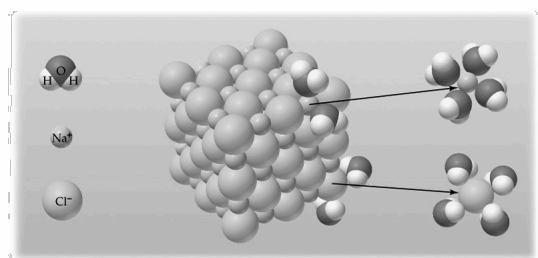


## Chapter 11: Solutions and Colloids



Chemistry 222  
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<http://mhchem.org/222>

Last update:  
7/7/25

MAR

### 11.1 - The Dissolution Process

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

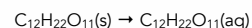
**Solutions, Colloids and Suspensions** are all examples of homogeneous mixtures

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

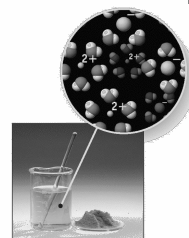
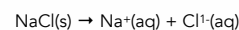
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A **solution** has a **solute** (substance being dissolved) and a **solvent** (substance dissolving the solute)

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



Ionic molecules that dissolve **dissociate**:



### 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.



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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)

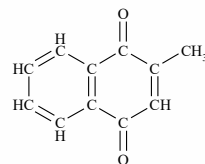
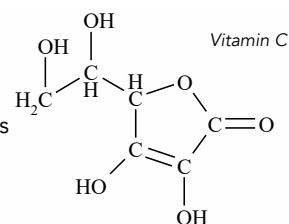
Ex: **oil** (ID-ID) dissolves in **kerosene** (ID-ID)

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

### Soluble Examples

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

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



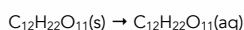
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Example: is **Vitamin K3** water soluble?

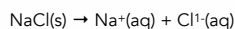
**No** - most of molecule nonpolar, C=O groups cancel, so Vitamin K3 is not water soluble (Vitamin K3 is fat soluble)

### 11.2 - Electrolytes

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

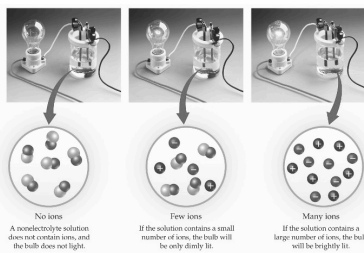


**Ionic** molecules that dissolve **dissociate**:



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

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**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*)

### Strong Electrolytes

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

**Five strong acids:**

HCl

HBBr

HI

HNO<sub>3</sub>

HClO<sub>4</sub>

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hydrochloric acid

hydrobromic acid

hydroiodic acid

nitric acid

perchloric acid

SOLUBLE COMPOUNDS	
Almost all salts of $\text{Na}^+$ , $\text{K}^+$ , $\text{NH}_4^+$	Solubility guide: <a href="http://mhchem.org/sol/">http://mhchem.org/sol/</a>
Salts of nitrate, $\text{NO}_3^-$ ; chlorate, $\text{ClO}_3^-$ ; perchlorate, $\text{ClO}_4^-$ ; acetate, $\text{CH}_3\text{CO}_2^-$	
Almost all salts of $\text{Cl}^-$ , $\text{Br}^-$ , $\text{I}^-$	
Compounds containing $\text{F}^-$	
EXCEPTIONS	
Halides of $\text{Ag}^+$ , $\text{Hg}_2^{2+}$ , $\text{Pb}^{2+}$	SOLUBLE COMPOUNDS
Fluorides of $\text{Mg}^{2+}$ , $\text{Ca}^{2+}$ , $\text{Sr}^{2+}$ , $\text{Ba}^{2+}$ , $\text{Pb}^{2+}$	
Salts of sulfate, $\text{SO}_4^{2-}$	
Sulfates of $\text{Ca}^{2+}$ , $\text{Sr}^{2+}$ , $\text{Ba}^{2+}$ , $\text{Pb}^{2+}$	
EXCEPTIONS	
Most salts of carbonate, $\text{CO}_3^{2-}$ ; phosphate, $\text{PO}_4^{3-}$ ; sulfide, $\text{S}^{2-}$ ; chromate, $\text{CrO}_4^{2-}$	INSOLUBLE COMPOUNDS
Salts of $\text{NH}_4^+$ and the alkali metal cations	
Most metal sulfides, $\text{S}^{2-}$	
Most metal hydrides and oxides	
$\text{Ba}(\text{OH})_2$ is soluble	

