

Chapter 13. Properties of Solutions

Common Student Misconceptions

- Students often confuse dilute and concentrated with weak and strong.
- Students do not realize that crystallization is the reverse of dissolution.
- Students often forget that calculations of molality require the mass of *solvent*, not *solution*.

Lecture Outline

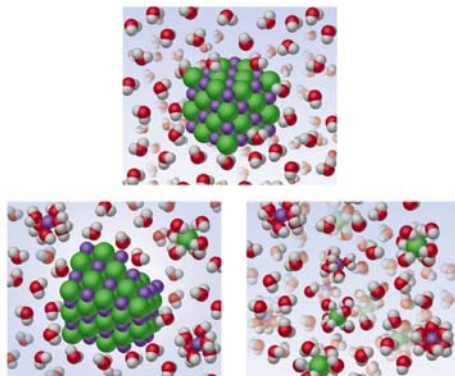
13.1 The Solution Process

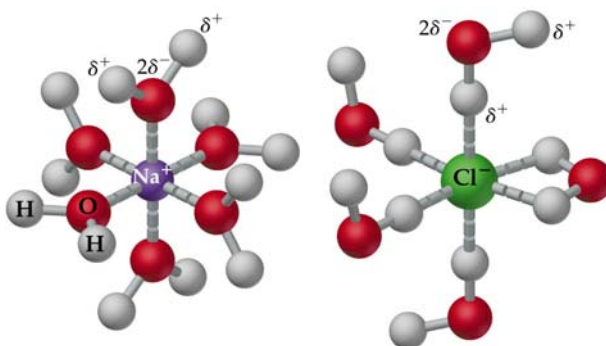
- A solution = **homogeneous mixture** of **solute** and **solvent**.
- may be gases, liquids, or solids
- Each substance present is a component of the solution:
 - solvent = component present in the largest amount
 - solutes = other components
- Intermolecular forces become rearranged in the process of making solutions with condensed phases.

TABLE 13.1 Examples of Solutions

State of Solution	State of Solvent	State of Solute	Example
Gas	Gas	Gas	Air
Liquid	Liquid	Gas	Oxygen in water
Liquid	Liquid	Liquid	Alcohol in water
Liquid	Liquid	Solid	Salt in water
Solid	Solid	Gas	Hydrogen in palladium
Solid	Solid	Liquid	Mercury in silver
Solid	Solid	Solid	Silver in gold

- Consider NaCl (solute) dissolving in water (solvent):
 - Water molecules orient themselves on the NaCl crystals.
→ H-bonds between the water molecules have to be broken.
 - NaCl dissociates into Na^+ and Cl^- .
→ Ion-dipole forces form between the Na^+ and the negative end of the water dipole.
 - Similar ion-dipole interactions form between the Cl^- and the positive end of the water dipole.
 - Such an interaction between solvent and solute is called **solvation**.
If water is the solvent, the interaction is called **hydration**.





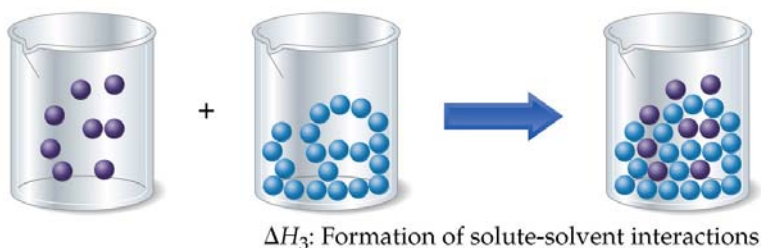
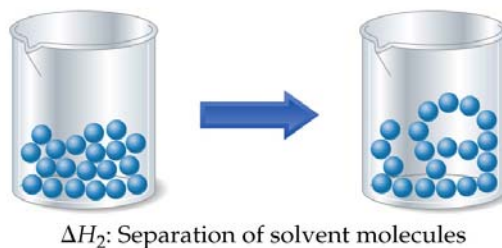
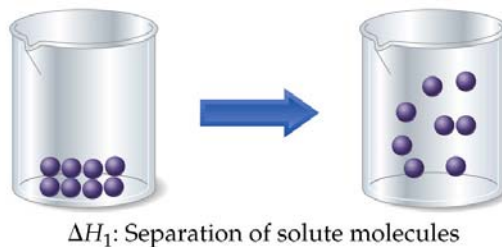
Energy Changes and Solution Formation

- There are three steps involving energy in the formation of a solution:
 - Separation of solute molecules (ΔH_1),
 - Separation of solvent molecules (ΔH_2), and
 - Formation of solute-solvent interactions (ΔH_3).

- We define the enthalpy change in the solution process as:

$$\Delta H_{\text{soln}} = \Delta H_1 + \Delta H_2 + \Delta H_3$$

- ΔH_{soln} can either be positive or negative depending on the intermolecular forces.



