Analytical Chemistry

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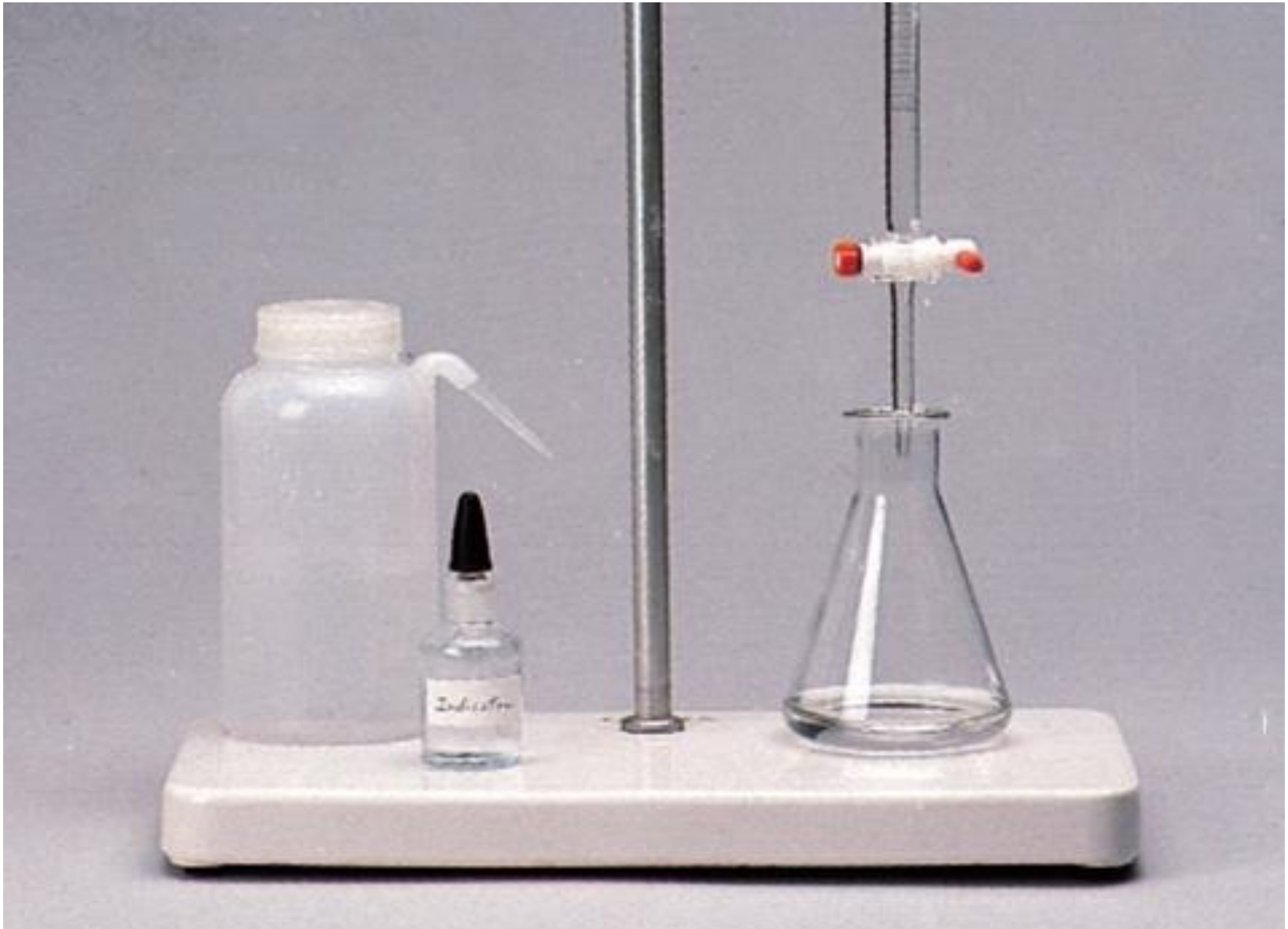
Acid-Base Titrations

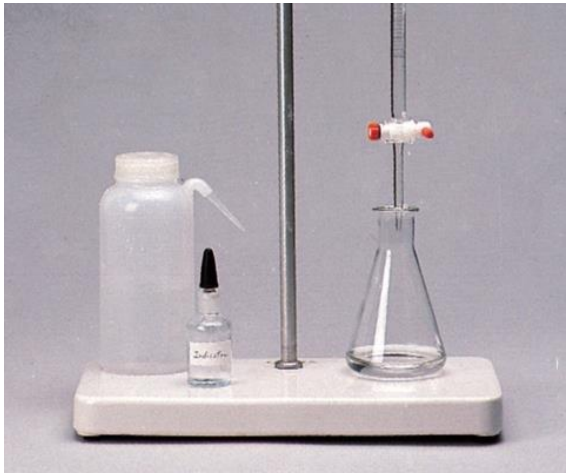


Titrations

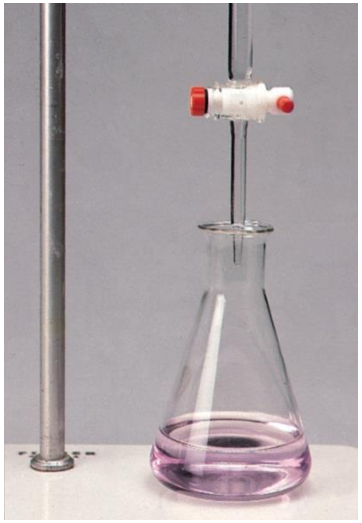
In a ***titration,*** a solution of accurately **known concentration** is gradually added to another solution of **unknown concentration** until the chemical reaction between the two solutions is complete.

