

Chapter 8. Solutions

8.1 Characteristics of Solutions

A solution has a broader meaning than something dissolved in water. It is a homogeneous mixture of two or more substances in a single phase. Many materials exist as **homogeneous mixtures**: air (g)-mixture of N_2 and O_2 mainly, sea water (aq)-mixture of salt and water, gasoline (l)-mixture of hydrocarbons, stainless steel (s)-alloy of iron and carbon. Water solutions are called aqueous and given the (aq) to identify it in solution process in the dissolved substance in water.

Solution Components: Solvent and solute

- Solution** - homogeneous mixture of two or more substances; often comprised of solvent.
Solute - minor component mixed with solvent. e.g., gases or solids are solutes when dissolved in a liquid such as water.
Solvent - component with same phase as solution; substance present in excess in liquid-liquid mixtures

8.2 Solubility

Solubility refers to amount of solute that goes in to a solvent when mixing of two or more substances. Solubility of a substance could be given in grams or moles of solute per mL of solution. Not all substances will dissolve in all types of solvents.

Solubility is affected by type of attractions forces that are created between solute and the solvent molecules. These solute-solvent interactions given in examples below will provide a basis for understanding the solubility of a compound. Following statements summarize the idea "Like dissolves like!" applied in deciding solubility of a compound.

- 1) **Polar** compound dissolve in **polar** solvents.
- 2) **Nonpolar** compounds dissolve in **nonpolar** solvents.
- 3) Presence of similar intermolecular forces between the solute and solvent also need to be taken into account: London dispersive, dipolar, hydrogen bond.

Polar compounds

Ionic compounds: Such as NaCl and many other salts are made up of ions. Solubility of salt is summarized by the **solubility rules** that are discussed in section 8.4 later.

Polar molecules: H_2O and ammonia with unbalanced polar bonds in the molecule.

a) **Is carbon tetrachloride, CCl_4 , soluble in water?**

Insoluble: CCl_4 is **nonpolar** only has London dispersive and H_2O is **polar** which all there London dispersive, dipolar and hydrogen bond.

Compound	Intermolecular Forces	"Unlike does not dissolve unlike"
H_2O	London dispersive, dipolar, hydrogen bond	
CCl_4	London dispersive	

b) Is dimethyl ether, CH_3OCH_3 , soluble in water? Soluble: CH_3OCH_3 is polar that can form hydrogen bonding attraction with O-H dipoles in water which is more polar molecule.

Compound	Intermolecular Forces	"Like dissolves like"
H_2O	dispersive, dipolar, hydrogen bond	
CH_3OCH_3	dispersive, dipolar, can hydrogen bond with water (but not itself)	

c) Is ethanol, $\text{CH}_3\text{CH}_2\text{OH}$, soluble in water? Soluble: $\text{CH}_3\text{CH}_2\text{OH}$ is polar that can form hydrogen bonding attraction with O-H dipoles in water which is more polar molecule.

Compound	Intermolecular Forces	"Like dissolves like"
H_2O	dispersive, dipolar, hydrogen bond	
$\text{CH}_3\text{CH}_2\text{OH}$	dispersive, dipolar, hydrogen bond	

d) Is hydrogen fluoride, HF , soluble in water? Soluble: HF is polar that can form hydrogen bonding attraction with O-H dipoles in water which is more polar molecule.

Compound	Intermolecular Forces	"Like dissolves like"
H_2O	dispersive, dipolar, hydrogen bond	
HF	dispersive, dipolar, hydrogen bond	

e) Is propane, $\text{CH}_3\text{CH}_2\text{CH}_3$, soluble in water? Insoluble: $\text{CH}_3\text{CH}_2\text{CH}_3$ is nonpolar that cannot form hydrogen bonding attraction with O-H dipoles in water which is more polar molecule.

