

Chapter 13. Properties of Solutions

Common Student Misconceptions

- Students often confuse dilute and concentrated; weak and strong are often confused.
- Students often do not appreciate the *driving forces* behind the formation of a solution.
- Students often confuse dissolution with melting.
- Students do not realize that crystallization is the reverse of dissolution.
- Many student think that solutions can only be made either by mixing two liquids or dissolving a solid in a liquid.
- Students often confuse solution with solvation.
- Students often think that every mixture is a solution.
- Students often do not appreciate how *unusual* water is.
- Students often forget that calculations of molality require the mass of *solvent*, not *solution*.
- Students often do not realize that colloids, like solutions, can occur in all three states of matter.

Teaching Tips

- Remind students that even the so-called *insoluble* compounds dissolve to *some* extent in water.
- The differences between the definitions of molarity (M) and molality (m) and between their respective notations and pronunciations must be emphasized to avoid confusion.
- Remind students that, for more concentrated solutions, the density of the solution is needed to relate molarity and molality.
- The assumption that the density of a dilute aqueous solution is identical to that of pure water (1 g/mL) is a valid at normal temperatures (to two significant figures).
- It may be helpful to note the similarity between Raoult's law ($P_A = X_A P_A^\circ$) and the expression for partial pressures in a mixture of gases derived from Dalton's law ($P_A = X_A P_{\text{tot}}$).

Lecture Outline

13.1 The Solution Process^{1,2,3,4}

- A *solution* is a homogeneous mixture of solute and solvent.
- Solutions may be gases, liquids, or solids,
- Each substance present is a *component* of the solution.
 - The *solvent* is the component present in the largest amount.
 - The other components are the *solutes*.
 - We will be particularly interested in *aqueous solutions* which contain water as the solvent.

The Natural Tendency Toward Mixing

- Consider the formation of a gaseous solution of $\text{O}_2(\text{g})$ and $\text{Ar}(\text{g})$.
 - Initially they are separated by a barrier.
 - When the barrier is removed, the gases mix to form a homogeneous mixture, or solution.
 - The mixing of gases is a *spontaneous* process.
 - It occurs without input of energy from the surroundings.

¹ "Dissolution of NaCl in Water" Animation from Instructor's Resource CD/DVD

² Figure 13.3 from Transparency Pack

³ "Sodium Chloride (1 × 1 Unit Cell)" 3-D Model from Instructor's Resource CD/DVD

⁴ "Dissolution of KMnO_4 in Water" Movie from Instructor's Resource CD/DVD

- *Entropy* is the thermodynamic quantity that measures the extent of the spreading of the molecules and their associated kinetic energies.
 - The mixing that occurs as the solution is formed represents an increase in entropy.
 - Formation of a solution is favored by the increase in entropy that accompanies mixing.

The Effect of Intermolecular Forces on Solution Formation^{5,6,7,8,9}

- Intermolecular forces become rearranged in the process of making solutions with condensed phases.
- Intermolecular forces operate between solute and solvent particles in a solution.
 - Three kinds of intermolecular interactions are involved in solution formation:
 - *Solute-solute* interactions between solute particles.
 - These must be overcome in order to disperse the particles through the solvent.
 - *Solvent-solvent* interactions between solvent particles.
 - These must be overcome to make room for the solute particles in the solvent.
 - *Solvent-solute* interactions between solvent and solute particles.
 - These occur as the particles mix.
 - 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**.

Energetics of Solution Formation¹⁰

- There are three steps involving energy in the formation of a solution:
 - Separation of solute molecules (ΔH_{solute}),
 - Separation of solvent molecules ($\Delta H_{\text{solvent}}$), and
 - Formation of solute-solvent interactions (ΔH_{mix}).
- We define the enthalpy change in the solution process as:

$$\Delta H_{\text{soln}} = \Delta H_{\text{solute}} + \Delta H_{\text{solvent}} + \Delta H_{\text{mix}}$$
- ΔH_{soln} can either be positive or negative, depending on the intermolecular forces.
 - To determine whether ΔH_{soln} is positive or negative, we consider the strengths of all solute-solute, solvent-solvent, and solute-solvent interactions:
 - Breaking attractive intermolecular forces is always endothermic.
 - ΔH_{solute} and $\Delta H_{\text{solvent}}$ are both positive.
 - Forming attractive intermolecular forces is always exothermic.
 - ΔH_{mix} is always negative.
 - It is possible to have either $\Delta H_{\text{mix}} > (\Delta H_{\text{solute}} + \Delta H_{\text{solvent}})$ or $\Delta H_{\text{mix}} < (\Delta H_{\text{solute}} + \Delta H_{\text{solvent}})$.
 - Examples:
 - MgSO_4 added to water has $\Delta H_{\text{soln}} = -91.2 \text{ kJ/mol}$.
 - NH_4NO_3 added to water has $\Delta H_{\text{soln}} = +26.4 \text{ kJ/mol}$.

⁵ “Water” 3-D Model from Instructor’s Resource CD/DVD

⁶ “Pentane” 3-D Model from Instructor’s Resource CD/DVD

⁷ “Acetone” 3-D Model from Instructor’s Resource CD/DVD

⁸ “Chloroform” 3-D Model from Instructor’s Resource CD/DVD

⁹ “Ethanol” 3-D Model from Instructor’s Resource CD/DVD

¹⁰ Figure 13.4 from Transparency Pack

- 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.
 - Consider the process of 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.
 - Result: NaCl doesn't dissolve to any great extent in gasoline.
 - Consider the process of mixing a polar liquid solute (water) with a nonpolar liquid solvent (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.
 - Result: water and octane do not mix.

Solution Formation and Chemical Reactions^{11,12,13,14}

- Some solutions form by physical processes and some by chemical processes.
 - Consider:

$$\text{Ni}(s) + 2\text{HCl}(aq) \rightarrow \text{NiCl}_2(aq) + \text{H}_2(g)$$
 - Note that the chemical form of the substance being dissolved has changed during this process ($\text{Ni} \rightarrow \text{NiCl}_2$)
 - When all the water is removed from the solution, no Ni is found, only $\text{NiCl}_2 \cdot 6\text{H}_2\text{O}$ remains.
 - Therefore, the dissolution of Ni in HCl is a chemical process.
 - By contrast:

$$\text{NaCl}(s) + \text{H}_2\text{O}(l) \rightarrow \text{Na}^+(aq) + \text{Cl}^-(aq)$$
 - When the water is removed from the solution, NaCl is found.
 - Therefore, NaCl dissolution is a physical process.

FORWARDS REFERENCES

- Hydrolysis of metal ions will be brought up again in Chapter 16 (section 16.11).
- Thermodynamics of processes will be further discussed throughout Chapter 19.
- Rust is a hydrate (Chapter 20, section 20.8).

13.2 Saturated Solutions and Solubility^{15,16,17}

- As a solid dissolves, a solution forms:
 - Solute + solvent \rightarrow solution
- The opposite process is **crystallization**.
 - Solution \rightarrow solute + solvent
- If crystallization and dissolution are in equilibrium with undissolved solute, the solution is **saturated**.
 - There will be no further increase in the amount of dissolved solute.

¹¹ “The Use of Dots in Chemical Formulas” from Further Readings

¹² “Hydrated Magnesium Cation” 3-D Model from Instructor’s Resource CD/DVD

¹³ “Deprotonated Hydrated Aluminum Cation” 3-D Model from Instructor’s Resource CD/DVD

¹⁴ “Hydrated Aluminum Cation” 3-D Model from Instructor’s Resource CD/DVD

¹⁵ “Crystallization from a Supersaturated Solution of Sodium Acetate” from Further Readings

¹⁶ “Supersaturation” from Live Demonstrations

¹⁷ “Crystallization from Supersaturated Solutions of Sodium Acetate” from Live Demonstrations

- **Solubility** is the amount of solute required to form a saturated solution.
 - A solution with a concentration of dissolved solute that is less than the solubility is said to be **unsaturated**.
 - A solution is said to be **supersaturated** if more solute is dissolved than in a saturated solution.

FORWARDS REFERENCES

- Solubility equilibria will be further discussed in Chapter 17 (section 17.4).
- Various crystalline substances (anhydrous and hydrated) for transition metal compounds will be described throughout Chapter 23.