Determining whether ΔH_{soln} is positive or negative:

Consider the strengths of all solute-solute, solvent-solvent and solute-solvent interactions:

- Breaking attractive IMFs is always endothermic.
→ ΔH_1 and ΔH_2 are both positive (i.e. separation of solute molecules and separation of solvent molecules)
- Forming attractive IMFs is always exothermic.
→ ΔH_3 is always negative (i.e. formation of solute-solvent interactions)

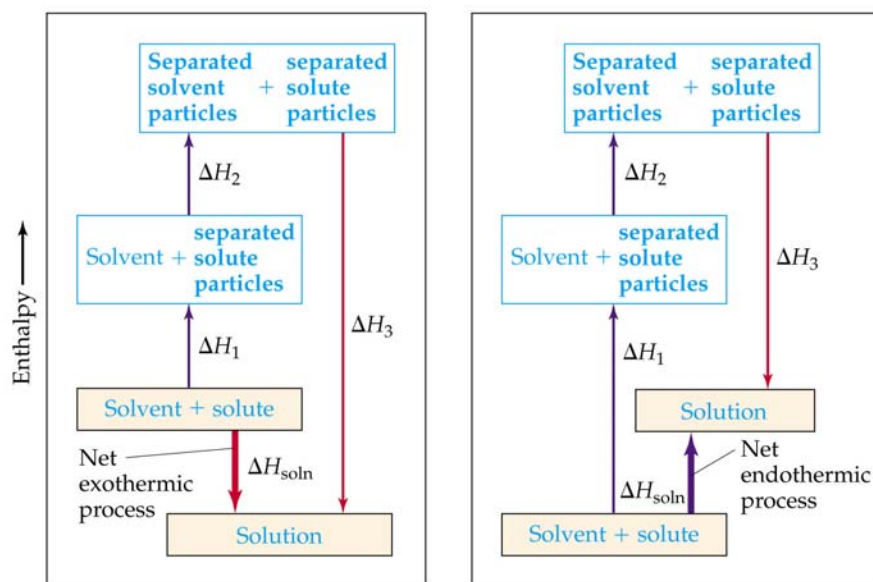
Look for whether $\Delta H_1 + \Delta H_2$ is greater or less than ΔH_3

Examples:

- MgSO_4 added to water has $\Delta H_{\text{soln}} = -91.2 \text{ kJ/mol}$.
- NH_4NO_3 added to water has $\Delta H_{\text{soln}} = +26.4 \text{ kJ/mol}$.

Practical applications:

MgSO_4 is often used in instant heat packs and NH_4NO_3 is often used in instant cold packs.



How can we predict if a solution will form?

- In general, solutions form if the ΔH_{soln} is **negative**.
- If ΔH_{soln} is too endothermic a solution will not form.
- "Rule of thumb": Polar solvents dissolve polar solutes.
Nonpolar solvents dissolve nonpolar solutes.
- E.g. mixing NaCl in gasoline.
 - Only weak interactions are possible because gasoline is nonpolar.
 - These interactions do not compensate for the separation of ions from one another.
→ NaCl doesn't dissolve to any great extent in gasoline.
- Consider the process of mixing water in octane (C_8H_{18}).
 - Water has strong H-bonds.
 - The energy required to break these H-bonds is not compensated for by interactions between water and octane.
→ Water and octane do not mix.

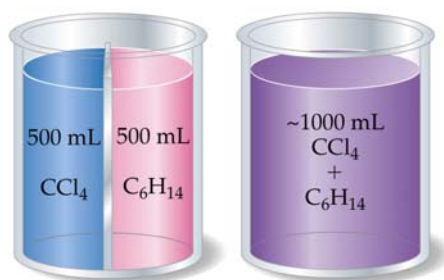
Solution Formation, Spontaneity, and Disorder

- A spontaneous process occurs without outside intervention.

When the E of the system ↓ (e.g., dropping a book and allowing it to fall to a lower PE), the process is spontaneous.

- Spontaneous** processes tend to be **exothermic**, but not all.
- There is another component = **entropy** = the amount of randomness or disorder of the system.
Mixing → ↑ entropy → solution formation

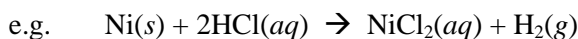
e.g. mixture of CCl_4 and C_6H_{14} is less ordered than the two separate liquids.
→ spontaneously mix even though ΔH_{soln} is very close to zero.



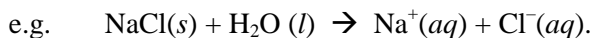
i.e. A solution will form unless the solute-solute or solvent-solvent interactions are too strong relative to solute-solvent interactions.

Solution Formation and Chemical Reactions

- Solutions form by **physical processes** and by **chemical processes**.



- Note that the chemical form of the substance being dissolved has changed during this process ($\text{Ni} \rightarrow \text{NiCl}_2$)
- Remove water → no Ni is found, only $\text{NiCl}_2 \cdot 6\text{H}_2\text{O}$ remains.
→ the dissolution of Ni in HCl is a chemical process.



- Remove water → NaCl is found.
→ NaCl dissolution is a physical process.

13.3 Factors Affecting Solubility

- The tendency of a substance to dissolve in another depends on:
 - The nature of the solute.
 - The nature of the solvent.
 - The temperature.
 - The pressure (for gases).

Solute-Solvent Interactions

- IMFs = important factor in determining solubility of a solute in a solvent.
 - ↑ attraction between solute and solvent molecules → ↑ solubility
 - e.g. **polar liquids tend to dissolve in polar solvents**, because of favorable dipole-dipole interactions (solute-solute, solvent-solvent, and solute-solvent).
- Pairs of liquids that mix in any proportions are said to be **miscible**.
e.g. Ethanol and water are miscible liquids.
- In contrast, **immiscible** liquids do not mix significantly.
e.g. Gasoline and water are immiscible.