Compound	Intermolecular Forces	"Unlike does not dissolve unlike"
H_2O	London dispersive, dipolar, hydrogen bond	
$\text{CH}_3\text{CH}_2\text{CH}_3$	London dispersive	

Question: Keeping in mind that "Like Dissolves Like", which of the following compounds would be the most soluble in a nonpolar solvent such as carbon tetrachloride, CCl_4 ?

- a) H_2O **Most polar**
- b) CH_3OH **somewhat polar**
- c) $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{OH}$ (**Correct Answer**) **Least polar** molecule Molecular has a significant nonpolar component therefore most soluble in nonpolar CCl_4 .
- d) $\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$ **polarity is decreasing**

Nonpolar compounds

All nonpolar molecules with no total polarity in the molecule are nonpolar compounds. Some of the concepts of polarity of molecules based on bond polarity using electronegativity and the molecular geometry were introduced in chapter 5.

Non-polar solute - Non-polar solvent: Solid iodine (I_2) dissolves in liquid bromine (Br_2).

Non-polar solute - Polar solvent: O_2 , N_2 , H_2 , and CO_2 are not very soluble in water.

Polar Solute - Polar Solvent- Solute (NH_3 and solvent (H_2O) dissolves readily.

Ionic Solute - Polar Solvent- NaCl dissolves in water readily

Solution terminology

Miscible

Liquids that dissolve in each other.

Immiscible

Liquids that do not dissolve.

Saturated solution

A solution that contains as much it can hold at a certain temperature. A solution that is in equilibrium with undissolved solute is said to be saturated. Additional solute will not dissolve if added to a saturated solution. The amount of solute needed to form a saturated solution in a given quantity of solvent is known as the solubility of that solute.

Unsaturated solution

A solution that contains less than maximum amount at a certain temperature.

Supersaturated solution

A solution that contains more than maximum amount at a certain temperature. Under suitable conditions it is possible to form solutions that contain a greater amount of solute than that needed to form a saturated solution. Such solutions are said to be supersaturated.

Effect of temperature on solubility

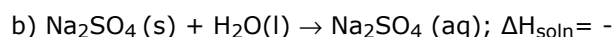
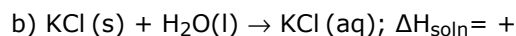
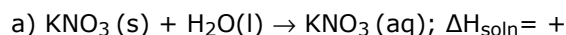
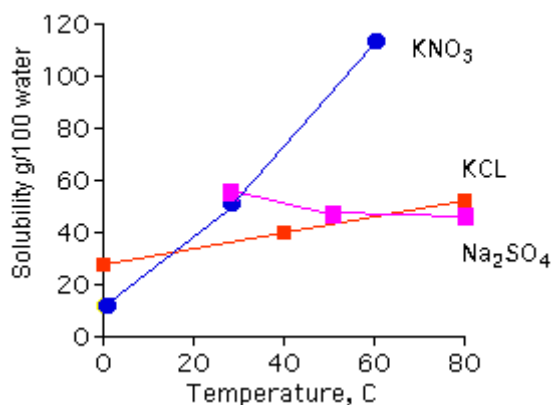
When a liquid or solid solute is dissolved in water the temperature of the resulting solution called **heat of solution** (ΔH_{soln}) either goes up **exothermic** (released) or goes down **endothermic** (absorbed). For example when ammonium nitrate-NH₄NO₃ used as fertilizer it is dissolved in water and the water become colder because heat of solution of is endothermic (absorbed). The solubility of most solids increase with increasing temperature since most solids have endothermic (ΔH_{soln}).

Liquid or solid solutes

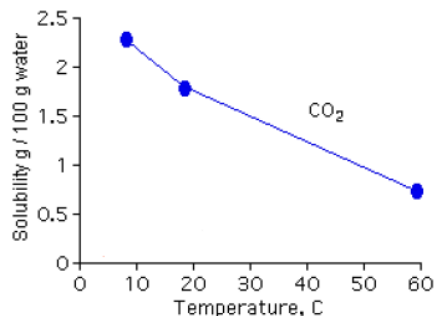
1) Exothermic (-) Heat of solution (ΔH_{soln}):	Decreasing the temperature increase the solubility.
2) Endothermic (+) Heat of solution (ΔH_{soln}):	Increasing the temperature increase the solubility.

