Vapor pressure of water at 30°C is 31.8 mmHg.

A: 69.6 g

Solutions Of Electrolytes

Colligative properties:

- Physical properties of solutions that depend on the number of solute particles but ***NOT*** on the identity of the solute.
- Ex: Boiling Point, Freezing Point, Osmotic Pressure

van't Hoff factor, *i*

- Used in solution equations for colligative properties (ex. Boiling point elevation: $\Delta T_b = iK_bm$)
- Nonelectrolytic solutions: *i* = 1.
- Electrolyte solutions: *i* = the number of ions the solute will dissociate into
- Ex: Hexane, *i* = 1
  NaCl, *i* = 2
  $Pb(NO_3)_2$, *i* = 3

Boiling Point Elevation

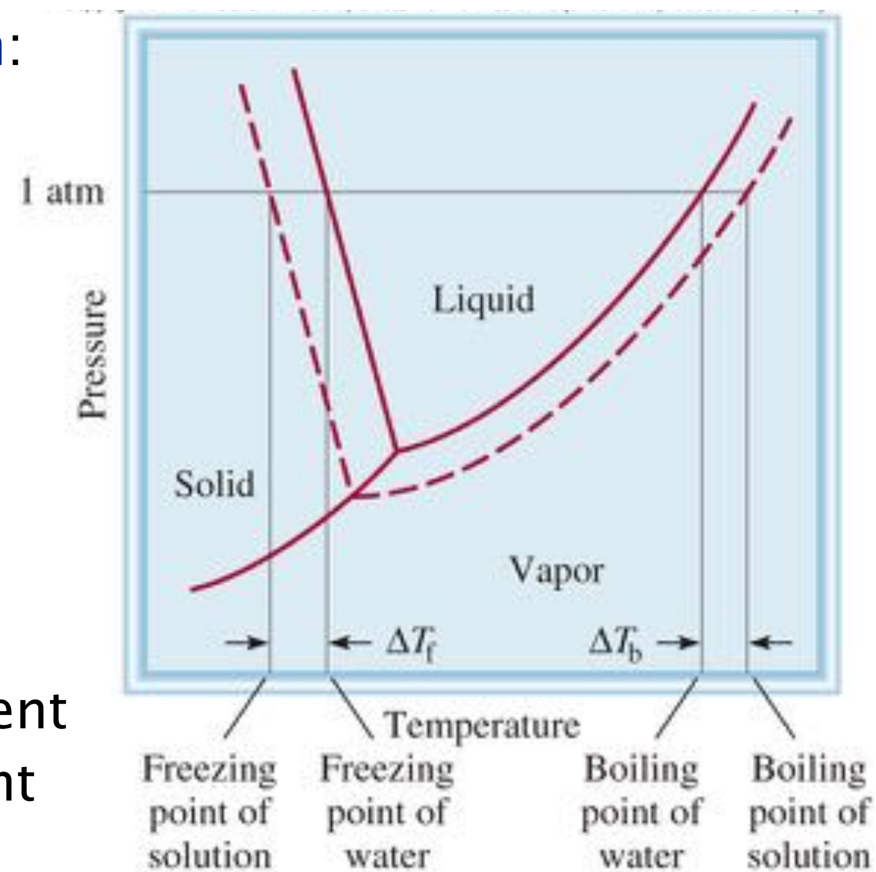
Vapor pressure above a solution is always less than vapor pressure above pure solvent.

1. Higher temperature needed for vapor pressure to hit 1 atm.
2. Boiling point of solution higher than boiling point of pure solvent
3. Boiling Point Elevation depends on:
   - Type of solvent
   - # of solute particles

$$\Delta T_b = iK_b m$$

$$T_b = T^\circ_b + \Delta T_b$$

m = solute molality
T_b = boiling point of solution
T°_b = boiling point of pure solvent
K_b = constant – based on solvent
i = van't Hoff factor



