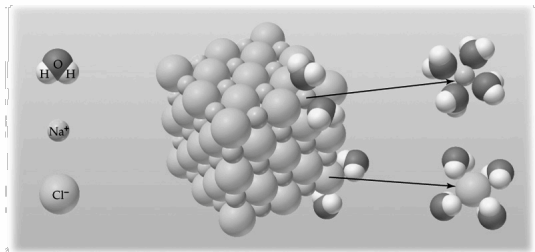


Solutions and Their Behavior

Chapter 11



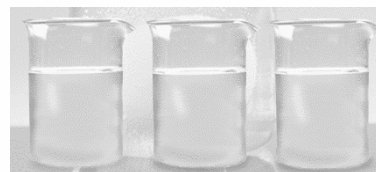
Chemistry 222
Professor Michael Russell

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Last update:
4/28/24

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Solution Behavior



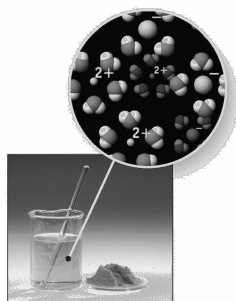
Why does a raw egg swell or shrink when placed in different solutions?

Some Definitions

A solution is a **HOMOGENEOUS** mixture of 2 or more substances in a single phase.

Solutions, Colloids and Suspensions are all examples of homogenous mixtures

Solutions most important to chemists



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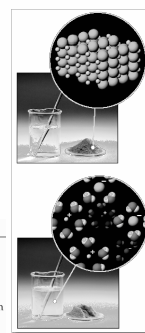
Some Definitions

A **SOLUTION** has at least two constituents

One constituent is usually regarded as the **SOLVENT** and the others as **SOLUTES**

So a **SOLUTION = SOLUTE + SOLVENT**

State of Solution	State of Solvent	State of Solute	Example
Gas	Gas	Gas	Air
Liquid	Liquid	Gas	Oxygen in water
Liquid	Liquid	Liquid	Alcohol in water
Liquid	Liquid	Solid	Salt in water
Solid	Solid	Gas	Hydrogen in palladium
Solid	Solid	Liquid	Mercury in silver
Solid	Solid	Solid	Silver in gold



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Definitions

Solutions can be classified as **unsaturated** or **saturated**.



A **saturated solution** contains the maximum quantity of solute that dissolves at that temperature.



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Definitions

SUPERSATURATED SOLUTIONS contain more than is possible and are unstable.

Unstable solutions; crystallization can usually be stimulated by adding a "seed crystal" or scratching the side of the flask.

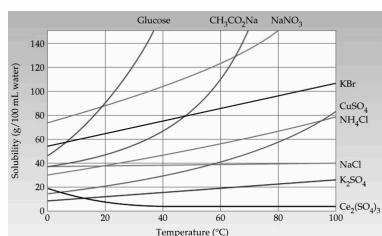


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Solution Definitions

Miscible: Two (or more) liquids that are completely soluble in each other

Solubility: A measure of how much solute will dissolve in a solvent at a specific temperature



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Factors Affecting Solubility

Chemists use the axiom "like dissolves like"

- The closer the intermolecular forces of two phases of matter, the more likely they will dissolve to make a solution
- Polar substances tend to dissolve in polar solvents.
- Nonpolar substances tend to dissolve in nonpolar solvents.

Alcohol	Solubility in H ₂ O	Solubility in C ₆ H ₁₄
CH ₃ OH (methanol)	∞	0.12
CH ₃ CH ₂ OH (ethanol)	∞	∞
CH ₃ CH ₂ CH ₂ OH (propanol)	∞	∞
CH ₃ CH ₂ CH ₂ CH ₂ OH (butanol)	0.11	∞
CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ OH (pentanol)	0.030	∞
CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ OH (hexanol)	0.0088	∞

*Expressed in mol alcohol/100 g solvent at 20 °C. The infinity symbol (∞) indicates that the alcohol is completely miscible with the solvent.

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Factors Affecting Solubility

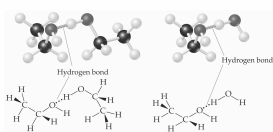
"Like dissolves like"

Example: water (polar, H-bonds) dissolves in acetone (polar)

Example: benzene (ID-ID) dissolves in toluene (ID-ID)

Example: benzene (ID-ID) does not dissolve very well in water (polar, H-bonds)

See [Solubility Handout](#)



We will discuss solubility more in CH 223

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The solubility of liquids and solids does not change appreciably with pressure.