When ammonium chloride, NH₄Cl dissolves in water the solution becomes colder

When sodium hydroxide, NaOH dissolves in water the solution heats up.



Gaseous solutes in liquids: Decreasing the temperature always increase the solubility.



**Effect of pressure on solubility
Liquid or solid solutes in liquids**

Pressure has no effect on solubility of liquid or solid solutes.

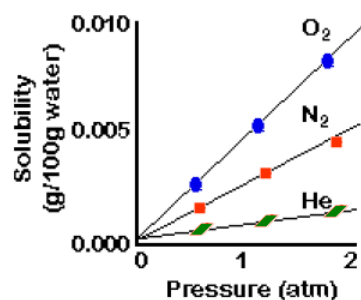
Gaseous solutes in liquids

Pressure always increases the solubility of dissolving gases in liquids.

Gas solubility is directly proportional to the partial pressure of the gas above the solution is known as

Henry's Law

$S_g = K_H P_g$; where S_g is the solubility of the gas, K_H is Henry's the Law constant, P_g is partial pressure of gas.



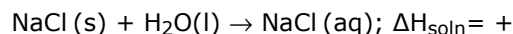
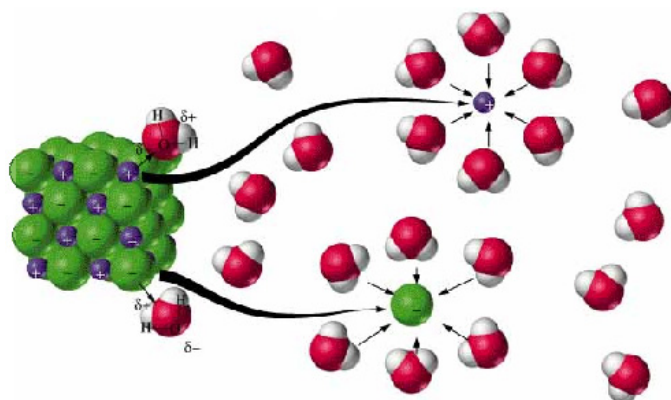
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Summary of observations during solution process

- Materials with similar polarity are soluble in each other. Dissimilar ones are not.
- Polar substances with similar forces are likely to be soluble in each other.
- Non-polar solutes dissolve in non-polar solvents.
- Stronger solute-solvent attractions favor solubility, stronger solute-solute or solvent-solvent attractions reduce solubility.

8.3 Solution Formation

Interactions between solute and solvent molecules are known as solvation. When the solvent is water, the interactions are known as **hydration**. Sodium chloride dissolves in water because the water molecules have a sufficient attraction for the Na⁺ and Cl⁻ ions to overcome the attraction of these two ions for one another in the ionic solid. To form an aqueous solution of NaCl, water molecules must also separate from one another to form spaces in the solvent that will be occupied by the Na⁺ and Cl⁻ ions.



The formation of a solution can be either exothermic or endothermic. For example, when sodium hydroxide, NaOH, is added to water, the resultant solution gets quite warm:

Properties of Solutions

The majorities of chemical reactions, and virtually all important organic and biochemical processes, occur in solution. A solution is composed of one or more solutes dissolved in a solvent. In aqueous solutions, the solvent is water.

Transparent to Light

Liquid solutions are clear and transparent with no visible particles of solute. They may be colored or colorless, depending on the properties of the solute and solvent.

Electrical Conductivity of solutions

In solutions of **electrolytes**, the solutes are ionic compounds that have dissociated; the solution conducts electricity. Solutions of **nonelectrolytes** are nonconducting.

