# 9 Solutions

A solution is a homogeneous mixture, which means that it is uniform throughout in appearance and composition. It consists of two classes of components:

- Solute – the component(s) in smaller quantity; and
- Solvent – the component in major quantity.

If both of these components are infinitely soluble in each other and are in roughly equal proportions, then the distinction between solute and solvent becomes irrelevant and the identification of solute and solvent becomes arbitrary.

Aqueous solutions are solutions in which the solvent is water. In an earlier part of this course, we learned that when ionic solutes are dissolved in water, they break down into their component ions. Thus, for example, when mercury (I) phosphite dissolves in water, we do not see at the microscopic level any particle like mercury (I) phosphite, $(Hg_2)_3(PO_3)_2$ . What we see at the microscopic level are mercury (I) and phosphite ions

$$(Hg_2)_3(PO_3)_2\ (aq) \qquad 3\ Hg_2^{2+}\ (aq)\ +\ 2\ PO_3^{3}\ (aq)$$

This does not mean to imply that water only dissolves ionic compounds. On the contrary, it dissolves many other substances such as sugar, which is non-ionic. It does not dissolve large molecules such as oils.

There are two things we always want to know in chemistry. They are:

- The reactions going on; and
- The quantity of each substance present.

We have studied some of the reactions that can occur in solutions qualitatively. Such reactions include precipitation reactions, acid-base reactions and redox reactions. In this section, we will be examining ways we can express the quantity of each substance in the solution.

## 9.1 Solution Composition

For every solution, there is a limit to the amount of solute that can be dissolved in the solvent. For example, if we add salt to water, it will dissolve. If we add more salt to the water, it too will dissolve. However, at a certain point, no more salt will dissolve and any additional salt will simply fall to the bottom of the container. At this point, the solution is said to be **saturated**. A **saturated solution** is one which has the maximum amount of solute dissolved in the solvent. A solution that has not reached the maximum limit of dissolved solute is said to be **unsaturated**. An unsaturated solution can dissolve more solute, so solute added to an unsaturated solution will dissolve and not settle to the bottom. Generally, if we want to create a saturated solution, we will add solute to the solvent until no more solute will dissolve and then filter off the saturated liquid into a clean, dry, container. Note that the container must be dry, or the solution will not longer be saturated as the few drops of liquid on the surfaces of the container can dissolve more solute.

We can refer to the amount of solute dissolved in the solvent qualitatively. We say that a solution is **concentrated** of it has a large amount of solute. A solution which has only a relatively small amount of dissolved solute is termed **dilute**.

We now turn our attention to quantitative expressions for the amount of solute dissolved in the solvent.

Solutions are homogeneous entities. This means that they have a uniform composition. We make use of this by expressing the relative amount of solute molecules to solvent molecules. Thus we can carry out calculations based on relative quantities of solute and solvent – only concerning ourselves about the total amount of material when it is absolutely necessary.

Three common ways of expressing the relative amounts of solute to solvent are as follow:

- Mole fraction (or mole percent);
- Mass fraction (or mass percent); and

- Molarity.

**Example:** If 50.0 g of a solution of benzene in toluene is 0.45% by mass of toluene, calculate the mass of toluene and the mole fraction of toluene.

