

Slide 1

Colligative Properties

Practical uses of solutions

Slide 2


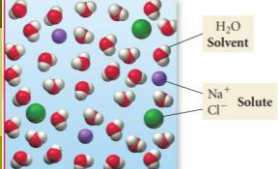
Solution

- homogeneous mixtures
 - composition may vary from one sample to another
 - appears to be one substance, though really contains multiple materials
- most homogeneous materials we encounter are actually solutions
 - e.g., air and sea water
- nature has a tendency toward spontaneous mixing
 - generally, uniform mixing is more energetically favorable

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Solutions

- **solute** is the dissolved substance



- solutions in which the solvent is water are called aqueous solutions

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dissolve NaCl.MOV

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Solubility

- when one substance (solute) dissolves in another (solvent) it is said to be **soluble**
 - salt is soluble in water
 - bromine is soluble in methylene chloride
- when one substance does not dissolve in another it is said to be **insoluble**
 - oil is insoluble in water
- the solubility of one substance in another depends on two factors – nature's tendency towards mixing, and the types of intermolecular attractive forces

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Spontaneous Mixing

Spontaneous Mixing

When the barrier is removed, spontaneous mixing occurs, producing a solution of uniform concentration.

(a) (b)

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Solubility

- there is usually a limit to the solubility of one substance in another
 - gases are **always** soluble in each other
 - two liquids that are mutually soluble are said to be **miscible**
 - alcohol and water are miscible
 - oil and water are immiscible
- the maximum amount of solute that can be dissolved in a given amount of solvent is called the **solubility**
- the solubility of one substance in another varies with temperature and pressure

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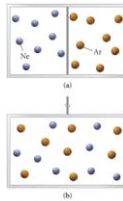
Mixing and the Solution Process

Entropy

formation of a solution does not necessarily lower the potential energy of the system

- the difference in attractive forces between atoms of two separate ideal gases vs. two mixed ideal gases is negligible
- yet the gases mix spontaneously

- the gases mix because the energy of the system is lowered through the release of **entropy**
- **entropy** is the measure of energy dispersal throughout the system
- energy has a spontaneous drive to spread out over as large a volume as it is allowed



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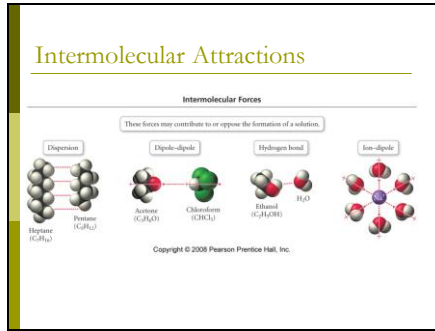
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Intermolecular Forces and the Solution Process

Enthalpy of Solution

- energy changes in the formation of most solutions also involve differences in attractive forces between particles
- must overcome solute-solute attractive forces
 - endothermic
- must overcome some of the solvent-solvent attractive forces
 - endothermic
- at least some of the energy to do this comes from making new solute-solvent attractions
 - exothermic

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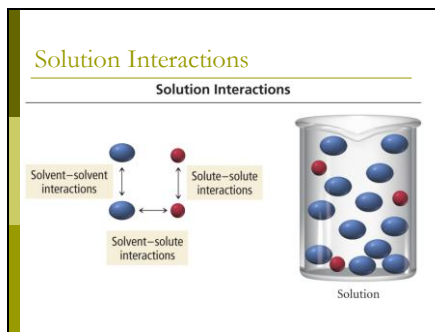


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Relative Interactions and Solution Formation

Solute-to-Solvent	>	Solute-to-Solute + Solvent-to-Solvent	Solution Forms
Solute-to-Solvent	=	Solute-to-Solute + Solvent-to-Solvent	Solution Forms
<p>□ when the solute-to-solvent attractions are weaker than the sum of the solute-to-solute and solvent-to-solvent attractions, the solution will only form if the energy difference is small enough to be overcome by the entropy.</p>			

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Will It Dissolve?

Chemist's Rule of Thumb –

Like Dissolves Like

- a chemical will dissolve in a solvent if it has a similar structure to the solvent
- when the solvent and solute structures are similar, the solvent molecules will attract the solute particles at least as well as the solute particles to each other

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Classifying Solvents

Solvent	Class	Structural Feature
Water, H ₂ O	polar	O-H
Methyl Alcohol, CH ₃ OH	polar	O-H
Ethyl Alcohol, C ₂ H ₅ OH	polar	O-H
Acetone, C ₃ H ₆ O	polar	C=O
Toluene, C ₇ H ₈	nonpolar	C-C & C-H
Hexane, C ₆ H ₁₄	nonpolar	C-C & C-H
Diethyl Ether, C ₄ H ₁₀ O	nonpolar	C-C, C-H & C-O, (nonpolar > polar)
Carbon Tetrachloride	nonpolar	C-Cl, but symmetrical

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Example 12.1a – predict whether the following vitamin is soluble in fat or water

The 4 OH groups make the molecule highly polar and it will also H-bond to water.

