Announcements

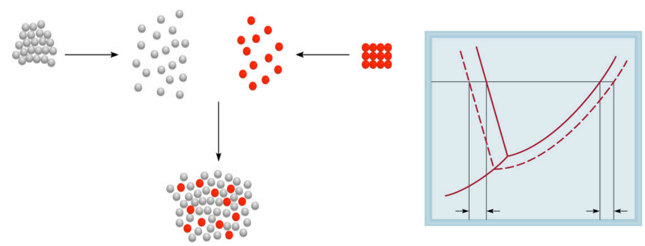
Chem 11 Exam 1
December 18
6PM - 7:30PM
Chapter 12, 13 & 16
Format (40MC + 2 long questions)

PS/recitation classes for all the CH11 sections:

A (Chem majors) - 4:30-6:00 pm Wednesday, C114 (Val Miclat)
B/C - 1:30-2:30 pm Friday, C114 (Albert Chen)
D/E - 11:00 am-12:00 pm Thursday, G304 (Albert Chen)
F - 11:00 am-12:00 pm Tuesday, G304 (Albert Chen)
G - 12:30-1:30 pm Monday, C114 (Albert Chen)

Chapter 13: Start problems in Chapter 13

Chapter 13 The Properties of Mixtures: Solutions and Colloids



The Properties of Mixtures: Solutions and Colloids

13.1 Types of Solutions: IMF's and Solubility

13.2 IMF's and Biological Macromolecules

13.3 Why Substances Dissolve (Chapter 20)

13.4 Solubility as an Equilibrium Process

13.5 Quantitative Ways of Expressing Concentration

13.6 Colligative Properties of Solutions

13.7 The Structure and Properties of Colloids

A *solution* is a homogenous mixture of a *solute* and a *solvent* that has recognizable properties.

Solute = substance(s) present in the smaller amount(s)

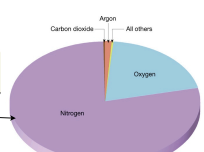
Solvent = substance present in the largest amount in which the solute dissolves.

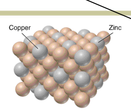
Mass of Solute + Mass of Solvent = Mass Solution

- homogeneous uniform composition, typically transparent
- Contain soluble particles
- Does not separate on standing;
- Colloids and Suspensions are larger particles that can be separated physically and which scatter light

Any phase of matter can be a *solute* in another more abundant phase (*solvent*).

Kind of Solution	Example
Gas in gas	Air (O_2, N_2, Ar, and other gases)
Gas in liquid	Carbonated water (CO_2 in water)
Gas in solid	H_2 in palladium metal
Liquid in liquid	Gasoline (mixture of hydrocarbons)
Liquid in solid	Dental amalgam (mercury in silver)
Solid in liquid	Seawater (NaCl and other salts in water)
Solid in solid	Metal alloys, such as sterling silver (92.5% Ag, 7.5% Cu)

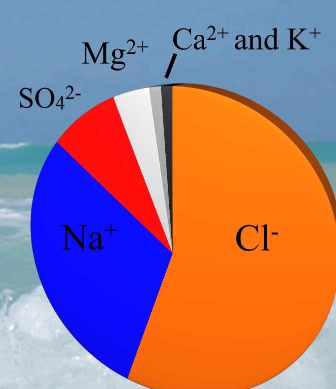




Solutes in Seawater (seawater = solution)

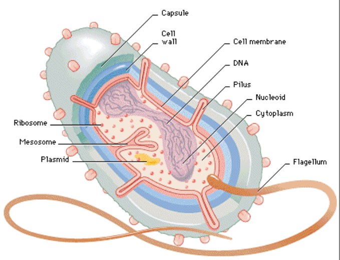
70+ dissolved components but 6 make up >99%

35 grams of dissolved salts per kilogram of seawater

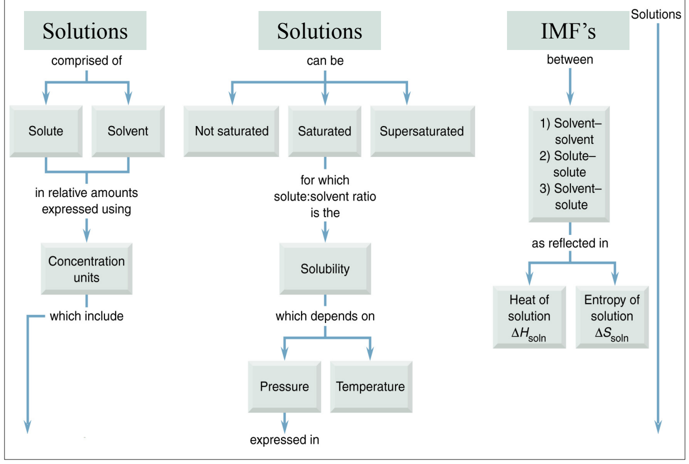


Composition of A Bacterium

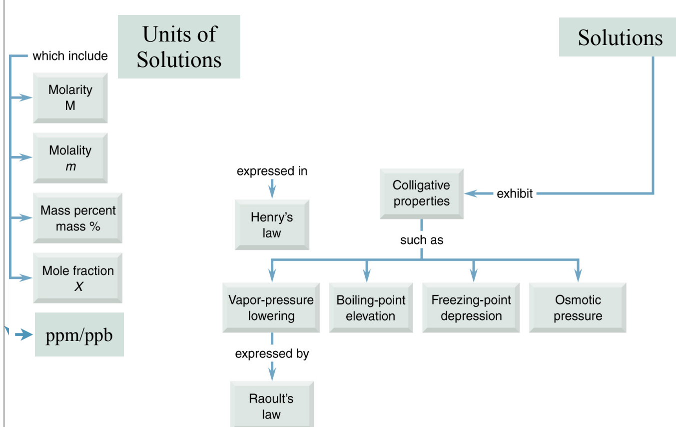
Substance	% Mass of Cell	Types	Number of
Water	70	1	5×10^{10}
Ions	1	20	?
Sugars	3	200	3×10^{8}
Amino Acids	0.4	100	5×10^{7}
Lipids	2	50	3×10^{7}
Nucleotides	0.4	200	1×10^{7}
Small molecules	0.2	~200	?
Large Molecules	23	~5000	6×10^{6}



Solutions



Solutions



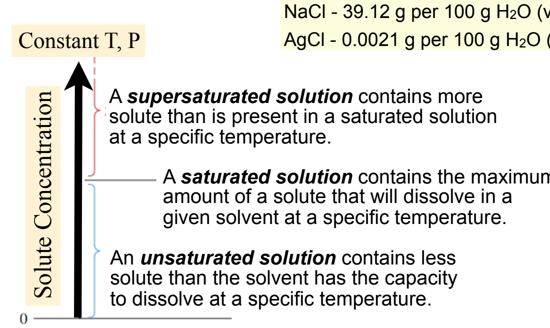
Solubility (S) **is the maximum amount of a solute that can completely dissolve in a fixed quantity of a solvent at a specified temperature. (Units of g solute/100 g water)**

For example:

Sucrose (sugar) - 203 g per 100 g H_2O

NaCl - 39.12 g per 100 g H_2O (very soluble)

AgCl - 0.0021 g per 100 g H_2O (insoluble)



A ***supersaturated solution*** contains more solute than is present in a saturated solution at a specific temperature.

A ***saturated solution*** contains the maximum amount of a solute that will dissolve in a given solvent at a specific temperature.

An ***unsaturated solution*** contains less solute than the solvent has the capacity to dissolve at a specific temperature.

Sodium acetate crystallizing from a supersaturated solution.

A small crystal of solid sodium acetate

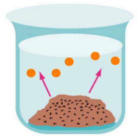
Saturated Solution of sodium acetate

Spontaneous crystallization of the saturated solution.

Science often uses confusing terminology.

Term	Parts of Solvent Required for 1 part of Solute
Very soluble	Less than 1 part
Freely soluble	1-10 parts
Soluble	10-30 parts
Sparingly Soluble	30-100 parts
Slightly soluble	100-1000 parts
Very slightly soluble	1000-10,000 parts
Insoluble	> 10,000 parts

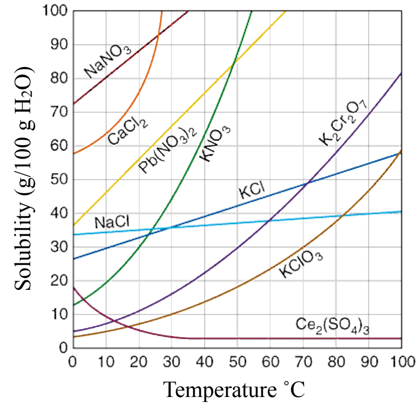
The dissolution process in a solvent is an *equilibrium process*.



