Chapter 12. Properties of Solutions

What we will learn:

- Types of solutions
- Solution process
- Interactions in solution
- Types of concentration
- Concentration units
- Solubility and temperature
- Solubility and pressure
- Colloids

Solution

• A homogeneous mixture of two or more substances

Saturated solution

• Contains the maximum amount of a solute that will dissolve in a solvent

Supersaturated solution

• Contains more amount of a solute that will dissolve in a solvent (an unstabile state)

Unsaturated solution

• Contains less amount of a solute that will dissolve in a solvent

Crystalization

• The process in which a dissolved solute comes out of solution and forms cristals

Types of solutions

Comp. 1	Comp. 2	State	Examples
Gas	Gas	Gas	Air
Gas	Liquid	Liquid	Soda water
Gas	Solid	Solid	H_{2} in palladium
Liquid	Liquid	Liquid	Ethanol in water
Solid	Liquid	Liquid	NaCl in water
Solid	Solid	Solid	Brass (Cu/Zn)

Molecular view of solution process

During the solution process, solvent molecules interact with solute molecules. If the solvent-solute interactions are stronger than the solvent-solvent and solute-solute interactions, then the solute is dissolved in the solvent. (**Energy** driven process)

Another driving force of the solution process is an increase of dissorder on the final system. If the dissorder of the solution is bigger then dissorder of the solvent and dissorder of the solute, then the solute is dissolved in the solvent. (**Entropy** driven process)

Three steps of solution process (energy driven)

- Heat of separation of solute molecules ΔH_1
- Heat of separation of solvent molecules
- Heat of mixing of solute and solvent molecules

$$\Delta H_1$$

 ΔH_2

 ΔH_{z}

Heat of solution

$$\Delta H_{sol} = \Delta H_1 + \Delta H_2 + \Delta H_3$$

Exothermic

(a favorable process)

$$\Delta H_{sol} < 0$$

Endothermic (a non favorable process)

$$\Delta H_{sol} > 0$$

Example

The endothermic process - solution of sodium chloride in water



Solubility

• Measure of how much a solute will dissolve in a solvent at a particular temperature

Like dissolves like

• A solute is better dissolved in a solvent, which has similar intermolecular force types and magnites to a solute

lonic molecules are better dissolved in polar solvents

Nonpolar molecules are better dissolved in nonpolar solvents

Examples

- H₂O is a good solvent for NaCl
- CCI_4 is a good solvent for C_6H_6



Hydrogen bondings between water and methanol are responsible for solubility of methanol in water



Water molecules are stabilizing Na⁺ and Cl⁻ ions in NaCl solution



Dispersion interactions between benzene and CCl_4 are responsible for solubility of CCl_4 in benzene

Miscible liquids

• Two liquids which are completely soluble in each other in all proportions

Example

- Methanol and water
- Ethanol and water
- 1,2-ethylene glycol and water





Solvation

• The process in which a solute molecule (ion) is surrounded by solvent molecules





Problem

Predict the relative solubility in the following cases:

a) Bromine (Br_2) in benzene (C_6H_6) and in water b) KCI in carbon tetrachoride (CCI_4) and in liquid ammonia (NH_3) c) Formaldehyde (CH_2O) in carbon disulfide (CS_2) and in water



Concentration units

- **Percent by mass** (units: %), is independent of temperature and does not require calculations of molar masses
- **Mole fraction** (no units); is used in calculating partial pressures of gases
- **Molarity** (units: mol/L); used practically in laboratories because a volume of a solution is easy to measure, but is dependent of temperature
- Molality (units: mol/kg); is independent of temperature because the concentration is expressed in solute number of moles and mass of solvent

$$= \frac{mass of solute (m_{s'})}{mass of solute (m_{s'}) + mass of solvent (m_{s''})} \times 100\% \quad (units \%)$$

Example

What is the percent by mass of 0.892 g KCI dissolved in 54.6 g water.

Data

```
m_{sl} - mass of solute = 0.892 g
m_{sv} - mass of solvent = 54.6 g
```

Solution

percent by mass = 0.892 g / (0.892 g + 54.6 g) X 100 %

= 1.61 %

Mole fraction

moles of solute
$$(n_{sl})$$

X =

(no units)

moles of solute (n_{sv}) + moles of solvent (n_{sv})

Example

What is the mole fraction of 12.5 g sodium chloride (NaCl) dissolved in 80.5 g water.

