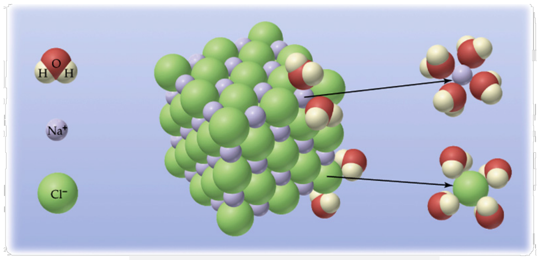
Chapter 11: Solutions and Colloids



Chemistry 222

Professor Michael Russell

http://mhchem.org/222

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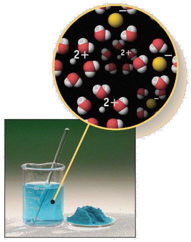
11.1 - The Dissolution Process

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; *most substances we encounter are solutions (ocean, air, coffee, etc.)*



A **solution** has a **solute** (substance being dissolved) and a **solvent** (substance dissolving the solute)

In water, covalent molecules (ie sugar) that dissolve are *hydrated*:

$C_{12}H_{22}O_{11}(s) \rightarrow C_{12}H_{22}O_{11}(aq)$

Ionic molecules that dissolve *dissociate*:

$NaCl(s) \rightarrow Na^+(aq) + Cl^{1-}(aq)$

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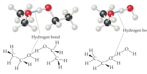
Soluble and Insoluble

When one substance dissolves in another, they are said to be **soluble** *(salt and water)*

When one substance does not dissolve in another, they are said to be **insoluble** *(oil and water)*

Solubility depends on IM forces, etc.



Chemists use "**like dissolves like**"

The closer the IM forces of two phases of matter, the more likely they will dissolve to make a solution

Polar substances dissolve in polar solvents; nonpolar substances dissolve in nonpolar solvents.

Ex: **water** (polar, H-bonds) dissolves in **ethanol** (polar, H-bonds)

Ex: **oil** (ID-ID) dissolves in **kerosene** (ID-ID)

Ex: **oil** (ID-ID) does not dissolve very well in **water** (polar, H-bonds)

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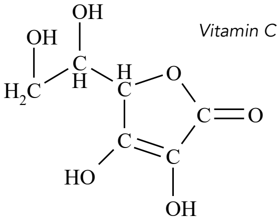
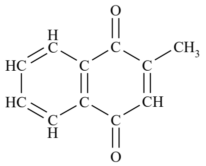
Soluble Examples

Example: is **Vitamin C** water soluble?

Yes - the four -OH groups display H-bonding (like water, so Vitamin C is water soluble



Vitamin C



Vitamin K_3

Example: is **Vitamin** $\mathbf{K_3}$ water soluble?

No - most of molecule nonpolar, C=O groups cancel, so Vitamin K_3 is not water soluble

(Vitamin K_3 is *fat soluble*)

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11.2 - Electrolytes

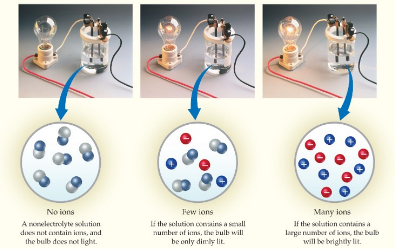
In water, **covalent** molecules (ie sugar) that dissolve are *hydrated*:

$C_{12}H_{22}O_{11}(s) \rightarrow C_{12}H_{22}O_{11}(aq)$

Ionic molecules that dissolve *dissociate*:

$NaCl(s) \rightarrow Na^+(aq) + Cl^{1-}(aq)$

Dissociated species are **ionic** - they have a charge - and **can conduct electricty**



Nonelectrolyte: covalent molecules dissolved in water *(no ions)*

Weak Electrolyte: compounds partially dissociated in water *(few ions)*

Strong Electrolyte: ionic compounds dissolved in water *(many ions)*

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Strong Electrolytes

Strong electrolytes are **strong acids, strong bases,** and **ionic compounds that dissolve in water**

Five strong acids:

HCl	**hydrochloric acid**
HBr	**hydrobromic acid**
HI	**hydroiodic acid**
$\mathbf{HNO_3}$	**nitric acid**
$\mathbf{HClO_4}$	**perchloric acid**

