

Colligative properties are those properties of solutions that depend on the number of dissolved particles in solution, but not on the identities of the solutes. For example, the freezing point of salt water is lower than that of pure water, due to the presence of the salt dissolved in the water. To a good approximation, it does not matter whether the salt dissolved in water is sodium chloride or potassium nitrate; if the molar amounts of solute are the same and the number of ions are the same, the freezing points will be the same. For example, AlCl_3 and K_3PO_4 would exhibit essentially the same colligative properties, since each compound dissolves to produce four ions per formula unit. The four commonly studied colligative properties are freezing point depression, boiling point elevation, vapor pressure lowering, and osmotic pressure. Since these properties yield information on the number of solute particles in solution, one can use them to obtain the molecular weight of the solute.

Freezing Point Depression

The presence of a solute lowers the freezing point of a solution relative to that of the pure solvent. For example, pure water freezes at 0°C ; if one dissolves 10 grams of sodium chloride (table salt) in 100 grams of water, the freezing point goes down to -6.4°C . If one uses sucrose (table sugar) instead of sodium chloride, 10 grams in 100 grams of water gives a solution with a freezing point of -0.54°C . The reason that the salt solution has a lower freezing point than the sugar solution is that there are more particles in 10 grams of sodium chloride than in 10 grams of sucrose. Since sucrose, $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ has a molecular weight of 342.34 grams per mole and sodium chloride has a molecular weight of 58.44 grams per mole, 1 gram of sodium chloride has almost six times as many sodium chloride units as there are sucrose units in a gram of sucrose. In addition, each sodium chloride unit comes apart into two ions (a sodium cation and a chloride **anion**) when



The freezing point depression of a solution containing a dissolved substance, such as salt dissolved in water, is a colligative property.

dissolved in water. Sucrose is a nonelectrolyte, which means that the solution contains whole $C_{12}H_{22}O_{11}$ molecules. In predicting the expected freezing point of a solution, one must consider not only the number of formula units present, but also the number of ions that result from each formula unit, in the case of ionic compounds. One can calculate the change in freezing point (ΔT_f) relative to the pure solvent using the equation:

$$\Delta T_f = i K_f m$$

where K_f is the freezing point depression constant for the solvent ($1.86^\circ\text{C}\cdot\text{kg}/\text{mol}$ for water), m is the number of moles of solute in solution per kilogram of solvent, and i is the number of ions present per formula unit (e.g., $i = 2$ for NaCl). This formula is approximate, but it works well for low solute concentrations.

Because the presence of a solute lowers the freezing point, many communities put salt on their roads after a snowfall, to keep the melted snow from refreezing. Also, the antifreeze used in automobile heating and cooling systems is a solution of water and ethylene glycol (or propylene glycol); this solution has a lower freezing point than either pure water or pure ethylene glycol.

Boiling Point Elevation

The boiling point of a solution is higher than that of the pure solvent. Accordingly, the use of a solution, rather than a pure liquid, in antifreeze serves to keep the mixture from boiling in a hot automobile engine. As with freezing point depression, the effect depends on the number of solute particles present in a given amount of solvent, but not the identity of those particles. If 10 grams of sodium chloride are dissolved in 100 grams of water, the boiling point of the solution is 101.8°C ; which is 1.8°C higher than the boiling point of pure water). The formula used to calculate the change in boiling point (ΔT_b) relative to the pure solvent is similar to that used for freezing point depression:

$$\Delta T_b = i K_b m ,$$

where K_b is the boiling point elevation constant for the solvent ($0.52^\circ\text{C}\cdot\text{kg}/\text{mol}$ for water), and m and i have the same meanings as in the freezing point depression formula. Note that ΔT_b represents an increase in the boiling point, whereas ΔT_f represents a decrease in the freezing point. As with the freezing point depression formula, this one is most accurate at low solute concentrations.

Vapor Pressure Lowering

The vapor pressure of a liquid is the **equilibrium** pressure of gas molecules from that liquid (i.e., the results of evaporation) above the liquid itself. A glass of water placed in an open room will evaporate completely (and thus never reach equilibrium); however, if a cover is placed on the glass, the space above the liquid will eventually contain a constant amount of water vapor. How much water vapor is present depends on the temperature, but not on the amount of liquid that is present at equilibrium (provided some liquid is present at equilibrium). (At room temperature, the vapor pressure of pure water is about 20 Torr, which is about one-fortieth of the total atmospheric pressure on a "normal" day at sea level.)