Freezing Point Depression

Solution freezes at lower temperature than pure solvent

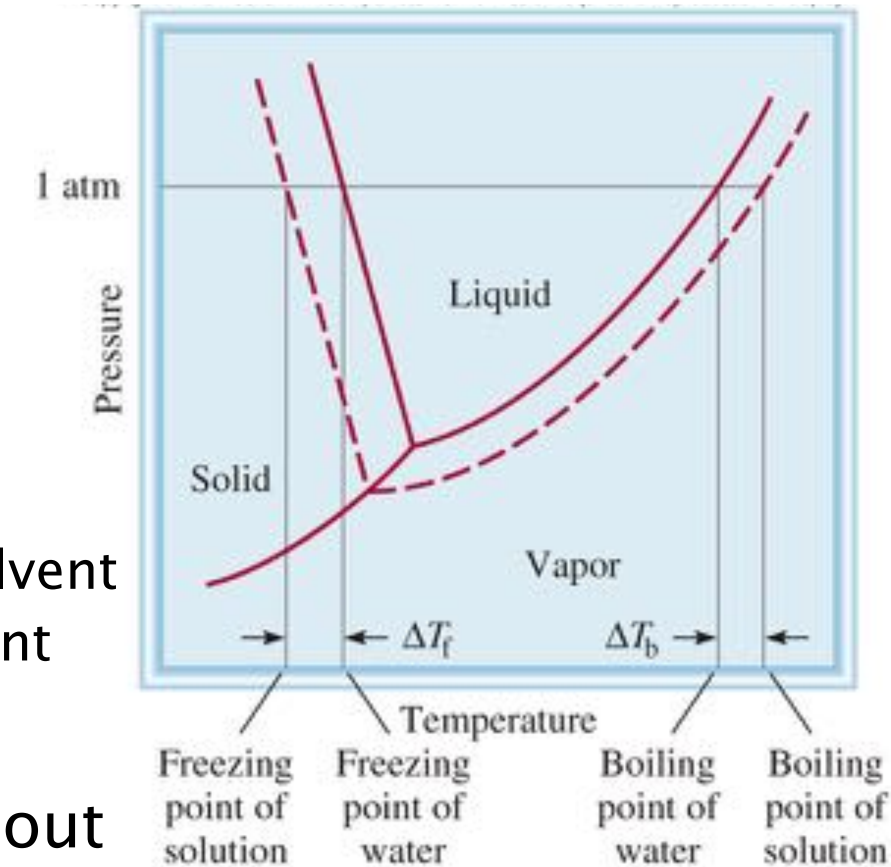
Freezing Point Depression depends on

- Type of solvent
- # of solute particles

$$\Delta T_f = iK_f m$$

$$T_f = T°_f - \Delta T_f$$

m = solute molality
T_f = freezing point of solution
$T°_f$ = freezing point of pure solvent
K_f = constant – based on solvent
i = van’t Hoff factor



- Only the pure solvent freezes out
- Pure substances:"sharp" melting point
- Impure materials: broad melting point range

What is the boiling point of a 0.886 m solution of pentane in benzene? For benzene: bpt = 80.1°C, K_b = 2.53°C/m

A: 82.3°C

What is the melting point of a 1.00 m solution of $CaCl_2$ in water? For water: mpt = 0.0°C, K_f = 1.86°C/m

A: - 5.6°C

A solution of 2.366 g solute in 82.10 g cyclohexane freezes at 2.65°C. Determine the molar mass of the solute. For cyclohexane, $K_f = 20.0°C/m$ and $T°_f = 6.55°C$.

Have freezing point of solution & pure solvent – can use $\Delta T_f = iK_f m$ to determine the molality of the solution

Use mass of solvent & molality to calculate moles solute

Given mass solute, so can calculate molar mass of solute

A: 148 g/mol

Constants

Will be provided on exams if needed

TABLE 13.2 Molal Boiling-Point Elevation and Freezing-Point Depression Constants of Several Common Liquids

Solvent	Normal Freezing Point (°C)*	K_f (°C/*m*)	Normal Boiling Point (°C)*	K_b (°C/*m*)
Water	0	1.86	100	0.52
Benzene	5.5	5.12	80.1	2.53
Ethanol	−117.3	1.99	78.4	1.22
Acetic acid	16.6	3.90	117.9	2.93
Cyclohexane	6.6	20.0	80.7	2.79

Osmotic Pressure

Semi-permeable membranes: Materials with tiny pores that only allow solvent molecules to pass

- solute cannot get through the membrane

Osmosis: Net flow of solvent molecules through a semi-permeable membrane from dilute to concentrated solution.