Vitamin C is water soluble

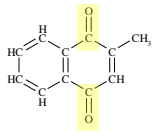
Vitamin C

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Example 12.1b – predict whether the following vitamin is soluble in fat or water

The 2 C=O groups are polar, but their geometric symmetry suggests their pulls will cancel and the molecule will be nonpolar.

Vitamin K₃ is fat soluble



Vitamin K₃

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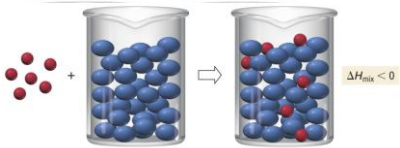
Energetics of Solution Formation

- overcome attractions between the solute particles – **endothermic**
- overcome some attractions between solvent molecules – **endothermic**
- for new attractions between solute particles and solvent molecules – **exothermic**
- the overall ΔH depends on the relative sizes of the ΔH for these 3 processes

$$\Delta H_{\text{sol'n}} = \Delta H_{\text{solute}} + \Delta H_{\text{solvent}} + \Delta H_{\text{mix}}$$

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Solution Process



$\Delta H_{\text{mix}} < 0$

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3. form new solute-solvent attractions, releasing energy
2. add energy in to overcome some solvent-solvent attractions

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Energetics of Solution Formation

if the total energy cost for breaking attractions between particles in the pure solute and pure solvent is **greater than** the energy released in making the new attractions between the solute and solvent, the overall process will be **endothermic**

The diagram illustrates the energetics of solution formation. It shows the energy levels for solute and solvent particles in their pure states and in solution. The process involves breaking solute-solute and solvent-solvent attractions (endothermic) and forming solute-solvent attractions (exothermic). The overall process is endothermic if the energy cost of breaking attractions is greater than the energy released in forming new attractions.

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Heats of Hydration

for aqueous ionic solutions, the energy added to overcome the attractions between water molecules and the energy released in forming attractions between the water molecules and ions is combined into a term called the **heat of hydration**

- attractive forces in water = H-bonds
- attractive forces between ion and water = ion-dipole
- $\Delta H_{\text{hydration}}$ = heat released when 1 mole of gaseous ions dissolves in water

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Heat of Hydration

Heat of Hydration

The diagram shows the energy levels for the hydration of potassium fluoride. It starts with KF(s) at a low energy level. The energy cost of breaking KF(s) into K+(g) + F-(g) is +821 kJ/mol. The energy released in forming K+(aq) + F-(aq) is -819 kJ/mol. The overall heat of hydration is +2 kJ/mol.

$\Delta H_{\text{diss}} = -\Delta H_{\text{lattice}}$
 $\Delta H_{\text{diss}} = +821 \text{ kJ/mol}$

$\Delta H_{\text{hydration}} = -819 \text{ kJ/mol}$

$\Delta H_{\text{hydr}} = +2 \text{ kJ/mol}$

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Solubility Limit

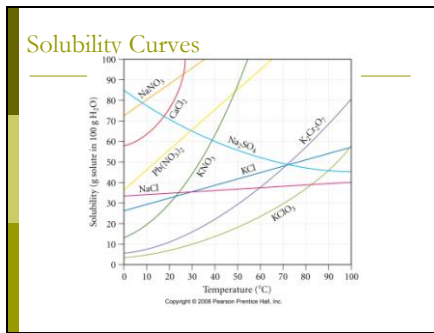
- a solution that has the maximum amount of solute dissolved in it is said to be **saturated**
 - depends on the amount of solvent
 - depends on the temperature
 - and pressure of gases
- a solution that has less solute than saturation is said to be **unsaturated**
- a solution that has more solute than saturation is said to be **supersaturated**

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Temperature Dependence of Solubility of Solids in Water

- solubility is generally given in grams of solute that will dissolve in 100 g of water
- for **most** solids, the solubility of the solid increases as the temperature increases
 - when $\Delta H_{\text{solution}}$ is endothermic
- solubility curves can be used to predict whether a solution with a particular amount of solute dissolved in water is saturated (on the line), unsaturated (below the line), or supersaturated (above the line)

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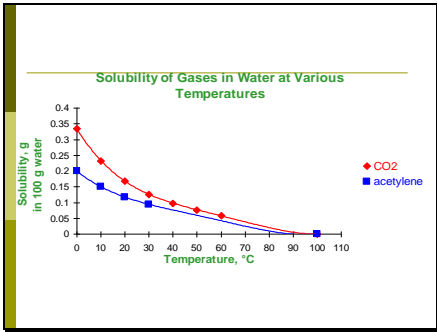