Equivalence point – the point at which the reaction is complete

Indicator – substance that changes color (this is called the end point). Indicators must be carefully chosen so that the end point occurs at the equivalence point in the titration.



Slowly add base to unknown acid UNTIL the indicator changes color



- **The equivalence point:** the point when the reactants are done reacting.

The equivalence point is the ideal point for the completion of titration. At the equivalence point the correct amount of standard solution must be added to fully react with the unknown concentration.

- **The end point:** it indicates once the equivalence point has been reached. It is indicated by some form of indicator which varies depending on what type of titration being done. For example, if a color indicator is used, the solution will change color when the titration is at its end point.

Equivalence point & End point **ARE NOT** necessarily equal.

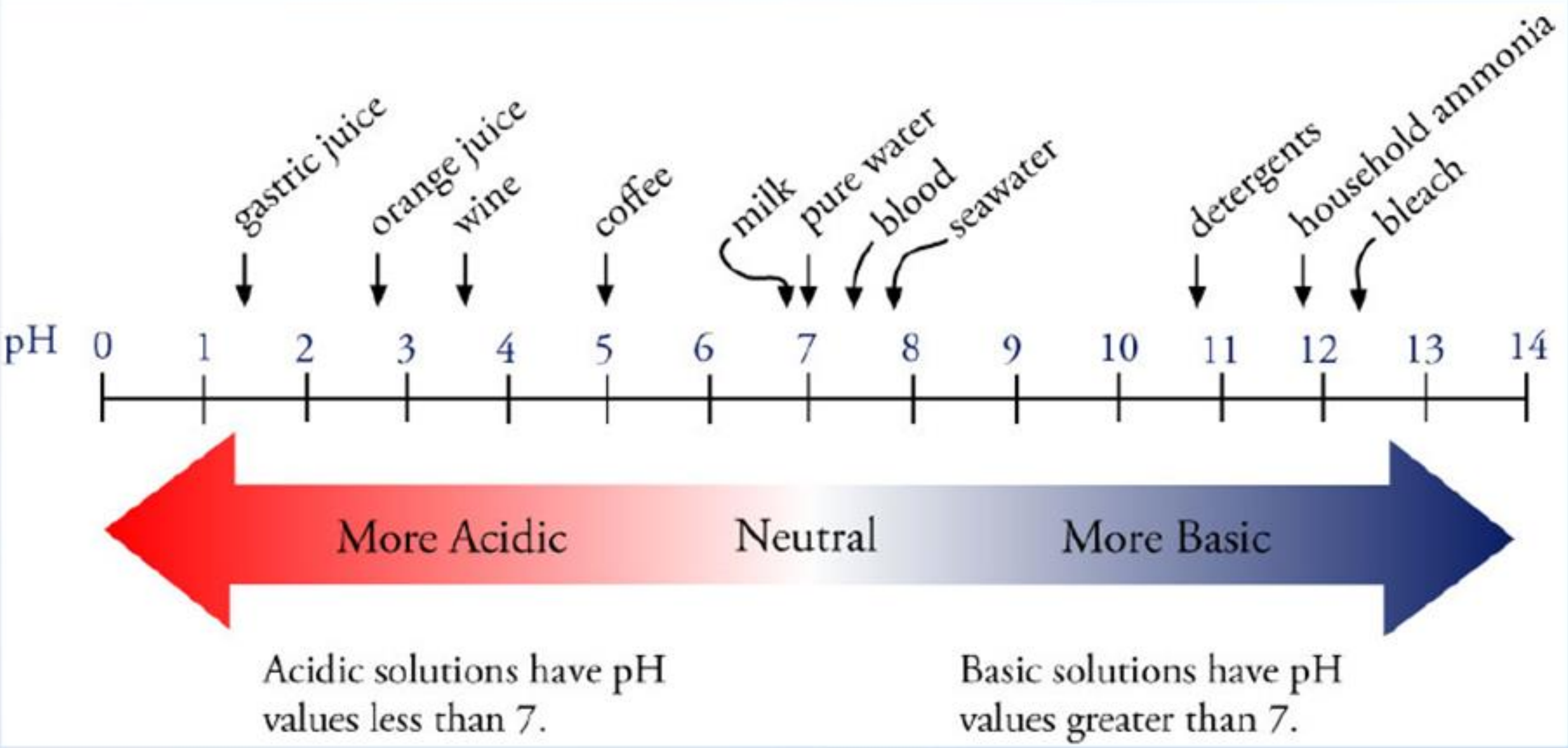
- An endpoint is indicated by some form of indicator at the end of a titration.

- An equivalence point is when the moles of a standard solution (titrant) equal the moles of a solution of unknown concentration (analyte).

pH

- Acidic solutions have pH values less than 7, and the more acidic the solution is, the lower its pH.
- Basic solutions have pH values greater than 7, and the more basic the solution is, the higher its pH.

pH Range


gastric juice
orange juice
wine
coffee
milk
pure water
blood
seawater
detergents
household ammonia
bleach
pH 0 1 2 3 4 5 6 7 8 9 10 11 12 13 14
More Acidic
Neutral
More Basic
Acidic solutions have pH values less than 7.
Basic solutions have pH values greater than 7.

TYPES Of Titrations . . .

- There are many types of titrations with different procedures and goals.

