7.2 ACID-BASE TITRATIONS

Objective

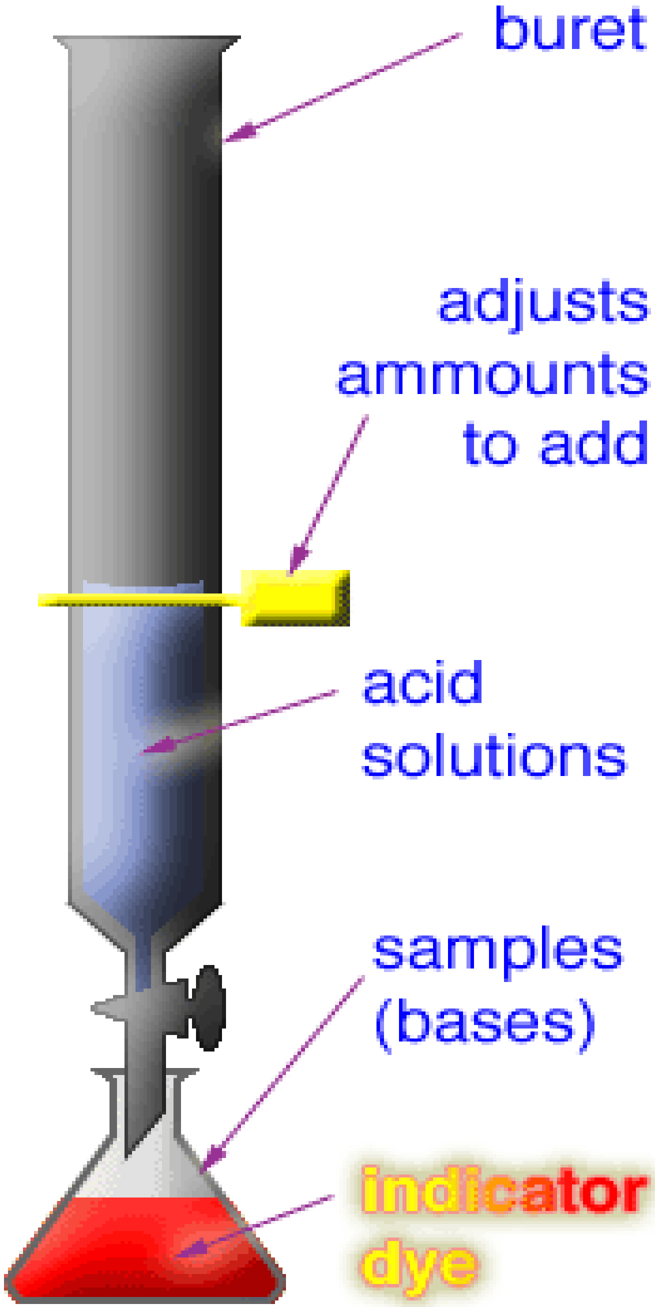
At the end of this topic, students should be able to:

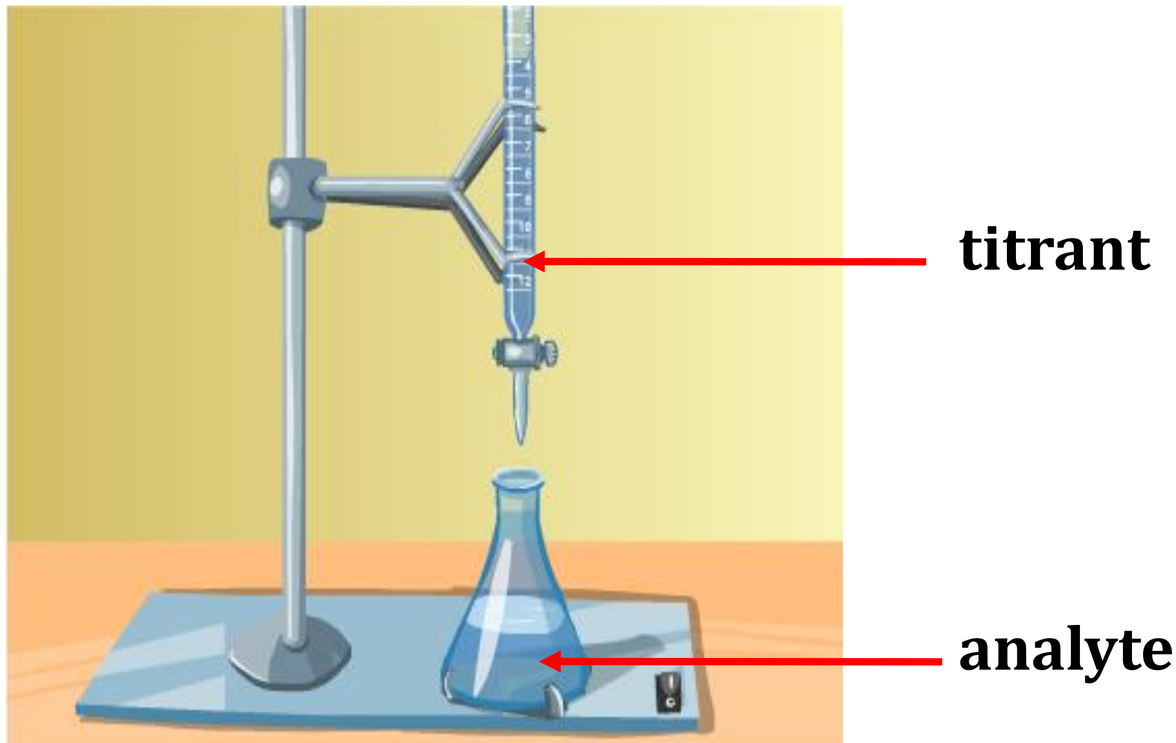
a) Describe the titration process and distinguish between the end point and equivalent point.
b) Perform calculations involving titration between a strong acid and a strong base.
c) Sketch and interpret the variation of pH against titre value for titrations between:
   - strong acid and strong base
   - strong acid and weak base
   - weak acid and strong base
d) Identify suitable indicators for acid-base