**Three strong bases:**

NaOH

KOH

LiOH

sodium hydroxide

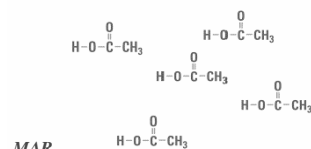
potassium hydroxide

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:

$\text{CH}_3\text{CO}_2\text{H}$	acetic acid
$\text{H}_2\text{CO}_3$	carbonic acid
$\text{H}_3\text{PO}_4$	phosphoric acid
$\text{HF}$	hydrofluoric acid
$\text{NH}_3$	ammonia (base)
$\text{C}_5\text{H}_5\text{N}$	pyridine (base)
$\text{C}_2\text{H}_5\text{NH}_2$	ethylamine (base)
$\text{N}_2\text{H}_4$	hydrazine (base)

plus many more!

## 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**.



**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

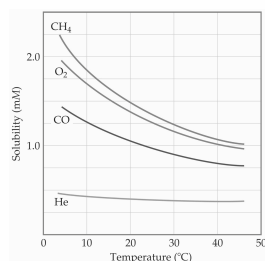
Gases usually more soluble in liquids at lower temperatures

Carbonated soft drinks are more "bubbly" if stored in the refrigerator; warm lakes have less dissolved  $\text{O}_2$  than cool lakes.

The solubility of a gas in a liquid dictated by **Henry's Law**:

MAR  $S_g = kP_g$

$S_g$  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

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

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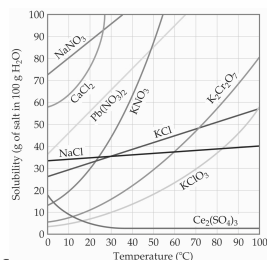


$k$  for  $\text{O}_2 = 1.66 \times 10^{-6} \text{ M/mm Hg}$

When  $P_g$  drops, solubility drops.

## Other Solutions

**Solids in liquids** generally become more soluble as temperature increases



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



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

oil and water, immiscible



## 11.4 - Colligative Properties

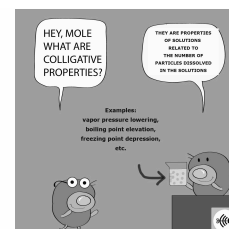
On adding a **solute** to a **solvent**, the **solvent** properties are modified.

- 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

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

## Mass Percent:

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

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Molality,  $m$ 

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

see [Solutions Calculations](#) handout

## Concentration Calculations

Dissolve 62.1 g (1.00 mol) of ethylene glycol in 250. g of  $\text{H}_2\text{O}$ . Calculate mol fraction, molality, and mass % of ethylene glycol.

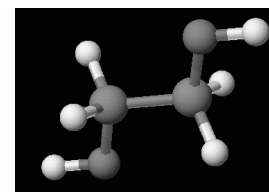
$$250. \text{ g } \text{H}_2\text{O} = 13.9 \text{ mol}$$

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

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

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

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ethylene glycol

Concentration Calculations (*continued*)

Dissolve 62.1 g (1.00 mol) of ethylene glycol in 250. g of  $\text{H}_2\text{O}$ . Calculate mol fraction, molality, and mass % of ethylene glycol.

Calculate molality

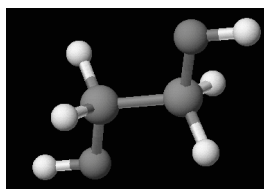
$$\text{molality} = \frac{1.00 \text{ mol glycol}}{0.250 \text{ kg } \text{H}_2\text{O}} = 4.00 \text{ m}$$

Calculate mass%

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

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



ethylene glycol

## Concentration Calculations and Molarity

Converting between molality ( $m$ ), mole fraction ( $X$ ) 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  $\text{g/mL}$  or  $\text{g/cm}^3$ ) 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  $X$ , mass% and  $M$  if the solution density =  $1.33 \text{ g/cm}^3$ .

Solution:

Assume 1 kg solvent and 10.7 mol solute

1 kg water = 1000 g  $\text{H}_2\text{O}$  = 55.5 mol  $\text{H}_2\text{O}$ 10.7 mol NaOH \* 40.0 g  $\text{mol}^{-1}$  = 428 g NaOH

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

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

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

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

Concentration Calculations and Molarity (*continued*)

Example: Convert a 10.7 m aqueous NaOH solution to  $X$ , mass% and  $M$  if the solution density =  $1.33 \text{ g/cm}^3$ .

