ALMUSTAQBAL UNIVERSITY

College of Engineering and Engineering Techniques

Stage : Second year students

Subject : Chemistry 1 - Lecture 4

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Normality (N)

Represents the number of equivalents contained in one liter solution or the number of milli equivalents of solute contained in one milliliter of solution .

e.g: 0.2 N HCl solution contains 0.2 equivalents (eq) of HCl in liter solution or 0.2 milli equivalent (meq) of HCl in each mL of solution .

Normality (N) $= \frac{\textbf{number of equivalents}(\textbf{solute})}{\textbf{VL}(\textbf{solution})}$

Number of equivalents(eq) $= \frac{\textbf{wt }(\textbf{g})}{\textbf{eq.wt}(\textbf{g})}$

Normality (N) $= \frac{\frac{\textbf{wt}}{\textbf{eq.wt}}}{\textbf{V}(\textbf{liter})}$

Normality (N) $= \frac{\frac{\textbf{wt}}{\textbf{eq.wt}}}{\frac{\textbf{V}(\textbf{mL})}{\textbf{1000}}}$

Normality (N) $= \frac{\textbf{wt x 1000}}{\textbf{eq.wt x V}(\textbf{mL})}$

Exercise: proof that **Normality (N)** $= \frac{wt\ x\ 1000}{eq.wt\ \ x\ \ V(mL)}$

الجواب : نكتب الاشتقاق (الخطوات الاربعه اعلاه)

Eq.wt = $\frac{Mwt}{\eta}$

Normality (N) = $\frac{wt\ x\ 1000}{\frac{Mwt}{\eta}\ x\ V(mL)}$

Normality (N) = $\boldsymbol{\frac{wt\ x\ 1000}{\frac{Mwt\ xV(mL)}{\eta}}}$

Normality (N) $=$ $\boldsymbol{(\frac{wt\ x1000}{Mwt\ x\ V(mL)})\ \eta}$

Normality (N) = Molarity (M) . η , or Molarity(M) = Normality(N) / η

I. Equivalent mass in neutralization reaction:

A) Equivalent mass of acids (Eq):-

Is the mass that either contribute or reacts with one mole of hydrogen ion in the reaction.

Eq = $\frac{Mwt}{number\ of\ H}$

1. Monoprotic acid e.g: (HCl , HNO_3 , CH_3COOH) η=1

Eq = $\frac{Mwt}{1}$

Eq = $\frac{36.5}{1} = 36.5\ for\ HCl$

$Eq = \frac{63}{1} = 63 \; for \; HNO_3$

2. Diprotic acid e.g: (H_2SO_4, H_2S, H_2SO_3) η= 2

Eq = $\frac{Mwt}{2} = \frac{98}{2} = 49$ for H_2SO_4

Eq = $\frac{34}{2} = 17$ for H_2S

Eq = $\frac{82}{2} = 41$ for H_2SO_3

B) Equivalent mass of Bases:

Is the mass that either contribute or reacts with one mole of OH in the reaction.

Eq = $\frac{Mwt}{number\ of\ OH}$

1. Monohydroxy base e.g: (η=1)

e.g: NaOH (40 g/mole)

Eq. = $\frac{Mwt}{1} = \frac{40}{1}$ =40

e.g: KOH (56 g/mole)

Eq. = $\frac{Mwt}{1} = \frac{56}{1}$ = 56

2. Dihydroxy base (η=2)

e.g: $Ca(OH)_2$ (74 g / mole)

Eq. = $\frac{Mwt}{2} = \frac{74}{2}$ = 37

$Zn(OH)_2$ (99.4 g /mole)

Eq. = $\frac{Mwt}{2} = \frac{99.4}{2}$ =49.7

$Ba(OH)_2$ (171.35 g / mole)

Eq. = $\frac{Mwt}{2} = \frac{171.35}{2} = 85.67$

II. Equivalent mass in (oxidation – reduction) reaction (Redox):

The equivalent mass of a participant in an (oxidation–reduction) reaction is that mass which directly produce or consume one mole of electron.

