

In this Chapter, we talk about colligative properties. "Colligative" is not a common word, but it is etymologically related to "collective", which is a useful way to think about it. Colligative properties are solution properties which depend on the collective concentrations of all solutes.

43.1 Preliminaries

Ideally, colligative properties are proportional to the total concentration of all solutes, regardless of identity. The solutes can be molecular or ionic. For ionic solutes, the assumed, ideal condition is full dissociation. Thus, a solution of 0.01 mol of NaCl dissolved in one kg H₂O has a concentration of 0.02 m for total ions.

Pure water boils at 100.0 °C and freezes at 0.00 °C. Salt water boils above 100.0 °C and freezes below 0.00 °C. These changes in boiling point and in freezing point are two examples of colligative properties. At higher concentrations of salt, the boiling point is higher and the freezing point is lower. That's the concentration effect: more concentration gives more effect. Identity doesn't matter. Ideally, a solution of NaCl(aq) and a solution of KBr(aq) will have the same boiling point and freezing point as long as their concentrations are the same.

Colligative properties follow mathematical relationships but, as we had seen for gases in Chapter 17, there's ideal behavior and there's real behavior. The real behavior of colligative properties can easily deviate from ideality, depending on the type of solute. For neutral, molecular solutes, the colligative properties for solutions are reasonably ideal up to ~0.5 m. For ionic solutes, however, there are substantial departures from ideality even at 0.01 m for simple 1+/1- salts and the situation is worse for salts with ions of higher charge. Thus, calculations for ionic solutes are not very meaningful. The primary reason for the greater nonideality of ionics in solution is that the interactions of ions are so strong; they interact strongly with solvent and they interact with other solvated ions nearby. Due to the large deviations from ideality for ionics, we limit calculations to neutral, molecular solutes. Nevertheless, the general qualitative trend for ionics still applies: higher concentrations have greater effects.

Before getting into this, there is one thing which you must understand about this colligative business. Colligative properties are about the solvent and they are not about the solute. Much of our prior discussions focused on the solute. That won't be the case here: we shift our focus to the solvent of the solution. We are comparing the properties of a pure liquid to the properties of that same liquid acting as a solvent in a solution. This is a subtle but important distinction. Be careful on this point. For example, the thermodynamic properties of water when pure are different from the thermodynamic properties of water in an aqueous solution. This works for any solvent. And, under ideal conditions, the identity of the solute in the solution does not matter.

Let me address some of the thermodynamic aspects. In general, all solutions have a natural preference to take up more solvent and to mix if they are able to do so. This can be manifested in several ways. For example, if you add pure H₂O(l) carefully to a container of salt water without a lot of stirring, then the pure water and the solution will eventually diffuse and become fully mixed on their own. This is simple dilution. If you add a solution of 1 m NaCl(aq) to a container of 2 m NaCl(aq) without a lot of stirring, then the solutions will also mix (eventually) on their own and they will achieve some uniform, middle concentration. For all of these cases and under ideal conditions, the system gains entropy upon mixing and this then favors mixing. This gain in entropy is true regardless of solute identity.

Overall then, entropy favors mixing. Our emphasis here will be simple dilution: the scenario will be that of a pure liquid mixing into a solution which has that liquid as the solvent. We will not deal with the scenario of two solutions of different concentrations mixing together. For our purposes here, we can say that entropy favors gaining more solvent; conversely, entropy opposes loss of solvent. There will be a catch to this, however, depending on how the additional, pure solvent is accessible. In one variation, the solution can be in contact with the gas phase of the solvent; for example, a solution of sugar water can be in contact with H₂O(g) in the air which is overhead. In another variation, the solution can be in contact with the solid phase of the solvent compound; for example, a solution of sugar water can be in contact with an ice cube of pure H₂O(s). In a third variation, the solution can be in contact with the liquid phase of the pure solvent but separated by means of a membrane. For any of these cases, although solution entropy favors the solution to take up more solvent and to become more dilute, that must now be tempered by the entropy differences of all phases which are present. Thus, opposing factors will operate and an equilibrium will be involved.

