Solution Concentration

O Solution Concentration is a statement of the amount of solute present in a solution relative to either the amount of solvent or the amount of solution.

L **Weight Percentage**

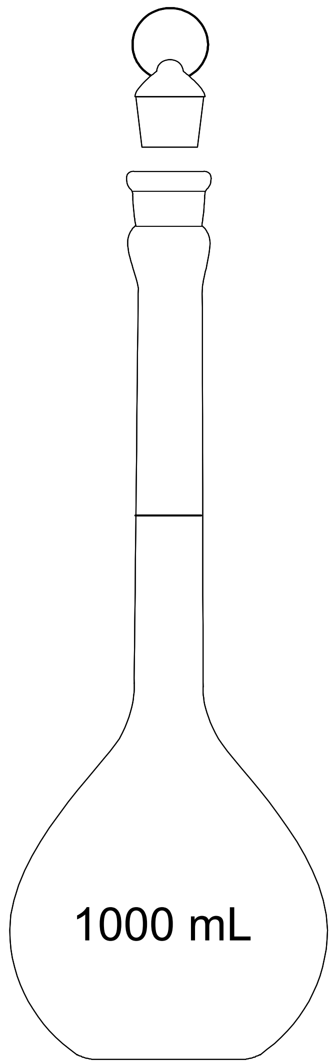
$$\text{weight \%} = \left(\frac{\text{solute weight}}{\text{solution weight}} \right) \times 100\%$$

L **Molarity (M)**

$$\text{Molarity} = \text{M} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

$$\text{Molarity} = \text{M} = \frac{\text{millimoles of solute}}{\text{milliliters of solution}}$$

Volumetric Flask



Molarity and Volume

L A given volume of a solution with a given molarity contains a specific amount of solute in moles or millimoles.

$$M \times V_{L} \;\text{'}\; \left(\frac{\text{mol solute}}{\text{L soln}} \right) (\text{L soln}) \;\text{'}\; \text{mol solute}$$

$$M \times V_{mL} \;\text{'}\; \left(\frac{\text{mmol solute}}{\text{mL soln}} \right) (\text{mL soln}) \;\text{'}\; \text{mmol solute}$$

Preparing a Solution by Diluting a More Concentrated Solution

L When a solution of a particular molarity is diluted, by adding more solvent, the number of moles of solute does not change.

$$M_{initial}V_{initial} = M_{final}V_{final}$$

or more simply

$$M_iV_i = M_fV_f$$

Analytical Concentration vs. Actual Concentration

- If a strong electrolyte breaks up completely on dissolving, then there is none of the undissociated solute in solution; i.e., its actual concentration is zero!

- Analytical concentration (*C*) is the number of moles per liter (or other concentration term) of solute put in solution without regard to what happens to it in solution.

- Actual concentrations of the species that exist in solution are indicated by writing their symbols in square brackets; e.g., $[Na^+] = 0.010$ M.

Analytical vs. Actual Concentrations
Examples

Strong electrolyte: NaCl, $C = 0.10$ M

	NaCl(*s*)	6	Na^+(*aq*)	+	Cl^-(*aq*)
Add	0.10 M		0 M		0M
Result	0 M		0.10 M		0.10 M