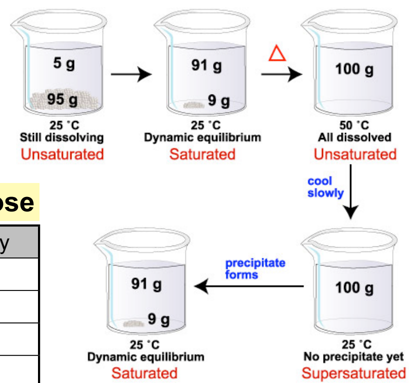
At saturation:

Solute + Solvent $\underset{\text{Crystallize}}{\overset{\text{Dissolve}}{\rightleftharpoons}}$ Solution

The *solubility of all substances depends on temperature*----typically increasing as temperature increases.



Solubility of Glucose vs Temperature



Solubility of Glucose

Temperature	Solubility
25˚C	91
30˚C	125
50˚C	244
70˚C	357
90˚C	556

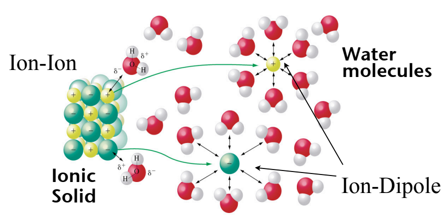
Science uses three different ways to help us understand how solutes dissolve in solvents.

1. **Practical Answer** (the one that comes in handy in real-life)
2. **Macroscopic Answer** (thermodynamic approach--cool if you like thermo---no need for molecules). COVER THIS IN CHAPTER 20 SECTION 13.2.
3. **Microscopic Answer** (molecular approach using IMF's--helps one understand the practical)

2. Molecular Focus: The interplay of IMF's between solute and solvent molecules determines a solute's solubility.

- weak solvent-solvent forces favors solubility
- weak solute-solute forces favors solubility
- strong solvent-solute interaction forces favor solubility

Dissolution will only occur if ion-dipole forces are stronger than ion-ion forces



The "likes dissolves likes" principles is explained by considering whether solute-solvent forces may exceed solute-solute forces.

1. Practical Answer: "Likes dissolve likes". This means:

Forces between solute and solvent are comparable to IMF's between solute particles.

- Non-polar molecules are soluble in non-polar solvents.

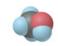
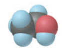
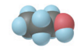
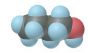
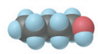
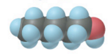
  CCl_4, gasoline, kerosene, hexane, oils, are soluble (miscible) in non-polar solvent benzene C_6H_6

- Polar molecules and ionic compounds are soluble in polar solvents

  C_2H_5OH or NH_3 or CH_3COOH are soluble in H_2O

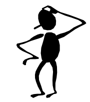
Note how the solubility decreases in a polar solvent water as the carbon chain grows.

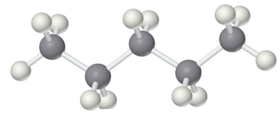
Table 13.2 Solubility* of a Series of Alcohols in Water and Hexane

Alcohol	Model	Solubility in Water	Solubility in Hexane
CH_3OH (methanol)		∞	1.2
CH_3CH_2OH (ethanol)		∞	∞
$CH_3(CH_2)_2OH$ (1-propanol)		∞	∞
$CH_3(CH_2)_3OH$ (1-butanol)		1.1	∞
$CH_3(CH_2)_4OH$ (1-pentanol)		0.30	∞
$CH_3(CH_2)_5OH$ (1-hexanol)		0.058	∞

*Expressed in mol alcohol/1000 g solvent at 20°C.

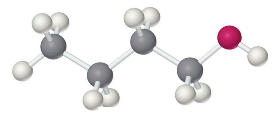
Pentane, C_5H_{12} and 1-butanol, C_4H_9OH have similar molecular masses. Which of the two would be more soluble in water and why?





Pentane

C_5H_{12}



1-Butanol

C_4H_9OH

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For sections D/E and F, since the only available room is in Gonzaga, I will have to write the Fine Arts Department first to ask for permission to use the room before the reservation can be finalized. All room reservations will be finalized by next week.

Problems Chapter 13:
See Blog

Predict which solvent will dissolve more of the given solute:

(a) Sodium chloride in methanol (CH_3OH) or in propanol ($CH_3CH_2CH_2OH$)

(b) Ethylene glycol ($HOCH_2CH_2OH$) in hexane ($CH_3CH_2CH_2CH_2CH_2CH_3$) or in water.

(c) Diethyl ether ($CH_3CH_2OCH_2CH_3$) in water or in ethanol (CH_3CH_2OH)

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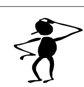

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PLAN: Consider the intermolecular forces which can exist between solute molecules and consider whether the solvent can provide such interactions and thereby substitute.

SOLUTION:

(a) Methanol - NaCl is ionic and will form ion-dipoles with the -OH groups of both methanol and propanol. However, propanol is subject to the dispersion forces to a greater extent.

(b) Water - Hexane has no dipoles to interact with the -OH groups in ethylene glycol. Water can H bond to the ethylene glycol.

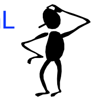
(c) Ethanol - Diethyl ether can interact through a dipole and dispersion forces. Ethanol can provide both while water would like to H bond.

As with the ideal gas law, we have an idealized and non-ideal situations when discussing liquid-phase phenomena.

✳ ***IDEAL SOLUTIONS***: Substances dissolve when when IMF's between a solute and solvent are of the same type.

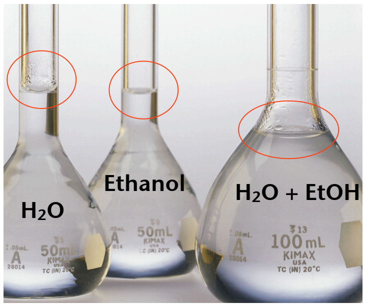
✳ ***NON-IDEAL SOLUTIONS***: occurs when IMF's **between *solute or between solvent exceed those IMF's between solute-solvent*** then properties are hard to predict.

Non-ideal solutions lead to non-intuitive results.



What is the volume of solution if we take 50.00 mL water and add it to 50.00 mL of ethanol?

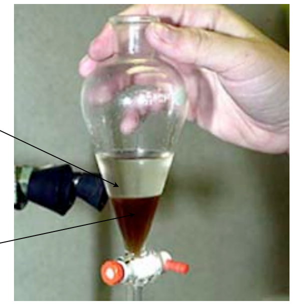
IMF's alter the combined volume of water and ethanol. The system is non-ideal.



Liquids are ***miscible*** if they mix to an appreciable extent. ***Immiscible liquids*** separate into different layers or phases.

Lowest density liquid rises to the top

Highest density liquid sinks to bottom



We can use miscibility as a means to separate substances that are soluble in polar and non-polar solvents.

Predict whether each of the following substances is more likely to dissolve in carbon tetrachloride or in water.

(a) heptane C_7H_{16}
(b) sodium sulfate
(c) HCl(g)
(d) I_2

Predict whether each of the following substances is more likely to dissolve in carbon tetrachloride or in water.

(a) heptane C_7H_{16} -> no polar bonds -->no dipole --> non-polar --> LDF => CCl_4
(b) sodium sulfate ------ ion-ion => polar => water

(c) HCl(g) ----polar HCl => water water

(d) I_2 CCl_4 ----- LDF => non-polar => CCl_4

Arrange the following substances in order of increasing solubility in water.

(a) $CH_3CH_2CH_2CH_2CH_3$
(b) $OH\text{-}CH_2CH_2CH_2CH_2CH_2OH$
(c) $CH_3CH_2CH_2CH_2CH_2OH$
(d) $CH_3CH_2CH_2CH_2CH_2Cl$

What IMF forces are present in each?

Which compound should have the lowest surface tension?

Which compound should hydrogen bond?

Arrange the following substances in order of increasing solubility in water.