We will consider four colligative properties here: vapor pressure lowering, boiling point elevation, freezing point depression, and osmotic pressure. The first two will receive only qualitative coverage; the latter two will receive both qualitative and quantitative coverage.

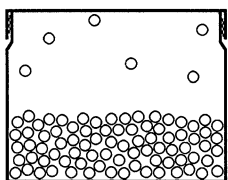
43.2 Vapor pressure lowering and boiling point elevation

Before describing these two properties, let me note that we will limit our discussion of all four colligative properties to solutions with nonvolatile solutes. "Nonvolatile" means not volatile. The adjective "volatile" applies to a compound which vaporizes readily. There is no quantitative definition for this term. "Vaporizes readily" can mean the compound has a relatively high EVP at normal temperature or a relatively low $\Delta H_{\text{vap}}^{\circ}$ or a relatively low nbp. For example, diethyl ether, $\text{C}_2\text{H}_5\text{OC}_2\text{H}_5$, is generally regarded as very volatile. Methanol, CH_3OH , is less volatile than diethyl ether but more volatile than water. Flip back to the end of Chapter 35 and look at the numbers for those compounds. The noun volatility refers to the property itself. Diethyl ether has high volatility; water's volatility is much less. On the other hand, the adjective nonvolatile means not vaporizing readily, which can be taken to mean not vaporizing to a significant extent. For example, food oils are nonvolatile. Sublimable solids can also be characterized in these terms but their volatility is much less than that of a liquid. Nevertheless, we can regard mothballs as volatile if compared to rocks which are nonvolatile.

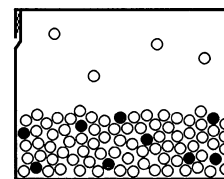
Why do we need this?

Limiting our coverage to nonvolatile solutes allows us to focus on the gas phase of the solvent only. The solutes do not vaporize, so they contribute no vapor pressure of their own. We focus only on the vaporization of the solvent and we do not worry about any vaporization of solute. On the other hand, your instructor may choose to include volatile solute cases. Heed if there's need.

I already mentioned one case of boiling point elevation when I said upstairs that salt water boils above $100.0\text{ }^{\circ}\text{C}$. Before we cover bp elevation, let me do vapor pressure lowering. They're related.



In Chapter 35, we talked extensively about the equilibrium of vaporization and about EVP. Vaporization was favored by entropy; condensation was favored by enthalpy (IFs). That discussion applied for a single compound by itself, as depicted at left. When that compound is the solvent in a solution, however, it already has a bit more entropy so the entropy gained by vaporization is slightly diminished. Vaporization is still favored, but now not as much. For this reason, the point of equilibrium balance for the system has a lower concentration (pressure) in the gas phase, as depicted at right. (The dark circles are the solute again.) In other words, the EVP for the solvent in the solution is less than the EVP for the pure solvent by itself. With more and more solute dissolved in solution, the effect is greater and the EVP is even lower. For example, the EVP of $\text{H}_2\text{O}(g)$ over pure $\text{H}_2\text{O}(l)$ is 17.5 Torr at $20\text{ }^{\circ}\text{C}$. The EVP of $\text{H}_2\text{O}(g)$ over a solution of 0.1 m sucrose, $\text{C}_{12}\text{H}_{22}\text{O}_{11}(aq)$, is less and the EVP over a solution of 0.2 m $\text{C}_{12}\text{H}_{22}\text{O}_{11}(aq)$ is even lower. This is the colligative property known as vapor pressure lowering. All of this is a result of the diminished tendency of a solvent to escape from solution to the gas phase.



Our second colligative property, the elevation of the boiling point, is a natural consequence of the lowering of the vapor pressure. As also seen in Chapter 35 for a pure liquid, EVP increases with temperature and there will be some T for which EVP is one atm. That's the normal boiling point for the pure liquid. For a solution, the EVP is less to start with, so it takes even higher temperature to hit one atm. Thus, the temperature needed for boiling the solvent in a solution is higher. With more and more solute dissolved in solution, the effect is even greater. For example, the nbp of pure H_2O is $100.0\text{ }^{\circ}\text{C}$; the boiling point of a solution of 0.1 m $\text{NaCl}(aq)$ is higher than this, and the boiling point of a solution of 0.2 m $\text{NaCl}(aq)$ is even higher.

