Chapter 11 – Solutions and Colloids

Before we get started with new material, we should begin by reviewing some material you covered early last semester. Recall that a <u>solution</u> is any homogeneous mixture and that the component present in largest amount is called the <u>solvent</u>, while all other components of the solution are called <u>solutes</u>. There is a basic review of solutions in Sections 1.2 (p. 18) and 3.3 & 3.4 (p. 151 - 164) that you may find useful. Solutions exist in all three phases (Table 11.1).

11.1   <u>The Dissolution Process</u>

When a crystalline solid dissolves in water, molecules or ions leave the lattice and are eventually dispersed evenly throughout the solvent. Since lattice energies for ionic compounds are so large (Chapter 7.5), why does this occur instead of the solid just sitting there and not dissolving?

The answer lies in the intermolecular forces discussed in Chapter 10. There are three interactions we must consider. The ions or molecules in a lattice are collectively attracted to one another. Likewise, the molecules in a sample of pure solvent are attracted to one another.

When an ion or molecule breaks away from the crystal, solvent molecules surround it. All of the solvent molecules are attracted to it to some degree (London dispersion forces at a minimum.). This process is called <u>solvation</u>. (When the solvent is water, the process is called <u>hydration</u>.) The solid will dissolve if the energy released from the intermolecular attractions is greater than the lattice energy of the crystal from which it came and the attraction for the solvent molecules for each other.

If the heat of solution (ΔH_{soln}, heat released as a consequence of solvation) is sufficiently large, the solution will become warm. An example of this from around the house is the use of crystal Drano®. If you have ever used this product on a sink, you may have noticed the drainpipe gets very hot. The heat is the result of dissolving the Drano® and is what removes grease clogs (the grease melts and is flushed away).

If the heat of solvation is positive, but not very large, the solid may still dissolve. This can

happen for two reasons. First, thermal energy from the surroundings (the solvent) may overcome the energy of attraction in the lattice. The second reason arises because, from one energy perspective (called entropy), it is always slightly favorable for the ions to disperse into solution. This term can be very important when ΔH is small, which is frequently the case when examining the dissolving of solids. We'll talk about entropy in greater detail in Chapter 19. When such salts dissolve, the solution becomes cooler. Instant cold packs work this way. A salt packet (frequently ammonium nitrate because it dissolves rapidly) is placed in water and when you knead the pack, the salt packet breaks and dissolves.

Ignore the book's brief discussion of gaseous solutions. We'll discuss that in Chapter 16.

11.2   Electrolytes

All solutes that cause aqueous solutions to conduct electricity are called underline electrolytes. Nonelectrolyte, aqueous solutions do not conduct electricity. Compounds induce water to conduct electricity by generating ions. They can do so by one of two pathways. Dissociation occurs when an ionic compound breaks up into ions in solution. Here the ions already existed, the water merely acts to separate them. Ionization occurs when a molecular compound breaks up into ions. In this case, the water causes a chemical change that results in the molecule splitting into parts carrying charges. These terms mean different things and you should know the difference between them.

Molecular compounds such as acetic acid (CH_3COOH) and hydrogen chloride (HCl) are electrolytes. Hydrogen chloride behaves much like ionic compounds in that it breaks up completely into ions. i.e. It completely ionizes. In contrast, only some acetic acid molecules separate into ions. Thus, acetic acid only partially ionizes.

$$CH_3COOH_{(aq)} \rightleftharpoons CH_3COO^-_{(aq)} + H^+_{(aq)}$$

The double arrow means "in equilibrium with." Here, the reaction moves in both directions simultaneously and the amount of each species doesn't change over time. The total concentration of each species remains constant. You'll spend a lot of time in Chapter 13 learning about this concept.