Solution (*continued*):We have 1000 g of solvent ( $\text{H}_2\text{O}$ ) and 428 g of solute (NaOH) or1428 g of solution (NaOH +  $\text{H}_2\text{O}$ )

$$\text{Volume of solution} = 1428 \text{ g} * (\text{cm}^3 / 1.33 \text{ g}) = 1070 \text{ mL} = 1.07 \text{ L}$$

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

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

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

### 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 ( $\chi$ ):** assume 1 mol total;  $\chi$  = mol solute, and  $(1 - \chi)$  = mol solvent
- **If starting with mass%:** assume 100 g total, then mass% = g solute and  $100 - \text{mass\%} = \text{g solvent}$
- **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.

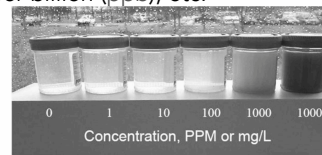
**1.0 ppm = 1.0 g of a substance in a sample with a total mass of 1.0 million ( $10^6$ ) g**

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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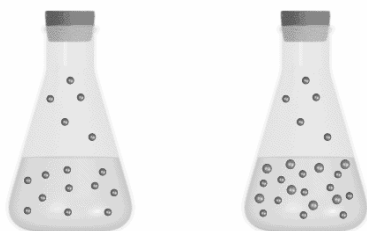
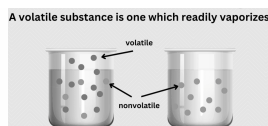


### Colligative Properties

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)

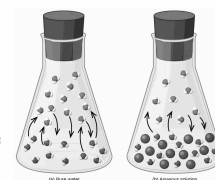
Vapor Pressure of  $\text{H}_2\text{O}$  over a solution ( $P_{\text{solvent}}$ ) depends on the number of  $\text{H}_2\text{O}$  molecules per solute molecule ( $X_{\text{solvent}}$ )

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

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

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

Vapor Pressure of *solvent* over *solution* ( $P_{\text{solvent}}$ ) = (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.  $X_{\text{glycol}} = 0.0672$  (earlier example) What is the vapor pressure of water in the solution at 30 °C? (The vapor pressure of pure  $\text{H}_2\text{O}$  is 31.8 mm Hg)**

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

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

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 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:

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

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

**Solution:** This is a two component system.

If  $X_{\text{heptane}} = 0.15$ , then

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

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

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

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

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The total vapor pressure in the mixture is 14.1 torr

## Changes to Boiling & Freezing Points

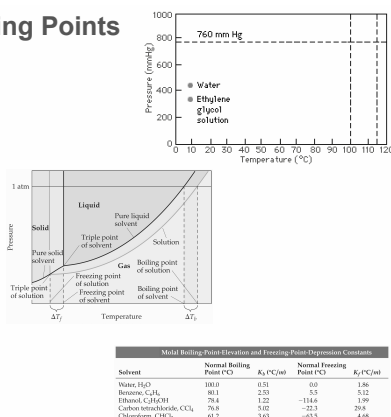
Solutes block solvent particles from evaporating (**boiling**) and condensing (**freezing**) in solutions

The change to the boiling or freezing point depicted as  $\Delta T$  (always positive)

$\Delta T_b$  = difference between boiling point pure solvent and solution

$\Delta T_f$  = difference between freezing point pure solvent and solution

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## Boiling Point Elevation

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

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

$\Delta T_b$  = difference between boiling point pure solvent and solution

$K_b$  = boiling point constant for the solution

$m$  = molality of the solution

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

$K_b = 0.512^\circ\text{C/molal}$  for water

Solution

1. Calculate molality = 4.00 m

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

$$\Delta T_b = 0.512^\circ\text{C/molal} (4.00 \text{ m})$$

$$\Delta T_b = 2.05^\circ\text{C}$$

$$\text{Boiling point} = 100.00 + 2.05$$

$$\text{Boiling point} = 102.05^\circ\text{C}$$

## Freezing Point Lowering

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

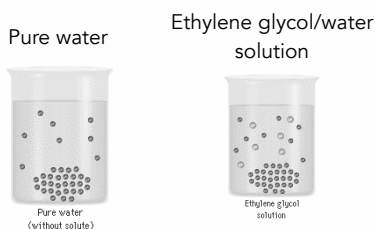
$$\Delta T_f = K_f \cdot m$$

$\Delta T_f$  = difference between freezing point pure solvent and solution

$K_f$  = freezing point constant for the solution

$m$  = molality of the solution

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

Notice how the animations incorrectly use "dense" ice!