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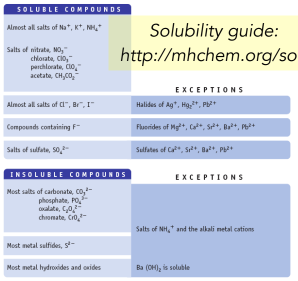


SOLUBLE COMPOUNDS	EXCEPTIONS
Almost all salts of Na^+, K^+, NH_4^+	
Salts of nitrate, NO_3^- chlorate, ClO_3^- perchlorate, ClO_4^- acetate, $CH_3CO_2^-$	
Almost all salts of Cl^-, Br^-, I^-	Halides of Ag^+, Hg_2^{2+}, Pb^{2+}
Compounds containing F^-	Fluorides of Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}
Salts of sulfate, SO_4^{2-}	Sulfates of Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}

INSOLUBLE COMPOUNDS	EXCEPTIONS
Most salts of carbonate, CO_3^{2-} phosphate, PO_4^{3-} oxalate, $C_2O_4^{2-}$ chromate, CrO_4^{2-}	Salts of NH_4^+ and the alkali metal cations
Most metal sulfides, S^{2-}	
Most metal hydroxides and oxides	Ba $(OH)_2$ is soluble

Solubility guide: http://mhchem.org/sol/

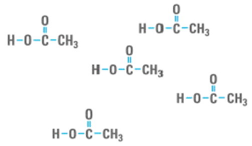
Three strong bases:

NaOH	**sodium hydroxide**
KOH	**potassium hydroxide**
LiOH	**lithium hydroxide**

Weak and Non-Electrolytes

Weak electrolytes are **weak acids and weak bases**

Nonelectrolytes include sugar(aq), coffee(aq), ethanol(aq), *many more* - all covalent solutes



Weak Electrolyte Examples:

CH_3CO_2H	**acetic acid**
H_2CO_3	**carbonic acid**
H_3PO_4	**phosphoric acid**
HF	**hydrofluoric acid**
NH_3	**ammonia** *(base)*
C_5H_5N	**pyridine** *(base)*
$C_2H_5NH_2$	**ethylamine** *(base)*
N_2H_4	**hydrazine** *(base)*

plus many more!

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11.3 - Solubility

The **solubility** of a solute is the maximum concentration that can be achieved in the given solvent and temperature

When a solute's concentration is equal to its solubility, the solution is said to be **saturated**; at lesser concentrations, the solution is **unsaturated**.

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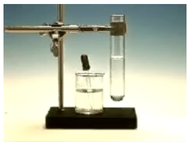
unsaturated



saturated

super saturated

Supersaturated Solutions contain more than is possible and are unstable.

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

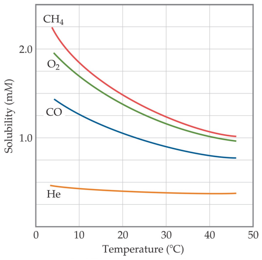
Solubility of Gases in Liquids

Gases usually more soluble in liquids at lower temperatures

Carbonated soft drinks are more "bubbly" if stored in the refrigerator; warm lakes have less dissolved O_2 than cool lakes.

The solubility of a gas in a liquid dictated by **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.

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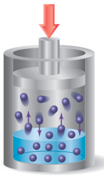
Henry's Law

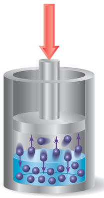
Henry's Law:

$$S_g = kP_g$$

S_g is the **solubility** of the gas **(M)**
k is the **Henry's Law constant** for that gas
P_g is the **partial pressure of the gas** above the liquid.



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



k for O_2 = 1.66 x 10^{-6} M/mm Hg

When P_g drops, solubility drops.