Solubility of alcohols in water

- Broken hydrogen bonds in both pure liquids are re-established in the mixture.

Why are not all alcohols are miscible with water?

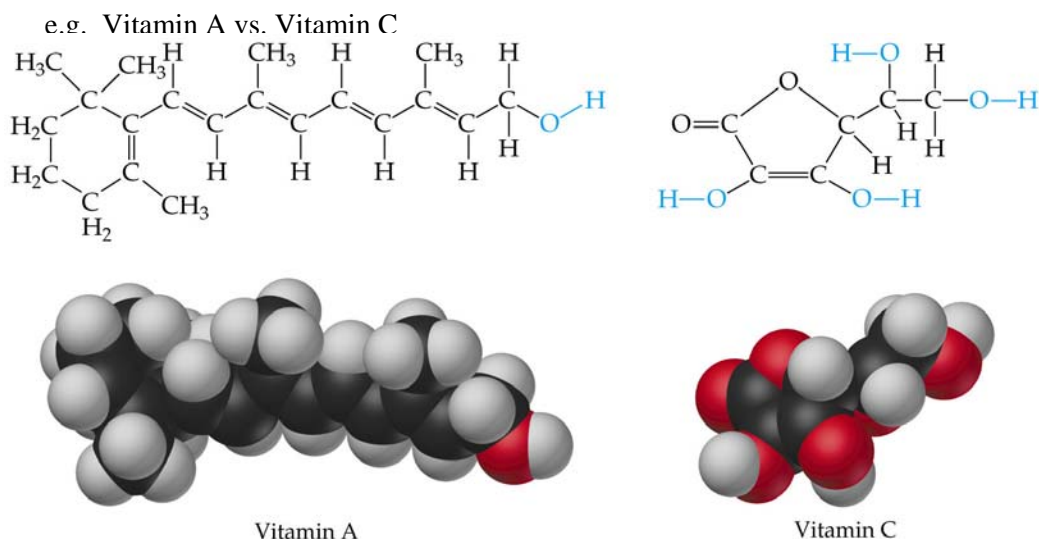
- number of carbon atoms in a chain
 - ↑ carbons in the chain → molecule behaves more like a hydrocarbon.
→ ↑ C atoms in the alcohol → ↓ solubility in water.
- ↑ number of -OH groups within a molecule → ↑ solubility in water.
- ↑ number of -OH groups along the chain → ↑ solute-water H-bonding

TABLE 13.3 Solubilities of Some Alcohols in Water and in Hexane

Alcohol	Solubility in H ₂ O ^a	Solubility in C ₆ H ₁₄
CH ₃ OH (methanol)	∞	0.12
CH ₃ CH ₂ OH (ethanol)	∞	∞
CH ₃ CH ₂ CH ₂ OH (propanol)	∞	∞
CH ₃ CH ₂ CH ₂ CH ₂ OH (butanol)	0.11	∞
CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ OH (pentanol)	0.030	∞
CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ OH (hexanol)	0.0058	∞
CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ OH (heptanol)	0.0008	∞

- **Generalization: “Like dissolves like”.**

- Substances with similar attractive IMFs tend to be soluble in one another.
 - ↑ polar bonds in the molecule → better dissolution in a polar solvent.
 - Less polar → less likely molecule will dissolve in a polar solvent, and more likely it will dissolve in a nonpolar solvent.



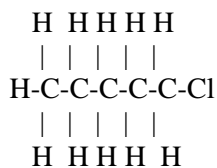
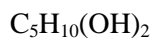
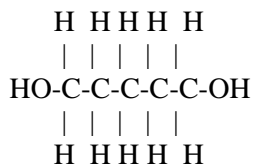
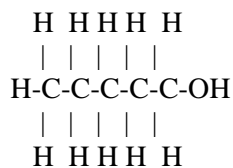
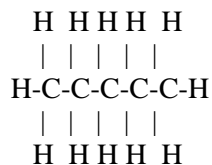
- Network solids do not dissolve because the strong intermolecular forces in the solid are not reestablished in any solution.

Sample Exercise 13.1 (p. 495)

Predict whether each of the following substances is more likely to dissolve in carbon tetrachloride (CCl_4) or in water: C_7H_{16} , Na_2SO_4 , HCl , and I_2 .

Practice Exercise 13.1

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



Pressure Effects



- The solubility of a **gas** in a **liquid** is a function of the **P of the gas over the solution**.
 - Solubilities of solids and liquids are not greatly affected by pressure.
- \uparrow gas P \rightarrow \uparrow molecules of gas close to the surface of the solution \rightarrow \uparrow probability of a gas molecule striking the surface \rightarrow \uparrow entering of gas into the solution \rightarrow \uparrow P \rightarrow \uparrow solubility of gas in liquid.
- \downarrow P \rightarrow \downarrow number molecules of gas close to the surface of the solution resulting in a \downarrow solubility of gas in liquid.

Henry's Law: The solubility of a gas is directly proportional to the partial pressure of the gas above the solution.

