Introduction

We've discussed the properties of pure liquids, solids, and gases. The properties of pure liquids and solids depended primarily on the type and strength of the forces that held these phases together - ionic bonds, intermolecular forces, covalent bonds, etc. We will now examine what happens when we mix substances together - when we **dissolve** one material in another to form **solutions**.

Nomenclature

Before we discuss solutions, we need to first define some terms related to solutions.

- A **solution** is a **homogeneous mixture** of two or more substances. In other words, a solution is a mixture of substances that is uniform in composition throughout.
- The **solvent** is the component of the solution **present in the greatest amount**. By convention, when we deal with a solid or gas dissolved in a liquid, the liquid is always considered the solvent.
- A **solute** is the component of the solution **present in the smallest amount**, assuming we're talking about a solution containing two components. Otherwise, solutes are components present in amounts smaller than the solvent.

As a simple example, consider a solution of salt water (a small amount of sodium chloride dissolved in a quantity of water). The sodium chloride is the solute, while the water is the solvent. For another example, consider most alcoholic drinks - a mixture containing some ethanol mixed with water. In a drink like a beer or wine (with a small ethanol content), the ethanol is the solute and the water is the solvent.

Types of solution

When you think of a solution, the picture that probably comes immediately to mind is the picture used above as the first example of a solution - a solid dissolved in a quantity of liquid. We can have solutions where the solvent is **any** phase of matter - solid, liquid, or gas. Here are some common types of solution, grouped by solvent type:

1) **Gas solvents**: Gases will dissolve other gases. Remember that gases are mostly empty space, so any gas will mix with any other gas in all proportions. No matter how much of one gas and how much of another you have, you can make a solution. Air is a good example of a gas solution.

Gases are **miscible** - which is another way of saying that they mix in all proportions.

2) **Liquid solvents**: Liquids are versatile solvents. You can dissolve solids, other liquids, or gases into liquid solvents. In fact, solid solutes and liquids solvents are what usually comes to mind when we say the word "solution". However, the other solutions are common as well. We mentioned ethanol/water above - that's a liquid/liquid solution. A good example of a gas dissolved in a liquid would be carbonated water - CO_2 gas dissolved in water. (This is assisted by a chemical reaction between CO_2 and water).

3) **Solid solvents**: Liquids and other solids may dissolve in solid solvents, though most non-scientists don't usually think of them as "solutions". A common example of a solution with a solid solvent is brass - which is a solution of zinc (about 15%) in copper. We commonly call these solid/solid solutions **alloys**.

Expressing the concentration of a solution

Now that we've covered the definition of a solution, we need to discuss how you would describe a solution. You can tell a person to make a solution of salt in water (or to make a solution of hydrochloric acid in water) but unless you tell them something about the **concentration** of the solution that you want, you may be in for a nasty surprise when they bring you the solution - especially with the hydrochloric acid!

We'll now briefly review and discuss concentration units. Some of these you've seen before - both inside chemistry class and out. We'll start with one you've seen **outside** of class as well as inside:

1) **Mass percentage:** The mass percentage is exactly what it sounds like - a percentage of the solute by mass. To apply this solution, just remember that the **total mass** is in this case the mass of the solution. Here's the equation:

$$\text{Mass percentage} = \frac{(\text{Mass solute})}{(\text{mass } \textbf{solution})} \times 100\,\%$$

As you can see, this is a simple percentage. As a quick example, let's look at the mass percentage of a solution of 10.00g NaCl in 1000.0 g water. The mass of the solute (NaCl, the component present in smaller amount) is 10.00 g. The mass of the solution is the mass of water plus the mass of the NaCl, or 1010.0 g. The mass percentage of NaCl in the solution is:

$$\text{Mass percentage} = \frac{(10.0\,\text{g NaCl})}{(1010.0\,\text{g } \textbf{solution})} \times 100\,\% = 0.9901\,\%\ \text{NaCl}$$

In the chemistry lab, you will often see mass percentage used as the concentration unit on bottled of concentrated acid and base bought from chemical suppliers. Mass percentage is also common in industry.

2) **Molarity:** Molarity, abbreviated with a capital **M** (sometimes underlined - **<u>M</u>**), is overall the most practical concentration unit in the laboratory - at least, if you're a chemist. You've no doubt noticed that in our lab, nearly all of the reagent bottles are labeled with molarity units. Molarity is defined as:

$$\text{Molarity}(\underline{\text{M}}) = \frac{(\text{moles of solute})}{(\text{liters of } \textbf{solution})}$$

If you had a solution that contained 3.00 mol of HCl in 500.0 mL of water, you would have a 6.00 M HCl solution.

$$\frac{3.00\,\text{mol HCl}}{0.5000\,\text{L water}} = 6.00\,\underline{\text{M}}\,\text{HCl}$$

3) **Molality:** Molality, abbreviated with a lower case *m*, is similar to molarity even in name. We can define molality as follows:

$$\text{Molality}(m) = \frac{(\text{moles of solute})}{(\text{kilograms of } \textbf{solvent})}$$

If you remember that the density of water at room temperature is approximately 1 kg/L, you'll realize that (for solutions with density similar to that of water - e.g. dilute aqueous solutions at room temperature), molality is approximately equal to molarity. So why do we need molality?

