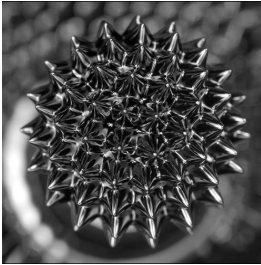


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Chapter 12

Physical Properties of Solutions

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Solutions

A **solution** is a homogenous mixture of 2 or more substances.

The **solute** is(are) the substance(s) present in the smaller amount(s).

The **solvent** is the substance present in the larger amount.

Table 12.1 Types of Solutions

Component 1	Component 2	State of Resulting Solution	Examples
Gas	Gas	Gas	Air
Gas	Liquid	Liquid	Soda water (CO ₂ in water)
Gas	Solid	Solid	H ₂ gas in palladium
Liquid	Liquid	Liquid	Ethanol in water
Solid	Liquid	Liquid	NaCl in water
Solid	Solid	Solid	Brass (Cu/Zn), solder (Sn/Pb)

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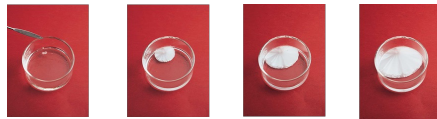
Saturation

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.

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

Sodium acetate crystals rapidly form when a seed crystal is added to a supersaturated solution of sodium acetate.



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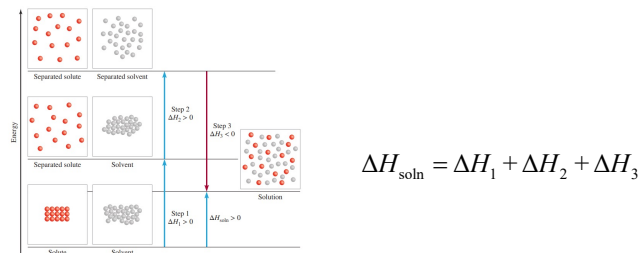
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Formation of a Solution

Three types of interactions in the solution process:

- **solvent-solvent** interaction
- **solute-solute** interaction
- **solvent-solute** interaction

Molecular view of the formation of solution



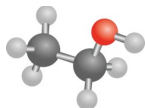
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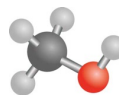
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“Like Dissolves Like”



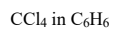
C_2H_5OH



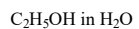
CH_3OH

Two substances with similar *intermolecular* forces are likely to be soluble in each other.

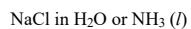
non-polar molecules are soluble in non-polar solvents



polar molecules are soluble in polar solvents



ionic compounds are more soluble in polar solvents



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Example 12.1

Predict the relative μ solubilities in the following cases:

- Bromine (Br_2) in benzene ($C_6H_6, \mu = 0$ D) and in water ($\mu = 1.87$ D)
- KCl in carbon tetrachloride ($CCl_4, \mu = 0$ D) and in liquid ammonia ($NH_3, \mu = 1.46$ D)
- Formaldehyde (CH_2O) in carbon disulfide ($CS_2, \mu = 0$ D) and in water

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Example 12.1 ₂

Strategy

In predicting solubility, remember the saying: Like dissolves like. A nonpolar solute will dissolve in a nonpolar solvent; ionic compounds will generally dissolve in polar solvents due to favorable ion-dipole interaction; solutes that can form hydrogen bonds with the solvent will have high solubility in the solvent.

Solution

- a) Br_2 is a nonpolar molecule and therefore should be more soluble in C_6H_6 , which is also nonpolar, than in water. The only intermolecular forces between Br_2 and C_6H_6 are dispersion forces.

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Example 12.1 ₃

- b) KCl is an ionic compound. For it to dissolve, the individual K^+ and Cl^- ions must be stabilized by ion-dipole interaction. Because CCl_4 has no dipole moment, KCl should be more soluble in liquid NH_3 , a polar molecule with a large dipole moment.
- c) Because CH_2O is a polar molecule and CS_2 (a linear molecule) is nonpolar, the forces between molecules of CH_2O and CS_2 are dipole-induced dipole and dispersion. On the other hand, CH_2O can form hydrogen bonds with water, so it should be more soluble in that solvent.

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Concentration Units 1

The **concentration** of a solution is the amount of solute present in a given quantity of solvent or solution.