Data

$$m_{sl}$$
 - mass of solute = 12.5 g
 m_{sv} - mass of solvent = 80.5 g

Solution

Step 1. Calculation of moles:

n_{sv} = 4.46 mol

Number of water moles:

Group Period	1	2		3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H 1.008																		2 He 4.0026
2	3 Li 6.94	4 Be 9.0122												5 B 10.81	6 C 12.011	7 N 14.007	8 O 15.999	9 F 18.998	10 Ne 20.180
3	11 Na 22.990	12 Mg 24.305												13 Al 26.982	14 Si 28.085	15 P 30.974	16 S 32.06	17 CI 35.45	18 Ar 39.948
4	19 K 39.098	20 Ca 40.078		21 Sc 44.956	22 Ti 47.867	23 V 50.942	24 Cr 51.996	25 Mn 54.938	26 Fe 55.845	27 Co 58.933	28 Ni 58.693	29 Cu 63.546	30 Zn 65.38	31 Ga 69.723	32 Ge 72.63	33 As 74.922	34 Se 78.96	35 Br 79.904	36 Kr 83.798
5	37 Rb 85.468	38 Sr 87.62		39 Y 88.906	40 Zr 91.224	41 Nb 92.906	42 Mo 95.96	43 Tc [97.91]	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 126.90	54 Xe 131.29
6	55 Cs 132.91	56 Ba 137.33	*	71 Lu 174.97	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 OS 190.23	77 ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 1 204.38	82 Pb 207.2	83 Bi 208.98	84 Po [208.98]	85 At [209.99]	86 Rn [222.02]
7	87 Fr [223.02]	88 Ra [226.03]	**	103 Lr [262.11]	104 Rf [265.12]	105 Db [268.13]	106 Sg [271.13]	107 Bh [270]	108 HS [277.15]	109 Mt [276.15]	110 DS [281.16]	111 Rg [280.16]	112 Cn [285.17]	113 Uut [284.18]	114 FI [289.19]	115 Uup [288.19]	116 LV [293]	117 Uus [294]	118 Uuo [294]
				57	58	59	60	61	62	63	64	65	66	67	68	69	70		
*Lai	nthanoi	ds	*	La 138.91	Ce 140.12	Pr 140.91	Nd 144.24	Pm [144.91]	Sm 150.36	EU 151.96	Gd 157.25	Tb 158.93	Dy 162.50	HO 164.93	Er 167.26	T M 168.93	Yb 173.05		
**/	Actinoid	S	**	89 Ac [227.03]	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np [237.05]	94 Pu [244.06]	95 Am [243.06]	96 Cm [247.07]	97 Bk [247.07]	98 Cf [251.08]	99 ES [252.08]	100 Fm [257.10]	101 Md [258.10]	102 No [259.10]		

Step 2. Calculation of moles

n_{si} = 0.21 mol

Number of NaCl moles:



n = 1 mol x 12.5 g / 58.44 g = 0.21 mol

Step 3. Calculation of mole fraction

mole fraction = 0.21 mol / (0.21 mol + 4.46 mol)

= 0.04

Molarity

moles of solute (n_{sl})

M =

volume of solution (V_{sn})

(units: mol / L)

Example

Calculate molarity of a 1.2 L of HNO₃ solution containing 24.4 g of HNO₃.

Data

```
m_{sl} - mass of solute = 24.4 g
V<sub>sn</sub> - volume of solution = 1.2 L
```

Solution

Step 1. Calculation of moles

n_{si} = 0.38 mol

Step 2. Calculation of molarity

molarity = 0.38 mol / 1.2 L

= 0.31 mol/L

Number of HNO₃ moles

n = 1 mol x 24.4 g / 63.02 g = 0.38 mol

Molality

=

moles of solute
$$(n_{st})$$

m

mass of solvent (m_{sv})

(units: mol / kg)

Example

Calculate the molality of H_2SO_4 solution, containing 24.4 g of H_2SO_4 in 198 g of water.

Data

```
m<sub>sl</sub> - mass of solute = 24.4 g
m<sub>sv</sub> - mass of solvent = 198 g = 0.98 kg
```

Solution

Step 1. Calculation of moles:

n_{sl} = 0.24 mol

Step 2. Calculation of molality:

n_{si} = 0.24 mol /0.98 kg

= 1.26 mol / kg

Number of H_2SO_4 moles

n = 1 mol x 24.4 g / 98.10 g = 0.24 mol

Density of solution

mass of solution (m_{sn})

d =

volume of solution (V_{sn})

(units: g / mL)

Examples

$$m_{sn}$$
 - mass of solution = 128 g
V_{sn} - volume of solution = 1.45 L = 1450 mL

d = 128 g / 1450 mL = 0.08 g/mL

Problem

The molarity if a aqueous solution of methanol is 2.45 M. The density of this solution is 0.976 g/mL. What is the molality of the solution.