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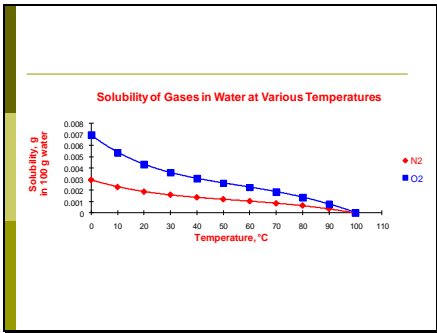
Temperature Dependence of Solubility of Gases in Water

- solubility is generally given in moles of solute that will dissolve in 1 Liter of solution
- generally lower solubility than ionic or polar covalent solids because most are nonpolar molecules
- for **all** gases, the solubility of the gas decreases as the temperature increases
 - the $\Delta H_{\text{solution}}$ is exothermic because you do not need to overcome solute-solute attractions

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Pressure Dependence of Solubility of Gases in Water

- the larger the partial pressure of a gas in contact with a liquid, the more soluble the gas is in the liquid

The diagram shows three stages of a gas cylinder submerged in water. In the first stage, labeled 'Equilibrium', the cylinder is partially filled with gas and the water level is at a certain height. In the second stage, labeled 'Pressure is Increased, More O₂ Dissolves', the gas pressure is increased, and more gas molecules are shown entering the water, raising the water level. In the third stage, labeled 'Equilibrium restored', the system has reached a new equilibrium with a higher water level.

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Henry's Law

- the solubility of a gas (S_{gas}) is directly proportional to its partial pressure, (P_{gas})

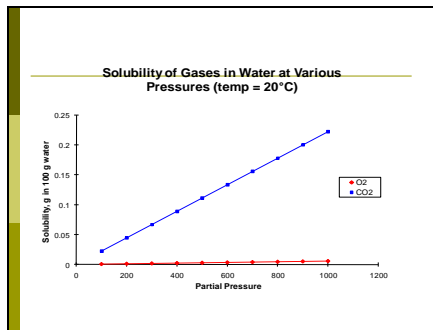
$$S_{gas} = k_H P_{gas}$$

- k_H is called **Henry's Law Constant**

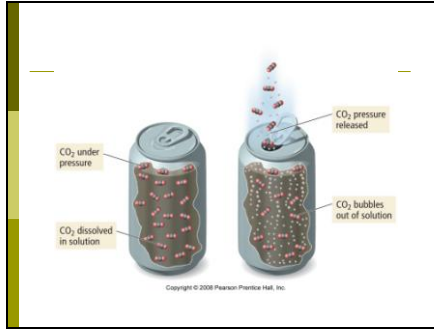
Gas	k_H (M/atm)
O ₂	1.3×10^{-3}
N ₂	6.1×10^{-4}
CO ₂	3.4×10^{-2}
NH ₃	5.8×10^1
He	3.7×10^{-4}

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Ex 12.2 – What pressure of CO₂ is required to keep the [CO₂] = 0.12 M at 25°C?

Given:	$S = [\text{CO}_2] = 0.12 \text{ M}$
Find:	P of CO ₂ , atm
Concept Plan:	$[\text{CO}_2] \xrightleftharpoons[S = k_H P]{P = \frac{S}{k_H}}$
Relationships:	$S = k_H P, k_H = 3.4 \times 10^{-2} \text{ M/atm}$
Solve:	$P = \frac{S}{k_H} = \frac{0.12 \text{ M}}{3.4 \times 10^{-2} \text{ M} \cdot \text{atm}^{-1}} = 3.5 \text{ atm}$
Check:	the unit is correct, the pressure higher than 1 atm meets our expectation from general experience

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Units of Concentration

Whatever units you use, the goal is the same: specify the quantity of 1 component (the solute_s) relative to the quantity of another component (the solvent).

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Common Units

- % by mass
- % by volume
- Mole %
- Molarity (M)
- Molality (m)

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Common Units

- % by mass – g solute/100 g solution
- % by volume
- Mole %
- Molarity (M)
- Molality (m)

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Common Units

- % by mass – g solute/100 g solution
- % by volume – Liters solute/100 L solution
- Mole % - moles solute/100 moles solution
- Molarity (M)
- Molality (m)

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Common Units

% by mass – g solute/100 g solution

% by volume – Liters solute/100 L solution

Mole % - moles solute/100 moles solution

Molarity (M) – moles solute/ L solution

Molality (m)

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Common Units

% by mass – g solute/100 g solution

% by volume – Liters solute/100 L solution

Mole % - moles solute/100 moles solution

Molarity (M) – moles solute/ L solution

Molality (m) – moles solute/ kg **solvent**

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Solute as part of a solution

- Note that, with the exception of molality, all of the units of concentration are expressed as some amount of solute compared to some amount of solution.
- All the units of concentration are easily convertible, although sometimes you may need to know another piece of information (molar mass, density, etc.)

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Molarity

Molarity is probably the most common unit of concentration in chemistry.

Why is that?

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Molarity

Molarity is probably the most common unit of concentration in chemistry.