Percent by Mass

$$\begin{aligned}\% \text{ by mass} &= \frac{\text{mass of solute}}{\text{mass of solute} + \text{mass of solvent}} \times 100\% \\ &= \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%\end{aligned}$$

Mole Fraction (X)

$$X_A = \frac{\text{moles of A}}{\text{sum of moles of all components}}$$

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Example 12.2 1

A sample of 0.892 g of potassium chloride (KCl) is dissolved in 54.6 g of water.

What is the percent by mass of KCl in the solution?

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Example 12.2 ₂

Strategy

We are given the mass of a solute dissolved in a certain amount of solvent. Therefore, we can calculate the mass percent of KCl using Equation (12.1).

Solution

We write

$$\begin{aligned} \text{percent by mass of KCl} &= \frac{\text{mass of solute}}{\text{mass of soln}} \times 100\% \\ &= \frac{0.892 \text{ g}}{0.892 \text{ g} + 54.6 \text{ g}} \times 100\% \\ &= 1.61\% \end{aligned}$$

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Concentration Units ₂

Molarity (*M*)

$$M = \frac{\text{moles of solute}}{\text{liters of solution}}$$

Molality (*m*)

$$m = \frac{\text{moles of solute}}{\text{liters of solvent (K.g)}}$$

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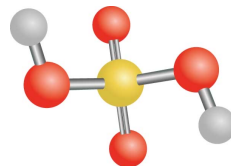
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Example 12.3 ₁

Calculate the molality of a sulfuric acid solution containing 24.4 g of sulfuric acid in 198 g of water.

The molar mass of sulfuric acid is 98.09 g.



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Example 12.3 ₂

Strategy

To calculate the molality of a solution, we need to know the number of moles of solute and the mass of the solvent in kilograms.

Solution

The definition of molality (m) is

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

First, we find the number of moles of sulfuric acid in 24.4 g of the acid, using its molar mass as the conversion factor

$$\begin{aligned} \text{moles of H}_2\text{SO}_4 &= 24.4 \text{ g } \cancel{\text{H}_2\text{SO}_4} \times \frac{1 \text{ mol H}_2\text{SO}_4}{98.09 \text{ g } \cancel{\text{H}_2\text{SO}_4}} \\ &= 0.249 \text{ mol H}_2\text{SO}_4 \end{aligned}$$

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Example 12.3 ₃

The mass of water is 198 g, or 0.198 kg. Therefore,

$$m = \frac{0.249 \text{ mol H}_2\text{SO}_4}{0.198 \text{ kg H}_2\text{O}}$$
$$= 1.26 m$$

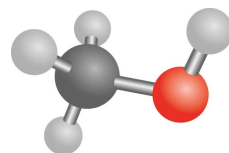
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Example 12.4 ₁

The density of a 2.45 M aqueous solution of methanol (CH₃OH) is 0.976 g/mL. What is the molality of the solution? The molar mass of methanol is 32.04 g.



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Example 12.4₂

Strategy

To calculate the molality, we need to know the number of moles of methanol and the mass of solvent in kilograms. We assume 1 L of solution, so the number of moles of methanol is 2.45 mol.

$$m \text{ (want to calculate)} = \frac{\text{moles of solute (given)}}{\text{mass of solvent (kg) (need to find)}}$$

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Example 12.4₃

Solution

Our first step is to calculate the mass of water in 1 L of the solution, using density as a conversion factor. The total mass of 1 L of a 2.45 *M* solution of methanol is

$$1 \cancel{\text{ L soln}} \times \frac{1000 \cancel{\text{ mL soln}}}{1 \cancel{\text{ L soln}}} \times \frac{0.976 \text{ g}}{1 \cancel{\text{ mL soln}}} = 976 \text{ g}$$

Because this solution contains 2.45 moles of methanol, the amount of water (solvent) in the solution is

Mass of H₂O = mass of soln – mass of solute

$$\begin{aligned} &= 976 \text{ g} - \left(2.45 \cancel{\text{ mol CH}_3\text{OH}} \times \frac{32.04 \text{ g CH}_3\text{OH}}{1 \cancel{\text{ mol CH}_3\text{OH}}} \right) \\ &= 898 \text{ g} \end{aligned}$$

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Example 12.4 ₄

The molality of the solution can be calculated by converting 898 g to 0.898 kg:

$$\begin{aligned}\text{molality} &= \frac{2.45 \text{ mol CH}_3\text{OH}}{0.898 \text{ kg H}_2\text{O}} \\ &= 2.73 \text{ m}\end{aligned}$$

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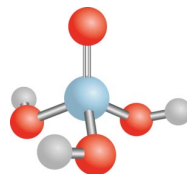
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Example 12.5 ₁

Calculate the molality of a 35.4 percent (by mass) aqueous solution of phosphoric acid (H₃PO₄).