=

Data

$$\rm d_{sn}~$$
 - density of solution = 0.976 g/mL $\rm M_{sn}~$ - molarity of solution = 2.45 M

Solution

 $M = \frac{\text{moles of solute } (n_{sl})}{\text{volume of solution } (V_{sn})} d$

moles of solute (n_{sl})

m

=

mass of solvent (m_{sv})

mass of solution (m_{sn})

volume of solution (V_{sn})

Step 1. Calculation of mass of solute

2.45 moles - 1 liter of solution

79.49 g - 1 liter of solution

Mass of CH₃OH:

n = 2.45 mol x 32.04 g / 1mol = 79.49 g

Step 2. Calculation of mass of water

0.976g of solution - 1 mliter of solution

976 g of solution - 1 liter of solution

mass of water = 976 g - 76.49 g = 898 g **Step 3.** Calculation of molality:

- m = 2.45 mol / 0.898 kg
 - = 2.73 *m*

Problem

The density of an aqueous solution containing 10 % of ethanol (C_2H_2OH) by mass, is 0.984 g/mL. Calculate the molality of this solution. Calculate the molarity

Data

$$d_{sn}$$
 - density of solution = 0.984 g/mL

P - percent by mass of solution = 10%

 $M = \frac{\text{moles of solute } (n_{sl})}{\text{volume of solution } (V_{sn})} \qquad d = \frac{\text{mass of solution } (m_{sn})}{\text{volume of solution } (V_{sn})}$ $m = \frac{\text{moles of solute } (n_{sl})}{\text{moles of solute } (n_{sl})}$

mass of solvent (m_{sv})

Step 1. Calculation of mass of water in 100 g of solution

10g (E) - 100g (S) m (W) = 100 g (S) - 10 g(E)

= 90 g = 0.09 kg

Step 2. Calculation of ethanol moles in 100 g of solution

10g(E)(1 mol) / (46.07 g) = 0.217 mol(E)

Step 3. Calculation of molality of 100 g of solution

- m = (0.217 mol) / (0.09 kg)
 - = 2.41 mol/kg

Moles of ethanol:



n = 10 g x 1 mol / 46.07 g = 0.217 mol **Step 4.** Calculation of a volume of 100 g of solution

100 g (1 mL) / (0.984 g) = 102 mL = 0.102 L

Volume of solution:



V = 100 g x 1 mL / 0.984 g = 102 mL

Step 5. Calculation molarity of 100 g of solution

M = (0.217 mol) / (0.102 L) = 2.13 mol/L

Effect of temperature on solubility

- Solubility of solid solutes
- Solubility of gas solutes

Solid solubility

Generally, the solubility of solids is bigger at the higher temperature

There is no correlation between the solubility of a solute at a particular temperature, and its heat of solution ΔH_{sol}

CaCl₂ exothermic
 NH₄NO₃ endothermic

Solubility of both compounds increases with the temperature

Solubility

• Maximal concentration of a solute in a solvent



Fractional crystalization

• Separation of a solution into pure solutes using their different solubilities

By lowering temperature we can extract from a solution a pure solute.

If we have a solution of a mixture of two solutes, by using fractional crystalization we can separate those two solutes.



Gas solubility

Generally, the solubility of gasses is smaller at the higher temperatures

Fish prefer to be deeper in water, where the temperature is lower, and the oxygen concentration is bigger

Global warming decreses the concentration of oxygen in water, which is harmful to its living animals



Effect of pressure on gas solubility

Henry's law

• The solubility of a gas in a liquid is proportional to the pressure of the gas over the liquid

$$c = kP$$

- c gas concentration (molarity)
- k constant
- P pressure (partial)

Dynamic equilibration

• A number of molecules going from the gas to the liquid is the same as the number of molecules going from the liquid to the gas at the same time



Practically, in soft drinks, the pressure of CO_2 in a bottle is big which makes bigger the concentration CO_2 in the liquid

Problem

The solubility of pure nitrogen gas in water at 25 C and at 1 atm, is 6.8 x 10^{-4} mol/L. What is the concentration (in molarity) of nitrogen dissolved in water under atmospheric conditions? The partial pressure of nitrogen gas in the atmosphere is 0.78 atm.

