

Lecture Two

⇒ Examples

Example 1 : Calculate the weight of one mole of $\text{CaSO}_4 \cdot 7\text{H}_2\text{O}$. ?

Solve:

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

$$\text{Ca} = 40.08, \quad \text{S} = 32.06, \quad \text{O} = 16, \quad \text{H} = 1$$

$$\begin{aligned} \text{M.wt.} &= 40.08 \cdot 1 + 32.06 \cdot 1 + 16 \cdot 4 + 7 \cdot 2 \cdot 1 + 7 \cdot 16 \\ &= 262.25 \text{ g/mol} \end{aligned}$$

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

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

$$\begin{aligned} \text{Weight (g)} &= 1 \text{ mol} \cdot 262.25 \text{ g/mol} \\ &= 262.25 \text{ g} . \end{aligned}$$

Example 2: Calculate the number of moles in 500 mg Na_2WO_4 (sodium tungstate). M.wt. of Na_2WO_4 $293.8 \frac{\text{mg}}{\text{mmol}}$.

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

$$= 0.00170 \text{ mol}$$

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⇒ 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 .

$$\begin{aligned}\text{Moles} &= (\text{moles/liter}) \times \text{liters} \\ &= \text{molarity} \times \text{liters}\end{aligned}$$

Example

A solution is prepared by dissolving 1.26 g AgNO_3 in a 250-mL volumetric flask and diluting to volume. Calculate the molarity of the silver nitrate solution. How many mill moles AgNO_3 were dissolved? **M.wt. of AgNO_3 169.9 g/mol .**

Solution

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

Then,

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



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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

$$M = n \text{ (mmol)} / V \text{ (ml)}$$

$$0.100 \text{ M} = n \text{ (mol)} / 500 \text{ ml}$$

$$n(\text{mmol}) = 0.100 \text{ M} * 500 \text{ ml} = 50 \text{ mmol}$$

$$n(\text{mmol}) = \text{weight (g)} / \text{M.wt (g/mol)}$$

$$\text{M.wt (Na}_2\text{SO}_4) = 23*2 + 32.1 *1+ 16*4 = 142 \text{ mg/mmol}$$

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

$$= 7.1 \text{ g}$$

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 $\text{K} = 39.098 \text{ g/mmol}$

Solve: $\text{KCl} \rightarrow \text{K}^+ + \text{Cl}^-$, $\text{K}_2\text{SO}_4 \rightarrow 2 \text{K}^+ + \text{SO}_4^-$

$$n_{\text{K}^+} = n_{\text{KCl}} + 2* n_{\text{K}_2\text{SO}_4}$$

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

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

$$n_{\text{K}_2\text{SO}_4} = M_{\text{K}_2\text{SO}_4} * V_{\text{K}_2\text{SO}_4}$$

$$= 0.100 \text{ M} * 200 \text{ ml} = 20 \text{ mmol}$$

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

$$V_{\text{K}^+} = 200 \text{ ml} + 100 \text{ ml} = 300 \text{ ml}$$

$$\text{M of K}^+ = n_{\text{K}^+} / V_{\text{K}^+}$$

$$= (65 \text{ mmol} * 39.1 \text{ mg/mmol} * 0.001\text{g/mg}) / 300\text{ml} * 0.001 \text{ ml/ l} = 8.47 \text{ g /L}$$

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⇒ 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.

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

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

Or

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

ρ : density of solution

% : concentration of subtenant

⇒ Dilutions—Preparing The Right Concentration

We often must prepare dilute solutions from more concentrated stock solutions. For example, we may prepare a dilute HCl solution from concentrated HCl to be used for titrations . Or, we may have a stock standard solution from which we wish to prepare a series of more dilute standards. The millimoles of stock solution taken for dilution will be identical to the millimoles in the final diluted solution, remember, $C_1 V_1 = C_2 V_2$.

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Example : You wish to prepare a calibration curve for the spectrophotometric determination of permanganate. You have a stock 0.100 M solution of KMnO_4 and a series of 100-mL volumetric flasks. What volumes of the stock solution will you have to pipet into the flasks to prepare standards of a) 1×10^{-3} M ,b) 2×10^{-3} M, c) 5×10^{-3} M, d) 10.0×10^{-3} M KMnO_4 solutions?