The molar mass of phosphoric acid is 97.99 g.



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Example 12.5₂

Strategy

In solving this type of problem, it is convenient to assume that we start with a 100.0 g of the solution. If the mass of phosphoric acid is 35.4 percent, or 35.4 g, the percent by mass and mass of water must be $100.0\% - 35.4\% = 64.6\%$ and 64.6 g.

Solution

From the known molar mass of phosphoric acid, we can calculate the molality in two steps, as shown in Example 12.3. First, we calculate the number of moles of phosphoric acid in 35.4 g of the acid

$$\begin{aligned} \text{moles of H}_3\text{PO}_4 &= 35.4 \cancel{\text{g H}_3\text{PO}_4} \times \frac{1 \text{ mol H}_3\text{PO}_4}{97.99 \cancel{\text{g H}_3\text{PO}_4}} \\ &= 0.361 \text{ mol H}_3\text{PO}_4 \end{aligned}$$

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Example 12.5₃

The mass of water is 64.6 g, or 0.0646 kg. Therefore, the molality is given by

$$\begin{aligned} \text{molality} &= \frac{0.361 \text{ mol H}_3\text{PO}_4}{0.0646 \text{ kg H}_2\text{O}} \\ &= 5.59 \text{ m} \end{aligned}$$

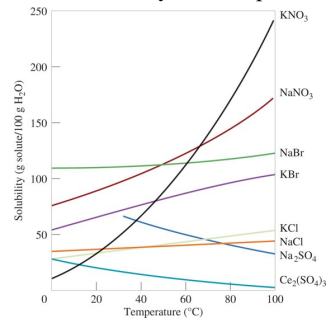
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Temperature and Solubility 1

Solid solubility and temperature



solubility increases with increasing temperature

solubility decreases with increasing temperature

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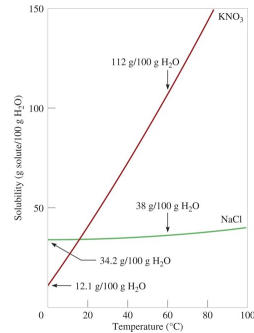
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Fractional Crystallization

Fractional crystallization is the separation of a mixture of substances into pure components on the basis of their differing solubilities.



Suppose you have 90 g KNO₃ contaminated with 10 g NaCl

Fractional crystallization:

1. Dissolve sample in 100 mL of water at 60°C
2. Cool solution to 0°C
3. All NaCl will stay in solution ($s = 34.2 \text{ g}/100 \text{ g}$)
4. 78 g of PURE KNO₃ will precipitate ($s = 12 \text{ g}/100 \text{ g}$). $90 \text{ g} - 12 \text{ g} = 78 \text{ g}$

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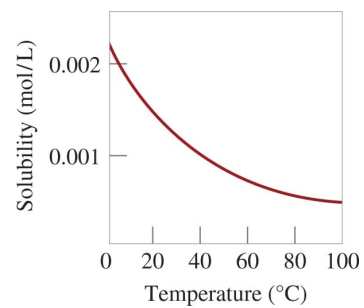
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Temperature and Solubility 2

O₂ gas solubility and temperature



**solubility usually
decreases with
increasing temperature**

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Pressure and Solubility of Gases

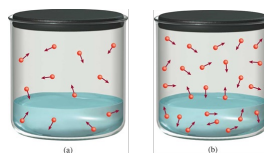
The solubility of a gas in a liquid is proportional to the pressure of the gas over the solution (**Henry's law**).

$$c = kP$$

c is the concentration (M) of the dissolved gas

P is the pressure of the gas over the solution

k is a constant for each gas ($\text{mol/L} \cdot \text{atm}$) that depends only on temperature



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Example 12.6 ₁

The solubility of nitrogen gas at 25°C and 1 atm is 6.8×10^{-4} mol/L.

What is the concentration (in molarity) of nitrogen dissolved in water under atmospheric conditions?

The partial pressure of nitrogen gas in the atmosphere is 0.78 atm.

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Example 12.6 ₂

Strategy

The given solubility enables us to calculate Henry's law constant (k), which can then be used to determine the concentration of the solution.

