Al-Mustaqbal-College University

Chemical Engineering and Petroleum Industry Department

Analytical chemistry

First class / first term

Lecture Two part 1

By

Asst. lect. Ban Ali Hassan

Lecture Two

⇒ Examples

Example 1 : Calculate the weight of one mole of $CaSO_4 \cdot 7H_2O$. ?

Solve:

One mole is the formula weight expressed in grams. The atomic weight is :

Ca= 40.08 , S =32.06 , O =16 , H = 1

M.wt. = 40.08*1+32.06*1+ 16*4 + 7*2*1 +7*16

= 262.25 g/mol

Moles = $\frac{\text{weight grams}}{\text{formula weight (g/mol)}}$

1 mol = $\frac{\text{weight (g)}}{262.25\ \text{g/mol}}$

Weight (g) = 1 mol * 262.25 g/mol

= 262.25 g .

Example 2: Calculate the number of moles in 500 mg Na_2WO_4 (sodium tungstate).

Mol = $\frac{500\ \text{mg}}{293.8\frac{mg}{mmol}}$ **× 0.001 mol/mmol**

= 0.00170 mol

Lecture Two

⇒Molarity

The mole concept is useful in expressing concentrations of solutions, especially in analytical chemistry, where we need to know the volume ratios in which solutions of different materials will react. A one-molar solution is defined as one that contains one mole of substance in each liter of a solution,
It is prepared by dissolving one mole of the substance in the solvent and diluting to a final volume of one liter in a volumetric flask; or a fraction or multiple of the mole may be dissolved and diluted to the corresponding fraction or multiple of a liter. Molar is abbreviated as *M* .

$$\text{Moles} = (\text{moles/liter}) \times \text{liters}$$
$$= \text{molarity} \times \text{liters}$$

Example

A solution is prepared by dissolving 1.26 g AgNO3 in a 250-mL volumetric flask and diluting to volume. Calculate the molarity of the silver nitrate solution. How many mill moles AgNO3 were dissolved?

Solution

$$M = \frac{1.26\text{ g}/169.9\text{ g/mol}}{0.250\text{ L}} = 0.0297\text{ mol/L (or 0.0297 mmol/mL)}$$

Then,

$$\text{Millimoles} = (0.0297\text{ mmol/mL})(250\text{ mL}) = 7.42\text{ mmol}$$

Lecture Two

Example : How many grams Na_2SO_4 should be weighed out to prepare 500mL of a 0.100 *M* solution?
Atomic mass for : Na = 23 , S= 32.1 , O = 16

Solve :

Molarity = mole / volume

M = n (mmol) / V (ml)

0.100 M = n (mol) / 500 ml

n(mmol) = 0.100 M * 500 ml = 50 mmol

n(mmol) = weight (g) / M.wt (g/mol)

M.wt (Na_2SO_4) = 23*2 + 32.1 *1+ 16*4 = 142 mg/mmol

weight (g) = 50 mmol * 142 $\frac{\mathbf{mg}}{\mathbf{mmol}} * \frac{1\ g}{1000\ mg}$

= 7.1 g

Example : Calculate the concentration of potassium ion (K^+) in grams per liter after mixing 100mL of 0.250 M KCl and 200mL of 0.100 M K_2SO_4. , M.wt $_K$ = 39.098m g/mmol

Solve: $KCl \rightarrow K^+ + Cl^-$, $K_2SO_4 \rightarrow 2\ K^+ + SO_4^-$

$n_{K+} = n_{KCl} + 2 * n_{K2SO4}$

$n_{KCl} = M_{KCl} * V_{KCl}$

n_{KCl} = 0.250 M * 100 ml = 25 mmol

$n_{K2SO4} = M_{K2SO4} * V_{K2SO4}$

= 0.100 M * 200 ml = 20 mmol

n_{K+} = 25 mmol + 2*20 mmol = 65 mmol in total volume (300 ml) .

V_{k+} = 200 ml + 100 ml = 300 ml

weight of K^+ = n_{K+} / V_{K+}

= (65 mmol * 39.1 mg/mmol * 0.001g/mg) / 300ml * 0.001 ml/ l = 8.47 g /L

Lecture Two

⇒ Normality

Although molarity is widely used in chemistry, some chemists use a unit of concentration in quantitative analysis called normality (N). A one-normal solution contains one equivalent per liter.

$$N = \frac{wt.(g) \times 1000}{eq.wt. \times V\ ml}$$

Eq. wt. is explained in lecture one (General principles to calculate the equivalent weight) page : 7

Or

$$N = \frac{\rho \times \% \times 10}{eq.wt.}$$

ρ : density of solution

% : concentration of subtenant