

# Section 11.1

## The Dissolution Process



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# Learning Objectives



- Describe the basic properties of solutions and how they form
- Predict whether a given mixture will yield a solution based on molecular properties of its components
- Explain why some solutions either produce or absorb heat when they form

# Solution Vocabulary



- **Solutions** are homogeneous mixtures of two or more substances.
- The **Solvent** is the component of the solution with significantly higher concentration.
- **Solutes** are components of the solution present in relatively smaller concentrations.
- When a solution forms solute is distributed evenly throughout the solvent.

# Solution Phases



- Components of a solution can be any phase.
- **Aqueous** solutions (aq) have a water solvent.
- **Alloys** are solid solutions of two metals.
- Air is solution of gases with a nitrogen solvent.

Solution	Solute	Solvent
air	$O_2(g)$	$N_2(g)$
soft drinks <sup>1</sup>	$CO_2(g)$	$H_2O(l)$
hydrogen in palladium	$H_2(g)$	$Pd(s)$
rubbing alcohol	$H_2O(l)$	$C_3H_8O(l)$ (2-propanol)
saltwater	$NaCl(s)$	$H_2O(l)$
brass	$Zn(s)$	$Cu(s)$

# Solution Properties



- They are homogeneous
  - After a solution is mixed, its composition is uniform.
- The physical state of a solution is typically the same as that of the solvent.
- The components of a solution are dispersed on a molecular scale
- They consist of a mixture of separated solute particles
  - Molecules, atoms, and/or ions are each closely surrounded by solvent species.

# Solution Properties



- The dissolved solute in a solution will not settle out or separate from the solvent.
- The composition of a solution, or the concentrations of its components, can be varied continuously.
  - Within limits determined by the solubility

# Spontaneity

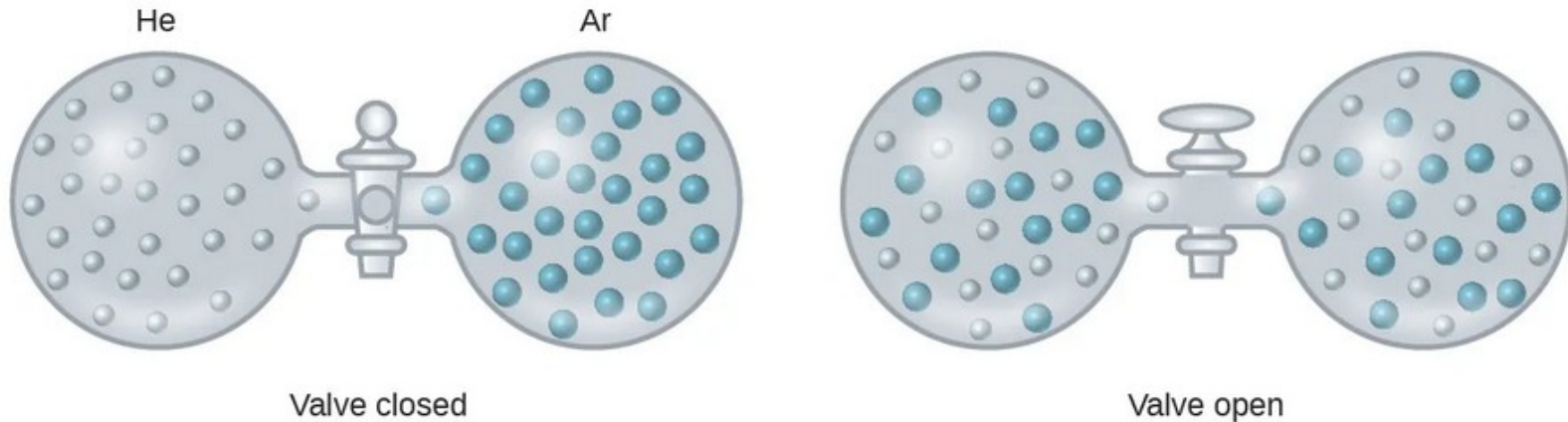


- The formation of a solution is **spontaneous process**
  - Occurs under specified conditions without requiring the input of external energy.
  - Decrease the internal energy of the system, exothermic,  $\Delta H < 0$ .
  - Increases the dispersal of matter in the system,  $\Delta S > 0$ .
- Heating and stirring may *speed up* the dissolution process. But they are not *necessary*.
  - Stirring will only speed up dissolution, it will not increase or decrease the amount of solute dissolved.
  - Heating can speed or slow the dissolution process, it also effects the amount of solute dissolved.

# Ideal Solutions



- **Ideal Solutions** occur when the strengths of the intermolecular forces of attraction between solute and solvent species are the same the IMFs present in the separated components
  - The solution is formed with no accompanying energy change,  $\Delta H = 0$ .





# Ideal Solutions



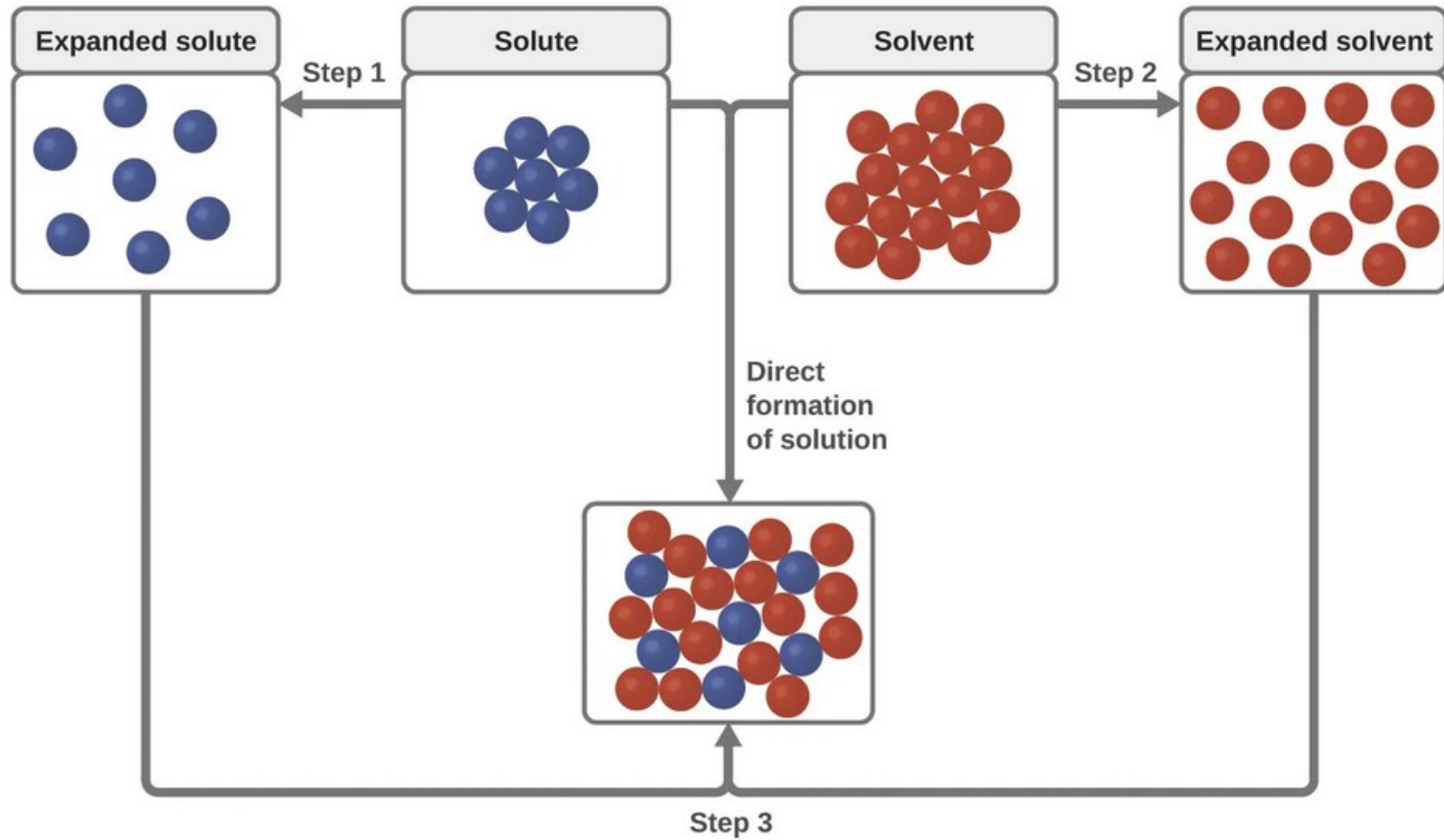
- Ideal solutions may also form when structurally similar liquids are mixed.
- Some ideal solutions do experience intermolecular attractive forces. But since the molecules of the two substances being mixed are structurally very similar, the intermolecular attractive forces between like and unlike molecules are essentially the same.

