

Ways of Expressing Concentrations of Solutions



Mole Fraction (X)

$$X_A = \frac{\text{moles of A}}{\text{total moles in solution}}$$

- In some applications, one needs the mole fraction of *solvent*, not solute—make sure you find the quantity you need!

$$\text{Mole fraction of component} = \frac{\text{moles of component in solution}}{\text{total moles of solution}}$$



Molarity (M)

$$\text{Molarity} = \frac{\text{moles solute}}{\text{liters of solution}}$$

- Because volume is temperature dependent, molarity can change with temperature.

Molality (m)

$$\text{Molality, } m = \frac{\text{moles solute}}{\text{kg of solvent}}$$

Because both moles and mass do not change with temperature, molality (unlike molarity) is *not* temperature dependent.



Parts Solute in Parts Solution

- Parts can be measured by mass or volume
- Parts are generally measured in same units
 - by mass in grams, kilogram, lbs, etc.
 - by volume in mL, L, gallons, etc.
 - mass and volume combined in grams and mL
- Percentage = parts of solute in every 100 parts solution
 - if a solution is 0.9% by mass, then there are 0.9 grams of solute in every 100 grams of solution
 - or 0.9 kg solute in every 100 kg solution
- Parts per million = parts of solute in every 1 million parts solution
 - if a solution is 36 ppm by volume, then there are 36 mL of solute in 1 million mL of solution



Percent Concentration

$$\text{Percent} = \frac{\text{Part (solute)}}{\text{Whole (solution)}} \times 100\%$$

$$\text{Mass Percent} = \frac{\text{Mass of Solute}}{\text{Mass of Solution}} \times 100\%$$

Mass of Solute + Mass of Solvent = Mass of Solution

$$\text{Percent Mass/Volume} = \frac{\text{Mass of Solute, g}}{\text{Volume of Solution, mL}} \times 100\%$$

Mass of Solute + Volume of Solvent \neq Volume of Solution

$$\text{Volume Percent} = \frac{\text{Volume of Solute}}{\text{Volume of Solution}} \times 100\%$$

Volume of Solute + Volume of Solvent \neq Volume of Solution



Parts Per Million Concentration

$$\text{PPM} = \frac{\text{Part (solute)}}{\text{Whole (solution)}} \times 10^6$$

$$\text{PPM by Mass} = \frac{\text{Mass of Solute}}{\text{Mass of Solution}} \times 10^6$$

Mass of Solute + Mass of Solvent = Mass of Solution

$$\text{PPM by Mass/Volume} = \frac{\text{Mass of Solute, g}}{\text{Volume of Solution, mL}} \times 10^6$$

Volume of Solvent \approx Volume of Solution for dilute solutions

$$\text{PPM by Volume} = \frac{\text{Volume of Solute}}{\text{Volume of Solution}} \times 10^6$$

Volume of Solvent \approx Volume of Solution for dilute solutions



PPM

- grams of solute per 1,000,000 g of solution
- mg of solute per 1 kg of solution
- 1 liter of water = 1 kg of water
 - for aqueous solutions we often approximate the kg of the solution as the kg or L of water
 - for dilute solutions, the difference in density between the solution and pure water is usually negligible

$$\text{PPM} = \frac{\text{Amount of Solute}}{\text{Amount of Solution}} \times 10^6$$

$$\text{PPM} = \frac{\text{mg of Solute}}{\text{kg of Solution}} = \frac{\text{mg of Solute}}{\text{L of Solution}}$$



Parts Per Billion Concentration

$$\text{PPB} = \frac{\text{Part (solute)}}{\text{Whole (solution)}} \times 10^9$$

$$\text{PPB by Mass} = \frac{\text{Mass of Solute}}{\text{Mass of Solution}} \times 10^9$$

Mass of Solute + Mass of Solvent = Mass of Solution

$$\text{PPB by Mass/Volume} = \frac{\text{Mass of Solute, g}}{\text{Volume of Solution, mL}} \times 10^9$$

Volume of Solvent \approx Volume of Solution for dilute solutions

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Using Concentrations as Conversion Factors

