Lecture 6

Concentrations

Composition of Solutions

Mass of a solute per mass of solvent or solution

- Mole fraction and Mole percent,
- Mass fraction and Mass percent,
- Molality

Mass of a solute per Volume of solution

- Molarity
- Normality

Mole fraction

- *The ratio of the no. of moles of a given component in a solution to the total no. of moles of all components*

Unitless

$$X_A = \frac{n_A}{n_A + n_B + n_C + \dots + n_Z}$$

$$\boxed{\sum X_i = 1}$$

Mole Percent *is the mole fraction multiplied by 100*

Mass fraction

- *The ratio of the mass of a given component in a solution to the total mass of all components*

$$(\text{mass fraction})_A = \frac{m_A}{m_A + m_B + m_C + \dots + m_Z}$$

Unitless

$$\boxed{\sum (\text{mass fraction})_i = 1}$$

Mass Percent *is the mass fraction multiplied by 100*

A solution labeled "0.9 % NaCl" means 0.9 g NaCl is dissolved in 99.1 g H_2O or 100 g of solution

Molality (Molal concentration)

The no. of moles of solute dissolved in 1 kg of solvent

$$\boxed{m = \frac{n_{solute}}{mass\ of\ solvent\ (kg)}}$$

$$\boxed{m = \frac{w_{solute}}{Mwt_{solute}} \times \frac{1000}{w_{solvent}\ (g)}}$$

n_{solute} : moles of solute

Mwt_{solute} : molar mass of solute

W_{solute} : mass of solute

$W_{solvent}$: mass of solvent

Molarity (Molar concentration)

- *The no. of moles of solute dissolved in 1 liter of solution*

$$\boxed{M = \frac{n_{solute}}{Volume\ of\ solution\ (L)} = \frac{w_{solute}}{Mwt_{solute} \times V\ (L)}}$$

n_{solute} : no. of moles of solute

Mwt_{solute} : molar mass of solute

W_{solute} : mass of solute

V : Volume of solution in liter

$$\boxed{w_{solute}(g) = M(mol/L) \times Mwt_{solute}(g/mol) \times V\ (L)}$$

Normality

- The number of equivalents of solute per liter of solution

$$No.\,Equivalent = \frac{mass_{solute}(g)}{Equivalent\ mass_{solute}\ (g/equivalent)}$$

$$= \frac{w_{solute}}{Eq.wt_{solute}}$$

The **definition** of the **"equivalent mass"** and **"equivalent"** depends on the reaction taking place in the solution.

Equivalent mass

Acid–base reactions

$$Eq.wt = \frac{molar\ mass}{No\ of\ replacable\ H^{+} or\ OH^{-}}$$

An **"equivalent"** is the mass of acid or base that can furnish or accept exactly 1 mole of protons (H^+ ions).

Acid/Base	Mwt, g/mol	Equivalent mass	M/N relationship
HCl	36.5	36.5	1M = 1N
H_2SO_4	98	98/2 = 49	1M = 2N
NaOH	40	40	1M = 1N
$Ca(OH)_2$	74	74/2 = 37	1M = 2N

Equivalent mass

Oxidation–reduction reactions

The "equivalent" is defined as the quantity of oxidizing or reducing agent that can accept or furnish 1 mole of electrons.

$$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$$

For example, MnO_4^- reacts in acidic solution absorbing five electrons to produce Mn^{2+}:

$$\text{Equivalent mass of } KMnO_4 = \frac{\text{molar mass}}{5} = \frac{158\,g}{5} = 31.6\,g$$

Equivalent mass

Salts reactions

$$\text{Equivalent mass of a salt} = \frac{\text{molar mass}}{\text{no. of ions} \times \text{its valency}}$$

NaCl: Eq. wt = MWt/1

Na_3PO_4: Eq.Wt = Mwt / 3

Normality

$$N = \frac{no.\,equivalents_{solute}}{Volume\; of\; solution\;(L)} = \frac{w_{solute}}{Eq.wt_{solute} \times V\;(L)}$$

n_{solute} : no. of moles of solute
$Eq.wt_{solute}$: Equivalent mass of solute

W_{solute} : mass of solute
V: Volume of solution in liter

$$w_{solute}(g) = N\;(equivelents/L) \times Eq.wt_{solute}(g/equivalents) \times V\;(L)$$

Exercise

A solution is prepared by mixing 1.00 g ethanol (C_2H_5OH) with 100.0 g water to give a final volume of 101 mL. Calculate the molarity, mass percent, mole fraction, and molality of ethanol in this solution.

Solution

molar mass of ethanol = 46.07 g/mol

$$M = \frac{w_{solute}}{Mwt_{solute} \times V\;(L)}$$

$$= \frac{1.0\;g}{46.07\;g\;mol^{-1} \times 101\;mL} \times \frac{1000\;mL}{1\;L} = 0.215\;mol/L$$

$$= 0.215\;M$$

$$mass\ percent\ of\ ethanol = \frac{m_{ethanol}}{m_{ethanol} + m_{water}} \times 100$$

$$= \frac{1.0}{1.0 + 100.0} \times 100 = 0.990\,\%$$

$$X_{ethanol} = \frac{n_{ethanol}}{n_{ethanol} + n_{water}} =$$

$$\frac{1.0\,g \big/ 46.07\,g\,mol^{-1}}{1.0\,g \big/ 46.07\,g\,mol^{-1} + 100.0\,g \big/ 18.0\,g\,mol^{-1}}$$

$$= 0.00389$$

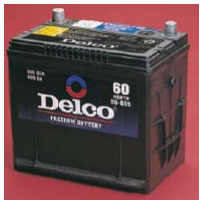
$$m = \frac{w_{ethanol}}{Mwt_{etanol}} \times \frac{1000}{w_{water}\,(g)}$$

$$= \frac{1.0\,g}{46.07\,g\,mol^{-1}} \times \frac{1000}{100.0\,(g)} = 0.217\ mol\,kg^{-1} = 0.217\ m$$

Exercise

molar mass of sulfuric acid= 98 g/mol

The electrolyte in automobile lead storage batteries is a 3.75 M sulfuric acid solution that has a density of 1.230 g/mL. Calculate the mass percent, molality, and normality of the sulfuric acid?



Solution

- *d = 1.230 g/mL = 1230 g/L → means we have 1230 g solution ($H_2SO_4 + H_2O$) in every liter of solution ($H_2SO_4 + H_2O$)*

- *Molarity of H_2SO_4 = 3.75 M → means we have 3.75 mole (368 g) of H_2SO_4 in every liter of solution ($H_2SO_4 + H_2O$)*

1L Solution

- 1230 g $H_2SO_4 + H_2O$
- 368 g H_2SO_4

$$mass\ of\ water\ in\ 1L\ solution = 1230 - 368 = 862\ g$$

$$mass\ percent\ of\ ethanol = \frac{m_{sulfuric\ acid}}{m_{sulfuric\ acid} + m_{water}} \times 100 = \frac{368}{368 + 862} \times 100 = 29.9\ \%$$

$$m = \frac{w_{sufuric\ acid}}{Mwt_{sufuric\ acid}} \times \frac{1000}{w_{water}\ (g)}$$

$$= \frac{368\ g}{98\ g\ mol^{-1}} \times \frac{1000}{862\ (g)} = 4.35\ m$$

$$N = M \times 2 = 3.75 \times 2$$

$$= 7.5\ equivalents\ per\ liter\ (N)$$