(a) $CH_3CH_2CH_2CH_2CH_3$
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(c) $CH_3CH_2CH_2CH_2CH_2OH$
(d) $CH_3CH_2CH_2CH_2CH_2Cl$

$a < d < c < b$

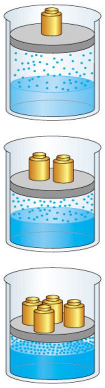
Which compound should have the lowest surface tension? A

What IMF forces are present in each? A=>LDF

B,C=> LDF, HB, D=>LDF, DD

Which compound should hydrogen bond? B,C

The ***solubility of a gas*** **in a liquid depends on the gas, the liquid (solvent), the pressure and temperature.**



1. A gas (say CO_2) is put over the liquid and sealed.

2. Increasing the external pressure increases the solubility of the gas in the liquid.

3. Container is sealed and under pressure until opened.

4. It is observed that the concentration of gas that dissolves at some specified temperature T is proportional to the gas partial pressure above the liquid.

HENRY'S LAW

$$C_{gas} \propto P_{gas}$$

$$C_{gas} = k_H P_{gas}$$

Henry's law: **At constant temperature, the solubility of a gas (C_{gas}) in a liquid is proportional to the partial pressure of the gas over the solution.**

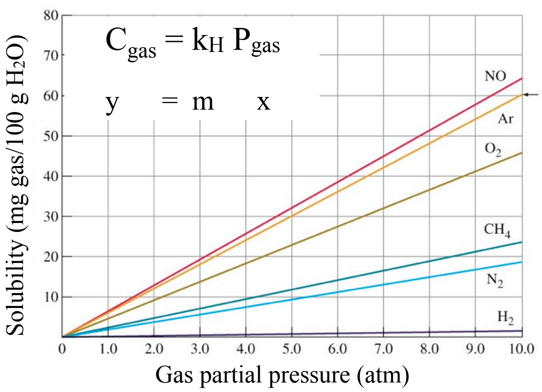
$$C_{gas} = k_H P_{gas}$$

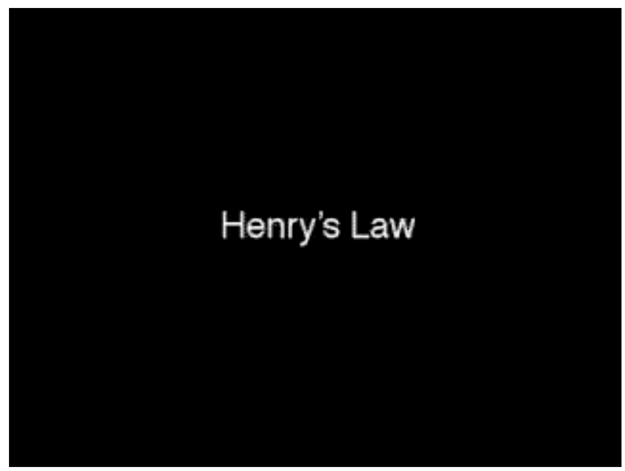
Molarity of the dissolved gas in solution.

a constant (mol/L•atm) that depends only on T, substance and the solvent.

P is the partial vapor pressure **of the gas** above the solution.

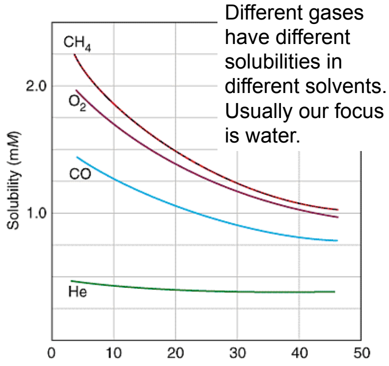
If we plot a gas's solubility vs *gas partial pressure* over the solvent liquid we get a straight line. The slope is the the Henry Law constant for that gas.





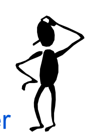
The *solubility of a gas* in a solvent *depends on* the nature of the gas, temperature, the nature of the solvent and *the pressure of the same gas above the solution*.

Solubility of a gas decreases with increasing temperature.



Different gases have different solubilities in different solvents. Usually our focus is water.

Given that the partial pressure of O_2 in the atmosphere is 0.21 atm and the Henry constant @ 25˚C for O_2 in water is 1.3 x 10^{-3} mol/L atm, determine the concentration of O_2 in a fresh water stream in equilibrium with air at 25˚C and 1 atm. Express the answer in ppm (mg/L).



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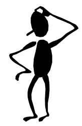


$$C_{gas} = k_H P_{gas}$$

$$C_{O_2} = 1.3 \times 10^{-3} \frac{mol\ O_2}{L\ atm} \times 0.21\ atm = 2.7 \times 10^{-4}\ M$$

Now convert Molarity to ppm. Recall ppm means in this case mg/L.

$$ppm\ O_2 = 2.73 \times 10^{-4} \frac{mol\ O_2}{L} \times \frac{31.98\ g\ O_2}{1\ mol\ O_2} \times \frac{10^3\ mg}{1\ g} = 8.7\ ppm$$

The partial vapor pressure of CO_2 gas inside a bottle of liquid Coke is 4 atm at 25˚C. What is the solubility of CO_2 at this pressure, and also when the cap on the coke is removed in ppm? The Henry Law constant k for CO_2 in water is 3.3 X 10^{-2} mol/L atm at 25˚C, and the partial pressure of CO_2 in the atmosphere is 0.00033 atm.

The partial vapor pressure of CO_2 gas inside a bottle of liquid Coke is 4 atm at 25˚C. What is the solubility of CO_2 at this pressure and also when the cap on the coke is removed in ppm? The Henry Law Constant k for CO_2 in water is 3.3 X 10^{-2} mol/L atm at 25˚C, and the partial pressure of CO_2 in the atmosphere is 0.00033 atm.

$$C_{CO2} = k_H P_{CO2}$$

$$C_{CO_2} = 3.3 \times 10^{-2} \frac{mol}{L\ atm} \times 4\ atm = 0.1 \frac{mol}{L}$$

$$ppm\ CO2 = \frac{0.1\ mol\ CO_2}{L} \times \frac{44.01\ CO_2}{1\ mol\ CO_2} \times \frac{10^3\ mg}{1\ g} = 4401\ ppm = 4000\ ppm$$

$$C_{CO_2} = 3.3 \times 10^{-2} \frac{mol}{L\ atm} \times 0.00033\ atm = 1.1 \times 10^{-5} \frac{mol}{L}$$

$$ppm\ CO2 = \frac{1.1 \times 10^{-5}\ mol\ CO_2}{L} \times \frac{44.01\ CO_2}{1\ mol\ CO_2} \times \frac{10^3\ mg}{1\ g} = .48\ ppm$$

Chemists use different definitions to quantify the concentration of solutes in solutions.

Molarity *(M)* = $\frac{\text{moles solute}}{\text{liters solution}}$

Issue: volume changes with temperature and therefore so does molarity.

Molality *(m)* = $\frac{\text{moles solute}}{\text{kg solvent}}$

mass does not change with temperature

KEY: mass of solute + mass solvent = mass solution

More Concentration Units

% by mass = $\frac{\text{mass solute}}{\text{mass of solute + mass of solvent}}$ x 100%

= $\frac{\text{mass of solute}}{\text{mass of solution}}$ x 100%

% by volume = $\frac{\text{volume of solute}}{\text{volume of solute + volume of solvent}}$ x 100%

= $\frac{\text{volume of solute}}{\text{volume of solution}}$ x 100%

% Mass/Volume = $\frac{\text{mass of solute}}{\text{volume of solution}}$ x 100%

More Concentration Units

Mole Fraction *(X_A)* = $\frac{\text{moles of A}}{\text{sum of moles of all components}}$

Mole % *(X_A)* = $\frac{\text{moles of A}}{\text{sum of moles of all components}}$ x 100%

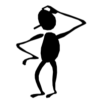
Low solute concentrations **are reported as** ***"parts per thousand--ppt"*** **or** ***"parts per million--ppm"*** **or** ***"parts per billion--ppb"*** **typically by weight.....but sometimes m/v.**

ppt = 1 gram solute in 10^3 grams of solution = 1 mg/g
ppm = 1 gram solute in 10^6 grams of solution = 1 µg/g
ppb = 1 gram solute in 10^9 grams of solution = 1 ng/g

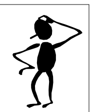
If the solvent is water we can use the density of water 1 g/ml as a conversion factor.