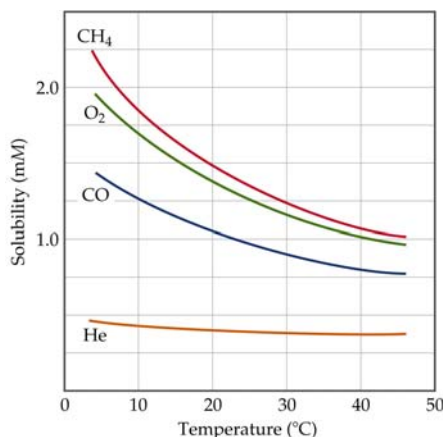
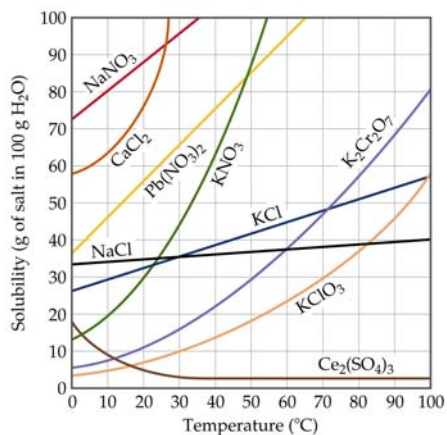
- $S_g = kP_g$ where S_g = solubility of gas, P_g = partial pressure, k = Henry's law constant.
 - Note: The Henry's law constant differs for each solute-solvent pair and differs with T

Application of Henry's law: preparation of carbonated soda.

- Carbonated beverages are bottled under $P_{CO_2} > 1$ atm.
- As the bottle is opened, $P_{CO_2} \downarrow \rightarrow$ solubility of $CO_2 \downarrow \rightarrow$ escape of bubbles of CO_2 from solution.

Temperature Effects

- Sugar dissolves better in warm water than in cold water.
 - $\uparrow T \rightarrow \uparrow$ solubility of solids in liquids
 - Sometimes solubility \downarrow as T \uparrow (e.g., $Ce_2(SO_4)_3$).
- Carbonated beverages go flat as they get warm.
 - \rightarrow Gases: \uparrow T of liquid $\rightarrow \downarrow$ solubility of gas in liquid
- Environmental application:
 - \uparrow T of lakes $\rightarrow \downarrow$ solubility CO_2 and O_2 in water \rightarrow fish suffocate



Sample Exercise 13.2 (p. 496)

Calculate the concentration of CO_2 in a soft drink that is bottled with a partial pressure of CO_2 of 4.0 atm over the liquid at 25°C . The Henry's Law constant for CO_2 in water at this temperature is $3.1 \times 10^{-2} \text{ mol/K-atm}$.

(0.12 M)

Practice Exercise 13.2

Calculate the concentration of CO_2 in a soft drink after the bottle is opened and equilibrates at 25°C under a CO_2 partial pressure of $3.0 \times 10^{-4} \text{ atm}$.

($9.3 \times 10^{-6} \text{ M}$)

13.4 Ways of Expressing Concentration

- All methods involve quantifying the amount of solute per amount of solvent (or solution).
- Concentration may be expressed qualitatively or quantitatively.
 - The terms *dilute* and *concentrated* are qualitative ways to describe concentration.
 - A dilute solution has a relatively small concentration of solute.
 - A concentrated solution has a relatively high concentration of solute.
- Quantitative expressions of concentration require specific information regarding such quantities as masses, moles, or liters of the solute, solvent, or solution.
 - The most commonly used expressions for concentration are:
 - Mass percentage
 - Mole fraction
 - Molarity
 - Molality

Mass percentage, ppm, and ppb

- **Mass percentage** is one of the simplest ways to express concentration.
 - By definition:

$$\text{Mass \% of component} = \frac{\text{mass of component in soln}}{\text{total mass of solution}} \times 100$$

- Similarly, **parts per million (ppm)** can be expressed as the number of mg of solute per kilogram of solution.
 - By definition:

$$\text{Parts per million (ppm) of component} = \frac{\text{mass of component in soln}}{\text{total mass of solution}} \times 10^6$$

- If the density of the solution is 1g/ml, then 1 ppm = 1 mg solute per liter of solution.
- We can extend this again!
 - **Parts per billion (ppb)** can be expressed as the number of μg of solute per kilogram of solution.
 - By definition:

$$\text{Parts per billion (ppb) of component} = \frac{\text{mass of component in soln}}{\text{total mass of solution}} \times 10^9$$

- If the density of the solution is 1g/ml, then 1 ppb = 1 μg solute per liter of solution.

Sample Exercise 13.3 (p. 498)

- a) A solution is made by dissolving 13.5 g glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) in 0.100 kg of water. What is the mass percentage of solute in this solution? (11.9%)
- b) A 2.5-g sample of groundwater was found to contain 5.4 μg of Zn^{2+} . What is the concentration of Zn^{2+} in parts per million? (2.2 ppm)

Practice Exercise 13.3

- a) Calculate the mass percentage of NaCl in a solution containing 1.50 g of NaCl in 50.0 g of water. (2.91%)
- b) A commercial bleaching solution contains 3.62 mass % sodium hypochlorite, NaOCl. What is the mass of NaOCl in a bottle containing 2500 g of bleaching solution? (90.5 g NaOCl)

Mole Fraction, Molarity, and Molality

- Common expressions of concentration are based on the number of moles of one or more components.
- Recall that mass can be converted to moles using the molar mass.
- Recall:

$$\text{Mole fraction of component, } X = \frac{\text{moles of component}}{\text{total moles of all components}}$$

- Note: Mole fraction has no units.
- Note: Mole fractions range from 0 to 1.