13.3 Factors Affecting Solubility^{18,19,20,21}

- 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^{22,23,24,25,26,27,28,29,30,31}

- Intermolecular forces are an important factor in determining solubility of a solute in a solvent.
 - The stronger the attraction between solute and solvent molecules, the greater the solubility.
 - For example, polar liquids tend to dissolve in polar solvents.
 - Favorable dipole-dipole interactions exist (solute-solute, solvent-solvent, and solute-solvent).
- Pairs of liquids that mix in any proportions are said to be **miscible**.
 - Example: Ethanol and water are miscible liquids.
- In contrast, **immiscible** liquids do not mix significantly.
 - Example: Gasoline and water are immiscible.
- Consider the solubility of alcohols in water.
 - Water and ethanol are miscible because the broken hydrogen bonds in both pure liquids are re-established in the mixture.
- However, not all alcohols are miscible with water.
 - Why?
 - The number of carbon atoms in a chain affects solubility.
 - The greater the number of carbons in the chain, the more the molecule behaves like a hydrocarbon.
 - Thus, the more C atoms in the alcohol, the lower its solubility in water.

¹⁸ “Vitamin C (ascorbic acid)” 3-D Model from Instructor’s Resource CD/DVD

¹⁹ “Vitamin A” 3-D Model from Instructor’s Resource CD/DVD

²⁰ “Alanine” 3-D Model from Instructor’s Resource CD/DVD

²¹ “Ibuprofen” 3-D Model from Instructor’s Resource CD/DVD

²² “Nonadditivity of Volumes” from Live Demonstrations

²³ Figure 13.11 from Transparency Pack

²⁴ “Polarity, Miscibility, and Surface Tension” from Further Readings

²⁵ “An Analogy to Illustrate Miscibility of Liquids” from Further Readings

²⁶ “Solubility of Alcohols” from Live Demonstrations

²⁷ “Using Computer-Based Visualization Strategies to Improve Students’ Understanding of Molecular Polarity and Miscibility” from Further Readings

²⁸ “Applications of Solubility Data” from Further Readings

²⁹ “Why Don’t Water and Oil Mix?” from Live Demonstrations

³⁰ “Cyclohexane” 3-D Model from Instructor’s Resource CD/DVD

³¹ “Glucose” 3-D Model from Instructor’s Resource CD/DVD

- Increasing the number of –OH groups within a molecule increases its solubility in water.
 - The greater the number of –OH groups along the chain, the more solute-water H-bonding is possible.
- Generalization: “like dissolves like”.
 - Substances with similar intermolecular attractive forces tend to be soluble in one another.
 - The more polar bonds in the molecule, the better it dissolves in a polar solvent.
 - The less polar the molecule the less likely it is to dissolve in a polar solvent and the more likely it is to dissolve in a nonpolar solvent.
- Network solids do not dissolve because the strong intermolecular forces in the solid are not reestablished in any solution.

Pressure Effects^{32,33,34,35}

- The solubility of a gas in a liquid is a function of the partial pressure of the gas over the solution.
 - Solubilities of solids and liquids are not greatly affected by pressure.
- With higher gas pressure, more molecules of gas are close to the surface of the solution and the probability of a gas molecule striking the surface and entering the solution is increased.
 - Therefore, the higher the pressure, the greater the solubility.
- The lower the pressure, the smaller the number of molecules of gas close to the surface of the solution resulting in a lower solubility.
 - The solubility of a gas in a liquid solvent is directly proportional to the partial pressure of the gas above the solution.
 - This statement is called **Henry's law**.
 - Henry's law may be expressed mathematically as:

$$S_g = kP_g$$
 - Where S_g is the solubility of gas, P_g the partial pressure, k = Henry's law constant.
 - Note that the Henry's law constant differs for each solute-solvent pair and differs with temperature.
- An application of Henry's law is the preparation of carbonated soda.
 - Carbonated beverages are bottled under $P_{\text{CO}_2} > 1$ atm.
 - As the bottle is opened, P_{CO_2} decreases and the solubility of CO_2 decreases.
 - Therefore, bubbles of CO_2 escape from solution.