Data

 c_{N2} - solubility of nitrogen gas = 6.8 x 10⁻⁴ mol/L P_{N2} - partial pressure of nitrogen gas = 0.78 atm P - atmospheric pressure = 1 atm

$$c = kP$$
 $k = c/P$

 $k = 6.8 \times 10^{-4} \text{ mol/L atm}$

- $c = (6.8 \times 10^{-4} \text{ mol/L atm}) (0.78 \text{ atm})$
 - = 5.3 x 10⁻⁴ mol/L

Colligative (collective) properties of nonelectrolyte solution

• Properties which depend on the number of solute particles in solution and not on the nature of solute

Vapor-pressure lowering

• The vapor pressure of its solution is always smaller than that of the pure solvent

For the solute, which is nonvolatile (it does not have a vapor pressure)



Raoult's law

 $P_{1} = X_{1}P_{1}^{0}$

 P_1 - vapor pressure of a solvent over a solution P_1^{0} - vapor pressure of the pure solvent X_1 - mole fraction of the solvent

 $X_1 = 1 - X_2$

 X_2 - mole fraction of the solute

Problem

Calculate the vapor pressure of a solution containing 218 g glucose in 460 mL of water at 30 C. The vapor pressure of pure water at 30 C is 31.82 mmHg. The density of the solution is 1 g/mL.

Data

 P_1^{0} - vapor pressure of pure water = 31.82 mmHg

$$m_2 - mass of the solute = 218 gV_1 - volume of water = 460 mLd_1 - density of water = 1 g / mL$$

 m_1 - mass of water = 460 g

Step 1. Calculate the number of moles

$$n_1 = 25.5 \text{ mol} n_2 = 1.21 \text{ mol}$$

Moles of water:

$$1 \text{ mol} = 18.02 \text{ g}$$

 $n_1 \text{ mol} = 460 \text{ g}$

n₁ = 1 mol x 460 g/18.02 g = 25.5 mol

Moles of glucose:



n₂ = 1 mol x 218 g/180.2 g = 1.21 mol Step 2. Calculation of the water mole fraction

$$X_1 = 25.5 \text{ mol} / (25.5 \text{ mol} + 1.21 \text{ mol})$$

 $X_1 = 0.955$

Step 3. Calculation of the vapor pressure

$$P_1 = X_1 P_1^0$$

- $P_1 = 0.955 \text{ x } 31.82 \text{ mmHg}$
 - = 30.4 mmHg

Raoult's law (a more practical equation)

$$\Delta P = X_2 P_1^{o}$$

 ΔP - vapor pressure lowering over a solution P_1^{0} - vapor pressure of the pure solvent X_2 - mole fraction of the solute

Problem

At 25 C the vapor pressure of pure water is 23.76 mmHg and that of seawater is 22.98 mmHg. Assuming that seawater contains only NaCl, estimate its molal concentration (molality).

$$\Delta P = X_2 P_1^{0}$$

 P_1^{0} - vapor pressure of pure water = 23.76 mmHg P_1^{0} - vapor pressure of solution = 22.98 mmHg

$$\Delta P = X_2 P_1^{0}$$

$$X_2 = \Delta P / P_1^0$$

= (23.76 mmHg - 22.98 mmHg) / (23.76 mmHg)

= 0.033

assuming that we have 1000 g of water as the solvent

$$X_2 = n_2 / (n_1 + n_2)$$

$$n_2 / (55.49 + n_2) = 0.033$$

 $n_2 = 1.88 \text{ mol}$

Since NaCl dissociates to form two ions, the number of moles of NaCl is half of the above results

Moles NaCl = 0.94 mol

The molality of the solution

m = (0.94 mol) / (1 kg) = 0.94 mol/kg

Vapor pressure over a solution of two liquids

Raoult's law

$$P_A = X_A P_A^{\ 0}$$

 $P_{B} = X_{B}P_{B}^{0}$

 P_A^{-} - vapor pressure of a liquid A over a solution P_A^{-0} - vapor pressure of the pure liquid A P_B^{-} - vapor pressure of a liquid B over a solution P_B^{-0} - vapor pressure of the pure liquid B X_A^{-} - mole fraction of the liquid A X_B^{-} - mole fraction of the liquid B

$$P_{T} = P_{A} + P_{B}$$

 $\mathsf{P}_{_{\!\mathrm{T}}}$ - total vapor pressure over the solution

Two volatile liquids (both having a vapor pressure)

Example

Benzene and toluene

A - benzene

B - toluene

Case 1. \triangle H of solution is zero (ideal solution)

The forces between A and B molecules are the same as the forces between A and A, and between B and B.

