Different ways to describe concentration

All of them are essentially

$$\frac{\text{Amount of solute}}{\text{Amount of everything (solvent)}}$$

Mole Fraction $$\chi_i = \frac{\text{moles of i}}{\text{total moles}}$$

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

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

Demo

What is the key effect of adding the salt to the water?

A. the salt dissolving is endothermic so the temperature drops

B. the salt dissolving is exothermic so it melts the ice

C. the salt dissolving increases the entropy of the solution

D. the salt dissolving increases the entropy of the solid ice

Why does the temperature drop?

A. the salt dissolving requires energy (endothermic)

B. the salt dissolving releases energy (exothermic)

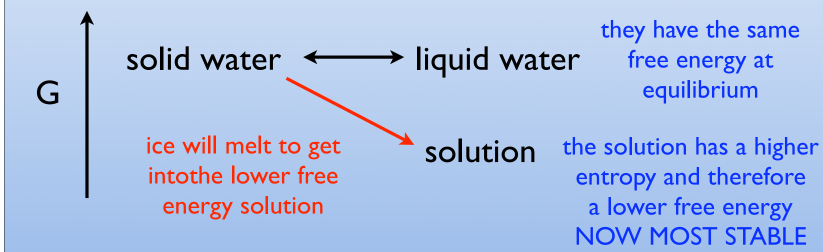
C. the ice melting releases energy (exothermic)

D. the ice melting requires energy (endothermic)

Solutions

The main effect of making a solution is that the entropy of the solution is higher than the separate solvent and solute

T = 0°C and P = 1 atm



This effect depends on the entropy of the solution which depends on how much "stuff" is dissolved but not what the "stuff" is

Colligative Properties
depend on the concentration of the solution but not what is actually dissolved
(note: this is approximate as it assumes and ideal solution)

The only thing that matters is the number of moles of "stuff"

Somethings dissolve into ions making more moles

1 M sugar solution = 1 moles of sugar in 1 L solution

1 M NaCl solution = 1 moles of Na^+ in 1 L solution
1 mole of Cl^- in 1 L solution

2 moles of "stuff"

Van't Hoff Number

$$i = \frac{\text{moles of "particles" in solution}}{\text{moles of solute dissovled}}$$

TABLE 17.6 Expected and Observed Values of the van't Hoff Factor for 0.05 *m* Solutions of Several Electrolytes

Electrolyte	*i* (expected)	*i* (observed)
NaCl	2.0	1.9
$MgCl_2$	3.0	2.7
$MgSO_4$	2.0	1.3
$FeCl_3$	4.0	3.4
HCl	2.0	1.9
Glucose*	1.0	1.0

*A nonelectrolyte shown for comparison.

Effect of making the solution

Boiling Point Elevation

Solution now more stable than vapor. Therefore the boiling point goes up

Freezing Point Depression

Solution now more stable than solid. Therefore the freezing point goes down

constant that depends on solvent

Boiling Point Elevation $\Delta T = K_b m_{solute}$

molality solute

Remember the number of particles is what matters

$$\Delta T = iK_b m_{solute}$$

constant that depends on solvent

Freezing Point Depression $\Delta T = -K_f m_{solute}$

molality solute

Remember the number of particles is what matters

$$\Delta T = -iK_f m_{solute}$$

TABLE 17.5 Molal Boiling-Point Elevation Constants (K_b) and Freezing-Point Depression Constants (K_f) for Several Solvents

Solvent	Boiling Point (°C)	K_b (°C kg/mol)	Freezing Point (°C)	K_f (°C kg/mol)
Water (H_2O)	100.0	0.51	0.	1.86
Carbon tetrachloride (CCl_4)	76.5	5.03	−22.99	30.
Chloroform ($CHCl_3$)	61.2	3.63	−63.5	4.70
Benzene (C_6H_6)	80.1	2.53	5.5	5.12
Carbon disulfide (CS_2)	46.2	2.34	−111.5	3.83
Ethyl ether ($C_4H_{10}O$)	34.5	2.02	−116.2	1.79
Camphor ($C_{10}H_{16}O$)	208.0	5.95	179.8	40.

Which would you expect to have the lowest freezing point

A. 2 M sugar solution

B. 0.5 M NaCl solution

C. 1 M NaCl solution

D. 1 M $MgCl_2$ solution ←

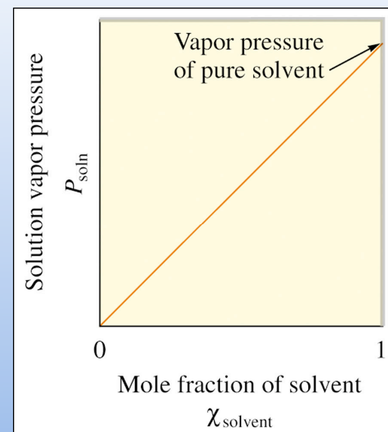
If the boiling point is higher, what is the vapor pressure of the solution?

