Solutions

– homogeneous mixtures
– solvent, solute(s)

4.6 Molarity (*M*)

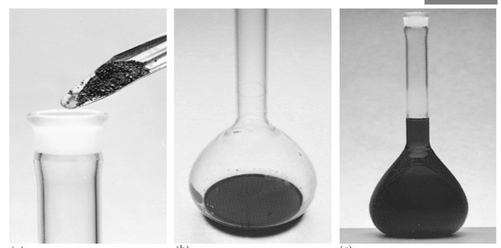
- Measure of the solute concentration

$$M = \left(\frac{\text{amount of solute (mol)}}{\text{volume of solution (L)}}\right) \quad \text{or} \quad M = \frac{n}{V}$$

- Units - molar (M) **1 M = 1 mol/L**

- Preparation of solutions with known molarity
  – transfer a known mass of solute in a volumetric flask
  – dissolve in small amount of water
  – add water to the mark

Fig. 4.10



(a) (b) (c)

- **Example:** Calculate the molarity of a solution prepared by dissolving **5.33 g NaOH** in water using a **100.0 mL** volumetric flask.

⇒**convert the mass to moles:**

$$5.33 \text{ g NaOH} \times \left(\frac{1 \text{ mol NaOH}}{40.00 \text{ g NaOH}}\right) = 0.133 \text{ mol NaOH}$$

⇒**convert volume to liters: 100.0 mL = 0.1000 L**

⇒**divide moles by solution volume:**

$$\frac{0.133 \text{ mol NaOH}}{0.1000 \text{ L solution}} = 1.33 \text{ mol NaOH / L} \rightarrow 1.33 \text{ M NaOH}$$

- Calculations using molarity

$M = n/V$ $n = M \times V$ $V = n/M$

Example: Calculate the volume of **1.33 M NaOH** solution that contains **5.00 mol NaOH**.

$$V = \frac{n}{M} = \frac{5.00 \text{ mol NaOH}}{1.33 \text{ mol NaOH / L}} = 3.76 \text{ L}$$

or use molarity as a conversion factor:

$$5.00 \text{ mol NaOH} \times \left(\frac{1 \text{ L}}{1.33 \text{ mol NaOH}}\right) = 3.76 \text{ L}$$

- Preparation of solutions with given molarity

Example: Calculate the mass of **NaOH** needed to prepare **250. mL 1.33 M** solution.

⇒**calculate the # of moles of NaOH needed:**

use $n = M \times V$ or the conversion method

$$0.250 \text{ L} \times \left(\frac{1.33 \text{ mol NaOH}}{1 \text{ L}}\right) = 0.332 \text{ mol NaOH}$$

⇒**convert the moles of NaOH to grams:**

$$0.332 \text{ mol NaOH} \times \left(\frac{40.00 \text{ g NaOH}}{1 \text{ mol NaOH}}\right) = 13.3 \text{ g NaOH}$$

4.7 Dilution

- Reducing the concentration of the solute by adding more solvent
- Stock solutions - concentrated solutions used to store reagents
- Dilution Procedure
  – use a pipette to measure a small volume of the concentrated solution and transfer it to a volumetric flask
  – add solvent to fill the volumetric flask to the mark

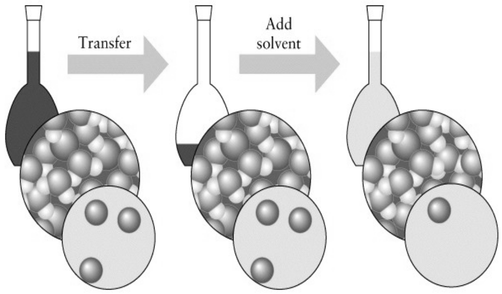


Fig. 4.12

- Dilution calculations

$$n = M \times V$$

  - dilution doesn't change the total # of moles of solute in the solution

$$\boldsymbol{n_i = n_f} \qquad \boldsymbol{M_i \times V_i = M_f \times V_f}$$

- **Example:** Calculate the molarity of a solution prepared by dilution of **5.00 mL 2.0 M HCl** stock solution to **100.0 mL**.

$$M_f = \frac{M_i \times V_i}{V_f} = \frac{2.0\ \text{M} \times 5.00\ \text{mL}}{100.0\ \text{mL}} = 0.10\ \text{M}$$

4.8 Titrations

- Use measurements of volumes - volumetric methods of analysis
- Based on stoichiometric reactions between the analyzed solution (analyte) and a solution with known concentration (titrant)
- Equivalence point - the amount of titrant added is stoichiometrically equivalent to the amount of analyte present in the sample
- Indicators - change color at the equivalence point (signal the end of the titration)

