

Solutions

Definitions pertaining to solutions

Solute is the substance that is dissolved. It is usually present in the smaller amount.

Solvent is the substance that does the dissolving. It is usually present in the larger amount.

Solubility is the limit to which something dissolves. It is the maximum amount of solute for a given amount of solvent.

A **saturated** solution is one that is at the solubility limit.

An **unsaturated** solution is one that is less than saturated.

Concentration Units

Molarity

Molarity (M) is defined as the ratio of the number of moles of solute to the total volume of the solution in liters.

$$M = \frac{\text{moles of solute}}{\text{Liters of solution}}$$

Percent units

There are three percent units that are commonly used, percent by mass, percent by volume and percent mass-volume. Percent by mass is defined as the ration of the mass of the solute to the mass of the solution multiplied by 100. Percent by volume is defined as the ratio of the volume of the solute to the volume of the solution multiplied by 100. Percent mass-volume is defined as the ratio of the grams of solute to milliliters of solution multiplied by 100.

$$\%(m/m) = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

$$\%(v/v) = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100$$

$$\%(m/v) = \frac{\text{grams of solute}}{\text{mL of solution}} \times 100$$

Related to the percent units are the units parts per million, parts per billion and parts per trillion. These are calculated the same way as the percent units (part per hundred) except we multiply by a million, billion or trillion.

Mole fraction

Mole fraction (X) is defined as the ratio of the moles of solute to the total number of moles in the solution.

$$X_i = \frac{n_i}{n_T}$$

Molality

Molality (m) is defined as the ratio of the number of moles of solute to the mass of the **solvent** in kilograms.

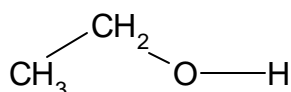
$$m = \frac{\text{moles of solute}}{\text{kg of solvent}}$$

You should understand how to convert between these various concentration units.

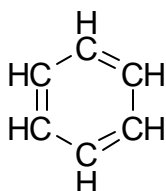
The solution process

Molecular solutes

A molecular substance will dissolve in a particular solvent depending on the types of intermolecular interactions present in the solute and the solvent. If the solute and solvent have similar interactions they will tend to dissolve in one another. If they do not they will not dissolve. Ethanol



has a structure similar to that of water. It also has similar intermolecular forces. Ethanol will dissolve in water in any proportion. However, benzene



only has London forces. It does not dissolve in water. The reason for this that is often given is that there are no attractions between the benzene molecules and the water molecules. This is not true. There are significant attractions between the two molecules. What makes benzene insoluble in water is that the attractions between individual water molecules are far stronger than the attractions between benzene molecules and water molecules.

Ionic solutes

All ionic solids dissolve in water to some extent. The ionic solids that are generally considered to be insoluble are only very slightly soluble (less than 0.01 g of solute per 100 mL of solution). What determines if an ionic solid is very soluble or only very slightly soluble? There are two competing interactions that determine this. The first is the strength of the ionic bond holding the ions together. The second is the strength of the ion-dipole interaction between the ions and the water. If the ionic bond is stronger the substance is only slightly soluble in water. If the ion-dipole interaction is stronger the substance is more soluble in water. Ions with higher charges will tend to have stronger ionic bonds. Smaller ions will tend to have larger ion-dipole interactions.

Things that can affect the solubility of a compound

Temperature

For solid solutes, in general, as we increase the temperature the solubility of the compound increases. There are a few exceptions to this (such as $\text{Ca}_3(\text{SiO}_4)_2$). The reason for this trend is that as the temperature is increased the kinetic energy of the solute particles increases which, in general, makes it easier for them to enter into solution.

For gaseous solutes the opposite is true. As the temperature is increased, gases become less soluble. This is because of the increase in kinetic energy of the gas particles. With more kinetic energy they are more likely to leave the solution.

Pressure

Pressure does not affect the solubility of solid solutes to any great extent.

Pressure does affect the solubility of gases though. If we increase the pressure of a gas above a liquid, the gas will become more soluble in that liquid. We must increase the partial pressure of the gas. Increasing the total pressure by adding another gas will not increase the solubility of the original gas.

Colligative Properties

Vapor pressure of a solution

The vapor pressure of a solution containing a non-volatile non-electrolyte solute is proportional to the mole fraction of the solvent and the vapor pressure of the pure solvent:

$$P_{sol'n} = X_{solv} P_{solv}^{\circ}$$

Because the mole fraction of the solvent in a solution is always less than one, we see that the vapor pressure of a solution is always less than the vapor pressure of the pure solvent.