## Freezing Point Lowering - Example

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

$$K_f = 1.86^\circ\text{C/molal}$$

Solution

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

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

$$\Delta T_f = 7.44^\circ\text{C}$$

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

Many textbooks use negative  $K_f$  values. If using negative  $K_f$ ,  $\Delta T_f$  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:

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

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  $\text{Na}^+$  and  $\text{Cl}^-$ , two particles per NaCl

$i = 3$  for  $\text{MgCl}_2$  ( $\text{Mg}^{2+}$  and 2  $\text{Cl}^-$ )

$i = 4$  for  $\text{AlCl}_3$  ( $\text{Al}^{3+}$  and 3  $\text{Cl}^-$ )

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Example: How much NaCl must be dissolved in 4.00 kg of water to lower FP to  $-10.00^\circ\text{C}$ ? ( $K_f = 1.86^\circ\text{C/m}$ )

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

$$\Delta T_f = 10.00^\circ\text{C} \ (\Delta T \text{ always positive})$$

$$i = 2 \text{ for NaCl (Na}^+ \text{ and Cl}^-)$$

$$m = \Delta T_f / K_f \cdot i = 10.00 / 1.86 \cdot 2$$

$$m = 2.68817... \text{ mol NaCl / kg (* 4.00 kg)}$$

$$= 10.7526... \text{ mol NaCl * 58.44 g/mol}$$

$$= 628 \text{ g NaCl}$$

## Colligative Properties and Molar Mass

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

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

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

Solution:

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

$$i = 1 \text{ (covalent)}$$

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

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

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

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

$$\text{MM} = (\text{grams solute} \cdot K) / (\Delta T \cdot \text{kg solvent})$$

useful for both bp and fp calculations



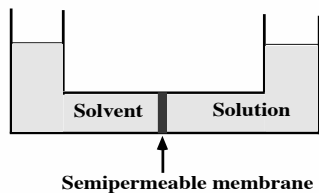
## Osmosis / Osmotic Pressure

**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

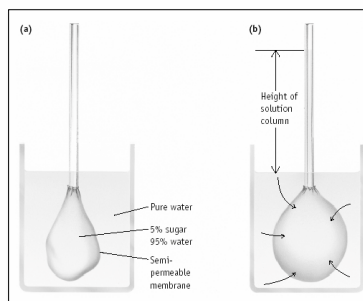
**Net result:** solvent molecules travel from dilute concentration to more concentrated solution

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

## Osmosis



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

## Osmotic Pressure, $\pi$

Equilibrium is reached when pressure - the **Osmotic Pressure,  $\pi$**  - produced by extra solution counterbalances pressure of solvent molecules moving through the membrane.

$$\pi = i c R T$$

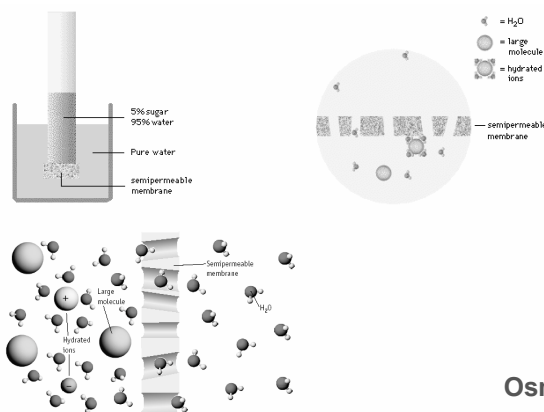
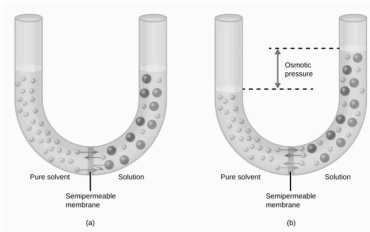
$c$  = concentration in mol/L

$i$  = the van't Hoff factor

$R = 0.082057 \text{ (L atm)/(mol K)}$

$T$  = Kelvin Temperature

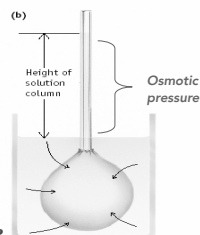
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Osmosis

