

Titration

◆ Titration

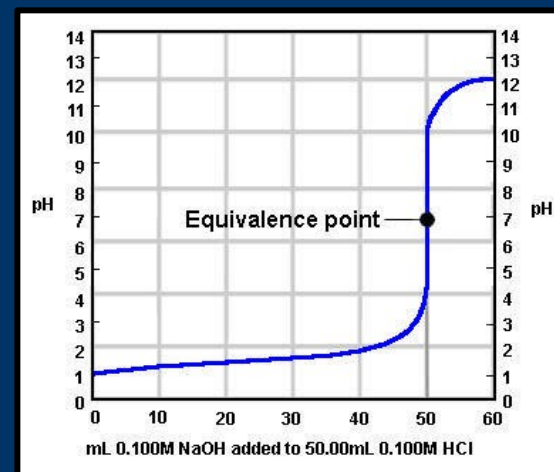
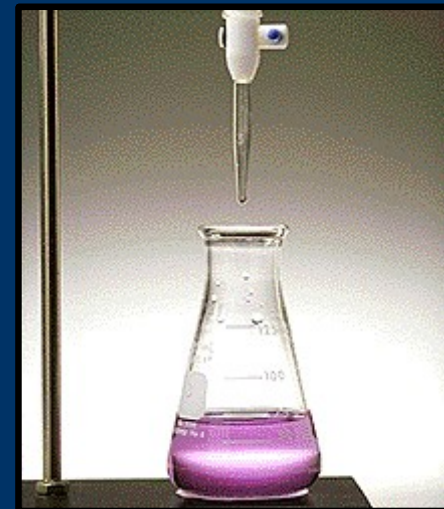
- Analytical method in which a standard solution is used to determine the concentration of an unknown solution.
- A standard solution (titrant or titrator) is **a solution containing a precisely known concentration of compound or a substance.**



Titration

◆ Equivalence point (endpoint)

- Point at which equal amounts of H_3O^+ and OH^- have been added.
- Determined by...
 - indicator color change
 - dramatic change in pH
- ❖ indicator changes color at its endpoint
- ❖ equivalence point is the stoichiometric point where neutralization takes place



Titration



moles H_3O^+ = moles OH^-

$$\mathbf{M \cdot V \cdot n = M \cdot V \cdot n}$$

M: Molarity

V: volume

n: # of H^+ ions in the
acid

or OH^- ions in the

Titration

- ◆ 42.5 mL of 1.3M KOH are required to neutralize 50.0 mL of H_2SO_4 . Find the molarity of H_2SO_4 .

H_3O^+	OH^-	$\text{MV}\# = \text{MV}\#$
$M = ?$ $V = 50.0$ mL $n = 2$	$M = 1.3\text{M}$ $V = 42.5$ mL $n = 1$	$M(50.0\text{mL})(2)$ $= (1.3\text{M})(42.5\text{mL})$ (1) $M = 0.55\text{M}$

METHODS OF EXPRESSING CONCENTRATION

1. Molarity (M) :

The molarity of a solution expresses the number of moles (gram-molecular weights) of solute contained in 1000 ml (1 Litre) of solution. A solution containing 1 mole of solute in each litre of total solution is said to be a one molar (M) solution.

2. Molality (m) :

The molality of a solution expresses as the number of moles of a solute contained in 1000 gm of a solvent. This method of denoting concentration is used in many equations to express thermodynamic properties of solutions.

3. Normality (N) :

The normality of a solution expresses the number of equivalents (gram-equivalent weight) of the solute in one litre of solution. This is generally a much more useful expression since it is directly related to reactive concentrations of various species in solution.

Normality

- Normality = $\frac{\text{Number of Equivalent Weights}}{\text{Liters of Solution}}$

$$N = M \times n$$

$$M = \frac{\text{mol}}{\text{L}}$$



Example: What is the normality of a solution containing 0.40 mol of H_2SO_4 in 600 mL of solution?

$$\#Eq = 0.40 \text{ mol} \times \left(\frac{2 \text{ eq}}{1 \text{ mol}} \right) = 0.80 \text{ eq}$$