Solve :

$$\text{a) } C_1 = 0.1 \text{ M} , V_1 = \text{ml?} , C_2 = 1 \times 10^{-3} \text{ M} , V_2 = 100 \text{ ml}$$

$$C_1 V_1 = C_2 V_2$$

$$0.1 \text{ M} * V_1 = 1 \times 10^{-3} \text{ M} * 100 \text{ ml}$$

$$V_1 = 1 \times 10^{-3} \text{ M} * 100 \text{ ml} / 0.1 \text{ M}$$

$$V_1 = 1 \text{ ml}$$

$$\text{b) } C_1 = 0.1 \text{ M} , V_1 = \text{ml?} , C_2 = 2 \times 10^{-3} \text{ M} , V_2 = 100 \text{ ml}$$

$$C_1 V_1 = C_2 V_2$$

$$0.1 \text{ M} * V_1 = 2 \times 10^{-3} \text{ M} * 100 \text{ ml}$$

$$V_1 = 2 \times 10^{-3} \text{ M} * 100 \text{ ml} / 0.1 \text{ M}$$

$$V_1 = 2 \text{ ml}$$

c) Home work

d) Home work

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Example : You are analyzing for the manganese content in an ore sample by dissolving it and oxidizing the manganese to permanganate for spectrophotometric measurement. The ore contains about 5% Mn. A 5-g sample is dissolved and diluted to 100 mL, following the oxidation step. By how much must the solution be diluted to be in the range of the calibration curve is prepared about 3×10^{-3} M permanganate? M.wt_{Mn} = 55 g/gmol

Solve:

The solution contains manganese (Mn) = $0.05 * 5 \text{ g} = 0.25 \text{ g}$

$$C_1 = n(\text{mol}) / V(\text{ml})$$

$$n(\text{mol}) = \text{weight}(\text{g}) / \text{M.wt.}(\text{g/gmol}) = (0.25 \text{ g} / 55 \text{ g/gmol}) = 4.5 * 10^{-3} \text{ mol}$$

$$C_1 = 4.5 * 10^{-3} \text{ mol} / 100 \text{ ml} = 4.5 * 10^{-5} \text{ M}$$

$$(C_1 V_1)_{\text{before dilute}} = (C_2 V_2)_{\text{after dilute}}$$

$$4.5 * 10^{-5} \text{ M} * V_1 = 3 * 10^{-3} \text{ M} * 100 \text{ ml}$$

$$V_1 = 66.7 * 10^2 \text{ ml}$$

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Example : What volume of 0.40 M Ba(OH)₂ must be added to 50mL of 0.30 M NaOH to give a solution 0.50 M in OH⁻ ?

Solve :

Volumes of dilute aqueous solutions can be assumed to be additive, i.e., if x mL of Ba(OH)₂ is added to 50mL NaOH, the total volume is going to be 50 + x mL.

We can use a modified form of $C_1V_1 = C_2V_2$ where all the initial solution components are added in this manner and these sum up to the final solution components: $\sum C_{in} V_{in} = \sum C_{fin} V_{fin}$

In the present case,

$$M_{NaOH} * V_{NaOH} + 2 * M_{Ba(OH)_2} * V_{Ba(OH)_2} = M_{OH^-} * V_{fin}$$

that 1 M Ba(OH)₂ is 2 M in OH⁻. (Ba(OH)₂ → Ba⁺ + 2OH⁻)

Thus 0.30 M × 50 mL + 2 × 0.40 M × x mL = 0.50 M × (50 + x) mL.

