Introduction

We've discussed previously how solids and liquids are put together (and how we can get information on the properties of a liquid or solid starting from as little information as a chemical formula). Now, we will turn our attention to how solutions are put together. What makes one substance dissolve in another one?

A definition of solubility

We already know a little bit about solubility. In the previous course, we talked about the **solubility rules**, which we used to guess whether or not a precipitate would form when we mixed some salt solutions together.

We said something was **soluble** if it would dissolve in water, and something was **insoluble** if it would not. That's a simplistic view that doesn't take into account other solvents or the fact that sometimes only a certain amount of one substance will dissolve in another. This is where the concept of **solubility** comes in.

Solubility is **quantitative** - it is **the amount of one substance that dissolves in another** and is usually expressed in units like grams per milliliter (g/mL) or grams per liter (g/L). We say a substance is "soluble" if the solubility is large (example - sodium chloride in water). On the other hand, a substance is "slightly soluble" or "insoluble" if the solubility is very small (example - a substance like silver chloride in water).

How and why things dissolve

For this section, we will primarily look at more familiar solutions like solids in liquids and liquids in liquids, though our discussion here will be valid for everything except gas/gas solutions. (There's no real solution **process** in a gas, since gas molecules act as if no other molecules are there!).

We start off by asking ourselves this: "What happens when a substance dissolves?" Consider a simple example like sucrose (table sugar, $C_{12}H_{22}O_{11}$) dissolving in water. A chemical equation for the process might look like this:

$$C_{12}H_{22}O_{11}(s) \leftrightarrow C_{12}H_{22}O_{11}(aq)$$

(The double-headed arrow, "↔", indicates that this reaction can proceed either way - to the left **or** right)

Graphically, we might picture the solution process this way:

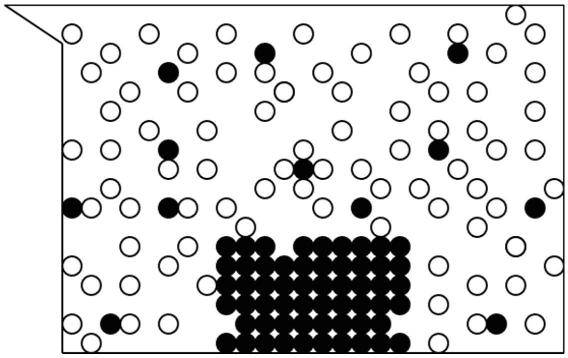

Illustration 1 - Sugar in water

- The sugar molecules are represented by black spheres, while water molecules are represented by white spheres.
- During the solution process, the water molecules pull the sugar out of the sugar crystals and into solution.
- The sugar and water molecules interact in some way.

From the picture (which follows from our pictures of the pure solid and pure liquid), we can see that the liquid and solid interact. Since what holds a liquid together in the first place are intermolecular forces, we would reasonably expect that **intermolecular forces** also play a role in the solution process. We can see this role in several sample situations:

1. Oil and water don't mix: Oil is nonpolar, so it's held together by strictly London forces. Water is different because it is polar **and** hydrogen bonds.
2. Sugar and water **do** mix: Sugar is a polar molecule, as is water.
3. Ethanol and water mix **extremely** well. Ethanol (CH_3CH_2OH - draw a Lewis structure) and water are both polar and both hydrogen bond.
4. Sodium chloride and water mix: Sodium chloride is ionic, while water is polar.

From these four examples and many others, we can see that **substances with similar intermolecular forces can dissolve each other**. In short, **"like dissolves like"**. Polar substances tend to dissolve in other polar substances. Nonpolar substances dissolve in other nonpolar substances. Ionic substances (held together by interactions of full positive and negative charges) may dissolve in polar substances.

<u>Saturation and supersaturation</u>

For some sets of substances, you can dissolve any amount of one substance in the other. The substances are **miscible**. Gases are completely **miscible** in each other - it doesn't matter how much of each gas is present - you can make a solution from all of it. Certain mixtures of liquids are also miscible.

For most sets of substances, though, you reach a point where you can't dissolve any more solute in the solvent. This point is called the **saturation** point, and the solution is said to be **saturated**.

Another way to define the saturation point is to say that the substances in solution are in **equilibrium**. For a system of sucrose in water, the solution is saturated when the rate the solid sucrose dissolves is **equal to** the rate that the aqueous sucrose **precipitates**. We know that these processes occur all the time (like phase transitions) because we can

observe crystals become larger in a saturated solution over time - the crystals get larger when the solute precipitates on the outside of crystals already there. Small crystals get smaller and disappear, and large crystals get larger. We will discuss equilibrium more in detail later on.

Before saturation, we refer to a solution as **unsaturated**, meaning that we could dissolve more solute in the solvent if we wished.

It is possible (under certain conditions) to form a solution which has **more solute** than would be able to dissolve normally in a solvent. These solutions are called **supersaturated**, because solute is present at a concentration greater than the concentration at the saturation point. These solutions are **unstable**, though, and even tapping a beaker containing a supersaturated solution can cause the "extra" solute to crash out of solution almost immediately.

<u>Outside factors that affect solubility</u>

We know that intermolecular forces effect solubility - we saw that earlier in this note pack. But some external factors can also affect solubility, and we will discuss two of these:

1) **Temperature:** Temperature has a varied effect on solubility. Increased temperature will either increase solubility, decrease it, or not affect it significantly. Luckily, though, there are a few rules of thumb you can use:

- *Most (but not all) ionic solids* **increase in solubility** in water when temperature is increased. A few decrease in solubility or show no significant change.
- *Gases* tend to be **less soluble** at higher temperatures. Consider what warm Coke tastes like. It tastes flat because CO_2 (responsible for the "bite" in soft drinks) is less soluble in water at room temperature than at the temperature of your fridge.

So, is the solution process endothermic or exothermic? Both! For some solutions (like ammonium chloride in water), the solution process requires energy and is endothermic. For others (like sulfuric acid in water), the solution process releases heat and is exothermic.

2) **Pressure:** Pressure primarily affects the solubility of gases in liquids. When you increase the pressure, the solubility of the gas increases. In fact, the solubility of a gas is **directly proportional to its partial pressure**. This is called **Henry's Law**:

$$\text{gas solubility} = k_H \times P$$

... where k_H is the Henry's Law constant for the gas/liquid solution and P is the partial pressure of the solute gas.

Summary

In this note pack we discussed the solution process. You should understand that solubility is not an on/off property - we quantify it be talking about how much solute will dissolve in a solvent. You should also understand some of the factors (both the types of forces **and** the external ones like temperature and pressure) that influence solubility. You should also be familiar with basic terms like solubility, saturation, and supersaturation.