Eq = $\frac{Mwt}{\eta}$

η= change in oxidation state number

η= numbers of electrons participate in oxidation - reduction processes (Redox)

Example :

$2KMnO_4 + 10FeSO_4 + 8H_2SO_4 \rightarrow 5Fe_2(SO_4)_3 + 2MnSO_4 + K_2SO_4 + 8H_2O$

$\mathbf{2MnO_4^- + 10Fe^{2+} + 8H^+ \rightleftharpoons 10Fe^{3+} + 2MnSO_4}$ **(acidic medium)**

Mn^{7+} → Mn^{2+} (5 e gain – reduction)

Fe^{2+} → Fe^{3+} (1 e loss – oxidation)

Eq. of $KMnO_4 = \frac{Mwt}{5} = \frac{157.9}{5} = 31.6$

III. Equivalent mass for salts:

Eq= $\frac{Mwt}{\eta}$

(η) = Σ [no. of cations x its valency(cation charge)]

e.g: $AgNO_3$ (170 g/mole)

($AgNO_3 \rightarrow Ag+ + NO_3^-$)

(η = Ag^+ (1) x 1= 1)

Eq. = $\frac{\mathbf{Mwt}}{\mathbf{1}} = \frac{\mathbf{170}}{\mathbf{1}}$ **=170**

e.g: Na_2CO_3 (106 g/mole)

($Na_2CO_3 \rightarrow 2\ Na^+ + CO_3^-$)

(η= Na^+ (2) x 1= 2)

Eq. = $\frac{Mwt}{2} = \frac{106}{2}$ =53

e.g: **$BaSO_4$ (233 g/mole)**

($Ba^{2+} + SO_4^{2-} \leftrightarrow BaSO_4$)

η = Ba^{2+} (1) x (2+) =2

Eq. = $\frac{Mwt}{2} = \frac{233}{2}$ =116.5

e.g: $La(IO_3)_3$ (663.6 g/mole)

($La(IO_3)_3 \rightarrow La^{3+} + 3\ IO_3^-$)

(η = La^{3+} (1) x 3 = 3)

Eq. = $\frac{Mwt}{3} = \frac{663.6}{3}$ =221.1

e.g: **$KAl(SO_4)_2$ (258 g/mole)**

(**η) = Σ [no. of cations x its valency(cation charge)]**

no. of cations = 1 K^+ + 1 Al^{3+}

η= K^+ (1) x (1+) **+** Al^{3+}(1) x (3+)= 4

Eq. =$\frac{M.wt}{4} = \frac{258}{4}$ = 64.5

Example

Find the Normality of the solution containing 5.3 g/L of Na_2CO_3 (106 g/mol).

Solution:

To find η for Na_2CO_3 **(η) = Σ [no. of cations x its valency(cation charge)]**

No. of cations =2Na+ while the cation charge for Na^+ =1 ,

Then (η) = 2 x 1 = 2

Eq. of Na_2CO_3 = $\frac{\boldsymbol{Mwt}}{\mathbf{2}} = \frac{\mathbf{106}}{\mathbf{2}}$ **=53.0 gm**

Normality (N) = $\frac{wt}{Eq.\ x\ VL}$

Normality (N) = $\frac{5.3\ g}{53.0\ x\ 1L}$ = 0.1N

Second method:

Normality (N) = $\left(\frac{\boldsymbol{wt\ x1000}}{\boldsymbol{Mwt\ x\ V(mL)}}\right)$ η

Normality (N) = $\left(\frac{\mathbf{5.3\ x1000}}{\mathbf{106\ x\ 1000(mL)}}\right)$ 2 = 0.1 N

Example;

Convert the following Molarities to Normalities.

a. 2.5 M HCl b. 1.4 M H_2SO_4 c. 1.0 M NaOH d. 0.5 M $Ca(OH)_2$

Answer:

a. Normality (N) of 2.5M HCl = M . η =2.5 x 1 = 2.5 N HCl,

b. Normality (N) of 1.4 M H_2SO_4 = M . η = 1.4 x 2 = 2.8 N H_2SO_4

c. Normality (N) of 1M NaOH= M . η = 1 x 1 = 1 N NaOH

d. Normality (N) of 0.5 M $Ca(OH)_2$ = M . η = 0.5 x 2 = 1 N $Ca(OH)_2$

Molality(m):

The number of moles of solute per kilogram of solvent.

