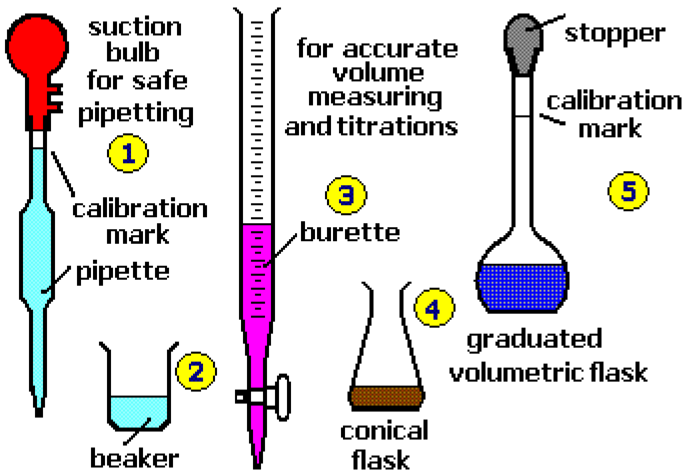
Dilution of solutions



A solution of known normality is frequently prepared from a more concentrated one of known strength by quantitative dilution.

$$N_{conced}\ x\ V_{conced} = N_{dil.}\ X\ V_{dil.}$$

Ex/What volume of 0.1500 N reagent is needed for the preparation of 500.0 ml ,0.100 N solution?

Solution/

$$N_{conced}\ x\ V_{conced} = N_{dil}.\ X\ V_{dil}$$

$$V_{conced} = \frac{Ndil\ x\ Vdil}{Nconced} = \frac{0.100\cancel{N}\ X\ 500.0\ ml}{0.1500\cancel{N}} = 333.3\text{ ml}$$

Ex/ Describe the preparation of 100 ml 6.0 M HCl from a concentrated solution that has a specific gravity of 1.18 and is 37%(w/w)HCl?

Solution/

$$[HCl] = \frac{sp.gr\ x\ \%\ 1000}{M.wt}$$

$$= \frac{1.18\ X\ 1000 gm\ reagent\ x\ 37 gm\ HCl\ X\ 1\ mol\ HCl}{L\ reagent\ x\ 100\ gm\ reagent\ x\ 36.5\ gm\ HCl}$$

M = 12.0 mol/L = 12.0 M

No.mol HCl = 100 ~~ml~~ x 6.0 mmol/ ~~ml~~

$$= 600.0\ \cancel{mmol} \times \frac{1\ mol}{1000\cancel{mmol}} = 0.6\ mol$$

$$V_{conced}\ reagent = 0.6\ \cancel{mol} \times \frac{1\ l\ reagent}{12.0\ \cancel{mol}} = 0.05\ \cancel{L} \times 1000ml/\cancel{L}$$

$$= 50\ ml$$

Thus dilute 50 ml of the concentrated reagent to 100 ml.

طريقة اخرى للحل

$$M_{conced} \times V_{conced} = M_{dil.} \times V_{dil.}$$

$$12.0\ M \times V_{conced} = 6.0\ M \times 100\ ml$$

$$V_{conced} = \frac{6.0\ \cancel{M} \times 100ml}{12.0\ \cancel{M}} = 50\ ml$$

Ex/ What volume of H_2SO_4 reagent is needed for the preparation of 200.0 ml,0.3 N solution that has a specific gravity of 1.84 and is 98%(w/w)H_2SO_4?

