Review of Fundamental Concepts

Formula Weight

It is assumed that you can calculate the formula or molecular weights of compounds from respective atomic weights of the elements forming these compounds. The formula weight (FW) of a substance is the sum of the atomic weights of the elements from which this substance is formed from.

The formula weight of CaSO$_4$.7H$_2$O is

<table>
<thead>
<tr>
<th>Element</th>
<th>Atomic weight</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ca</td>
<td>40.08</td>
</tr>
<tr>
<td>S</td>
<td>32.06</td>
</tr>
<tr>
<td>11 O</td>
<td>11x16.00</td>
</tr>
<tr>
<td>14 H</td>
<td>14x1.00</td>
</tr>
<tr>
<td>FWt</td>
<td>$\Sigma = 262.14$</td>
</tr>
</tbody>
</table>

The Mole

The mole is the major word we will use throughout the course. The mole is defined as gram molecular weight which means that:

<table>
<thead>
<tr>
<th>Mole</th>
<th>Grams</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 mol H$_2$</td>
<td>2.00 g</td>
</tr>
<tr>
<td>1 mol O$_2$</td>
<td>32.00 g</td>
</tr>
<tr>
<td>1 mol O</td>
<td>16.00 g</td>
</tr>
<tr>
<td>1 mol NaCl</td>
<td>58.5 g</td>
</tr>
<tr>
<td>1 mol Na$_2$CO$_3$</td>
<td>106.00 g</td>
</tr>
</tbody>
</table>

Assuming approximate atomic weights of 1.00, 16.00, 23.00, 35.5, and 12.00 atomic mass units for hydrogen, oxygen, sodium, chlorine atom, and carbon, respectively.

The number of moles contained in a specific mass of a substance can be calculated as:

$$\text{mol} = \frac{\text{g substance}}{\text{FW substance}}$$
The unit for the formula weight is g/mol.

In the same manner, the number of mmol of a substance contained in a specific weight of the substance can be calculated as:

\[ \text{mmol} = \frac{\text{mol}}{1000} \]

Or,

\[ \text{mmol} = \frac{\text{mg substance}}{\text{FW substance}} \]

**Look at this calculation**

The number of mmol of Na₂WO₄ (FW = 293.8 mg/mmol) present in 500 mg of Na₂WO₄ can be calculated as:

\[ ? \text{ mmol of Na₂WO₄} = \frac{500 \text{ mg}}{293.8 \text{ mg/mmol}} = 1.70 \text{ mmol} \]

The number of mg contained in 0.25 mmol of Fe₂O₃ (FW = 159.7 mg/mmol) can be calculated as:

\[ ? \text{ mg Fe₂O₃} = 0.25 \text{ mmol Fe₂O₃} \times 159.7 \text{ mg/mmol} = 39.9 \text{ mg} \]

Therefore, either the number of mg of a substance can be obtained from its mmols or vice versa.

**Methods of Expressing the Concentration of a Solution**

Concentration of solution is the amount of solute dissolved in a known amount of the solvent or solution. The concentration of solution can be expressed in various ways as discussed below,

1. **Percentage**: It refers to the amount of the solute per 100 parts of the solution. It can also be called as parts per hundred (pph). It can be expressed by any of following four methods,
   1. **Weight to weight percent**
      \[ \% \ W/W = \frac{\text{Wt. of solute}}{\text{Wt. of solution}} \times 100 \]
      
      Example: 10% Na₂CO₃ solution w/w means 10g of Na₂CO₃ is dissolved in 100g of the solution. (It means 10g of Na₂CO₃ is dissolved in 90g of H₂O)
   2. **Weight to volume percent**
      \[ \% \ W/V = \frac{\text{Wt. of solute}}{\text{Volume of solution}} \times 100 \]
      
      Example: 10% Na₂CO₃ (w/v) means 10g of Na₂CO₃ is dissolved in 100ml of solution.
(iii) **Volume to volume percent**

\[ \% \text{ v/v} = \frac{\text{Vol. of solute}}{\text{Vol. of solution}} \times 100 \]

Example: 10% ethanol (v/v) means 10 cc of ethanol dissolved in 100 cc of solution.

(iv) **Volume to weight percent**

\[ \% \text{ v/w} = \frac{\text{Vol. of solute}}{\text{Wt. of solution}} \times 100 \]

Example: 10% ethanol (v/w) means 10 cc of ethanol dissolved in 100 g of solution.