The answer to that question relates to **thermal expansion**. While it's difficult to change the volume of a liquid or solid with pressure, there is often a noticeable change in the volume of a liquid or solid with **temperature**. This means that **molarity depends on temperature, since the volume of a solution also depends on temperature!** So in certain situations, we need to use a temperature-independent unit like molality. Since molality is based on masses and moles (which do not vary with temperature), molality is **not** temperature dependent.

Quick example: you make a solution that contains 3.00 mol of HCl in 500.0 g of water. That solution is a 6.00 *m* HCl solution.

$$\frac{3.00\,\text{mol HCl}}{0.5000\,\text{kg water}} = 6.00\,m\,\text{HCl}$$

Tips to prevent unit confusion: *If you have been using "m" as an abbreviation for "moles",* ***stop.*** *The proper abbreviation for moles is "mol". Use "mol" or you'll get confused! Also, the proper abbreviation for molarity is CAPITAL M - make sure you can distinguish your capital "M" from your small "m" when you write it! Otherwise, you'll confuse yourself when dealing with moles (mol), molarity (M), and molality (m)!*

4) **Mole fraction:** Mole fraction is abbreviated with an X. It's typically used in calculations involving vapor-liquid equilibrium and gas compositions. We won't get too deeply into either topic, but some problems involving Raoult's Law use the mole fraction. We can define mole fraction very simply:

$$X_A = \frac{(\text{moles of substance A})}{(\text{total moles of solution})}$$

The **sum** of all the mole fractions of the materials in a solution (including the solvent) will equal 1. As a simple example, consider a solution of two gases - 1.5 moles of oxygen gas and 3.7 moles of nitrogen gas.

$$X_{O_2}=\frac{1.5\,\text{moles}\;O_2}{5.2\,\text{moles total}}=0.29$$

X_{N2}, then, is equal to 1.00-0.29 = 0.71 .

Converting between concentration units for solutions

Often, we may have concentration expressed in terms of one unit and need it in another. You will need to practice converting between the different units, as you'll be expected to be able to do this without much difficulty on exams and in the lab. Here are a few tips to consider:

- If you are converting from a volume-based unit (molarity) to a mass-based unit (molality. mass percentage), the **density** of the solution is required. The reverse is also true - you need the density of the solution to convert from a mass-based unit to a volume-based unit.
- If you are converting from a volume based unit (molarity) to a mole-based unit (mole fraction), you again need the **density** of the solution.
- If you are converting between mass-based units and mole-based units or you're converting between two different mass-based units, you do **not** need to know the solution density.

As an example, we will look at a very common situation in the chemical laboratory. We have ordered a solution of sulfuric acid, H_2SO_4, from a vendor. The vendor ships us a solution labeled "Sulfuric acid, 69.00% by mass". We'd like to prepare solutions of a certain **molarity**, since all our chemical equations are in moles and volumes are easy to measure out. To prepare solutions of a given molarity, we must first find out the molarity of the initial solution:

Find the molar concentration of a 69.00% H_2SO_4 solution. The density of this solution (obtained from the supplier's label) is 1.5989 g/mL.

First, compare the units we have with the units we want. We **have** mass percentage, which is *mass* of solute over *mass* of solution (expressed as a percentage). We **want** the molarity, which is the *moles* solute over *liters* solution. So, our task is to convert mass of sulfuric acid to moles, and then convert mass of the sulfuric acid to volume.

Converting the mass H_2SO_4 to moles: It's simplest to assume an amount of sulfuric acid solution and work the problem based on that assumption. We will assume we are dealing with *exactly* 100 g of solution. This makes the amount of sulfuric acid in the solution easy to calculate: 69.00% of 100 g is 69.00 g. (It doesn't matter what number you assume for amount of solution - so we pick something easy.) Then, convert the calculated mass of sulfuric acid to moles.

$$69.00\,\text{g } H_2SO_4 \times \frac{\text{mol } H_2SO_4}{98.01\,\text{g } H_2SO_4} = 0.7040\text{ mol } H_2SO_4$$

[98.01 g/mol is the formula weight of H_2SO_4]

We've solved half the problem! Now, we just need the volume of the solution (in liters).

Finding the volume of the solution: We assumed we had 100g of solution, so we will simply use the **density** to convert that amount to volume.

$$100\,\text{g solution} \times \frac{\text{mL solution}}{1.5989\,\text{g solution}} \times \frac{10^{-3}\,\text{L}}{\text{mL}} = 0.06254\,\text{L solution}$$

We now know the moles H_2SO_4 present in 0.06254 L of solution. All we have to do to find molarity is use the definition!

Finding the molarity: We use the simple definition of molarity - moles per liter.

$$\frac{(0.7040\,\text{mol } H_2SO_4)}{(0.06254\,\text{L solution})} = 11.26\,\underline{\text{M}}\,H_2SO_4$$

So, the bottle of stock solution was **11.26 <u>M</u> sulfuric acid.**

Practice

See the example problems on page 520-523 (7^{th} edition). These are worked out for you. Also, try some of the odd-numbered problems in the "Solution concentration" section on pages 547 and 548. Answers (but not complete worked-out solutions) for these are in the back of the book.

Summary

This note pack covered the basic units and terminology of solutions. You should know words like solution, solute, and solvent. You should also understand that solutions are more than just solids dissolved in liquids - the solvent can be any phase, and so can the solute! Finally, you should know the common units of solution concentration and be able to convert between them