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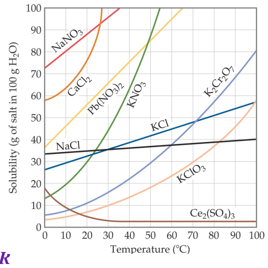
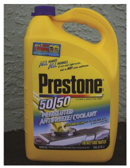
Other Solutions

Solids in liquids generally become more soluble as temperature increases

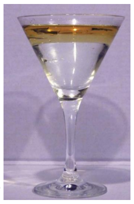


Liquids in liquids can be infinitely soluble (assuming IM forces similar) - these are **miscible** solutions (ethanol and water, water and antifreeze, etc.)



miscible

Immiscible solutions usually have different IM forces (oil in water, etc.)



oil and water, immiscible

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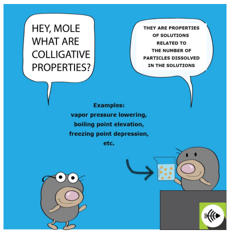
11.4 - Colligative Properties

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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New 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

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

Mass Percent:

$$\text{Mass \%} = \frac{\text{grams solute}}{\text{total grams}} * 100\%$$

see Solutions Calculations handout

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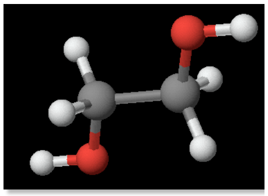
Concentration Calculations



ethylene glycol

Dissolve 62.1 g (1.00 mol) of ethylene glycol in 250. g of H_2O. Calculate mol fraction, molality, and mass % of ethylene glycol.

250. g H_2O = 13.9 mol

$$X_{glycol} = \frac{1.00 \text{ mol glycol}}{1.00 \text{ mol glycol} + 13.9 \text{ mol } H_2O}$$

$X_{glycol} = 0.0672$

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

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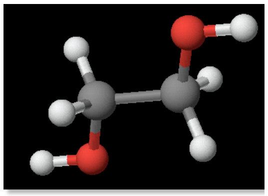
Concentration Calculations *(continued)*



ethylene glycol

Dissolve 62.1 g (1.00 mol) of ethylene glycol in 250. g of H_2O. Calculate mol fraction, molality, and mass % of ethylene glycol.

Calculate molality

$$\text{molality} = \frac{1.00 \text{ mol glycol}}{0.250 \text{ kg } H_2O} = 4.00 \text{ m}$$

Calculate mass%

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

%water = 100% - 19.9% = 80.1%

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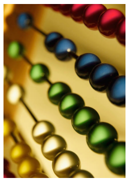
Concentration Calculations and Molarity



Converting between molality (m), mole fraction (χ) and mass percent (mass%) *relatively* straightforward

Molarity (M) = mol solute per *Liter of solution*

Liter is a volume, so need a solution density (in units of g/mL *or* g/cm³) when converting to and from molarity (M)

Recall: **solution = solute + solvent**

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Concentration Calculations and Molarity (example)

Example: **Convert a 10.7 m aqueous NaOH solution to χ, mass% 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 H_2O = 55.5 mol H_2O

10.7 mol NaOH * 40.0 g mol^{-1} = 428 g NaOH

χ_{NaOH} = 10.7 / (10.7 + 55.5) = 0.162

mass%$_{NaOH}$ = 428 / (1000 + 428) * 100%

mass%$_{NaOH}$ = 30.0%

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

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Concentration Calculations and Molarity *(continued)*

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

Solution (continued):

We have 1000 g of solvent (H_2O) and 428 g of solute (NaOH) ***or***

1428 g of solution *(NaOH + H_2O)*

Volume of solution = ***1428 g * (cm³ / 1.33 g)* = 1070 mL = 1.07 L**

Molarity = 10.7 mol NaOH / 1.07 L solution

Molarity NaOH = 10.0 M

Molality (m) value usually greater than molarity (M)

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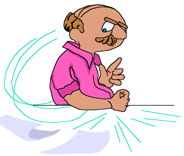
Tips on Converting Concentrations:



- **Need g solute** *and* **g solvent;** g solute + g solvent = g solution, use density solution to get to volume
- **If starting with molality (m):** assume 1 kg solvent, then molality = mol solute
- **If starting with mole fraction (χ):** assume 1 mol total; χ = mol solute, and (1 - χ) = mol solvent
- **If starting with mass%:** assume 100 g total, then mass% = g solute and 100 - mass% = g solvent
- **If starting with molarity (M):** assume 1 L solvent, then mol solute = molarity, turn into g solute
- *See Solutions Calculations handout*

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Parts Per Million (ppm)

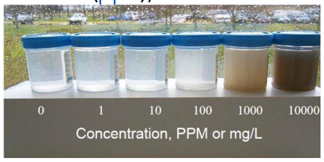
Expresses very dilute concentrations; used by environmental chemists, biologists, geologists, etc.