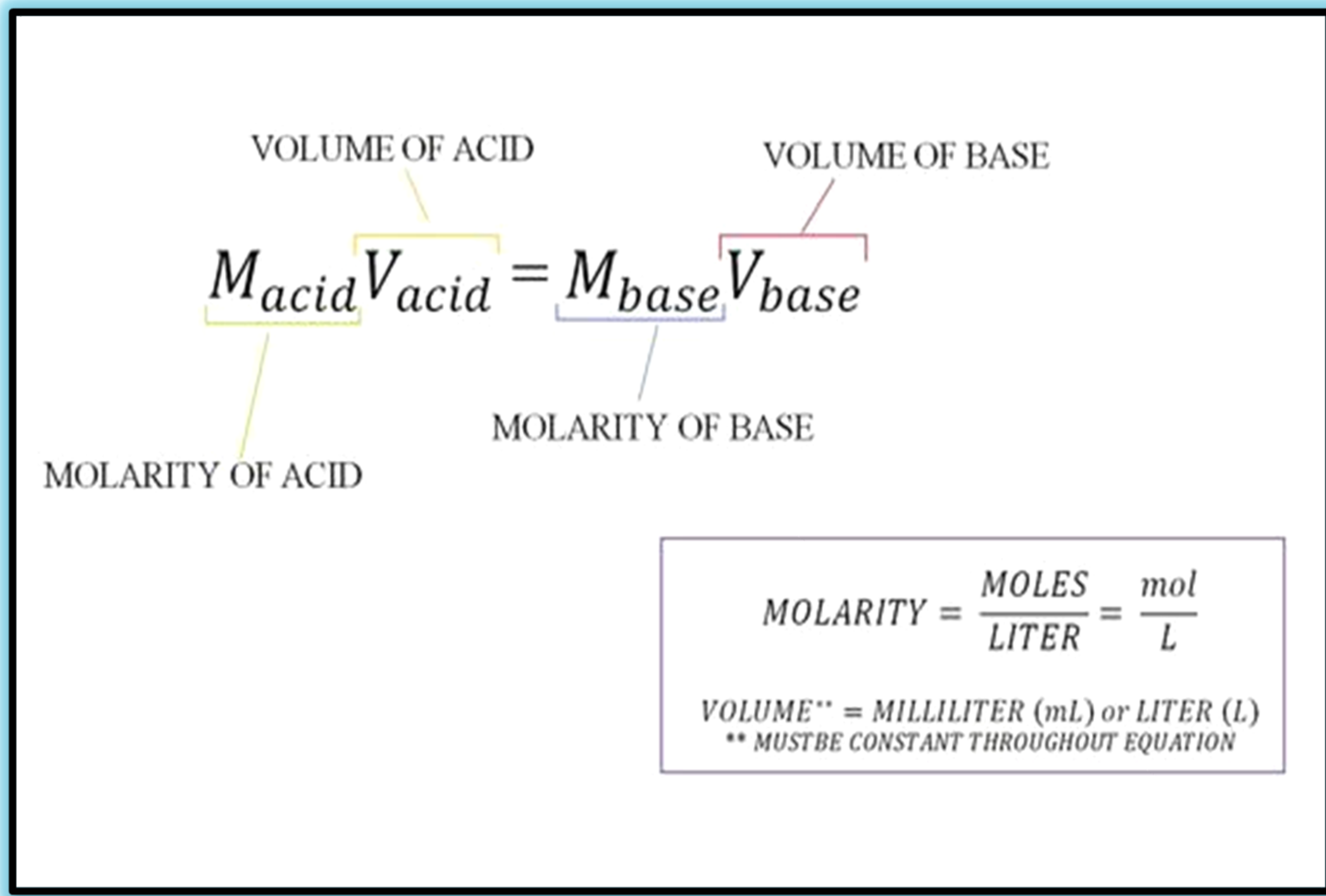
- Acid – Base titration
- Redox titration
- Gas phase titration
- Complexometric titration
- Back titration
- Karl Fischer titration (Potentiometric)

Acid – Base titration:

Acid-base titrations depend on the **neutralization between an acid and a base when mixed in solution.**

- In addition to the sample, an appropriate indicator is added to the titration chamber, reflecting the pH range of the equivalence point.
- The acid-base indicator indicates the endpoint of the titration by changing color.

The final solution after titration should be neutralized and contain equal moles of hydroxide and hydrogen ions. So the moles of acid should equal the moles of base:



Some common Indicators used in acid – base titration:

Indicator	Color on acidic side	Range of color change	Color on basic side
Methyl violet	Yellow	0.0–1.6	Violet
Bromophenol blue	Yellow	3.0–4.6	Blue
Methyl orange	Red	3.1–4.4	Yellow
Methyl red	Red	4.4–6.3	Yellow
Litmus	Red	5.0–8.0	Blue
Bromothymol blue	Yellow	6.0–7.6	Blue
Phenolphthalein	Colorless	8.3–10.0	Pink
Alizarin yellow	Yellow	10.1–12.0	Red

ACID-BASE TITRATION

a Acid + b Base → Salt + Water

$$\frac{M_a V_a}{M_b V_b} = \frac{a}{b}$$

- $M_a V_a$: Molarity and volume of acid
- $M_b V_b$: Molarity and volume of base
- a: Mole of acid
- b: Mole of base

Titration Formula

$$M_A \times V_A = M_B \times V_B$$

Molarity Acid x Volume Acid = Molarity Base x Volume Base

Example: A 24.00 ml sample of an **unknown acid** is titrated to a phenolphthalein end point, which required 31.46 ml of a 0.104 M solution of NaOH. What is the molar concentration of the unknown acid?

i. At the end point (equivalence point) the moles of acid = moles of added base:

ii. $M_A V_A$ = moles acid = moles base = $M_B V_B$

iii. $$M_A = \frac{M_B V_B}{V_A} = \frac{0.104\ \text{M} \times 31.46\ \text{ml}}{24.00\ \text{ml}} = 0.136\ \text{M}$$

What volume of a 1.420 *M* NaOH solution is required to titrate 25.00 mL of a 4.50 *M* H_2SO_4 solution to the eq. point?