ACID – BASE TITRATION

TERMS	DEFINITIONS
Equivalence point	Point at which both acid & base exactly neutralise each other (no.of moles of OH^- = no. of moles of H^+)
End point	Point at which the indicators changes colour
Indicator	Organic dyes whose colour depends on the pH of the solution
Titrant	Solution added from the burette

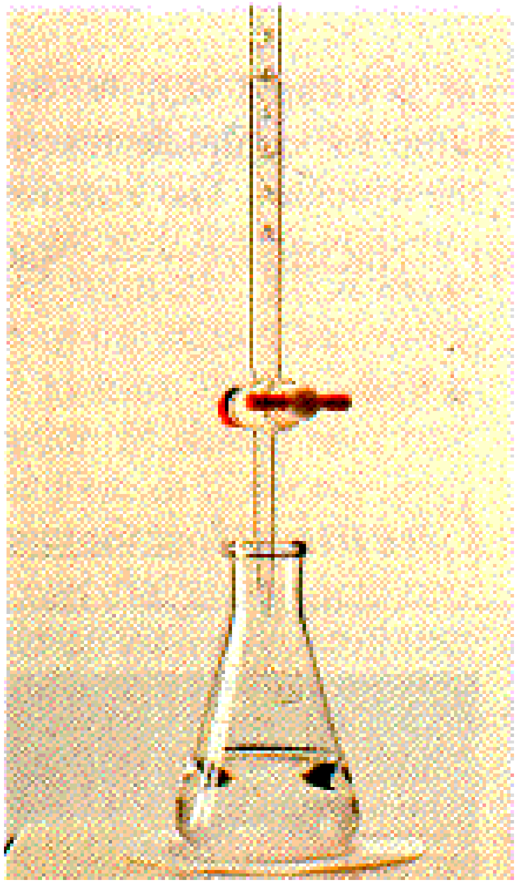
ACID – BASE TITRATION



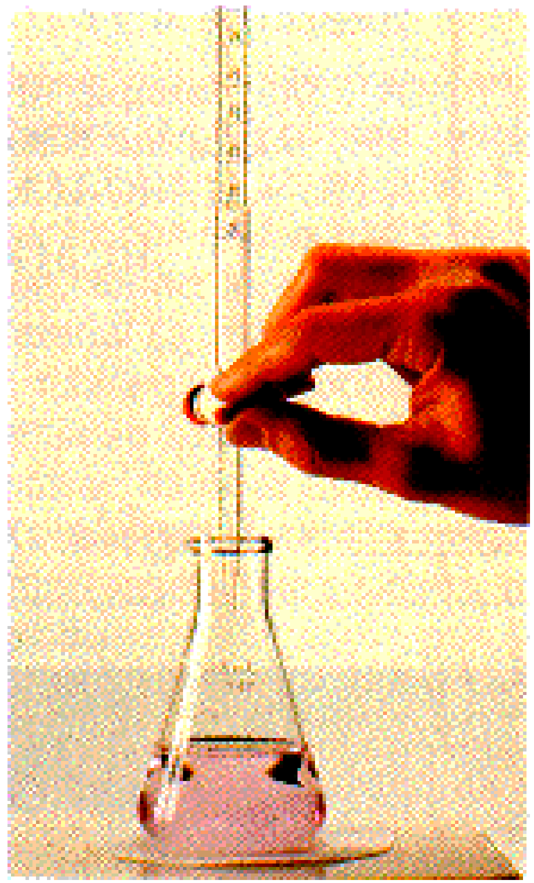


Titration apparatus

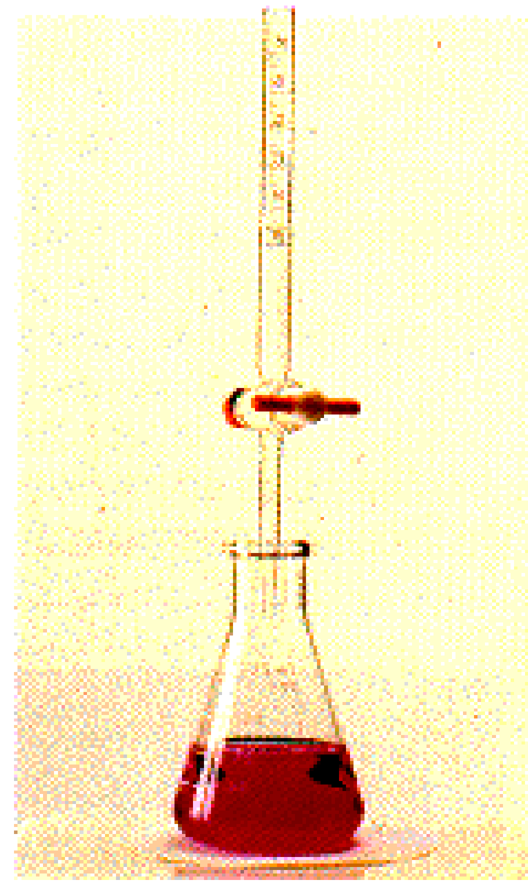
Titration Endpoint



before endpoint



endpoint



past endpoint

- **Titration :** analytical technique to determine the volume of an acid (or base) of known concentration that is necessary to exactly neutralise a sample of base (or acid)
- The **equivalence point** can be determined by :
  - **pH measurement**
  - **indicator**
- **2 types of** acid-base **indicators** :
  - **weak acid indicator**
  - **weak base indicator**