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**Example:** Dissolve 35.0 g of hemoglobin in enough water to make 1.00 L of solution.  $\pi$  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.  $\pi$  in atmospheres  
 $\pi = 10.0 \text{ mm Hg} \cdot (1 \text{ atm} / 760 \text{ mm Hg})$   
 $\pi = 0.0132 \text{ atm}$

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

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

(c) Calc. molar mass  
 $\text{Conc} = 5.40 \times 10^{-4} \text{ mol/L} (* 1 \text{ L})$   
 $\text{Molar mass} = 35.0 \text{ g} / 5.40 \times 10^{-4} \text{ mol}$   
 $\text{Molar mass} = 64,800 \text{ g/mol}$

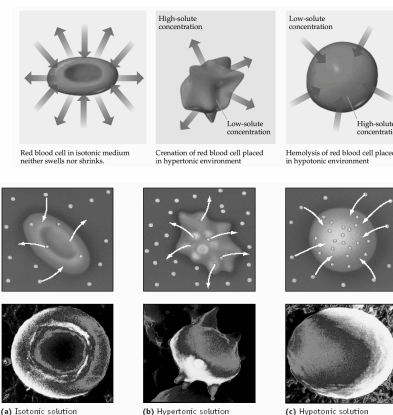
## Osmosis

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

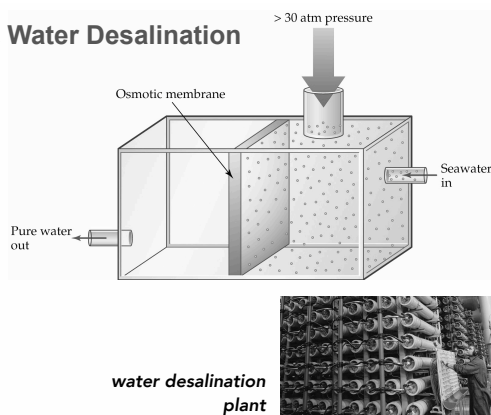


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## Reverse Osmosis & Water Desalination

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



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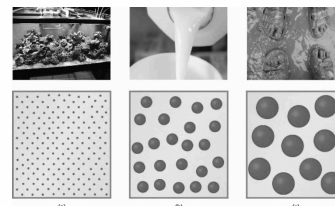
## 11.5 - Colloids (and Suspensions)

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.)



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## Colloids

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

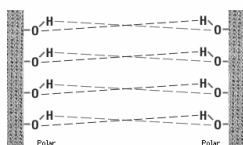
Type	D Medium	D Phase	Examples
aerosol	gas	liquid	fog, clouds
aerosol	gas	solid	smoke, viruses
foam	liquid	gas	shaving cream
emulsion	liquid	liquid	milk, mayonnaise
sol	liquid	solid	mud, paint, ink
foam	solid	gas	foam rubber, sponge
gel	solid	liquid	jelly, cheese, yogurt
solid sol	solid	solid	steel,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

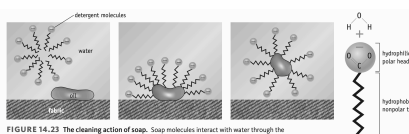


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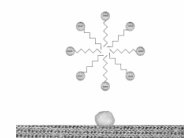
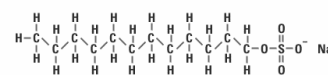
## Soaps

**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.



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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

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:

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

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

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

Solution Concentrations:

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

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

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

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

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

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## 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 <sub>2</sub> H <sub>5</sub> OH	_____	5	_____
C <sub>12</sub> H <sub>10</sub> O <sub>4</sub>	_____	_____	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/cm<sup>3</sup>, calculate the molality and the weight percent of HCl in the solution.
- A soda can has an aqueous CO<sub>2</sub> concentration of 0.0506 M. What is the pressure of CO<sub>2</sub> 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 CCl<sub>4</sub> at 65 °C. Given that the vapor pressure of CCl<sub>4</sub> 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 CHCl<sub>3</sub> (which boils at 61.70 °C) and 0.515 g of the nonvolatile solute acenaphthene, C<sub>12</sub>H<sub>10</sub>?
- 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 <sub>2</sub> H <sub>5</sub> OH	1.1	5.00	.020
C <sub>12</sub> H <sub>10</sub> O <sub>4</sub>	0.15	4.90	.0027

- 16.2 m, 37.1%
- 1130 mm Hg
- 444 mm Hg
- 62.51 °C
- 6.0 x 10<sup>3</sup> g/mol

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