Solution

The first step is to calculate the quantity k in Equation (12.3):

$$c = kP$$

$$6.8 \times 10^{-4} \text{ mol/L} = k(1 \text{ atm})$$

$$k = 6.8 \times 10^{-4} \text{ mol/L} \cdot \text{atm}$$

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Example 12.6 ₃

Therefore, the solubility of nitrogen gas in water is

$$\begin{aligned} c &= (6.8 \times 10^{-4} \text{ mol/L} \cdot \text{atm})(0.78 \text{ atm}) \\ &= 5.3 \times 10^{-4} \text{ mol/L} \\ &= 5.3 \times 10^{-4} M \end{aligned}$$

The decrease in solubility is the result of lowering the pressure from 1 atm to 0.78 atm.

Check

The concentration ratio $[(5.3 \times 10^{-4} M / 6.8 \times 10^{-4} M) = 0.78]$ should be equal to the ratio of the pressures (0.78 atm/1.0 atm = 0.78).

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Chemistry In Action: The Killer Lake

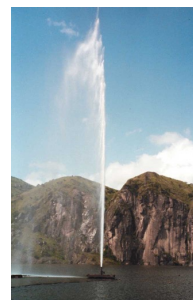
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CO₂ Cloud Released

1700 Casualties

Trigger?

- earthquake
- landslide
- strong winds



Lake Nyos, West Africa

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Source: Bill Evans, U.S. Geological Survey

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Colligative Properties of Nonelectrolyte Solutions ₁

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

Vapor-Pressure Lowering

$$P_1 = X_1 P_1^0 \quad P_1^0 = \text{vapor pressure of pure solvent}$$

Raoult's law $X_1 = \text{mole fraction of the solvent}$

If the solution contains only one solute:

$$X_1 = 1 - X_2$$

$$P_1^0 - P_1 = \Delta P = X_2 P_1^0 \quad X_2 = \text{mole fraction of the solute}$$

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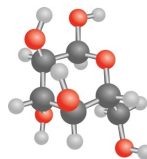
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Example 12.7 ₁

Calculate the vapor pressure of a solution made by dissolving 218 g of glucose (molar mass = 180.2 g/mol) in 460 mL of water at 30°C.

What is the vapor-pressure lowering?

The vapor pressure of pure water at 30°C is given in Table 5.3 (p. 199). Assume the density of the solvent is 1.00 g/mL.



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Example 12.7₂

Strategy

We need Raoult's law [Equation (12.4)] to determine the vapor pressure of a solution. Note that glucose is a nonvolatile solute.

Solution

The vapor pressure of a solution (P_1) is

$$P_1 (\text{want to calculate}) = X_1 (\text{need to find}) P_1^\circ (\text{given})$$

First we calculate the number of moles of glucose and water in the solution:

$$n_1 (\text{water}) = 460 \text{ mL} \times \frac{1.00 \text{ g}}{1 \text{ mL}} \times \frac{1 \text{ mol}}{18.02 \text{ g}} = 25.5 \text{ mol}$$

$$n_2 (\text{glucose}) = 218 \text{ g} \times \frac{1 \text{ mol}}{180.2 \text{ g}} = 1.21 \text{ mol}$$

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Example 12.7₃

The mole fraction of water, X_1 , is given by

$$X_1 = \frac{n_1}{n_1 + n_2}$$

$$= \frac{25.5 \text{ mol}}{25.5 \text{ mol} + 1.21 \text{ mol}} = 0.955$$

From Table 5.3, we find the vapor pressure of water at 30° C to be 31.82 mmHg. Therefore, the vapor pressure of the glucose solution is

$$P_1 = 0.955 \times 31.82 \text{ mmHg}$$

$$= 30.4 \text{ mmHg}$$

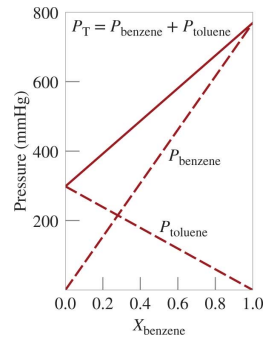
Finally, the vapor-pressure lowering (ΔP) is (31.82 – 30.4) mmHg, or 1.4 mmHg.