# Solvation



- When a solution is formed, three steps occur
  - Solute-solute attractions are overcome (endothermic)
  - Solvent-solvent attractions are overcome (endothermic)
  - Solvent-solute attractions are formed (exothermic)
- The process of forming solvent-solute attractions is called **solvation**.
- Solutions form when the energy released during solvation is greater than the energy required to overcome solute-solute and solvent-solvent interactions.

# Solvation



# Endothermic Dissolution



- Most spontaneous solvation reactions are exothermic.
- It is possible to have an endothermic solvation reaction.
- These reactions are driven entirely by the increase in disorder caused by dissolution.



# Section 11.2

## Electrolytes



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# Learning Objectives



- Define and give examples of electrolytes
- Distinguish between the physical and chemical changes that accompany dissolution of ionic and covalent electrolytes
- Relate electrolyte strength to solute-solvent attractive forces

# Electrolytes



- **Electrolytes** produce ions when dissolved in water.
- **Nonelectrolytes** do not produce ions when dissolved in water.
  - Sugar, most gases
- **Strong Electrolytes** ionize completely in solution.
  - Sodium Chloride, Potassium Bromide
- **Weak Electrolytes** ionize partially in solution.

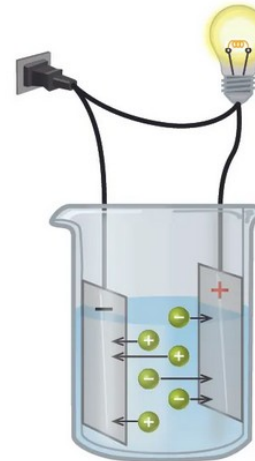
# Conductivity of Solutions



- When ions are dissolved in the water they become mobile conducting electricity.
  - Water itself is not conductive
- The degree of conductivity allows us to determine the strength of an electrolyte.



ethanol  
No conductivity



KCl  
High conductivity



acetic acid solution  
Low conductivity

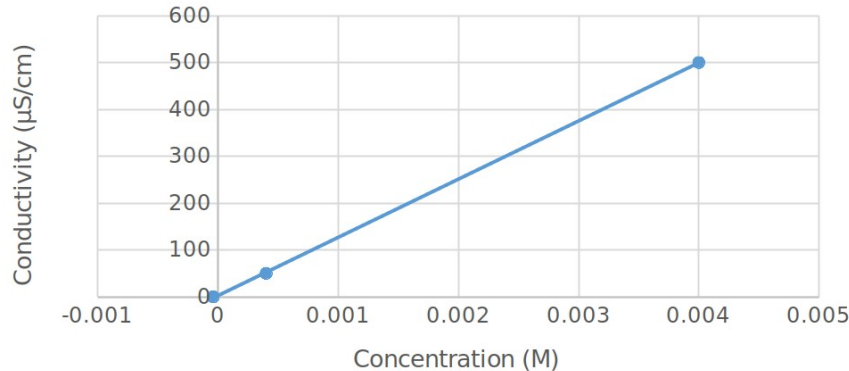


# Quantitative Conductivity

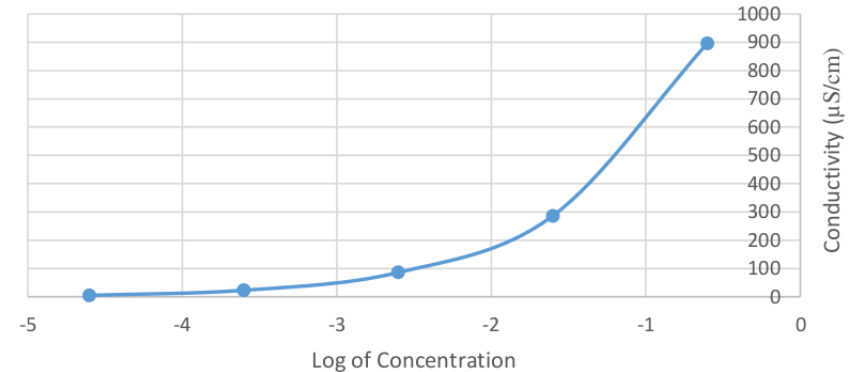


- Strong Electrolytes will have a linear relationship between concentration and conductivity
- Weak electrolytes will have nonlinear relationship.

