Al-Mustaqbal University College

Department of Medical Device Technologies

Medical Chemistry

First Stage

Lecture Two

Solutions and Their Concentrations

Solution: is a homogeneous mixture of two or more substances. A minor species in a solution is called solute and the major

species is the solvent.

Solute (minor species) + Solvent (major species) $\rightarrow$ Solution

$$NaCl_{(s)} + H_2O_{(l)} \longrightarrow \text{Salt Solution}$$

There are four fundamental ways of expressing solution concentration

a) Molar concentration.

It is defined as the number of moles of that species that is contained in 1 liter of the solution (not 1 L of the solvent).

$$\text{molar concentration} = \frac{no.moles\ solute}{volume\ in\ liters}$$

Symbolized by M, which has the dimensions of mol/L, or $mol\ L^{-1}$.

$$1M = 1\,mol\,L^{-1} = 1\,\frac{mol}{L} = 1\,mmol\,L^{-1} = 1\,\frac{mmol}{L}$$

Example: Calculate the molar concentration of ethanol in an aqueous solution that contains 2.30 g of C_2H_5OH (46.07 g/mol) in 3.50 L of solution.

Solution:
To calculate molar concentration, we must find both the amount of ethanol and the volume of the solution. The volume is given as 3.50 L, so all we need to do is convert the mass of ethanol to the corresponding amount of ethanol in moles.

$$n = 2.30\,g \times \frac{1\,mol\ of\ ethanol}{46.07\,g\ ethanol} = 0.04992\,mol\ ethanol$$

To obtain the molar concentration, we divide the amount by the volume. Thus,

$$M = \frac{0.04992\,mol}{3.50\,L}$$

$$= 0.0143\ \text{mol}/\text{L}$$

There are two ways of expressing molar concentration:

a) Molar analytical concentration (Analytical concentration). A solution gives the total number of moles of a solute in 1 liter of the solution (or the total number of millimoles in 1 mL). Note

In the above example, the molar concentration that we calculated is also the molar analytical concentration (cC2H5OH = 0.0143 M) because the solute ethanol molecules are intact following the solution process.

b) Molar equilibrium concentration (Equilibrium concentration). Refers to the molar concentration of a particular species in a solution at equilibrium.

To specify the molar equilibrium concentration of a species, it is necessary to know how the solute behaves when it is dissolved in a solvent. Equilibrium molar concentrations are usually symbolized by placing square brackets around the chemical formula for the species. For example:

Solution of H_2SO_4 with an analytical concentration of (cH_2SO_4 = 1.0 M), we write

$$[H_2SO_4] = 0.00\ M \qquad [H^+] = 1.01\ M$$

$$[HSO_4^-] = 0.99\ M \qquad [SO_4^{2-}] = 0.01\ M$$

Example: Calculate the analytical and equilibrium molar concentrations of the solute species in an aqueous solution that contains 285 mg of trichloroacetic acid, Cl_3CCOOH (163.4 g/mol), in 10.0 mL (the acid is 73% ionized in water).

Solution:

Cl_3CCOOH= HA.

$$amount\ of\ HA = n_{HA} = 285\ \cancel{mg}\ HA \times \frac{1\cancel{g}\ HA}{1000\ \cancel{mg}\ HA} \times \frac{1\ mol\ HA}{163.4\ \cancel{g}\ HA} = 1.744 \times 10^{-3}\ mol\ HA$$

The molar analytical concentration, C_{HA}, is then

$$C_{HA} = \frac{1.744\times10^{-3} mol\ HA}{10\ mL} \times \frac{1000\ mL}{1L} = 0.174 \frac{mol\ HA}{L} = 0.174\ M$$

In this solution, 73% of the *HA* dissociates, giving H^+ and A^-:

$$HA \leftrightarrow H^+ + A^-$$

The equilibrium concentration of *HA* is then 27% of $C_{HA.}$ Thus,

$$[HA] = C_{HA} \times \frac{100-73}{100} = 0.174 \times 0.27 = 0.047 \frac{mol}{L} = 0.047\ M$$

The equilibrium concentration of A^- is equal to 73% of the analytical concentration of HA, that is,

$$[A^-] = \frac{73\ mol\ A^-}{100\ mol\ HA} \times 0.174 \frac{mol\ HA}{L} = 0.127\ M$$

Because 1 mole of H^+ is formed for each mole of A^-, we can also write

$$[H^+] = [A^-] = 0.127\ M$$

and

$$C_{HA} = [HA] + [A^-] = 0.047 + 0.127 = 0.174\ M$$

Example: Describe the preparation of 2.00 L of 0.108 M $BaCl_2$ from $BaCl_2.2H_2O$ (244.3 g/mol).