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Partial Pressure and Mole Fraction



$$P_A = X_A P_A^0$$

$$P_B = X_B P_B^0$$

$$P_T = P_A + P_B$$

$$P_T = X_A P_A^0 + X_B P_B^0$$

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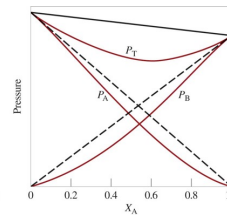
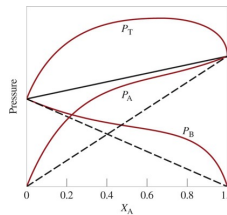
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Raoult's Law

P_T is greater than predicted by Raoult's law

P_T is less than predicted by Raoult's law



Force A-B < Force A-A & B-B

Force A-B > Force A-A & B-B

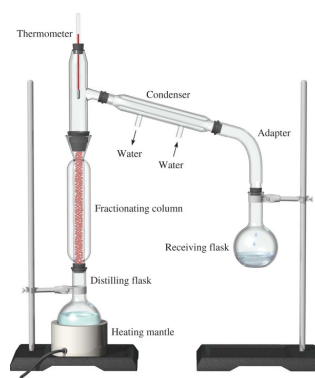
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Fractional Distillation Apparatus



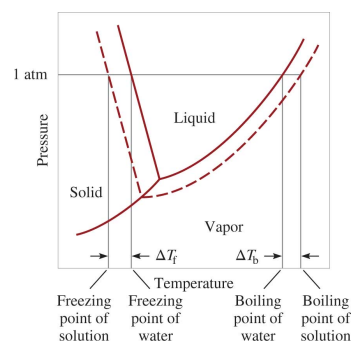
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Boiling-Point Elevation



$$\Delta T_b = T_b - T_b^0$$

T_b^0 is the boiling point of the pure solvent

T_b is the boiling point of the solution

$$T_b > T_b^0 \quad \Delta T_b > 0$$

$$\Delta T_b = K_b m$$

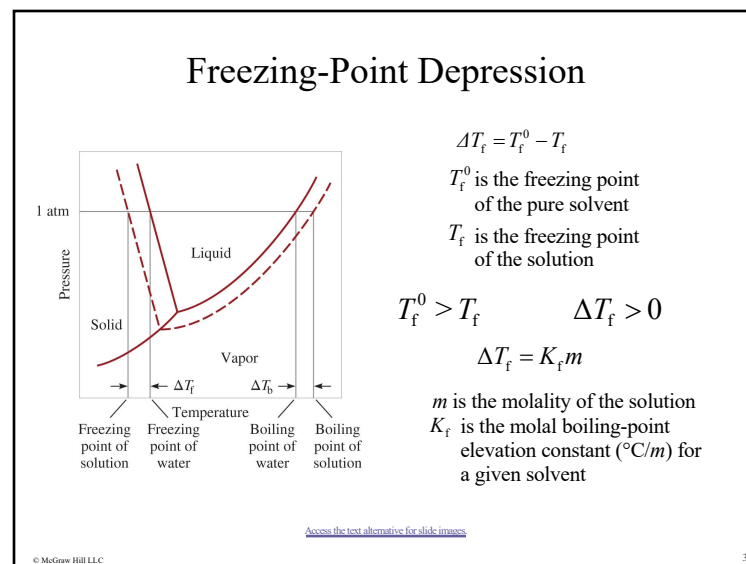
m is the molality of the solution
 K_b is the molal boiling-point elevation constant ($^{\circ}\text{C}/m$) for a given solvent

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Boiling-Point Elevation and Freezing-Point Depression Constants

Table 12.2 Molal Boiling-Point Elevation and Freezing-Point Depression Constants of Several Common Liquids

Solvent	Normal Freezing Point ($^{\circ}\text{C}$)*	K_f ($^{\circ}\text{C}/m$)	Normal Boiling Point ($^{\circ}\text{C}$)*	K_b ($^{\circ}\text{C}/m$)
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.

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Example 12.8 ₁

Ethylene glycol (EG), $\text{CH}_2(\text{OH})\text{CH}_2(\text{OH})$, is a common automobile antifreeze. It is water soluble and fairly nonvolatile (b.p. 197°C).

Calculate the freezing point of a solution containing 651 g of this substance in 2505 g of water.

Would you keep this substance in your car radiator during the summer?

The molar mass of ethylene glycol is 62.01 g.

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Example 12.8 ₂

Strategy

This question asks for the depression in freezing point of the solution.

$$\Delta T_f = K_f m$$

← want to calculate
← constant
← need to find

The information given enables us to calculate the molality of the solution and we refer to Table 12.2 for the K_f of water.