Temperature Effects^{36,37,38}

- Experience tells us that sugar dissolves better in warm water than in cold water.
 - The solubility of most solid solutes in water increases as the solution temperature increases.
 - Sometimes solubility decreases as temperature increases (e.g., $\text{Ce}_2(\text{SO}_4)_3$).
- Experience tells us that carbonated beverages go flat as they get warm.
 - The solubility of gas in water decreases with increasing temperature.
- An environmental application of this is thermal pollution.
 - *Thermal pollution*: if lakes get too warm, CO_2 and O_2 become less soluble and are not available for plants or animals.
 - Fish suffocate.

³² “Effect of Temperature and Pressure on the Solubility of Gases in Liquids” from Live Demonstrations

³³ “Henry's Law and Noisy Knuckles” from Further Readings

³⁴ “Henry's Law: A Retrospective” from Further Readings

³⁵ “Henry's Law” Animation from Instructor's Resource CD/DVD

³⁶ Figure 13.18 from Transparency Pack

³⁷ Figure 13.19 from Transparency Pack

³⁸ “Soft Drink Bubbles” from Further Readings

FORWARDS REFERENCES

- The dynamic equilibrium between a solid solute and its solution will be mentioned in Chapter 14 (section 14.7).
- Factors affecting solubility will be discussed in detail in Chapter 17 (section 17.5).
- Temperature effects on solubility of NaCl will be mentioned in Chapter 19 (section 19.7).
- Reactions involving CO₂, HCO₃⁻ and H₂CO₃ will be discussed in Chapter 22 (section 22.9).
- Solubility of organic substances in polar solvents will be mentioned in Chapter 24 (section 24.1).

13.4 Expressing Solution Concentration³⁹

- 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 ppb⁴⁰

- **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:

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

- The density of a very dilute aqueous solution is similar to that of pure water (1g/mL)
 - 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.

³⁹ “Caffeine” 3-D Model from Instructor’s Resource CD/DVD

⁴⁰ “Candy Sprinkles to Illustrate One Part Per Million” from Further Readings

Mole Fraction, Molarity, and Molality^{41,42,43}

- 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 that mole fraction has no units.
- Note that mole fractions range from 0 to 1.
- Recall:

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

- Note that 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 that converting between molarity (M) and molality (m) requires density.
- The molarity and molality of dilute solutions are often very similar.

FORWARD REFERENCES

- Molar concentrations will be used in rate law expressions in Chapter 14.
- Molar concentrations will be used in equilibrium constant and reaction quotient expressions in Chapters 15, 16, 17, 19, and 20.
- Concentrations (ppm) of trace constituents in mixtures will be re-introduced in Chapter 18 (section 18.1) and used throughout Chapter 18.

13.5 Colligative Properties

- Colligative properties depend on number of solute particles.
- There are four colligative properties to consider:
 - vapor pressure lowering (Raoult's law).
 - boiling point elevation.
 - freezing point depression.
 - osmotic pressure.

Vapor-Pressure Lowering^{44,45}

- Consider a *volatile* liquid in a closed container.
 - After a period of time, an equilibrium will be established between the liquid and its vapor.
 - The partial pressure exerted by the vapor is the *vapor pressure*.
- *Nonvolatile* solutes (with no measurable vapor pressure) reduce the ability of the surface solvent molecules to escape the liquid.
 - Therefore, vapor pressure is lowered.

⁴¹ “An Alternative Introduction to the Mole Fraction” from Further Readings

⁴² “Mole Fraction Analogies” from Further Readings

⁴³ Figure 13.20 from Transparency Pack

⁴⁴ “Raoult’s Law” Animation from Instructor’s Resource CD/DVD

⁴⁵ “One Cool Chemist” from Further Readings

- The amount of vapor pressure lowering depends on the amount of solute.
- **Raoult's law** quantifies the extent to which a nonvolatile solute lowers the vapor pressure of the solvent.
 - If P_{solution} is the vapor pressure of the solution, $P_{\text{solvent}}^{\circ}$ is the vapor pressure of the pure solvent, and X_{solute} is the mole fraction of solute, then

$$P_{\text{solution}} = X_{\text{solvent}} P_{\text{solvent}}^{\circ}$$
 - The vapor-pressure lowering, ΔP , is directly proportional to the mole fraction of the solute, X_{solute} .

$$\Delta P = X_{\text{solute}} P_{\text{solvent}}^{\circ}$$
- An **ideal solution** is 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 much greater or weaker than solute-solvent intermolecular forces.