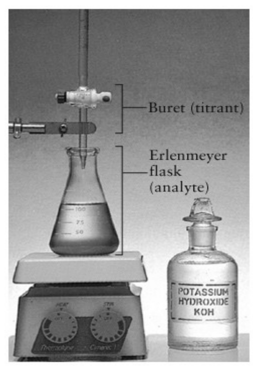


Fig. 4.14

- Types of titrations
  - neutralization, redox, precipitation, etc.
- The titrant is added slowly until the indicator changes color
- The concentration of the analyte is calculated from the measured volumes of the solutions and the titrant concentration

- Stoichiometric calculations involving solutions

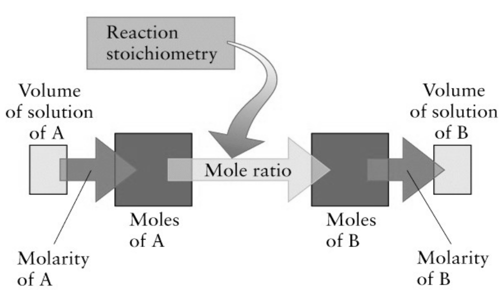


Fig. 4.16

- Titration calculations

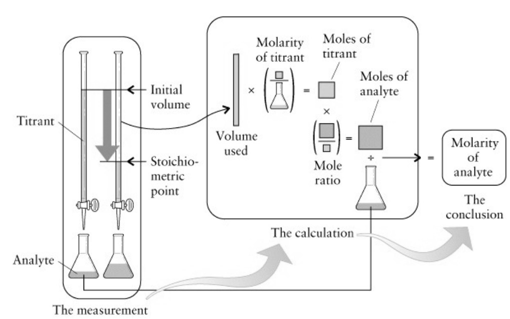


Fig. 4.17

- **Example:** A **25.0 mL H_2SO_4** solution is titrated with **16.4 mL 0.255 M KOH** solution. What is the molarity of the acid solution.

⇒**balanced equation:**

$2KOH_{(aq)} + H_2SO_{4(aq)} \rightarrow K_2SO_{4(aq)} + 2H_2O_{(l)}$

⇒**mole ratio:** **[1 mol H_2SO_4/2 mol KOH]**

$$16.4 \times 10^{-3}\ \text{L} \times \left(\frac{0.255\ \text{mol KOH}}{1\ \text{L}}\right) \times \left(\frac{1\ \text{mol } H_2SO_4}{2\ \text{mol KOH}}\right) =$$

$$= 2.09 \times 10^{-3}\ \text{mol } H_2SO_4$$

$$\frac{2.09 \times 10^{-3}\ \text{mol } H_2SO_4}{25.0 \times 10^{-3}\ \text{L}} = 8.36 \times 10^{-2}\ \text{M } H_2SO_4$$

- Calculations of the mass of the analyte

Example: A 0.202 g sample of iron ore is dissolved in HCl and all of its **Fe** content is converted to **Fe^{2+}**. The resulting solution is titrated with **16.7 mL 0.0108 M $KMnO_4$** solution. Determine the **mass%** of Fe in the sample, if the equation of the redox reaction is:

$5Fe^{2+} + MnO_4^- + 8H^+ \rightarrow$

$\rightarrow 5Fe^{3+} + Mn^{2+} + 4H_2O_{(l)}$

⇒**mole ratio:** **[5 mol Fe^{2+}/1 mol MnO_4^-]**

⇒**calculate the mass of Fe:**

$$16.7 \times 10^{-3}\ \text{L} \times \left(\frac{0.0108\ \text{mol } MnO_4^-}{1\ \text{L}}\right) \times \left(\frac{5\ \text{mol } Fe^{2+}}{1\ \text{mol } MnO_4^-}\right) \times$$

$$\times \left(\frac{55.85\ \text{g } Fe^{2+}}{1\ \text{mol } Fe^{2+}}\right) = 0.0504\ \text{g } Fe^{2+} \rightarrow 0.0504\ \text{g Fe}$$

⇒**calculate the mass%:**

$$\textit{Mass}\%\ \text{Fe} = \frac{0.0504\ \text{g Fe}}{0.202\ \text{g sample}} \times 100\% = 25.0\%$$

Assignments:

- Homework: Chpt. 4/1, 3, 7, 13, 15, 19, 23, 27, 31, 33, **37, 39, 43, 45, 49, 51, 73**
- Student Companion: 4.2, 4.4, **4.5**