Electrolytes can be divided into 2 categories. <u>Strong electrolytes</u> dissociate or ionize completely (100%) when dissolved in water. All ionic compounds (e.g. NaCl, $Ca(NO_3)_2$) and some molecular compounds (e.g. HCl, H_2SO_4) are strong electrolytes. <u>Weak electrolytes</u> ionize less than 100% when dissolved in water. At first, this may seem like a poor definition. After all, wouldn't it be hard to differentiate between 99% and 100% ionization? It turns out that nearly all weak electrolytes ionize less than 10% and almost none over 50% under typical conditions. <u>This results in no close calls.</u> While few types of molecular compounds ionize, most of the ones that do ionize are weak electrolytes (e.g. CH_3COOH (HOAc), NH_3).

It is important to remember that solubility and ion production are not related. The weak electrolyte acetic acid mixes completely with water, while the strong electrolyte magnesium hydroxide is only slightly soluble. In fact, this is why it can be used as an antacid (Milk of Magnesia). If it were very soluble, its solution would damage the tissue in your mouth and throat. Since it is insoluble, it goes straight to your stomach where it reacts with stomach acid.

11.3   <u>Solubility</u>

All *solid* solutes have a maximum amount that will dissolve in a given solvent. The <u>solubility</u> of a substance is the maximum amount of a solute that will dissolve in a solvent at a given temperature. Solubility is both solvent and temperature dependent and is usually listed in units of mass per volume (e.g. g/mL, g/L, etc.). For example, much more table salt (NaCl) will dissolve in water than in rubbing alcohol. Likewise, more sodium chloride will dissolve in hot water than cold. It is therefore necessary to provide both the solvent and temperature when discussing solubility.

A <u>saturated</u> solution contains the maximum amount of a solute at a particular temperature. *Any amount* of solute less than this amount yields an <u>unsaturated</u> solution. Another common misconception is confusing concentrated/dilute with saturated/unsaturated. The terms are not related.

When dissolving a solute in a solvent, the solute begins to enter solution. As the amount

increases, some of the dissolved solute will crystallize back to the solid phase. In an unsaturated solution, the rate of dissolving will always exceed the rate of crystallization. The result is the net conversion of solid solute into dissolved solute. As the solution becomes more concentrated, the rate of crystallization will accelerate, until, at saturation, the rate of dissolving equals the rate of crystallization. Is this reasonable?

Imagine a beaker of stirring water and a large crystal of some salt, MX. Drop the crystal into the water. The moment before the crystal hits the water, there are zero M^+ or X^- ions in solution. Thus, _every_ particle that contacts the crystal surface at $t = 0$ is a water molecule. Some of these water molecules will have sufficient energy to dislodge and solvate an ion (i.e. the ion dissolves). Let's look at the beaker one second later. Most of the crystal will still be there, but some ions will have dissolved. At this point, most of the particles colliding with the crystal will still be water molecules, but occasionally an M^+ or X^- ion will bump into it. When this happens one of two things will occur. If the ion is zipping along with a relatively large amount of energy, it will simply bounce off the crystal surface. On the other hand, if it is plodding along with little energy, it will re-adhere to the crystal (crystallization) when the incoming ion collides with a ion of the opposite charge.

Clearly, the rate at which ions collide with the crystal surface (and hence re-adhere) increases as more ions dissolve. Also, as time progresses, the rate at which the crystal dissolves will decrease. (There are two reasons for this, can you figure out what they are?) Eventually, for every salt, there will be a point where the rates become equal and the solution becomes saturated in the salt. As you know, not all salts dissolve to the same extent (have the same solubility). For example, table salt (NaCl) is much more soluble than limestone ($CaCO_3$). Why is this? The lattice energy of NaCl is much smaller than that of $CaCO_3$, so it is much more difficult to bring the ions in calcium carbonate into solution. Lattice energy was discussed in Chapter 7.5 (p. 371).

The solubilities of solids generally increase with increasing temperature. This is because energy is required to break-up the crystal lattice. Most of that energy is provided by solvation, but some comes from ambient thermal energy. At higher temperatures, there is more heat available

for separating the ions from the lattice. Table 11.6 (p. 617) shows this general trend, but also the wide range of variability of the effects of temperature on the solubility of solids.