- Concentrations show the relationship between the amount of solute and the amount of solvent
 - 12%(m/m) sugar(*aq*) means 12 g sugar \equiv 100 g solution
 - or 12 kg sugar \equiv 100 kg solution; or 12 lbs. \equiv 100 lbs. solution
 - 5.5%(m/v) Ag in Hg means 5.5 g Ag \equiv 100 mL solution
 - 22%(v/v) alcohol(*aq*) means 22 mL EtOH \equiv 100 mL solution
- The concentration can then be used to convert the amount of solute into the amount of solution, or vice-versa



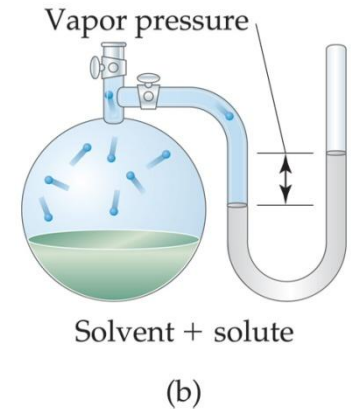
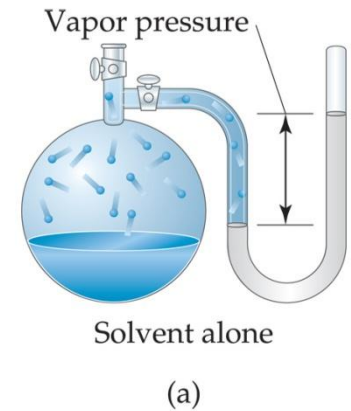
Colligative Properties

- Changes in colligative properties depend only on the *number* of solute particles present, not on the *kind* of the solute particles.
- colligative properties are
 - Vapor pressure
 - Boiling point elevation
 - Melting point depression
 - Osmotic pressure



Vapor Pressure

- Because of solute-solvent intermolecular attraction, higher concentrations of nonvolatile solutes make it harder for solvent to escape to the vapor phase.
- Therefore, the vapor pressure of a solution is lower than that of the pure solvent.



Raoult's Law

$$P_A = X_A P_A^\circ$$

where

- X_A is the mole fraction of compound A
- P_A° is the normal vapor pressure of pure solvent



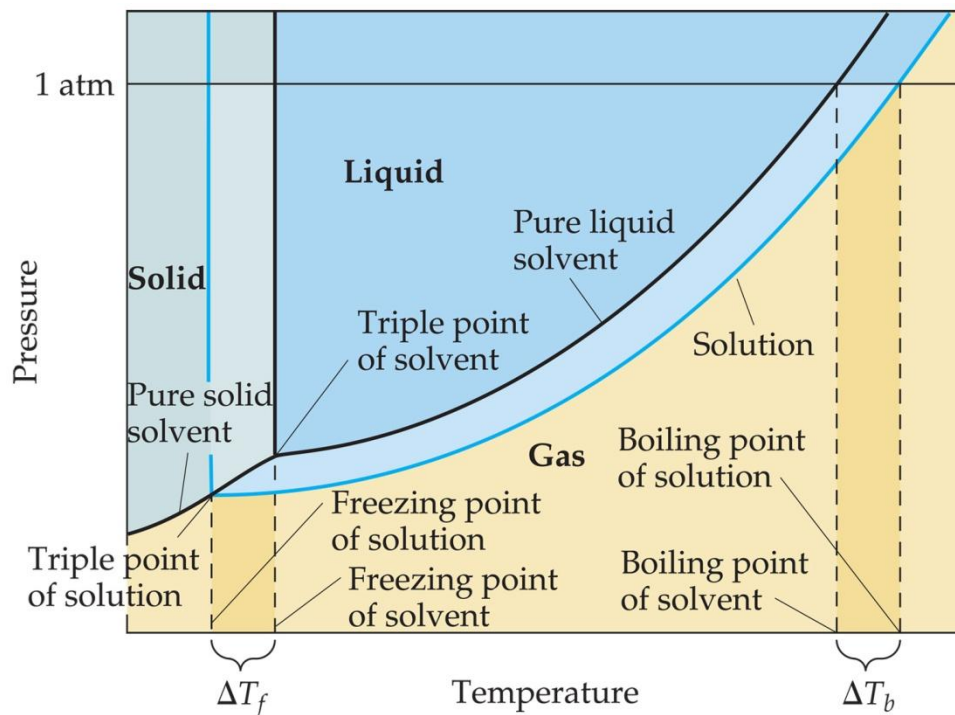
Raoult's Law - Example

At 27 °C the vapor pressure of benzene is 104 torr.

