Chapter 4
Solution Stoichiometry

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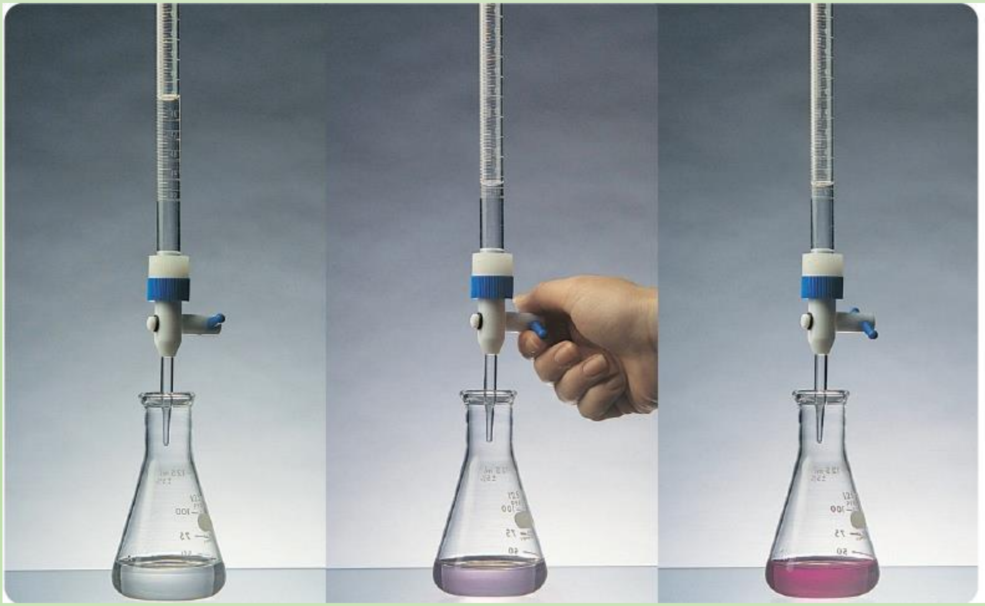
Solution Stoichiometry

- In solution stoichiometry you have to presume that soluble ionic compounds dissociate completely in solution.
- Then using mole ratios we can calculate the concentration of all species in solution.
- There are three common types of stoichiometric calculations
  - **Quantitative Analysis:** The determination of the amount of a substance or species present in a material.
  - **Volumetric Analysis:** A type of quantitative analysis based on titration.
  - **Gravimetric Analysis:** A type of quantitative analysis in which the amount of a species in a material is determined by converting the species to a product that can be isolated completely and weighed.

Volumetric Analysis - Titrations

A procedure for determining the amount of substance A by adding a carefully measured volume with a known concentration of B until the reaction of A and B is just complete. This can be for precipitation, neutralization or redox reactions.

- ***Standardization*** is the determination of the exact concentration of a solution.
- ***Equivalence point*** represents completion of the reaction.
- ***Endpoint*** is where the titration is stopped.
- An ***indicator*** is used to signal the endpoint.



Gravimetric Analysis

- In gravimetric analysis precipitation reactions are carried out.
- After the reaction the product is precipitated and collected in a crucible or filter paper.
- The precipitate is weighed and then using mole ratios we can calculate the concentration of all species in original solution.

The figure below shows the reaction of $Ba(NO_3)_2$ with K_2CrO_4 forming the yellow $BaCrO_4$ precipitate.

The $BaCrO_4$ precipitate is being filtered. It can then be dried and weighed. Then concentration of Ba^{2+} ions can be calculated

Example: Find the concentration of all species in a 0.25 *M* solution of $MgCl_2$

$$MgCl_2 \rightarrow Mg^{2+} + 2Cl^-$$

Given: $MgCl_2$ = 0.25 *M*

$[Mg^{2+}]$ = 0.25 *M* (1:1 ratio)

$[Cl^-]$ = 0.50 *M* (1:2 ratio)

Example: A soluble silver compound was analyzed for the percentage of silver by adding sodium chloride solution to precipitate the silver ion as silver chloride. If 1.583 g of silver compound gave 1.788 g of silver chloride, what is the mass percent of silver in the compound?