INDICATOR

- To be an **effective indicator**, the acid and it's conjugate base must **have distinctive colour**
- General example :

$$HIn\,(aq) \rightleftharpoons H^+\,(aq) + In^-\,(aq)$$

acid
(X colour)

conjugate base
(Y colour)

- **in acidic medium :**
  - $[H^+]$ is high
  - equilibrium position shift to the left
  - **x colour** appear

- **in basic medium :**
  - $[H^+]$ is low
  - equilibrium position shift to right
  - **y colour** appear

CHOOSING AN INDICATOR

- Main objective of titration :
  - **to match** the end point with the equivalence point
  - therefore the determination of solution's molarity is accurate
- Matching is achieved by choosing a suitable indicator i.e. colour changes of the indicator occurs over a pH range which includes the pH of the equivalence point

TYPES OF TITRATION

TITRATION	*EQUIVALENCE POINT*	*RANGE of pH*	*SUITABLE INDICATOR*
Strong acid-strong base	pH 7	3 – 10	Any indicator
Strong acid-weak base	pH < 7	3 – 11	Methyl orange Methyl red
Weak acid-strong base	pH > 7	7 – 11	Phenolphthalein Thymol blue
Weak acid-weak base	Not obvious	Not obvious	Not obvious

pH Calculation for Acid-Base Titration

1. Strong Acid – Strong Base Titrations

Example 1:

Consider the addition of 0.10 M NaOH solution (from a burette) to an Erlenmeyer flask containing 25.00 mL of 0.10 M HCl. Calculate the pH of the solution :

a) before the titration begin (before the addition of NaOH)

b) after the additon of 24.00 mL of 0.10M NaOH

c) after the addition of 25.00 mL of 0.10M NaOH

d) after the addition of 35.00 mL of 0.10M NaOH

Solution:

a) Before the titration begins, only HCl contained in the Erlenmeyer flask. HCl is a strong acid, therefore it ionizes completely.

Initial concentration of HCl = 0.10 M

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

$[H^+] = 0.10$ M

b) After addition of 24.00 mL 0.10 M NaOH

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

n_o			
n_Δ			
n_{final}			
$[\ ]_{final}$			

V_{total} =

=

The pH solution is calculated from the amount of HCl left after partial neutralization

c) After addition of 25.00 mL 0.10 M NaOH

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

n_o			
n_Δ			
n_{final}			
$[\ ]_{final}$			

V_{total} =

=

The calculation involves a complete neutralization reaction:

d) After addition of 35.00 mL 0.10 M NaOH

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

n_o			
n_Δ			
n_{final}			
$[\ ]_{final}$			

V_{total} =

=

pH solution is determined from the amount of NaOH left;

Example 2:

A 25.00 mL sample of 0.10 M HCl is titrated with 0.1 M NaOH.

Calculate the pH of the solution:

i. before the addition of NaOH

ii. after the addition of 10.0 mL of NaOH

iii. after the addition of 24.9 mL of NaOH

iv. at the equivalence point

v. after the addition 25.1 mL of NaOH

vi. after the addition of 35.0 mL of NaOH

Sketch the titration curve

Solution:

i. pH before the addition of 0.10 M NaOH

Dissociation equation of HCl :

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

ii. pH after the addition of 10.0 mL of 0.10 M NaOH

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

n_o			
n_Δ			
n_{final}			
$[\ ]_{final}$			

V_{total} =

=

The pH solution is calculated from the amount of HCl left after partial neutralization;

iii. pH after the addition of 24.9 mL of 0.10 M NaOH

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

n_o			
n_Δ			
n_{final}			
$[\ \ ]_{final}$			

V_{total} =
=

The pH solution is calculated from the amount of HCl left after partial neutralization;

iv. pH at the equivalence point

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

n_o			
n_Δ			
n_{final}			
$[\ ]_{final}$			

V_{total} =

=

The calculation involves a complete neutralization reaction.

$$NaCl_{\,(aq)} \rightarrow \underbrace{Na^+_{(aq)} + Cl^-_{(aq)}}_{}$$

(does not undergo hydrolysis)

The pH solution is calculated from the dissociation of water.

v. pH after the addition of 25.1 mL of 0.10 M NaOH

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

n_o			
n_Δ			
n_{final}			
$[\ ]_{final}$			

V_{total} =

=

The pH solution is determined from the amount of NaOH left.