Solution

To solve for the molality of the solution, we need to know the number of moles of EG and the mass of the solvent in kilograms.

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Example 12.8₃

We find the molar mass of EG, and convert the mass of the solvent to 2.505 kg, and calculate the molality as follows:

$$651 \cancel{\text{g EG}} \times \frac{1 \text{ mol EG}}{62.07 \cancel{\text{g EG}}} = 10.5 \text{ mol EG}$$

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

$$= \frac{10.5 \text{ mol EG}}{2.505 \text{ kg H}_2\text{O}} = 4.19 \text{ mol EG/kg H}_2\text{O}$$

$$= 4.19 \text{ } m$$

From Equation (12.7) and Table 12.2 we write

$$\Delta T_f = K_f m$$

$$= (1.86^\circ\text{C}/m)(4.19 \text{ } m)$$

$$= 7.79^\circ\text{C}$$

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Example 12.8₄

Because pure water freezes at 0°C , the solution will freeze at $(0 - 7.79)^\circ\text{C}$ or -7.79°C .

We can calculate boiling-point elevation in the same way as follows:

$$\Delta T_b = K_b m$$

$$= (0.52^\circ\text{C}/m)(4.19 \text{ } m)$$

$$= 2.2^\circ\text{C}$$

Because the solution will boil at $(100 + 2.2)^\circ\text{C}$, or 102.2°C , it would be preferable to leave the antifreeze in your car radiator in summer to prevent the solution from boiling.

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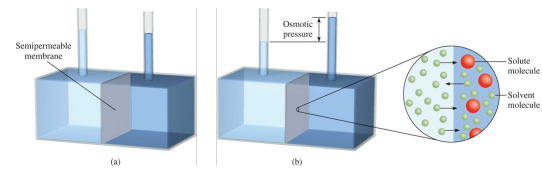
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Osmotic Pressure (π)₁

Osmosis is the selective passage of solvent molecules through a porous membrane from a dilute solution to a more concentrated one.

A **semipermeable membrane** allows the passage of solvent molecules but blocks the passage of solute molecules.

Osmotic pressure (π) is the pressure required to stop osmosis.



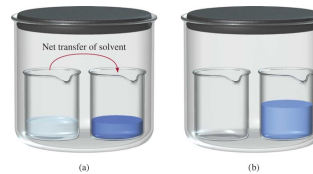
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Osmotic Pressure (π)₂



High P Low P

$$\pi = MRT$$

M is the molarity of the solution

R is the gas constant

T is the temperature (in K)

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A cell in an:

● Water molecules
● Solute molecules

(a)

(b)

(c)

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Colligative Properties of Nonelectrolyte Solutions ₂

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

Vapor-Pressure Lowering $P_1 = X_1 P_1^0$

Boiling – Point Elevation $\Delta T_b = K_b m$

Freezing – Point Depression $\Delta T_f = K_f m$

Osmotic Pressure (π) $\pi = MRT$

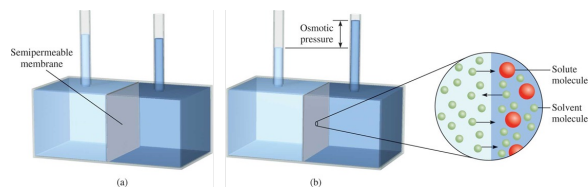
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Example 12.9₁

The average osmotic pressure of seawater, measured in the kind of apparatus shown in Figure 12.11, is about 30.0 atm at 25°C. Calculate the molar concentration of an aqueous solution of sucrose ($C_{12}H_{22}O_{11}$) that is isotonic with seawater.



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Example 12.9₂

Strategy

When we say the sucrose solution is isotonic with seawater, what can we conclude about the osmotic pressures of these two solutions?

Solution

A solution of sucrose that is isotonic with seawater must have the same osmotic pressure, 30.0 atm. Using Equation (12.8).

$$\begin{aligned}\pi &= MRT \\ M &= \frac{\pi}{RT} = \frac{30.0 \text{ atm}}{(0.0821 \text{ L} \cdot \text{atm} / \text{K} \cdot \text{mol})(298 \text{ K})} \\ &= 1.23 \text{ mol} / \text{L} \\ &= 1.23 \text{ M}\end{aligned}$$

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Example 12.10₁

A 7.85-g sample of a compound with the empirical formula C_5H_4 is dissolved in 301 g of benzene.