انتبه هنا استخدم وزن المذيب وليس المحلول

(المولاليه =عدد مولات المذاب في الكيلوغرام من المذيب)

المذيب = solvent و المحلول = solution والمذاب = Solute

Molality(m) = $\frac{number\ of\ moles(solute)}{mass\ \ of\ solvent\ (Kg)}$

Molality(m) = $\frac{number\ of\ moles(solute)}{mass\ \ of\ solvent\ (\frac{g}{1000})} = \frac{number\ of\ moles(solute)x\ 1000}{mass\ of\ solvent(g)}$

Example :

Determine the molality of a solution prepared by dissolving 75 g of solid $Ba(NO_3)_2$ (261.32 g/mole) into 374 g of water.

<u>Solution:</u>

Molality(m) = $\frac{number\ of\ moles(solute)x\ 1000}{mass\ of\ solvent\ (g)}$

No of moles(solute) = $\frac{wt}{M.wt} = \frac{75\ g}{261.32\ g\ /mol}$ **= 0.287 moles**

Molality(m) = $\frac{number\ of\ moles(solute)x\ 1000}{mass\ of\ solvent(g)} = \frac{0.287\ mol\ x\ 1000}{374\ \ g}$

Molality(m) = 0.76

Example:

The mass of an aqueous solution that contains 11.7 g of NaCl (58.5 g/mole) is 551.7 g . Calculate the molality of the solution.

Solution ;

Mass of solution = mass of solute + mass of solvent

Mass of solution = mass of solute (NaCl) + mass of solvent (H_2O)

Mass of solvent (H_2O) = Mass of solution - mass of solute(NaCl)

Mass of solvent (H_2O) = 551.7 g – 11.7 g = 540 g

No . of moles of NaCl = $\frac{mass(\ g)}{M.wt}$

No . of moles of NaCl = $\frac{11.7}{58.5}$ **= 0.2 mole**

Molality (m) = $\frac{number\ of\ moles(solute)x\ 1000}{mass\ of\ solvent(g)}$

Molality (m) = $\frac{0.2\ mol\ x\ 1000}{540\ \ g}$ **= 0.37**

Exercise:

7.45 g of potassium chloride KCl (74.5 g/ mole) was dissolved in 100 g of water. Calculate the molality of the solution.

Example :

A 11.11 g of urea NH_2CONH_2 (60 g/ mole) was dissolved in 100 g of water (d= 1 g /mL). Calculate the molarity and molality of the solution.

Solution:

No . of moles of urea = $\frac{mass(\ g)}{M.wt}$

No . of moles of urea = $\frac{11.11(\ g)}{60}$ **= 0.1852 mole**

Volume of water = $\frac{mass\ of\ water(g)}{density\ (\frac{g}{mL})}$

Volume of water = $\frac{100(g)}{1\ (\frac{g}{mL})}$ **= 100 mL = 0.1 L**

Molarity (M) = $\frac{\text{No .of moles of urea}}{\text{Volume of water }(L)}$

Molarity (M) = $\frac{0.1852}{0.1\ (L)}$ **= 1.852 M**

Or Molarity (M) = $\frac{wt\ x\ 1000}{Mwt\ x\ V(mL)}$ **=** $\frac{11.11\ x\ 1000}{60\ x\ 100(mL)}$ **= 1.852 M**

Molality (m) = $\frac{number\ of\ moles(solute)x\ 1000}{mass\ of\ solvent(g)}$

Density of water = 1 g /mL

Mass of solvent (H_2O) = 1 g /mL x volume of water (mL)= 1x100 = 100 g

Molality (m) = $\frac{0.1825\ molx1000}{100\ g}$ **= 1.852**

Example:
The weight of 10 g KCl (74.5 g / mol) is dissolved in 1000 g of water. If the density of the prepared solution is 0.997 g mL^{-1}, calculate :
a) molarity and b) molality of the solution.