$$15 + 0.8 x = 25 + 0.50 x$$

$$0.8 x - 0.50 x = 25 - 15$$

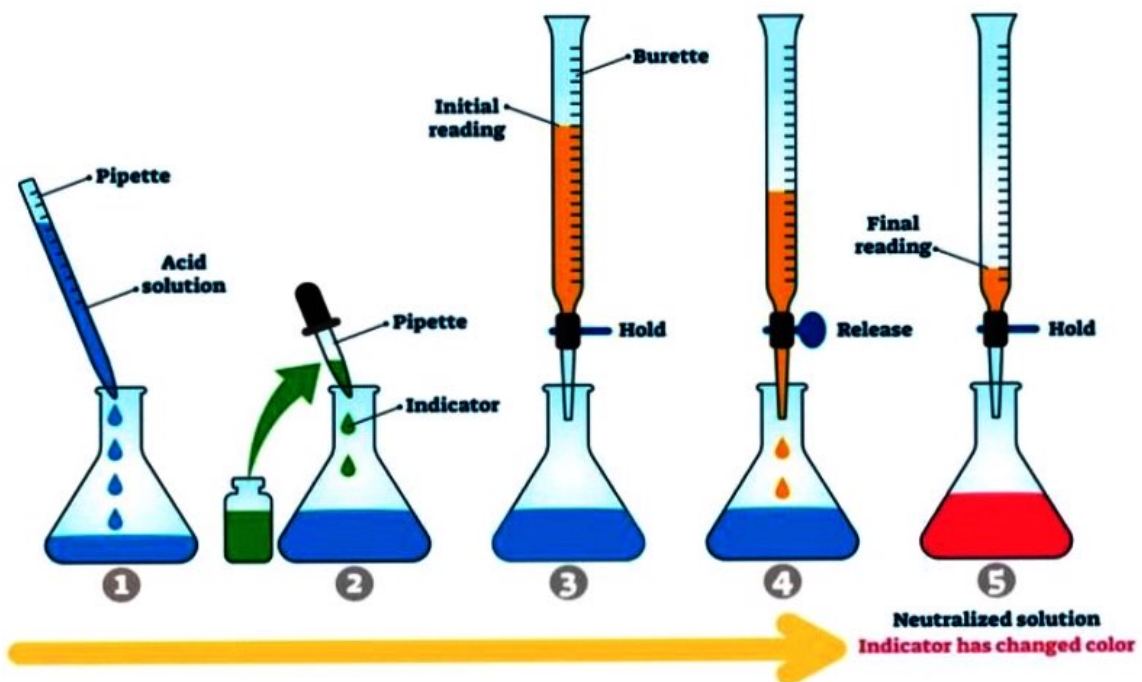
$$0.3 x = 10$$

X = 10 / 0.3 = 33.3 ml volume of Ba(OH)₂ must be added NaOH

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⇒ Titration

In a titration, the test substance (analyte) reacts with an added reagent of known concentration, generally instantaneously. The reagent of known concentration is referred to as a standard solution. It is typically delivered from a buret; the solution delivered by the buret is called the titrant. (In some instances, the reverse may also be carried out where a known volume of the standard solution is taken and it is titrated with the analyte of unknown concentration as the titrant.) The volume of titrant required to just completely react with the analyte is measured. Since we know the reagent concentration as well as the reaction stoichiometry between the analyte and the reagent, we can calculate the amount of analyte,



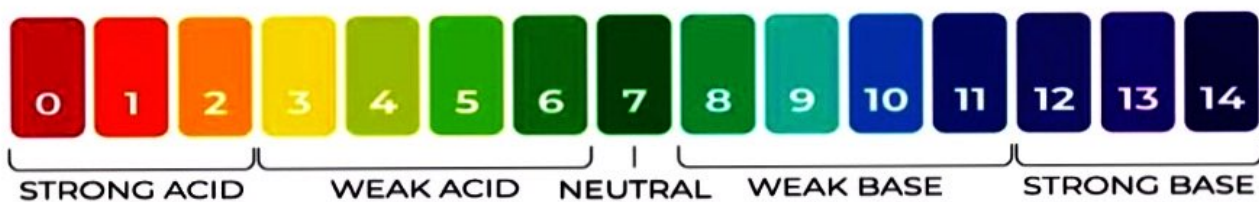
The equivalence point : is the theoretical end of the titration where the number equivalents of the analyte exactly equals the number of equivalents of the titrant added

The end point : is the observed end of the titration.

