Solutions in the liquid phase

A few definitions

A solution is a homogeneous mixture of 2 or more components in the same physical state.

One or more are in small concentration: **solute**(s)

One is the most abundant component: **solvent**

- We have already done exercises on the solutions in the gas phase : **gaseous mixtures**
- We are not going to do exercises in the solid phase: **metallic alloys**
- We are going to study extensively the solutions in the liquid phase, or better **aqueous solutions**:
  - Gas compound(s) + liquid compound(s)
  - Liquid + liquid
  - Solid + liquid

where “liquid” is usually the solvent and, for our purposes, is H_2O

Classification of solutes in aqueous solutions

- Water is the solvent, it is polar: hence it can only dissolve polar molecules
- The polar solutes can be divided into 2 classes:
  - Electrolytes (molecules bound by ionic bonds)
  - Non-electrolytes (molecules bound by polar covalent bonds)

- Electrolytes in solution readily dissociate into solvated ions
- Non-electrolytes do not dissociate once dissolved

Properties of aqueous solutions

- The physico-chemical properties of solutions depend on the actual quantities of molecules included in them
- The quantitative composition of a given solution is called **Concentration**
- There are several **unit of measurements** of concentration: some are dependent on T, some are T-independent

Classification of concentration unit of measurements

T-independent

- Based on the ratio between masses or moles

- Weight %
- Mole fraction
- Molality

T-dependent

- Based on the ratio between mass and volume

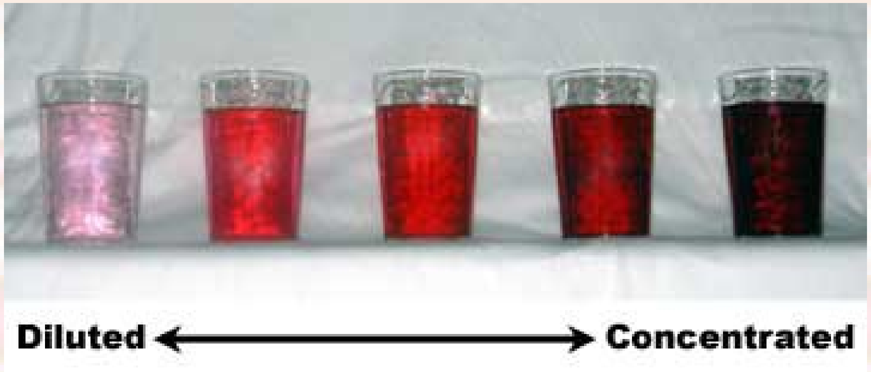
- Volume %
- Molar concentration
- Normality

T-independent units of measurements

- **Weight %**: indicates the ratio betweeen the mass of solute in 100g of solution (%w=g/100g)
- **Mole fraction**: ratio between moles of solute and moles of solution (same as for gas, $x_1=n_1/n_{tot}$)
- **Molality**: the amount (in mol) of solute divided by the mass of the solvent in Kg units ($m = n_i / m_{solvent}$)

T-dependent units of measurements

- **Volume %**: indicates the ratio betweeen the mass of solute in 100ml of solution (%v= g/100ml)
- **Molar concentration**: ratio between the amount (in moles) of solute and volume of solution in L units ($C_1 = n_1 / V_{tot}$ = M =mol/L)
- **Normality**: ratio between equivalents of solute and volume of solution in L units; ratio between molar concentration and equivalence factor ($N = eq._i / V_{tot} = C_1 / f_{eq}$)



Dilutions

- When we add a solvent to a given solution the number of moles of solutes remains unchanged
- Hence the product between the molar concentration and the volume is unchanged

$$C_1V_1=C_2V_2$$

Similarly, if we mix two solutions, the number of moles of solutes are simply the sum of what was present before mixing

$$C_3*(V_1+V_2)=(C_1*V_1)+(C_2*V_2)$$

Exercise 1

- In which volume do we need to dissolve 20g of sodium hydrogen carbonate (baking soda, $NaHCO_3$) so that the solution is 0.5M?