Why is that?

UNITS! UNITS! UNITS!

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Molarity

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

This is both a chemically relevant unit and a practically relevant unit.

L solution is very easy to measure in the lab!

MOLES! MOLES! MOLES!

Reaction occur based on the relative number of moles.

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% by mass

This is one of the easiest units to use.

Why?

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% by mass

This is one of the easiest units to use.

Why?

UNITS! UNITS! UNITS!

Mass is easy to measure in the lab!

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% by mass

$\% \text{ by mass} = \frac{\text{g solute}}{100 \text{ g solution}}$

This is easy to use, but not as chemically relevant as Molarity. After all, it's all about the MOLES! MOLES! MOLES!

Is this a problem?

Not at all, it is easy to convert these units, but you need to know a few other things.

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Example - How would you prepare 250.0 g of 5.00% by mass glucose solution (normal glucose)?

Given: 250.0 g solution
Find: g Glucose
Equivalence: 5.00 g Glucose = 100 g solution
Solution Map: $\text{g solution} \xrightarrow{\frac{5.00 \text{ g Glucose}}{100 \text{ g solution}}} \text{g Glucose}$

Apply Solution Map:
 $250.0 \text{ g solution} \times \frac{5.00 \text{ g glucose}}{100 \text{ g solution}} = 12.5 \text{ g glucose}$

Answer:
Dissolve 12.5 g of glucose in enough water to total 250.0 g

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Converting M to % by mass

% by mass = $\frac{\text{g solute}}{100 \text{ g solution}}$

M = $\frac{\text{moles solute}}{\text{L solution}}$

What do we need to know to be able to convert from one to the other?

UNITS! UNITS! UNITS! They shall lead the way.

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Converting M to % by mass

Compare the units:

$\frac{\text{g solute}}{100 \text{ g solution}}$

$\frac{\text{Moles solute}}{\text{Liter solution}}$

What do you see?
Both numerators refer only to solute.
Both denominators refer to the whole solution.
PLEASE! PLEASE! PLEASE! Pay attention to the entire unit - it's not "grams" it's "grams of something"

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Converting M to % by mass

$\frac{\text{g solute}}{100 \text{ g solution}}$

$\frac{\text{Moles solute}}{\text{Liter solution}}$

If I want to convert the numerators, what do I need to do?

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Converting M to % by mass

$\frac{\text{g solute}}{100 \text{ g solution}}$ $\frac{\text{Moles solute}}{\text{Liter solution}}$

If I want to convert the numerators, what do I need to do?
(Do I need to say UNITS! UNITS! UNITS! again)?

g solute → moles solute

What's the conversion factor?

$\frac{\text{g solute}}{\text{moles solute}}$ Or $\frac{\text{moles solute}}{\text{g solute}}$

What do we call this?

Molar mass!

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Converting M to % by mass

$\frac{\text{g solute}}{100 \text{ g solution}} \times \frac{\text{moles solute}}{\text{g solute}} = \frac{\text{moles solute}}{100 \text{ g solution}}$

Or, if you're going the other way:

$\frac{\text{moles solute}}{\text{L solution}} \times \frac{\text{g solute}}{\text{moles solute}} = \frac{\text{g solute}}{\text{L solution}}$

What about the denominator?

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Converting M to % by mass

$\frac{\text{moles solute}}{100 \text{ g solution}}$ $\frac{\text{g solute}}{\text{Liter solution}}$

100 g solution \rightarrow L solution

What's the conversion factor?

$\frac{\text{g solution}}{\text{L solution}}$ OR $\frac{\text{L solution}}{\text{g solution}}$.

What's this called?
Density! We need to know the density of the solution and the molar mass of the solute to be able to convert M to % by mass.

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Sample problem

You have a 1.2 M solution of NaCl in water.
What is its % by mass given that the density of 1.2 M NaCl is 1.08 g/mL?

What do we do?

Simply convert the units!

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Sample problem

You have a 1.2 M solution of NaCl in water.
What is its % by mass given that the density of 1.2 M NaCl is 1.08 g/mL?

$\frac{1.2 \text{ mol NaCl}}{\text{L solution}}$ \rightarrow $\frac{\text{g NaCl}}{100 \text{ g solution}}$

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Sample problem

You have a 1.2 M solution of NaCl in water.
What is its % by mass given that the density of 1.2 M NaCl is 1.08 g/mL?

$$\frac{1.2 \text{ mol NaCl}}{\text{L solution}} \times \frac{58.453 \text{ g NaCl}}{\text{mol NaCl}} = \frac{70.143 \text{ g NaCl}}{\text{L solution}}$$

$$\frac{70.143 \text{ g NaCl}}{\text{L solution}} = \frac{? \text{ g NaCl}}{100 \text{ g solution}}$$

We have density in g/mL, but we have volume in L, so...?