Answer: Write the chemical equation:

$$H_2SO_4 + 2NaOH \longrightarrow 2H_2O + Na_2SO_4$$

M_AV_A = moles acid* (H^+) = moles base (OH^-) = M_BV_B

Note that H_2SO_4 is diprotic, so moles H^+ = 2 x moles H_2SO_4

$$\text{volume acid} \xrightarrow[\text{acid}]{M} \text{moles acid} \xrightarrow[\text{ratio}]{\text{mol}} \text{moles base} \xrightarrow[\text{base}]{M} \text{volume base}$$

$$25.00\ \text{mL} \times \frac{4.50\ \text{mol}\ H_2SO_4}{1000\ \text{mL soln}} \times \frac{2\ \text{mol NaOH}}{1\ \text{mol}\ H_2SO_4} \times \frac{1000\ \text{ml soln}}{1.420\ \text{mol NaOH}} = 158.5\ \text{mL}$$

Question 1:

25.0 cm^3 of sulphuric acid is neutralised by 34.0 cm^3 of 0.1 mol of dm^{-3} NaOH. Calculate the concentration of sulphuric acid in:

(a) mol dm^{-3}
(b) g dm^{-3}

[relative atomic mass; H:1, S:32, O:16]

Solution:

Method 1

Step 1 : write down chemical equation

$$NaOH + H_2SO_4 \rightarrow Na_2SO_4 + H_2O$$

Step 2 : find the number of mole NaOH

n=MV

Moles of NaOH

= molarity X Volume (dm^3)

= 0.1 X 0.034

= 0.0034 mol

Step 3 : from the chemical reaction, the ratio of

$$\frac{number\ of\ moles\ of\ H_2SO_4}{number\ of\ moles\ of\ NaOH} = \frac{1}{2}$$

Step 4 : find the number of moles of H_2SO_4 reacted

2 mole of NaOH = 1 mole of H_2SO_4

0.0034 mole of NaOH = $\frac{0.0034 \times 1}{2} mol$

Step 5 : find the concentration of H_2SO_4 in mol dm^{-3}

Concentration = mol/volume
= 0.0017/ 0.025
= 0.068 mol dm^{-3}

Step 6 : find the concentration of H_2SO_4 in g dm^{-3}

Molar mass H_2SO_4 = 1(2) + 32+ 16(4)
= 98 g mol^{-}

Concentration = concentration in mol dm^{-3} x molar mass H_2SO_4
= 0.068 x 98
= 6.664 g dm^{-3}

Method 2:

Step 1 : write down chemical equation

$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$

Step 2 : find the concentration of H_2SO_4 in mol dm^{-3}

$M_a = ?$

$M_b = 0.1$ mol dm^{-3}

$V_a = 25\ cm^3$

$V_b = 34\ cm^3$

$$\frac{M_aV_a}{M_bV_b} = \frac{a}{b}$$

$$\frac{M_a\ (0.025)}{(0.1)\ (0.034)} = \frac{1}{2}$$

$$M_a = 0.068 \text{ mol dm}^{-3}$$

Step 3 : find the concentration of H_2SO_4 in g dm^{-3}

Molar mass H_2SO_4 = 1(2) + 32+ 16(4)
= 98 g mol^{-}

Concentration = concentration in mol dm^{-3} x molar mass H_2SO_4
= 0.068 x 98
= 6.664 g dm^{-3}

Questions

1. What volume of 0.20 mol dm^{-3} nitric acid is required to neutralise 0.14 g of potassium hydroxide? [relative atomic mass: O: 16, K:39, H:1]

2. A 15cm3 of an acid with the formula HaX of 0.1 mol dm-3 required 30 cm3 0f 0.15 mol dm-3 sodium hydroxide solution to complete neutralisation. Calculate the value of **a**.

Ans: 12.5 cm^3/0.0125 dm^3

Example Calculation

A 100 mL portion of wastewater was titrated with the chemical that will be used to neutralize the wastewater. 3.6 mL of the chemical were used. What rate should a chemical feed pump be set at in gallons/minute to neutralize the wastewater at a flow of 45,000 gallons/day?

$$\frac{3.6\text{ mL chemical}}{100\text{ mL wastewater}} = \frac{3.6\text{ gallons chemical}}{100\text{ gallons wastewater}} = \frac{0.036\text{ gal chemical}}{1\text{ gallon wastewater}}$$

$$\frac{45{,}000\text{ gal wastewater}}{\text{day}} \times \frac{0.036\text{ gal chemical}}{\text{gal wastewater}} = \frac{1620\text{ gal chemical}}{\text{day}}$$

$$\frac{1620\text{ gal}}{\text{day}} \times \frac{1\text{ day}}{1440\text{ min}} = 1.125\text{ gal/min}$$

Practice Problem

A 500 mL portion of wastewater was titrated with the chemical that will be used to neutralize the wastewater. 5.8 mL of the chemical were used. What rate should a chemical feed pump be set at in gallons/minute to neutralize the wastewater at a flow of 15,000 gallons/day?

$$\frac{5.8 \text{ mL chemical}}{500 \text{ mL wastewater}} = \frac{5.8 \text{ gallons chemical}}{500 \text{ gallons wastewater}} = \frac{0.0116 \text{ gal chemical}}{1 \text{ gallon wastewater}}$$

$$\frac{15{,}000 \text{ gal wastewater}}{\text{day}} \times \frac{0.0116 \text{ gal chemical}}{\text{gal wastewater}} = \frac{174 \text{ gal chemical}}{\text{day}}$$

$$\frac{174 \text{ gal}}{\text{day}} \times \frac{1 \text{ day}}{1440 \text{ min}} = 0.12 \text{ gal/min}$$

You Try It!