Although there are mathematical relationships for ideal conditions for vapor pressure lowering and for boiling point elevation, we will not cover those aspects. Your instructor may choose to add the math. As always, that is their prerogative. Here, we simply stick to qualitative coverage for these two colligative properties.

43.3 Freezing point depression

Freezing point depression is the lowering of the freezing point of the solvent in a solution, relative to the freezing point of the pure solvent compound. Pure water freezes at $0.00\text{ }^{\circ}\text{C}$ but salt water freezes below $0.00\text{ }^{\circ}\text{C}$. This is freezing point depression. This is also the reason for salting roads in the winter

time. Roads are salted so that water on the road stays liquid below 0.00 °C. If temperatures drop too far, then the water still freezes. That's life in the colder climes.

We limit our discussion here to solutes which are solids in their native phase. Under these conditions, as the solution freezes, it is the solvent (not the solute) which is freezing. There are some subtleties to this process which need elaboration. Let me describe some of this.

If you chill an aqueous solution of whatever solute to 0.00 °C, ice will not form. When ice forms, the amount of H₂O(l) in the solution decreases; that loss of solvent from the solution results in an increase in the concentration of solute. These results are opposed by entropy. If you keep dropping the temperature below 0.00 °C, however, you will hit a point at which some water will start to freeze to ice. How cold does it have to be to start freezing? That depends on the concentration of the dissolved solute. For example, pure H₂O(l) freezes at 0.00 °C, but the water in 0.1 m C₁₂H₂₂O₁₁(aq) freezes below this and the water in 0.2 m C₁₂H₂₂O₁₁(aq) freezes even lower. As you freeze any aqueous solution, some water freezes to ice and, if you stopped there, you would have an equilibrium system of ice sitting in a solution at some temperature below 0.00 °C. Since the concentration has increased in the meantime, then further freezing requires dropping the temperature even more. In other words, the temperature must continue dropping in order to freeze out more and more solvent. Here's an important point about freezing a solution: you don't get a single freezing point like you get for a pure liquid by itself. For a solution, the freezing point itself is changing; it's going down as the solution gets more and more concentrated. You get a freezing range instead of a single freezing point.

But that only goes so far. Eventually, the solution gets so concentrated that the solute reaches saturation. Now, as more solvent is removed by freezing, the solute crystallizes as its own solid phase. Note the distinction: the solute never "freezes"; it crystallizes to a solid. If you freeze the whole mixture to a solid, then you have a heterogeneous, solid mixture of frozen solvent and crystallized, solid solute.

For our colligative coverage, we won't be going that far. We're just covering the start of the freezing process of the solvent. Thus, the freezing point here is technically the initial freezing point: the first temperature at which freezing begins.

Let's take a look at some of the mathematics involved in this process. The solution first starts to freeze at some T lower than the usual freezing point, fp , of the pure solvent. The difference in those temperatures is called the freezing point depression, designated ΔT_f .

$$\Delta T_f = fp \text{ of pure solvent} - fp \text{ of actual solution}$$

Note that Δ here is not your usual final-minus-initial thermo thing. ΔT_f is just the drop in freezing point in °C (or K) for the solution relative to the pure solvent compound, expressed as a positive number.

This freezing point depression is proportional to the molal concentration of the solute. We can write this proportionality as the following.

$$\Delta T_f \propto m$$

Now, we introduce a proportionality constant which is labeled K_f ; this is called the freezing point constant or the cryoscopic constant.

$$\Delta T_f = K_f \times m$$

The units of K_f are °C/m or K/m. It doesn't matter if you use °C or K because a ΔT is the same in both scales (as noted in Section 18.5).