- Recall:

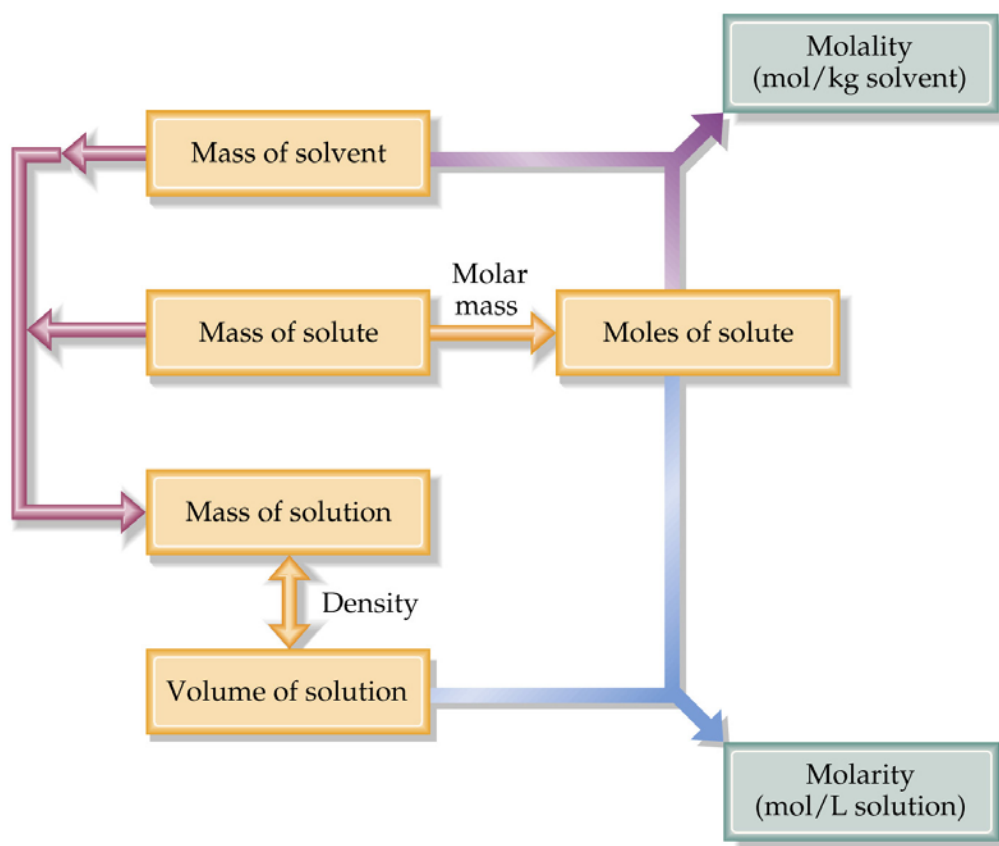
$$\text{Molarity, } M = \frac{\text{moles of solute}}{\text{liters of solution}}$$

- Note: Molarity will change with a change in temperature (as the solution volume increases or decreases).

- We can define **molality** (m), yet another concentration unit:

$$\text{Molality, } m = \frac{\text{moles of solute}}{\text{kilograms of solvent}}$$

- Molality does not vary with temperature.
 - Note: converting between molarity (M) and molality (m) requires density.



Sample Exercise 13.4 (p. 500)

A solution is made by dissolving 4.35 g glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) in 25.0 mL of water. Calculate the molality of glucose in the solution. (0.964 *m*)

Practice Exercise 13.4

What is the molality of a solution made by dissolving 36.5 g of naphthalene (C_{10}H_8) in 425 g of toluene (C_7H_8)? (0.670 *m*)

Sample Exercise 13.5 (p. 500)

A solution of hydrochloric acid contains 36% HCl by mass.

a) Calculate the mole fraction of HCl in the solution. (0.22)

b) Calculate the molality of HCl in the solution. (15 *m*)

Practice Exercise 13.5

A commercial bleach solution contains 3.62 mass % NaOCl in water.

Calculate

a) the molality and (0.505 *m*)

b) the mole fraction of NaOCl in the solution. (9.00×10^{-3})

Sample Exercise 13.6 (p. 501)

A solution contains 5.0 g of toluene (C_7H_8) and 225 g of benzene and has a density of 0.876 g/mL. Calculate the molarity of the solution. (0.21 M)

Practice Exercise 13.6

A solution containing equal masses of glycerol ($\text{C}_3\text{H}_8\text{O}_3$) and water has a density of 1.10 g/mL. Calculate

a) the molality of glycerol; (10.9 *m*)

b) the mole fraction of glycerol; (0.163)

c) the molarity of glycerol in the solution. (5.97 M)

13.5 Colligative Properties

- Colligative properties depend on **number of solute particles**.

Colligative properties are **physical** properties.

- There are four colligative properties to consider:
 - Vapor pressure lowering (Raoult's Law).
 - Boiling point elevation.
 - Freezing point depression.
 - Osmotic pressure.

1. Lowering the Vapor Pressure

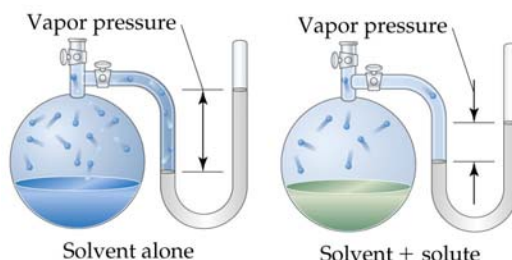
↑ # particles → ↓ vapor pressure

e.g. pure H₂O vs. H₂O + NaCl (salt water)

higher V.P. lower V.P.