Strong Electrolyte (Ammonium Chloride):  
Conductivity vs. Concentration



Weak Electrolyte (Acetic Acid): Conductivity vs.  
Log of Concentration

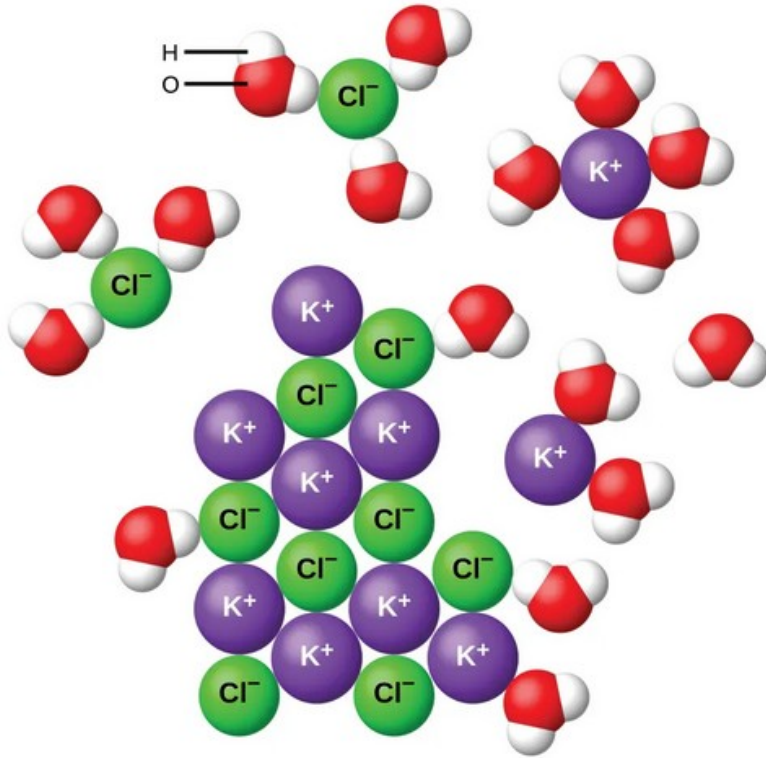


# Ionic Electrolytes



- The electrostatic attraction between an ion and a molecule with a dipole is called an **ion-dipole attraction**.
- Ions in the solid separate and disperse uniformly throughout the solution, **dissociate**, because water molecules surround and solvate the ions.
- Most ionic solutes are strong electrolytes.
- Even sparingly, soluble ionic compounds are strong electrolytes, since the small amount that does dissolve will dissociate completely.

# Solvating Electrolytes



- As potassium chloride (KCl) dissolves in water, the ions are hydrated. The polar water molecules are attracted by the charges on the  $K^+$  and  $Cl^-$  ions. Water molecules in front of and behind the ions are not shown.

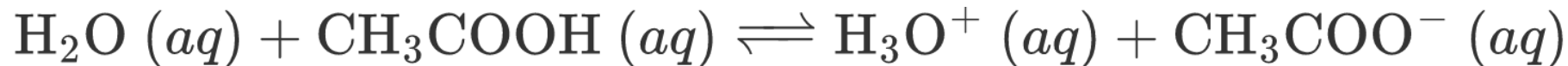
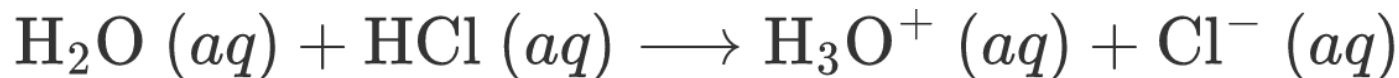
# Covalent Electrolytes



- Pure water is a poor conductor because it does not dissociate much in itself.



- Some covalent solvents will react with water to form ions.
  - For example, acids like HCl or CH<sub>3</sub>COOH



# Section 11.3

## Solubility



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# Learning Objectives



- Describe the effects of temperature and pressure on solubility
- State Henry's law and use it in calculations involving the solubility of a gas in a liquid
- Explain the degrees of solubility possible for liquid-liquid solutions

# Solubility



- **Solubility** of a solute in a particular solvent is the maximum concentration that may be achieved, under given conditions, when the dissolution process is at *equilibrium*.
- A solution whose solute concentration is equal to its solubility is **saturated**.
- If the solute concentration is less than its solubility, the solution is **unsaturated**.
- Solutions with relatively low solute concentration is **dilute**.
- Solutions with relatively high solute concentration is **concentrated**.

# Supersaturated



- Solutions with solute concentration exceeding its solubility is called **supersaturated**.
  - This is a nonequilibrium condition
- Solute will leave the solution until equilibrium is reached.



# Solutions of Gases in Liquids

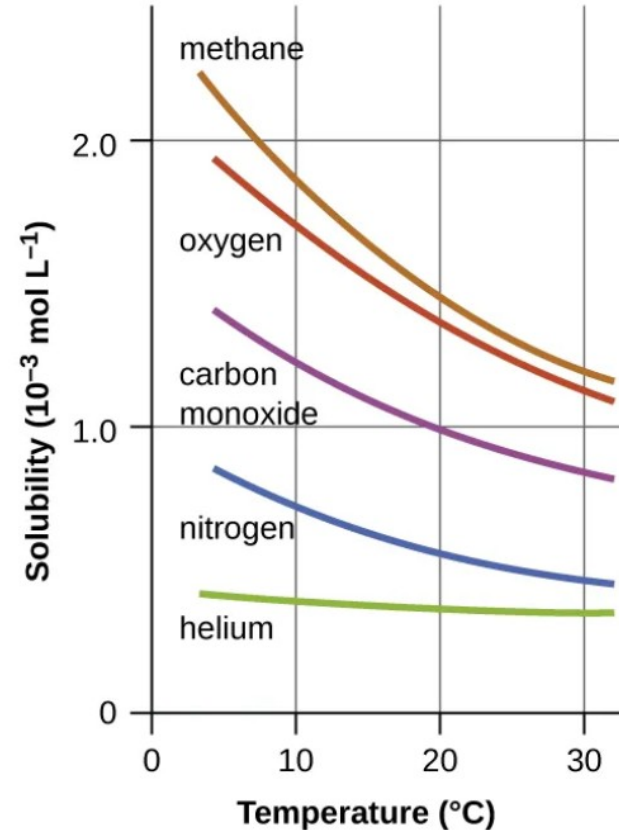


- Gases have only very weak intermolecular forces, therefore there is little or no solute-solute interaction to overcome during solvation.
- When the solvent-solute interactions exceed solvent-solvent interactions in strength, solvation will occur.
- Consequently, gases are more soluble in solvents with weak IMFs and gases capable of forming stronger IMFs are more soluble.

# Solubility of Gases and Temperature



- Typically, gases are more soluble in colder liquids.
- The decreased solubility of oxygen, carbon dioxide, and other gases in water is a major consequence of climate change.

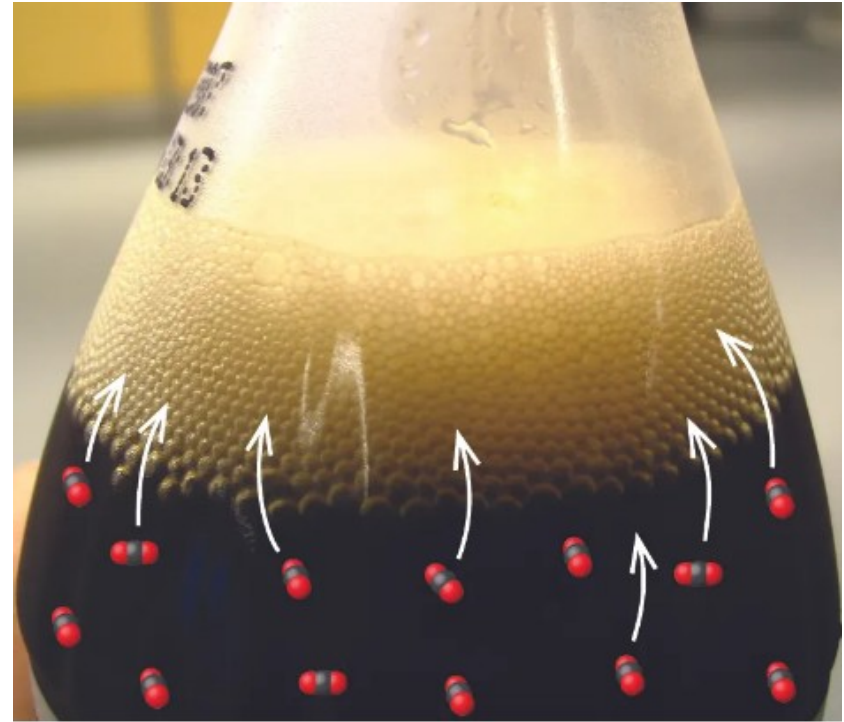


# Gas Solubility and Partial Pressure



- The solubility of gases increase with the partial pressure of the gas above a liquid.
- Many gases under most conditions will follow *Henry's Law*.