1 ppm = 1 gram in 10^6 ml of water = 1 mg/L
1 ppb = 1 gram in 10^9 ml of water = 1 µg/L

What is the molality of a solution prepared by dissolving 32.0 g of $CaCl_2$ in 271 g of water?



What is the molality of a solution prepared by dissolving 32.0 g of $CaCl_2$ in 271 g of water?

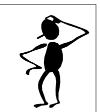


SOLUTION:

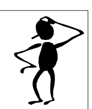
$$32.0 \text{ g } CaCl_2 \times \frac{1 \text{ mol } CaCl_2}{110.98 \text{ g } CaCl_2} = 0.288 \text{ mol } CaCl_2$$

$$\text{molality} = \frac{0.288 \text{ mol } CaCl_2}{271 \text{ g } H_2O \times \frac{\text{kg}}{10^3 \text{ g}}} = 1.06\ m\ CaCl_2$$

What is the molality of a 5.86 *M* ethanol (C_2H_5OH) solution whose density is 0.927 g/mL (MM C_2H_5OH = 46.1 g/mol)?



What is the molality of a 5.86 *M* aqueous ethanol (C_2H_5OH) solution having a density of 0.927 g/mL (MM EtOH = 46.1 g/mol)?



$$m = \frac{\text{moles of solute}}{\text{mass of solvent (kg)}} \qquad M = \frac{\text{moles of solute}}{\text{liters of solution}}$$

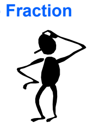
mass of solvent = mass of solution – mass of solute

Assume 1 L of 5.86M ethanol solution:
g ethanol = 5.86 mol x 46.1 g EtOH/mol = 270 g ethanol
g EtOH solution in 1 Liter = 1000 mL x 0.927 g/mL = 927 g

mass of solvent = mass of solution – mass of solute
= 927 g – 270 g = 657 g = 0.657 kg

$$m = \frac{\text{moles of solute}}{\text{mass of solvent (kg)}} = \frac{5.86 \text{ moles } C_2H_5OH}{0.657 \text{ kg solvent}} = 8.92\ m$$

Expressing Concentration in Parts by Mass, Parts by Volume, and Mole Fraction



(a) Find the concentration of calcium (in ppm) in a 3.50-g pill that contains 40.5 mg of Ca.

(b) The label on a 0.750-L bottle of Italian chianti indicates "11.5% alcohol by volume." How many liters of alcohol does the wine contain?

(c) A solution of rubbing alcohol contains 142 g of isopropyl alcohol (C_3H_7OH) and 58.0 g of water. What are the mole fractions of alcohol and water?

Expressing Concentration in Parts by Mass, Parts by Volume, and Mole Fraction

(a) Find the concentration of calcium (in ppm) in a 3.50-g pill that contains 40.5 mg of Ca.

$$\text{ppm} = \frac{\text{g Ca}}{\text{g pill}} \times 10^6 = \frac{40.5 \text{ mg Ca}}{3.5 \text{ g} \times \frac{10^3 \text{ mg}}{\text{g}}} \times 10^6 = 1.16\text{x}10^4 \text{ ppm Ca}$$

(b) The label on a 0.750-L bottle of Italian chianti indicates "11.5% alcohol by volume (v/v)." How many liters of alcohol does the wine contain?

$$\text{L EtOH} = 0.750 \text{ L chianti} \times \frac{11.5 \text{ L EtOH}}{100. \text{ L chianti}} = 0.0862 \text{ L}$$

Expressing Concentration in Parts by Mass, Parts by Volume, and Mole Fraction

(c) A solution of rubbing alcohol contains 142 g of isopropyl alcohol (C_3H_7OH) and 58.0 g of water. What are the mole fractions of IPA and water?

$$\text{mol IPA} = 142 \text{ g IPA} \times \frac{1 \text{ mol IPA}}{60.09 \text{ g IPA}} = 2.36 \text{ mol } C_3H_7OH$$

$$\text{mol } H_2O = 58.0 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} = 3.22 \text{ mol } H_2O$$

$$X_{C_3H_7OH} = \frac{2.36 \text{ mol } C_2H_8O_2}{2.36 \text{ mol } C_3H_7OH + 3.22 \text{ mol } H_2O} = 0.423$$

$$X_{H2O} = \frac{3.22 \text{ mol } H_2O}{2.36 \text{ mol IPA} + 3.22 \text{ mol } H_2O} = 0.577$$

H_2O_2 is a powerful oxidizing agent used in in rocket fuels and as a hair bleach. An aqueous solution H_2O_2 is 30.0% by mass and has a density of 1.11 g/mL. Calculate its **(a)** Molality, **(b)** Mole fraction **(c)** Molarity of H_2O_2

$$70.0 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} = 3.88 \text{ mol } H_2O$$

$$X_{H_2O_2} = \frac{0.882 \text{ mol } H_2O_2}{0.882 \text{ mol } H_2O_2 + 3.88 \text{ mol } H_2O} = 0.185$$

$M\ H_2O_2 = ?$

$$100.0 \text{ g soln} \times \frac{1 \text{ mL}}{1.11 \text{ g}} = 90.1 \text{ mL soln}$$

$$\frac{0.882 \text{ mol } H_2O_2}{90.1 \text{ mL soln} \frac{\text{L}}{10^3 \text{ mL}}} = 9.79\ M\ H_2O_2$$

- The density of a 25.0 % w/w solution of sulfuric acid (H_2SO_4) in water is 1.1783 g/mL at 25.0ºC. What is the molarity of this solution?

- The density of a 25.0 % w/w solution of sulfuric acid (H_2SO_4) in water is 1.1783 g/mL at 25.0ºC. What is the molarity of this solution?
  - Convert 25 g H_2SO_4 into moles:
    25.0 g H_2SO_4/100 g sol. × 1 mol H_2SO_4/98.1 g H_2SO_4
    = 0.255 mol/100 g sol.
  - Volume = 100 g sol. × 1 mL/1.1783 g sol.
    = 84.87 mL = 0.08487 L
  - Molarity = moles H_2SO_4/liters of solution
    = 0.255 mol H_2SO_4/0.08487 L = 3.00 M

A 0.750 M solution of H_2SO_4 in water has a density of 1.046 g/mL at 20ºC. What is the concentration in (a) mole fraction, (b) mass percent, (c) molality (MM = 98.086 g/mol) ?

A 0.750 M solution of H_2SO_4 in water has a density of 1.046 g/mL at 20ºC. What is the concentration in (a) mole fraction, (b) mass percent, (c) molality (MM = 98.086 g/mol) ?

(a) Since the solution is 0.750 mol/L and has a density of 1.046 g/mL (or 1.046 kg/L) density, 1.0 L solution contains 0.750 mol (or 73.6 g) H_2SO_4 and has a mass of 1.046 kg:

Mass of H_2O in 1 L solution = 1.046 kg – 0.0736 kg = 0.972 kg

0.972 kg H_2O = 972 g × 1 mol/18.0 g = 54.0 mol H_2O

For H_2SO_4, X = 0.750 mol H_2SO_4/(0.750 mol H_2SO_4 + 54.0 mol H_20) = 0.0137

A 0.750 M solution of H_2SO_4 in water has a density of 1.046 g/mL at 20ºC. What is the concentration in (a) mole fraction, (b) mass percent, (c) molality (MM = 98.086 g/mol) ?

(b) Mass % H_2SO_4 = 0.0736 kg H_2SO_4/1.046 kg total = 7.04%

(c) Since 0.972 kg water has 0.750 mol H_2SO_4 in it, 1 kg water would have 0.772 mol H_2SO_4 dissolved in it:

1.00 kg H_2O × 0.750 mol H_2SO_4/0.972 kg H_2O = 0.772 mol H_2SO_4

Thus, molality of sulfuric acid is 0.772 *m*

Household vinegar is a 5% w/w aqueous solution of acetic acid (Density = 1.01 g/mL) at 20ºC and has a pH = 2.4. Deterimine vinegar's concentration in (a) molarity (b) ppm (c) mole fraction (d) molality (MM = 60.05 g/mol) ?

Translate the symbols

5% w/w => $\frac{\text{5 g solute}}{\text{100 g solution}}$ => $\frac{\text{5 g solute}}{\text{5 g acetic + 95 g water}}$

These are key pieces of information you have to SEE! They will splice together with the other concentration definitions of molarity, ppm, molality and mole fraction.