Different solvents behave differently in this regard, so different solvents all have different values of K_f . Here are cryoscopic constants and freezing points for four compounds.

	water H ₂ O	benzene C ₆ H ₆	cyclohexane C ₆ H ₁₂	naphthalene C ₁₀ H ₈
fp	0.00 °C	5.49 °C	6.59 °C	80.26 °C
K_f	1.86 °C/m	5.07 °C/m	20.8 °C/m	7.45 °C/m

Naphthalene is a solid at room temperature, so we don't usually think of it as a solvent. Nevertheless, it has found use in these applications when heated above its melting point.

Let's illustrate a sample calculation. Let's take the aqueous sugar solution from the last Chapter and find its freezing point depression and its freezing point.

“ Consider a sugar solution which contains 2.84 g sucrose (table sugar), $C_{12}H_{22}O_{11}$, in 100.0 g water. Calculate the molality of the sugar in the solution. ”

You need that molality from Chapter 42 and you will need K_f for water from the above list. Your setup for the molality from last Chapter should have come out in the end to be the following.

$$\frac{2.84/342.30}{0.1000} \text{ m}$$

We enter this into the equation for the freezing point depression.

$$\Delta T_f = K_f \times m$$

$$\Delta T_f = \frac{1.86 \text{ }^\circ\text{C}}{\text{m}} \times \frac{2.84/342.30}{0.1000} \text{ m} = 0.154 \text{ }^\circ\text{C}$$

Now, that's how much the freezing point went down. The actual freezing point for the solution is calculated from that. Since

$$\Delta T_f = \text{fp of pure solvent} - \text{fp of actual solution}$$

then

$$\text{fp of actual solution} = \text{fp of pure solvent} - \Delta T_f$$

so that

$$\text{fp of actual solution} = 0.00 \text{ }^\circ\text{C} - 0.154 \text{ }^\circ\text{C} = -0.15 \text{ }^\circ\text{C}$$

and that is your answer. The solution of 2.84 g sucrose in 100.0 g water would start freezing at $-0.15 \text{ }^\circ\text{C}$.

Your turn.

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Example 1. Consider a solution of 1.26 g of S_8 dissolved in 100.0 g benzene. What is the freezing point depression (in $^\circ\text{C}$) for the solution? What is the freezing point (in $^\circ\text{C}$) of the solution?

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The depression is $0.249 \text{ }^\circ\text{C}$ to give a freezing point of $5.24 \text{ }^\circ\text{C}$.

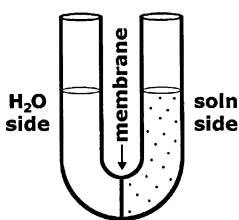
OK, time to chill out on freezing point depression. We now turn to our final colligative property.

43.4 Osmotic pressure

Osmotic pressure is based on osmosis. This colligative is a bit more involved than the prior colligatives so let me do some explaining. We need to bring in the notion of a semipermeable membrane. A semipermeable membrane is one which is only partly permeable. Permeable here is used in the usual dictionary sense related to permeate which means to pass through. A semipermeable membrane is a membrane which allows some things to pass through but not others. In other words, it leaks but it leaks selectively. Membranes can be chosen which allow solvent to pass through but block other things from permeating. For example, some membranes can block ions while some other membranes can block large solute molecules while letting small water molecules through.

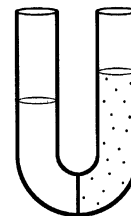
Osmosis involves the passage of solvent from one solution into another solution through such a semipermeable membrane. Due to entropy, the solutions naturally want to mix but solute cannot pass through the membrane. Solvent can pass through, however, so there is a net flow of solvent from the solution of low concentration to the solution of high concentration. One "solution" is often chosen to be just plain solvent, so its solute concentration is zero. In this application, pure solvent will pass through the semipermeable membrane into the solution on the other side. Thus, the solution gains solvent and this constitutes dilution of the solution. As more solvent passes through the semipermeable membrane, the solution gets more dilute but this continues only to a certain point as we shall see.