vi. pH after the addition of 35.0 mL of 0.10 M NaOH

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

n_o			
n_Δ			
n_{final}			
$[\ \ ]_{final}$			

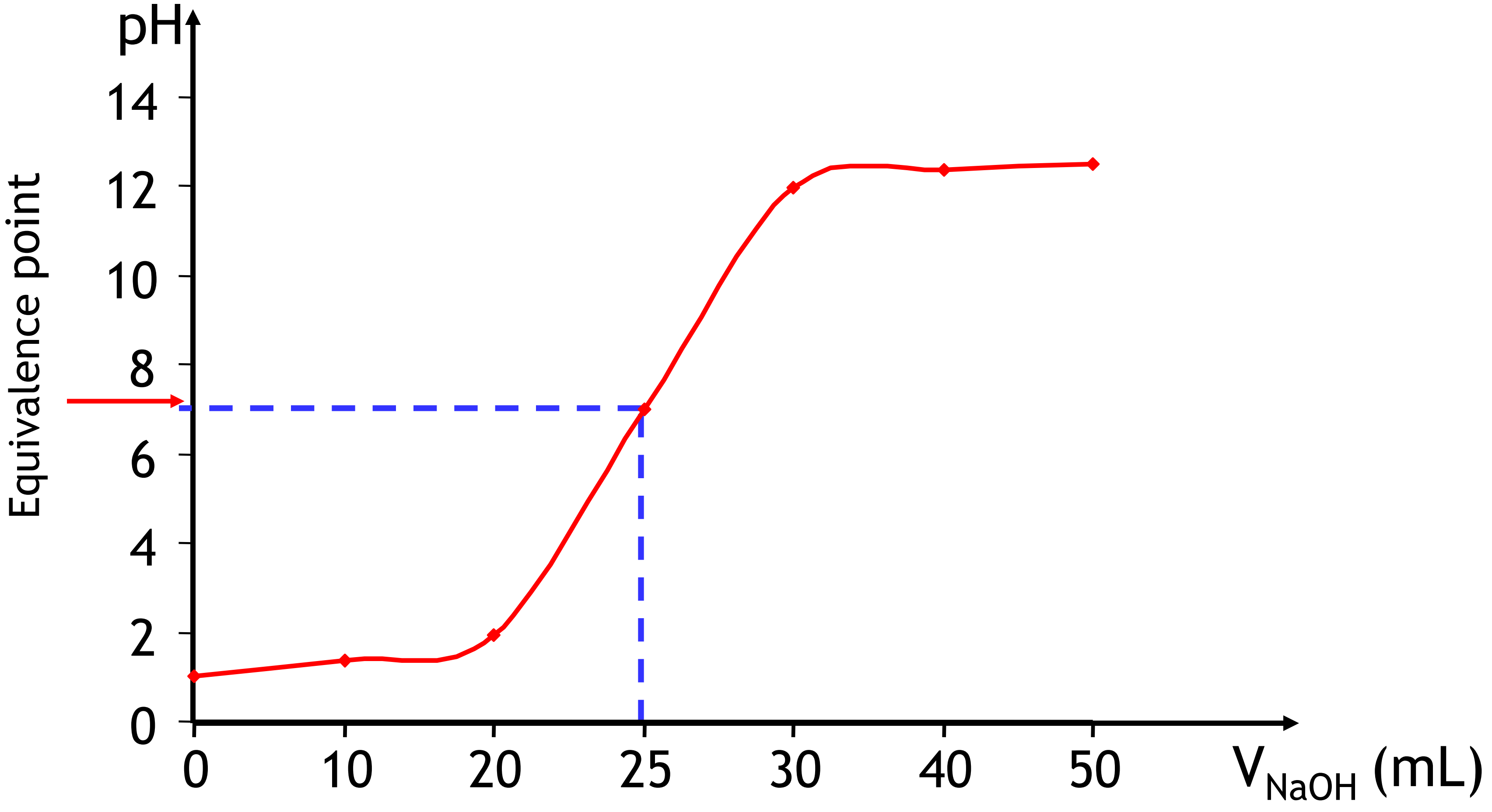
V_{total} =

=

The pH solution is determined from the amount of NaOH left.

The titration curve for strong acid-strong base titration

NaOH (mL)	0	10	24.9	25	25.1	35
pH	1.0	1.37	3.7	7	10.3	12.2



THE ACID-BASE TITRATION CURVE

- Is a graph of pH versus volume of the titrant
- Steps in sketching a titration curve
  - Calculate the pH of the solution in conical flask
  - Calculate the volume of titrant required to neutralise the solution in the flask
    - Can be determined from the chemical equation for the neutralisation reaction
  - Determine the type of titration, therefore the pH change at equivalence point can be stated
  - Draw a sharp vertical line to show the sudden pH change at end point