Solution/

$$[H_2SO_4] = \frac{1.84 \times 1000\cancel{gm\ reagent} \times 98\cancel{gm\ H2SO4} \times 1eq\ H2SO4}{L\ reagent \times 100\ \cancel{gm\ reagent} \times 49.0\ \cancel{gm\ H2SO4}}$$

$$= 36.8\ eq / L = 36.8\ N$$

$$No.eq\ H_2SO_4 = 200.0\ \cancel{ml} \times 0.3\ meq/\cancel{ml}$$

$$= 60.0\ meq$$

$$V_{conced} reagent = 60.0\ \cancel{meq} \times \frac{1\ ml}{36.8\cancel{meq}\ H2SO4} = 1.63\ ml$$

Dilute 1.63 ml of the concentrated reagent to 200.0ml

Analysis of samples by titration with standard solution

Titrimetric methods include a large and powerful group of quantitative procedures that are based upon measuring the amount of a reagent of known concentration that is consumed by the analyte.

Volumetric titrimetry involves measuring the volume of a solution of known concentration that is needed to react essentially completely with the analyte.

Standard solution (standard titrant) : is a reagent of known concentration that is used to carry out a titrimetric analysis.

The equivalence point in titration is reached the amount of added titrant is chemically equivalent to the amount of analyte in the sample.

For example: the equivalence point in the titration of sodium chloride with silver nitrate occurs after exactly 1 mol of silver ion has been added for each mol of chloride ion in the sample.

The equivalence point in the titration of sulfuric acid with sodium hydroxide is reached after introduction of 2 mol of base for each mol of acid.

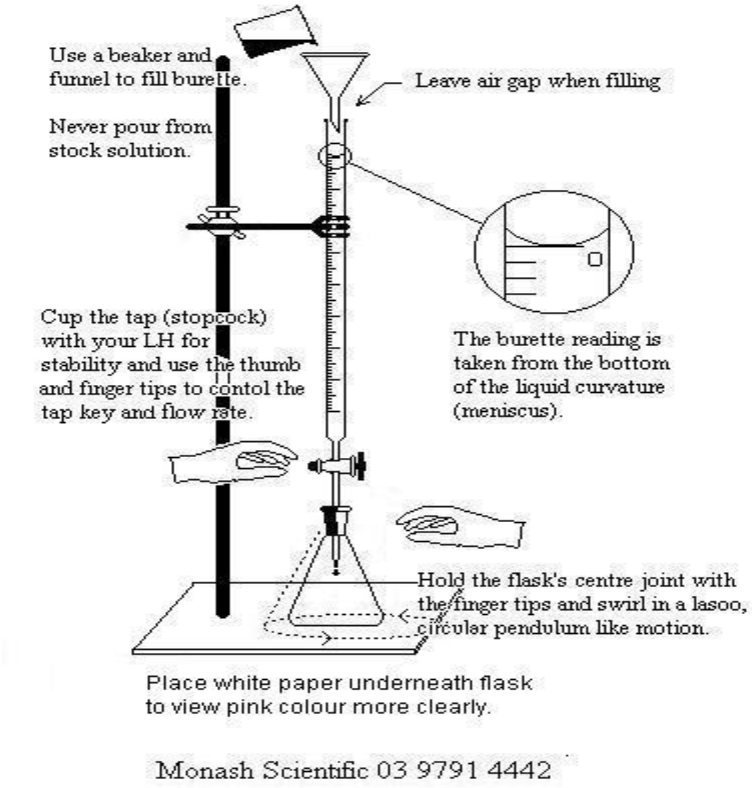
Acid and base samples are analyzed by titration with a standard solution.A weighed portion of sample is dissolved in water and standard acid or base is added to the proper end point. From the volume of reagent used and the weight of sample, the percentage putity of the sample is computed.

The basis for all computions dealing with normalities of solutions is the simple relation that the number of equivalents or milliequivalents of one reaction is equal to the number for the other reacteat thus

A reacts with B

Equivalents A = Equivalents B

Milliequivalents A = Milliequivalents B



Ex/A sample of impure calcite ($CaCO_3$)(100.1 gm /mol) which weighs 0.4950 gm is dissolved in 50.00 ml of standard acid and the excess acid is titrated with 5.25 ml standard base; 1.00 ml of acid is equivalent to 0.005300 gm sodium carbonate; 1.050 ml acid = 1.00 ml base.Calculate the percentage of calcium carbonate in the sample.