Solubility and Equilibrium

Insoluble salts which barely dissolve and always have solid precipitate present. A typical example is



This describes AgCl(s) dissolved in solution to silver (Ag^+) and chloride (Cl^-) ions the aqueous phase (aq), in equilibrium with the sparingly soluble solid (s) salt AgCl. This is an example of solubility equilibrium.

8.4 Solubility Rules

Solubility of ionic compound in water

There are a series of guidelines in your book. On the next page 197, I provide you with another presentation of these same rules:

Solubility Rules

1. All compounds containing Group IA: Na^+ , K^+ etc.
2. All compounds containing ammonium- (NH_4^+), nitrates- (NO_3^-) and acetate- ($\text{C}_2\text{H}_3\text{O}_2^-$) ions are soluble in water.
3. Most halides, chlorides (Cl^-), bromides (Br^-), and iodides (I^-) are soluble except: silver, lead and mercury.
4. All sulfates (SO_4^{2-}) are soluble except: Ca^{2+} , Sr^{2+} , Ba^{2+} , and Pb^{2+} .
5. Most carbonates (CO_3^{2-}), phosphates (PO_4^{3-}), sulfides (S^{2-}), and hydroxides (OH^-) are insoluble in water. Important exceptions are those of Group IA: Na^+ , K^+ , etc. and NH_4^+ .

Which of the following salts (ionic compounds) is soluble/insoluble in water?

- a) NaCl b) Li_2CO_3 c) AgCl d) PbBr_2 e) NH_4NO_3 f) $\text{Ca}(\text{NO}_3)_2$
g) CaSO_4 h) CaCO_3 i) $\text{Mg}_3(\text{PO}_4)_2$ j) MnO_2 k) $\text{Al}(\text{OH})_3$ l) BaSO_4
m) $\text{CH}_3\text{CO}_2\text{Na}$

Answer

Compound	Solubility	Reason	Compound	Solubility	Reason
NaCl	Soluble	Group IA	CaCO_3	Insoluble	CO_3^{2-}
Li_2CO_3	Soluble	Group IA	$\text{Mg}_3(\text{PO}_4)_2$	Insoluble	PO_4^{3-}
PbBr_2	Insoluble	Ag, Hg, Pb halide	MnO_2	Insoluble	Metal oxide
NH_4NO_3	Soluble	NH_4^+ and NO_3^-	$\text{Al}(\text{OH})_3$	Insoluble	Metal hydroxide
$\text{Ca}(\text{NO}_3)_2$	Soluble	NO_3^-	BaSO_4	Insoluble	Sulfate of Ba
CaSO_4	Insoluble	Sulfate of Ca	$\text{CH}_3\text{CO}_2\text{Na}$	Soluble	Na^+ and $\text{C}_2\text{H}_3\text{O}_2^-$

8.5 Solution Concentration Units

The amount of solute dissolved in a given amount of solution is the solution concentration. The more widely used percentage-based concentration units are:

Concentration of Solutions: Percentage

The amount of solute dissolved in a given amount of solution is the solution concentration. The more widely used percentage-based concentration units are:

- Mass/mass percent.
- Volume/volume percent
- Mass/volume percent.

a) Mass/mass percent- (m/m%).

$$\% (m/m) = \frac{\text{grams of solute}}{\text{grams of solution}} \times 100\%$$

What is the %(m/m) concentration of I_2 in a solution prepared by mixing 50.0 g solid I_2 with 80.0 g of liquid Br_2 .

$$\% (m/m) = \frac{\text{grams of solute}}{\text{grams of solution}} \times 100\%$$

$$\% (m/m) = \frac{50.0 \text{ g}}{50.0 \text{ g } \text{I}_2 + 80.0 \text{ g } \text{Br}_2} \times 100\% = 38.5 \%$$

b) Volume/volume percent- (v/v%).