**Mass/Weight Percentage or Per cent by Mass/Weight : (W/W)%**

It is defined as the amount of solute in grams present in 100 grams of the solution.

\[
\text{Mass Percentage} = \frac{\text{Mass of Solute}}{\text{Mass of Solution}} \times 100
\]

\[
= \frac{\text{Mass of Solute}}{\text{Mass of Solute} + \text{Mass of Solvent}} \times 100
\]

\[
= \frac{\text{Volume of Solution} + \text{Density of Solution}}{\text{Volume of Solution}} \times 100
\]

- The ratio mass of solute to the mass of solvent is termed as **mass fraction**.
- Thus, Mass percentage of solute = Mass fraction × 100
- 10% solution of sugar by mass means that 10 grams of sugar is present in 100 grams of the solution, i.e., 10 grams of sugar has been dissolved in 90 grams of water.

**Example 1:**

**Question:**

What is the weight percentage of urea solution in which 10 gm of urea is dissolved in 90 gm water.

**Solution**

Weight percentage of urea = (weight of urea/ weight of solution) × 100

\[ = \frac{10}{90+10} \times 100 = 10\% \text{ urea solution (w/W)} \]

**Volume Percentage : (V/V)%**

- It is defined as the volume of solute in mL present in 100 mL solution.

\[
\text{Volume Percentage} = \frac{\text{Volume of Solute}}{\text{Volume of Solution}} \times 100
\]
10% solution of HCl by volume means that 10 mL of liquid HCl is present in 100 mL of the solution.

**Mass by Volume Percentage (W/V)%**

- It is defined as the mass of solute present in 100 mL of solution.

\[
\text{Mass by Volume Percentage} = \frac{\text{Mass of Solute}}{\text{Volume of Solution}} \times 100
\]

- A 10% mass by volume solution means that 10 gm solute is present in 100 mL of solution.

**Molarity (M)**

- The molarity of a solution gives the number of gram molecules of the solute present in one litre of the solution.

\[
\text{Molarity (M)} = \frac{\text{Number of moles of solute}}{\text{Volume of Solution in L}}
\]

**Example 2:**

**Question:**

3.65 gm of HCL gas is present in 100 mL of its aqueous solution. What is the molarity?

**Solution**

\[
M = \frac{\text{mass of solute}}{\text{volume of solution in L}} = \frac{3.65}{1000} = 0.00365 \text{ M}
\]

**Molality (m)**

- Molality of a solution is defined as the number of moles of solute dissolved in 1 Kg of the solvent.

\[
\text{Molality (m)} = \frac{\text{Number of moles of solute}}{\text{Mass of Solvent in kg}}
\]
Lecture One  Concentration of the Solutions  Asst. Prof. Dr. Azhar A. Ghali

- Thus, if one gram molecule of a solute is present in 1 kg of the solvent, the concentration of the solution is said to be one molal.

- Units of molarity: mol kg\(^{-1}\)
- Molality is the most convenient method to express the concentration because it involves the mass of liquids rather than their volumes. It is also independent of the variation in temperature.
- Molality and solubity are related by the following relation.

\[
\text{Molality} = \frac{\text{Solubility} \times 10}{\text{Molecular mass of the solute}}
\]

[Solubility = Mass of solute in grams/Mass of solvent in grams \times 100]

**Relationship Between Molality and Molarity:**

\[
\frac{Molarity}{\text{Molality}} = \frac{\text{Number of moles of solute} \times \text{Mass of solvent in Kg}}{\text{Volume of solution in L} \times \text{Number of moles of solute}}
\]

Let the density of the solution be d.  Unit= g mL\(^{-1}\)

Mass of solution = V \times d

Mass of solute = number of moles \times \text{molecular mass of solute} = n m_A

Mass of solvent, W = mass of solution – mass of solute = V \times d – n \times m_A

Thus,

\[
\frac{Molarity}{\text{Molality}} = \frac{V \times d - n \times m_A}{V}
\]

Where m_A is molecular mass of solvent.

**Normality: (N)**

- The normality of a solution gives the number of gram equivalents of the solute present in one litre of the solution.

\[
\text{Normality (N)} = \frac{\text{Number of gram equivalents of solute}}{\text{Volume of Solution in L}}
\]

Number of gram equivalent of solute = Mass of solute in gram/ equivalent weight of solute

Equivalent weight of solute (E) = Molar mass of solute/ Valence factor
- Valence factor for base = acidity of base
- Valence factor for acid = basicity of acid
- Valence factor for element = valency

An equivalent is defined as the weight of substance giving an Avogadro’s number of reacting units. Reacting units are either protons (in acid base reactions) or electrons (in oxidation reduction reactions). For example, HCl has one reacting
unit (H⁺) when reacting with a base like NaOH but sulfuric acid has two reacting units (two protons) when reacting completely with a base. Therefore, we say that the equivalent weight of HCl is equal to its formula weight and the equivalent weight of sulfuric acid is one half its formula weight. In the reaction where Mn(VII), in KMnO₄, is reduced to Mn(II) five electrons are involved and the equivalent weight of KMnO₄ is equal to its formula weight divided by 5.