There are numerous physical setups which can be used for an osmotic pressure experiment. I'm going to keep it basic here and consider a simple U-tube setup. We will also work only with water as solvent. We will use a U-tube with a semipermeable membrane which prevents bulk contact between the



left compartment and the right compartment. Although the membrane prevents bulk contact, the membrane does allow water molecules to slowly pass through, either way. As shown at left, we will start the process with pure water in the left compartment and with a solution whose concentration is to be measured in the right compartment. For purposes of this illustration, we start with both sides at the same liquid height. Gravity prefers equal heights for the two sides. (This assumes the density of the solution is the same density as the water, which is a good assumption for very dilute solutions, as in our application here.)

Although water can pass either way through the membrane, more will pass from left to right than will pass from right to left. This will give a net transfer of water from left to right, thus fulfilling the entropy preference for the solution to gain solvent. This causes the levels of the two sides to become unequal, as shown at right. That introduces a conflict, however, because that opposes gravity. Thus, we have opposing processes operating and this is an equilibrium system. At equilibrium, the two opposing processes are in dynamic balance; water still flows through the membrane, but now at equal rates in both directions.



The difference in liquid heights causes a small pressure difference between the two sides due to gravity. That pressure difference is the colligative property known as osmotic pressure. Osmotic pressure is given by capital, Greek pi, Π .

For ideal solutions, the extent of the pressure difference depends on the concentration of the solute in the solution at equilibrium. We use the concentration unit of molarity in this application. Arithmetically, the colligative relationship is

$$\Pi = M \times R \times T$$

where R is once again the gas constant in the same version as used back in Chapter 16. In fact, this equation looks hauntingly like the concentration expression for the ideal gas equation in Chapter 16 but there is no simple direct connection. By the way, this equation is good for any solvent and not just for water.

Osmotic pressure is much more sensitive as a colligative property than those discussed earlier. The greater sensitivity allows for the use of very dilute solutions. Let's see how this goes with a sample calculation.

We'll refer back to the sucrose solution which we've been using since Chapter 42.

“ Consider a sugar solution which contains 2.84 g sucrose (table sugar), $C_{12}H_{22}O_{11}$, in 100.0 g water. Calculate the molality of the sugar in the solution. ”

But now we modify this with two changes. Instead of 2.84 g sucrose, we'll use one-thousand times less: 2.84 mg. We'll also specify 100. mL of solution since we need molarity in this calculation. So here's the question: what is the osmotic pressure (at equilibrium) at 25 °C for 100. mL of solution which contains 2.84 mg $C_{12}H_{22}O_{11}$?

Set it up.

$$\Pi = M \times R \times T$$

For molarity, you need the mass of the sucrose as 0.00284 g, you need its molar mass (342.30 g/mol) and you need the solution volume as 0.100 L soln.

$$\Pi = \frac{\overbrace{0.00284}^M \text{ mol}}{\overbrace{0.100}^L} \times \overbrace{0.08206}^R \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \overbrace{298}^T \text{ K}$$

Plug in, punch out, round off. You get 0.00203 atm. That may not sound like much pressure, but it would give a difference in height of 21.0 mm for the two sides of the U-tube. When you punched out the numbers for the molarity, you may have noticed that the solution concentration was 8.30×10^{-5} M. That's very dilute compared to typical circumstances. Thus, osmotic pressure measurements are very sensitive and can provide an easily measured result for a very dilute solution. You could not reliably measure freezing point depression on a solution which is this dilute.

Dilute solutions become important for solutes with very large molar masses, such as for various proteins and polymers. For these, molar masses can get into the thousands or tens-of-thousands (or even higher) in terms of g per mol. Here's an Example for you.

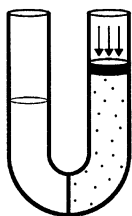
Example 2. Consider a protein which has a molar mass of 12,300 g/mol. For 100. mL of a solution which contains 26.3 mg of the protein at 18 °C, what is the osmotic pressure (in atm)?

You should get 5.11×10^{-4} atm.

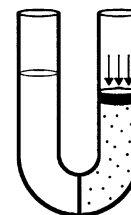
Having looked at several calculations for freezing point depression and osmotic pressure, you can look back and see that all of these calculations involved the molar mass of the solute. Historically, these methods were actually used to determine molar masses of compounds. One would experimentally measure the freezing point depression or the osmotic pressure; with this result and the experimental parameters (masses, volumes, etc.), one could then calculate the molar mass. While that was important historically, other methods are used for molar mass determination nowadays.