Strategy: g AgCl → mol AgCl → mol Ag → g Ag → % Ag

Molar mass of silver chloride (AgCl) = 143.32 g

$$1.788\text{ g AgCl} \times \frac{1\text{mol AgCl}}{143.32\text{ g AgCl}} \times \frac{1\text{mol Ag}}{1\text{mol AgCl}} \times \frac{107.9\text{ g Ag}}{1\text{mol Ag}} = 1.346\text{ g Ag in the compound}$$

$$\frac{1.346\text{ g Ag}}{1.583\text{ g silver compound}} 100\% = 85.03\%\text{ Ag}$$

Example: Zinc sulfide reacts with hydrochloric acid to produce hydrogen sulfide gas:

$$ZnS(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2S(g)$$

How many milliliters of 0.0512 M HCl are required to react with 0.392 g ZnS?

Strategy: g ZnS → mol ZnS → mol HCl (mol ratio from eq) → vol HCl (using Molarity)

Molar mass of ZnS = 97.45 g

$$0.392\ \text{g ZnS} \times \frac{1\,\text{mol ZnS}}{97.45\ \text{g ZnS}} \times \frac{2\,\text{mol HCl}}{1\,\text{mol ZnS}} \times \frac{1\,\text{L solution}}{0.0512\,\text{mol HCl}}$$

= 0.157 L = 157 mL HCl solution

Example: A dilute solution of hydrogen peroxide is sold in drugstores as a mild antiseptic. A typical solution was analyzed for the percentage of hydrogen peroxide by titrating it with potassium permanganate:

$$5H_2O_2(aq) + 2KMnO_4(aq) + 6H^+(aq) \rightarrow 8H_2O(l) + 5O_2(g) + 2K^+(aq) + 2Mn^{2+}(aq)$$

What is the mass percent of H_2O_2 in a solution if 57.5 g of solution required 38.9 mL of 0.534 M $KMnO_4$ for its titration?

Strategy: mols $KMnO_4$ → mols H_2O_2 → mass H_2O_2 → % H_2O_2

Molar mass of H_2O_2 = 34.01 g

$$38.9 \times 10^{-3} \text{L} \times \frac{0.534 \text{ mol KMnO}_4}{1 \text{L}} \times \frac{5 \text{ mol H}_2\text{O}_2}{2 \text{ mol KMnO}_4} \times \frac{34.01 \text{ g H}_2\text{O}_2}{1 \text{ mol H}_2\text{O}_2} = \mathbf{1.77\ g\ H_2O_2}$$

$$\frac{1.77 \text{ g H}_2\text{O}_2}{57.5 \text{ g solution}} 100\% = \mathbf{3.07\%\ H_2O_2}$$

Example: A student measured exactly 15.0 mL of an unknown monoprotic acidic solution and placed in an Erlenmeyer flask. An indicator was added to the flask. At the end of the titration the student had used 35.0 mL of 0.12 *M* NaOH to neutralize the acid. Calculate the molarity of the acid.

Strategy: mols NaOH → mols acid (from eq) → Molarity of acid

$$0.035 \text{ L NaOH x } \frac{0.12 \text{ mol NaOH}}{1 \text{ L}} \text{ x } \frac{1 \text{ mol acid}}{1 \text{ mol base}} = 0.0042 \text{ mol acid}$$

$$\text{Molarity} = \frac{0.0042 \text{ mol}}{0.015 \text{ L}} = \mathbf{0.28\ M\ acid}$$

Example: Calculate the molarity of 25.0 mL of a monoprotic acid if it took 45.50 mL of 0.25 *M* KOH to neutralize the acid.

$$\frac{0.25 \text{ mol KOH}}{\text{L}} \times 0.04550 \text{ L} \times \frac{1 \text{ mol acid}}{1 \text{ mol KOH}} = 0.01338 \text{ mol acid}$$

$$\frac{0.01338 \text{ mol acid}}{0.0250 \text{ L}} = 0.455\, M$$

Key Words/Concepts

- Molarity (mol/L)
- Solution stoichiometry
    - Volumetric analysis (titration)
        - End point
        - Equivalence point
    - Gravimetric analysis