If, instead of pure water, an **aqueous solution** is placed in the glass, the equilibrium pressure will be lower than it would be for pure water. Raoult's law states that the vapor pressure of the solvent over the solution is proportional to the fraction of solvent molecules in the solution; that is, if two-thirds of the molecules are solvent molecules, the vapor pressure due to the solvent is approximately two-thirds of what it would be for pure solvent at that temperature. If the solute has a vapor pressure of its own, then the total vapor pressure over the solution would be:

$$P_{\text{vap}} = \frac{2}{3} (\text{pure solvent vapor pressure}) + \frac{1}{3} (\text{pure solute vapor pressure})$$

Generally, one expects that solutes which are liquids in their pure form (such as ethyl alcohol) will have some vapor pressure of their own, whereas ionic compounds (such as sodium chloride) will not contribute to the total vapor pressure over the solution.

One consequence of this lowering in vapor pressure may be observed in a spilled can of soda. As the water evaporates, the soda becomes more sugar and less water, until the vapor pressure of the water is so low that it barely evaporates. As a result, the spilled soda remains sticky for a long time. Contrast this behavior with that of a water spill.

Osmotic Pressure

Osmosis is the process whereby a solvent passes through a semipermeable membrane from one solution to another (or from a pure solvent into a solution). A semipermeable membrane is a barrier through which some substances may pass (e.g., the solvent particles), and other species may not (e.g., the solute particles). Important examples of semipermeable membranes are the cell walls in cells of living things (plants and animals). Osmosis tends to drive solvent molecules through the semipermeable membrane from the low solute concentrations to the high solute concentrations; thus, a "complete"

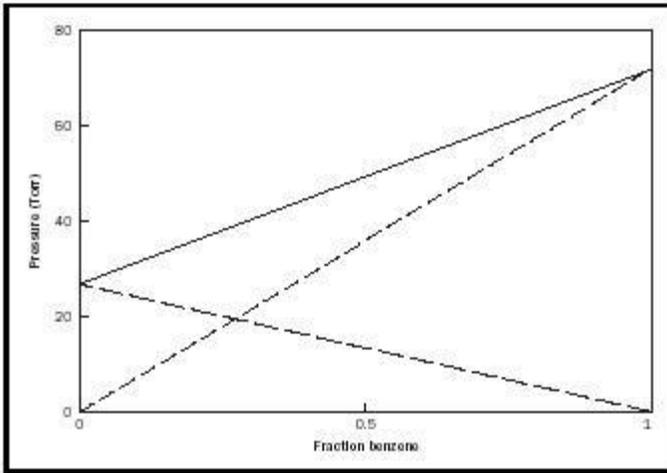


Figure 1. Vapor pressure over a benzene-toluene solution, plotted as a function of the fraction of benzene molecules in the solution. The solid curve is the total vapor pressure, while the short-dashed and long-dashed curves are the vapor pressures from the benzene and toluene, respectively. Note that the two dashed curves add up to the solid curve.

osmosis process would be one that ends with the solute concentrations being equal on both sides of the membrane. Osmotic pressure is the pressure that must be applied on the high concentration side to stop osmosis.

Osmosis is a very useful process. For example, meats can be preserved by turning them into jerky: The meat is soaked in a very concentrated salt solution, resulting in dehydration of the meat cells. Jerky does not spoil as quickly as fresh meat, since bacteria on the surface of the salty meat will fall victim to osmosis, and shrivel up and die. This process thus extends the life of the meat without the use of refrigeration.

There are times when one wishes to prevent osmosis when two solutions (or a pure solvent and a solution) are on opposite sides of a semipermeable membrane. Osmosis can be prevented by applying pressure to the more concentrated solution equal to the osmotic pressure on the less concentrated side. This can be accomplished either physically, by applying force to one side of the system, or chemically, by modifying a solute concentration so that the two solute concentrations are equal. (If one applies a pressure greater than the osmotic pressure to the higher concentration solution, one can force solvent molecules from the concentrated solution to the dilute solution, or pure solvent. This process, known as reverse osmosis, is often used to purify water.) A hospital patient receiving fluids intravenously receives an intravenous (IV) solution that is isotonic with (i.e., at the same solute concentration as) his or her cells. If the IV solution is too concentrated, osmosis will cause the cells to shrivel; too dilute a solution can cause the cells to burst. Similar problems would be experienced by freshwater fish swimming in salt water, or saltwater fish swimming in freshwater. The osmotic pressure, like other colligative

properties, does not depend on the identity of the solute, but an electrolyte solute will contribute more particles per formula unit than a nonelectrolyte solute.

SEE ALSO [SOLUTION CHEMISTRY](#).

Wayne B. Bosma

Bibliography

Atkins, Peter, and de Paula, Julio (2001). *Physical Chemistry*, 7th edition. New York: W.H. Freeman.

Brown, Theodore L.; LeMay, H. Eugene; Bursten, Bruce E.; and Burdge, Julia R. (2002). *Chemistry: The Central Science*, 9th edition. Upper Saddle River, NJ: Prentice Hall.

Read more: <http://www.chemistryexplained.com/Ce-Co/Colligative-Properties.html#ixzzOXQfZYsZ8>