What is the vapor pressure of a solution that contains 0.100 mol of naphthalene in 9.90 mol of benzene?

Boiling Point Elevation and Freezing Point Depression

Nonvolatile solute-solvent interactions also cause solutions to have higher boiling points and lower freezing points than the pure solvent.



Boiling Point Elevation

The change in boiling point is proportional to the molality of the solution:

Solvent	Normal Boiling Point (°C)	K_b (°C/m)	Normal Freezing Point (°C)	K_f (°C/m)
Water, H ₂ O	100.0	0.51	0.0	1.86
Benzene, C ₆ H ₆	80.1	2.53	5.5	5.12
Ethanol, C ₂ H ₅ OH	78.4	1.2	-114.6	.99
Carbon tetrachloride, CCl ₄	76.8	5.02	-22.3	29.8
Chloroform, CHCl ₃	61.2	3.6	-63.5	.68

$$\Delta T_b = K_b \cdot m$$

ΔT_b is *added to* the normal boiling point of the solvent.

where K_b is the molal boiling point elevation constant, a property of the solvent.



Freezing Point Depression

- The change in freezing point can be found similarly:

$$\Delta T_f = K_f \cdot m$$

Solvent	Normal Boiling Point (°C)	K_b (°C/m)	Normal Freezing Point (°C)	K_f (°C/m)
Water, H ₂ O	100.0	0.51	0.0	1.86
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- Here K_f is the molal freezing point depression constant of the solvent.

ΔT_f is subtracted from the normal freezing point of the solvent.



Boiling Point Elevation and Freezing Point Depression

Note that in both equations, ΔT does not depend on *what the solute is*, but only on *how many particles are dissolved*.

$$\Delta T_b = K_b \cdot m$$

$$\Delta T_f = K_f \cdot m$$



De-icing of Airplanes is Based on Freezing-Point Depression

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De-icing Airplanes

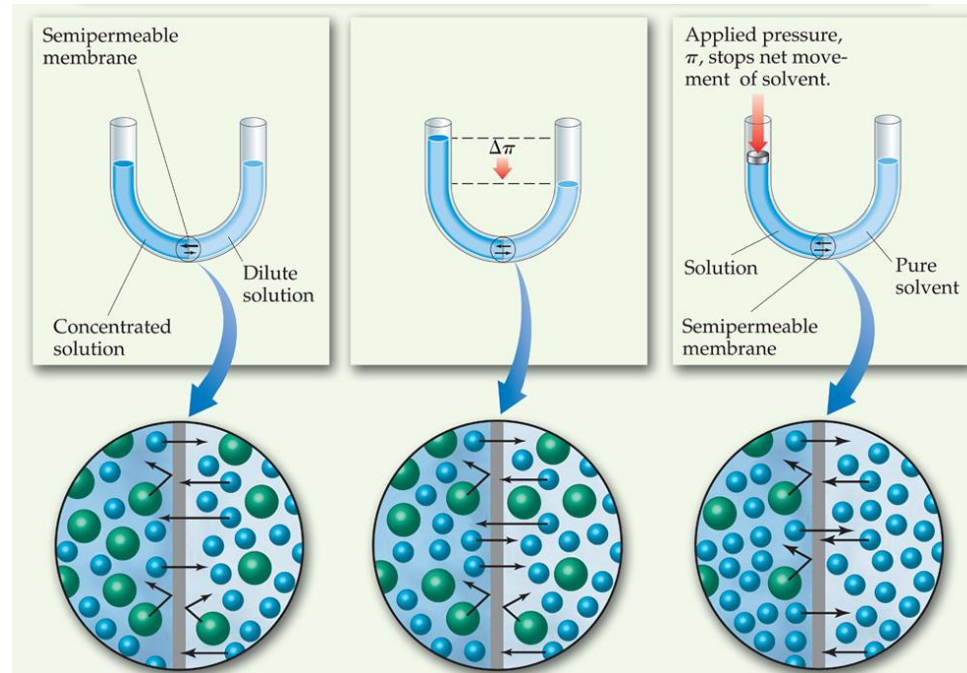


Osmosis

- Osmosis is the **diffusion** of water through a semi-permeable **membrane**
- **Some substances form semipermeable membranes, allowing some smaller particles to pass through, but blocking other larger particles.**
- In biological systems, most semipermeable membranes allow water to pass through, but solutes are not free to do so.