Household vinegar is a 5.0% w/w aqueous solution of acetic acid (Density = 1.01 g/mL) at 20ºC and has a pH = 2.4. Deterimine vinegar's concentration in (a) molarity (b) ppm (c) mole fraction (d) molality (MM = 60.05 g/mol) ?

Translate the symbols by understanding what they mean!

5% w/w => $\frac{\text{5 g solute}}{\text{100 g solution}}$ => $\frac{\text{5 g solute}}{\text{5 g acetic + 95 g water}}$

$$M = \frac{5.0\ g\ \mathrm{CH_3CO_2H}}{100\ g\ soln} \times \frac{1\ mol\ \mathrm{CH_3CO_2H}}{60.05\ g\ \mathrm{CH_3CO_2H}} \times \frac{1.01\ g\ soln}{1\ mL\ soln} \times \frac{1000\ mL}{1\ L} = 0.84M$$

$$ppm = \frac{5.0\ g\ \mathrm{CH_3CO_2H}}{100\ g\ soln} \times \frac{1000\ mg\ \mathrm{CH_3CO_2H}}{1\ g\ \mathrm{CH_3CO_2H}} \times \frac{1.01\ g\ soln}{1\ mL\ soln} \times \frac{1000\ mL}{1\ L} = 50,500\ ppm$$

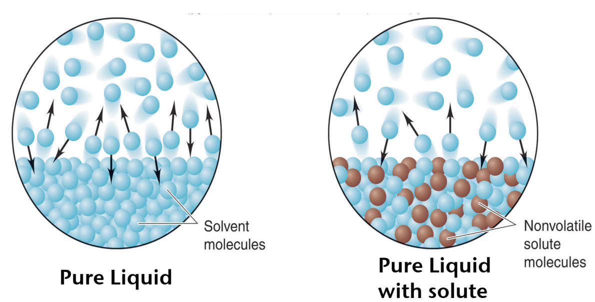
$$mol\ frac = \frac{5.0\ g\ \mathrm{CH_3CO_2H}}{5\ g\ \mathrm{CH_3CO_2H} + 95\ g\mathrm{H_2O}} \Rightarrow \frac{(5.0\ g\ \mathrm{CH_3CO_2H}/60.05\ g/mol)}{(5\ g\ \mathrm{CH_3CO_2H}/60.05\ g/mol) + (95\ g\mathrm{H_2O}/18.0\ g/mol\ \mathrm{H_2O})} = 0.016$$

$$molality = \frac{mol\ solute}{kg\ solvent} = \frac{5\ g\ \mathrm{CH_3CO_2H}}{95\ g\ \mathrm{H_2O}} \times \frac{1\ mol\ \mathrm{CH_3CO_2H}}{60.05\ g\ \mathrm{CH_3CO_2H}} \times \frac{1000\ g\ \mathrm{H_2O}}{1\ kg\ \mathrm{H_2O}} = 0.88m$$

An ethanol-water solution is prepared by dissolving 10.00 mL of etOH, C_2H_5OH (d=0.789 g/mL) in a sufficient volume of water to produce 100.0 mL of a solution of a with a density of 0.982 g/mL. What is the concentration of etOH n the solution expressed as a) volume % b) mass % c) mole fraction e) molarity g) molality h) ppm?

1. Identify the solute and the solvent
2. Understand that "solution" is solute + solvent
3. Understand and practice the definitions on the previous slide
4. Plug and chug

Colligative properties **are physical properties of solutions that arise because of the** ***number of solute molecules*** **dissolved in solution *(and not on the kind of solute particles dissolved in solution).**



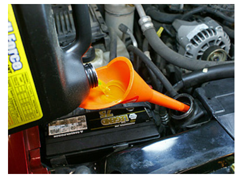
We distinguish between colligative properties of non-electrolytes and electrolytes (ionic compounds). They are different!

The Four-Colligative Properties (Non-Ionic)

Vapor-Pressure Lowering	$P_{solution} = X_{solvent} P^{\circ}_{solvent}$
Boiling-Point Elevation	$\Delta T_b = K_b\, m_{solution}$
Freezing-Point Depression	$\Delta T_f = -K_f\, m_{solution}$
Osmotic Pressure (π)	$\pi = MRT$

The above equations work only for Non-ionic non-volatile solutes BUT NOT IONIC SOLUTES because such solutions are non-ideal!

Colligative properties of solutions is used in many applications.



adding antifreeze to a car's radiator

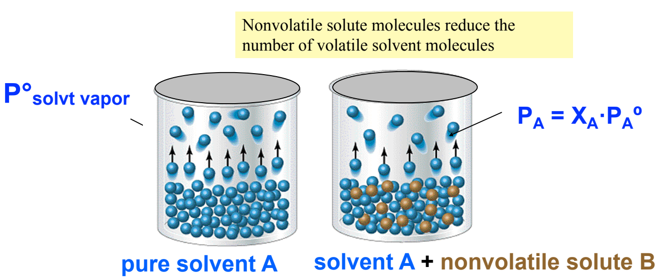
spreading salt on ice

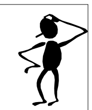
de-iceing airplanes

Vapor-Pressure Lowering: *Raoult's Law*: A non-volatile, non-ionic solute added to a pure solvent will lower the vapor pressure of the pure solvent.

$$P_{solution} = \chi_{solvent}\, P^{\circ}_{solvent}$$

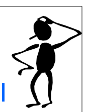
- $P_{solution}$: Vapor pressure of solution
- $\chi_{solvent}$: Mole Fraction of Solvent
- $P^{\circ}_{solvent}$: Vapor Pressure of Pure Solvent





At 100˚C ***what is the vapor pressure of*** a 50/50% (v/v) solution of ethylene glycol, $C_2H_6O_2$, in water at 1 atm? (Molar mass $C_2H_6O_2$ = 62.06 g/mol *d*($C_2H_6O_2$) = 1.1155 g/mL, *d*(H_2O) = 1.0000 g/mL, *d*(50v/50v) solution = 1.069 g/mL

A student should realize that the other possible question is what is the vapor pressure lowering relative to the pure solvent? Both can be obtained by knowing one equation.



At 100ºC what is the vapor pressure of a 50/50% (v/v) solution of ethylene glycol, $C_2H_6O_2$, in the solvent water at 1 atm? (MM $C_2H_6O_2$ = 62.06 g/mol *d*($C_2H_6O_2$) = 1.1155 g/mL, *d*(H_2O) = 1.0000 g/mL, *d* (50/50) = 1.069 g/mL

$$P_{solution} = \chi_{H2O}\, P^{\circ}_{H2O}$$

$$\text{Moles } C_2H_6O_2 = 500.\text{ mL} \times \frac{1.1155\text{ g } C_2H_6O_2}{\text{mL}} \times \frac{1\text{ mol}}{62.06\text{ g}} = 8.99\text{ mol}$$

$$\text{Moles } H_2O = 500.\text{ mL} \times \frac{1.000\text{ g } H_2O}{\text{mL}} \times \frac{1\text{ mol}}{18.02\text{ g}} = 27.8\text{ mol}$$

Mole Fraction H_2O = 27.8 mol/ (27.8 mol + 8.99 mol) = 0.7556

$$\mathbf{P_{H2O} = \chi_{H2O}\, P^{\circ}_{H2O}} = .7556 \times 760.\text{ torr} = 574.\text{ torr}$$

Vapor-pressure lowering can be recast in terms of the mole fraction of solute χ_{Solute}.

$$P_{Solution} = \chi_{solvent}\, P^{\circ}_{solvent} \qquad (1)$$

$$\chi_{Solute} + \chi_{Solvent} = 1$$

$$\chi_{Solvent} = 1 - \chi_{Solute}$$

substituting for $\chi_{solvent}$

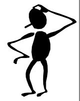
$$P_{Solution} = (1 - \chi_{Solute})\, P^{\circ}_{solvent} \qquad (2)$$

exapanding

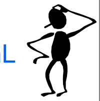
$$P_{Soln} = P^{\circ}_{solvent} - P^{\circ}_{solvent}(\chi_{Solute}) \qquad (3)$$

$$\mathbf{\Delta P = (P^{\circ}_{solvent} - P_{soln}) = P^{\circ}_{solvent}(\chi_{Solute})}$$

Calculate the vapor pressure lowering, ΔP, when 10.0 mL of glycerol ($C_3H_8O_3$) is dissolved in 500. mL of water at 50.ºC. At this temperature, the vapor pressure of pure water is 92.5 torr and its density is 0.988 g/mL. The density of glycerol is 1.26 g/mL.