Under the right conditions, an unexpected result sometimes occurs for a dissolving solute. When more than the maximum amount of solute is dissolved in a solvent, the solution is called supersaturated. This may at first strike you as an oxymoron, but it can happen. Remember that a crystal is an ordered array. Now imagine saturating a solvent with a solute and then warming the solution. If more solute is dissolved and the solution then cools back to the original temperature, the solute should crystallize from solution. In fact, this usually happens. But a solution is inherently disordered, so if the conditions are appropriately manipulated, crystallization can be delayed. Supersaturated solutions are inherently unstable and will eventually crystallize on their own. They can be pushed into crystallizing by dropping in a crystal of solute, stirring the container, or scratching a wall inside the flask. In the first case, a template is provided onto which ions can attach. In the other cases, sites are provided where ordering may more easily occur.

We find that the crystallization rate will exceed the dissolving rate when the solution becomes supersaturated. This occurs when the temperature of a saturated solution drops or when some of the solvent evaporates.

Solutions of Gases in Liquids

While solutions of gases in liquids aren't commonly encountered in the lab, they have obvious, significant biological importance.

The solubilities of gases generally decrease with increasing temperature. Gases are typically small, low mass molecules. Most are nonpolar or only slightly polar. As a result they move rapidly at ambient temperature. In addition, their small size means they will exhibit negligible London dispersion forces. Thus, as temperature increases, the molecules accelerate (and more easily overcome the intermolecular forces operating in the solution). Once a gas molecule breaks free from the solvent, it is typically gone to the atmosphere.

As you might guess, when the pressure of a gas above a liquid increases, more of the gas dissolves in the liquid. (e.g. Consider a sealed container half-filled with water and oxygen gas

above it. Now put in more gas. Does the amount of oxygen in the water change? If so, how?)

The amount of gas that dissolves is directly proportional to pressure. That is, if the pressure doubles, the amount of dissolved gas doubles. The actual solubility depends on the gas itself. The result is Henry's law:

$$C_g = kP_g$$

where C_g is the solubility of the gas, k is the Henry's law constant which measures the inherent solubility of the gas in that solvent (and is temperature dependent), and P_g is the partial pressure of the gas.

You can see this effect whenever you open a soda. Carbonated soft drinks are packed under pressure. When their containers are opened, the pressure above the liquid immediately drops to atmospheric pressure. When this happens much of the dissolved CO_2 is released from solution and bubbles out. Not surprisingly, pressure rarely has a significant effect on the solubility of either solids or liquids.

Solutions of Liquids in Liquids

When two liquids mix completely in any ratio, they are said to be miscible. If they do not mix at all, they are immiscible. *N.B.*: The words "soluble" and "dissolve" are not usually used in connection to the combination of two liquids. Thus, 5 g of liquid *A* will mix into solvent *B*.

Alcohols (organic compounds with -OH groups) provide an example of the progression from miscible to nearly immiscible with water for a group of compounds. With the addition of each -CH_2- group the molecule becomes less polar and less soluble in water. There is a certain subjectivity to the term immiscible. The first three alcohols CH_3OH, CH_3CH_2OH, and $CH_3CH_2CH_2OH$ are all miscible with water. Since $CH_3(CH_2)_3OH$ is only about 0.1% soluble by mass, most people would consider it immiscible.

11.4   Colligative properties

In CHM 211, you learned about 3 concentration units in Sections 3.3 and 3.4: molarity, parts per million, and mass percent. Of these, molarity is the only one you've worked with. There are

two more units of concentration we need to discuss. Both are specialized, but necessary in the places they are used. The first is molality and the other mole fraction, both are defined below.

