Solution Stoichiometry

SOLUTION STOICHIOMETRY

- the methods used to calculate the quantities of substances in solution.
- involves molar concentrations and the volumes of solutions.

Sample Problem

Solutions of ammonia and phosphoric acid are used to produce ammonium hydrogen phosphate fertilizer. What volume of 14.8 mol/L $NH_{3(aq)}$ is needed for the ammonia to react completely with 10.0 L of 12.9 mol/L $H_3PO_{4(aq)}$ to produce fertilizer?

$$2NH_{3(aq)} + H_3PO_{4(aq)} \longrightarrow (NH_4)_2HPO_{4(aq)}$$

Titrations

In solution stoichiometry, sometimes you don't have enough information t
solve the problem on paper

Ex. 10 mL of acetic acid reacts with a 0.202 mol/L NaOH solution. What is the concentration of the acetic acid?

$C = \frac{n}{V}$

$CH_3COOH_{(aq)} + NaOH_{(aq)} \rightarrow H_2O_{(l)} + NaCH_3COO_{(aq)}$

x mol/L c = 0.202 mol/L

v = 0.0100 L **v = ???** *****

**** you need this volume in order to solve the problem

titration= procedure used to find the volume of substances to help calculate concentration

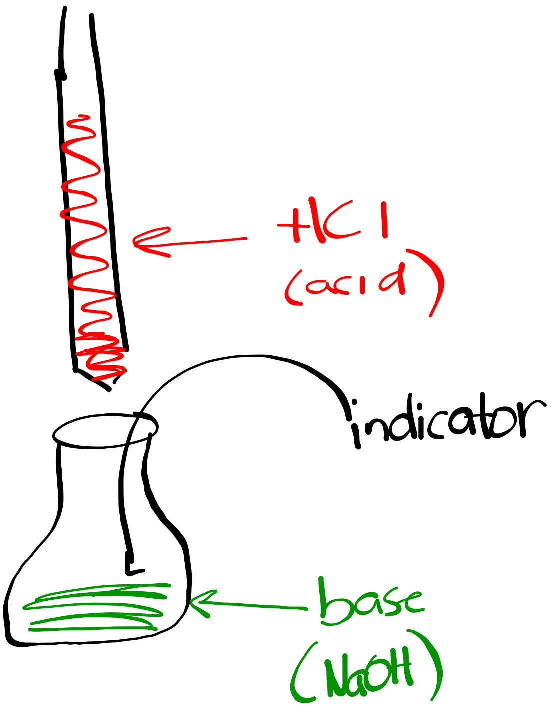
a solution (titrant) is transferred from a precisely marked tube called a buret to a flask containing another sample and an indicator

an indicator (eg. methyl orange, bromothymol blue) is used because a sudden change in colour indicates the completion of the reaction

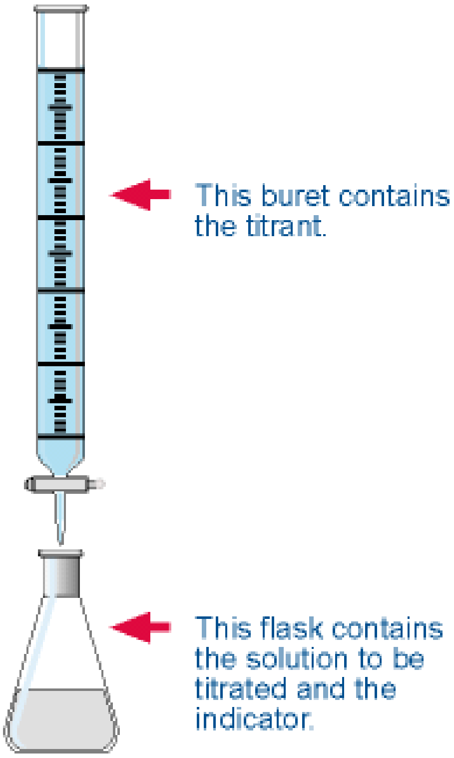
endpoint- the point where the titrant reacts completely with the sample

equivalence point- volume needed to reach the endpoint

A minimum of 3 trials is needed to ensure results are accurate


HCl
(acid)
indicator
base
(NaOH)

http://www.youtube.com/watch?v=YDzzMcrdyB4


This buret contains the titrant.
This flask contains the solution to be titrated and the indicator.