A. higher than the pure solvent

B. lower than the pure solvent ←

C. the same as the pure solvent

Raoult's Law

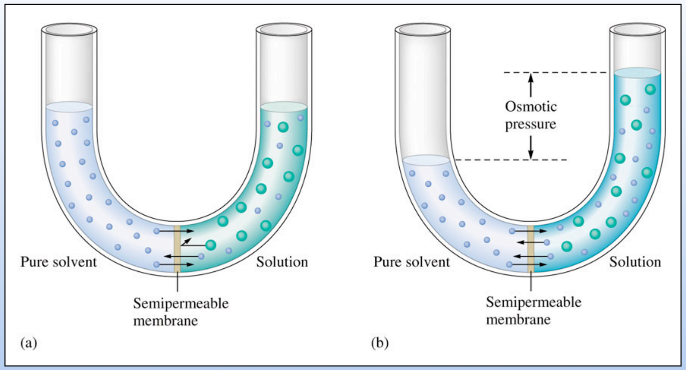


vapor pressure of pure solvent

$$P_{solvent} = X_{solvent}P°$$

mole fraction of solvent!

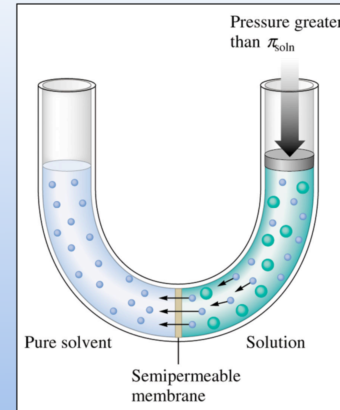
Osmosis



Solvent can pass through the membrane but the solute can't

Solution is lower in free energy so pure solvent moves to the solution side

Reverse Osmosis



Osmotic Pressure

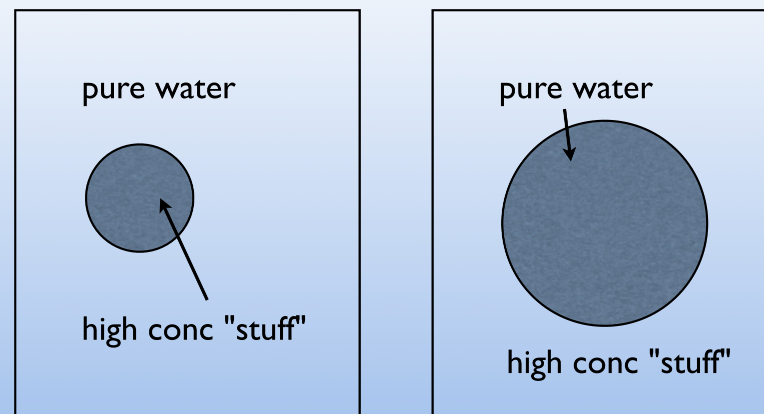
$$\Pi = MRT$$

Osmotic pressure → Π

Molarity of solution! → M

$$\Pi = iMRT$$

Cells



Examples

Solution of 100 g of sugar (sucrose MW 342 g mol^{-1}) in 1 L of water.

(100 g)/(342 g mol^{-1}) = 0.292 mol sugar

1 L water is approx. 1 kg

(1000 g)/(18 g mol^{-1}) = 55.6 moles

Mole fraction sugar of solution

χ_{sugar} = (0.292 mol)/(0.292 + 55.6) = .00522

Mole fraction water of solution

χ_{water} = (55.6 mol)/(0.292 + 55.6) = 0.995 (or 1- χ_{sugar})

Molality

m = (.292 mol)/(1 kg) = 0.292 mol kg^{-1}

Molarity

M = (.292 mol)/(1 L) = 0.292 mol L^{-1}

Freezing point depressions (given K_f for water is 1.86)

$\Delta T = -iK_f m$= -(1)(1.86)(.292) = -0.543 °C

Boiling point elevation (given K_b for water is 0.51)

$\Delta T = -iK_b m$= -(1)(0.51)(.292) = +0.15 °C

Osmotic Pressure (at 25°C)

Π = MRT= (1 mol L^{-1})(0.08206 L atm K^{-1} mol^{-1})(298.15 K) = 24.5 atm

Vapor Pressure (given pure vapor pressure of water at 25°C is 23.76 Torr)

$P_{H2O} = \chi_{H2O}P°$ = (.995)(23.76) = 23.64 Torr