1.0 ppm = 1.0 g of a substance in a sample with a total mass of 1.0 million (10^6) 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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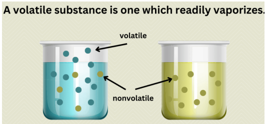
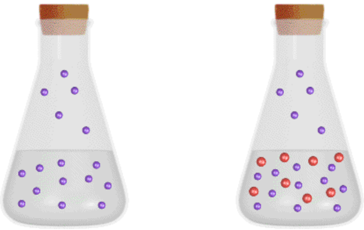
Colligative Properties



Adding a nonvolatile solute to a solvent *lowers* the solvent's vapor pressure (most solvents are volatile)

The result: the solvent evaporates more slowly (if open container)

To understand colligative properties, study the **Liquid-Vapor Equilibrium**:



The red molecules are nonvolatile solutes

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Raoult's Law (Vapor Pressure Lowering)

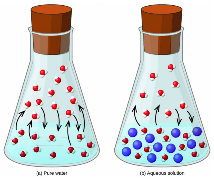
Vapor Pressure of H_2O over a solution ($P_{solvent}$) depends on the number of H_2O molecules per solute molecule ($X_{solvent}$)

$P_{solvent}$ *proportional to* $X_{solvent}$ *or*

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

where **$P^{\circ}_{solvent}$** = *vapor pressure of pure solvent*

Vapor Pressure of *solvent* over *solution* ($P_{solvent}$) = (mole 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 Example

Non-volatile ethylene glycol (62.1 g) is placed in 250. g of water. X_{glycol} = 0.0672 *(earlier example)* **What is the vapor pressure of water in the solution at 30 ˚C?** *(The vapor pressure of pure H_2O is 31.8 mm Hg)*

Solution: $P_{solvent} = X_{solvent} \bullet P^{\circ}_{solvent}$

X_{glycol} = 0.0672 and so **X_{water} = ?**

Because **$X_{glycol} + X_{water} = 1$**

X_{water} = 1.000 - 0.0672 = **0.9328**

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

P_{water} = 29.7 mm Hg

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Raoult's Law With a Volatile Solute(s)

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

$$P_{total} = X_A \bullet P^{o}_{A} + X_B \bullet P^{o}_{B} + X_C \bullet P^{o}_{C} + \ldots$$

Example: At 25 ˚C, heptane (P^{o} = 31 torr) and octane (P^{o} = 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_{octane} = 1 - 0.15 = **0.85**

$P_{total} = X_{hept} \bullet P^{o}_{hept} + X_{oct} \bullet P^{o}_{oct}$

P_{total} = 0.15 • 31 torr + 0.85 • 11 torr

P_{total} = 4.7 + 9.4 = **14.1 torr**

The total vapor pressure in the mixture is 14.1 torr

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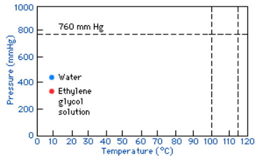
Changes to Boiling & Freezing Points

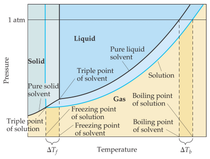
Solutes block solvent particles from evaporating (**boiling**) and condensing (**freezing**) in solutions

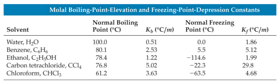
The change to the boiling or freezing point depicted as **ΔT** *(always positive)*

ΔT_b = difference between boiling point pure solvent and solution

ΔT_f = difference between freezing point pure solvent and solution







Molal Boiling-Point-Elevation and Freezing-Point-Depression Constants

Solvent	Normal Boiling Point (°C)	K_b (°C/m)	Normal Freezing Point (°C)	K_f (°C/m)
Water, H_2O	100.0	0.51	0.0	1.86
Benzene, C_6H_6	80.1	2.53	5.5	5.12
Ethanol, C_2H_5OH	78.4	1.22	−114.6	1.99
Carbon tetrachloride, CCl_4	76.8	5.02	−22.3	29.8
Chloroform, $CHCl_3$	61.2	3.63	−63.5	4.68