Osmotic Pressure (π): Pressure needed to stop osmotic flow

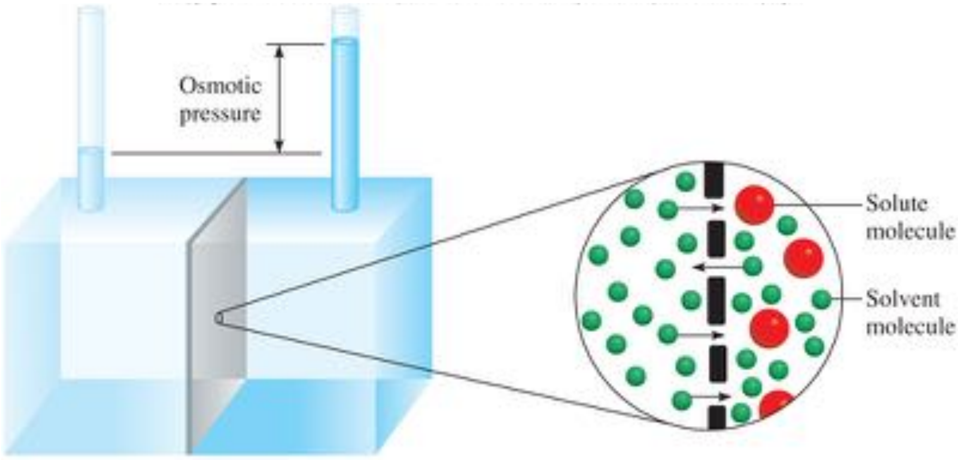
- π depends on:
  - concentration
  - temperature

π = iMRT

M= molarity

R= gas constant $\left(0.0821 \frac{\text{L atm}}{\text{mol K}}\right)$

T= temperature



(b)

What is the osmotic pressure of a 0.238 M aqueous solution at 25°C?

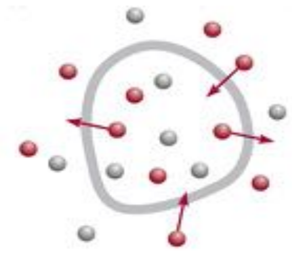
A: 5.83 atm

What is the molar mass of a solute if 397.2 g of the compound dissolved in water produces 592.4 mL of a solution that has an osmotic pressure of 1.98 atm at 35°C?

A: 8570 g/mol

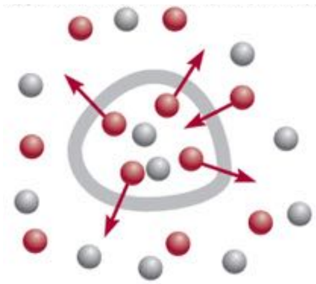
Practical Applications Of Osmosis

Isotonic Solution: Organ Transplants



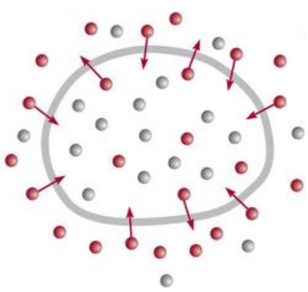
- Same concentration on both sides of membrane
- Organs would burst if stored in water

Hypertonic Solution: Food Preservation



- Higher concentration outside the membrane
- Salt pulls water from microbes and kills them

Hypotonic Solution: Tree Growth



- Lower concentration outside the membrane
- Water pulled through sap to top of trees

Reverse osmosis: Water Purification

- Reversing the net flow of solvent through a membrane by applying pressure greater than osmotic pressure.
- Obtain water with low solute concentration