Boiling-Point Elevation^{46,47,48}

- A nonvolatile solute lowers the vapor pressure of a solution.
- At the normal boiling point of the pure liquid, the solution has a vapor pressure less than 1 atm.
 - Therefore, a higher temperature is required to reach a vapor pressure of 1 atm for the solution (ΔT_b).
- The **molal boiling-point-elevation constant**, K_b , expresses how much ΔT_b changes with molality, m :

$$\Delta T_b = K_b m$$
- The nature of the solute (electrolyte vs. nonelectrolyte) will impact the colligative molality of the solute.

Freezing-Point Depression^{49,50,51}

- 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.
 - Therefore, the triple point occurs at a lower temperature because of the lower vapor pressure for the solution.
- The melting-point (freezing-point) curve is a vertical line from the triple point.
 - Therefore, the solution freezes at a lower temperature (ΔT_f) than the pure solvent.
 - The decrease in freezing point (ΔT_f) is directly proportional to molality.
- K_f is the **molal freezing-point-depression constant**.

$$\Delta T_f = K_f m$$
 - Values of K_f and K_b for most common solvents can be found in Table 13.4.

⁴⁶ Figure 13.23 from Transparency Pack

⁴⁷ "Boiling-Point Elevation and Freezing-Point Depression" Activity from Instructor's Resource CD/DVD

⁴⁸ "Adrenaline" 3-D Model from Instructor's Resource CD/DVD

⁴⁹ "Antifreeze Solutions: The Colligative Properties of Antifreeze" from Further Readings

⁵⁰ "Freeze-Proof Bugs" from Further Readings

⁵¹ Figure 13.24 from Transparency Pack

Osmosis^{52,53,54,55,56,57,58}

- *Semipermeable* membranes permit passage of some components of a solution.
 - Often they permit passage of water but not larger molecules or ions.
 - Examples of semipermeable membranes are cell membranes and cellophane.
- **Osmosis** is the net movement of a solvent from an area of low solute concentration to an area of high solute concentration.
- 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.
 - There is movement of solvent in both directions across a semipermeable membrane.
 - 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:

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

- 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.
- We can illustrate this with a biological system: red blood cells.
 - Red blood cells are surrounded by semipermeable membranes.
 - If red blood cells are placed in a hypertonic solution (relative to intracellular solution), there is a lower solute concentration in the cell than the surrounding tissue.
 - Osmosis occurs and water passes through the membrane out of the cell.
 - The cell shrivels up.
 - This process is called *crenation*.
 - If red blood cells are placed in a hypotonic solution, there is a higher solute concentration 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 include:
 - 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.

⁵² “A Simple Demonstration Model of Osmosis” from Live Demonstrations

⁵³ “Osmotic Pressure of a Sugar Solution” from Live Demonstrations

⁵⁴ “Osmosis Through the Membrane of an Egg” from Live Demonstrations

⁵⁵ “Osmosis and Osmotic Pressure” Animation from Instructor’s Resource CD/DVD

⁵⁶ “Osmosis and the Egg Membrane” from Live Demonstrations

⁵⁷ “Seawater Gets Fresh” from Further Readings

⁵⁸ Figure 13.25 from Transparency Pack

- 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.

Determination of Molar Mass^{59,60,61,62}

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

FORWARDS REFERENCES

- Desalination via reverse osmosis will be described in Chapter 18 (section 18.4).

13.6 Colloids^{63,64,65}

- **Colloids** or **colloidal dispersions** are suspensions in which the suspended particles are larger than molecules but too small to separate out of the suspension due to gravity.
 - Particle size: 5 to 1000 nm.
 - A colloid particle may consist of a single giant molecule.
 - Example: hemoglobin has molecular dimensions of $6.5 \times 5.5 \times 5.0$ nm and a molar mass of 64,500 g/mol.
- There are several types of colloids:
 - aerosol: gas + liquid or solid (e.g., fog and smoke),
 - foam: liquid + gas (e.g., whipped cream),
 - emulsion: liquid + liquid (e.g., milk),
 - sol: liquid + solid (e.g., paint),
 - solid foam: solid + gas (e.g., marshmallow),
 - solid emulsion: solid + liquid (e.g., butter),
 - solid sol: solid + solid (e.g., ruby glass).
- The **Tyndall effect** is the ability of colloidal particles to scatter light.
 - The path of a beam of light projected through a colloidal suspension can be seen through the suspension.