Solution/

1 ml of acid ≡ 0.005300 gm Na_2CO_3

$$\text{N of acid} = \frac{5.300\cancel{mg}}{53.00\ \cancel{mg}\ /meq} \times \frac{1}{1\ ml} = 0.1000\ \text{meq / ml}$$

$$\text{Net volume acid required for titration of sample} = 50.00 - \left(5.25 \times \frac{1.050}{1.000}\right)$$

$$= 44.49\text{ml}$$

Milliequivalents acid = Milliequivalents $CaCO_3$

$= 44.49\ \cancel{ml} \times 0.100\ meq\ /\ \cancel{ml} = 4.449\ meq$

$$\text{Milliequivalents } CaCO_3 = \frac{weight\ of\ solute}{eq.wt}$$

Weight $CaCO_3$ = Milliequivalents X eq.wt

$$= 4.449\ \cancel{meq} \times \frac{100.1\ mg/\cancel{mmol}}{2\cancel{meq}\ /\cancel{mmol}}$$

$= 222.7$ mg

$$\text{Percentage } CaCO_3 \text{ in sample} = \frac{222.7}{495.0} \times 100 = 44.99\%$$

Ex/What must be the normality of sodium hydroxide solution if the volume in milliliters used for the titration of a 0.500 gm sample represents that percentage of acetic acid in the sample?

Solution/ 1 eq acid ≡ 1eq base

1 ml of NaOH ≡ 0.5 gm % acetic acid

$$1 \text{ ml of NaOH} \equiv \frac{0.5gm\ x\ 1000mg/gm}{100}$$

1 ml of NaOH ≡ 5.00 mg acetic acid

$CH_3COOH + NaOH \longrightarrow CH_3COONa + H_2O$

$$\text{N of NaOH} = \frac{5.00\cancel{mg}\ /ml}{40\ \ 60.05\ \cancel{mg}\ /meq} = 0.0833\ meq\ /ml$$

Ex/ What is the normality of hydrochloric acid solution having a sodium carbonate titer of a 5.00 mg per milliliter?

Solution/

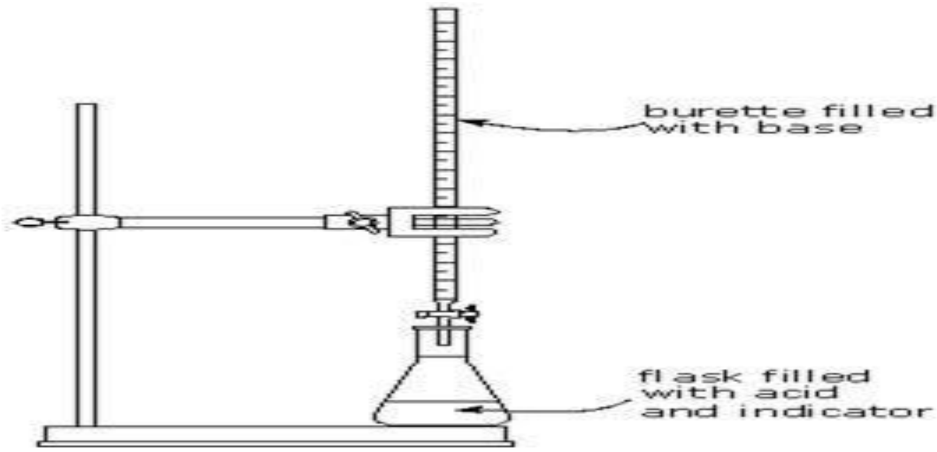
N of base ≡ N ofacid

I ml of acid ≡ 5 mg Na_2CO_3

$$N = \frac{wt}{eq.wt} \times \frac{1}{volume(ml)}$$

$$N = \frac{50\,\cancel{mg}}{\frac{106\,\cancel{mg}/\cancel{mmol}}{2\,meq\,/\,\cancel{mmol}}} \times \frac{1}{1ml} = 0.094 \text{ meq / ml}$$

<u>Titrations in the volumetric analysis</u>



<u>Titration</u> : is an analytical procedure that allows us to measure the amount of a solution reagent of known concentration that is consumed by the analyte.