Example

A 10.00 ml sample of hydrochloric acid was titrated with a standardized solution of 0.685 mol/L NaOH. Bromothymol blue indicator was used and it changes from yellow to green at the endpoint. What is the concentration of hydrochloric acid?

Note: hydrochloric acid "is titrated with" NaOH
flask buret

Titration Results:

Trial	1 overshoot	2	3	4
Final Volume (mL)	11.30	22.25	33.05	44.05
Initial Volume (mL)	0.20 1	11.30	22.25	33.05
Volume ~~NaOH~~ HCl (mL)	11.10	10.95	10.80	11.00
Endpoint Colour	blue	blue	blue	blue

Avg. volume: $\frac{10.95\,mL + 10.80\,mL + 11.00\,mL}{3}$

$= \underline{\underline{10.92\,mL}}$

$\frac{0.100\,mol\,\cancel{HCl}}{1\cancel{L\,HCl}} \times 0.01092\,\cancel{L\,HCl} \times \frac{1\,mol\,NaOH}{1\,\cancel{mol\,HCl}} \times \frac{1}{0.0100L\,NaOH}$

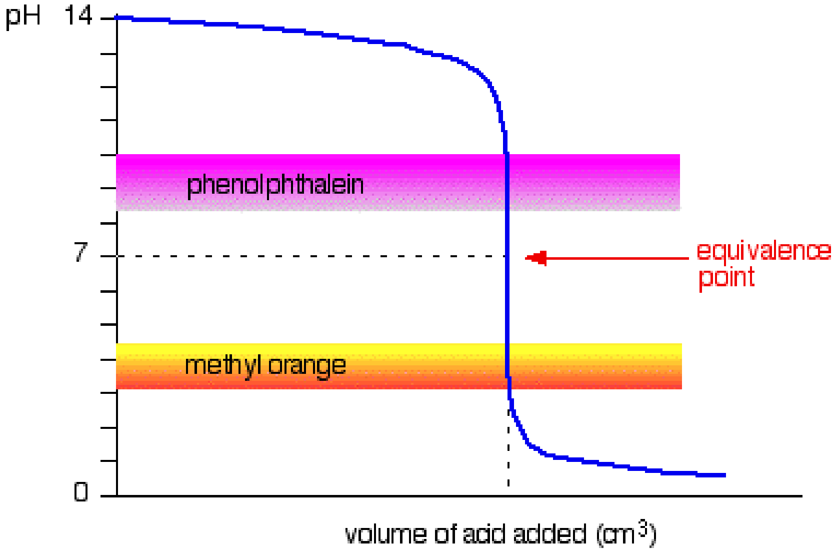
$= \boxed{0.192\,mol/L\,NaOH}$

EXAMPLE: Calculating Molarity from Titration Data

Titration reveals that 11.6 mL of 3.0 M sulfuric acid are required to neutralize the sodium hydroxide in 25.00 mL of NaOH solution. What is the molarity of the NaOH solution?

Solution:

$$H_2SO_{4(aq)} + 2NaOH_{(aq)} \longrightarrow 2H_2O_{(l)} + Na_2SO_{4(aq)}$$


pH
14
7
0
phenolphthalein
methyl orange
equivalence point
volume of acid added (cm^3)

Strong Acids (100% rxn)

pH = ?

$HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$

0.300M 0.300M

$[H^+] = [H_3O^+]$

$pH = -\log[H^+]$

Weak Acids

pH = ?

$HF_{(aq)} + H_2O_{(l)} \rightleftharpoons F^-_{(aq)} + H_3O^+_{(aq)}$

0.300M

$K_a = \frac{[F^-][H_3O^+]}{[HF]}$, $[F^-] = [H_3O^+]$

$K_a = \frac{[H_3O^+]^2}{[HF]}$

$pH = -\log[H_3O^+]$

Strong Bases (OH^-)

pH = ?

$$NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

0.300 M 0.300 M 0.300 M

$pOH = -\log[OH^-]$ $pH + pOH = 14.00$

Weak Bases

$$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH^+_{4(aq)} + OH^-_{(aq)}$$

0.300 M

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}, \quad [NH_4^+] = [OH^-]$$

$$K_b = \frac{[OH^-]^2}{[NH_3]}$$

$K_a K_b = K_w$

$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{}$

$pOH = -\log[OH^-]$

$pH + pOH = 14.00$

$$HSO_3^- + H_2O \rightleftharpoons SO_3^{2-} + H_3O^+$$

$$HSO_3^- + H_2O \rightleftharpoons H_2SO_3 + OH^-$$