Any solution that obeys this relationship is called an ideal solution. Most solutions do not exactly obey this relationship, but all dilute solutions ($X_{\text{solv}} > 0.90$) can be considered ideal.

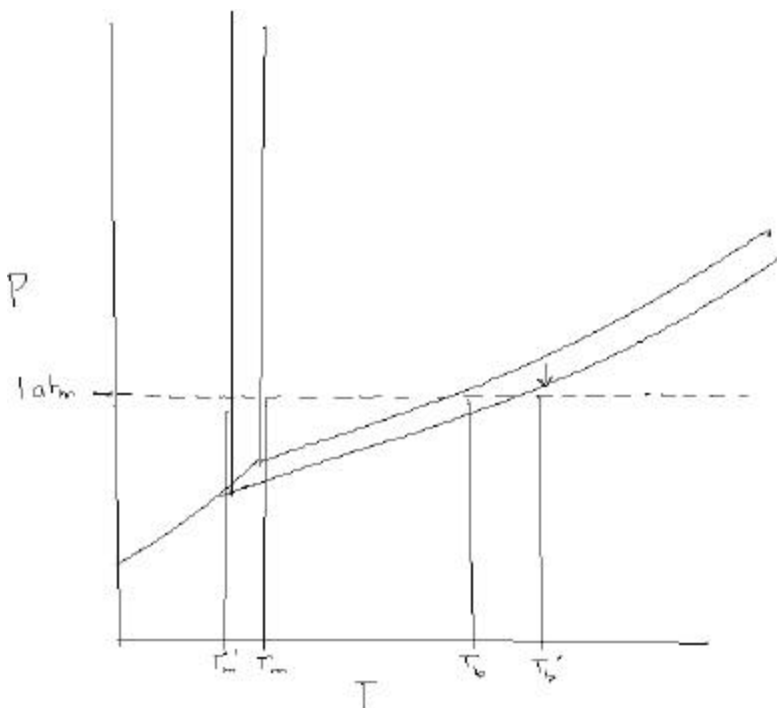
A solution may also contain a volatile solute, for example ethanol in water. In this case the vapor pressure of the solution is a combination of the vapor pressure of the solute and the solvent. We must calculate the contribution of each with the above equation and add them together to get the vapor pressure of the solution. For an ideal solution with two volatile components (A and B), the vapor pressure is given by:

$$P_{\text{sol'n}} = X_A P_A^\circ + X_B P_B^\circ$$

This lowering of the vapor pressure of a solution has implications also for the solution's melting and boiling points.

Boiling point elevation and freezing point depression

When a solute is added to a solvent the vapor pressure of the solution decreases. This is the same as lowering the boiling point curve on the phase diagram.



When the boiling point curve lowers the melting point curve shifts to the left. This implies that the boiling point has increased and the melting point has decreased (see phase diagram).

The amount of the increase or decrease can be calculated if we know the molality of the solution. This calculation using molality is an approximation that works for dilute

solutions. A more accurate value can be obtained by using the mole fraction of the solution.

The amount by which a solution's boiling point increases is given by the formula:

$$\Delta T_b = K_b mi$$

ΔT_b is the change in the boiling point, K_b is the boiling point elevation constant and depends only on the solvent used, m is the molality of the solution, and i is a value called the van't Hoff factor. The van't Hoff factor is a value that shows how many ions come from each formula unit. For molecular solutes i equals 1.

A similar formula gives the depression of the freezing point:

$$\Delta T_f = K_f mi$$

Here K_f is the freezing point depression constant and, again, depends only on the solvent.

Osmotic Pressure

Osmosis is the process by which a solvent flows through a **semi-permeable membrane** to equalize the concentration on both sides of the membrane. A **semi-permeable membrane** is a membrane that will allow some molecules, such as solvent molecules, to pass through, but not others. We can stop the osmotic flow through the membrane by applying pressure to counteract the flow. The amount of pressure we need to stop the osmosis is called the **osmotic pressure** and is given by:

$$\Pi = MRT$$

This equation is related to the Ideal Gas Law. Π is the osmotic pressure, M is the molarity of the solution, R is the ideal gas constant in pressure-volume units ($L \text{ atm mol}^{-1} \text{ K}^{-1}$ for example), and T is the Kelvin temperature. Because molarity is the same as moles over volume the Osmotic pressure equation can be rewritten as:

$$\Pi = \frac{n}{V} RT = \frac{nRT}{V}$$

This is the same as the Ideal Gas Law with Π being pressure.

Applications of Colligative Properties

The colligative properties of a solution can be used to calculate the molar mass of a solute. They can also be used to calculate the percent dissociation of a weak electrolyte. Examples of these problems are in your textbook.