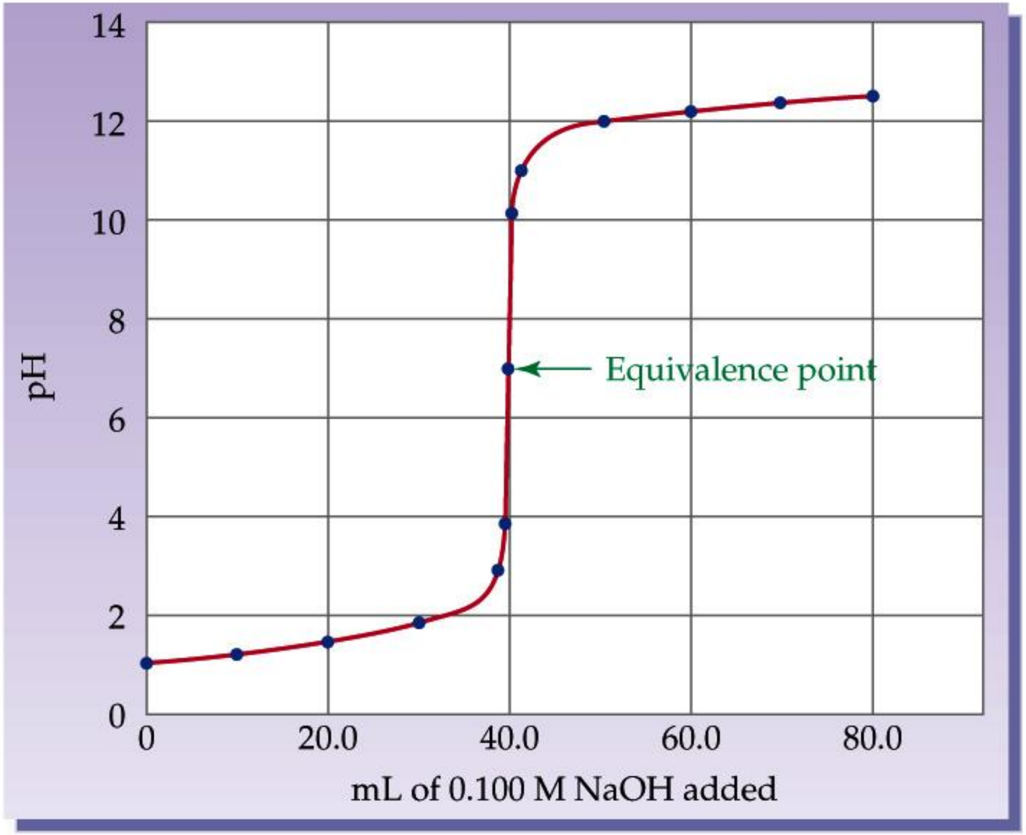
1. STRONG ACID-STRONG BASE TITRATION

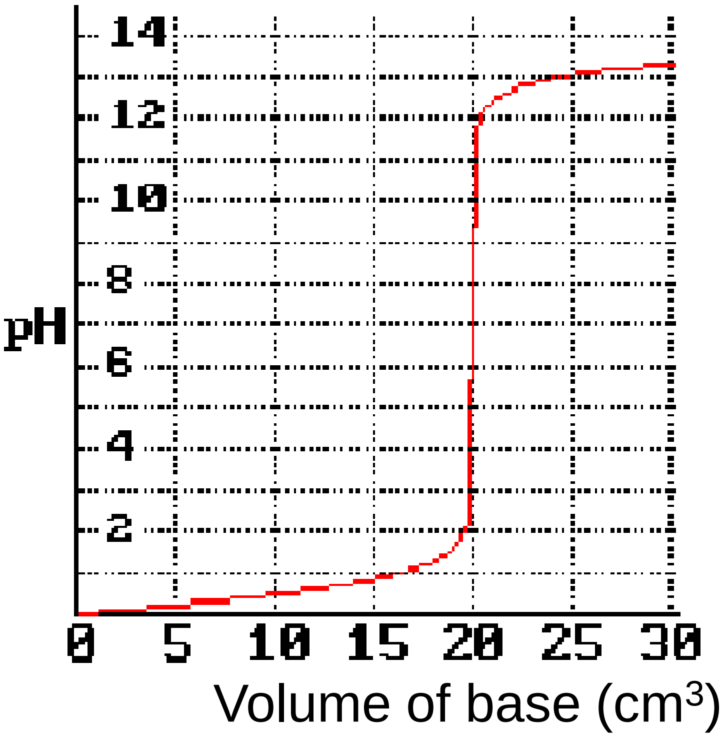
$$NaOH_{(aq)} + HCl_{(aq)} \longrightarrow H_2O_{(l)} + NaCl_{(aq)}$$

The net ionic equation,

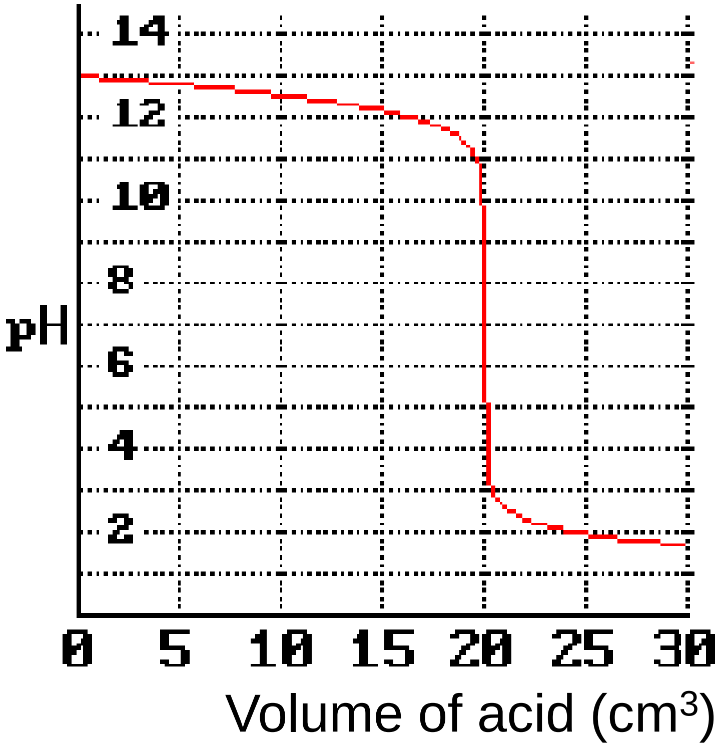
$$H_2O_{(l)} \longrightarrow H^+_{(aq)} + OH^-_{(aq)}$$