Osmosis



In osmosis, there is net movement of solvent from the area of **higher solvent concentration** (*lower solute concentration*) to the **lower solvent concentration** (*higher solute concentration*).



Osmotic Pressure

- The pressure required to stop osmosis, known as osmotic pressure, π , is

$$\pi = \left(\frac{n}{V} \right) RT = MRT$$

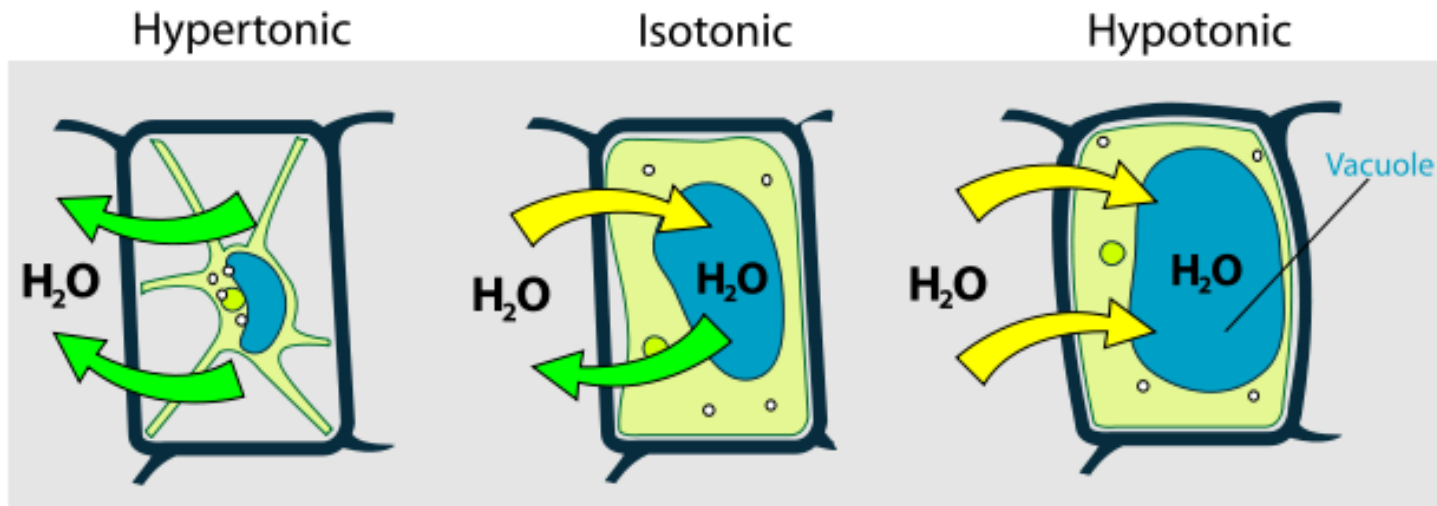
where M is the molarity of the solution

If the osmotic pressure is the same on both sides of a membrane (i.e., the concentrations are the same), the solutions are isotonic.



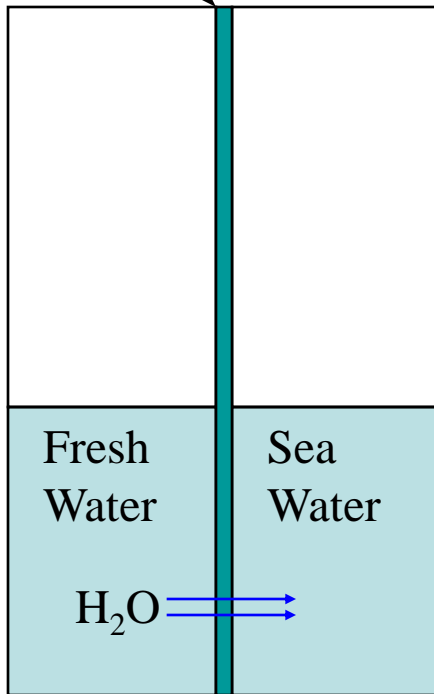
Osmosis in Blood Cells

- If the solute concentration outside the cell is greater than that inside the cell, the solution is hypertonic.
- If the solute concentration outside the cell is less than that inside the cell, the solution is hypotonic.