Volume/volume percent.

$$\% (V/V) = \frac{\text{mL solute}}{\text{mL solution}} \times 100\%$$

What is the %(V/V) concentration of a aqueous acetic acid solution prepared by adding 20.0 mL of acetic acid to water to a final volume of 2.50 L?

$$\% (V/V) = \frac{20.0 \text{ mL acetic acid}}{\left(2.50 \text{ L soln} \times \frac{10^3 \text{ mL soln}}{1 \text{ L soln}}\right)} \times 100\%$$

$$\% (V/V) = 0.800 \% \text{ acetic acid}$$

c) Mass/volume percent- (m/v%).

Calculating weight/volume percent we use following equation:

$$\% (m/v) = \frac{\text{grams of solute}}{\text{milliliters of solution}} \times 100\%$$

What is the % (m/v) of a 476 L Ar in CCl₄ solution prepared by dissolving 50.0 g.

$$\% (m/v) = \frac{\text{grams of solute}}{\text{milliliters of solution}} \times 100\%$$

$$\% (m/v) = \frac{50.0 \text{ g Ar}}{\left(476 \text{ L solution} \times \frac{10^3 \text{ mL solution}}{1 \text{ L solution}}\right)} \times 100\%$$

$$\% (m/v) = 10.5 \times 10^{-3} \%$$

How many grams of NaOH is required prepare 2.0 L of 1.0 % (m/v) solution?

Solving for mass,

$$\text{g solute} = \frac{[\% (m/v) (\text{mL solution})]}{100\%}$$

Substituting,

$$\text{g NaOH} = \frac{(1.0 \%) \left(2.0 \text{ L} \times \frac{10^3 \text{ mL}}{1 \text{ L}}\right)}{100\%} = 2.0 \times 10^1 \text{ g NaOH}$$

What is the volume of the solution if 10.0g of NaCl is dissolved to obtain 25 % (m/v) solution?

Solving for volume,

$$\text{mL solution} = \frac{\text{g solute}}{\% (m/v)} \times 100\%$$

Substituting,

$$\text{mL solution} = \frac{10.0 \text{ g NaCl}}{25.0 \% (m/v)} \times 100\% = 40.0 \text{ mL}$$

Concentration of Solutions: Moles per liters of solution

Molarity (M) is the number of moles of solute per liter of solution.

Molarity

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

What is the molarity of a solution prepared by dissolving 0.700 g of KCl in water to a solution with 1.00 mL solution?

$$M_{\text{KCl}} = \frac{\text{mol KCl}}{\text{L solution}}$$
$$M_{\text{KCl}} = \frac{\left(0.700 \text{ g KCl} \times \frac{1 \text{ mol KCl}}{74.55 \text{ g KCl}}\right)}{\left(1.00 \text{ mL solution} \times \frac{1 \text{ L}}{10^3 \text{ mL solution}}\right)} = 9.39 \text{ M KCl}$$

Using molarity as a conversion factor

How many grams of AgNO₃ are in 2.00 L of 0.500 M AgNO₃ solution?

$$M_{\text{AgNO}_3} = \frac{\text{mol AgNO}_3}{\text{L solution}}$$

Solving for moles of AgNO₃,

$$\text{mol AgNO}_3 = (M_{\text{AgNO}_3})(\text{L solution})$$

$$\text{mol AgNO}_3 = (0.500 \text{ M})(2.00 \text{ L}) = 1.00 \text{ mol AgNO}_3$$

Then, convert mol AgNO₃ → g AgNO₃

$$1.00 \text{ mol AgNO}_3 \times \frac{169.9 \text{ g AgNO}_3}{1 \text{ mol AgNO}_3} = 1.70 \times 10^2 \text{ g AgNO}_3$$