Calculate the vapor pressure lowering, ΔP, when 10.0 mL of glycerol ($C_3H_8O_3$) is dissolved in 500. mL of water at 50.ºC. At this temperature, the vapor pressure of pure water is 92.5 torr and its density is 0.988 g/mL. The density of glycerol is 1.26 g/mL.



PLAN: Find the mol fraction, χ, of glycerol in solution and multiply by the vapor pressure of water. $\Delta P = (P°_{solvent} - P_{solvent}) = P°_{solvent}(\chi_{Solute})$

SOLUTION:

$$10.0 \text{ mL } C_3H_8O_3 \times \frac{1.26 \text{ g } C_3H_8O_3}{\text{mL } C_3H_8O_3} \times \frac{\text{mol } C_3H_8O_3}{92.09 \text{ g } C_3H_8O_3} = 0.137 \text{ mol } C_3H_8O_3$$

$$500.0 \text{ mL } H_2O \times \frac{0.988 \text{ g } H_2O}{\text{mL } H_2O} \times \frac{\text{mol } H_2O}{18.02 \text{ g } H_2O} = 27.4 \text{ mol } H_2O$$

$$\Delta P = \frac{0.137 \text{ mol } C_3H_8O_3}{0.137 \text{ mol } C_3H_8O_3 + 27.4 \text{ mol } H_2O} \times 92.5 \text{ torr} = 0.461 \text{ torr}$$

$\chi = 0.00498$

Colligative properties **are properties that depend only on the** ***number of solute particles*** **in solution and not on the nature of the solute particles.**

The Four-Colligative Properties Non-Ionic

Vapor-Pressure Lowering	$P_{solution} = X_{solvent} P°_{solvent}$
Boiling-Point Elevation	$\Delta T_b = K_b\, m_{solution}$
Freezing-Point Depression	$\Delta T_f = -K_f\, m_{solution}$
Osmotic Pressure (π)	$\pi = MRT$

The above equations work only for Non-ionic non-volative solutes BUT NOT IONIC SOLUTES because such solutions are non-ideal!

2. Boiling Point Elevation: The dissolution of a nonvolatile non-ionic solute in a pure solvent ***increases the boiling the point of a solution*.**

$$\Delta T_b = T_{bp} - T°_{bp} = K_b\, m$$

Solution boiling point (T_{bp}); Pure solvent boiling point ($T°_{bp}$); Boiling point elevation (+) (ΔT_b); Molal boiling point elevation constant (˚C/m) (K_b); Molality (mol/kg) (m)

3. Freezing Point Depression: The addition of a non-volatile non-ionic solute dissolved in a pure solvent ***decreases*** **the freezing point of the solution.**

$$\Delta T_f = T_f - T°_{fp} = -K_f\, m$$

Pure solvent freezing point ($T°_{fp}$); Freezing point depression (-) (ΔT_f); Molal freezing point depression constant (˚C/m) (K_f); Solution Molality (mol/kg) (m)

The molal BP and FP constants must be provided

Molal Boiling-Point Eleveation and Freezing Point Depression Constants of Common Liquids

Solvent	Normal Freezing Point (°C)*	K_f (°C/m)	Normal Boiling Point (°C)*	K_b (°C/m)
Water	0	1.86	100	0.52
Benzene	5.5	5.12	80.1	2.53
Ethanol	−117.3	1.99	78.4	1.22
Acetic acid	16.6	3.90	117.9	2.93
Cyclohexane	6.6	20.0	80.7	2.79

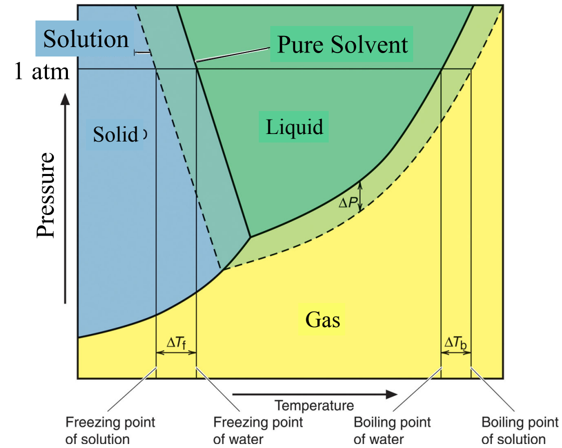
* Measured at 1 atm.

Table 13.6 Molal Boiling Point Elevation and Freezing Point Depression Constants of Several Solvents

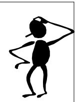
Solvent	Boiling Point (°C)*	K_b (°C/m)	Melting Point (°C)	K_f (°C/m)
Acetic acid	117.9	3.07	16.6	3.90
Benzene	80.1	2.53	5.5	4.90
Carbon disulfide	46.2	2.34	-111.5	3.83
Carbon tetrachloride	76.5	5.03	-23	30.
Chloroform	61.7	3.63	-63.5	4.70
Diethyl ether	34.5	2.02	-116.2	1.79
Ethanol	78.5	1.22	-117.3	1.99
Water	100.0	0.512	0.0	1.86

*at 1 atm.

We can compare the freezing and boiling pts of a pure solvent and a solution in a phase diagram.



Suppose you mix 1.00 kg of ethylene glycol antifreeze ($C_2H_6O_2$) to 4450 g of water to make an antifreeze/ antiboil solution for your car. Calculate the boiling and freezing points of the resulting solution?



You add 1.00 kg of ethylene glycol antifreeze ($C_2H_6O_2$) to your car radiator, which contains 4450 g of water. What are the boiling and freezing points of the resulting solution?

$$m\ C_2H_6O_2 = 1.00 \times 10^3\ g\ C_2H_6O_2 \times \frac{mol\ C_2H_6O_2}{62.07\ g\ C_2H_6O_2} = 16.1\ mol\ C_2H_6O_2$$

$$m\ C_2H_6O_2 = \frac{16.1\ mol\ C_2H_6O_2}{4.450\ kg\ H_2O} = 3.62\ m\ C_2H_6O_2$$

$$\Delta T_b = 0.512\ ^oC/m \times 3.62\ m = 1.85\ ^oC$$

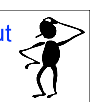
BP = 101.85 °C

$$\Delta T_f = 1.86\ ^oC/m \times 3.62\ m = +6.73\ ^oC$$

FP = -6.73 °C

In Upstate New York (where Rick was hatched) folks put a solution of 50% v/v of ethylene glycol (antifreeze) in water into the radiator of our automobiles. Does this make sense to do? At what temperature will this solution boil and freeze? The density of ethylene glycol 1.11 g/ml and that of water is 1.00 g/mL. The molar mass of ethylene glycol is 62.01 g.

In Upstate New York (where Rick was hatched) folks put a solution of 50% v/v of ethylene glycol (antifreeze) in water into the radiator of our automobiles. Does this make sense to do? At what temperature will this solution boil and freeze? The density of ethylene glycol 1.11 g/ml and that of water is 1.00 g/mL. The molar mass of ethylene glycol is 62.01 g.