The solubility of a gas in a liquid is directly proportional to its pressure.

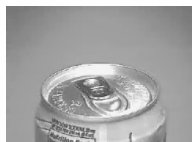
Henry's Law: $S_g = kP_g$

- S_g is the solubility of the gas (M)
- k is the Henry's Law constant for that gas in that solvent, and
- P_g is the partial pressure of the gas above the liquid.



Henry's Law

Dissolving Gases & Henry's Law



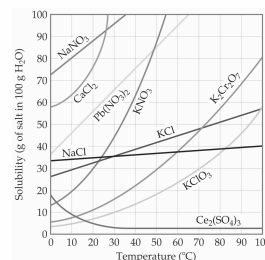
Gas solubility (M) = $S_g = k \cdot P_g$

k for O₂ = 1.66×10^{-6} M/mm Hg

When P_g drops, solubility drops.

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Temperature and Solubility



Generally, the solubility of solid solutes in liquid solvents **increases** with increasing temperature.

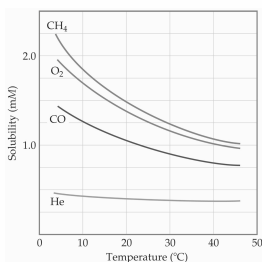
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Temperature and Solubility

The opposite is true for gases.

Carbonated soft drinks are more “bubbly” if stored in the refrigerator.

Warm lakes have less O₂ dissolved in them than cool lakes.



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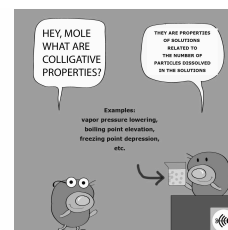
Colligative Properties

On adding a solute to a solvent, the *solvent* properties are modified.

- Vapor pressure **decreases**
- Melting point **decreases**
- Boiling point **increases**
- Osmosis is possible (osmotic pressure)

These changes are called **COLLIGATIVE PROPERTIES**.

They depend only on the **NUMBER of solute particles relative to solvent particles**, not on the **KIND** of solute particles. We need **new concentration units!** (more than molarity, M)



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Concentration Units

• MOLE FRACTION, X

For a mixture of A, B, and C

$$X_A = \text{mol fraction A} = \frac{\text{mol A}}{\text{mol A} + \text{mol B} + \text{mol C}}$$

• MOLALITY, m

$$m \text{ of solute} = \frac{\text{mol solute}}{\text{kilograms solvent}}$$

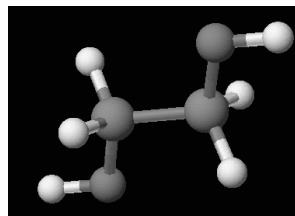
• WEIGHT % = grams solute per total g in solution

see [Concentration Units Handout](#)

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Calculating Concentrations

Dissolve 62.1 g (1.00 mol) of ethylene glycol in 250. g of H₂O. Calculate mol fraction, molality, and weight % of glycol.



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Calculating Concentrations

Dissolve 62.1 g (1.00 mol) of ethylene glycol in 250. g of H₂O. Calculate X, m, and % of glycol.

250. g H₂O = 13.9 mol

$$X_{\text{glycol}} = \frac{1.00 \text{ mol glycol}}{1.00 \text{ mol glycol} + 13.9 \text{ mol H}_2\text{O}}$$

$$X_{\text{glycol}} = 0.0672$$

$$X_{\text{water}} = 1 - 0.0672 = 0.9328$$

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Calculating Concentrations

Dissolve 62.1 g (1.00 mol) of ethylene glycol in 250. g of H₂O. Calculate X, m, and % of glycol.