8.6 Dilution

Concentrated solutions can be mixed with solvent to make weaker or dilute solutions. This is the kind of thing people do everyday with consumer products like fruit juice. Some concentrated solutions are used as "stock" solutions. Weaker solutions are typically used but the concentrated solutions require less storage space. In recent years accidents have occurred in the health care professions when dilutions were done incorrectly. Some of these errors have resulted in deaths or serious injuries. A number of health care facilities have abandoned the practice of "diluting" stock solutions because dilution

instructions were too hard to follow. They do not want to take the risks associated with errors in preparing diluted solutions.

$$(M_i)(V_i) = (M_f)(V_f)$$

M_i = initial molarity; V_i = initial volume; M_f = final molarity; V_f = final volume

What would be the volume of a 0.002 M solution prepared starting with 50.0 mL of 0.400 M solution?

$$M_i = 0.400 \text{ M}$$

$$V_i = 50.0 \text{ mL} \times (1\text{L}/10^3 \text{ mL}) = 5.00 \times 10^{-2} \text{ L}$$

$$M_f = 0.200 \text{ M}$$

$$V_f = ? \text{ L}$$

Solving for the final volume of 0.200 M sugar, V_f ,

$$V_f = \frac{(M_i)(V_i)}{(M_f)} = \frac{(0.400 \text{ M})(5.00 \times 10^{-2} \text{ L})}{(0.200 \text{ M})} = 0.1 \text{ L} = 100 \text{ mL}$$

How many mL of 2.00 M solution of HNO₃ are required with water to make a 250 mL of 1.50 M nitric acid solution?

This is a dilution problem. You have to add water to certain volume of HNO₃ and make a 250 mL solution with the molarity, 1.50 M HNO₃.

Equation is: $M_i V_i = M_f V_f$ M_i = initial molarity

V_i = initial volume

$$M_i = 2.00$$

M_f = final molarity

$$V_i = ?$$

V_f = final volume

$$M_f = 1.50$$

$$V_f = \frac{250}{1000} = 0.25 \text{ Liters}$$

$$V_i = \frac{M_f V_f}{M_i} = \frac{1.50 \times 0.25}{2.00} = 0.1875 \text{ L}$$

0.1875 L = 187.5 mL of 2.00 M solution.

Chemistry at a Glance: Solutions

8.7 Colloidal Dispersions

Depending on the particle sizes of solutes in solvents could make three types of homogenous mixtures.

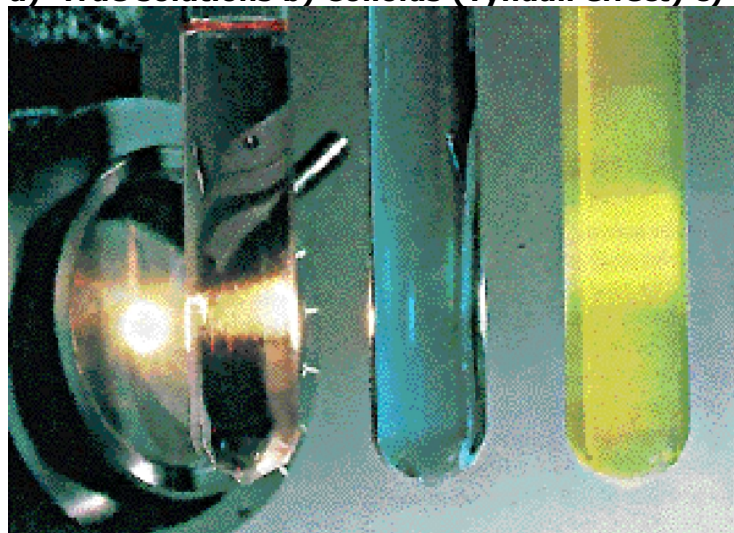
True solutions	Colloids	Suspensions
<ul style="list-style-type: none"> Particle sizes between 0.1 to 2 nanometers Light can usually pass through the solution and transparent. Solute particles do not 	<ul style="list-style-type: none"> Particle sizes between 2 to 1000 nanometers The particles are large enough to scatter light and solutions are opaque. Solute particles still 	<ul style="list-style-type: none"> Particle sizes greater than 1000 nanometer. The particles are large enough to scatter light and solutions are opaque. Solute particles are larger

separate and settle at the bottom on standing	smaller and do not separate and settle at the bottom on standing <ul style="list-style-type: none"> Shows Tyndall Effect. 	and separate and settle at the bottom on standing
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Tyndall effect