**Example**

Find the normality of the solution containing 5.300 g/L of Na₂CO₃ (FW = 105.99), carbonate reacts with two protons.

**Solution**

Normality is the number of equivalents per liter, therefore we first find the number of equivalents

\[
eq w t = \frac{F W}{2} = \frac{105.99}{2} = 53.00
\]

\[
eq = \frac{W t}{\text{eq wt}} = \frac{5.300}{53.00} = 0.1000
\]

\[
N = \frac{\text{eq/L}}{1 \text{L}} = 0.1000 \text{ N}
\]
The problem can be worked out simply as below:

\[ \text{eq Na}_2\text{CO}_3 /L = (5.300 \text{ g Na}_2\text{CO}_3 /L) \times (1 \text{ mol Na}_2\text{CO}_3/105.99 \text{ g Na}_2\text{CO}_3) \times \]
\[ (2 \text{ eq Na}_2\text{CO}_3/1 \text{ mol Na}_2\text{CO}_3) = 0.1 \text{ N} \]

The other choice is to find the molarity first and the convert it to normality using the relation

\[ N = n \times M \]

No of mol = \( \frac{5.300 \text{ g}}{(105.99 \text{ g/mol})} \)

\[ M = \frac{\text{mol/L}}{1L} = \frac{\left[\frac{5.300 \text{ g}}{(105.99 \text{ g/mol})}\right]}{1L} \]

\[ N = n \times M = 2 \times \frac{\left[\frac{5.300 \text{ g}}{(105.99 \text{ g/mol})}\right]}{1L} = 0.1000 \]

A further option is to find the number of moles first followed by multiplying the result by 2 to obtain the number of equivalents.

### Mole Fraction
- The mole fraction of any component in a solution is the ratio of the number of moles of that component to the total number of moles of all components.
- Total mole fraction of all the components of any solution is 1.
- For a binary solution of A and B

**Mole Fraction of A (X_A) =**
\[ \frac{n_A}{n_A + n_B} \]

**Mole Fraction of B (X_B) =**
\[ \frac{n_B}{n_A + n_B} \]

And, \( X_A + X_B = 1 \)

### Relation between mole fraction and Molality:

\[ X_A = \frac{n}{N+n} \text{ and } X_B = \frac{N}{N+n} \]

\[ X_A/X_B = \frac{n}{N} = \text{Moles of solute/Moles of solvent} = \frac{w_A}{m_B} = \frac{w_A}{w_B \times m_A} \]

\[ X_A \times 1000/X_B \times m_B = \frac{w_A \times 1000}{w_B \times m_A} = m \]

or
Parts per million (ppm):

- When a solute is present in trace quantities, it is convenient to express concentration in parts per million.
  \[ \text{ppm} = \frac{\text{Number of parts of the component}}{\text{Total number of parts of the components in the solution}} \times 10^6 \]
- In case of mass it may be expressed as: \((\text{Mass of solute}/\text{Mass of solution}) \times 10^6\)

- In case of volume it may be expressed as: \((\text{Volume of solute}/\text{Volume of solution}) \times 10^6\)
- So, concentration in parts per million can be expressed as mass to mass, volume to volume, and mass to volume form.
- Atmospheric pollution in cities is also expressed in ppm by volume. It refers to the volume of the pollutant in \(10^6\) units of volume. 10 ppm of \(\text{SO}_2\) in air means 10 mL of \(\text{SO}_2\) is present in \(10^6\) mL of air.

(2) **Parts per million (ppm) and parts per billion (ppb)**: When a solute is present in trace quantities, it is convenient to express the concentration in parts per million and parts per billion. It is the number of parts of solute per million \((10^4)\) or per billion \((10^5)\) parts of the solution. It is independent of the temperature.

\[
\begin{align*}
\text{ppm} &= \frac{\text{mass of solute component}}{\text{Total mass of solution}} \times 10^4 \\
\text{ppb} &= \frac{\text{mass of solute component}}{\text{Total mass of solution}} \times 10^8
\end{align*}
\]