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Sample problem

You have a 1.2 M solution of NaCl in water.
What is its % by mass given that the density of 1.2 M NaCl is 1.08 g/mL?

$$\frac{1.2 \text{ mol NaCl}}{\text{L solution}} \times \frac{58.453 \text{ g NaCl}}{\text{mol NaCl}} = \frac{70.143 \text{ g NaCl}}{\text{L solution}}$$

$$\frac{70.143 \text{ g NaCl}}{\text{L solution}} \times \frac{1 \text{ L solution}}{1000 \text{ mL solution}} \times \frac{1 \text{ mL solution}}{1.08 \text{ g solution}} = \frac{0.06495 \text{ g NaCl}}{\text{g solution}}$$

To get from g solution to 100 g solution???

Just multiply by 100.

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Sample problem

You have a 1.2 M solution of NaCl in water.
What is its % by mass given that the density of 1.2 M NaCl is 1.08 g/mL?

$$\frac{1.2 \text{ mol NaCl}}{\text{L solution}} \times \frac{58.453 \text{ g NaCl}}{\text{mol NaCl}} = \frac{70.143 \text{ g NaCl}}{\text{L solution}}$$

$$\frac{70.143 \text{ g NaCl}}{\text{L solution}} \times \frac{1 \text{ L solution}}{1000 \text{ mL solution}} \times \frac{1 \text{ mL solution}}{1.08 \text{ g solution}} = \frac{0.06495 \text{ g NaCl}}{\text{g solution}} \times 100$$

6.495 g NaCl/100 g solution OR 6.495% by mass
(6.5% by mass to the correct number of sig figs.)

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Slightly more complicated problem

You have a 2.5 M solution of NaCl in water. What is its % by mass given that the density of 1.0 M NaCl is 1.07 g/mL, the density of 3.0 M NaCl is 1.10 g/mL and the density of 5.0 M NaCl is 1.13 g/mL?

This is a more common real-life example. There are tables of solution densities, but they don't have an infinite number of values. So, what do we do...?

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Slightly more complicated problem

You have a 2.5 M solution of NaCl in water. What is its % by mass given that the density of 1.0 M NaCl is 1.07 g/mL, the density of 3.0 M NaCl is 1.10 g/mL and the density of 5.0 M NaCl is 1.13 g/mL?

2 options:

- Use the 3.0 M as being "close enough".
- Do a linear interpolation to determine the approximate value at 2.5 M.

What is a "linear interpolation"?

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Linear interpolation

Assume that the data should have a linear trend, plot the line of the data you do have and find the data you don't have?

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Linear interpolation

Assume that the data should have a linear trend, plot the line of the data you do have and find the data you don't have?

So, the density appears to be about 1.09 g/mL at 2.5 M.

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The rest of the problem is the same.

You have a 2.5 M solution of NaCl in water. What is its % by mass given that the density of 1.0 M NaCl is 1.06 g/mL, the density of 3.0 M NaCl is 1.09 g/mL and the density of 5.0 M NaCl is 1.13 g/mL?

$$\frac{2.5 \text{ mol NaCl}}{\text{L solution}} \times \frac{58.453 \text{ g NaCl}}{\text{mol NaCl}} = \frac{146.13 \text{ g NaCl}}{\text{L solution}}$$

$$\frac{146.13 \text{ g NaCl}}{\text{L solution}} \times \frac{1 \text{ L solution}}{1000 \text{ mL solution}} \times \frac{1 \text{ mL solution}}{1.09 \text{ g solution}} = \frac{0.134 \text{ g NaCl}}{\text{g solution}} \times 100$$

= 13.4 g NaCl = 13.4% NaCl by mass
100 g solution

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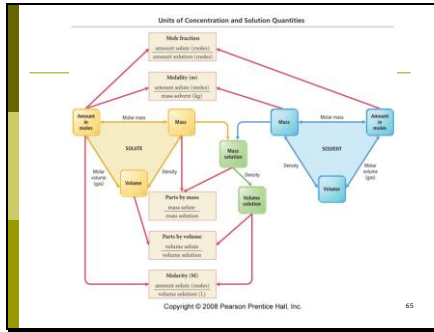
EXAMPLE 12.3 Using Parts by Mass in Calculations
What volume (in mL) of a soft drink that is 10.5% sucrose (C₁₂H₂₂O₁₁) by mass contains 78.3 g of sucrose? (The density of the solution is 1.04 g/mL.)

continued...

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Begin with 78.5 g $C_{12}H_{22}O_{11}$ and multiply by the conversion factors to arrive at the volume of solution.

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Colligative Property

A colligative property is a property that depends ONLY on the amount of the substance present NOT on the identity of the substance.

In other words, it doesn't matter if it is salt, sugar, gasoline, or tennis balls – it will behave the same way!

