

17.5: Boiling-Point Elevation and Freezing-Point Depression

Freezing point depression is a colligative property observed in solutions that results from the introduction of solute molecules to a solvent. The freezing points of solutions are all lower than that of the pure solvent and is directly proportional to the *molality* of the solute.

$$\begin{aligned}\Delta T_f &= T_f(\text{solvent}) - T_f(\text{solution}) \\ &= K_f \times m\end{aligned}$$

where ΔT_f is the freezing point depression, T_f (solution) is the freezing point of the solution, T_f (solvent) is the freezing point of the solvent, K_f is the freezing point depression constant, and m is the molality.

Introduction

Nonelectrolytes are substances with no ions, only molecules. **Strong electrolytes**, on the other hand, are composed mostly of ionic compounds, and essentially all soluble ionic compounds form electrolytes. Therefore, if we can establish that the substance that we are working with is uniform and is not ionic, it is safe to assume that we are working with a nonelectrolyte, and we may attempt to solve this problem using our formulas. This will most likely be the case for all problems you encounter related to freezing point depression and boiling point elevation in this course, but it is a good idea to keep an eye out for ions. It is worth mentioning that these equations work for both volatile and nonvolatile solutions. This means that for the sake of determining freezing point depression or boiling point elevation, the vapor pressure does not effect the change in temperature. Also, remember that a pure solvent is a solution that has had nothing extra added to it or dissolved in it. We will be comparing the properties of that pure solvent with its new properties when added to a solution.

Adding solutes to an ideal solution results in a positive ΔS , an increase in entropy. Because of this, the newly altered solution's chemical and physical properties will also change. The properties that undergo changes due to the addition of solutes to a solvent are known as **colligative properties**. These properties are dependent on the number of solutes added, not on their identity. Two examples of colligative properties are boiling point and freezing point: due to the addition of solutes, the boiling point tends to increase, and freezing point tends to decrease.

The freezing point and boiling point of a pure solvent can be changed when added to a solution. When this occurs, the freezing point of the pure solvent may become lower, and the boiling point may become higher. The extent to which these changes occur can be found using the formulas:

$$\Delta T_f = -K_f \times m \quad (17.5.1)$$

$$\Delta T_b = K_b \times m \quad (17.5.2)$$

where m is the solute *molality* and K values are proportionality constants; (K_f and K_b for freezing and boiling, respectively).

Molality

Molality is defined as the number of moles of solute per kilogram **solvent**. Be careful not to use the mass of the entire solution. Often, the problem will give you the change in temperature and the proportionality constant, and you must find the molality first in order to get your final answer.

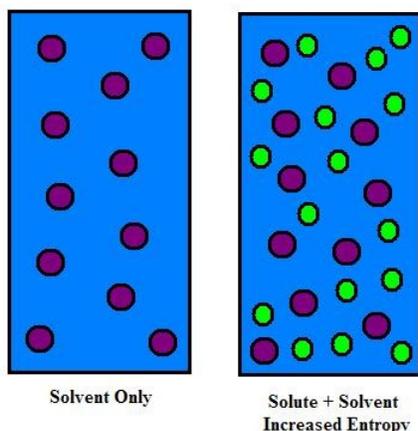
If solving for the proportionality constant is not the ultimate goal of the problem, these values will most likely be given. Some common values for K_f and K_b respectively, are in Table 17.5.1:

Table 17.5.1: Ebullioscopic and cryoscopic constants for select solvents. Note that the nature of the solute does not affect colligative property relations.

| Solvent | K_f | K_b |
|-------------|-------|-------|
| Water | 1.86 | .512 |
| Acetic acid | 3.90 | 3.07 |
| Benzene | 5.12 | 2.53 |

| Solvent | K_f | K_b |
|---------|-------|-------|
| Phenol | 7.27 | 3.56 |

The solute, in order for it to exert any change on colligative properties, must fulfill two conditions. First, it must not contribute to the vapor pressure of the solution, and second, it must remain suspended in the solution even during phase changes. Because the solvent is no longer pure with the addition of solutes, we can say that the **chemical potential** of the solvent is lower. Chemical potential is the molar Gibb's energy that one mole of solvent is able to contribute to a mixture. The higher the chemical potential of a solvent is, the more it is able to drive the reaction forward. Consequently, solvents with higher chemical potentials will also have higher vapor pressures.