No termal effects of the solution process

The Raoult's law is fulfilled

Example

A mixture of benzene and toluene



Case 2. Δ H of solution bigger than zero (endothermic process)

The forces between A and B molecules are smaller than the forces between A and A, and between B and B. Molecules of both solutes have a tendency to leave the solution and go to the gas phase. The total vapor pressure is bigger than the sum of partial vapor pressures.

The solution process requires the energy

The positive deviation of the Raoult's law

Example:

The system of chloroform (CHCl₃) and ethanol (CH₃CH₂OH)



Case 3. Δ H of solution smaller than zero (exothermic process)

The forces between A and B molecules are bigger than the forces between A and A, and between B and B. Molecules of both solutes have a tendency to be in the solution rather than go to the gas phase. The total vapor pressure is smaller than the sum of partial vapor pressures.

The solution process generates the energy

The negative deviation of the Raoult's law

Example:

The system of chloroform $(CHCl_3)$ and acetone (CH_3COCH_3)



Fractional distillation

• Separating a solution of two liquids into individual components, based on their different boiling points.

Practically used in chemical laboratories and industry (oil fractional distillation in petroleum refineries)

Boiling point

• The temperature at which the vapour pressure of a solution is equal to the external amospheric pressure

A solid solute usually lowers the vapor pressure of its solution, therefore the boiling point of the solution is bigger than the boiling point of the pure solvent

Boiling point elevation

• This is a difference between the boiling point of the solution and the boiling point of the pure solvent



$$\Delta T_{b} = T_{b} - T_{b}^{0}$$

- T_{b} boiling point of the solution
- T_{b}^{0} boiling point of the pure solvent

The boiling point of the solution depends on the solution concentration (molality)

$$\Delta T_{b} = K_{b} m$$

- $\rm K_{\rm \scriptscriptstyle b}~$ molar boiling-point elevation constant
- m molality concentration

Freezing point

• The temperature at which a liquid solution transforms into a solid

When the solution freezes, its separates into individual solvent components

Freezing point depression

• This is a difference between the freezing point of the solution and the freezing point of the pure solvent



$$\Delta T_f = T_f^o - T_f$$

- T_{f} freezing point of the solution
- T_{f}^{0} freezing point of the pure solvent

The freezing point of the solution depends on the solution concentration (molality)

$$\Delta T_{f} = K_{f} m$$

- K_r molar freezing-point depression constant
- *m* molality concentration

Problem

Calculate the freezing point of a solution containing 651 g of ethylene glycol in 2505 g of water. The molal freezing-point depression constant of water is 1.86 C/m.

Data

$$m_{su}$$
 - mass of solute = 651 g
 m_{sl} - mass of solvent = 2505 g
 K_{f} - molal freezing-point depression constant of solvent = 1.86 C/m

Solution

 $\Delta T_{f} = K_{f} m$

Moles of ethylene glycol:

n = 651 g x 1 mol / 62.01 g = 10.5 mol

- *m* = (moles of solute) / (mass of solvent)
- m = (10.5 mol) / (2.505 kg)
- m = 4.19 m
- $\Delta T_{f} = (1.86 \text{ C/m}) (4.19 \text{ m})$
 - = 7.79 C

Osmosis

• An effect of moving solvent molecules through a membrane from a less to a more concentrated solution

Solvent molecules are moving from a smaller dissorder state in a less concentrated solution, to a bigger dissorder state in a more concentrated solution

Entropy driven process



Osmotic pressure

• A pressure which stops the osmotic process



$$\pi = MRT$$

SOLVENT SOLUTION

- π osmotic pressure
- M molarity of the solution
- R constant (gas constant = 0.0821 L atm/K mol)
- T temperature

The osmotic pressure is responsible for moving water from tree roots to tree leafs

Problem

The osmotic pressure between two solutions of seawater is 30.0 atm. Calculate the molar concentration of sucrose that is isotonic with seawater. (Isotonic: two solutions having the same concentration, and therefore the same osmotic pressure)

Data

- π osmotic pressure = 30.0 atm
- T temperature = 25 C = 298 K
- R gas constant = 0.0821 (L atm)/(K mol)

- π = MRT
- $M = \pi / (RT)$
 - = 30.0 atm / [(0.0821 L atm/K mol) (298 K)]
 - = 1.23 mol/L

Determination of molar mass

The freezing-point depression or osmotic pressure, we can determine the molar mass of a solvent in a solution

Problem

A 7.85 g of an unknown compound is dissolved in 301 g of benzene. The freezing point of the solution is 1.05 C below that of pure benzene. What is molar mass of the compound ?