Calculate molality

$$\text{molality} = \frac{1.00 \text{ mol glycol}}{0.250 \text{ kg H}_2\text{O}} = 4.00 \text{ m}$$

Calculate weight %

$$\% \text{glycol} = \frac{62.1 \text{ g}}{62.1 \text{ g} + 250. \text{ g}} \times 100\% = 19.9\%$$

$$\% \text{water} = 100\% - 19.9\% = 80.1\%$$

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see [Concentration Units Handout](#)

M vs m, χ and wt%

Converting between molality (m), mole fraction (χ) and weight percent (wt%) relatively straightforward



Molarity (M) = mol solute per Liter of solution

Converting from m, χ or wt% to molarity (M) requires solution density in units of g/mL or g/cm³

Recall: solution = solute + solvent

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M vs m, χ and wt%

Example: Convert a 10.7 m aqueous NaOH solution to χ , wt% and M if the solution density = 1.33 g/cm³.

Solution:

Assume 1 kg solvent and 10.7 mol solute

1 kg water = 1000 g H₂O = 55.5 mol H₂O

10.7 mol NaOH * 40.0 g mol⁻¹ = 428 g NaOH

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M vs m, χ and wt%

Example: Convert a 10.7 m aqueous NaOH solution to χ , wt% and M if the solution density = 1.33 g/cm³.

$$\chi_{\text{NaOH}} = 10.7 / (10.7 + 55.5) = 0.162$$

$$\text{wt}\%_{\text{NaOH}} = 428 / (1000 + 428) * 100\%$$

$$\text{wt}\%_{\text{NaOH}} = 30.0\%$$

You do not need the density of the solution to calculate wt% and χ from molality!

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M vs m, χ and wt%

Example: Convert a 10.7 m aqueous NaOH solution to χ , wt% and M if the solution density = 1.33 g/cm³.

If we have 1000 g of solvent (H₂O), then we have 428 g of solute (NaOH) or 1428 g of solution (NaOH + H₂O)

$$\text{Volume of solution} = 1428 \text{ g} * (\text{cm}^3 / 1.33 \text{ g})$$

$$\text{Volume of solution} = 1070 \text{ mL} = 1.07 \text{ L}$$

$$\text{Molarity} = 10.7 \text{ mol NaOH} / 1.07 \text{ L solution}$$

$$\text{Molarity NaOH} = 10.0 \text{ M}$$

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see [Concentration Units Handout](#)**Parts Per Million (ppm)**

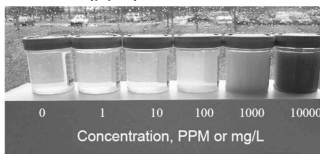
Expresses very dilute concentrations; used by environmental chemists, biologists, geologists, etc.

1.0 ppm = 1.0 g of a substance in a sample with a total mass of 1.0 million (10⁶) g

Density of water about 1.0 g/mL, units often interconverted

Also parts per thousand (ppt), parts per billion (ppb), etc.

Different than the lab parts per thousand, which is a measure of precision

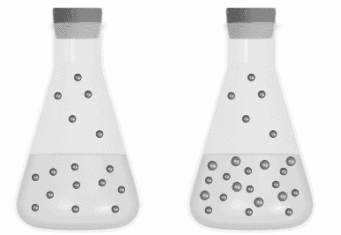


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Understanding Colligative Properties

To understand colligative properties, study the LIQUID-VAPOR EQUILIBRIUM for a solution.



The right animation assumes a non-volatile solute

Understanding Colligative Properties

Vapor Pressure of H₂O over a solution depends on the number of H₂O molecules per solute molecule.

$$P_{\text{solvent}} \text{ proportional to } X_{\text{solvent}}$$

or

$$P_{\text{solvent}} = X_{\text{solvent}} \cdot P^{\circ}_{\text{solvent}}$$

Vapor Pressure of solvent over solution = (Mol frac solvent) • (VP pure solvent)

RAOULT'S LAW: The vapor pressure of solvent over a solution is always LOWER than the pure solvent!

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Raoult's Law

Non-volatile ethylene glycol (62.1 g) is placed in 250. g of water. What is the vapor pressure of water in the solution at 30 °C? (The vapor pressure of pure H₂O is 31.8 mm Hg)

Solution

$$X_{\text{glycol}} = 0.0672 \text{ and so } X_{\text{water}} = ?$$

$$\text{Because } X_{\text{glycol}} + X_{\text{water}} = 1$$

$$X_{\text{water}} = 1.000 - 0.0672 = 0.9328$$

$$P_{\text{water}} = X_{\text{water}} \cdot P^{\circ}_{\text{water}} = (0.9328)(31.8 \text{ mm Hg})$$

$$P_{\text{water}} = 29.7 \text{ mm Hg}$$

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Raoult's Law

Previous examples assumed a non-volatile solute to calculate vapor pressure. If solute is volatile, then:

$$P_{\text{total}} = X_A \cdot P^{\circ}_A + X_B \cdot P^{\circ}_B + X_C \cdot P^{\circ}_C + \dots$$

Example: At 25 °C, heptane (P° = 31 torr) and octane (P° = 11 torr) are mixed such that X_{heptane} = 0.15. Calculate the total vapor pressure of the system.