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Name of indicator	Colour in acid solution	Colour in basic solution	pH range
Methyl Orange	Red	Orange – yellow	3.1 – 4.6
Bromophenol Blue	Yellow	Blue – violet	3.0 – 4.6
Methyl Red	Red	Yellow	4.2 – 6.3
Bromothymol Blue	Yellow	Blue	6.0 – 7.6
Thymol Blue	Yellow	Blue	8.0 – 9.6

Phenolphthalein	Colourless	Red	8.0 – 9.8
Thymolphthalein	Colourless	Blue	9.4 – 10.6



The requirements of a titration are as follows:

1. The reaction must be stoichiometric. That is, there must be a well-defined and known reaction between the analyte and the titrant. In the titration of acetic acid in vinegar with sodium hydroxide, for example, a well-defined reaction takes place:
$$\text{CH}_3\text{COOH} + \text{NaOH} \rightarrow \text{CH}_3\text{COONa} + \text{H}_2\text{O}$$
2. The reaction should be rapid. Most ionic reactions, as above, are very rapid
3. There should be no side reactions; the reaction should be specific.
4. There should be a marked change in some property of the solution when the reaction is complete. This may be a change in color of the solution or in some electrical or other physical property of the solution.
5. The point at which an equivalent or stoichiometric amount of titrant is added is called the equivalence point. The point at which the reaction is observed to be complete is called the end point, that is, when a change in some property of the solution is detected. The end point should coincide with the equivalence point or be at a reproducible interval from it.

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⇒ Analytical Solutions – standard solution

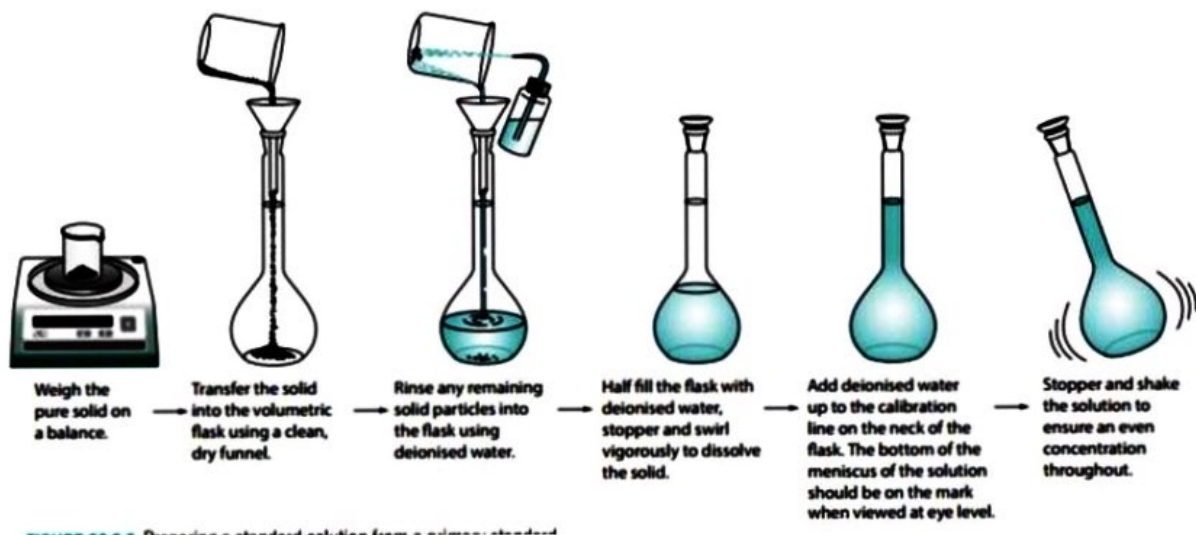
A solution of accurately known concentration, prepared by dissolving a known amount of solute in a known volume of solvent.

1- Primary standard solution

A solution of accurately known concentration prepared from primary standard material which has the following properties:

- 1- High purity (99.99%)
- 2- Stable toward air (not absorb H_2O or CO_2 from atmosphere).
- 3- Available in reasonable price (inexpensive)
- 4- Soluble in titration medium.
- 5- Has a large molecular weight (M.wt.) so that the relative error associated with weighing would be low.

Steps for preparing a standard solution



2- Secondary standard solution

This is usually prepared from non-primary standard substance and prepared in approximate concentration $\approx 0.1M$, and its concentration is then determined accurately by titration with a solution of primary standard material.