Solution:

To determine the number of grams of solute to be dissolved and diluted to 2.00 L, we note that 1 mole of the dihydrate yields 1 mole of $BaCl_2$. Therefore, to produce this solution we will need

$$1.00\cancel{L} \times \frac{0.108\ mol\ BaCl_2.2H_2O}{\cancel{L}} = 0.216\ mol\ BaCl_2.2H_2O$$

The mass of $BaCl_2.2H_2O$ is then

$$0.216\ \cancel{mol\ BaCl_2.2H_2O} \times \frac{244.3\ g\ BaCl_2.2H_2O}{\cancel{mol\ BaCl_2.2H_2O}} = 52.8\ g\ BaCl_2.2H_2O$$

Dissolve $52.8\ g\ BaCl_2.2H_2O$ in water and dilute to 2 L.

Example: Describe the preparation of 500 mL of 0.0740 M Cl^- solution from solid $BaCl_2.2H_2O$ (244.3 g/mol).

Solution:

$$mass\ BaCl_2.2H_2 = \frac{0.0740\ \cancel{mol\ Cl}}{\cancel{L}} \times 0.500\cancel{L} \times \frac{1\ mol\ \cancel{BaCl_2.2H_2O}}{2\ \cancel{mol\ Cl}} \times \frac{244.3\ g\ BaCl_2.2H_2O}{\cancel{mol\ BaCl_2.2H_2O}} =$$

$4.52\ g\ BaCl_2.2H_2O$

Dissolve $4.52\ g\ BaCl_2.2H_2O$ in water and dilute to 0.500L or 500 mL.

b) Percent concentration.

Three common methods are used to express a percent concentration.

$$weight\ percent(w/w) = \frac{weight\ solute}{weight\ solution} \times 100\%$$

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

$$weight/volume\ percent\ (w/v) = \frac{weight\ solute, g}{volume\ solution, mL} \times 100\%$$

Parts per Million and Parts per Billion.
For very dilute solutions, parts per million (ppm) and Parts per Billion (ppb) are convenient ways to express concentration:

$$C_{ppm} = \frac{mass\ of\ solute\ (g)}{mass\ of\ solution\ (g)} \times 10^6\ ppm$$

$$C_{ppm} = \frac{mass\ solute\ (mg)}{volume\ solution\ (L)}\ ppm$$

$$C_{ppb} = \frac{mass\ solute\ (g)}{mass\ solution\ (g)} \times 10^9\ ppb$$

$$C_{ppb} = \frac{mass\ solute\ (\mu g)}{volume\ solution\ (L)}\ ppb$$

Example: What is the molar concentration of K^+ in a solution that contains 63.3 ppm of $K_3Fe(CN)_6$ (329.3 g/mol)?

Solution

Because the solution is so dilute, it is reasonable to assume that its density is1.00 g/mL.

63.3 ppm $K_3Fe(CN)_6$ = 63.3 mg $K_3Fe(CN)_6$/L

$$\frac{no.mol\ K_3Fe(CN)_6}{L} = \frac{63.3\ \cancel{mg\ K_3Fe(CN)_6}}{L} \times \frac{\cancel{1\ g\ K_3Fe(CN)_6}}{1000\ \cancel{mg\ K_3Fe(CN)_6}} \times \frac{1\ mol\ K_3Fe(CN)_6}{\cancel{329.3\ g\ K_3Fe(CN)_6}} = 1.922 \times 10^{-4}\ M$$

$$[K^+] = \frac{1.922\times10^{-4}\ \cancel{mol\ K_3Fe(CN)_6}}{L} \times \frac{3\ mol\ K^+}{\cancel{1\ mol\ K_3Fe(CN)_6}} = 5.77 \times 10^{-4} \frac{mol\ K^+}{L} = 5.77 \times 10^{-4}\ M$$

c) Solution-Diluent Volume Ratios.

The composition of a dilute solution is sometimes specified in terms of the volume of a more concentrated solution and the volume of solvent used in diluting it. The volume of the former is separated from that of the latter by a colon.

For example:

1:4 HCl (i.e. solution contains four volumes of water for each volume of concentrated hydrochloric acid.)

1:1HCl (i.e. solution contains one volume of water for each volume of concentrated hydrochloric acid.)

3:2 H_3Po_4 (i.e. solution contains two volumes of water for three volume of concentrated H_3Po_4)

d) p-Functions.