Solution: This is a two component system.

If X_{heptane} = 0.15, then

$$X_{\text{octane}} = 1 - 0.15 = 0.85$$

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Raoult's Law

Problem: At 25 °C, heptane (P° = 31 torr) and octane (P° = 11 torr) are mixed such that X_{heptane} = 0.15 and X_{octane} = 0.85. Calculate the total vapor pressure of the system.

Solution: In a two component system,

$$P_{\text{total}} = X_{\text{hept}} \cdot P^{\circ}_{\text{hept}} + X_{\text{oct}} \cdot P^{\circ}_{\text{oct}}$$

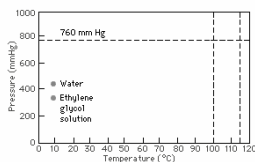
$$P_{\text{total}} = 0.15 \cdot 31 \text{ torr} + 0.85 \cdot 11 \text{ torr}$$

$$P_{\text{total}} = 4.7 + 9.4 = 14.1 \text{ torr}$$

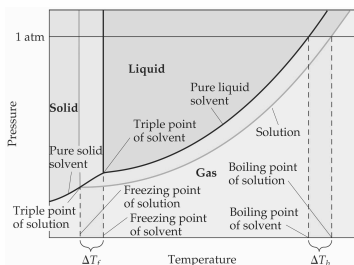
The total vapor pressure in the mixture is 14.1 torr

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The boiling point of a solution is higher than that of the pure solvent.



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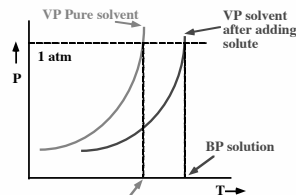


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Elevation of Boiling Point

$$\text{Elevation in BP} = \Delta T_{\text{BP}} = K_{\text{BP}} \cdot m$$

(where K_{BP} is characteristic of solvent)



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	1.22	-114.6	1.99
Carbon tetrachloride, CCl ₄	76.8	5.65	-23.3	29.8
Chloroform, CHCl ₃	61.2	3.63	-63.5	4.68

Change in Boiling Point

Dissolve 62.1 g of glycol (1.00 mol) in 250. g of water. What is the BP of the solution?

$K_{BP} = 0.512 \text{ }^\circ\text{C/molal}$ for water (see Textbook)

Solution

1. Calculate solution molality = 4.00 m
2. $\Delta T_{BP} = K_{BP} \cdot m$
 $\Delta T_{BP} = 0.512 \text{ }^\circ\text{C/molal} (4.00 \text{ molal})$
 $\Delta T_{BP} = 2.05 \text{ }^\circ\text{C}$
 $BP = 100.00 + 2.05 = 102.05 \text{ }^\circ\text{C}$

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Change in Freezing Point

Pure water

Ethylene glycol/water solution



The freezing point of a solution is **LOWER** than that of the pure solvent.

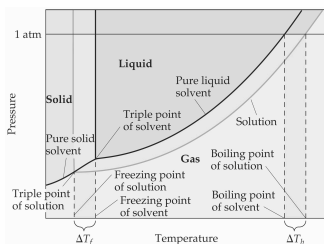
Notice how the animations incorrectly use "dense" ice!

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The freezing point of a solution is lower than that of the pure solvent.

$$\text{FP depression} = \Delta T_{FP} = K_{FP} \cdot m$$

* Molal Boiling-Point-Elevation and Freezing-Point-Depression Constants			
Solvent	Normal Boiling Point ($^\circ\text{C}$)	K_b ($^\circ\text{C}/m$)	K_f ($^\circ\text{C}/m$)
Water, H_2O	100.0	0.51	1.86
Benzene, C_6H_6	80.1	2.53	5.12
Ethanol, $\text{C}_2\text{H}_5\text{OH}$	78.4	1.22	1.99
Carbon tetrachloride, CCl_4	76.8	5.02	29.9
Chloroform, CHCl_3	61.2	3.63	63.5



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Freezing Point Depression

Calculate the FP of a 4.00 molal glycol/water solution.