Hydrophilic and Hydrophobic Colloids^{66,67}

- Focus on colloids in water.
 - Water-loving colloids are **hydrophilic**.
 - Water-hating colloids are **hydrophobic**.
- In the human body, large biological molecules such as proteins are kept in suspension by association with surrounding water molecules.

⁵⁹ “Determination of Molar Mass” Activity from Instructor’s Resource CD/DVD

⁶⁰ Figure 13.27 from Transparency Pack

⁶¹ “Salts Are Mostly NOT Ionized” from Further Readings

⁶² “J.H. van’t Hoff” from Further Readings

⁶³ “The Impact of Colloid Science” from Further Readings

⁶⁴ “Colloidal Systems” from Further Readings

⁶⁵ “Color of the Sunset: The Tyndall Effect” from Live Demonstrations

⁶⁶ “Illustrating the Properties of Magnetic Sand” from Live Demonstrations

⁶⁷ “Put the Brakes on Wastewater Emulsions” from Further Readings

- These macromolecules fold up so that hydrophobic groups are away from the water (inside the folded molecule).
- Hydrophilic groups are on the surface of these molecules and interact with solvent (water) molecules.
 - Typical hydrophilic groups are polar (containing C–O, O–H, N–H bonds) or charged.
- Hydrophobic colloids need to be stabilized in water.
 - One way to stabilize hydrophobic colloids is to adsorb ions on their surface.
 - *Adsorption*: when something sticks to a surface, we say that it is adsorbed.
 - If ions are adsorbed onto the surface of a colloid, the colloid appears hydrophilic and is stabilized in water.
 - Consider a small drop of oil in water.
 - Add a small amount of sodium stearate.
 - Sodium stearate has a long hydrophobic hydrocarbon tail and a small hydrophilic head.
 - The hydrophobic tail can be absorbed into the oil drop, leaving the hydrophilic head on the surface.
 - The hydrophilic heads then interact with the water and the oil drop is stabilized in water.
 - A soap acts in a similar fashion.
 - Soaps are molecules with long hydrophobic tails and hydrophilic heads that remove dirt by stabilizing the colloid in water.
 - Most dirt stains on people and clothing are oil-based.
- Biological application of this principle:
 - The gallbladder excretes a fluid called bile.
 - Bile contains substances (bile salts) that form an emulsion with fats in our small intestine.
 - Emulsifying agents help form an emulsion.
 - Emulsification of dietary fats and fat-soluble vitamins is important in their absorption and digestion by the body.

Removal of Colloid Particles^{68,69,70,71}

- We often need to separate colloidal particles from the dispersing medium.
- This may be problematic.
 - Colloid particles are too small to be separated by physical means (e.g., filtration).
 - However, colloid particles often may be *coagulated* (enlarged) until they can be removed by filtration.
- There are various methods of coagulation.
 - Colloid particles move more rapidly when the colloidal dispersion is heated, increasing the number of collisions. The particles stick to each other when they collide.
 - Adding an electrolyte neutralizes the surface charges on the colloid particles.
- A biological application of another approach to separating colloidal particles from the suspending medium is dialysis.
 - In *dialysis*, a semipermeable membrane is used to separate ions from colloidal particles.
 - In kidney dialysis, the blood is allowed to pass through a semipermeable membrane immersed in a washing solution.
 - The washing solution is isotonic in ions that must be retained.
 - The washing solution does not have the waste products that are found in the blood.

⁶⁸ “Surfactants: The Ubiquitous Amphiphiles” from Further Readings

⁶⁹ “Clearly Cleaner” from Further Readings

⁷⁰ “Blood-Chemistry Tutorials: Teaching Biological Applications of General Chemistry Material” from Further Readings

⁷¹ “Chemistry and Physics in the Kitchen” from Further Readings

- Wastes therefore dialyze out of the blood (move from the blood into the washing solution).
- The "good" ions remain in the blood.