The concentration of a species frequently express in terms of p-function, p-value. The p-value is the negative logarithm (to the base 10) of the molar concentration of that species. Thus, for the species X, Density and Specific Gravity of Solutions Density of a substance is its mass per unit volume, unit (kg/L or g/mL).

$$pX = -\log[X]$$

Example: Calculate the p-value for each ion in a solution that is 2.00×10^{-3} M in NaCl and 5.4×10^{-4} M in HCl.

$$pH = -\log[H^+] = -\log[5.4 \times 10^{-4}] = 3.27$$

To obtain pNa , we write

$$pNa = -\log[Na^+] = -log[2.00 \times 10^{-3}] = 2.699$$

The total Cl^-concentration is given by sum of the concentration of the two solutes:

$$[Cl^-] = 2.00{\times}10^{-3}\,\mathrm{M} + 5.4{\times}10^{-4}\mathrm{M}$$

$$= 2.00 \times 10^{-3}M + 0.54 \times 10^{-3}\,M = 2.54 \times 10^{-3}M$$

$$pCl = -\log[Cl^-] = -\log 2.54 \times 10^{-3} = 2.595$$

Example: Calculate the molar concentration of Ag^+ in a solution that has a *pAg* of 6.372.

Solution:

$$pAg = -\log [Ag^+] = 6.372$$

$$log[Ag^+] = -6.372$$

$$[Ag^+] = 10^{-6.372} = 4.4246 \times 10^{-7} \approx 4.25 \times 10^{-7} M$$

Density and Specific Gravity of Solutions

Density of a substance is its mass per unit volume, unit (kg/L or g/mL).

$$\boldsymbol{Density} = \frac{\boldsymbol{mass}, (\boldsymbol{kg}, \boldsymbol{g})}{\boldsymbol{Volume}, (\boldsymbol{L}, \boldsymbol{mL})}$$

Specific gravity Specific gravity is the ratio of the mass of a substance to the mass of an equal volume of water. (Specific gravity is dimensionless).

Example: Calculate the molar concentration of HNO_3 (63.0 g/mol) in a solution that has a specific gravity of 1.42 and is 70.5% HNO_3 (w/w).

Solution:

Let us first calculate the mass of acid per liter of concentrated solution.

$$\frac{g\ HNO_3}{L\ reagent} = \frac{1.42\ \cancel{kg\ reagent}}{L\ reagent} \times \frac{\cancel{10^3\ g\ reagent}}{\cancel{kg\ reagent}} \times \frac{70.5\ gHNO_3}{100\ \cancel{g\ reagent}} = \frac{1001\ g\ HNO_3}{L\ reagent}$$

Then,

$$C_{HNO_3} = \frac{1001\ \cancel{g\ HNO_3}}{L\ reagent} \times \frac{1\ mol\ HNO_3}{63.0\ \cancel{g\ HNO_3}} = \frac{15.9\ mol\ HNO_3}{L\ reagent} \approx 16\ M$$

Example: Describe the preparation of 100 mL of 6.0 M HCl from a concentrated solution that has a specific gravity of 1.18 and is 37% (w/w) HCl (36.5 g/mol).

Solution

$$C_{HCl} = \frac{1.18 \times 10^{3}\ \cancel{g\ reagent}}{L\ reagent} \times \frac{37\ \cancel{g\ HCl}}{100\ \cancel{g\ reagent}} \times \frac{1\ mol\ HCl}{36.5\ \cancel{g\ HCl}} = 12\ M$$

The number of moles HCl required is given by

$$no.\ mol\ HCl = 100\ \cancel{ml} \times \frac{\cancel{1\ L}}{1000\ \cancel{mL}} \times \frac{6\ mol\ HCl}{\cancel{L}} = 0.600\ mol\ HCl$$

Finally, to obtain the volume of concentrated reagent we write

$$\boldsymbol{V_{concd} \times C_{concd} = V_{dil} \times C_{dil}}$$

$$V_{concd} \times \frac{12\ mol}{L} = 0.600\ mol\ HCl$$

$$V_{concd\ reagent} = 0.600\ mol\ HCl \times \frac{1\ L_{reagent}}{12\ mol\ HCl} = 0.0500 L\ or\ 50.0\ mL$$

$$\boldsymbol{V_{concd} \times C_{concd} = V_{dil} \times C_{dil}}$$

$$\boldsymbol{L_{concd} \times \frac{mol_{concd}}{L_{concd}} = L_{dil} \times \frac{mol_{dil}}{L_{dil}}}$$

$$\boldsymbol{mL_{concd} \times \frac{mmol_{concd}}{mL_{concd}} = mL_{dil} \times \frac{mmol_{dil}}{mL_{dil}}}$$

Therefore, dilute 50 mL of the concentrated reagent to 600 mL.