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Some Examples:

- Vapor Pressure Reduction
 - Related to boiling point
- Freezing Point Depression
 - Salt on the road
 - Anti-freeze in your radiator
- Boiling Point Elevation
 - Anti-freeze in your radiator
- Osmotic Pressure
 - Membrane diffusion
 - The Great Sugar Fountain!

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Vapor Pressure Reduction

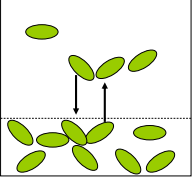
What is "vapor pressure"?

Vapor pressure is the amount (P_A for ideal gases) of gas A that is in equilibrium above the surface of liquid A.

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Vapor Pressure

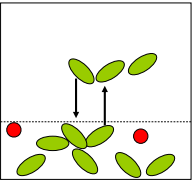
At equilibrium, the rate of evaporation (liquid to gas) equals the rate of condensation (gas to liquid). The amount of gas is the "vapor pressure"



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What if you add a solute?

At equilibrium, the rate of evaporation still must equal the rate of condensation. But at any given temperature, the # of solvent molecules at the surface is decreased and, therefore, so is the vapor pressure




Surface of liquid

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Vapor Pressure of Solutions

- the vapor pressure of a solvent above a solution is lower than the vapor pressure of the pure solvent
 - the solute particles replace some of the solvent molecules at the surface

At equilibrium, the rate of evaporation must equal the rate of condensation. The number of solvent molecules at the surface is reduced, so the rate of evaporation is lower and the vapor pressure will be lower



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Thirsty Solutions

- a concentrated solution will draw solvent molecules toward it due to the natural drive for materials in nature to mix
- similarly, a concentrated solution will draw pure solvent vapor into it due to this tendency to mix
- the result is reduction in vapor pressure

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Thirsty Solutions

When equilibrium is established, the liquid level in the solution beaker is higher than the liquid level in the pure solvent beaker—the thick film of the jar, grabs additional solvent vapor more effectively in the beakers.

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Raoult's Law

- the vapor pressure of a volatile solvent above a solution is equal to its mole fraction of its normal vapor pressure, P°

$$P_{\text{solvent in solution}} = \chi_{\text{solvent}} P^\circ$$
 - since the mole fraction is always less than 1, the vapor pressure of the solvent in solution will always be less than the vapor pressure of the pure solvent

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EX 12.5 – Calculate the vapor pressure of water in a solution prepared by mixing 99.5 g of $C_{12}H_{22}O_{11}$ with 300.0 mL of H_2O

Given:	99.5 g $C_{12}H_{22}O_{11}$, 300.0 mL H_2O
Find:	P_{H_2O}
Concept Plan:	$g C_{12}H_{22}O_{11} \rightarrow mol C_{12}H_{22}O_{11} \xrightarrow{\chi} P = \chi \cdot P^\circ$ $mol C_{12}H_{22}O_{11} \rightarrow \chi_{H_2O} \rightarrow P_{H_2O}$ $mol H_2O \rightarrow P_{H_2O}$
Relationships:	$P_{H_2O}^\circ = 23.8 \text{ torr}$, 1 mol $C_{12}H_{22}O_{11}$ = 342.30 g, 1 mol H_2O = 18.02 g $99.5 g C_{12}H_{22}O_{11} \times \frac{1 \text{ mol } C_{12}H_{22}O_{11}}{342.30 g} = 0.2907 \text{ mol } C_{12}H_{22}O_{11}$ $(0.2907 \text{ mol } C_{12}H_{22}O_{11}) \times \frac{18.02 g}{1 \text{ mol } H_2O} = 5.238 \text{ g } H_2O$ $300.0 \text{ mL } H_2O \times \frac{1.00 g}{1 \text{ mL } H_2O} = 300.0 \text{ g } H_2O$ $P_{H_2O} = \chi_{H_2O} \cdot P_{H_2O}^\circ = \left(\frac{300.0 g}{300.0 g + 5.238 g} \right) (23.8 \text{ torr})$ $P_{H_2O} = 23.4 \text{ torr}$

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Boiling Point Elevation

What is the "boiling point" of a liquid?

It is the temperature at which the vapor pressure equals the atmospheric pressure. So...

...if you decrease the vapor pressure, you must increase the boiling point – it will take a higher temperature to get enough gas molecules (vapor pressure) to equal the atmospheric pressure.

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**Boiling Pt. Elevation
Freezing Pt. Depression**

Birds of a Feather:

$\Delta T_b = T_{\text{boiling, solution}} - T_{\text{boiling, pure solvent}} = K_b m$

$\Delta T_f = T_{\text{freezing, solution}} - T_{\text{freezing, pure solvent}} = -K_f m$

m = molality of the solution
 K_b = boiling constant
 K_f = cryoscopic constant

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NOTE ABOUT THE SIGN CONVENTION

$\Delta T_b = T_{\text{boiling, solution}} - T_{\text{boiling, pure solvent}} = K_b m$

$\Delta T_f = T_{\text{freezing, solution}} - T_{\text{freezing, pure solvent}} = -K_f m$

Δ = change
 Δ = final state – initial state
Many texts do not put a minus sign before the cryoscopic constant and reverse the definition of ΔT_f instead.