$K_{FP} = 1.86 \text{ }^\circ\text{C/molal}$ (from Textbook)

Solution

$$\Delta T_{FP} = K_{FP} \cdot m$$

$$= (1.86 \text{ }^\circ\text{C/molal})(4.00 \text{ m})$$

$$\Delta T_{FP} = 7.44 \text{ }^\circ\text{C}$$

$$\text{So the FP} = 0 \text{ }^\circ\text{C} - 7.44 \text{ }^\circ\text{C} = -7.44 \text{ }^\circ\text{C}$$

Many textbooks use *negative* K_{FP} values
 If using *negative* K_{FP} , ΔT_{FP} is also negative
 Concentration (molality) *must* be positive!

Freezing Point Depression of Electrolytes

How much NaCl must be dissolved in 4.00 kg of water to lower FP to $-10.00 \text{ }^\circ\text{C}$?

Solution

Calculate required molality

$$\Delta T_{FP} = K_{FP} \cdot m$$

$$(0 - -10.00 \text{ }^\circ\text{C}) = (1.86 \text{ }^\circ\text{C/molal}) \cdot m$$

$$m = 5.38 \text{ molal}$$

Use $-10 \text{ }^\circ\text{C}$ for ΔT if $-1.86 \text{ }^\circ\text{C}/m$



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Freezing Point Depression of Electrolytes

How much NaCl must be dissolved in 4.00 kg of water to lower FP to $-10.00 \text{ }^\circ\text{C}$?

Solution

Concentration required = 5.38 molal

This means we need 5.38 mol of dissolved particles per kg of solvent.

Recognize that m represents the total concentration of *all* dissolved particles.

Recall that 1 mol NaCl(aq)

$\rightarrow 1 \text{ mol Na}^+(\text{aq}) + 1 \text{ mol Cl}^-(\text{aq})$... or: 2 mol of particles per 1 mol of ionic solute!

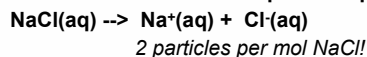
Freezing Point Depression of Electrolytes

How much NaCl must be dissolved in 4.00 kg of water to lower FP to -10.00 °C?

Solution

Concentration required = 5.38 molal

We need 5.38 mol of dissolved particles per kg of solvent.



To get 5.38 mol/kg of particles we need

$$5.38 \text{ mol} / 2 = 2.69 \text{ mol NaCl} / \text{kg}$$

$$2.69 \text{ mol NaCl} / \text{kg} \rightarrow 157 \text{ g NaCl} / \text{kg}$$

$$(157 \text{ g NaCl} / \text{kg}) \cdot (4.00 \text{ kg}) = 629 \text{ g NaCl}$$

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Boiling Point Elevation and Freezing Point Depression

$$\Delta T = k \cdot m \cdot i$$

A generally useful equation

i = van't Hoff factor = number of particles produced per formula unit.

Compound	Theoretical Value of i
glycol	1
NaCl	2
CaCl ₂	3

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We will use the theoretical value of i in CH 222

Colligative Properties and Molar Mass

Can use colligative properties to find molar mass of solute

Molar mass = grams solute / moles solute

Example: Find the molar mass of azulene if 0.640 g are dissolved in 99.0 g of benzene ($k_b = 2.53 \text{ }^\circ\text{C/m}$), and the observed boiling point is 80.230 °C. (normal bp = 80.100 °C)

Solution:

$$\Delta T = (80.230 - 80.100) = 0.130 \text{ }^\circ\text{C}$$

$$m = \Delta T / k_b = 0.130 \text{ }^\circ\text{C} / 2.53 \text{ }^\circ\text{C/m} = 5.14 \cdot 10^{-2} \text{ mol} / \text{kg}$$

$$0.0990 \text{ kg} \cdot 5.14 \cdot 10^{-2} \text{ mol} / \text{kg solvent} = 5.05 \cdot 10^{-3} \text{ mol}$$