Why?

- 1) Na⁺ + Cl⁻ particles (nonvolatile solutes) get surrounded by H₂O → fewer “free” H₂O particles left to vaporize at surface
- 2) ↑ IMFs → harder to get out



- **Raoult's law** quantifies the extent to which a nonvolatile solute lowers the vapor pressure of the solvent.
 - If P_A is the vapor pressure with solute, P_A° is the vapor pressure of the pure solvent, and X_A is the mole fraction of A, then

$$P_A = X_A P_A^\circ$$

- **Ideal solution:** one that obeys Raoult's law.
 - Real solutions show approximately ideal behavior when:
 - The solute concentration is low.
 - The solute and solvent have similarly sized molecules.
 - The solute and solvent have similar types of intermolecular attractions.
 - Raoult's law breaks down when the solvent-solvent and solute-solute intermolecular forces are >> or << solute-solvent IMFs.

Sample Exercise 13.7 (p. 504)

Glycerin (C₃H₈O₃) is a nonvolatile nonelectrolyte with a density of 1.25 g/mL at 25°C. Calculate the vapor pressure at 25°C of a solution made by adding 50.0 mL of glycerin to 500.0 mL of water. The vapor pressure of pure water at 25°C is 23.8 torr.

(23.2 torr)

Practice Exercise 13.7

The vapor pressure of pure water at 110°C is 1070 torr. A solution of ethylene glycol and water has a vapor pressure of 1.00 atm at 110°C. Assuming that Raoult's law is obeyed, what is the mole fraction of ethylene glycol in the solution?

(0.290)

2. Boiling Point Elevation

↑ # particles of nonvolatile solute → ↓ VP → ↑ boiling point

Since boiling point is when VP = atmospheric pressure, if the V.P. is lower, then **more energy** must be used to overcome atmospheric pressure = **higher T** at boiling

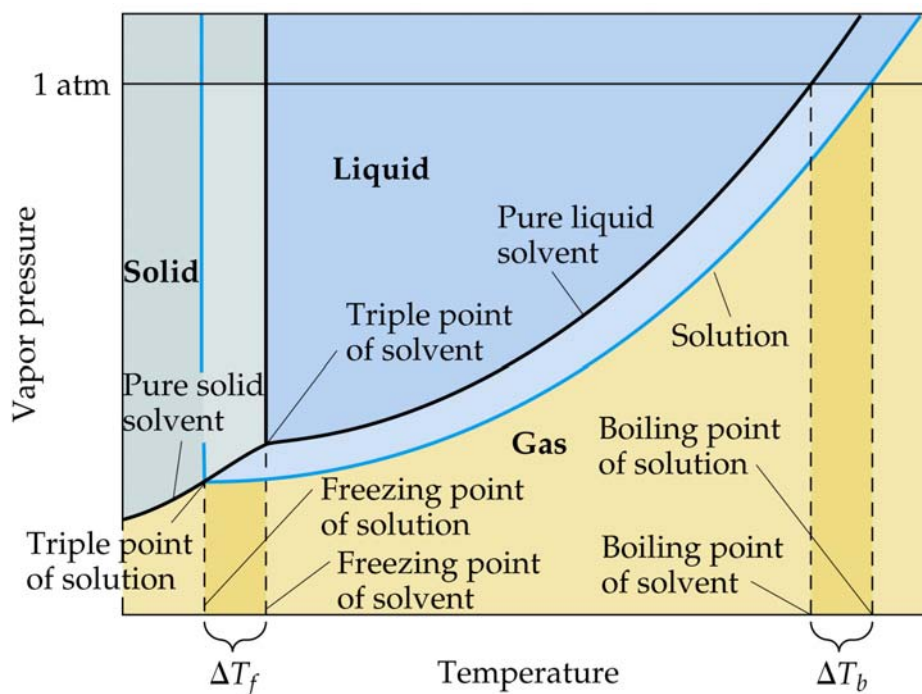
It takes **more energy to break apart solvent/solute bonds** so solvent can vaporize.

$$\text{H}_2\text{O}: \frac{1 \text{ mole particles}}{1 \text{ kg H}_2\text{O}} = + 0.512^\circ\text{C}$$

- molal boiling-point-elevation constant, K_b** , expresses how much ΔT_b changes with molality, m :

$$\Delta T_b = K_b m$$

- The boiling-point elevation is proportional to [solute particles].
- A 1 m solution of NaCl is 2 m in total solute particles.



3. Freezing Point Depression

↑ # particles → ↓ freezing point

- When a solution freezes, crystals of almost pure solvent are formed first.
- Solute molecules are usually not soluble in the solid phase of the solvent.
- the triple point occurs at a lower T because of the lower VP for the solution.
- The melting-point (freezing-point) curve is a vertical line from the triple point.
- the solution freezes at a lower temperature (ΔT_f) than the pure solvent.
- The ↓ in freezing point (ΔT_f) is directly proportional to **molality**, i.e. [solute particles].