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Boiling Point Elevation

Boiling points of solutions ***higher*** than the boiling points of pure solvent - **boiling point elevation**

$$\Delta T_b = K_b \bullet m$$

ΔT_b = difference between boiling point pure solvent and solution

K_b = boiling point constant for the solution

m = molality of the solution

Example: **Dissolve 62.1 g of ethylene glycol** (*1.00 mol*) in **250. g** of water. **What is the boiling point of the solution?**

K_b = 0.512 °C/molal *for water*

Solution

1. Calculate **molality = 4.00 m**

2. $\Delta T_b = K_b \bullet m$

ΔT_b = 0.512 °C/molal (4.00 m)

ΔT_b = 2.05 °C

Boiling point **= 100.00 + 2.05**

Boiling point **= 102.05 °C**

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Freezing Point Lowering

Freezing points of solutions ***lower*** than the freezing points of pure solvent - **freezing point lowering**

$$\Delta T_f = K_f \bullet m$$

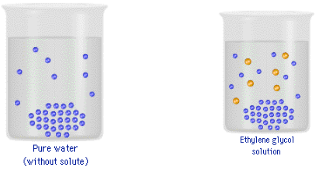
ΔT_f = difference between freezing point pure solvent and solution

K_f = freezing point constant for the solution

m = molality of the solution

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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Freezing Point Lowering - Example

Calculate the Freezing Point of a 4.00 m ethylene glycol/water solution.

K_f = 1.86 °C/molal

Solution

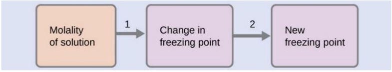
ΔT_f = $K_f \bullet m$

= (1.86 °C/molal)(4.00 m)

ΔT_f = 7.44 °C

So the FP = 0 °C - 7.44 °C = **-7.44 °C**

Many textbooks use negative K_f values. If using negative K_f, ΔT_f is also negative, concentration (molality) *must* be positive!



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Freezing Point Lowering with Electrolytes

Colligative properties depend on the number of solute particles.... so with ionic compounds that dissociate, this leads to:

$\Delta T_f = K_f \bullet m \bullet i$

where **i** = **van't Hoff factor** which equals the number of ions per compound

i = 1 (all covalent molecules)

i = 2 for NaCl since NaCl in water becomes Na^+ and Cl^-, two particles per NaCl

i = 3 for $MgCl_2$ (Mg^{2+} and 2 Cl^-)

i = 4 for $AlCl_3$ (Al^{3+} and 3 Cl^-)

Example: How much NaCl must be dissolved in 4.00 kg of water to lower FP to -10.00 °C? (K_f = 1.86 °C/m)

$\Delta T_f = K_f \bullet m \bullet i$

ΔT_f = 10.00 °C *(ΔT always positive)*

i = 2 for NaCl (Na^+ and Cl^-)

m = ΔT_f / $K_f \bullet$ i = 10.00 / 1.86 * 2

m = 2.68817... mol NaCl / kg (* 4.00 kg)

= 10.7526... mol NaCl * 58.44 g/mol

= **628 g NaCl**

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Colligative Properties and Molar Mass

The colligative properties offer several methods to find the molar mass of the solute (grams solute / moles solute)

$\Delta T_f = K_b \bullet m \bullet i$

Example: Find the molar mass of azulene if 0.640 g are dissolved in 99.0 g of benzene (K_b = 2.53 °C/m), and the observed boiling point is 80.230 °C. (normal bp = 80.100 °C)

Solution:

ΔT = (80.230 - 80.100) = **0.130 °C**

i = 1 (covalent)

m = $\Delta T/K_b$ = 0.130 °C/2.53 °C/m = **5.14*10^{-2}** mol / kg

0.0990 kg * 5.14*10^{-2} mol / kg solvent = **5.05*10^{-3} mol solute**

Molar mass = grams solute / moles solute

Molar mass = 0.640 g / 5.05*10^{-3} moles = **127 g mol^{-1}**

MM = (grams solute * K)/(ΔT*kg solvent)

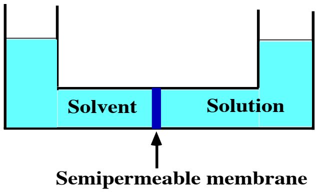
useful for both bp and fp calculations

Osmosis / Osmotic Pressure

Osmosis (or Osmotic Pressure) is a colligative property useful to biochemists and scientists in general