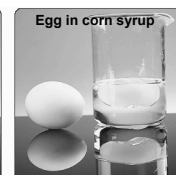
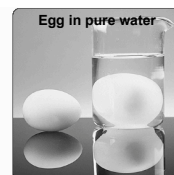
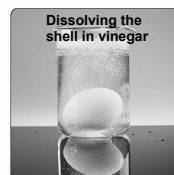
$$\text{Molar mass} = \text{grams solute} / \text{moles solute}$$

$$\text{Molar mass} = 0.640 \text{ g} / 5.05 \cdot 10^{-3} \text{ moles} = 127 \text{ g mol}^{-1}$$

MM = (grams solute * k) / (ΔT * kg solvent)
useful for both bp and fp calculations

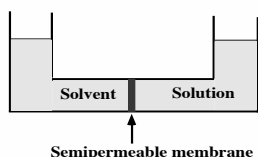
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Osmosis



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Osmosis



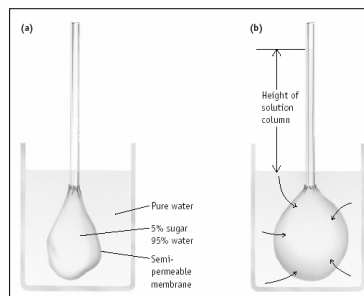
The semipermeable membrane allows only the movement of solvent molecules.

Solvent molecules move from pure solvent to solution in an attempt to make both have the same concentration of solute.

Driving force is entropy (CH 223)

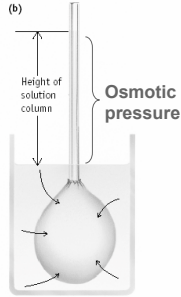
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Process of Osmosis



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Osmotic Pressure, Π

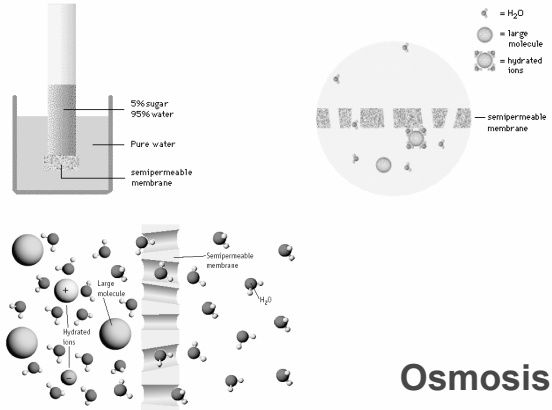


Equilibrium is reached when pressure - the **OSMOTIC PRESSURE, Π** - produced by extra solution counterbalances pressure of solvent molecules moving through the membrane.

$$\Pi = icRT$$

c = conc. in mol/L
 i = the van't Hoff factor
 $R = 0.082057 \text{ (L atm)/(mol K)}$
 T = Kelvin Temp.

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Osmosis

Osmosis Calculating a Molar Mass

Dissolve 35.0 g of hemoglobin in enough water to make 1.00 L of solution. π measured to be 10.0 mm Hg at 25 °C. Calc. molar mass of hemoglobin.

Solution

- (a) Calc. π in atmospheres
 $\pi = 10.0 \text{ mm Hg} \cdot (1 \text{ atm} / 760 \text{ mm Hg})$
 $\pi = 0.0132 \text{ atm}$

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Osmosis Calculating a Molar Mass

Dissolve 35.0 g of hemoglobin in enough water to make 1.00 L of solution. π measured to be 10.0 mm Hg at 25 °C. Calc. molar mass of hemoglobin.

Solution

- (b) Calc. concentration from $\pi = cRT$

$$\text{Conc} = \frac{0.0132 \text{ atm}}{(0.082057 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(298\text{K})}$$

$$\text{Conc} = 5.40 \times 10^{-4} \text{ mol/L} (* 1 \text{ L})$$

(c) Calc. molar mass
Molar mass = 35.0 g / 5.40 x 10⁻⁴ mol
Molar mass = 64,800 g/mol

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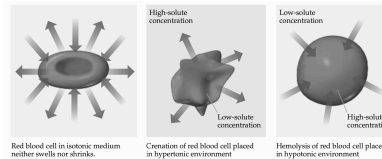
Osmosis

Osmosis of solvent from one solution to another can continue until the solutions are **ISOTONIC** - they have the same concentration.