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K is a SOLVENT property

Colligative properties don't care what the solute is – they only care how much solute there is.

The boiling point constant and cryoscopic constant are SOLVENT properties only.

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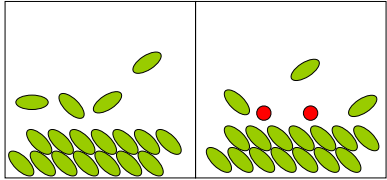
What is the Microscopic View of Freezing Point Depression?

These phase change properties arise from the attractive interactions between all molecules (Van der Waal's forces).

When you make a solution, you break some of the solvent-solvent interactions and replace them with solvent-solute interactions.

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Freezing is a dynamic equilibrium between melting and freezing.



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A simple problem

How many grams of sucrose ($C_{12}H_{22}O_{11}$) are needed to lower the freezing point of 100 g of water by $3^\circ C$?

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The Answer

$\Delta T_f = - K_f m$
We want to decrease the freezing point by $3^\circ C$

$-3^\circ C = -(1.86^\circ C/molal) m$
 $m = 1.61 \text{ molal} = 1.61 \text{ moles solute/kg solvent}$

NOTE: K_f is the WATER cryoscopic constant

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1.61 moles solute = x moles solute
1 kg solvent 0.100 kg water

0.161 moles sucrose * 342 g sucrose = 55.1 g sucrose
1 mole sucrose

342 g/mol = $11 \cdot 12.01 \text{ g/mol} + 22 \cdot 1.008 \text{ g/mol} + 11 \cdot 16 \text{ g/mol}$

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Slight Variation on the problem

How many grams of NaCl are needed to lower the freezing point of 100 g of water by 3 °C?

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The Answer

$\Delta T_f = - K_f m$
We want to decrease the freezing point by 3°C

-3° C = -(1.86 °C/molal) m
m=1.61 molal = 1.61 moles solute/kg solvent

NOTE: K_f is the WATER cryoscopic constant

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1.61 moles solute * 0.100 kg water = 0.161 moles solute
1 kg solvent

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K_f

TABLE 12.8 Freezing Point Depression and Boiling Point Elevation Constants for Several Liquid Solvents

Solvent	Normal Freezing Point (°C)	K_f (°C/m)	Normal Boiling Point (°C)	K_b (°C/m)
Benzene (C ₆ H ₆)	5.5	5.12	80.1	2.53
Carbon tetrachloride (CCl ₄)	-22.9	29.9	76.7	5.03
Chloroform (CHCl ₃)	-63.5	4.70	61.2	3.63
Ethanol (C ₂ H ₅ OH)	-114.1	1.99	78.3	1.22
Diethyl ether (C ₂ H ₅) ₂ O	-116.3	1.79	34.6	2.02
Water (H ₂ O)	0.00	1.86	100.0	0.512

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Ex 12.8 – What is the freezing point of a 1.7 m aqueous ethylene glycol solution, C₂H₄O₂?

Given:	1.7 m C ₂ H ₄ O ₂ (aq)
Find:	T _f , °C
Concept Plan:	$m \rightleftharpoons \Delta T_f$ $\Delta T_f = m \cdot K_f$
Relationships:	$\Delta T_f = m \cdot K_f$, K_f for H ₂ O = 1.86 °C/m, $FP_{H_2O} = 0.00^\circ C$
Solve:	$\Delta T_f = m \cdot K_{f,H_2O} \quad FP_{H_2O} - FP_{soln} = \Delta T_f$ $= (1.7 m) \left(1.86 \frac{^\circ C}{m} \right) \quad 0.00^\circ C - FP_{soln} = 3.2^\circ C$ $\Delta T_f = 3.2^\circ C \quad FP_{soln} = -3.2^\circ C$
Check:	the unit is correct, the freezing point lower than the normal freezing point makes sense

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Colligative Properties

- **colligative properties** are properties whose value depends only on the number of solute particles, and not on what they are
 - Vapor Pressure Depression, Freezing Point Depression, Boiling Point Elevation, Osmotic Pressure
- the **van't Hoff factor, i** , is the ratio of moles of solute particles to moles of formula units dissolved
- measured van't Hoff factors are often lower than you might expect due to ion pairing in solution

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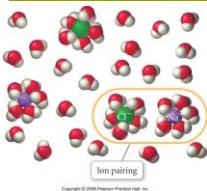


TABLE 12.9 Van't Hoff Factors at 0.05 m Concentration in Aqueous Solution

Solute	<i>i</i> Expected	<i>i</i> Measured
Nonelectrolyte	1	1
NaCl	2	1.9
MgSO ₄	2	1.3
MgCl ₂	3	2.7
K ₂ SO ₄	3	2.6
FeCl ₃	4	3.4