In osmosis, two solutions are separated by a **semipermeable membrane** which allows solvent molecules *only* to travel back and forth

Net result: solvent molecules travel from dilute concentration to more concentrated solution



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

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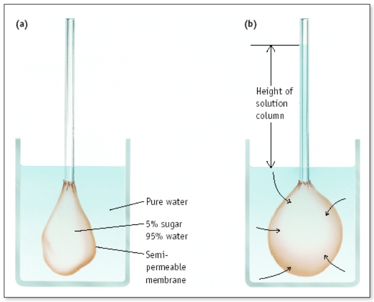
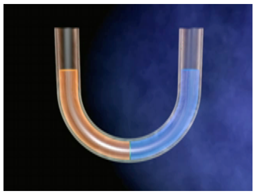
Osmosis





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

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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.

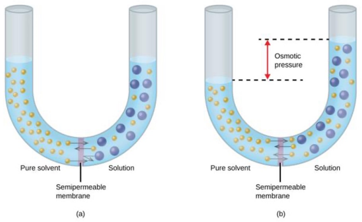
$$\pi = icRT$$

c = concentration in mol/L

i = the van't Hoff factor

R = 0.082057 (L atm)/(mol K)

T = Kelvin Temperature



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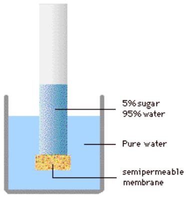

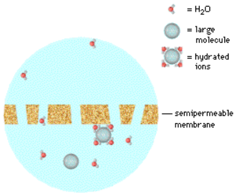

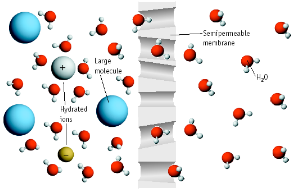

Osmosis

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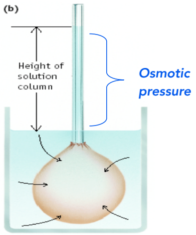
Example: Dissolve 35.0 g of hemoglobin in enough water to make 1.00 L of solution. π measured to be 10.0 mm Hg at 25 ˚C. Find the **molar mass of hemoglobin.**



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Osmosis and Molar Mass

(a) Calc. π in atmospheres

π = 10.0 mm Hg • (1 atm / 760 mm Hg)

π = **0.0132 atm**

(b) Calc. concentration from π = cRT

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

Conc = **5.40×10^{-4} mol/L** (* 1 L)

(c) Calc. molar mass

Molar mass = 35.0 g / 5.40×10^{-4} mol

Molar mass = 64,800 g/mol

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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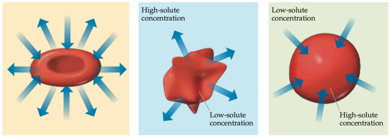
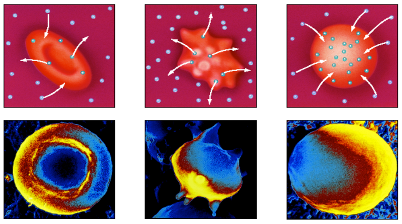


Red blood cell in isotonic medium neither swells nor shrinks.

Crenation of red blood cell placed in hypertonic environment

Hemolysis of red blood cell placed in hypotonic environment

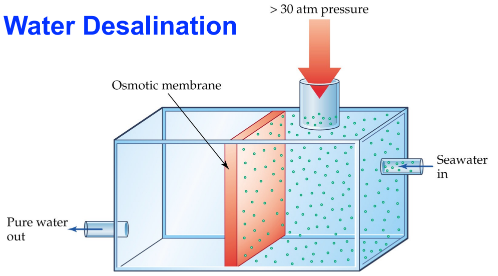


(a) Isotonic solution

(b) Hypertonic solution

(c) Hypotonic solution

Reverse Osmosis & Water Desalination



Applying a pressure to the solution greater than the osmotic pressure results in **reverse osmosis.**

Solvent molecules are pushed from the solution to the pure solvent.