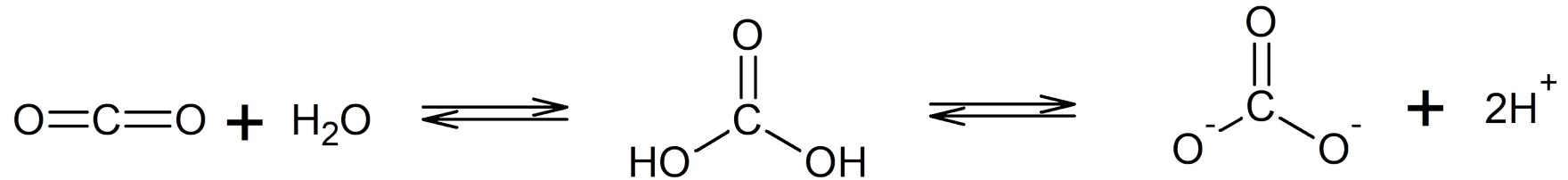
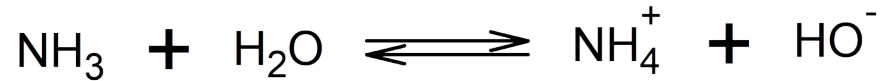
$$C_g = kP_g$$



# Deviations From Henry's Law



- The solubility of gases will deviate from Henry's Law when chemical reactions occur between the solute and solvent.



# Solutions of Liquids in Liquids



- Sometimes solutes and solvents can be mixed in any proportions i.e. they have **infinite solubility**.
- Liquid-liquid solutions that are infinite soluble are called **miscible**.
- Miscible liquids will have very similar properties.
  - Solute-solute, solvent-solvent, and solute-solvent IMFs will have comparable strength.
- Polar solvents will dissolve polar solvents and nonpolar solvents will dissolve nonpolar solvents.
  - *Like dissolves like*

# Solutions of Liquids in Liquids

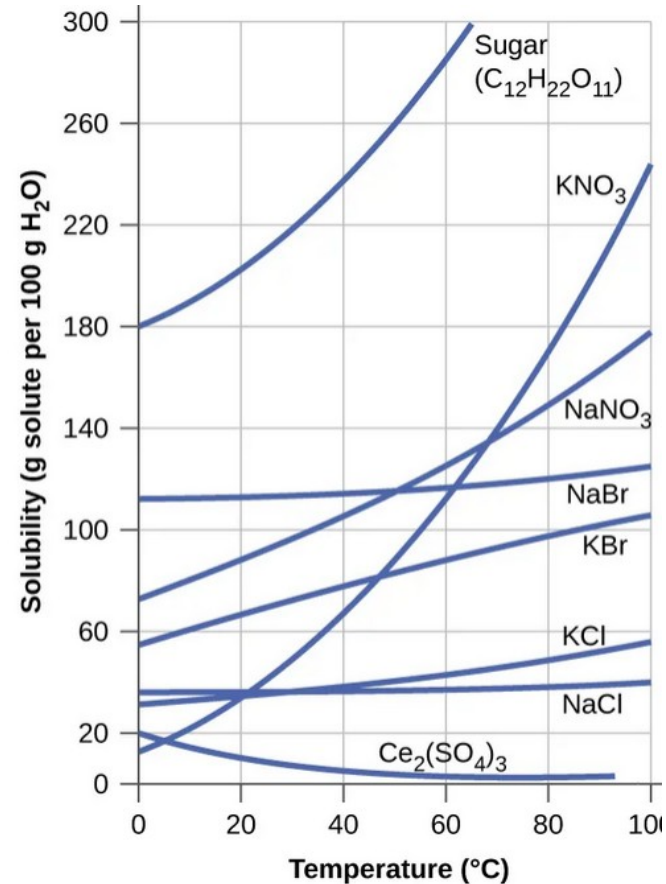


- When only a moderate portion of one liquid will dissolve in another the liquids are **partially miscible**.
- When two liquids do not mix they are called **immiscible**.
- Immiscible liquids will form two layers with the less dense liquid on top.
- It is rarely, if ever, the case that two liquids have zero mutual solubility.

# Solutions of Solids in Liquids



- Typically, the solubility of solids in liquids increases with temperature.
- Leveraging the temperature dependence of the solubility of solid-liquid solutions is the most common way of preparing supersaturated solutions.



# Section 11.4

## Colligative Properties



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# Learning Objectives



- Express concentrations of solution components using mole fraction and molality
- Describe the effect of solute concentration on various solution properties (vapor pressure, boiling point, freezing point, and osmotic pressure)
- Perform calculations using the mathematical equations that describe these various colligative effects
- Describe the process of distillation and its practical applications
- Explain the process of osmosis and describe how it is applied industrially and in nature



# Concentration Units

# Concentration Unit Review



- Molarity

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

- Molality

$$m = \frac{\text{mol solute}}{\text{kg solvent}}$$

- Mol Fraction

$$X_A = \frac{\text{mol A}}{\text{total mol of all components}}$$

# Converting from Molarity to Molality



$$M = \frac{\text{mols}_{\text{solute}}}{V_{\text{soln}}} \quad \text{where} \quad V_{\text{soln}} = 1 \text{ L}$$

$$V_{\text{soln}} \rho_{\text{soln}} = m_{\text{soln}} = m_{\text{solute}} + m_{\text{solvent}}$$

$$m_{\text{soln}} - m_{\text{solute}} = m_{\text{soln}} - (\text{mols}_{\text{solute}})M^W = m_{\text{solvent}}$$

$$m = \frac{\text{mols}_{\text{solute}}}{m_{\text{solvent}}} \quad \text{where } m_{\text{solvent}} \text{ has units of kg}$$