Solution:

Molarity = $\frac{wt\ x1000}{Mwt\ x\ V(mL)}$

Mass of solution = 10 g + 1000 g = 1010 g

Volume of solution = $\frac{mass\ of\ solution}{density}$ **=** $\frac{1010\ g}{0.997}$ **= 1013 mL**

$$\text{Molarity} = \frac{10 \ x \ 1000}{74.5 \ x \ 1013(mL)} = 0.1325 \text{ M}$$

$$\text{Molality(m)} = \frac{number \ of \ moles(solute)x1000}{mass \ \ of \ solvent \ (\ g)}$$

$$\text{No . of moles of KCl} = \frac{mass(\ g)}{M.wt}$$

$$\text{No . of moles of KCl} = \frac{10}{74.5} = 0.1342$$

$$\text{Molality(m)} = \frac{0.1342 \ x1000}{1000} = 0.1342$$

Mole fraction:

The number of moles of one component relative to the total number of moles of all components in the solution.

$$\text{Mole fraction of solute}(X_1) = \frac{\text{No.of moles of solute } (n_1)}{\text{mole of solute } (n_1) + \text{moles of solvent } (n_2)}$$

$$\text{Mole fraction of solvent}(X_2) = \frac{\text{No.of moles of solvent } (n_2)}{\text{moles of solute } (n_1) + \text{moles of solvent } (n_2 \ )}$$

1 = مجموع الكسور الموليه في المحلول

$$X_T = \sum X_i = 1$$

$$\boxed{X_1 + X_2 = 1}$$

Then $\boxed{X_1 = 1 - X_2}$

and $\boxed{X_2 = 1 - X_1}$

<u>Example:</u> calculate the mole fraction for each of solute and solvent in a solution if the solute is (2 mole) and the solvent in (3 mole) .

Solution:

$$X_1 = \frac{n_1}{n_1+n_2} = \frac{2}{2+3} = \frac{2}{5} = 0.4$$

$$X_2 = \frac{n_2}{n_1+n_2} = \frac{3}{2+3} = \frac{3}{5} = 0.6$$

$X_1 + X_2 = 0.4 + 0.6 = 1$

For 3 components mixture we have X_1 , X_2 , and X_3 Then:

$$X_1 = \frac{n1}{n1+n2+n3}$$

$$X_2 = \frac{n2}{n1+n2+n3}$$

$$X_3 = \frac{n3}{n1+n2+n3}$$

<u>Example:</u> Calculate the mole fraction for each component in a mixture that contains 1mole of A , 2 moles of B and 3 moles of C .

Total no of moles n_T = moles of A (n_A) + moles of B (n_B) + moles of C (n_C)

$n_T = n_A + n_B + n_C$

$n_T = 1 + 2 + 3 = 6$ **moles**

$$X_A = \frac{n_A}{n_T} = \frac{1}{6} = 0.17$$

$$X_B = \frac{n_B}{n_T} = \frac{2}{6} = 0.33$$

$$X_C = \frac{n_C}{n_T} = \frac{3}{6} = 0.5$$

$X_T = \sum X_i = 0.17 + 0.33 + 0.5 = 1$

Example:

A 4.6 mL of methanol (32 g/mol , d= 0.7952 g/mL) is dissolved in 25.2 g of water(18 g/mol). Calculate the mole fraction of methanol and water. Solution:

Mass of methanol (g)= Volume x density
Mass of methanol (g) = 4.6 mL x 0.7952 g mL^{-1} = 3.658 g

No . of moles of methanol = $\frac{mass(\ g)}{M.wt}$

No . of moles of methanol (n_1)= $\frac{3.658(\ g)}{32}$ **=0.1143**

No . of moles of water(n_2) = $\frac{25.2(\ g)}{18}$ **=1.4**

Total number of moles = n_1+ n_2 = 0.1143 + 1.4 = 1.5143 mol
Mole fraction of methanol (X_1) = $\frac{n_1}{n_1+n_2}$
Mole fraction of methanol (X_1) = $\frac{0.1143}{1.5143}$ **= 0.0755**

Mole fraction of water(X_2) = $\frac{n_2}{n_1+n_2}$

Mole fraction of solvent (water) = X_2 = $\frac{1.4}{1.5143}$ **= 0.9245**

Exercise:

The mass of an aqueous solution that contains 10.1 g of KNO_3 (101 g/mol) is

154.1 g . Calculate :

1. The molality of the solution.

2. The mole fraction of each of the solute(KNO_3) and solvent (H_2O)(18 g/mol).