The scattering of light by colloidal particles is called **Tyndall effect** which has been used to distinguish between true solutions and colloids. This effect will increase with increasing solute particle diameter of a colloid. For example, milk is opaque because of the higher diameters of the solute particles such as proteins in the milk.

a) True solutions b) Colloids (Tyndall effect) c) Suspensions.



Surfactants

Surfactants constitute the most important group of detergent components. Generally, these are water-soluble surface-active agents comprised of a hydrophobic portion, usually a long alkyl chain, attached to hydrophilic or water solubility enhancing functional groups. The hydrophilic end, which is either polar or ionic, dissolves readily in water. The hydrophobic, or non-polar, end, however, does not dissolve in water. In fact, the hydrophobic, or "water-hating" end will move as far away from water as possible. This phenomenon is called the hydrophobic effect. Because one end of a surfactant resists water and the other end embraces it, surfactants have very unique characteristics.

8.8 Colligative Properties of Solutions

Colligative properties are properties that depend on concentration rather than the identity of the solute. Four colligative properties are: **vapor pressure lowering,**

freezing-point depression, boiling-point elevation and **osmotic pressure**. Each has a number of very practical ramifications.

Four colligative solution properties are:

- **Vapor pressure lowering.** When a solute is added to a solvent, the vapor pressure of the solvent decreases in proportion to the concentration of the solute.
- **Freezing-point depression.** When a nonvolatile solute is added to a solvent, the freezing-point of the resulting solution decreases.
- **Boiling-point elevation.** When a nonvolatile solute is added to a solvent, the boiling point of the resulting solution increases.
- **Osmosis.** Osmosis is the movement of solvent from a dilute solution to a more concentrated solution through a semipermeable membrane.

8.9 Osmosis and Osmotic Pressure

Osmosis is the net movement of water across a selectively (semi) permeable membrane driven by a difference in solute concentrations on the two sides of the membrane. A semi permeable membrane is one that allows unrestricted passage of water, but not solute molecules or ions.

Osmotic pressure

Osmosis may be opposed by increasing the pressure in the region of high solute concentration with respect to that in the low solute concentration region in process called **reverse osmosis**. The force per unit area, or pressure, required to prevent the passage of water through a selectively-permeable membrane and into a solution of greater concentration is the osmotic pressure of the solution. Osmotic pressure is a colligative property, meaning that the property depends on the concentration of the solute but not on its identity.

Osmolarity (osmol), iM

Osmolarity is a term that has been used to calculate osmotic pressure of solutions that produce more particles or ions in the solution than the molarity would indicate. For example 1M NaCl would have 2 M of ions since NaCl would break up in solution to Na⁺ and Cl⁻ ions. For non-dissociating solutes like glucose molarity and osmolarity is the same.

For dissociating solutes, osmolarity is related to the molarity of the solution multiplied by a factor (i) to obtain the number of particles of particles in the solution. Osmotic pressure (π) is calculated from the equation:

$$\pi = i MRT.$$

i = number of particles produced per formula unit.

M = osmolarity of the solution.

T = Kelvin temperature of the solution.

Hypertonic solutions cause the outflow of water from a cell, resulting in cell collapse.
Hypotonic solutions cause water to flow into the cell, resulting in hemolysis (cell rupture).
Isotonic solutions have identical osmotic pressures.

Calculate the osmolarity of a 5.0×10^{-3} M glucose solution.

$C_6H_{12}O_6$ is a non-electrolyte; it does not dissociate.