# Converting from Molality to Molarity



$$m = \frac{\text{mols}_{\text{solute}}}{\text{m}_{\text{solvent}}} \quad \text{Where } \text{m}_{\text{solvent}} = 1 \text{ kg}$$

$$\text{m}_{\text{soln}} = \text{m}_{\text{solvent}} + \text{m}_{\text{solute}} = \text{m}_{\text{solvent}} + \text{mols}_{\text{solute}} M^{\text{W}}$$

$$V_{\text{soln}} = \frac{\text{m}_{\text{soln}}}{\rho_{\text{soln}}}$$

$$M = \frac{\text{mols}_{\text{solute}}}{V_{\text{soln}}}$$



# Molality to Mol Fraction

$$m_a = \frac{\text{mols A}}{m_{\text{solvent}}}$$

$$m_b = \frac{\text{mols B}}{m_{\text{solvent}}}$$

$$m_c = \frac{\text{mols C}}{m_{\text{solvent}}}$$

$$m_{\text{solvent}} = 1 \text{ kg}$$

$$\text{mols}_{\text{solvent}} = m_{\text{solvent}} M_{\text{solvent}}^W$$

$$X_a = \frac{\text{mols}_a}{\text{mols}_a + \text{mols}_b + \text{mols}_c + \text{mols}_{\text{solvent}}}$$



# Colligative Properties

# Colligative Properties



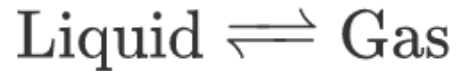
- Solution properties differ from those of pure solutes and solvents.
- Most of the time, solution properties depend on the identity of the solutes and solvents.
- **Colligative Properties**, however, depend only of the concentration of the solute.
- They include: vapor pressure lowering, boiling point elevation, freezing point depression, and osmotic pressure.



# Vapor Pressure Lowering



- Vapor pressure is the partial pressure of a liquid at equilibrium.



- When a solute is dissolved in a liquid the vapor pressure will be lowered.
- Solute molecules occupy sites at the liquids surface blocking solvent molecules access to the gas phase.
- Solvent molecules can, however, recombine with the liquid phase on top of solute molecules.
- The result is fewer solvent molecules in the gas phase.

# Raoult's Law



- The partial pressure exerted by any component of an ideal solution is equal to the vapor pressure of the pure component multiplied by its mole fraction in the solution.

$$P_A = X_A P_A^*$$

- Raoult's Law can be combined with Dalton's Law to yield an expression of the total pressure exerted by the solution on the gas phase.

$$P_{\text{solution}} = \sum_i P_i = \sum_i X_i P_i^*$$

# Nonvolatile Solutes



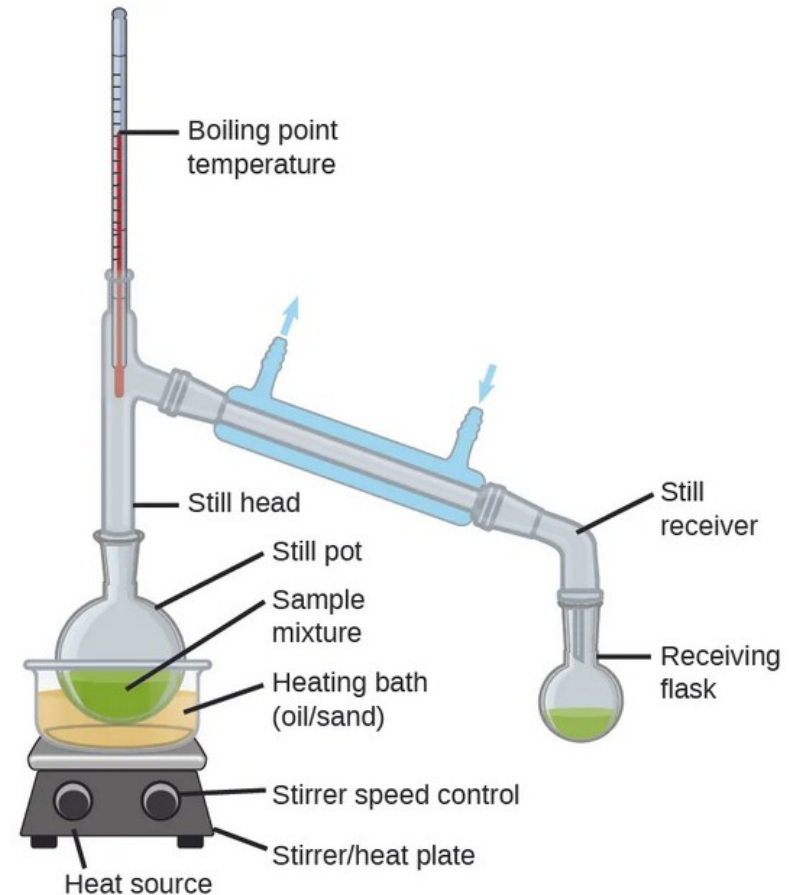
- When a solute's vapor pressure is essentially zero, it is said to be **nonvolatile**.
- When a solution contains only nonvolatile solutes, the solution's total pressure depends solely on the solvent.

$$P_{\text{solution}} = X_{\text{solvent}} P_{\text{solvent}}^*$$

# Distillation



- Raoult's Law can be used to prove that the vapor released by a solution will be enriched in the less volatile component.
- This explains why distillation is an effective means of separating the components of a solution.



# Boiling Point Elevation



- **Boiling Point Elevation** is the result of vapor pressure lowering.
- A higher temperature is required to reach a pressure equal to atmospheric pressure.
- The increase in temperature required is proportional to molality of the solution.

$$\Delta T_b = K_b m$$

- Where  $K_b$  is the **boiling point elevation constant**.
  - $K_b$  is unique to each solvent.

# Freezing Point Depression



- Solutions freeze at lower temperatures than pure liquids.
- Freezing Point Depression is directly proportional to the molality of the solution.