molarity: $\quad\text{molarity}_x = \left(\dfrac{\text{moles of component x}}{\text{volume of solution (in L)}}\right) = M_x$

parts per million: $\quad\text{ppm}_x = \left(\dfrac{\text{mass of component x}}{\text{total mass of solution}}\right) \times 10^6$

mass percent: $\quad\text{mass \% of component x} = \left(\dfrac{\text{mass of component x}}{\text{total mass of solution}}\right) \times 100\%$

molality: $\quad\text{molality}_x = \left(\dfrac{\text{moles of component x}}{\text{mass of solvent (in kg)}}\right) = m_x$

mole fraction: $\quad\text{mol fraction}_x = \left(\dfrac{\text{moles of component x}}{\text{total moles}}\right) = \chi_x$

The term *molality* is unfortunate because of its similar spelling to *molarity*. Nonetheless, it is an important unit because neither the mass nor the number of moles of a substance changes with temperature. Density is temperature dependent (as temperature increases, most substances expand), so molarity is temperature dependent. When studies are done at different temperatures, the use of molality allows convenient data comparison without having to adjust for density changes.

Mole fraction is a counting ratio. Thus, the numerator is the number of solute particles, while the denominator is the number of particles of everything in solution.

In your experience, until now, physical properties generally depended on the species under consideration. In other words, two substances melt at different temperatures because of differing strengths of intermolecular forces or they have different densities because although they are about the same size and one has a much larger mass than the other. **There is a class of solution properties, called <u>colligative properties</u>, that depends only on the number of solute particles present.** As we will see, there are some caveats to the "only on the number" rule, but it is substantially true that the number of solute particles determines the magnitude of the property. We will examine four colligative properties.

<u>Vapor Pressure Lowering</u>

You will find a discussion of vapor pressure in Chapter 10.3 (p. 541), if you need a review.

We begin with a couple of definitions, one of which has developed a meaning in common usage that comes from an overgeneralization. Neither is in your book, but both are commonplace in chemistry. A substance is said to be <u>volatile</u> if it exhibits a measurable vapor pressure under a given set of conditions. If it does not have a measurable vapor pressure the substance is <u>nonvolatile</u>. There are a few things worth noting here. First, the conditions matter. Water is volatile at 75 ºC, but not at -20 ºC. Next, when a chemist uses the term volatile, it usually means *significant* vapor pressure, unless otherwise stated. Finally, when listening to a newscast, you may hear of an explosion or fire resulting from a volatile substance. It is true that a liquid substance must be volatile to burn, but not all volatile substances (e.g. water vapor) burn. FYI, liquids don't burn, but gases do.

When a nonvolatile solute is added to a solvent, the vapor pressure of the solvent always drops. Remember that a vapor is a gas in the presence of the same material as a liquid (or solid).

A natural question that follows from the introductory paragraph is "Why does the choice of solute <u>not</u> matter?" To answer this question, we must first assume the solute is evenly distributed in the solvent. If so, imagine the surface of the liquid. In pure solvent, a solvent molecule occupies every location on the surface (Figure 1). Thus, all sites represent a place where a solvent molecule may leave the liquid phase and enter the gas phase. Now when a solute molecule occupies a site, no solvent molecule may enter the gas phase there. Thus, solute molecules act like a blanket covering the solvent and slowing the rate of evaporation (Figure 2).
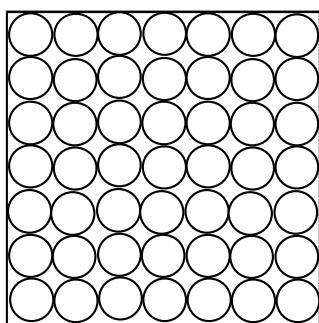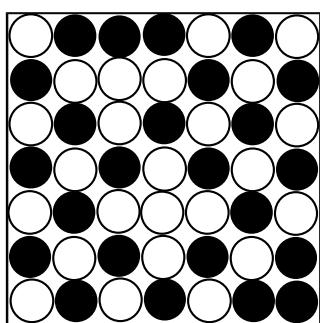
pure solvent   solution
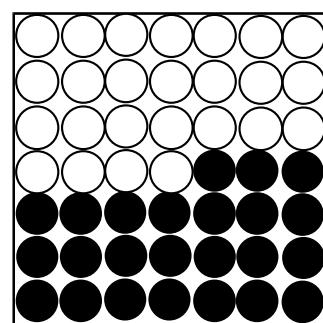
Figure 1   Figure 2