The freezing point of the solution is 1.05°C below that of pure benzene.

What are the molar mass and molecular formula of this compound?

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Example 12.10₂

Strategy

Solving this problem requires three steps. First, we calculate the molality of the solution from the depression in freezing point.

Next, from the molality we determine the number of moles in 7.85 g of the compound and hence its molar mass.

Finally, comparing the experimental molar mass with the empirical molar mass enables us to write the molecular formula.

Solution

The sequence of conversions for calculating the molar mass of the compound is

Freezing – point depression \rightarrow molality \rightarrow number of moles \rightarrow molar mass

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Example 12.10₃

Our first step is to calculate the molality of the solution. From Equation (12.7) and Table 12.2 we write

$$\text{molality} = \frac{\Delta T_f}{K_f} = \frac{1.05^\circ\text{C}}{5.12^\circ\text{C}/m} = 0.205\ m$$

Because there is 0.205 mole of the solute in 1 kg of solvent, the number of moles of solute in 301 g, or 0.301 kg, of solvent is

$$0.301\ \text{kg} \times \frac{0.205\ \text{mol}}{1\ \text{kg}} = 0.0617\ \text{mol}$$

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Example 12.10₄

Thus, the molar mass of the solute is

$$\begin{aligned} \text{molar mass} &= \frac{\text{grams of compound}}{\text{moles of compound}} \\ &= \frac{7.85\ \text{g}}{0.0617\ \text{mol}} = 127\ \text{g/mol} \end{aligned}$$

Now we can determine the ratio

$$\frac{\text{molar mass}}{\text{empirical molar mass}} = \frac{127}{64\ \text{g/mol}} \approx 2$$

Therefore, the molecular formula is $(\text{C}_5\text{H}_4)_2$ or C_{10}H_8 (naphthalene).

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Example 12.11 ₁

A solution is prepared by dissolving 35.0 g of hemoglobin (Hb) in enough water to make up 1 L in volume.

If the osmotic pressure of the solution is found to be 10.0 mmHg at 25°C, calculate the molar mass of hemoglobin.

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Example 12.11 ₂

Strategy

We are asked to calculate the molar mass of Hb. The steps are similar to those outlined in Example 12.10.

From the osmotic pressure of the solution, we calculate the molarity of the solution.

Then, from the molarity, we determine the number of moles in 35.0 g of Hb and hence its molar mass. What units should we use for π and temperature?

Solution

The sequence of conversions is as follows:

osmotic pressure \rightarrow molality \rightarrow number of moles \rightarrow molar mass

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Example 12.11 ₃

First we calculate the molarity using Equation (12.8)

$$\pi = MRT$$

$$M = \frac{\pi}{RT}$$

$$\begin{aligned} &= \frac{10.0 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}}{(0.0821 \text{ L} \cdot \text{atm} / \text{K} \cdot \text{mol})(298 \text{ K})} \\ &= 5.38 \times 10^{-4} M \end{aligned}$$

The volume of the solution is 1 L, so it must contain 5.38×10^{-4} mol of Hb.

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Example 12.11 ₄

We use this quantity to calculate the molar mass:

$$\begin{aligned} \text{moles of Hb} &= \frac{\text{mass of Hb}}{\text{molar mass of Hb}} \\ \text{molar mass of Hb} &= \frac{\text{mass of Hb}}{\text{moles of Hb}} \\ &= \frac{35.0 \text{ g}}{5.38 \times 10^{-4} \text{ mol}} \\ &= 6.51 \times 10^4 \text{ g/mol} \end{aligned}$$

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Colligative Properties of Electrolyte Solutions ₁

0.1 *m* NaCl solution → 0.1 *m* Na⁺ ions & 0.1 *m* Cl⁻ ions

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

0.1 *m* NaCl solution → 0.2 *m* ions in solution

van't Hoff factor (*i*) = $\frac{\text{actual number of particles in soln after dissociation}}{\text{number of formula units initially dissolved in soln}}$

	<i>i</i> should be
nonelectrolytes	1
NaCl	2
CaCl ₂	3

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Colligative Properties of Electrolyte Solutions ₂

Boiling-Point Elevation $\Delta T_b = iK_b m$

Boiling-Point Depression $\Delta T_f = iK_f m$

Osmotic Pressure (π) $\pi = iMRT$

Table 12.3 The van't Hoff Factor of 0.0500 *M* Electrolyte Solutions at 25°C

Electrolyte	<i>i</i> (Measured)	<i>i</i> (Calculated)
Sucrose*	1.0	1.0
HCl	1.9	2.0
NaCl	1.9	2.0
MgSO ₄	1.3	2.0
MgCl ₂	2.7	3.0
FeCl ₃	3.4	4.0

*Source is a nonelectrolyte. It is listed here for comparison only.

