13.4 Expressing Solute Concentration

- **Concentration** – the ratio of the quantity of solute to the quantity of solution (or solvent)

Table 13.5 Concentration Definitions

Concentration Term	Ratio
Molarity (M)	$\dfrac{\text{amount (mol) of solute}}{\text{volume (L) of solution}}$
Molality (m)	$\dfrac{\text{amount (mol) of solute}}{\text{mass (kg) of solvent}}$
Parts by mass	$\dfrac{\text{mass of solute}}{\text{mass of solution}}$
Parts by volume	$\dfrac{\text{volume of solute}}{\text{volume of solution}}$
Mole fraction (X)	$\dfrac{\text{amount (mol) of solute}}{\text{amount (mol) of solute + amount (mol) of solvent}}$

- **Molarity** (***M***) – the number of moles of solute per 1 liter of solution

$$M = \textbf{(mol of solute)/(liters of solution)}$$

  – ***M*** is affected by temperature (the solution volume changes with ***T***, so ***M*** changes too)

  – The solution volume is not a sum of the solvent and solute volumes (it must be measured after mixing)

- **Molality** (***m***) – the number of moles of solute per 1 kilogram of solvent

$$m = \textbf{(mol of solute)/(kilograms of solvent)}$$

  – ***m*** is not affected by temperature (the amounts of solute and solvent don't change with ***T***)

  – The solution volume is not needed; ***m*** can be calculated from the masses of solute and solvent

- ***M*** and ***m*** are nearly the same for dilute aqueous solutions since **1 L** of water is about **1 kg**, so (liters of solution) ≈ (kg of solvent)

Example: Calculate ***M*** and ***m*** for a solution prepared by dissolving **2.2 g** of NaOH in **55 g** of water if the density of the solution is **1.1 g/mL**.

$$\textit{mol solute} = \mathbf{2.2\ g\ NaOH \times \frac{1\ mol\ NaOH}{40\ g\ NaOH} = 0.055\ mol}$$

$$m = \mathbf{\frac{0.055\ mol\ NaOH}{0.055\ kg\ water} = 1.0\frac{mol}{kg}} \rightarrow \boxed{\mathbf{1.0\ m}}\ (\textit{molal})$$

$$\textit{Volume} = \frac{\textit{mass}}{\textit{density}} = \mathbf{\frac{2.2\ g + 55\ g}{1.1\ g/mL} = 52\ mL}$$

$$M = \mathbf{\frac{0.055\ mol\ NaOH}{0.052\ L\ solution} = 1.1\frac{mol}{L}} \rightarrow \boxed{\mathbf{1.1\ M}}\ (\textit{molar})$$

- Parts of solute by parts of solution

- **Parts by mass**

  – **Mass %** – grams of solute per 100 grams of solution → % (w/w)

$$\textit{Mass\%} = \mathbf{\frac{mass\ of\ solute}{mass\ of\ solution} \times 100\%}$$

  – **ppm** or **ppb** – grams of solute per **1 million** or **1 billion** grams of solution (for trace components)

- **Parts by volume**

  – **Volume %** – volume of solute per 100 volumes of solution → % (v/v)

$$\textit{Volume\%} = \mathbf{\frac{volume\ of\ solute}{volume\ of\ solution} \times 100\%}$$

– **ppmv** or **ppbv** – volume of solute per **1 million** or **1 billion** volumes of solution (used for trace gases in air)

- **Mole fraction** (***X***) – ratio of the # mol of solute to the total # mol (solute + solvent)

$$X = \frac{\textbf{mol of solute}}{\textbf{mol of solute + mol of solvent}}$$

Example: Calculate the *X* of NaOH in a solution containing **2.2 g** of NaOH in **55 g** of water.

$$\boxed{2.2\text{ g}\frac{1\text{ mol}}{40\text{ g}} = 0.055\text{ mol NaOH}} \quad \boxed{55\text{ g}\frac{1\text{ mol}}{18\text{ g}} = 3.1\text{ mol }H_2O}$$

$$X = \frac{0.055\text{ mol}}{0.055\text{ mol} + 3.1\text{ mol}} = 0.018$$

- Converting units of concentration

Example: A sample of water is $\mathbf{1.1\times10^{-6}}$ **M** in chloroform (CH_3Cl). Express the concentration of chloroform in **ppb**. (Assume density of 1.0 g/mL)

$$1.1\times10^{-6}\text{ M} \rightarrow 1.1\times10^{-6}\text{ mol }CH_3Cl\text{ per 1 L solution}$$

$$1.1\times10^{-6}\text{ mol }CH_3Cl\times\frac{50.5\text{ g }CH_3Cl}{1\text{ mol }CH_3Cl} = 5.6\times10^{-5}\text{ g }CH_3Cl$$

$$1\text{ L} \rightarrow 1000\text{ mL}\times1.0\frac{\text{g}}{\text{mL}} = 1.0\times10^{3}\text{ g } \textit{solution}$$

$$\frac{5.6\times10^{-5}\text{ g }CH_3Cl}{1.0\times10^{3}\text{ g } \textit{solution}}\times10^{9}\text{ ppb} = 56\text{ ppb}$$

Example: What is the molality of a solution of methanol in water, if the mole fraction of methanol in it is **0.250**?

Assume 1 mol of solution:

$$\rightarrow n_{\text{meth}} = 1\text{ mol}\times0.250 = 0.250\text{ mol}$$

$$\rightarrow n_{\text{water}} = 1 - 0.250 = 0.750\text{ mol}$$

$$0.750\text{ mol }H_2O\times\frac{18.0\text{ g }H_2O}{1\text{ mol }H_2O}\times\frac{1\text{ kg}}{10^{3}\text{ g}} = 0.0135\text{ kg }H_2O$$

$$m = \frac{0.250\text{ mol methanol}}{0.0135\text{ kg }H_2O} = 18.5\text{ m}$$

Example: What is the molality of a **1.83 M** NaCl solution with density of **1.070 g/mL**?

Assume 1 L (10^3 mL) of solution:

$$\rightarrow n_{\text{NaCl}} = 1.83\text{ mol}$$

$$\textit{mass of solution} = 10^{3}\text{ mL}\times\frac{1.070\text{ g}}{1\text{ mL}} = 1070\text{ g}$$

$$\textit{mass of NaCl} = 1.83\text{ mol}\times\frac{58.44\text{ g NaCl}}{1\text{ mol}} = 107\text{ g}$$

$$\textit{mass of water} = 1070\text{ g} - 107\text{ g} = 963\text{ g} = 0.963\text{ kg}$$

$$m = \frac{1.83\text{ mol NaCl}}{0.963\text{ kg }H_2O} = 1.90\text{ m}$$

Example: What is the molarity of a **1.20 m** KOH solution in water having density of **1.05 g/mL**?

Assume 1 kg (1000 g) of solvent (H_2O):

→ n_{KOH} = 1.20 mol

$$\textit{mass of KOH} = 1.20\ \text{mol} \times \frac{56.1\ \text{g KOH}}{1\ \text{mol}} = 67.3\ \text{g}$$

$$\textit{mass of solution} = 1000\ \text{g} + 67.3\ \text{g} = 1067\ \text{g}$$

$$\textit{volume of solution} = 1067\ \text{g} \times \frac{1\ \text{mL}}{1.05\ \text{g}} = 1016\ \text{mL} = 1.02\ \text{L}$$

$$M = \frac{1.20\ \text{mol KOH}}{1.02\ \text{L}\ \textit{solution}} = 1.18\ \text{M}$$