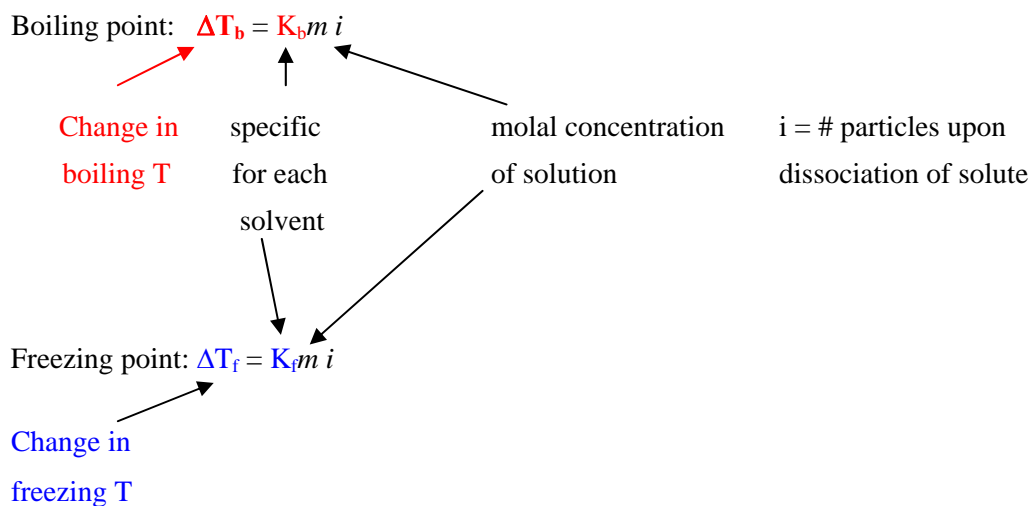
- K_f = molal freezing-point-depression constant.**

$$\Delta T_f = K_f m$$

TABLE 13.4 Molal Boiling-Point-Elevation and Freezing-Point-Depression Constants

Solvent	Normal Boiling Point (°C)	K_b (°C/m)	Normal Freezing Point (°C)	K_f (°C/m)
Water, H ₂ O	100.0	0.51	0.0	1.86
Benzene, C ₆ H ₆	80.1	2.53	5.5	5.12
Ethanol, C ₂ H ₅ OH	78.4	1.22	-114.6	1.99
Carbon tetrachloride, CCl ₄	76.8	5.02	-22.3	29.8
Chloroform, CHCl ₃	61.2	3.63	-63.5	4.68

Boiling Point/Freezing Point Calculations



For H₂O: $K_b = 0.512\text{ }^\circ\text{C/m}$
 $K_f = 1.86\text{ }^\circ\text{C/m}$

e.g. What is the boiling point of a solution that contains 1.25 mol CaCl₂ in 1400. g of water?

$$\frac{\text{mol CaCl}_2}{1000. \text{ g H}_2\text{O}} = \frac{1.25 \text{ mol CaCl}_2}{1400. \text{ g H}_2\text{O}} \times \frac{1000. \text{ g}}{1 \text{ kg}} = 0.8929 \text{ m} \times 3 = 2.679 \text{ m}$$

$i = \#$ particles when CaCl₂ dissociates in H₂O

$$\Delta T_b = \frac{0.512^\circ\text{C}}{\text{m}} \times 2.679\text{m} = 1.37^\circ\text{C} + 100.^\circ\text{C} = \boxed{101.37^\circ\text{C}}$$

Sample Exercise 13.8 (p. 506)

Automotive antifreeze consists of ethylene glycol ($\text{C}_2\text{H}_6\text{O}_2$), a nonvolatile nonelectrolyte. Calculate the boiling point and freezing point of a 25.0 mass % solution of ethylene glycol in water.

(102.7°C; -10.0°C)

Practice Exercise 13.8

Calculate the freezing point of a solution containing 0.600 kg of chloroform (CHCl_3) and 42.0 g of eucalyptol ($\text{C}_{10}\text{H}_{18}\text{O}$), a fragrant substance found in the leaves of eucalyptus trees. (See Table 13.4)

(-65.6°C)

Sample Exercise 13.9 (p. 507)

List the following aqueous solutions in order of their expected freezing points: 0.050 *m* CaCl_2 ; 0.15 *m* HCl ; 0.050 *m* $\text{HC}_2\text{H}_3\text{O}_2$; 0.10 *m* $\text{C}_{12}\text{H}_{22}\text{O}_{11}$.

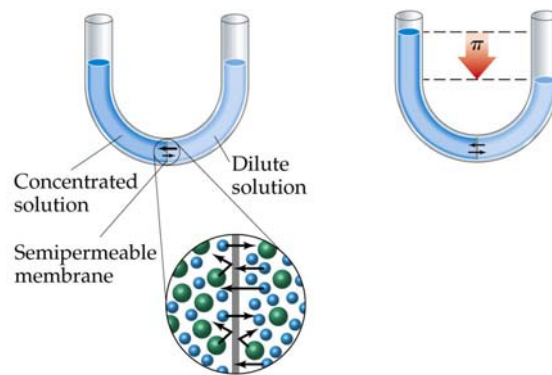
Practice Exercise 13.9

Which of the following solutes will produce the largest increase in boiling point upon addition of 1 kg of water: 1 mol $\text{Co}(\text{NO}_3)_2$, 2 mol of KCl , 3 mol of ethylene glycol ($\text{C}_2\text{H}_6\text{O}_2$)?