Molar concentration is defined as: $$C_i = \frac{n_i}{V_{tot}} = \frac{g_i}{(MW_i * V_{tot})}$$

Hence we can obtain the Volume, by simply re-shuffling the above equation:

$$V = \frac{g}{(MW * C_i)} = \frac{20}{(84 * 0.5)} = 0.476\,L = 467\,ml$$

Exercise 2

- Calculate the concentration in molality of a solution of sul-phoric acid (H_2SO_4) 11%w.

Weight % means that 11g of acid are in 100g of solution (acid +water)

Hence the solvent content is:

$G_{solvent} = g_{solution} - g_{solute} = 100 - 11 = 89$ g = 0.089Kg

$$m = \frac{n_{solute}}{Kg_{solvent}} = \frac{g_{solute}}{(MW_{solute} * Kg_{solvent})} = \frac{11}{(98 * 0.089)} = 1.26\,m$$

Exercise 3

- The density of Na^+ in human plasma is 3.4g/L. Which is the corresponding molar concentration?

d=g/V, hence we can rearrange the definition of Molar concentration as such:

$$C = \frac{n}{V} = \frac{g}{(MW * V)} = \frac{d}{MW} = \frac{3.4}{23} = 0.148\,M$$

Exercise 4

- Which is the molar concentration of a solution of ammonia (NH_3) such that 700ml of this solution added to 300ml of a solution 0.2M will give a final ammonia solution 0.12M?

This is the case where we are mixing two solutions to obtain a third solution

$$C_3*(V_1+V_2)=(C_1*V_1)+(C_2*V_2)$$

$$C_1=\frac{[C_3*(V_1+V_2)-(C_2*V_2)]}{V_1}=\frac{[(0.12*1)-(0.2*0.3)]}{0.7}=8.57\,10^{-2}\,M$$

Henry's law: solution of gas and liquid

- The mass of a given gas which can be dissolved into a liquid at a fixed T is proportional to the pressure of the gas onto the liquid

$$c=kP_i$$

  c= concentration of the gas in the solution
  k= solubility coefficient of the gas
  P_i= pressure of the gas over the liquid

- In case the liquid is in equilibrium with a gaseous mixture, the partial pressures law is valid: hence the solubility of each gaseous component is proportional to its partial pressure and is independent of the nature of the mixture

Exercise 5

- The mole fraction of N_2 in a gasoues mixture is 0.8. This mixture has a pressure of 3 atm over a liquid underneath. How many ml of N_2 are going to dissolve in the liquid pha-se, given that k_{N2}= 18.2 ml/atm?

We can apply Henry's law, but we need to calculate the partial pressure of N_2 first:

$P_{N2} = x_{N2} P_{tot} = 0.8*3 = 2.4$ atm

$c = kP_i = 18.2*2.4 = 43.68$ ml

Exercise 6

- At 25°C and 1 atm, $1.63 \cdot 10^{-2}$ g of dioxygen are dissolved in 400ml of water. Calculate how much dioxygen will dissolve at 0.3 atm.

We can use a proportion, given that the only changed variable is the pressure:

$$0.0163 \text{ g} : 1 \text{ atm} = x \text{ g} : 0.3 \text{ atm} \quad \rightarrow \quad x = 0.0163 * 0.3 = 4.89\ 10^{-3} \text{ g}$$

Normality

- We have just above defined normality as the ratio between molar concentration C and the equivalence factor f_c.
- There are **three** common areas where normality is used as a measure of reactive species in solution:
  - In **acid-base** chemistry, normality is used to express the concentration of protons (H+) or hydroxide ions (OH-) in a solution. Each solute can produce one or more equivalents of reactive species when dissolved.
  - In **redox reactions**, the equivalence factor describes the number of electrons that an oxidizing or reducing agent can accept or donate.
  - In **precipitation reactions**, the equivalence factor measures the number of ions which will precipitate in a given reaction.

Normality - continued

- $1/f_c$ is an integer number representing:
    - Number of H^+ released by an acid
    - Number of OH^- released by a base
    - Number of e^- exchanged in a redox reaction
    - Number of ions dissociated from an ionic compound

Ex.: $H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$ $1/f_c = 2$

$NaOH \rightarrow Na^+ + OH^-$ $1/f_c = 1$

$Fe_2(SO_4)_3 \rightarrow 2Fe^{+3} + 3SO_4^{2-}$ $1/f_c = 6$

Exercise 1

- Calculate the normality and the molar concentration of a solution of sulphuric acid obtained by dissolving 49g of acid in 1L of water.