Although freezing point depression and osmotic pressure are not used much experimentally, the processes themselves remain extremely important. The importance of freezing point depression was already noted for salting roads. Osmotic pressure also has applications, and the greatest applications are extremely big and extremely small.

The extremely big impact is potability, as applied to water for drinking purposes. This aspect utilizes the process of reverse osmosis (RO), which means running osmosis backwards relative to its natural direction. As seen above, the natural direction in osmosis is for net transfer of pure water through the membrane into the solution side. This sets up a pressure, imposed by gravity, which opposes further dilution. Now, what if we add to that opposing pressure?



As depicted at left, if we place a movable plug (piston) atop the solution and press down with extra, sufficient pressure, then this added pressure works against the natural dilution. Pure water is forced back through the membrane into the left side. The solution gets more concentrated, while the left side increases in the volume of pure water, as shown at right. As the solution gets more concentrated, it takes even more added pressure to have an effect. The net result is an increase in the volume of water on the left side. As such, this constitutes a purification of water: solvent water from the solution is forced through the membrane as purified water.



This process is operated on huge scales to provide drinkable water from a nondrinkable source. Reverse osmosis is one desalination method which is used to remove salts from brackish water and seawater. Historically, evaporation methods accounted for most global desalination efforts but reverse osmosis has grown in use tremendously. In these applications, the goal is to produce water that is "pure enough", which means reducing the concentration of salts down to potable and palatable levels. In many lands which are located near seawater, large reverse osmosis plants provide drinking water. The methods used are very advanced, unlike our simple U-tube illustrations here, but the underlying concepts are the same. Even in areas where municipal water sources don't require desalination, many homeowners opt to have RO units installed as a separate tap in their home to further purify their drinking water.

So why can't you drink seawater by itself? It will dry you out. This relates to the impact of osmosis on the extremely small scales. Small in size, but huge in consequences.

Cell membranes are actually semipermeable; they are permeable to a variety of substances (but not all), either directly through the bilayer or through the various channels which are built into the bilayer (as mentioned in Chapter 40). The water inside your cells exchanges with water outside your cells. The water inside your cells has a wide variety of salts, sugars, proteins, etc. dissolved in it. Likewise, the water outside your cells, such as in your blood, also has a wide variety of salts, sugars, proteins, etc. dissolved in it, although they may not be the same salts, sugars, proteins, etc. or they may be at different

concentrations of the individual components. In order to keep things in osmotic balance, the total concentrations of all solutes must be reasonably close inside and outside the cell.

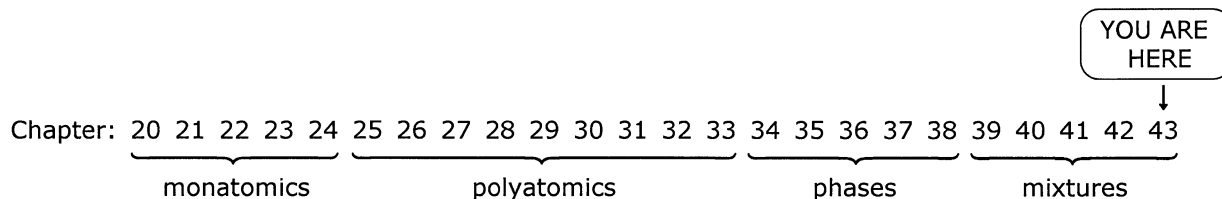
In biological or physiological applications, the term "tonicity" is used to indicate osmotic balance across a cell membrane. If a cell is placed into an "isotonic" solution, then that solution and the solution inside the cell are in osmotic balance and there will be equal flow of water both ways between them. (The prefix "iso-" again means equal or same, as seen for isoelectronic in Section 24.2.) If a cell is placed into a "hypotonic" solution, then that solution has a lower total solute concentration than the solution inside the cell. (The "hypo-" prefix means below or less than.) Under these conditions, the natural tendency is for water to flow from the outside solution (lower concentration of solutes) into the cell (higher concentration of solutes). If excessive, the additional water inside the cell can kill the cell and can cause it to burst open. Conversely, if a cell is placed into a "hypertonic" solution, then that solution is of higher total concentration than the solution inside the cell. (The prefix "hyper-" means over or beyond.) The natural tendency is for water to flow out of the cell, and the cell shrivels. If excessive, the cell can die.

