Solutions

Explore

What are the various colligative properties of solutions?

Colligative Properties

Colligative properties depend on the number of particles of **solute** present in a **solution**. They are not determined by the properties of the solute or the **solvent**, and always follow some specific patterns.
The primary property which is affected by colligative factors is vapor pressure. In a liquid, vapor pressure is caused by particles escaping and changing to a gas state. If the solution is in a closed container, they cannot escape the entire system, so they occupy the space above the solution and below the top of the container. As temperature increases, more particles evaporate and vapor pressure increases.
When a solute is added to a solvent, it becomes more difficult for the particles of the solvent to evaporate because there is less surface area available. In other words, some of the surface is blocked by the solute particles. The more solute particles there are, the more vapor pressure decreases.
The change in vapor pressure affects two specific colligative properties. First, when a solid solute is added to a liquid to create a solution, the *boiling point* of that solvent will *increase*. This **colligative property** is described as **boiling point elevation**. Because particles of solute on the surface of the solution make it more difficult for evaporation to occur, boiling requires extra energy.
Second, adding solid solute to a solvent will *decrease the* ***freezing*** *point* of the solvent used in the solution. This colligative property is known as **freezing point**

depression. Freezing point depression is also proportional to the number of solute particles in the solution and is independent of the type of particle. The solute particles interfere with the solvent's formation of crystals. The freezing process does not occur until the temperature is lowered further.

To calculate changes in boiling or freezing point, the concentration of the solution must be known. Molality is proportional to the change in boiling and freezing points, but the proportional relationship depends on the specific pure solvent or solution that is present. Constants for boiling point elevation (K_b) and freezing point depression (K_f) have been determined through research.

Based on this proportional relationship, the change in temperature required for boiling can be expressed by the equation

$\Delta T_b = iK_b m$

Similarly, the change in temperature required for freezing is calculated using the equation

$\Delta T_f = iK_f m$

The Greek symbol delta, Δ, represents “change.” The *m* in each equation is the molality. The van 't Hoff factor, *i*, is determined by the solute’s behavior in the particular solvent. For example, solutes that do not dissociate (or break apart into two components) in water, $i = 1$. For solutes that completely dissociate into two ions, $i = 2$. Sodium chloride, NaCl, completely dissociates into two ions, Na^+ and Cl^-. Some solutes completely dissociate into three ions, and for those $i = 3$. For example, $CaCl_2$ completely dissociates into three ions, one Ca^+, and two Cl^-. The values of *i* continue to increase as the number of ions created during dissociation increases.

Calculating Freezing Point: Sample Problem

A solution is made with 12.55 g of sodium chloride, NaCl, and 180.4 mL of water. The water is 35.0°C. What is the freezing point of this solution? Assume the NaCl

dissolves completely in the water The density of water at 35.0°C is 0.9941 g/mL, and the K_f of H_2O is 1.858°C kg/mol.

Solution

To find the freezing point of the solution, use the formula

$\Delta T_f = iK_fm$.

First, it is known that $i = 2$ because two ions (Na^+ and Cl^-) form when NaCl dissociates.

The value K_f of water is given as 1.858°C kg/mol.

The value of m must be calculated. This is the molal concentration of the solution, or

$$m = \frac{\text{moles of solute}}{\text{1 kg of solvent}}$$

To find the moles of NaCl, first use atomic masses from the periodic table to find the molar mass of NaCl.

Sodium (Na) = 22.99 g

Chlorine (Cl) = 35.45 g

$$\begin{aligned}\text{molar mass} &= 22.99 \text{ g/mol} + 35.45 \text{ g/mol}\\ &= 58.44 \text{ g/mol}\end{aligned}$$

One mole of NaCl has a mass of 58.44 g.

The solution contains 12.55 g of NaCl. This is divided by the molar mass to find the number of moles present.

$$\text{number of moles} = 12.55 \cancel{\text{g}} \times \frac{1 \text{ mol}}{58.44 \cancel{\text{g}}} = 0.214750 \text{ mol}$$

Notice that two extra digits are included because this is an intermediate calculation. Carrying extra digits can avoid round-off errors in the final answer. There are 0.214750 mol of solute present. Now the total moles of solute and solvent must be found. The volume of the water in the solution is 180.4 mL, and the density of the water is given as 0.9941 g/mL.

$$\text{mass of solvent} = 0.9941 \frac{\text{g}}{\cancel{\text{mL}}} \times 180.4 \cancel{\text{mL}} = 179.336 \text{ g}$$

Again, two extra digits are carried. Because the molality ratio uses the unit of kilograms, the mass of the water must be converted from grams.

$$\text{mass of solvent} = 179.3 \cancel{\text{g}} \times \frac{1 \text{ kg}}{1000 \cancel{\text{g}}} = 0.179336 \text{ kg}$$

The values can then be substituted into the molality ratio to find m.

$$m = \frac{\text{moles of solute}}{\text{kg solvent}} = \frac{0.214750\ \text{mol NaCl}}{0.179336\ \text{kg H}_2\text{O}} = 1.19747\ \text{mol/kg}$$

Now all of the values can be substituted into the formula to find the freezing point depression.

$\Delta T_f = iK_fm$

$$\Delta T_f = 2 \times 1.858°\text{C}\ \frac{\cancel{\text{kg}}}{\cancel{\text{mol}}} \times 1.19747\ \frac{\cancel{\text{mol}}}{\cancel{\text{kg}}}$$

$$= 4.44980°\text{C}$$

Rounding to 4 significant digits, ΔT_f = 4.450°C. The freezing point of pure water is 0°C. Apply the freezing point depression as calculated above to determine the new freezing point:

0 °C – 4.450°C = –4.450°C

Adding 12.55 g of sodium chloride to 180.4 mL of water will decrease the freezing point to more than 4.4 degrees lower than normal.

Osmotic Pressure

The transport of solutions through structures is part of many natural processes. Plant and animal cells, for example, rely on the movement of solutions through their bodies to transport nutrients, waste, and other biochemicals. These cells are surrounded by semipermeable membranes that control substances that enter and exit cells. Only solute particles of a certain size can pass through the membrane. Typically, this means that solutes are filtered out by the membrane, while the pure solvent passes through it. This movement is called osmosis.

Think about a solution of sugar and water that is passed through a semipermeable membrane. The sugar molecules will only be found on one side of the membrane (the solution side), and pure water will be on the other side (the “solvent side).

The water molecules move back and forth through the membrane, but more of them tend to move from the solvent side to the solution side. This creates **osmotic pressure** on the solution side of the membrane, which could eventually increase so that osmosis stops completely. Factors such as temperature also affect osmotic pressure, since the energy of the particles will affect how they move through the semipermeable membrane.

While osmosis is a complex process, the key is to remember that only water can pass through the “wall” because the sugar molecules are too large. The more sugar molecules in the solution, the more osmotic pressure increases.