(2 mol KCl)

Osmosis

- **Osmosis** = the net movement of a solvent from an area of low [solute] to an area of high [solute], e.g. cell membranes and cellophane.
- Consider a U-shaped tube with a two liquids separated by a semipermeable membrane.
 - One arm of the tube contains pure solvent.
 - The other arm contains a solution.
 - Solvent moves in both directions across a semipermeable membrane.
 - The rate of movement of solvent from the pure solvent to the solution is faster than the rate of movement in the opposite direction.
 - As solvent moves across the membrane, the fluid levels in the arms become uneven.
 - The vapor pressure of solvent is higher in the arm with pure solvent.
 - Eventually the pressure difference due to the difference in height of liquid in the arms stops osmosis.



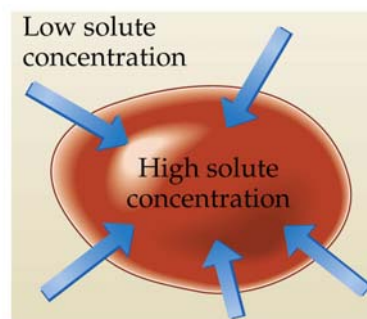
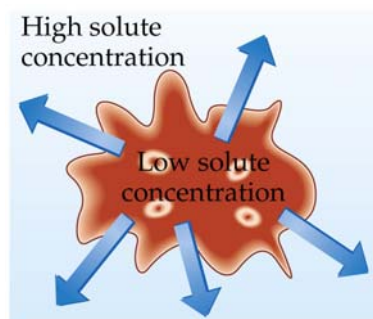
- **Osmotic pressure**, π , is the pressure required to prevent osmosis.
 - Osmotic pressure obeys a law similar in form to the ideal-gas law.
 - For n moles, V = volume, M = molarity, R = the ideal gas constant, and an absolute temperature, T , the osmotic pressure is:

$$\pi V = nRT$$

$$\pi = \left(\frac{n}{V} \right) RT = MRT$$

- Two solutions are said to be *isotonic* if they have the same osmotic pressure.
 - *Hypotonic* solutions have a lower π , relative to a more concentrated solution.
 - *Hypertonic* solutions have a higher π , relative to a more dilute solution.

Red Blood Cells



- If red blood cells are placed in a **hypertonic** solution (relative to intracellular solution), there is a lower [solute] in the cell than the surrounding tissue.
→ Osmosis occurs and water passes through the membrane out of the cell.
→ The cell shrivels up (*crenation*).
- If red blood cells are placed in a **hypotonic** solution, there is a higher [solute] in the cell than outside the cell.
→ Osmosis occurs and water moves into the cell.
→ The cell bursts (*hemolysis*).
- To prevent crenation or hemolysis, IV (intravenous) solutions must be **isotonic** relative to the intracellular fluids of cells.

Everyday examples of osmosis:

- If a cucumber is placed in NaCl solution, it will lose water to shrivel up and become a pickle.
- A limp carrot placed in water becomes firm because water enters via osmosis.
- Eating large quantities of salty food causes retention of water and swelling of tissues (*edema*).
- Water moves into plants, to a great extent, through osmosis.
- Salt may be added to meat (or sugar added to fruit) as a preservative.
 - Salt prevents bacterial infection: A bacterium placed on the salt will lose water through osmosis and die.
- *Active transport* is the movement of nutrients and waste material through a biological membrane against a concentration gradient.
 - Movement is from an area of low concentration to an area of high concentration.
 - Active transport is not spontaneous.
 - Energy must be expended by the cell to accomplish this.

Sample Exercise 13.10 (p. 509)

The average osmotic pressure of blood is 7.7 atm at 25°C. What concentration of glucose ($C_6H_{12}O_6$) will be isotonic with blood? (0.31 M)

Practice Exercise 13.10

What is the osmotic pressure at 20°C of a 0.0020 M sucrose ($C_{12}H_{22}O_{11}$) solution? (0.048 atm or 37 torr)

Determination of Molar Mass

- Any of the four colligative properties may be used to determine molar mass.

Sample Exercise 13.11 (p. 509)

A solution of an unknown nonvolatile nonelectrolyte was prepared by dissolving 0.250 g of the substance in 40.0 g of CCl_4 . The boiling point of the resultant solution was 0.357°C higher than that of the pure solvent. Calculate the molar mass of the solute. (88.0 g/mol)

Practice Exercise 13.11

Camphor ($\text{C}_{10}\text{H}_{16}\text{O}$) melts at 179.8°C , and it has a particularly large freezing-point-depression constant, $K_f = 40.0^\circ\text{C}/m$. When 0.186 g of an organic substance of unknown molar mass is dissolved in 22.01 g of liquid camphor, the freezing point of the mixture is found to be 176.7°C . What is the molar mass of the solute? (110 g/mol)

Sample Exercise 13.12 (p. 510)

The osmotic pressure of an aqueous solution of a certain protein was measured in order to determine its molar mass. The solution contained 3.50 mg of protein dissolved in sufficient water to form 5.00 mL of solution. The osmotic pressure of the solution at 25°C was found to be 1.54 torr. Calculate the molar mass of the protein. (8.45×10^3 g/mol)

Practice Exercise 13.12

A sample of 2.05 g of the plastic polystyrene was dissolved in enough toluene to form 0.100 L of solution. The osmotic pressure of this solution was found to be 1.21 kPa at 25°C . Calculate the molar mass of the polystyrene. (4.20×10^4 g/mol)