The Titration Curve





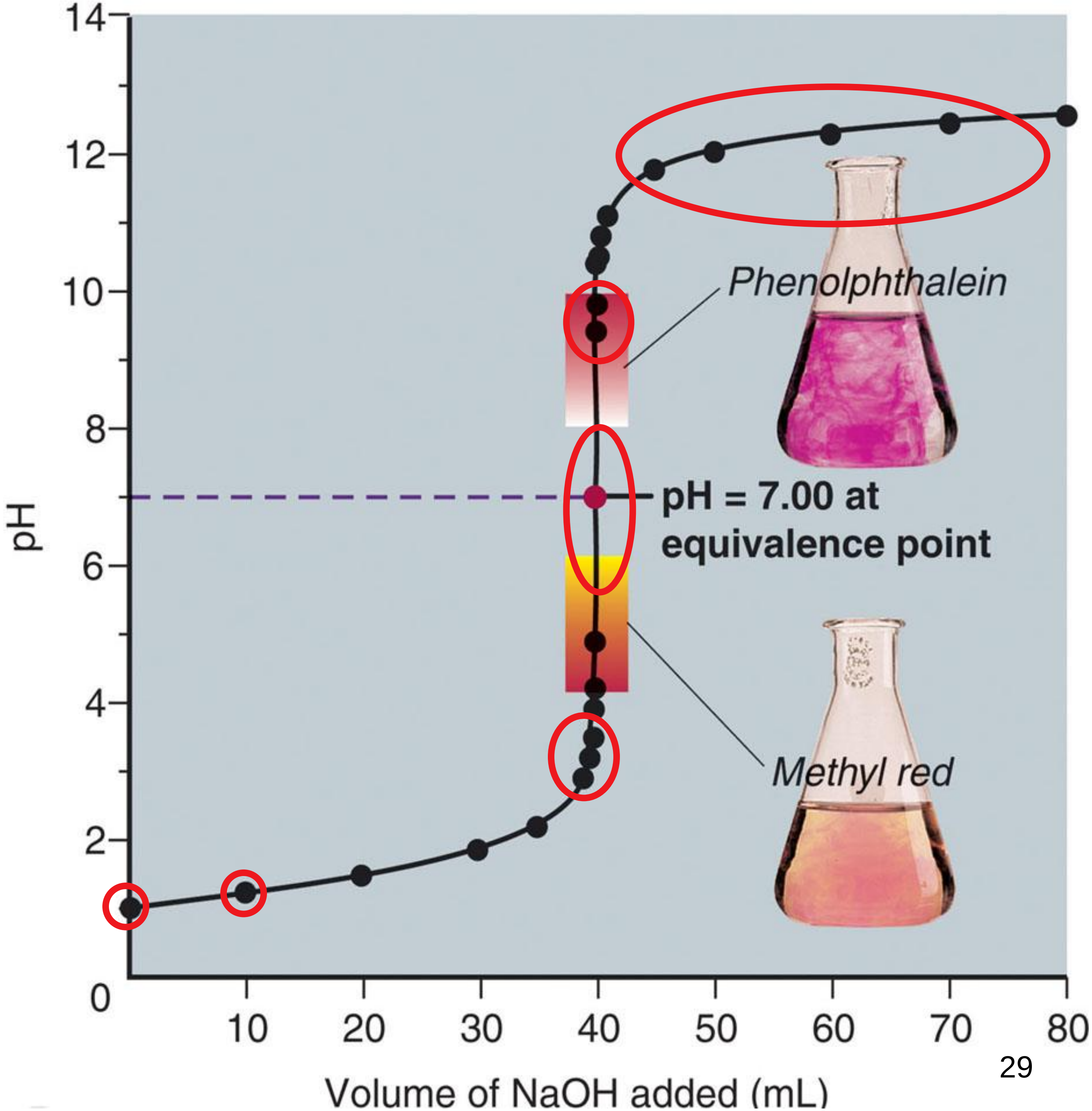
Strong base added to strong acid

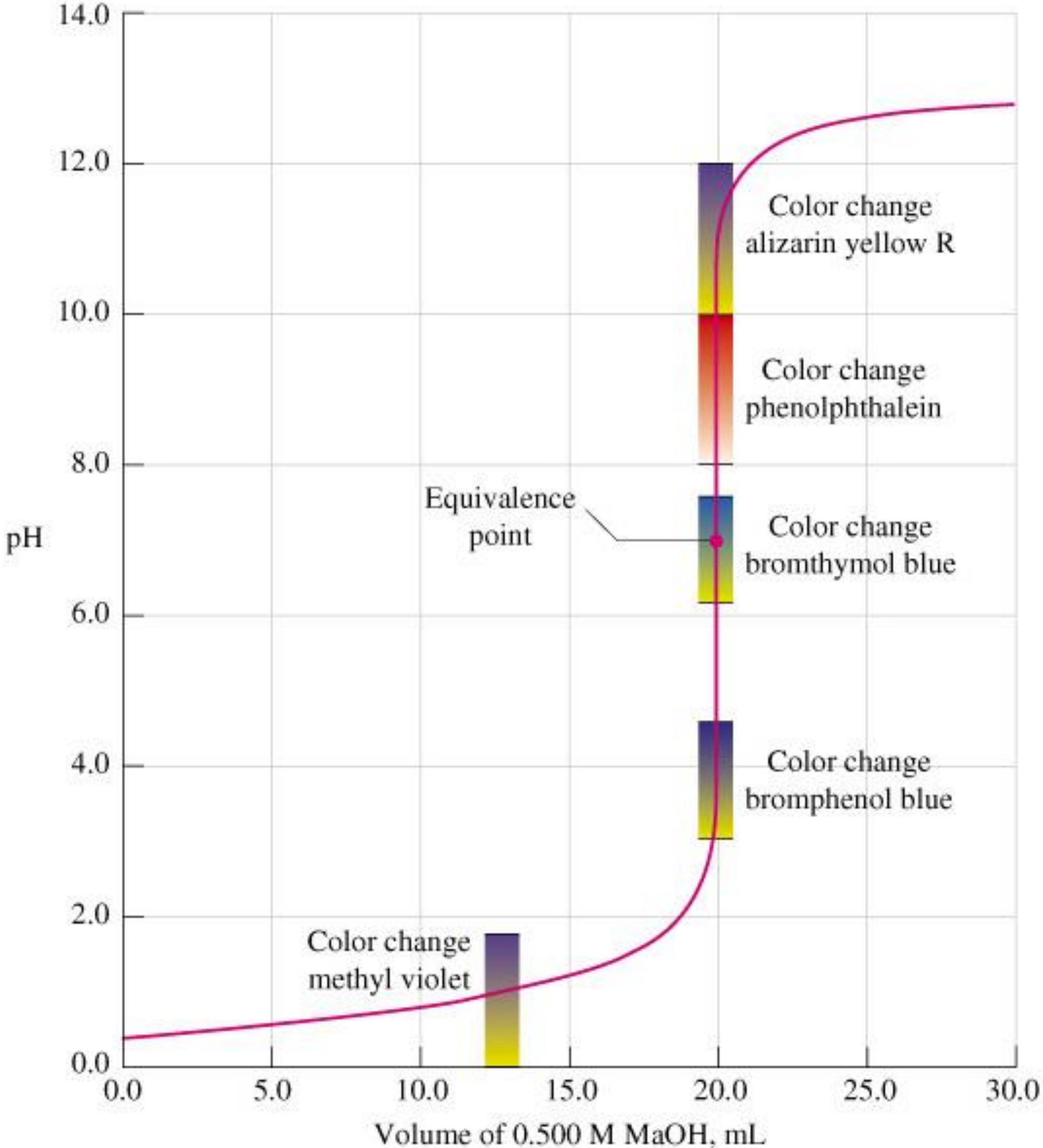


Strong acid added to strong base

Curve for a strong acid-strong base titration

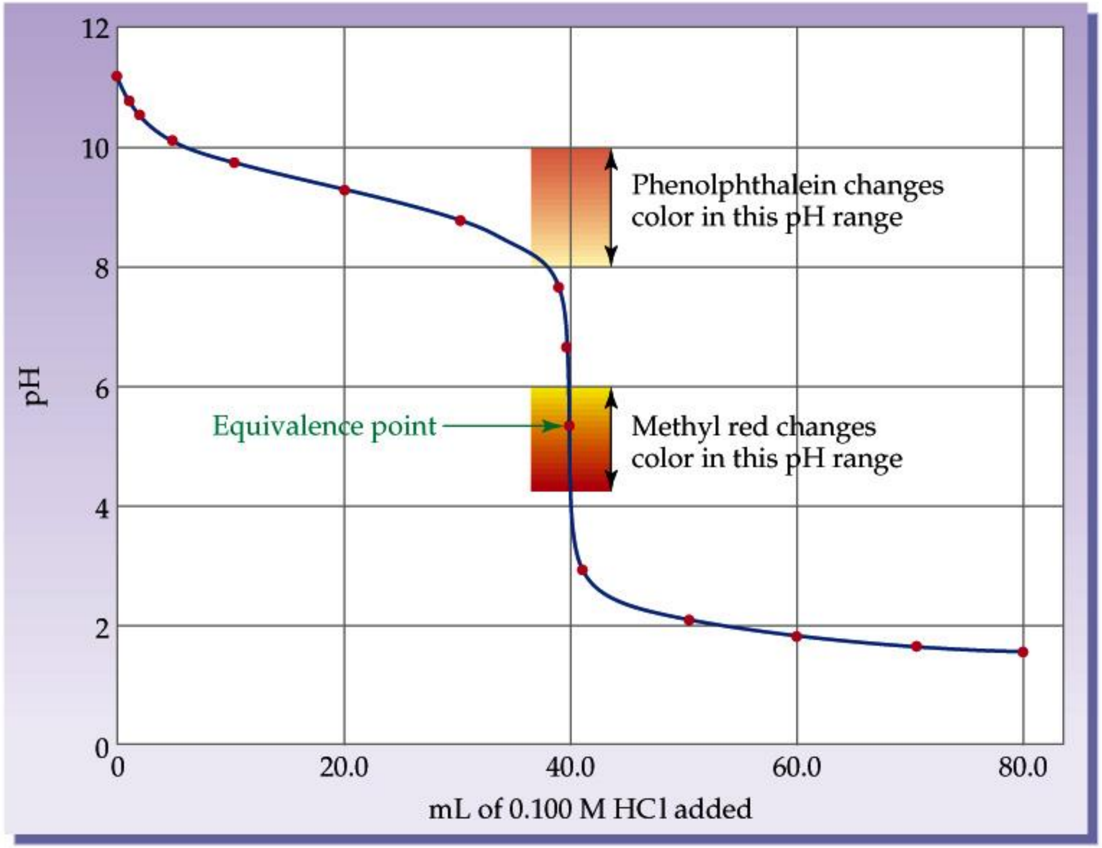
Volume of NaOH added (mL)	pH
00.00	1.00
10.00	1.22
20.00	1.48
30.00	1.85
35.00	2.18
39.00	2.89
39.50	3.20
39.75	3.50
39.90	3.90
39.95	4.20
39.99	4.90
40.00	7.00
40.01	9.40
40.05	9.80
40.10	10.40
40.25	10.50
40.50	10.79
41.00	11.09
45.00	11.76
50.00	12.05
60.00	12.30
70.00	12.43
80.00	12.52