A 50 ml sample of 0.2M HCl is neutralized by 75ml of NaOH. What is the conc. of the base?

<u>HCl</u> **<u>NaOH</u>**

$$0.2M \times 50ml = M_B \times 75ml$$

$M_B = .13M$

Let's Practice

- **Which substance is always a product when an Arrhenius acid in an aqueous solution reacts with an Arrhenius base in an aqueous solution?**

(1) HBr

(2) KBr

(3) H_2O

(4) KOH

- **Which word equation represents a neutralization reaction?**

(1) base + acid → salt + water.

(2) base +salt → water + acid

(3) salt + acid → base + water

(4) salt + water → acid + base

- **What are the products of a reaction between KOH(aq) and HCl(aq)?**

(1) H_2 and KClO

(2) KH and HClO

(3) H_2O and KCl.

(4) KOH and HCl

- **Sulfuric acid, $H_2SO_4(aq)$, can be used to neutralize barium hydroxide, $Ba(OH)_2(aq)$. What is the formula for the salt produced by this neutralization?**

(1) BaS

(3) $BaSO_3$

(2) $BaSO_2$

(4) $BaSO_4$

- **Which compound could serve as a reactant in a neutralization reaction?**

(1) NaCl

(2) CH_3OH

(3) KOH.

(4) CH_3CHO

Which equation represents a neutralization reaction?

(1) $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

(2) $2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell)$

(3) $HNO_3(aq) + KOH(aq) \rightarrow KNO_3(aq) + H_2O(\ell)$

(4) $AgNO_3(aq) + KCl(aq) \rightarrow KNO_3(aq) + AgCl(s)$

You Try It!

How much of a 0.1M H_2SO_4 solution is needed to neutralize 50 ml of a 0.05 KOH solution?

$\underline{H_2SO_4}$ **$\underline{KOH}$**

$$2 \times .1M \times V_A = .05M \times 50ml$$

$$V_A = 12.5\ ml$$

Let’s Practice

- **During which process can 10.0 milliliters of a 0.05 M HCl(aq) solution be used to determine the unknown concentration of a given volume of NaOH(aq) solution?**

(1) evaporation
(2) filtration
(3) distillation
(4) titration.

- **The data collected from a laboratory titration are used to calculate the**

(1) rate of a chemical reaction

(2) heat of a chemical reaction

(3) concentration of a solution.

(4) boiling point of a solution

A student completes a titration by adding 12.0 ml of NaOH(aq) of unknown concentration to 16.0 ml of 0.15 M HCl(aq).

What is the molar concentration of the NaOH(aq)?

(1) 0.11 M

(2) 1.1 M

(3) 0.20 M

(4) 5.0 M

Neutralization and Titration Problems

- What mass of $NaOH_{(s)}$ must be added to 300 mL of HCl 0.25 M in order to completely neutralize this acid? (hint…just find number of moles)

$$HCl + NaOH \longrightarrow NaCl + H_2O$$

0.25 mol = 1 L

X mol = 0.3 L

X = 0.075 mol of HCl in 300 mL of solution

0.075 mol of NaOH will neutralize 0.075 mol of HCl

1 mol of NaOH = 40 g

0.075 mol of NaOH = x g

x = 3 g

3 g of NaOH are used to neutralize 0.3 L of HCl 0.25 M

•During a lab, you mix 2 solutions: a 100 ml solution containing 0.40 g of NaOH and a 100 mL solution containing 0.73 g of HCl. What is the concentration of H^+ ions in the new solution?

$$HCl + NaOH \longrightarrow NaCl + H_2O$$

1 mol of NaOH will neutralize 1 mol of HCl

The total volume of the solution is = 200 mL

There will be an excess of HCl (0.40 g of NaOH to 0.73 g of HCl)

1 mol of NaOH = 40 g

x mol = 0.4 g

x = 0.01 mol

0.73 g of HCl
36.5 g = 1 mol
0.73 g = x mol
x = 0.02 mol

You will have 0.01 mol of HCl that will not have reacted with the NaOH

0.01mol = 0.2 L
x mol = 1 L
X = 0.05 mol
$[H^+]$ = 0.05 M

Calculate the volume of $Ba(OH)_2$ 0.20 M necessary to neutralize 300.0 mL of H_3PO_4 0.30 M.

$M_1V_1 = M_2V_2$

It is $Ba(OH)_2$ So you have the double of OH^- ions.

It is H_3PO_4 so you have three times the amount of H^+ ions 0.30 M.

$(0.2\ M \times 2) \times V_1 = (0.3\ M \times 3) \times 0.3\ L$

$V_1 = 0.675\ L$

The volume is 675 mL.

During an acid-base titration, 23.3 mL of HCl completely neutralized 19.5 mL of potassium hydroxide 0.315 M. Calculate the concentration of the hydrochloric acid.

$$M_a V_a = M_b V_b$$

$$M_a = (0.315\ M)\ (0.0195\ L)\ /0.0233\ L$$

$$= 0.264\ M$$

During an acid-base titration, 25 mL of 0.2 M NaOH were required to neutralize 20 mL of HCl. Calculate the concentration

$$M_1V_1 = M_2V_2$$

$$0.02\ L \times M_i = 0.2\ M \times 0.025\ L$$

$$M_i = 0.25\ M$$

Calculate the volume of 0.5 M NaOH necessary to neutralize 300 mL of 0.2 M HCl.