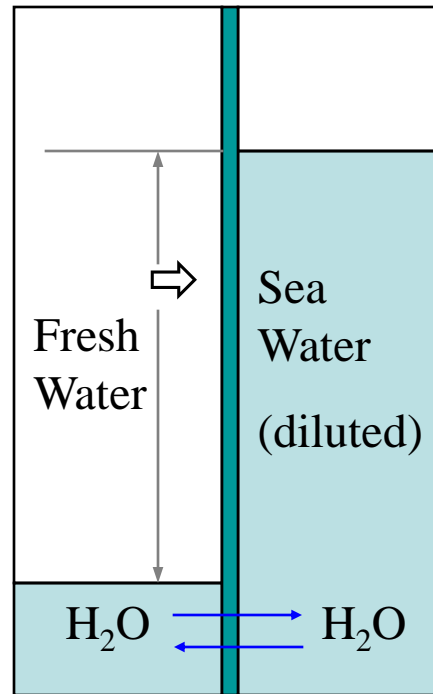


Theory of Osmosis

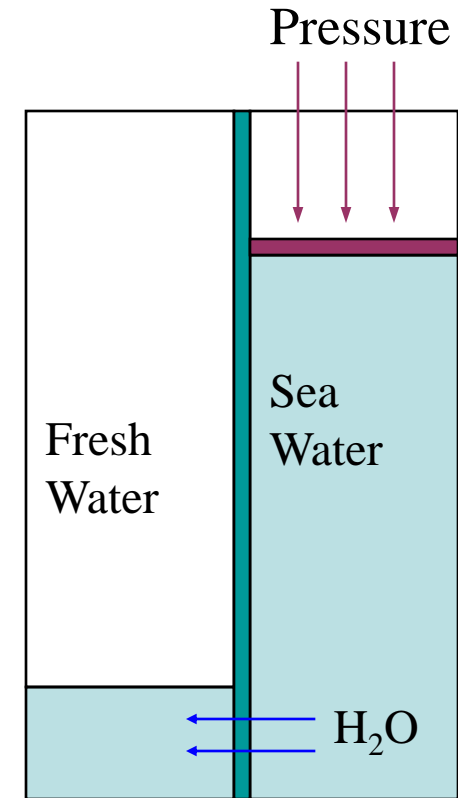
Semipermeable
Membrane



Initial Condition



Equilibrium



Reverse Osmosis

The Osmotic Pressure, π , is defined as: $\pi = MRT$
For sea water, π is about 35 psi.



$$\pi = cRT$$

where c is the molar concentration of the salt ions, $R = 0.082$ (liter·bar) / (deg·mol), is the gas constant, and $T = 300$ K is the ambient temperature on the absolute temperature scale (Kelvin).

The amount of salt in seawater is about 33 gram / liter. Seawater contains a variety of salts, but the calculation will be simplified by assuming that all the salt is sodium chloride (NaCl). The atomic weight of sodium is 23 gram, and of chlorine is 35.5 gram, so the molecular weight of NaCl is 58.5 gram. The number of NaCl moles in seawater is, therefore, $33 / 58.5 = 0.564$ mol / liter. (not familiar? see ["Chemical measures and units"](#)).

When NaCl salt dissolves in water it dissociates into Na^+ and Cl^- ions. There are two ions per salt molecule, so the ions' concentration is twice the molecules' concentration. $c = 2 \cdot 0.564 = 1.128$ mol / liter. Inserting the values into the van't Hoff formula yields the osmotic pressure:

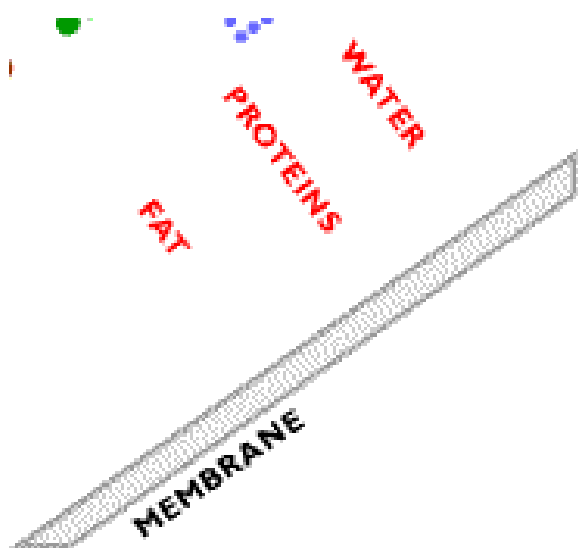
$$\pi = 1.128 \cdot 0.082 \cdot 300 = 27.8 \text{ bar}$$

or, 27.8 kilogram per square centimeter.