○ = solvent, ● = solute

In both Figure 2 drawings just under half of the molecules are solute molecules. Instead of random distribution (left side), if the solute molecules were collected together on the surface (right side), one can easily see how almost half of the surface is covered so that solvent can't evaporate in that region (rather like a blanket). In both cases, any solvent molecule in the gas phase may reenter the liquid phase by striking the surface of the liquid at any location. Thus, the rate of reentry is not slowed for the solution. In combination with the lower rate of evaporation of the solution, the result is a lower vapor pressure for the solution because there are fewer solvent molecules in the gas phase for it.

A simple experiment that you can do at home to demonstrate the lowering of vapor pressure follows. Pick two small glasses, preferably ones with clear, colorless glass. Shot glasses or something comparable work most quickly. They don't have to be the exactly same size, but should be similar. Prepare a saturated salt-water solution. (Adding salt to warm water and letting it cool is easiest.) Fill one of the glasses almost full with it. Fill a second glass ¾ full with distilled water, if you have it; tap water if you don't. Put a grease pencil mark at the liquid lines if you wish. Now put both in a container you can seal. Again, it's best if the walls of the container are transparent, so you can watch the fluid levels. Put the container in sunlight if you can to help speed the process and check it daily. If you can't see in the container, then wait about a week to check on it. You will find that the pure water level will drop and the saltwater level will rise until it actually spills out over the lip of the glass it is in. (If you have identical glasses and set the liquid heights exactly the same, the salt water will evaporate more slowly if you set them in the open.)

It is important to consider a second feature of the solute. Salts, strong acids, and strong bases break up into ions, as do some molecules (e.g. HCl). For species that break-up completely into ions, one must determine the total concentration of all ions. (i.e. a 0.5 M NaCl solution is 1.0 M in ions). Since the extent to which weak acids and bases dissociate depends on their concentrations, and concentrations change as solvent evaporates, we'll ignore them.

Vapor pressure reduction can be expressed quantitatively through <u>Raoult's law</u>.

$$P_A = \chi_A P_A{}^\circ$$

where P_A is the vapor pressure of the solvent above the solution, $P_A{}^\circ$ is the vapor pressure of the pure solvent, and χ_A is the mole fraction of the solvent in the solution.

Solutions that obey Raoult's law are called <u>ideal solutions</u>; much like the ideal gas law would describe an ideal gas. Solutions are most nearly ideal when the solute concentration is low, the solute and solvent particles are of similar sizes, and intermolecular forces between all species are of similar magnitude. Thus, the deviations from ideality (alluded to in the first paragraph of this section) are what you might expect. We'll discuss the deviations now.

Since all solute molecules will differ at least a little from the solvent, increasing the number of solute molecules will increase deviations. Solutes that are much larger than solvent molecules will block multiple sites making them seem to have a higher concentration. For example, consider $CH_3(CH_2)_{23}CH_3$ dissolved in $CH_3(CH_2)_3CH_3$. The solute is about 5 times longer than the solvent and presumably would block much more of the surface than a solvent molecule. It would yield results suggesting the solute concentration was higher than it actually was. Finally, if the solute associates more strongly with the solvent than the solvent does with itself, then it will be more difficult for a solvent molecule adjacent to a solute molecule to leave the surface and enter the gas phase. Also, a solvent molecule striking the liquid phase from the gas phase would be more likely to stick if it hits a solute molecule. Each would lower vapor pressure more than expected. (A similar reverse argument can be made for solutes that bind much less strongly to solvent.)

<u>Boiling Point Elevation and Melting Point Depression</u>

Whenever a nonvolatile solute is added to a solvent, the boiling point of the solution is higher than the boiling point of the pure solvent and the melting point of the solution is lower than the melting point of the pure solvent. The rationale for this derives from the change in vapor pressure.