Answer: V= 120 mL

•To standardize a solution of HCl, 0.1423 g of Na_2CO_3 (eq wt = 53.00) was dissolved in 50 mL of water and titrated with 27.82 mL of HCl. What is the normality of the HCl*(aq)*?

N = eq/L 27.82 mL x (1 L/1000 mL) = 0.02782 L HCl

0.1423 g Na_2CO_3 x (1 eq/53.00 g) x (1 eq acid/1 eq base) = 0.002684 eq

[HCl] = 0.002684 eq/0.02782 L = 0.09651 N

•To standardize a solution of NaOH, 0.5531 g of KHP (eq wt = 204.23) was dissolved in 50 mL of water. This solution required 31.11 mL of NaOH*(aq)* for neutralization. What is the normality of the NaOH*(aq)*?

[NaOH] = eq/L 31.11 mL x (1 L/1000 mL) = 0.03111 L NaOH 0.5531 g KHP x (1 eq/204.23 g) x (1 eq base/1 eq acid) = 0.002708 eq [NaOH] = 0.002708 eq/0.03111 L = 0.08705 N NaOH

A 0.9932 g sample of limestone was titrated with 15.67 mL of 0.113 N HCl, what is the percent of calcium carbonate in the sample?

%CaCO3 = (g CaCO3/g limestone) x 100 g limestone = 0.9932 g Eq wt CaCO3 = 100.09 g/mol x (1 mol/2 eq) = 50.04 g/eq
15.67 mL HCl x (1 L/1000 mL) x (0.113 eq/1 L) x (1 eq CaCO3/1 eq HCl) x (2 mol/1 eq) x (100.09 g/1 mol) = 0.0886 g

Or

15.67 mL HCl x (1 L/1000 mL) x (0.113 eq/1 L) x (1 eq CaCO3/1 eq HCl) x (50.04 g/1 eq) = 0.0886 g
%CaCO3 = 0.0886 g/0.9932 g x 100 = 8.92%

A 27.44 mL of 0.222 N $Ba(OH)_2$ was required to neutralize all the benzoic acid (C6H5COOH) in a 1.224 g sample of organic material. What was the percent benzoic acid in the sample?

BA = benzoic acid

% BA = (g BA/g sample) x 100 g sample = 1.224 g

Eq wt BA = 122.13 g/mol x (1 mol/1 eq) = 122.13 g/eq

27.44 mL $Ba(OH)_2$ x (1 L/1000 mL) x (0.222 eq/1 L) x (1 eq BA/1 eq $Ba(OH)_2$) x (122.13 g/1 eq) = 0.744 g BA

% BA = (0.744 g/1.224g) x 100 = 60.8%

The citric acid in a 0.541 g vitamin tablet was dissolved in 20 mL of 1.021 N NaOH. The excess base was titrated with 9.21 mL of 0.223 N HCl. How many mg of citric acid (eq wt = 64.0) was in the vitamin tablet?

CA = citric acid

Original amount of base = (0.02 L)(1.021 eq/L) = 0.02042 eq base
Excess base = 9.21 mL HCl x (1 L/1000 mL) x (0.223 eq/1L) x (1 eq base/1 eq acid) = 0.00205 eq Base that reacted with CA = 0.02042 – 0.00205 = 0.01837
0.01837 eq x (64.0 g/eq) x (1000 mg/1 g) = 1175 mg = 1.18 x 10^3 mg

A 0.1298 g sample of pure acid was titrated with 42.11 mL of 0.1234 N NaOH. What is the equivalent weight of the acid?

grams acid = 0.1298

Equivalent weight = grams per equivalent

42.11 mL NaOH x (1 L/1000 mL) x (0.1234 eq/L) x (1 eq acid/1 eq base) = 0.005196 eq

Eq wt = 0.1298 g/0.005196 eq = 24.98 g/eq

A 0.3349 g sample of pure base was titrated with 33.33 mL of 0.1673 N H_2SO_4. What is the equivalent weight of the base?

grams base = 0.3349

Equivalent weight = grams per equivalent

33.33 mL x (1 L/1000 mL) x (0.1673 eq/L) x (1 eq base/1 eq acid) = 0.005576 eq

Eq wt = 0.3349 g/0.005576 eq = 60.06 g/eq

•How many milliliters of 0.341 N H_2SO_4 are required to react with 30 mL of 0.333 N Cr(OH)3?

$$N_aV_a = N_bV_b$$

$$(0.341\ N)V = (0.333\ N)(30\ mL)$$

$V = 29.3\ mL$

•What is the normality of a solution of unknown base if 23.45 mL of 0.1236 N HCl are required to neutralize 33.5 mL of it?

$$N_aV_a = N_bV_b$$

$$(0.1236\ N)(23.45\ mL) = N(33.50\ mL)$$

$[base] = 0.08652\ N$

Convert the following to Normality:

a) 0.54 M NH_3

1 mol = 1 eq

0.54 mol/L x (1 eq/1mol) = 0.54 eq/L

0.54 M = 0.54 N

b) 0.43 M $H_2C_2O_4$

1 mol = 2 eq

0.43 M = 0.86 N

c) 0.32 M H_3AsO_4

1 mol = 3 eq

0.32 M = 0.96 N

•Convert the following to Equivalents

a) 0.12 mol H_2SO_4

for H_2SO_4 with 2 acidic protons (H^+), 1 mol = 2 eq

0.12 mol x (2 eq/1 mol) = 0.12 eq

0.12 mol = 0.24 eq

b) 0.34 mol $NaHCO_3$

for NaHCO3 which can accept one proton, 1 mol = 1 eq

0.34 mol = 0.34 eq

c) 0.56 mol $Cr(OH)_3$

for $Cr(OH)_3$ which can react with 3 protons, 1 mol = 3 eq

0.56 mol = 1.7 eq

- **Convert the following to Moles.**

a) 0.98 eq H_2S

for H_2S which has 2 acidic H^+, 1 mol = 2 eq

0.98 eq x (1 mol/2 eq) = 0.48 eq

0.98 eq = 0.48 mol

b) 0.763 eq CH_3COOH

acetic acid has one acidic proton (on the end) so

1 mol = 1 eq

0.763 eq = 0.763 mol

c) 0.5647 eq $Ba(OH)_2$

barium hydroxide reacts with 2 protons, 1 mol = 2 eq

0.5647 eq = 0.2824

- **<u>Convert the following to Molarity.</u>**

<u>a) 0.211 N CH_3NH_2</u>

1 mol = 1 eq

0.211 eq/L x (1 mol/1 eq) = 0.211 mol/L

0.211 N = 0.211 M

<u>b. 0.1951 N $Ca(OH)_2$</u>

1 mol = 2 eq

0.1951 N = 0.09755 M

<u>c) 0.334 N C6H5COOH</u>

1 mol = 1 eq

0.344 N = 0.344 M

• **What is the equivalent weight of each of the following substances?**

a) NH_3

Molar mass = 14.01 + 3(1.01) = 17.04 g/mol

For ammonia, 1 mol = 1 eq

17.04 g/mol x (1 mol/1 eq) = 17.04 g/eq

b. $Ca(OH)_2$

molar mass = 40.08 + 2(16.00 + 1.01) = 74.10 g/mol

for calcium hydroxide, 1 mol = 2 eq

74.10 g/mol x (1 mol/2 eq) = 37.05 g/eq

c. H_3AsO_4

molar mass = 3(1.01) + 74.92 + 4(16.00) = 141.95 g/mol for arsenic acid,

1 mol = 3 eq

141.95 g/mol x (1 mol/3 eq) = 47.32 g/eq

d. H_2S

molar mass = 2(1.01) + 32.07 = 34.09 g/mol

for hydro sulfuric acid, 1 mol = 2 eq

34.09 g/mol x (1 mol/2 eq) = 17.04 g/eq

What is the normality of a solution of unknown base if 23.45 mL of 0.1236 N HCl are required to neutralize 33.5 mL of it?

NaVa = NbVb

(0.1236 N)(23.45 mL) = N(33.50 mL) Vb

[base] = 0.08652 N

Equation for the titration

$$2HCl + Na_2CO_3 \rightarrow 2NaCl + H_2O + CO_2$$

Answer

- **Volume of the acid = 19.2 ml**
- **Factor for the acid = 2 (the number in front of HCl in the balanced equation)**
- **Molarity of the acid = ?**
- **Volume of the base = 20 ml**
- **Factor for the base = 1 (the number in front of sodium carbonate in the balanced equation)**
- **Molarity of the base = 0.048 M**

To make up an approximate solution of sodium hydroxide and standardize it (find its exact concentration) by titration with the standard hydrochloric acid solution above

Equation for the titration

$$NaOH + HCl \rightarrow NaCl + H_2O$$

Answer

- Volume of base = 20 ml
- Factor for the base = 1
- Molarity of the base = ?
- Volume of the acid = 19.8 ml
- Factor for the acid = 1
- Molarity of the acid = 0.1 M

M_a = 0.1 moles per litre
= 0.1 Molar = 0.1 M

Calculation

$$V_a M_a = V_b M_b$$

$$19.8 * 0.1 = 20 * M_b$$

$$M_b = \frac{19.8 * 0.1}{20}$$

- **M_b = 0.099 moles/L = 0.099 M**

To determine the percentage of ethanoic acid in vinegar.

Equation for titration

$$CH_3COOH + NaOH \longrightarrow CH_3COONa + H_2O$$

Results

- Volume of acid used = 13 ml
- Factor for the acid = 1
- Molarity of the acid = ?
- Volume of base used = 20 ml
- Factor for the base = 1
- Molarity = 0.1 M

Titration readings

Titration number	1	2	3	4
Final burette reading /cm^3	25.2	24.8	33.3	24.9
Initial burette reading /cm^3	0.0	0.0	7.4	0.1
Volume of NaOH used /cm^3	25.2	24.8 √	25.9 √	24.8
Best titration results (√)				

Titration of a known acid with an alkali

Suppose that in an experiment, you are asked to find the concentration of a solution of sulphuric acid by titrating <u>25.0 cm^3</u> of the acid against a standard solution of sodium hydroxide of concentration <u>0.100 mol/dm^3</u>, using phenolphthalein as an indicator

First set up the apparatus as shown in the diagram and then carry out the titration, repeating it as many times as necessary to obtain a set of consistent results

Results

Suppose the following readings are obtained:

Titration number	1	2	3	4
Final burette reading /cm^3	25.2	24.8	33.3	24.9
Initial burette reading /cm^3	0.0	0.0	7.4	0.1
Volume of NaOH used /cm^3	25.2	24.8	25.9	24.8
Best titration results (√)		√		√

Mean volume of sodium hydroxide used = <u>24.8 cm^3</u>

Titration of a known acid with an alkali

- You are then asked to calculate the concentration of the sulphuric acid from your results.
- The equation for the reaction is:

$$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$$

- From the equation,

$$\frac{\text{No. of moles of } H_2SO_4}{\text{No. of moles of NaOH}} = \frac{1}{2}$$

$$\frac{\text{Vol. of } H_2SO_4 \times \text{Conc. of } H_2SO_4}{\text{Vol. of NaOH} \times \text{Conc. of NaOH}} = \frac{1}{2}$$

$$\frac{25.0 \times \text{Conc. of } H_2SO_4}{24.8 \times 0.100 \text{ mol/dm}^3} = \frac{1}{2}$$