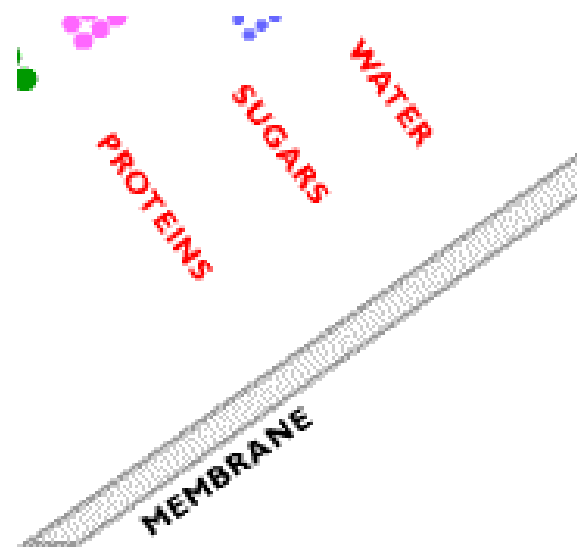


Water Treatment & Filtration Technologies

Micron Size	0.0001	0.001	0.01	0.1	1.0	10	100	1000
Examples	Metal Ions	Aqueous Salts	Colloids Viruses	Bacteria		Pollens	Beach Sand	
Filtration Tech						Particle Filtration		
					Microfiltration			
			Ultrafiltration					
		Nanofiltration						
	Hyperfiltration							



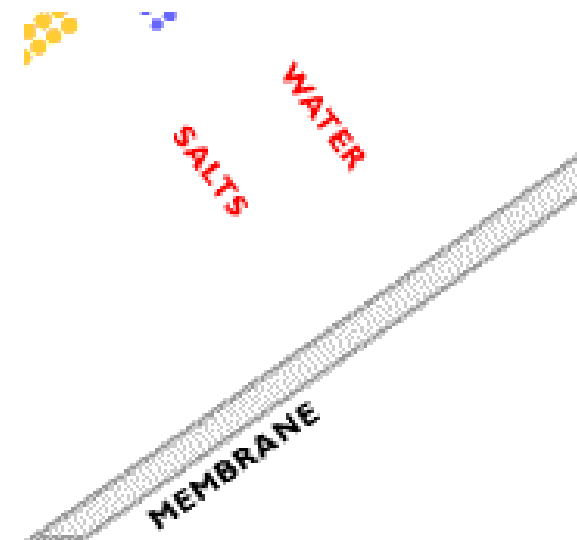
Microfiltration



Ultrafiltration



Nanofiltration



Reverse Osmosis



Colloids

- Suspensions of particles larger than individual ions or molecules, but too small to be settled out by gravity.
- Particle size: 10 to 2000 Å.

Phase of Colloid	Dispersing (solventlike) Substance	Dispersed (solutelike) Substance	Colloid Type	Example
Gas	Gas	Gas	—	None (all are solutions)
Gas	Gas	Liquid	Aerosol	Fog
Gas	Gas	Solid	Aerosol	Smoke
Liquid	Liquid	Gas	Foam	Whipped cream
Liquid	Liquid	Liquid	Emulsion	Milk
Liquid	Liquid	Solid	Sol	Paint
Solid	Solid	Gas	Solid foam	Marshmallow
Solid	Solid	Liquid	Solid emulsion	Butter
Solid	Solid	Solid	Solid sol	Ruby glass



Removal of Colloidal Particles

- Sodium stearate is one example of such a molecule.
- Removal process coagulation-filtration
- Colloid particles are too small to be separated by physical means (e.g. filtration).
- Colloid particles are coagulated (enlarged) until they can be removed by filtration.
- **Methods of coagulation:**
 - heating
 - adding an electrolyte