Strong electrolyte: Na_3PO_4, $C = 0.10$ M

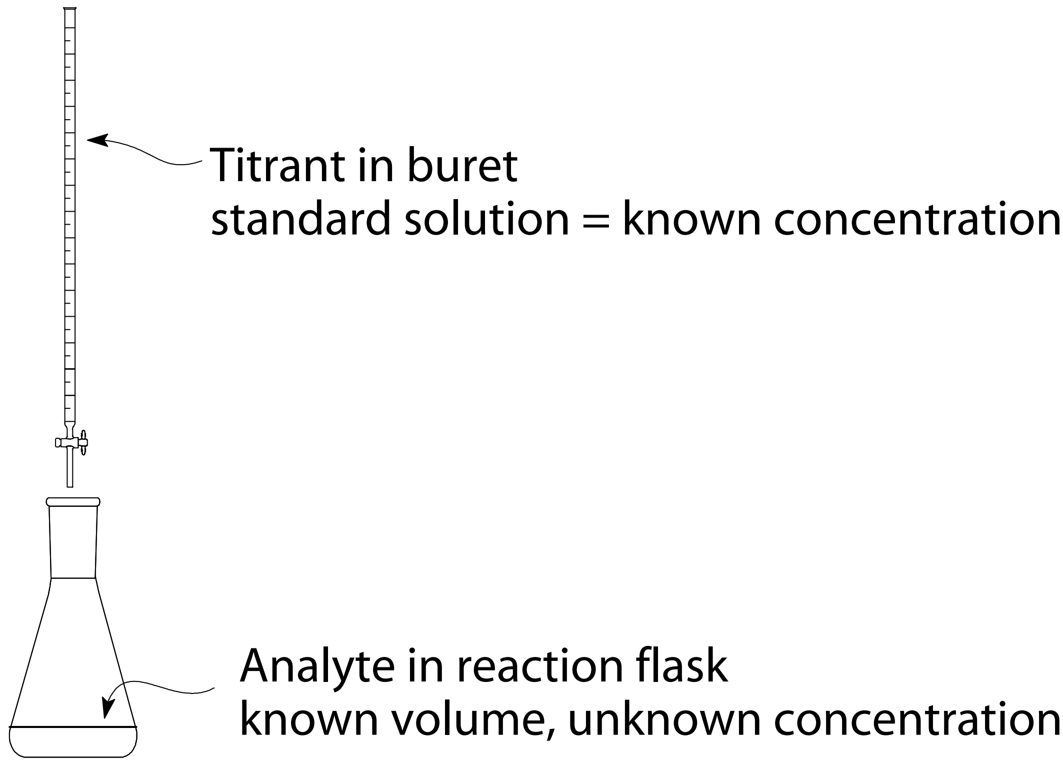
	Na_3PO_4(*s*)	6	$3Na^+$(*aq*)	+	PO_4^{3-}(*aq*)
Add	0.10 M		0 M		0 M
Result	0 M		0.30 M		0.10 M

Weak electrolyte: CH_3CO_2H, $C = 0.10$ M

	CH_3CO_2H(*l*)	º	H^+(*aq*)	+	$CH_3CO_2^-$(*aq*)
Add	0.10 M		0 M		0 M
Result	0.098 M		0.0013 M		0.0013 M

[At this point, we can't predict the actual concentrations for a weak electrolyte, but eventually we will be able to do so.]

Titration



n [Titrant] + m [Analyte] 6 products

Usually $m = 1$, and $n = 1$ or less often $n = 2, 3,$ or 4

Equivalence Point and End Point

- The **equivalence point** is the volume of added titrant necessary to completely neutralize the analyte.

- The **end point** is the volume of added titrant at which the color of an **indicator** changes.

- Ideally, the equivalence point and the end point are the same, but in reality there is usually a small difference between the two, which introduces some error.

Types of Acid-Base Titrations

Analyte	Titrant
Weak or strong base	Strong acid (e.g., HCl)
Weak or strong acid	Strong base (e.g., NaOH)

Mole Relationships at Equivalence

O If the analyte is monoprotic:

$$M_t V_t = M_a V_a$$

Example: NaOH = titrant, HCl = analyte

$$NaOH(aq) + HCl(aq)\ 6\ NaCl(aq) + H_2O(l)$$

$$M_{NaOH} V_{NaOH} \ \ '\ \ M_{HCl} V_{HCl}$$

O If the analyte is polyprotic (n = 2, 3, ...):

$$M_t V_t = n M_a V_a$$

Example: NaOH = titrant, H_2SO_4 = analyte

$$2NaOH(aq) + H_2SO_4(aq)\ 6\ Na_2SO_4(aq) + 2H_2O(l)$$

$$M_{NaOH} V_{NaOH} \ \ '\ \ 2 \times M_{H_2SO_4} V_{H_2SO_4}$$