Therefore,

$$\text{Conc. of } H_2SO_4 = \frac{1 \times 24.8 \times 0.100 \text{ mol/dm}^3}{2 \times 25.0}$$

$$= \underline{0.0496 \text{ mol/dm}^3}$$

Titration of an unknown acid with an alkali

Aim:

You are provided with a solution containing 5.00 g/dm^3 of the acid H_3XO_4. You are to find the relative molecular mass of the acid by titrating 25.0 cm^3 portions of the acid with the standard (0.100 mol/dm^3) sodium hydroxide solution, and hence find the relative atomic mass of element X.

The equation for the reaction is:

$$H_3XO_4 + 2NaOH \rightarrow Na_2HXO_4 + 2H_2O$$

Results:

Titration No.	1	2	3	4
Final reading/ cm^3	25.4	25.5	25.6	35.8
Initial reading/ cm^3	0.0	0.0	0.0	10.0
Volume of NaOH/ cm^3	25.4	25.5	25.6	25.8

- **Average volume of NaOH = 25.5 cm^3**

From the equation,

$$\frac{\text{No. of moles of } H_3XO_4}{\text{No. of moles of NaOH}} = \frac{1}{2}$$

$$\frac{\text{Vol. of } H_3XO_4 \times \text{Conc. of } H_3XO_4}{\text{Vol. of NaOH} \times \text{Conc. of NaOH}} = \frac{1}{2}$$

$$\frac{25.0 \times \text{Conc. of } H_3XO_4}{25.5 \times 0.100 \text{ mol/dm}^3} = \frac{1}{2}$$

$$\text{Therefore, conc. of } H_3XO_4 = \frac{1 \times 25.5 \times 0.100}{2 \times 25.0}$$

$$= \underline{0.0510 \text{ mol/dm}^3}$$

Since 1 dm^3 of the acid contains 5.00 g of the acid, therefore

$$0.0510 \times M_r \text{ of } H_3XO_4 = 5.00\text{ g}$$

$$M_r \text{ of } H_3XO_4 = \frac{5.00}{0.0510} = \underline{98.0}$$

Calculate the relative atomic mass of X:

$$1 \times 3 + X + 16 \times 4 = 98$$

$$X = 98 - 67$$

$$= \underline{31}$$

Quick Check

25.0 cm^3 samples of sodium hydroxide solution are titrated against hydrochloric acid which has a concentration of 0.225 mol/dm^3.

The results obtained are shown in the table below

Titration No.	1	2	3	4
Final burette reading/ cm^3	24.4	48.9	23.6	48.0
Initial burette reading/ cm^3	0.0	24.4	0.0	23.6
Volume of HCl/ cm^3				
Best titration result (√)				

(a) Complete the table above.

(b) Calculate the concentration of the sodium hydroxide solution

1. **A titration flask contains 25.0 cm^3 of sodium hydroxide and a few drops of phenolphthalein as indicator. It is titrated against hydrochloric acid contained in a burette. What colour change would you observe when the end point is reached?**

(A) colorless to light pink (B) light pink to colorless
(C) red to colorless

Solution to Quick check

Titration No.	1	2	3	4
Final burette reading/ cm^3	24.4	48.9	23.6	48.0
Initial burette reading/ cm^3	0.0	24.4	0.0	23.6
Volume of HCl/ cm^3	24.4	24.5	23.6	24.4
Best titration result (√)	√			√

Average volume of HCl used = 24.4 cm^3

Equation: $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

$$\frac{25.0 \times \text{Conc. of NaOH}}{24.4 \times 0.225 \text{ mol/dm}^3} = \frac{1}{1}$$

$$\text{Conc. of NaOH} = \frac{24.4 \times 0.225 \text{ mol/dm}^3}{25.0}$$

$$= \underline{0.220 \text{ mol/dm}^3}$$

Prepare 100mL of 1.0M hydrochloric acid, HCl from concentrated (1.21M) hydrochloric acid, HCl.

Solution:

$$M_1V_1 = M_2V_2$$

$$(12.1M)(V_1) = (1.0M)(100mL)$$

$$V_1 = 8.26 \text{ mL conc. HCl}$$

Thus, we need 8.26mL from stock of 1.21M HCl to prepared 100mL of a 1.0M of HCl

– **Q: A solution is labelled 0.15M HCl. How many grams of HCl are present in 1 litre?**

A: M_r of HCl = 36.5g

1M solution contains 36.5g / l

0.15M solution = 36.5 x 0.15 = 5.475g

– **Q: A solution contains 4.9g H_2SO_4 in 100cm^3. Calculate the concentration in mol/l.**

A: M_r of H_2SO_4 = 98g

4.9 ÷ 98 = 0.05mol in 100cm^3

0.05 x 10 = 0.5mol/l

- **Q: What volume of 18M HCl would be required to prepare 250cm³ of 0.5M HCl?**

A: $$\frac{V_{dil} \times M_{dil}}{1000} = \frac{V_{conc} \times M_{conc}}{1000}$$

$$\frac{250 \times 0.5}{\cancel{1000}} = \frac{V_{conc} \times 18}{\cancel{1000}}$$

$$\frac{250 \times 0.5}{18} = V_{conc}$$

$$6.94\,cm^3 = V_{conc}$$