$$\Delta T_f = K_f m$$

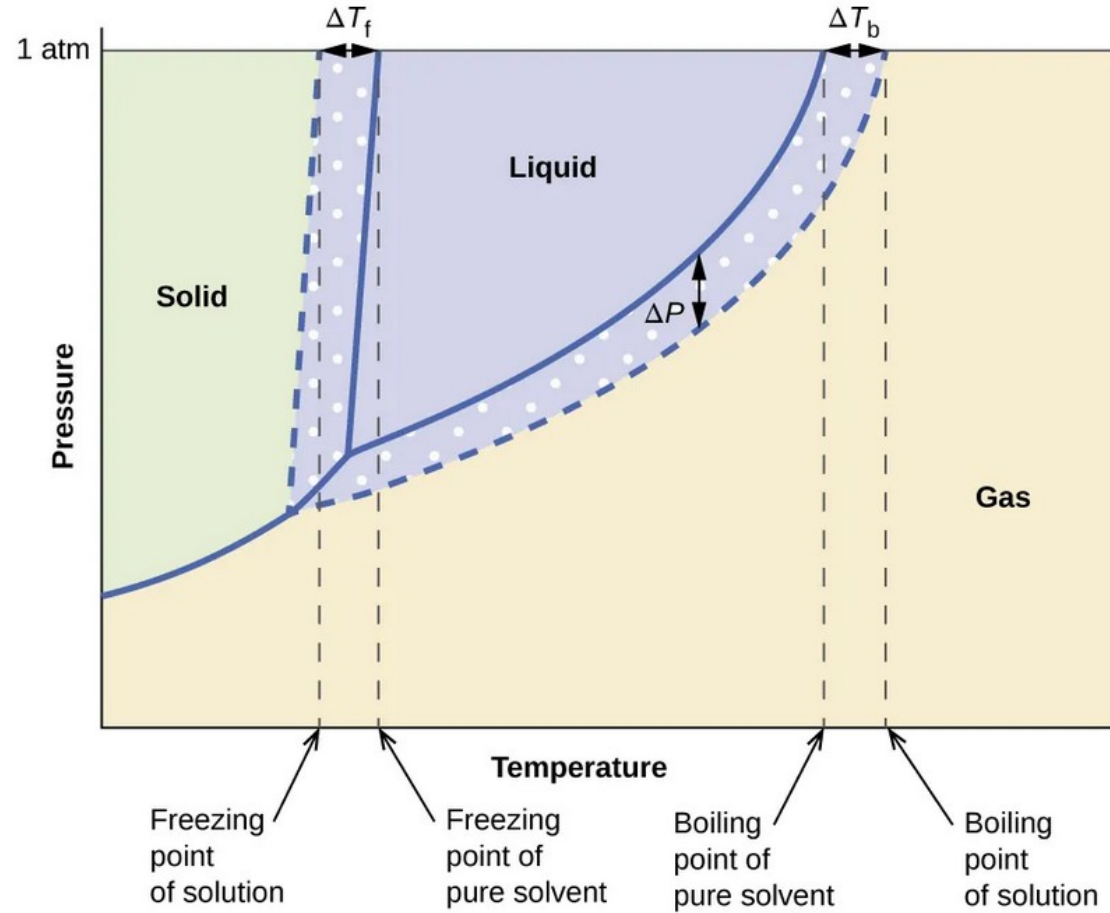
- $K_f$  is the **freezing point depression constant**.
  - $K_f$  is unique to each solvent.

# Common $K_b$ and $K_f$ Values



Solvent	Boiling Point ( $^{\circ}\text{C}$ at 1 atm)	$K_b$ ( $^{\circ}\text{Cm}^{-1}$ )	Freezing Point ( $^{\circ}\text{C}$ at 1 atm)	$K_f$ ( $^{\circ}\text{Cm}^{-1}$ )
water	100.0	0.512	0.0	1.86
hydrogen acetate	118.1	3.07	16.6	3.9
benzene	80.1	2.53	5.5	5.12
chloroform	61.26	3.63	-63.5	4.68
nitrobenzene	210.9	5.24	5.67	8.1

# Solution Phase Diagram





# Osmosis



- Some membranes only allow molecules or ions of a certain size, shape, polarity, etc. to pass through them.
- These membranes are called **semipermeable**.
- Osmosis occurs when
  - A semipermeable membrane separates two solutions
  - The same solvent exists on either side
  - Only the solvent molecules can permeate the membrane.
- Solvent molecules will diffuse across the membrane in both directions but at different rates. This is called **Osmosis**.

# Osmotic Pressure

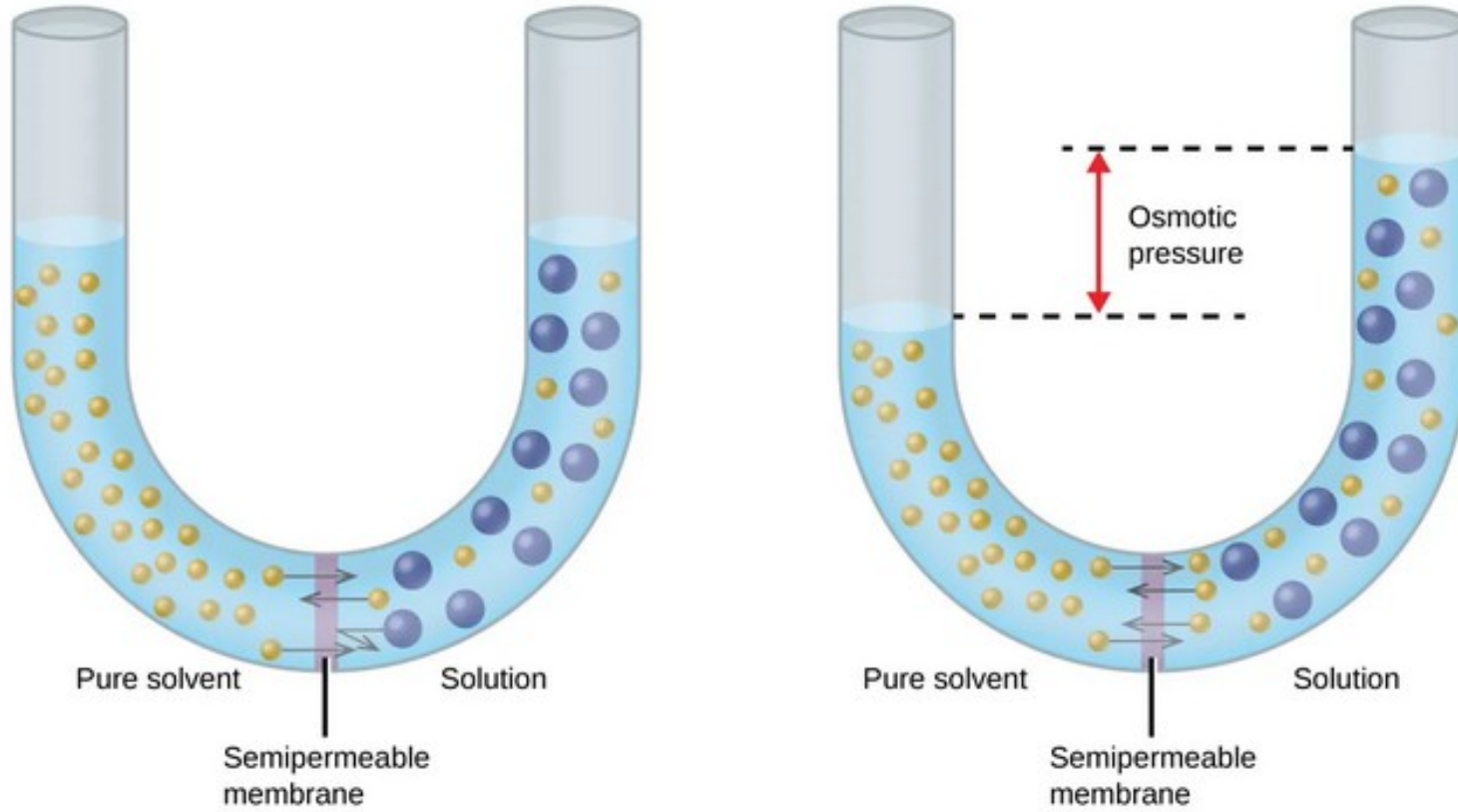


- The net effect of osmosis is solvent molecules accumulating on the side of the membrane with a more concentrated solution until one of two conditions is met.
  - The concentration of the two solutions is equal
  - The pressure differential across the membrane is equal to the osmotic pressure ( $\Pi$ ).

$$\Pi = MRT$$

- Where  $M$  is the molarity,  $T$  is the temperature, and  $R$  is the gas law constant.