FORWARDS REFERENCES

- Adsorption and absorption in heterogeneous catalysis will be mentioned in Chapter 14 (section 14.7).
- A model of hemoglobin in blood will be provided in Chapter 23 and (sections 23.3 and 23.6) and Chapter 24 (section 24.9).
- Surfactant organic molecules will be mentioned again in Chapter 24 (section 24.1).

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Chapter 13. Properties of Solutions

Media Resources

Figures and Tables in Transparency Pack:

Figure 13.3 Dissolution of an Ionic Solid in Water

Figure 13.4 Enthalpy Changes Accompanying the Solution Process

Figure 13.11 Hydrogen Bonding Involving OH Groups

Figure 13.18 Solubilities of Some Ionic Compounds in Water as a Function of Temperature

Figure 13.19 Solubilities of Four Gases in Water as a Function of Temperature

Figure 13.20 Calculating Molality and Molarity from Solute Mass, Solvent Mass, and Solution Density

Figure 13.23 Phase Diagram Illustrating Boiling-Point Elevation

Figure 13.24 Phase Diagram Illustrating Freezing-Point Depression

Figure 13.25 Osmosis

Figure 13.27 Ion Pairing and Colligative Properties

Section:

13.1 The Solution Process

13.1 The Solution Process

13.3 Factors Affecting Solubility

13.3 Factors Affecting Solubility

13.3 Factors Affecting Solubility

13.4 Expressing Solution Concentration

13.5 Colligative Properties

13.5 Colligative Properties

13.5 Colligative Properties

13.5 Colligative Properties

Activities:

Boiling-Point Elevation and Freezing-Point Depression

Determination of Molar Mass

Section:

13.5 Colligative Properties

13.5 Colligative Properties

Animations:

Dissolution of NaCl in Water

Henry's Law

Raoult's Law

Osmosis and Osmotic Pressure

Section:

13.1 The Solution Process

13.3 Factors Affecting Solubility

13.5 Colligative Properties

13.5 Colligative Properties

Movies:

Dissolution of KMnO_4 in Water

Section:

13.1 The Solution Process

3-D Models:

Water

Sodium Chloride (1×1 Unit Cell)

Hydrated Magnesium Cation

Deprotonated Hydrated Aluminum Cation

Hydrated Aluminum Cation

Pentane

Acetone

Chloroform

Ethanol

Vitamin C (ascorbic acid)

Vitamin A

Section:

13.1 The Solution Process

13.3 Factors Affecting Solubility

13.3 Factors Affecting Solubility

Cyclohexane
 Glucose
 Ibuprofen
 Alanine
 Caffeine
 Adrenaline

13.3 Factors Affecting Solubility
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 13.3 Factors Affecting Solubility
 13.3 Factors Affecting Solubility
 13.4 Expressing Solution Concentration
 13.5 Colligative Properties

Other Resources

Further Readings:

The Use of Dots in Chemical Formulas
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 Applications of Solubility Data
 Henry's Law and Noisy Knuckles
 Henry's Law: A Retrospective
 Soft Drink Bubbles
 Candy Sprinkles to Illustrate One Part Per Million
 An Alternative Introduction to the Mole Fraction
 Mole Fraction Analogies
 One Cool Chemist
 Antifreeze Solutions: The Colligative Properties of Antifreeze
 Freeze-Proof Bugs
 Seawater Gets Fresh
 Salts Are Mostly NOT Ionized
 J. H. van't Hoff
 The Impact of Colloid Science
 Colloidal Systems
 Put the Brakes on Wastewater Emulsions
 Surfactants: The Ubiquitous Amphiphiles
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Section:

13.1 The Solution Process
 13.2 Saturated Solutions and Solubility
 13.3 Factors Affecting Solubility
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 13.3 Factors Affecting Solubility
 13.4 Expressing Solution Concentration
 13.4 Expressing Solution Concentration
 13.4 Expressing Solution Concentration
 13.5 Colligative Properties
 13.6 Colloids
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13.1 The Solution Process
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A Simple Demonstration Model of Osmosis
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13.5 Colligative Properties
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13.6 Colloids
13.6 Colloids