Molar mass: Benzene = 78.11 Toluene = 92.14

$$0.45\% = \frac{m_T}{m_T + m_B}\ 100\% = \frac{m_T}{50.0\ g}\ 100\%$$

$$m_T = 0.45\%\ \frac{50.0\ g}{100\%} = 0.23\ g\ toluene$$

$$m_B = 50.0\ g \quad 0.23\ g\ toluene = 49.8\ g$$

$$n_T = \frac{0.23\ g}{92.14\ \mathrm{g\ mol^{-1}}} = 0.0024\ mol$$

$$n_T = \frac{49.8\ g}{78.11\ \mathrm{g\ mol^{-1}}} = 0.637\ mol$$

$$X_T = \frac{0.0024}{0.0024 + 0.637} = \frac{0.0024}{0.639} = 0.0038$$

### 9.1.2 CALCULATIONS INVOLVING MOLARITY

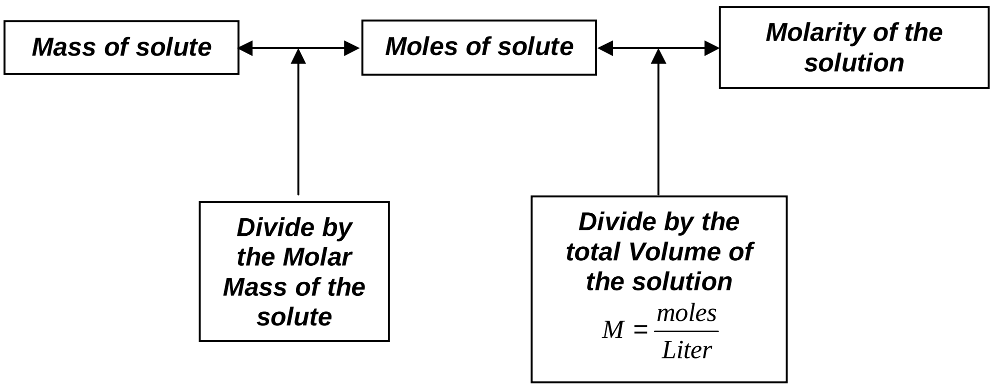
By far the most common method used chemists to express the concentration of a solution is through the **molarity** of the solution. The term, **molar concentration** is also used to refer to the molarity. According to its definition, the molarity is

Molarity: $$M = \frac{moles}{Volume} = \frac{n}{V}$$

In solution chemistry, all variations of this definition will be used:

- Calculation of the molarity of the solution: $M = \frac{n}{V}$
- Calculation of the number of moles of solute: $n = MV$
- Calculation of the volume of the solution: $V = \frac{n}{M}$

Of course, from the moles of solute, we can always calculate the mass of solute:




The reason that molarity is used is that usually in solution chemistry it is more convenient to measure volumes of solutions that to measure mass.

9.1.2.1 CALCULATING MOLARITY

Example: Calculate the molar concentration of a solution made by dissolving 42.3 g glucose in 250.00 mL water.

By definition, $M = \frac{moles\ glucose,\ mol}{Volume\ solution,\ L}$

<u>Calculate the number of moles of solute (glucose)</u>.

## 9.2 Dilution

Example: A Dawson lab technician estimates that she will need 5.00 L of 0.1015 M $H_2SO_4$ solution for an experiment that will be conducted by 475 students over the course of a week. The stock solution of concentrated $H_2SO_4$ is 18.3 M. How much of the stock solution is needed?

The question asks for the volume of the $H_2SO_4$ stock needed. To solve this problem, we will work backward. First we will calculate the number of moles of $H_2SO_4$

Calculate the concentrations of each substance

Note that the volumes are additive.

*V = 10.00 mL + 17.0 mL  = 27.0 mL = 0.0270 mL*

$$[\quad{}^{2}\,] \quad \frac{7.22 \quad 10^{4}}{0.0270} \quad 0.0267$$

3

3

3

3

$$[Ca^{2+}] = \frac{2.61\times10^{\ 3} mol}{0.03932\ L} = 0.0663\ M$$

$$[Na^{+}] = \frac{3.16\times10^{\ 3} mol}{0.03932\ L} = 0.0805\ M$$

$$[Cl\ \ ] = \frac{8.37\times10^{\ 3} mol}{0.03932\ L} = 0.213\ M$$

To recap:

## 9.4 STOICHIOMETRY OF REACTIONS IN SOLUTION

In stoichiometry problems, the balanced chemical equation allows us to relate the number of moles of reactants to the number of moles of product. In the section where stoichiometric calculations were introduced, first we were given the mass. From this mass, we calculated the moles which were then related to the moles of other substances in the chemical equation.

For the stoichiometry of reactions in solutions, the procedure is effectively the same. However, there is one change. Usually in solution chemistry we do not deal directly with masses – we deal with concentrations. Therefore, instead of converting mass to moles, in solution chemistry, we obtain the moles from the concentrations.

As usual for stoichiometric calculations the steps are:

1. Write the balanced chemical equation
2. Calculate the moles of reactants (from the concentrations)
3. Determine the Limiting Reactant
4. Calculate the amount of all compounds remaining at the end of the reaction. Express these amounts in the appropriate units (mass or concentration)

Example: Calculate the amount of $BaSO_4$ that will form when 15.2 mL of 0.125 M $Na_2SO_4$ are added to 14.0 mL of 0.145 M $Ba(NO_3)_2$.

Balanced Chemical Equation: $Na_2SO_4\ (aq)\ +\ Ba(NO_3)_2\ (aq) \quad BaSO_4\ (s)\ +\ 2\ NaNO_3\ (aq)$

Initial moles:

$$n_{Na\ SO} \quad M \quad L \quad mol$$

$$n_{Ba\ NO3\ 2} \quad M \quad L \quad mol$$

Calculate the number of moles given:

1 3

3 4

Thus, for a neutral solution, **pH = 7**.

Let us examine the relative pH's of acid solutions by comparing the pH of a pair of acid solutions:

Solution 1: $[H^+] = 0.0050\ M = 10$