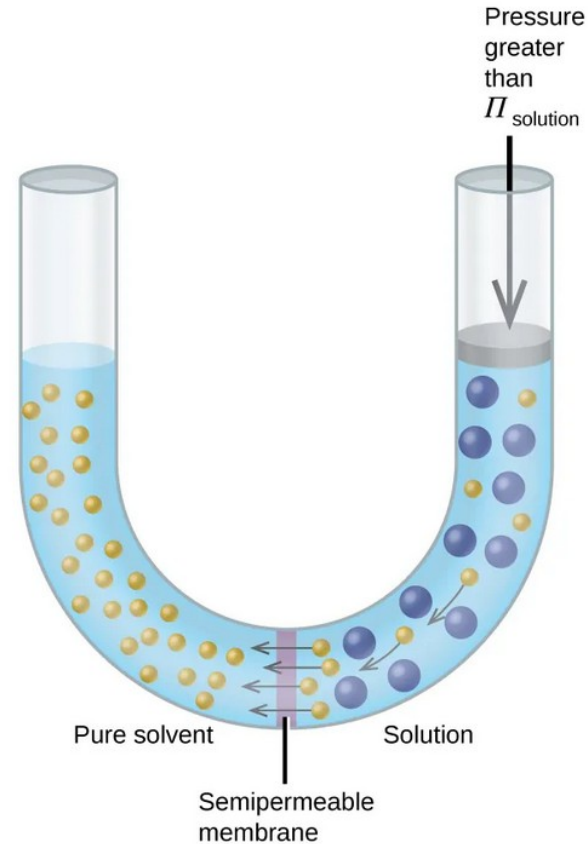
# Osmosis and Osmotic Pressure



# Reverse Osmosis



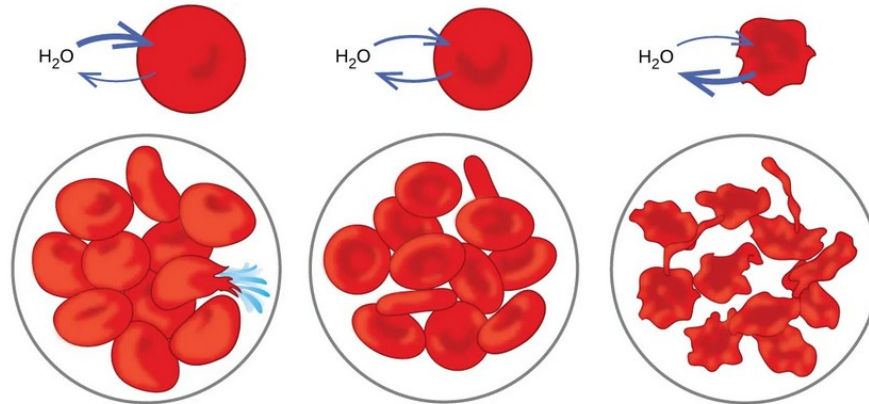
- When pressure greater than the osmotic pressure is applied to the more concentrated solution, the process can be reversed.



# Isotonic, Hypotonic, and Hypertonic



- When the osmotic pressure between the two solutions is zero they are called **isotonic**.
- When a solution is less concentrated than another it is called **hypotonic**.
- When a solution is more concentrated than another it is called **hypertonic**.



# Determination of Molar Mass



- Colligative Properties can be used to determine the molar mass of an unknown.

- 1) Add a known mass of solute to a known amount of solvent.
- 2) Measure the colligative property,  $\Delta T_f$ ,  $\Delta T_b$ , or  $\Pi$ .
- 3) Calculate the concentration ( $m$  or  $M$ ) of the solute.

$$m = \frac{\Delta T_f}{K_f} \quad m = \frac{\Delta T_b}{K_b} \quad M = \frac{\Pi}{RT}$$

- 4) Determine the number of mols present

$$\text{mols} = (m)(m_{\text{solvent}}) = (M)(V_{\text{soln}})$$

- 5) Calculate the Molar Mass

$$M^W = \frac{m_{\text{solute}}}{\text{mols}}$$

# Colligative Properties of Electrolytes



- When electrolytes dissolve they dissociate into ions.
- To account for this an factor called the **van't Hoff factor** is included in the colligative property equations.

$$i = \frac{\text{moles of particles in solution}}{\text{moles of formula units dissolved}}$$

$$\Delta T_f = iK_f m \qquad \Delta T_b = iK_b m$$

# Section 11.5

## Colloids



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# Learning Objectives



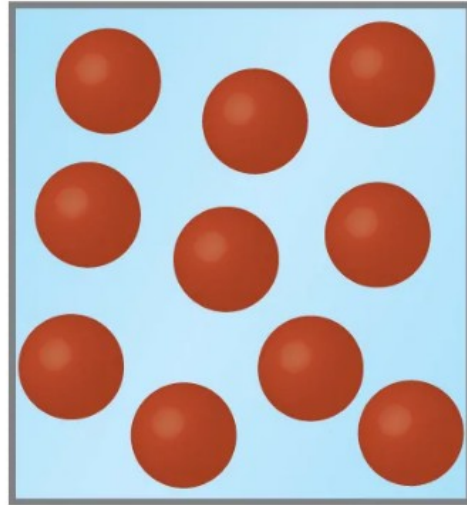
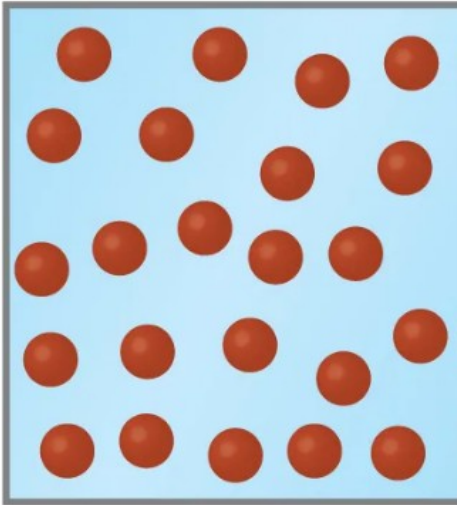
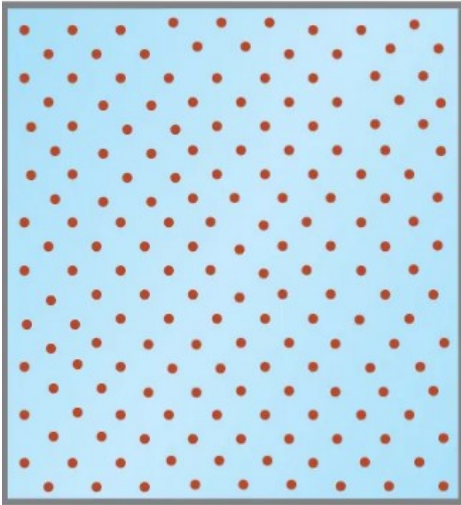
- Describe the composition and properties of colloidal dispersions
- List and explain several technological applications of colloids

# Suspensions and Colloids



- **Suspensions** are heterogeneous mixtures composed of relatively large particles that are visible.
  - They are cloudy
  - The suspended particles settle out after mixing
- **Colloids** have intermediate properties between suspensions and solutions.
  - Particles are larger than most molecules but smaller than suspension particles
  - The particles do not settle out upon standing

# Examples



# Tyndall Effect



- Particles in a colloid are large enough to scatter light
- This can make colloidal mixtures appear cloudy or opaque



# Colloid Facts



- The particles that make up colloids can be hundreds or thousands small molecules or single extremely large molecules.
- Components of a colloid can be any phase.
- Particle component present in relatively small amounts is called the **dispersed phase**.
- The substance or solution the particulate is dispersed in is called the **dispersion medium**.
- If particles aggregate, they can break the colloid.