The explanation for the boiling point elevation goes as follows. At the boiling point, the vapor pressure of the liquid equals the external pressure. (This is the definition of boiling.) Now adding

a nonvolatile solute reduces the vapor pressure of a solution containing it. The temperature of the solution must be raised above that of the pure solvent to get the same vapor pressure. Thus, the solution boils at a higher temperature than the pure solvent.

Melting point depressions occur for essentially the same reason, let's use an aqueous solution as an example. Imagine an ice cube immersed in water, both at 0 ºC. If heat is being removed from solution the cube grows; if heat is being added the cube shrinks. As long as ice is present, the temperature remains constant. If there is no heat change, the system is at equilibrium. What happens if we add a little salt to the ice water? The rate at which water molecules break free of the ice cube doesn't change, but now some of the sites adjacent to the ice are occupied by sodium or chloride ions, so the rate at which water molecules re-adhere to the ice slows. Lowering the temperature slows the rate of loss of water molecules from the ice and increases the rate of adding water molecules to the ice. When the rates become equal, the solution begins to freeze.

The equations for the processes are very similar:

$$\Delta T_b = K_b m \qquad\qquad \Delta T_f = K_f m$$

where ΔT_b is the boiling point elevation, K_b is the boiling-point-elevation constant, ΔT_f is the freezing point depression, K_f is the freezing-point-depression constant, and m is the **molal** solute concentration. *NB*: K_b almost never equals K_f. Also, remember, as with Raoult's law, one must take into account the number of ions released when salts dissolve.

Two practical areas for these properties are salting icy roads and using antifreeze in your car. For the former, the salt lowers the temperature at which the water freezes, so the ice melts making roads less hazardous in winter. A downside to salt use is that the runoff then makes it difficult for anything to grown by the roadside until rain washes the salt away. (Much like ancient Rome salted the ground at Carthage to prevent that civilization from rebuilding once the Romans departed.) That is why some places use sand (which doesn't remove ice) to increase traction on the ice. Antifreeze lowers the freezing point in your radiator so it doesn't freeze in winter. It also raises the boiling point, but this is less important because it is a sealed system. One common misconception is that salt is added to water when cooking to raise the boiling point. This is

functionally not true. Try the calculation for adding a tablespoon of salt (say 15 g) in 4 liters of water and see how much the boiling point of the solution is above that of pure water. The salt is added as a flavoring agent only. (You can also test this by preparing some pasta in plain water, then some in salted water. You'll probably prefer the taste of the pasta cooked in the salt water, even if you rinse it first.)

Skip solution phase diagrams.

Osmosis

Membranes are frequently used to wall regions of space off from each other. The cell walls in biological systems are membranes. While some membranes are impenetrable, most will allow molecules of a certain size or polarity pass through. These are called semipermeable membranes. For example, cell walls allow the ready passage of small molecules such as water, nutrients, and cell waste products, but stop cell structures (e.g. the nucleus) from leaving the cell.

Chemists and biologists find it useful to use membranes that permit molecules of a certain maximum size to pass through. Consider a U-tube with such a membrane in the middle of the U. On one side is placed a certain volume of pure solvent (e.g. water) and on the other is an equal volume of a solution (say sugar water). The solvent molecules are small enough to pass through the membrane, but the solute molecules (i.e. sugar) cannot. (See Figure 24 on pp. 630).

If no membrane was present, the solutions would slowly mix until a homogeneous solution with a sugar concentration half that of the original solution was generated. The membrane prevents this because the sugar molecules cannot cross the barrier. Solvent molecules can pass in both directions through and they do. However, solvent molecules pass through the membrane at a faster rate from the pure solvent into the solution than in the reverse direction. This dilutes the solution. Osmosis is the process of net solvent movement from a region of low solute concentration to a region of higher solute concentration through a semipermeable membrane.