$H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$ is a diprotic acid $\rightarrow 1/f_c = 2$

$$C = \frac{g}{(FW * V)} = \frac{49}{(98 * 1)} = 0.5\ M$$

$$N = \frac{C}{f_c} = 0.5 * 2 = 1\ N$$

Exercise 2

- Calculate the normality and the molar concentration of a solution of caustic soda (NaOH) obtained by dissolving 40g of base in 1L of water.

$NaOH \rightarrow Na^{+} + OH^{-}$ is a monobasic base $\rightarrow 1/f_c = 1$

$$C = \frac{g}{(FW * V)} = \frac{40}{(40 * 1)} = 1\,M$$

$$N = \frac{C}{f_c} = 1 * 1 = 1\,N$$

Exercise 3

- 11.72g of sulphuric acid are dissolved in 2L of water. Calculate the normality and the molar concentration of the solution.

$H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$ is a diprotic acid $\rightarrow 1/f_c = 2$

$$C = \frac{g}{(FW * V)} = \frac{11.72}{(98 * 2)} = 5.98\, 10^{-2}\, M$$

$$N = \frac{C}{f_c} = 5.98\, 10^{-2} * 2 = 0.12\, N$$

Exercise 4

- How many grams of KOH are needed to neutralize 100ml of HCl 0.8N?

Neutralization is a chemical reaction where an acid and a base react completely to form a salt.
This is possible if, and only if
Effective concentration of acid = Effective concentration of base
$\mathbf{N_{acid} = N_{base}}$

$[H^+] = NV = 0.8*0.1 = 0.08$ mol equivalent
$[OH^-] = [H^+] = 0.08$

$g = [OH^-]*FW = 0.08*56.1 = 4.49$ g

Exercise 5

- How many grams of $Ba(OH)_2$ are needed to prepare 100ml of 1N solution?

$$N = \frac{C}{f_c} = \frac{n_{eq}}{V} = \frac{n}{(f_c * V)} = \frac{g}{(FW * f_c * V)}$$

$1/f_c = 2$ because $Ba(OH)_2$ is a bibasic base

$$g = N * FW * f_c * V = 1 * 171.3 * (1/2) * 0.1 = 8.56\, g$$

Exercise 6

- 12 g of NaOH are able to neutralize 400ml of HCl. Which is the normality of HCl?

$$N_{NaOH} = \frac{C}{f_c} = \frac{n}{(f_c * V)} = \frac{g}{(FW * f_c * V)} = \frac{12}{(40 * 1 * 0.4)} = 0.75\,N$$

$N_{NaOH} = N_{HCl} = 0.75$ N

Exercise 7

- 50ml of a solution of ammonia (NH_3) 26%w are added to 0.5L of water. Which is the final molar concentration, given that the solution's density is 1.2 g/ml?

First, we need to convert %w to molar concentration, then to dilute the solution.

Hence we need to convert the mass of the solution into a corresponding volume:
26%w means 26 g of NH3 in 100 g of solution
d=g/V → V= g/d = 100/1.2 = 83.33 ml

$$C_1 = \frac{n}{V} = \frac{g}{(FW * V)} = \frac{26 * 1000}{(17 * 83.33)} = 18.35\,M$$

$C_1V_1 = C_2V_2$

$$C_2 = \frac{(C_1 * V_1)}{V_2} = \frac{(18.35 * 0.05)}{0.55} = 1.67\,M$$

Exercise 8

- During a kidney's checkup, the urine of 1h is collected. The measured content of inulin (a kidney marker) is 75mg. In the plasma inulin was 1mg%v. Calculate the volume of plasma filtrated by the kidneys in 1 min.

$[inulin]_{plasma}$ = 1mg%v = 1 mg/100 ml = 0.01 mg/ml

$[inulin]_{urine}$ = 75mg /1h = 75mg/60min = 1.25mg/min

From the ratio we can obtain V/min:

$$\frac{[inulin]_{plasma}}{[inulin]_{urine}} = \frac{1.25\text{mg}/min}{(0.01\text{mg}/ml)} = 125\,ml/min$$