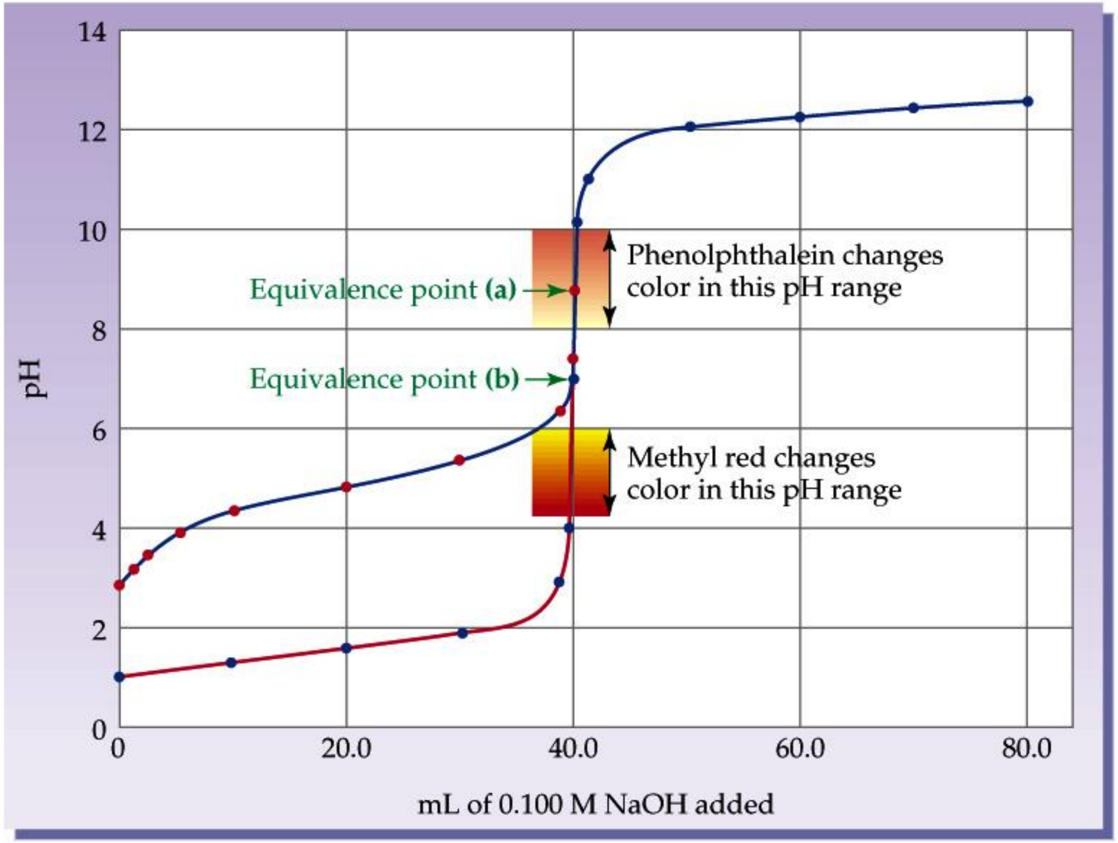


Suitable indicators for strong acid-strong base titration

2. WEAK BASE – STRONG ACID TITRATION



3. WEAK ACID – STRONG BASE TITRATION



Some Common Acid-Base Indicators

	Colour		
Indicator	**In Acid**	**In Base**	**pH Range ***
Methyl orange	Orange	Yellow	3.1-4.4
Bromophenol blue	Yellow	Bluish Purple	3.0-4.6
Methyl red	Red	Yellow	4.2-6.3
Litmus	Red	Blue	5.0 – 8.0
Bromothymol blue	Yellow	Blue	6.0-7.6
Cresol Red	Yellow	Red	7.2-8.8
Phenolphthalein	Colourless	Reddish pink	8.3-10.0
Alizarin yellow	Yellow	Red	10.1-12.0

***The pH range is defined as the range over which the indicator changes from the acid colour to the base colour.**

Summary

Type of Titration	End point pH Range
Strong Acid- Strong Base	3 – 10
Weak Acid - Strong Base	7 – 11
Strong Acid- Weak Base	3 – 7

example

Which indicator(s) would you use for a titration of HNO_2 with KOH ?

HNO_2 : Weak acid

KOH : Strong base

Titration between weak acid / strong base

∴ End point pH range

Table 16.1 Some Common Acid-Base Indicators

	Color		
Indicator	**In Acid**	**In Base**	**pH Range***
Thymol blue	Red	Yellow	1.2–2.8
Bromophenol blue	Yellow	Bluish purple	3.0–4.6
Methyl orange	Orange	Yellow	3.1–4.4
Methyl red	Red	Yellow	4.2–6.3
Chlorophenol blue	Yellow	Red	4.8–6.4
Bromothymol blue	Yellow	Blue	6.0–7.6
Cresol red	Yellow	Red	7.2–8.8
Phenolphthalein	Colorless	Reddish pink	8.3–10.0

PRACTICE EXERCISE

1. In an acid-base titration, 10 mL of 0.45 M HCl was added to 40 mL of 0.10 M NaOH. Calculate the pH of the solution. (ANS: 3.3)

2. What is the pH of a solution consisting of 9.60 mL of 0.1 M NaOH and 10.00 mL of 0.1 M HCl? (ANS: 4.4)

3. Define equivalence point and end point of a titration. Why must the end point the same as equivalence point for a titration ?

Exercise:

What is the colour of the solution when 3 drops of the below indicators are added seperately to water pH = 7?

Indicator	pH range	Colour Change
Phenolphthalein	**8.2 – 10.0**	**Colourless → Reddish pink**
Methyl orange	**3.2 – 4.2**	**Red → Yellow**
Bromothymol blue	**6.0 – 7.6**	**Yellow → Blue**
Phenol Red	**6.8 – 8.4**	**Yellow → Red**