 Accumulation of additional solvent on the solution side causes the liquid level on that side of the U-tube to rise, while the pure solvent side drops. This results in a greater downward pressure on the solution side of the membrane than the solvent side (since there is more fluid there). The

extra pressure increases the rate at which solvent molecules on the solution side travel back across the membrane to the pure-solvent side. At the same time, the rate of solvent movement from pure solvent to solution decreases as the solution becomes increasingly dilute. Solvent will continue net movement from pure-solvent side to solution side until the pressure becomes sufficiently great to offset the tendency to dilute the solution.

The pressure required to offset osmosis is called the osmotic pressure, π. Osmotic pressure obeys a law similar to the ideal gas law

$$\pi V = nRT \qquad \text{and} \qquad \pi = MRT$$

where M is the molar concentration and R is the ideal gas constant. As you can see, osmotic pressure increases with solution concentration. Both vapor pressure reduction and osmotic pressure provide convenient (with the proper equipment) methods of determining molecular weights.

Sample problem:

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 the protein dissolved in water to yield 5.00 mL of solution. Π for the solution was found to be 1.54 torr at 25 °C. Calculate the molar mass for the protein.

The first thing you must remember is to convert to the proper units: volume in liters, pressure in atmospheres, and temperature in Kelvins.

$$\left[(1.54 \text{ torr})\left(\frac{1 \text{ atm}}{760 \text{ torr}}\right)\right](5.00 \times 10^{-3} \text{ L}) = n(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(298 \text{ K})$$

$$n = 4.141 \times 10^{-7} \text{ mol}$$

$$MW = \frac{3.50 \times 10^{-3} \text{ g}}{4.141 \times 10^{-7} \text{ mol}} = 8.45 \times 10^3 \text{ g/mol}$$

Solutions, even if they include different solutes, with equal osmotic pressures are said to be isotonic. Eye drops are almost always isotonic with the average human eye to prevent discomfort. If two solutions possess different osmotic pressures, the solution with the higher osmotic pressure is said to be hypertonic with respect to the one with the lower osmotic pressure (which is

hypotonic).

The section, Colligative Properties of Electrolytes, touches on an important, but ultimately complicated topic. When you dissolve 1 mole of sugar in a container of water, there is one mole of sugar molecules in the solution, but when you dissolve 1 mole of sodium chloride in the same volume of water, there are two moles of ions in solution. Colligative properties behave most ideally for dilute solutions and this is especially true for salts dissolved in water. Going back to the sugar *vs*. salt solutions, one would expect to have twice the change in vapor pressure for NaCl than for sucrose in the example given. In practice, the change is smaller. In fact, except for very dilute solutions, the property change is always smaller than the simple math of counting ions would predict. For the purposes of this course, just assume that the solutions behave ideally and the number of ions computes as it would be expected to in an ideal solution.

11.5   Colloids

Colloids occur when fine particles of a solid do not dissolve into a solvent to form a solution, but never settle out of solution either. A familiar example of a colloidal dispersion is milk. No solid settles out of milk, yet it isn't a solution and you can prove this to yourself with a simple experiment. Obtain two clear, colorless, undistorted glasses. Place pure water into one. Now add a small amount of milk (a few drops) to the water and stir. The water never becomes transparent. (It will have a slightly cloudy look to it.) You can try adding water to the second glass and a small amount of the solution in glass one to the water. If you can see anything it will be cloudy, although there may be too little of the solid for you to see anything. Other examples of colloids include Jello®, mayonnaise, shaving cream, and porcelain.

Under normal conditions, undissolved large particles are pulled to the bottom of a container by gravity. In colloids, enough thermal energy exists to prevent this. (Brownian motion keeps the particles suspended much like the dust you see suspended in a sunbeam.) Colloids are opaque or translucent because the particles that comprise them are so large they scatter light rather than allowing it to pass straight through the sample. Yet simple filtration (e.g. as in your general

chemistry lab) won't remove the colloid particles because they are so small they pass through the filter pores. For some very large colloids, special filters can remove the particles. Colloid particles are typically removed by adding an agent that causes the particles to clump together. They then settle out of solution or can be filtered away.

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