As you can see, osmotic pressure can be a life or death situation. For humans, water balance is very important. Human blood plasma and cellular fluids are in the range of 0.3 M for total solutes. By far, most of the solutes in blood plasma are Na^+ and Cl^- ions (which are counted as dissociated, separate solutes). If you ever receive intravenous fluids in a hospital, then those fluids should be isotonic with blood plasma. Commonly these fluids are NaCl and/or glucose solutions, possibly with some nutrients and some medications thrown in. Even if you're seriously dehydrated, you cannot receive much pure water intravenously because it would dilute your plasma and render it hypotonic. That could rupture your blood cells which could kill you. As an opposite effect, this is why drinking a lot of seawater can be harmful. The ionic solutes in seawater (mostly Na^+ and Cl^- again) exceed 1 M, far too hypertonic for your system. Your cells and body fluids would lose water and you dehydrate. That's somewhat of an irony: you dehydrate by drinking seawater.

With this we close out on colligatives. In fact, we close out on a tremendous journey.

43.5 The journey never really ends.

After 24 Chapters, we reach the end of this particular journey which has dealt with the study of matter itself.



It has been a long journey. We began with the single atom, starting with humble H and its one electron, and then progressing into many-electron systems. Then we bonded, forming the wide diversity of covalent, ionic and metallic polyatomic entities; these ranged from simple linears through octahedral, all with their various hybridized (or not) orbitals or with their bonding and antibonding MOs. Then we came up to the observable scale by allowing IFs to kick in between molecules, thus resulting in liquid and solid phases. Finally, we saw why different compounds mix and we saw why some compounds refuse to mix. This journey has been all about matter. Your matter. My matter. The matter of the universe. From the simplest atom to some of the basics of life itself.

But the travels are not over. As noted at the end of Chapter 40, the next journey has already started. The trails have actually overlapped since Chapter 34 when we addressed entropy and its role alongside that of enthalpy in determining phases and then mixtures. The continuation of this journey takes us further into aspects dealing with enthalpy and entropy, applying these now to any process, including chemical reactions. We have some fundamental questions to address along the way.

Problems

1. True or false.
 - a. The EVP of 1 m $\text{KBr}(aq)$ is greater than the EVP of pure water.

- b. An aqueous solution of 0.1 m glucose, $C_6H_{12}O_6$, has a lower boiling point than an aqueous solution of 0.05 m sucrose, $C_{12}H_{22}O_{11}$.
 - c. When salt water freezes, only the water actually freezes.
 - d. Reverse osmosis can be used to purify water.
2. Consider a solution of 0.84 g of urea, $(NH_2)_2CO$, in 50.0 g H_2O .
 - a. What is the freezing point depression (in $^{\circ}C$) for the solution?
 - b. What is the freezing point (in $^{\circ}C$) of the solution?
 3. Erythritol, $C_4H_{10}O_4$, is a sweetener which is used as a sugar substitute in a variety of foods and drinks. In order to prepare a solution of erythritol in 100. g H_2O which freezes at $-0.75^{\circ}C$, how many grams of erythritol are needed?
 4. Polyvinyl alcohol (PVA) is a polymer used in a variety of applications including plastics and foods. The empirical formula of the polymer is C_2H_4O but the true molecular formulas can vary from 200 to 4,000 times the empirical formula unit, depending on the method of preparation. Typically, for any particular batch, a mixture of large molecules of somewhat different sizes is obtained. Thus, the molar mass is for the average of the batch. If you have a 125.0 mL solution containing 75.4 mg of PVA with an average molar mass of 14,100 g/mol, what would be the osmotic pressure (in atm) at $35^{\circ}C$?