$\Delta T_f = K_f\, m$ $\quad$ K_f water = 1.86 ^{0}C/m

$$m = \frac{\text{moles of solute}}{\text{mass of solvent (kg)}} = \frac{478\ g \times \frac{1\ mol}{62.01\ g}}{3.202\ kg\ solvent} = 2.41\ m$$

$$\Delta T_f = K_f\, m = 1.86\ ^0C/m \times 2.41\ m = 4.48\ ^0C$$

$$\Delta T_f = T_f^0 - T_f$$

$$T_f = T_f^0 - \Delta T_f = 0.00\ ^0C - 4.48\ ^0C = -4.48\ ^0C$$

Colligative properties **are properties that depend only on the** ***number of solute particles*** **in solution and not on the nature of the solute particles.**

The Four-Colligative Properties Non-Ionic

Vapor-Pressure Lowering	$P_{solution} = X_{solvent} P^{\circ}_{solvent}$
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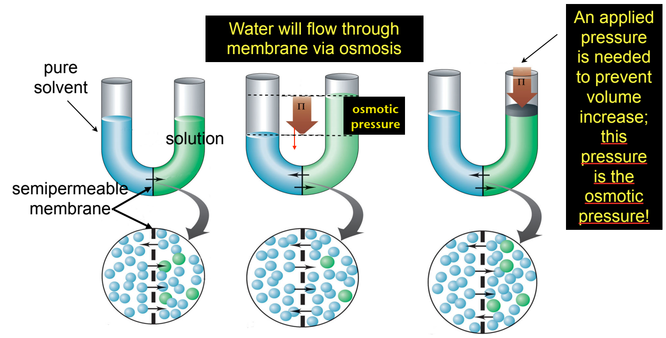
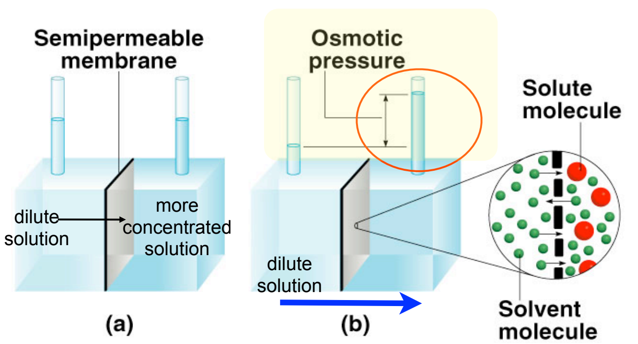
Osmotic Pressure (π)	$\pi = MRT$

The above equations work only for Non-ionic non-volatile solutes BUT NOT IONIC SOLUTES because such solutions are non-ideal!

Osmosis is the diffusion of a solvent (usually water) through a semi-permeable membrane, from a solution of low solute concentration (high water concentration) to a solution with high solute concentration (lower water concentration). "Osmosis tries to dilute the concentrate"



Osmosis **needs a membrane to select the passage of** ***solvent molecules*** **from a dilute solution to a more concentrated one (solutes can't cross).**



A pressure difference results from the net movement of solvent from a less-solute concentrated (*hypotonic*) to the more-solute concentrated (*hypertonic*) solution.

For dilute solutions of electrolytes the osmotic pressure is given by:

$$\pi = \mathrm{M\,R\,T}$$

M is the molarity of the solution

R is the gas constant 0.0821 L atm/M K

$$\pi = \frac{n}{V}\mathrm{R\,T}$$

T is the temperature (in Kelvin)

$$\Delta\pi = \Delta \mathrm{M\,R\,T}$$

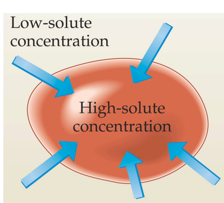
When there is a concentration "gradient" (difference in concentration)

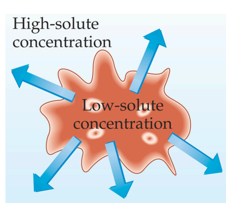
The driving force is due to the difference in concentration of the solutions on each side of the membrane.

Cell membranes **are semi-permeable "filters" that whereby water can pass freely but larger molecules can not.**

Movement of solvent (water) from dilute to concentrated side!

No Osmotic Pressure Concentrations Are the Same

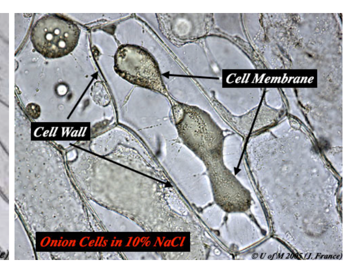




isotonic solution | ***hypotonic*** solution | ***hypertonic*** solution

Osmosis in an Onion Cell As It Really Looks

Normal onion cell as seen under microscope



Plasmolyzed cell (cell membrane has shrunk from the cell wall

Suppose we have a 0.020 molar solution of table sugar (sucrose) and a semi-permeable membrane not permeable to sucrose. What osmotic pressure in mm Hg and to what height could this pressure support a column of water (density Hg =13.6 g/mL and water = 1g/mL?

$\pi = M\,R\,T$

$\pi = 0.02\ M \times 0.0821\ L\ atm/mol\ K \times 298K = .49\ atm$

$\pi = .49\ atm \times 760\ torr/1\ atm = 371\ mm\ Hg$

$\pi = 371\ mm\ Hg \times 13.6 = 5.0$ meters!

A solution prepared by dissolving 20.0 mg of insulin in water and diluting to a volume of 5.00 mL gives an osmotic pressure of 12.5 torr at 300K. What is the molecular mass of the insulin?

$$M = \frac{12.5\ torr \times \frac{1\ atm}{760\ torr}}{0.0821\ L\ atm\ mol^{-1}\ K^{-1}\ 300\ K} = \frac{6.68 \times 10^{-4}\ mol\ insulin}{L} =$$

$$moles = 6.68 \times 10^{-4}\ \frac{mol\ insulin}{L} \times 0.005L = 3.33 \times 10^{-6} mol$$

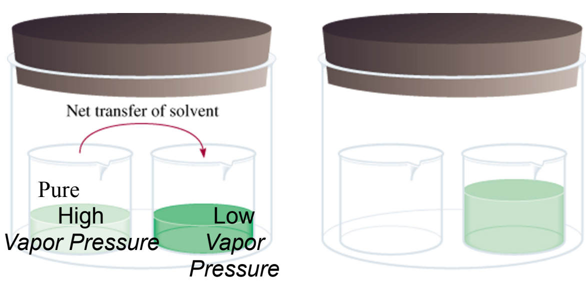
$$MolarMass = grams/mole = 0.020\ g/3.33 \times 10{-}6\ mol = 5988\ g/mol$$

Calculate molarity of a aqueous solution at 300K which is found to have an osmotic pressure of 3.00 atm. $\pi = M\ R\,T$

$$M = \frac{\pi}{R\,T} = \frac{3.00\ atm}{0.0821 L\ atm\ mol^{-1} K^{-1}\ 300\ K} = 0.122\ M$$

Now this is wild: Anology to Osmosis

In a closed container the solution with the highest vapor pressure will completely transfer to the container of lower vapor pressure until the mole fractions of solvent are equal in both! Cool.........



We can calculate the molar mass of a solute using any of the four *colligative properties*.

Because osmotic pressure changes are much much larger with solute concentrations we osmotic pressure is used to measure molar mass.

Biochemists have discovered more than 400 mutant varieties of hemoglobin (Hb), the blood protein that carries oxygen throughout the body. A physician studying a form of Hb associated with a fatal disease first finds its molar mass (M). She dissolves 21.5 mg of the protein in water at 5.0 ºC to make 1.50 mL of solution and measures an osmotic pressure of 3.61 torr. What is the molar mass of this Hb mutant?

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PLAN: We know Π as well as R and T. Convert Π to atm and T to Kelvin. Use the Π equation to find the molarity M and then the amount and volume of the sample to calculate M.

SOLUTION:

$$M = \frac{\Pi}{RT} = \frac{3.61\ torr \times \frac{atm}{760\ torr}}{(0.0821\ L \cdot atm/mol \cdot K)(278.15\ K)} = 2.08 \times 10^{-4}\ M$$

$$\frac{2.08 \times 10^{-4}\ mol}{L} \times 1.50\ mL \times \frac{L}{10^3\ mL} = 3.12 \times 10^{-7}\ mol$$

mol = g/M

$$21.5\ mg \times \frac{g}{10^3\ mg} \times \frac{1}{3.12 \times 10^{-7}\ mol} = 6.89 \times 10^4\ g/mol$$

Raoult's Law applies to mixtures of solutes and solvents that are volatile.

– For a two component mixture of volatile liquids we may apply Raoult's Law for both components.