<u>Titrant</u> : is the solution reagent in buret.

<u>Titrand</u> : is the analyte in beaker.

<u>End point</u> : is the point in which the color of the indicator changes.

<u>Indicator</u> : is a substance (acid or base organic compound) that has one

<u>Calculation of molarities from neutralization reaction (acid-base)</u>

Ex/Exactly 50.00 ml of an HCl solution required 29.71 ml of 0.0193M $Ba(OH)_2$ to reach an end point with bromocresol green indicator. Calculate the molarity of the HCL.

Solution:

$Ba(OH)_2$ +2Hcl $\longrightarrow$ $BaCl_2$ +2H_2O

1mmole of $Ba(OH)_2$≡2mmole of Hcl

Stoichiometric ratio = $\frac{2mmole\ Hcl}{1\ mmol\ Ba(OH)2}$

No.moles $Ba(OH)_2$ = 29.71 ~~ml~~ x 0.01963 mmol / ~~ml~~

=0.583 mmol

No.mmoles HCl = 0.583~~mmol $Ba(OH)_2$~~ x $\frac{2mmol\ HCl}{1\cancel{mmol\ Ba(OH)2}}$

= 1.166 mmol HCl

M_{HCl} = $\frac{1.166mmol\ HCl}{50.0\ ml\ HCl}$ = 0.0233 mmol / ml = 0.0233 M

Ex/Ttration of a sample of an drug was analyzed for aspirin amonoprotic acid ($HC_9H_7O_4$) of 0.500 gm sample of the drug required 21.50 ml of 0.100M NaOH for complete neutralization.

What percentage by mass of the drug was aspirin?

Solution/ $HC_9H_7O_4$ + NaOH $\longrightarrow$ $NaC_9H_7O_4$ + H_2O

1mmol of $HC_9H_7O_4$ ≡ 1mmol of NaOH

No.mmoles NaOH = 21.50 ~~ml~~ x 0.100 mmol / ~~ml~~

= 2.15 ~~mmol~~ x 10^{-3} mol/~~mmol~~ = 2.15 x 10^{-3} mol

1 mol of $HC_9H_7O_4$ ≡ 1 mol NaOH

1 mol $HC_9H_7O_4$ ≡ 180 gm

Mass of aspirin = 2.15 x 10^{-3} ~~mol~~ x 180 gm / ~~mol~~

=0.387 gm

% Aspirin = $\frac{0.387\ \cancel{gm}}{0.500\ \cancel{gm}}$ x 100 =77.4 %

Ex/ Calculate the molarity of the $Ba(OH)_2$ solution if 31.76 ml were needed to neutralize 46.25 ml of 0.1280 M H_2SO_4.

Solution/ $H_2SO_4 + Ba(OH)_2 \longrightarrow BaSO_4 + 2H_2O$

1 mmol $H_2SO_4 \equiv$ 1mmol $Ba(OH)_2$

$$\text{Stoichiometric ratio} = \frac{1\ mmol\ Ba(OH)2}{1\ mmol\ H2SO4}$$

No.mmoles H_2SO_4 = 46.24 ~~ml H_2SO_4~~ X 0.1280 1mmol H_2SO_4 / ~~ml H_2SO_4~~

= 5.92 mmol H_2SO_4

No.mmoles $Ba(OH)_2$ = 5.92 ~~mmol H_2SO_4~~ X $\frac{1\ mmol\ Ba(OH)2}{1\cancel{mmol\ H2SO4}}$

= 5.92 mmol $Ba(OH)_2$

$$M_{Ba(OH)2} = \frac{5.92\ mmol}{31.76\ ml} = 0.1864\ mmol / ml = 0.1864\ M$$