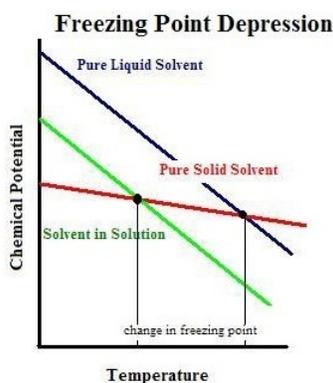


The boiling point is reached when the chemical potential of the pure solvent, a liquid, reaches that of the chemical potential of pure vapor. Because of the decrease in the chemical potential of mixed solvents and solutes, we observe this intersection at higher temperatures. In other words, the boiling point of the impure solvent will be at a higher temperature than that of the pure liquid solvent. Thus, **boiling point elevation** occurs with a temperature increase that is quantified using

$$\Delta T_b = K_b m \quad (17.5.3)$$

where K_b is known as the *ebullioscopic constant* and m is the molality of the solute.

Freezing point is reached when the chemical potential of the pure liquid solvent reaches that of the pure solid solvent. Again, since we are dealing with mixtures with decreased chemical potential, we expect the freezing point to change. Unlike the boiling point, the chemical potential of the impure solvent requires a colder temperature for it to reach the chemical potential of the pure solid solvent. Therefore, a **freezing point depression** is observed.



✓ Example 17.5.1

2.00 g of some unknown compound reduces the freezing point of 75.00 g of benzene from 5.53 to 4.90 °C. What is the molar mass of the compound?

Solution

First we must compute the molality of the benzene solution, which will allow us to find the number of moles of solute dissolved.

$$\begin{aligned} m &= \frac{\Delta T_f}{-K_f} \\ &= \frac{(4.90 - 5.53)^\circ C}{-5.12^\circ C/m} \\ &= 0.123m \end{aligned}$$

$$\begin{aligned} \text{Amount Solute} &= 0.07500 \text{ kg benzene} \times \frac{0.123 \text{ m}}{1 \text{ kg benzene}} \\ &= 0.00923 \text{ m solute} \end{aligned}$$

We can now find the molecular weight of the unknown compound:

$$\begin{aligned} \text{Molecular Weight} &= \frac{2.00 \text{ g unknown}}{0.00923 \text{ mol}} \\ &= 216.80 \text{ g/mol} \end{aligned}$$

The freezing point depression is especially vital to aquatic life. Since saltwater will freeze at colder temperatures, organisms can survive in these bodies of water.

? Exercise 17.5.1

Benzophenone has a freezing point of 49.00°C. A 0.450 molal solution of urea in this solvent has a freezing point of 44.59°C. Find the freezing point depression constant for the solvent.

Answer

$$9.80^\circ C/m$$

Applications

Road salting takes advantage of this effect to lower the freezing point of the ice it is placed on. Lowering the freezing point allows the street ice to melt at lower temperatures. The maximum depression of the freezing point is about $-18^\circ C$ ($0^\circ F$), so if the ambient temperature is lower, NaCl will be ineffective. Under these conditions, $CaCl_2$ can be used since it dissolves to make three ions instead of two for NaCl.



Figure 17.5.1: Workers manually spreading salt from a salt truck in Milwaukee, Wisconsin. (CC BY-SA 2.0 Generic; [Michael Pereckas](#) via [Wikipedia](#))

References

1. Atkins, Peter and de Paula, Julio. *Physical Chemistry for the Life Sciences*. New York, N.Y.: W. H. Freeman Company, 2006. (124-136).

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