Osmotic pressure important in living systems



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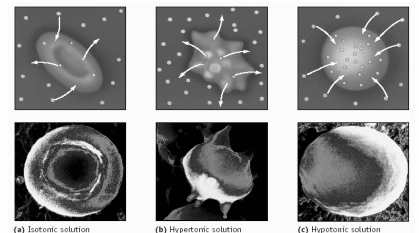


Red blood cell in isotonic medium neither swells nor shrinks.

Crenation of red blood cell placed in hypertonic environment

Hemolysis of red blood cell placed in hypotonic environment

Osmosis & Living Cells



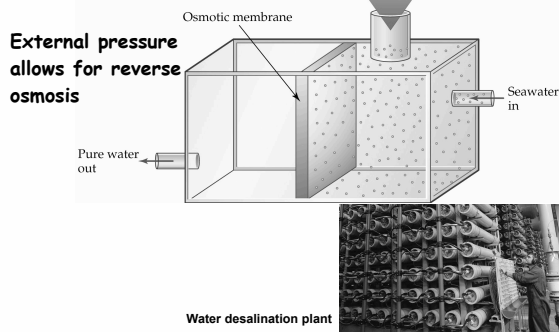
(a) Isotonic solution

(b) Hypertonic solution

(c) Hypotonic solution

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Reverse Osmosis Water Desalination



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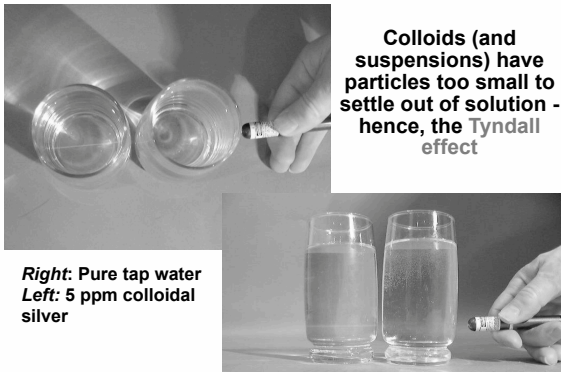
Other Homogeneous Mixtures

Homogeneous mixtures: more than just solutions!

- **Solutions** - most important, particles in the range 0.2 - 2 nm in diameter (salt water, acids, etc.)
- **Colloids** - a "solution" with particles in the range 2 - 1000 nm (milk, fog, etc.)
- **Suspensions** - contain particles greater than 1000 nm in diameter (blood, paint, aerosol sprays, etc.)

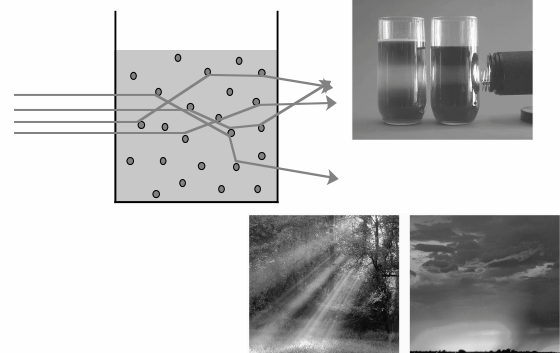
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The Tyndall Effect



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The Tyndall Effect



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Colloids

Colloids have a dispersed phase ("solute") and a dispersing medium ("solvent"). Many types of colloids:

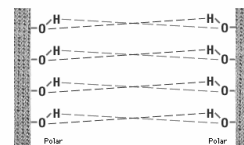
Type	D Medium	D Phase	Examples
aerosol	gas	liquid	fog, clouds
aerosol	gas	solid	smoke, viruses
foam	liquid	gas	shaving cream
emulsion	liquid	liquid	milk, mayonnaise
sol	liquid	solid	mud, paint, ink
foam	solid	gas	foam rubber, sponge
gel	solid	liquid	jelly, cheese, yogurt
solid sol	solid	solid	steel, bronze, pearls

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Surfactant Emulsifying Agents

Surfactants (or Surfactant Emulsifying Agents) promote the formation of colloids by coating dispersed phase, preventing the formation of large particles.

Surfactants used for cleaning called detergents.