Data

- m_{su} mass of a solute = 7.85 g
- m_{sl} mass of a solvent = 301 g
- ΔT_{f} freezing-point depression = 1.05 C
- K_{f} freezing-point depression constant of benzene = 5.12 C/m

Step 1

 $\Delta T_{f} = K_{f} m \qquad 0.205 \text{ mol} = 1 \text{ kg}$ $m = \Delta T_{f} / K_{f} \qquad n \text{ mol} = 0.301 \text{ kg}$ $= 1.05 \text{ C/5.12 (C/m)} \qquad n = 0.301 \text{ kg / 0.205 mol}$ $= 0.205 m \qquad = 0.0617 \text{ mol}$

Number of moles:

Step 2

- molar mass = (grams of compund) / (moles of compound)
 - = 7.85 g / 0.0617 mol
 - = 127 g/mol

Problem

A solution is prepared by dissolving 35 g of hemoglobin (Hb) in enough water to make up 1 L in volume. The osmotic pressure of the solution is 10 mmHg at 25 C. What is the molar mass of hemoglobin.

Data

- m_{su} mass of the solute = 35 g
- V_{sn} volume of the solution = 1 L
- π osmotic pressure = 10 mmHg = 0.0132 atm
- T temperature 25 C = 298 K
- R gas constant = 0.0821 (L atm)/(K mol)

π = MRT

- $M = \pi / (RT)$
 - = (0.0132 atm) / [(0.0821 L atm / K mol) (298 K)]
 - = 5.38 x 10⁻⁴ M

The volume of the solution is 1 L, therefore

- 1 L (solution) \rightarrow 5.38 x 10⁻⁴ mol Hb
- moles of Hb = (mass of Hb) / (molar mass of Hb)
- molar mass of Hb = (mass of Hb) / (moles of Hb)
 - = (35 g) / (5.38 x 10⁻⁴ mol)

= 6.51 x 10⁴ g / mol

Colligative (collective) properties of electrolyte solution

• Properties which depend on the number of solute particles after solute dissociation in solution, and not on the nature of solute

van't Hoff factor

actual number of particles after dissociation

i =

number of formula units initially dissolved

For most non-electrolytes dissolved in water, the van't Hoff factor is essentially 1

For most ionic compounds dissolved in water, the van 't Hoff factor is equal to the number of discrete ions in a formula unit of the substance.

Example

NaCl i = 2Na₂SO₄ i = 3

$$\Delta T_{f} = i K_{f} m$$
$$\Delta T_{b} = i K_{b} m$$
$$\pi = i MRT$$

Problem

The osmotic pressure of a 0.010 m potassium iodide (KI) solution at 25 C is 0.465 atm. Calculate the van't Hoff factor for KI at this concentration?

Data

- π osmotic pressure = 0.465 atm
- M molarity of the solution = 0.010 m
- T temperature = 25 C = 298 K
- R gas constant = 0.082 L atm/K mol

Solution

i

- π = iMRT
 - = π/MRT
 - = 0.465 atm/ [(0.010 m) (0.0821 Latm/Kmol) (298K)]
 - = 1.90

The theoretical van't Hoff's factor is sometimes different from the experimentally observed

Electrolyte	i(measured)	i(calculated)
HCI	1.9	2
NaCl	1.9	2
MgSO₄	1.3	2
MgCl ₂	2.7	3
FeCl	3.4	4

Ionic pair

lons of opposite signs which are held together be electrostatic forces

Electrostatic interactions between ions form ionic pairs. The presence of ionc pairs reduces the effective number of particles in solution

Colloids

• A dispersion of particles of one substance through a dispersing medium of another substance

Water colloids

Hydrophilic (water-loving)

Hydrophobic (water-fearing)



How soap works

Soap has a hydrophilic head, and a hydrophobic tail. Hydrophobic tails of soap molecules interact with a dust particle, making the dust particle soluble in water

HYDROPHILIC HYPROPHOBIC HEAD TRIL