# Colloid Examples



Dispersed Phase	Dispersion Medium	Common Examples	Name
solid	gas	smoke, dust	—
solid	liquid	starch in water, some inks, paints, milk of magnesia	sol
solid	solid	some colored gems, some alloys	—
liquid	gas	clouds, fogs, mists, sprays	aerosol
liquid	liquid	milk, mayonnaise, butter	emulsion
liquid	solid	jellies, gels, pearl, opal ( $\text{H}_2\text{O}$ in $\text{SiO}_2$ )	gel
gas	liquid	foams, whipped cream, beaten egg whites	foam
gas	solid	pumice, floating soaps	—

# Preparing Colloids



- Colloids are typically prepared by one of two means
  - 1) Dispersion methods: breaking down larger particles. For example, paint pigments are produced by dispersing large particles by grinding in special mills.
  - 2) Condensation methods: growth from smaller units, such as molecules or ions. For example, clouds form when water molecules condense and form very small droplets.

# Emulsions



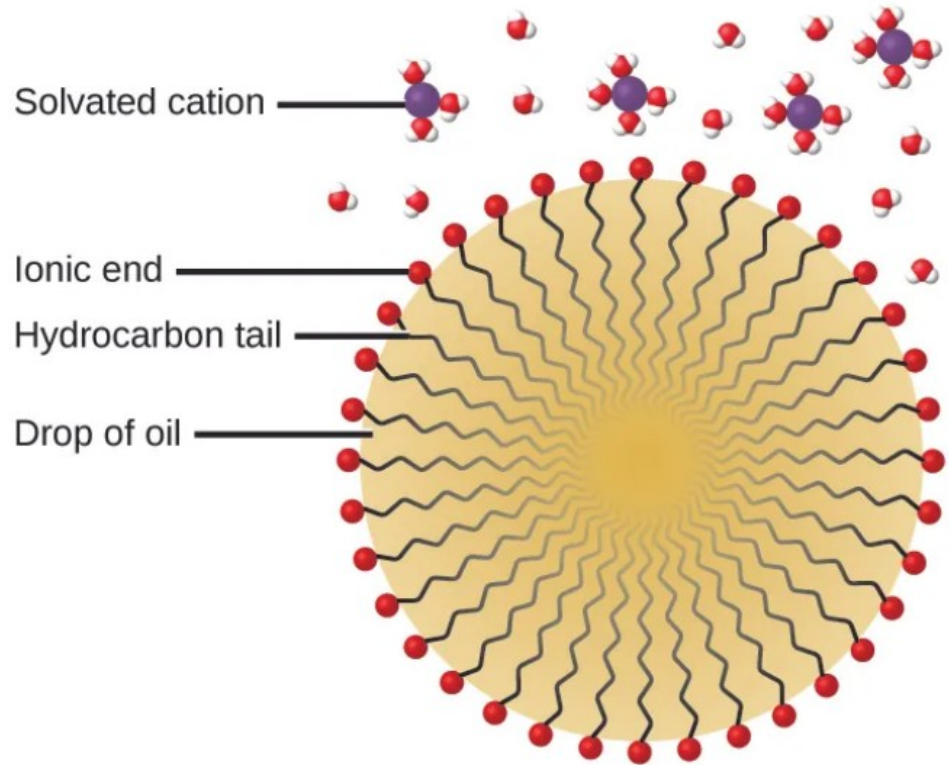
- **Emulsions** are colloids of two immiscible liquids
  - May be prepared by shaking together or blending the liquids.
  - One liquid breaks into droplets of colloidal size, which then disperse throughout the other liquid.
  - The droplets will typically coalesce and reform two phases.
- **Emulsifying Agents** can be used to stabilize emulsions.



# Soaps and Detergents



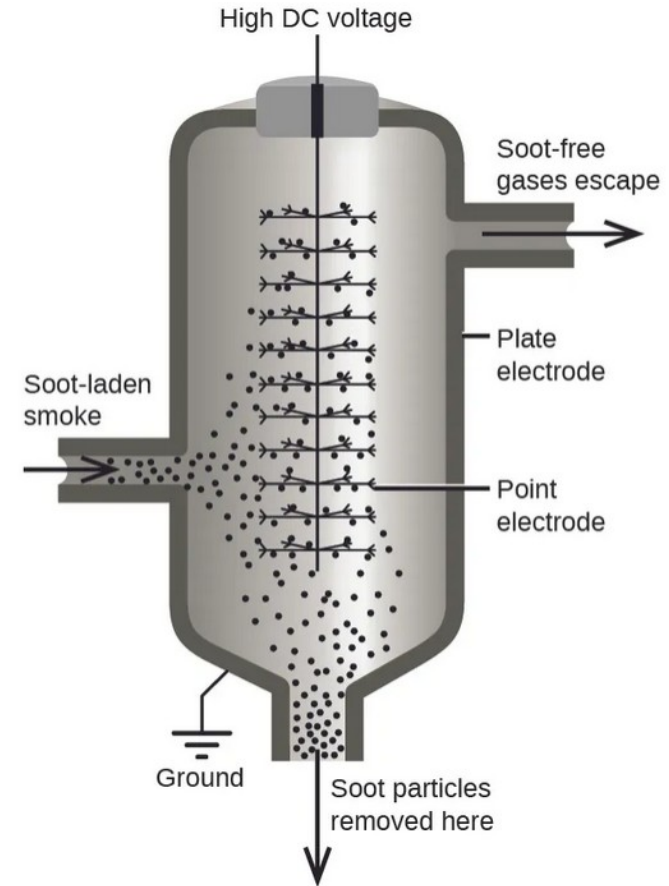
- Molecules with both hydrophobic and hydrophilic ends (**amphiphilic**) can be used as detergents.
- These molecules are capable of trapping dirt in particle formations known as **micelles**.



# Electrical Properties of Colloids



- Dispersed colloidal particles are often electrically charged.
- This helps keep them dispersed because particles containing like charges repel each other.
- The charged nature of some colloidal particles may be exploited to remove them from a variety of mixtures.



# Gels



- A **gel** is a colloidal dispersion of a liquid phase throughout a solid phase.
- The fibers of the dispersing medium form a complex three-dimensional network, the interstices being filled with the liquid medium or a dilute solution of the dispersing medium.