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It's all about the # of particles

1.61 moles solute * 0.100 kg water = 0.161 moles solute
1 kg solvent

BUT NaCl is an electrolyte:

$$\text{NaCl} \rightleftharpoons \text{Na}^+ + \text{Cl}^-$$

You get 2 moles of solute per mole NaCl

$$0.161 \text{ moles solute} * \frac{1 \text{ mol NaCl}}{2 \text{ mol solute}} * \frac{58.45 \text{ g NaCl}}{1 \text{ mole NaCl}} = 4.7 \text{ g NaCl}$$

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Osmotic Pressure

- Osmotic pressure is the "funky" colligative property, but it is very important biologically
- Osmotic pressure is the pressure required to overcome the natural pressure exerted by a solution by virtue of having a concentration.
- Osmotic pressure looks just like the ideal gas law:
$$\Pi V = nRT$$

where Π is osmotic pressure
$$\Pi = (n/V) RT = M RT$$

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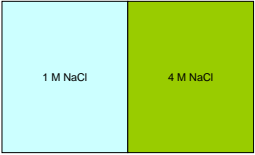
Why is it important?

Consider two solutions separated by a semi-permeable membrane (meaning water can pass through but nothing else) at room temperature (298 K).

On one side of the membrane is a 1 M NaCl solution. On the other side is a 4 M NaCl solution.

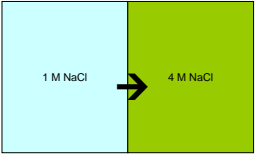
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What Happens?



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The water moves from the 1 M side to the 4 M side. Why?



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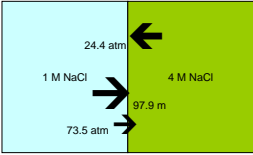
OSMOTIC PRESSURE
 $\Pi = M RT$

On the 1 M side, the pressure is:
 $\Pi = M RT = 1 M (0.0821 \text{ Latm/mol K})(298 \text{ K})$
 $\Pi = 24.4 \text{ atm}$

On the 4 M side, the pressure is:
 $\Pi = M RT = 4 M (0.0821 \text{ Latm/mol K}) (298 \text{ K})$
 $\Pi = 97.9 \text{ atm}$

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Both solutions “push” on the membrane
The bigger push wins!



Note the direction of the arrows. Osmotic pressure is pushing AGAINST the solution.

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It's all about sweating!

□ What happens when you sweat?

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It's all about sweating!

What happens when you sweat?

Your body loses water and some dissolved salts through your skin which cools your body.

What is the price you pay for this?

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The cost of sweating

Your cells have lost precious water and salts.

What do you do about it?

Drink water?

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The problem with water

Your cells are slightly dehydrated from sweating. The salt concentration is actually higher than it should be.

Water has no salt concentration. It flows into the cells rapidly (osmotic pressure) and fills the cells. But the cells are now diluted.

Further, since the pure water will always have lower osmotic pressure and flow into the cell until either you run out of water or the cell bursts.

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Gatorade – Man’s answer to Osmotic Pressure

Gatorade is flavored sweat!

By adding salts, it matches its osmotic pressure to what the cell should be. This allows for smooth rehydrating of the cell without dilution or lysing.

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EXAMPLE 12.4 Calculating Concentrations

A solution is prepared by dissolving 17.2 g of ethylene glycol ($C_2H_6O_2$) in 0.500 kg of water. The final volume of the solution is 513 mL. For this solution, calculate each of the following:

(a) molarity (b) molality (c) percent by mass (d) mole fraction (e) mole percent

continued...

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(e) To calculate percent by mass, divide the mass of the solute by the sum of the masses of the solute and solvent and multiply the ratio by 100%.

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EXAMPLE 12.5 Calculating the Vapor Pressure of a Solution Containing a Nonionic and Nonvolatile Solute

Calculate the vapor pressure at 25 °C of a solution containing 99.3 g sucrose ($C_{12}H_{22}O_{11}$) and 300.0 mL water. The vapor pressure of pure water at 25 °C is 23.8 torr. Assume the density of water to be 1.00 g/mL.

continued...

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Solve Calculate the number of moles of each solution component.

Component	Mass (g)	Molar Mass (g/mol)	Moles (mol)

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EXAMPLE 12.6 Calculating the Vapor Pressure of a Solution Containing an Ionic Solute

A solution contains 0.102 mol $Ca(NO_3)_2$ and 0.927 mol H_2O . Calculate the vapor pressure of the solution at 35 °C. The vapor pressure of pure water at 35 °C is 42.2 torr.

continued...

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EXAMPLE 12.11 Van't Hoff Factor and Freezing Point Depression
The freezing point of an aqueous 0.850 m CaCl_2 solution is -6.27°C . What is the van't Hoff factor (i) for CaCl_2 at this concentration? How does it compare to the predicted value of i ?