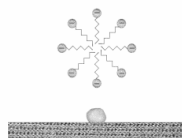
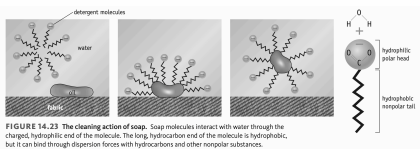
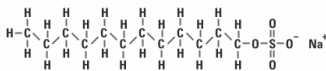


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Surfactant Emulsifying Agents

Soaps and detergents: surfactants with both a hydrophobic (nonpolar) and a hydrophilic (polar) end. Can remove nonpolar oils and polar compounds.

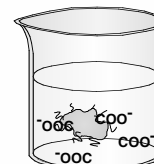
Soaps often grouped in micelles with like-polarities grouped together.



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Oil and Water

Oil and water do not mix due to differences in polarity; water *excludes* oil due to stronger intermolecular forces
Soap removes oil readily using nonpolar end of micelle



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when someone asks you to elevate the boiling point



See:

- [Chapter Eleven Study Guide](#)
- [Chapter Eleven Concept Guide](#)
- [Important Equations \(following this slide\)](#)
- [End of Chapter Problems \(following this slide\)](#)

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End of Chapter 11



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Important Equations, Constants, and Handouts from this Chapter:

- **solution = solute + solvent**
- **see [Concentration Units Handout](#)**

Solution Concentrations:

$$\text{Molarity (M)} = \frac{\text{mol Solute}}{\text{L Solution}}$$

$$\text{molality (m)} = \frac{\text{mol Solute}}{\text{kg Solvent}}$$

$$\text{weight \%} = \frac{\text{mass Solute}}{\text{Total mass}}$$

$$\text{mole fraction (}\chi\text{)} = \frac{\text{mol A}}{\text{total mol}}$$

$$\text{ppm} = \frac{1.0 \text{ g of substance}}{1.0 \text{ million g sample}}$$

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Henry's Law:

$$S_g = k \cdot P_g$$

Raoult's Law / Vapor Pressure Depression:

$$P_{\text{solvent}} = \chi_{\text{solvent}} \cdot P_{\text{solvent}}^o$$

Boiling Point Elevation / Freezing Point Depression:

$$\Delta T_{\text{BPF/FP}} = K_{\text{BPF/FP}} \cdot \left(\frac{\text{mol Solute}}{\text{kg Solvent}} \right) \cdot i$$

Osmosis:

$$\pi = i \left(\frac{\text{mol Solute}}{\text{L Solvent}} \right) RT$$

$$R = 0.082057 \text{ L atm mol}^{-1} \text{ K}^{-1}$$

$i = \text{van't Hoff factor}$

End of Chapter Problems: Test Yourself

1. Fill in the blanks in the table. All solutions are aqueous.

Compound	Molality	Weight Percent	Mole Fraction
NaI	0.15	—	—
C ₂ H ₅ OH	—	5	—
C ₁₂ H ₂₂ O ₁₁	—	—	0.0027

- Hydrochloric acid is sold as a concentrated aqueous solution. If the molarity of commercial HCl is 12.0 and its density is 1.18 g/cm³, calculate the molality and the weight percent of HCl in the solution.
- A soda can has an aqueous CO₂ concentration of 0.0506 M. What is the pressure of CO₂ gas in the can? ($k_H = 4.48 \times 10^{-5} \text{ M/mm Hg}$)
- Pure nonvolatile iodine (105 g) is dissolved in 325 g of CCl₄ at 65 °C. Given that the vapor pressure of CCl₄ at this temperature is 531 mm Hg, what is the vapor pressure of the solution at 65 °C?
- What is the boiling point of a solution composed of 15.0 g of CHCl₃ (which boils at 61.70 °C) and 0.515 g of the nonvolatile solute acenaphthene, C₁₂H₁₀?
- An aqueous solution containing 1.00 g of bovine insulin (a protein, not ionized) per liter has an osmotic pressure of 3.1 mm Hg at 25 °C. Calculate the molar mass of bovine insulin.

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End of Chapter Problems: Answers

1. Completed table:

Compound	Molality	Weight percent	Mole fraction
NaI	0.15	2.29	.0027
C ₂ H ₅ OH	1.1	5.00	.020
C ₁₂ H ₂₂ O ₁₁	0.15	4.90	.0027

- 16.2 m, 37.1%
- 1130 mm Hg
- 444 mm Hg
- 62.51 °C
- $6.0 \times 10^3 \text{ g/mol}$

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