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Example 12.12₁

The osmotic pressure of a 0.010 *M* potassium iodide (KI) solution at 25°C is 0.465 atm.

Calculate the van't Hoff factor for KI at this concentration.

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Example 12.12₂

Strategy

Note that KI is a strong electrolyte, so we expect it to dissociate completely in solution. If so, its osmotic pressure would be

$$2(0.010 M)(0.0821 \text{ L} \cdot \text{atm} / \text{K} \cdot \text{mol})(298 \text{ K}) = 0.49 \text{ atm}$$

However, the measured osmotic pressure is only 0.465 atm. The smaller than predicted osmotic pressure means that there is ion-pair formation, which reduces the number of solute particles (K^+ and I^- ions) in solution.

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Example 12.12 ₃

Solution

From Equation (12.12) we have

$$\begin{aligned}
 i &= \frac{\pi}{MRT} \\
 &= \frac{0.465 \text{ atm}}{(0.010 \text{ M})(0.0821 \text{ L} \cdot \text{atm} / \text{K} \cdot \text{mol})(298 \text{ K})} \\
 &= 1.9
 \end{aligned}$$

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Colloids

A *colloid* is a dispersion of particles of one substance throughout a dispersing medium of another substance.

Colloid versus solution

- colloidal particles are much larger than solute molecules
- colloidal suspension is not as homogeneous as a solution
- colloids exhibit the Tyndall effect



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Types of Colloids

Table 12.4 Types of Colloids

Dispersing Medium	Dispersed Phase	Name	Example
Gas	Liquid	Aerosol	Fog, mist
Gas	Solid	Aerosol	Smoke
Liquid	Gas	Foam	Whipped cream
Liquid	Liquid	Emulsion	Mayonnaise
Liquid	Solid	Sol	Milk of magnesia
Solid	Gas	Foam	Plastic foams
Solid	Liquid	Gel	Jelly, butter
Solid	Solid	Solid sol	Certain alloys (steel), opal

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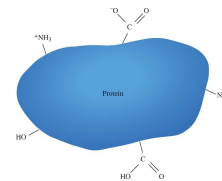
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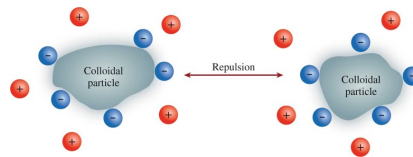
Hydrophilic and Hydrophobic Colloids

Hydrophilic: water-loving

Hydrophobic: water-fearing



Stabilization of a hydrophobic colloid

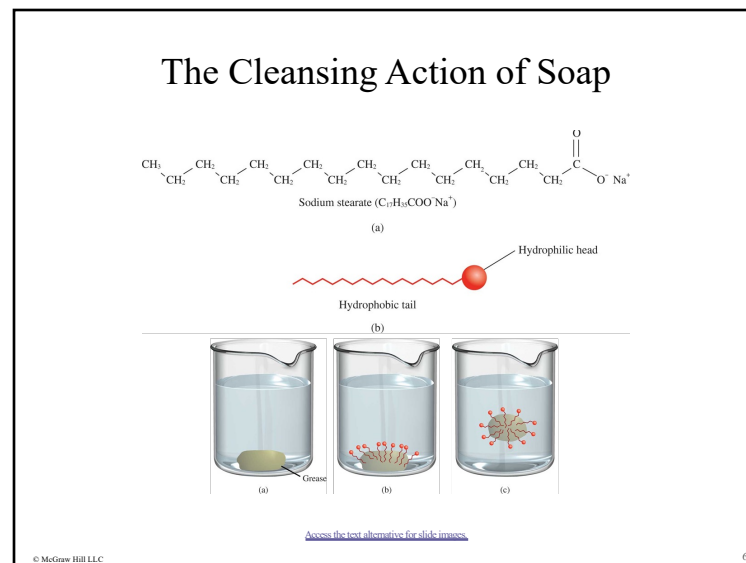


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