Reverse osmosis used to desalinate sea water

water desalination plant

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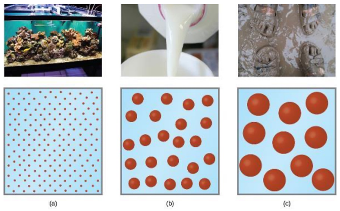
11.5 - Colloids (and Suspensions)

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.) *Homogeneous*

Suspensions - contain particles greater than 1000 nm in diameter (blood, paint, aerosol sprays, etc.) *Heterogeneous*



(a) A salt water **solution** (aquarium)
(b) A **colloid** (milk) - *particles do not settle*
(c) A **suspension** (mud) - *particles can settle*

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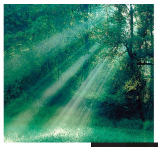
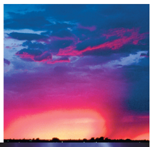
The Tyndall Effect

In a colloid, light is scattered by particles, resulting in the **Tyndall Effect**

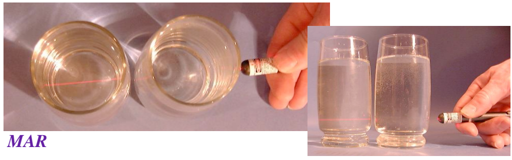
Colloids have particles too small to settle out of solution but can be confirmed by light scattering





Searchlight beams made visible when light is scattered by colloidal-size particles in the air (fog, smoke, etc.)

Right: Pure tap water
Left: 5 ppm colloidal silver



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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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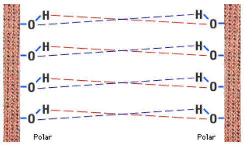
Surfactant Emulsifying Agents

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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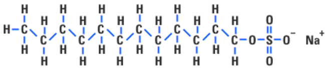
Soaps

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.



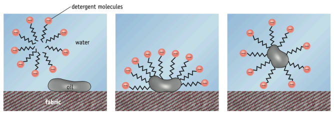
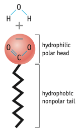
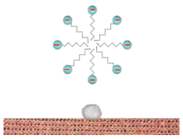


FIGURE 14.23 The cleaning action of soap. Soap molecules interact with water through the charged, hydrophilic end of the molecule. The long, hydrocarbon end of the molecule is hydrophobic, but it can bind through dispersion forces with hydrocarbons and other nonpolar substances.





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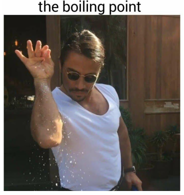
End of Chapter 11

See:

- *Chapter Eleven Study Guide*
- *Chapter Eleven Concept Guide*
- *Solution Calculations (handout)*
- *Important Equations (following this slide)*
- *End of Chapter Problems (following slide)*

when someone asks you to elevate the boiling point



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Important Equations, Constants, and Handouts from this Chapter:

- **solution = solute + solvent**
- **see Concentration Units Handout**

Solution Concentrations:

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

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

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

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

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

Henry's Law:

$$S_g = k \bullet P_g$$

Raoult's Law / Vapor Pressure Depression:

$$P_{solvent} = \chi_{solvent} \bullet P^o_{solvent}$$

Boiling Point Elevation / Freezing Point Depression:

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

Osmosis:

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

R = 0.082057 L atm mol^{-1} K^{-1}
i = van't Hoff factor

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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_2H_5OH	_____	5	_____
$C_{12}H_{22}O_{11}$	_____	_____	0.0027

2. 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.
3. A soda can has an aqueous CO_2 concentration of 0.0506 M. What is the pressure of CO_2 gas in the can? ($k_H = 4.48 \times 10^{-5}$ M/mm Hg)
4. Pure nonvolatile iodine (105 g) is dissolved in 325 g of CCl_4 at 65 °C. Given that the vapor pressure of CCl_4 at this temperature is 531 mm Hg, what is the vapor pressure of the solution at 65 °C?
5. What is the boiling point of a solution composed of 15.0 g of $CHCl_3$ (which boils at 61.70 °C) and 0.515 g of the nonvolatile solute acenaphthene, $C_{12}H_{10}$?
6. 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	2.20	.0027
C_2H_5OH	1.1	5.00	.020
$C_{12}H_{22}O_{11}$	0.15	4.90	.0027

2. 16.2 m, 37.1%
3. 1130 mm Hg
4. 444 mm Hg
5. 62.51 °C
6. 6.0×10^3 g/mol

MAR