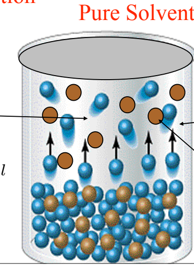
$P_A = \chi_A\ P_A°$ $\qquad$ $P_B = \chi_B\ P_B°$

Vapor pressure from A; Mole Fraction of A; Vapor Pressure of Pure Solvent

Dalton's Law

$P_{Total} = P_A + P_B$

$$P_A = \frac{n_A}{n_{total}}\ P_{total} = \chi_A\ P_{total}$$

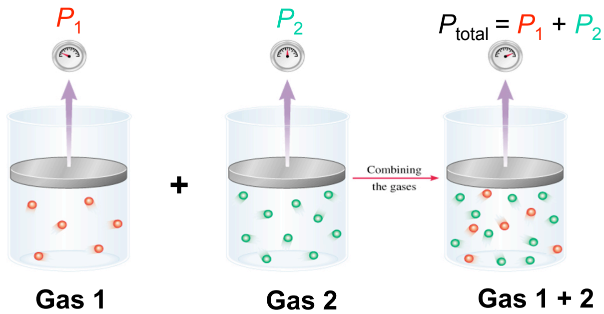


Raoult's Law

$p_A = X_A \cdot p_A°$

$p_B = X_B \cdot p_B°$

Dalton's Law...**the total pressure of a mixture of gases is equal to the sum of the partial pressure of each gas in the mixture (at constant V and T).**



Dalton's Law

$$P_A = \frac{n_A}{n_{total}} P_{total} = \chi_A \; P_{total}$$

Partial Pressure of Vapor A or Above Solution

Mole Fraction of Gas A in the Vapor Phase!

Total Pressure of Vapor or Gas Above Solution

$$P_T = \frac{n_A RT}{V_T} + \frac{n_B RT}{V_T} + \frac{n_C RT}{V_T}$$

$$P_T = (n_A + n_B + n_C)\frac{RT}{V_T}$$

$$\frac{P_A}{P_T} = \frac{n_A \frac{RT}{V_T}}{n_T \frac{RT}{V_T}}$$

$$\frac{P_A}{P_T} = \frac{n_A}{n_T} = \frac{n_A}{n_1 + n_2 + n_3 + ...}$$

=

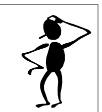
We define and call the fraction; $\frac{n_A}{n_T}$ *the mole fraction* X_A

$$P_A = \frac{n_A}{n_T} P_T$$

$$P_B = \frac{n_B}{n_T} P_T$$

$$X_A + X_B = 1$$

A solution is prepared in which the mole fractions of benzene and toluene are both 0.500. What are the partial pressures of the benzene and toluene above this solution? What is the total vapor pressure? The vapor pressures of pure benzene and toluene at 25°C are 95.1 and 28.4 mm Hg, respectively.



$P_{benzene} = \chi_{benzene} P°_{benzene} = (0.500)(96.1 \text{ mm Hg}) = 47.6 \text{ mm Hg}$

$P_{toluene} = \chi_{toluene} P°_{toluene} = (0.500)(28.4 \text{ mm Hg}) = 14.2 \text{ mm Hg}$

$P_{total} = P_{benzene} + P_{toluene} = 61.8 \text{ mm Hg}$

What is the composition of the vapor in equilibrium with the benzene-toluene solution?

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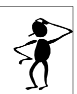

Using Dalton's Law we can calculate what the vapor composition is:

$$P_A = \frac{n_A}{n_{total}} P_{total} = \chi_A \; P_{total}$$

$\chi_{toluene} = P_{toluene}/P_{total} = 14.2 \text{ mm Hg}/61.89 \text{ mm Hg} = 0.230$

$\chi_{benzene} = P_{benzene}/P_{total} = 47.6 \text{ mm Hg}/61.89 \text{ mm Hg} = 0.770$

The colligative properties of solutions containing *ionic solutes* is different than non-ionic solutes!

For non-electrolyte or non-ionic solutions, the solute dissolves to give 1 mole of particles in the solution.

For electrolyte or ionic solutions, the solute dissolves to give a larger number of particles in solution.

Under ideal conditions **the van' Hofft factors are the total number of moles of ions per mole of substance.**

Example: 1 mol sugar => 1 mol particle
1 mol NaCl => 2 mol particle
1 mol $MgCl_2$ => 3 mol particle

Because ionic solutes have more particles than non-ionic ones colligative property equations are modified to reflect this.

0.1 *m* nonelectrolytes solution → 0.1 *m* in solution
0.1 *m* NaCl solution → 0.2 *m* ions in solution
0.1 *m* $CaCl_2$ solution → 0.3 *m* ions in solution
0.2 *m* Na_3PO4 solution → 0.8 *m* ions in solution

We modify the non-ionic equations by multiplying by the van't Hoff factor, *i*

Vapor Pressure Lowering	$P_{solution} = i\, X_{solvent} P^{\circ}_{solvent}$ $\Delta P_{solution} = i\, X_{solute} P^{\circ}_{solvent}$
Boiling-Point Elevation	$\Delta T_b = i\, K_b\, m$
Freezing-Point Depression	$\Delta T_f = i\, K_f\, m$
Osmotic Pressure	$\pi = i\, M\, R\, T$

The van't Hoff factor, *i*, tells us what the "effective" number of ions are in the solution. It's just a table summarizing experimental measurements.

van't Hoff factor (i)

$$i = \frac{\text{measured value for electrolyte solution}}{\text{expected value for nonelectrolyte solution}}$$

van't Hoff Factors for 3-solutes as 5 different

MOLALITY

Solute	1.0	0.10	0.010	0.0010	...	Inf. dil.[a]
NaCl	1.81	1.87	1.94	1.97	...	2
$MgSO_4$	1.09	1.21	1.53	1.82	...	2
$Pb(NO_3)_2$	1.31	2.13	2.63	2.89	...	3

[a]The limiting values: $i = 2, 2$, and 3 are reached when the solution is infinitely dilute. Note that a solute whose ions are singly charged (for example, NaCl) approaches its limiting value more quickly than does a solute whose ions carry higher charges. Interionic attractions are greater in solutes with more highly charged ions.

The Four-Colligative Equations For Both Ionic and Non-Ionic

Vapor Pressure Lowering	$P_{solution} = i\, X_{solvent} P^{\circ}_{solvent}$ $\Delta P_{solution} = i\, X_{solute} P^{\circ}_{solvent}$
Boiling-Point Elevation	$\Delta T_b = i\, K_b\, m$
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van't Hoff Factors for 3-solutes as 5 different

	MOLALITY					**IDEAL**
Solute	1.0	0.10	0.010	0.0010	...	Inf. dil.[a]
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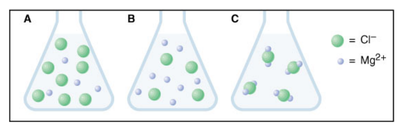
Sample Problem 13.9 **Finding Colligative Properties from Molecular Scenes**

PROBLEM: A 0.952-g sample of magnesium chloride is dissolved in 100. g of water in a flask.



(a) Which scene depicts the solution best?
(b) What is the amount (mol) represented by each green sphere?
(c) Assuming the solution is ideal, what is its freezing point (at 1 atm)?

PLAN: (a) Consider the formula for magnesium chloride, an ionic compound.
(b) Use the answer to part (a), the mass given, and the mol mass.
(c) The total number of mols of cations and anions, mass of solvent, and equation for freezing point depression can be used to find the new freezing point of the solution.

SOLUTION: (a) The formula for magnesium chloride is $MgCl_2$; therefore the correct depiction must be A with a ratio of 2 Cl^-/ 1 Mg^{2+}.

Sample Problem 13.9 **Finding Colligative Properties from Molecular Scenes**

SOLUTION: (b)

$$\text{mols } MgCl_2 = \frac{0.952\text{ g } MgCl_2}{\frac{95.21\text{ g } MgCl_2}{\text{mol } MgCl_2}} = 0.0100\text{ mol } MgCl_2$$

$$\text{mols } Cl^- = 0.0100\text{ mol } MgCl_2 \times \frac{2\text{ mol } Cl^-}{1\text{ mol } MgCl_2} = 0.0200\text{ mol } Cl^-$$

$$\text{mols/sphere} = \frac{0.0200\text{ mol } Cl^-}{8\text{ spheres}} = 2.50 \times 10^{-3}\text{ mol/sphere}$$

(c)

$$\text{molality } (m) = \frac{0.0100\text{ mol } MgCl_2}{100.\text{ g} \times \frac{1\text{ kg}}{10^3\text{ g}}} = 0.100\text{ m } MgCl_2$$

Assuming this is an IDEAL solution, the van't Hoff factor, ***i***, should be 3.

$$\Delta T_f = \boldsymbol{i}\,(K_f m) = 3(1.86^{\circ}C/m \times 0.100\ m) = 0.558^{\circ}C$$

$$T_f = 0.000^{\circ}C - 0.558^{\circ}C = -0.558^{\circ}C$$