$i = 1$ and since $C_6H_{12}O_6$ is a non-electrolyte; it does not dissociate.

Osmolarity is = $M = a 5.0 \times 10^{-3} M$

What is the osmotic pressure in atm of a 5.0×10^{-3} M glucose solution?

We found that the solution was 5.0×10^{-3} osmol; substituting this value for M (M represents osmolarity) in the expression:

$$\pi = MRT$$

$$\pi = 5.0 \times 10^{-3} \frac{\text{mol particles}}{\text{L solution}} \times \frac{0.0821 \text{ L} \cdot \text{atm}}{\text{K} \cdot \text{mol}} \times (25 + 273\text{K})$$

$$\pi = 0.12 \text{ atm}$$

Using our definition of osmotic pressure, $\pi = MRT$

M must be represented as osmolarity.

$$M = \frac{0.50 \text{ mol sucrose}}{\text{L}} \times \frac{1 \text{ mol particles}}{1 \text{ mol sucrose}} = \frac{0.50 \text{ mol particles}}{\text{L}}$$

Now, substituting in our osmotic pressure equation:

$$\pi = (0.50 \text{ mol particles/L}) \times (0.0821 \text{ Latm/Kmol particles}) \times 298 \text{ K} = 12 \text{ atm}$$

(Note, two significant figures for the answer; the sucrose concentration has two significant figures)

What is the osmolarity of 0.1 M Na_2SO_4 at 25°C?

Na_2SO_4 would break up in solution to $2Na^+$ and SO_4^{2-} ions which adds up to 3 ions and $i = 3$.
Osmolarity = $i \times M = 3 \times 0.1 = 0.3 M$

What is the osmotic pressure of 0.1 M Na_2SO_4 at 25°C?

$$\pi = i MRT, \quad i=3, \quad M= 0.1, \quad R = 0.0821 \text{ and } T = 298.15 \text{ K}$$

$$\pi = 3 \times 0.1 \times 0.0821 \times 298.15 = 7.34 \text{ atm}$$

Hypotonic, Hypertonic and Isotonic Solutions

When thinking about osmosis, we are always comparing solute concentrations between two solutions, and some standard terminology is commonly used to describe these differences:

Isotonic: The solutions being compared have equal concentration of solutes.

Hypertonic: The solution with the higher concentration of solutes.

Hypotonic: The solution with the lower concentration of solutes.

Chemistry at a Glance: Summary of Colligative Properties

8.10 Dialysis

Dialysis works on principles similar to osmosis through the diffusion of certain solutes and ultrafiltration of fluid across a semi-permeable membrane. Blood flows by one side of a semi-permeable membrane, and a fluid that flows on the opposite side collecting the smaller solutes and extra fluid passing through the membrane. The blood flows in one direction and the fluids (dialysate) in the dialysis machine flows in the opposite. The counter-current flow of the blood and fluid maximizes the concentration gradient of solutes between the blood and dialysate, which helps to remove more **urea** and **creatinine** from the blood. The concentrations of solutes (for example potassium, phosphorus, and urea) are undesirably high in the blood in patients with kidney failure, but low or absent in the dialysis solution and constant replacement of the dialysate ensures that the concentration of undesired solutes is kept low on this side of the membrane. The dialysis solution has levels of minerals like potassium and calcium that are similar to their natural concentration in healthy blood. For another solute, bicarbonate, dialysis solution level is set at a slightly higher level than in normal blood, to encourage diffusion of bicarbonate into the blood, to act as a pH buffer to neutralize the metabolic acidosis that is often present in these patients. The levels of the components of dialysate are typically prescribed by a nephrologists according to the needs of the individual patient.

Chemical Connections: Factors Affecting Gas Solubility; Solubility of Vitamins; Controlled-Release Drugs: Regulating Concentration, Rate, and Location of Release; The Artificial Kidney: A Hemodialysis Machine