$$N = \frac{0.80 \text{ eq}}{.600 \text{ L}} = 1.3 \text{ N}$$



MOLARITY

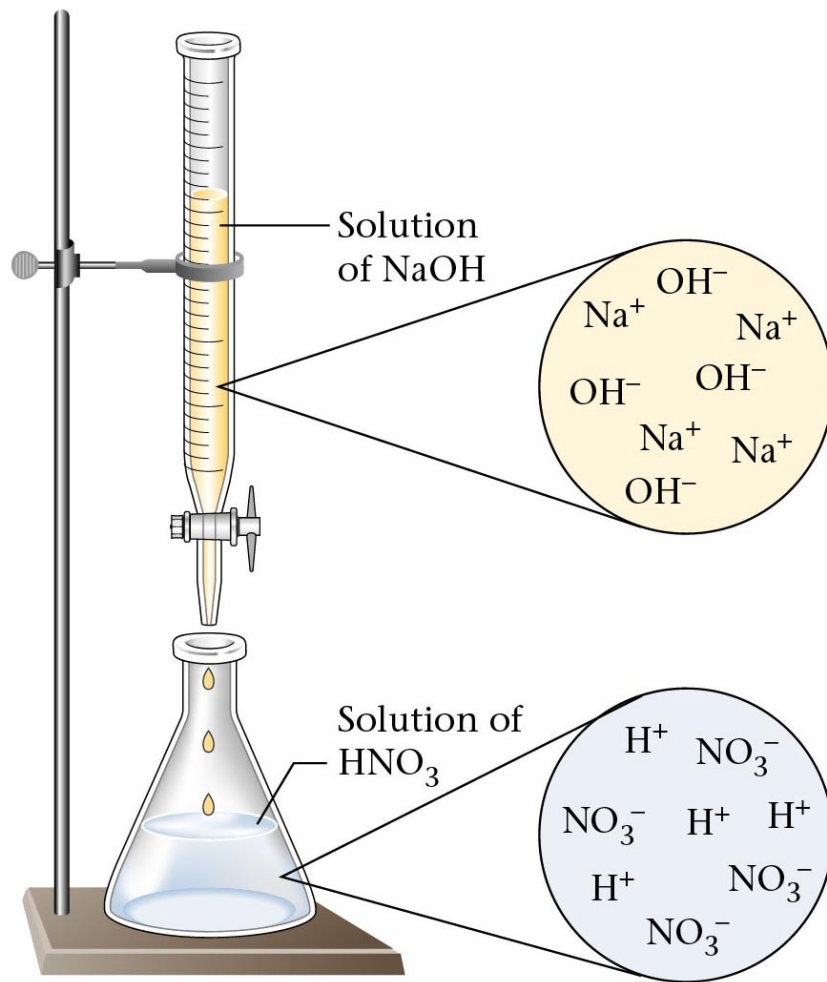
Number of moles of solute in a litre of solution.

$$\text{Molarity} = \frac{\text{moles solute}}{1 \text{ L of solution}}$$

OR

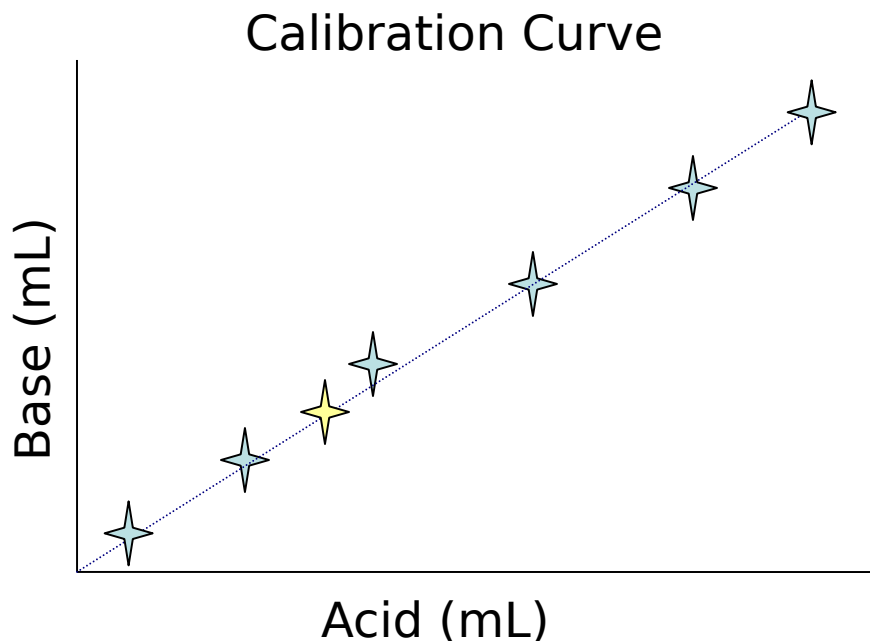
$$M = \frac{\text{mol}}{\text{L}}$$

Acid-Base Titration

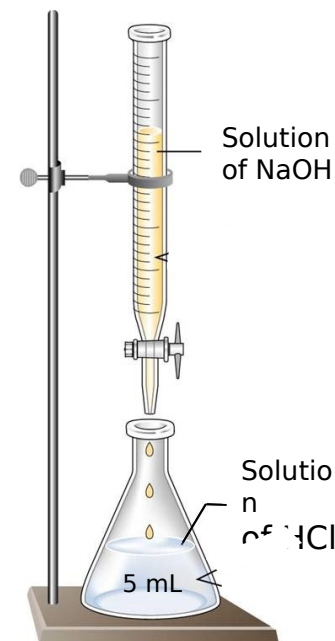


Data Table

<u>0.10 M HCl</u>	<u>? M NaOH</u>
0.00 mL	1.00 mL
1.00 mL	1.00 mL
2.00 mL	2.00 mL
4.00 mL	5.00 mL
9.00 mL	8.00 mL
17.00 mL	10.0 mL
27.00 mL	15.0 mL
48.00 mL	



- 1) Create calibration curve of six data points
- 2) Using [HCl], determine concentration of NH_3
- 3) Determine vinegar concentration using [NaOH] determined earlier in lab



Titration Curve

- A titration curve is a plot of pH vs. the amount of titrant added. Typically the titrant is a strong (completely) dissociated acid or base. Such curves are useful for determining endpoints and dissociation constants of weak acids or bases.

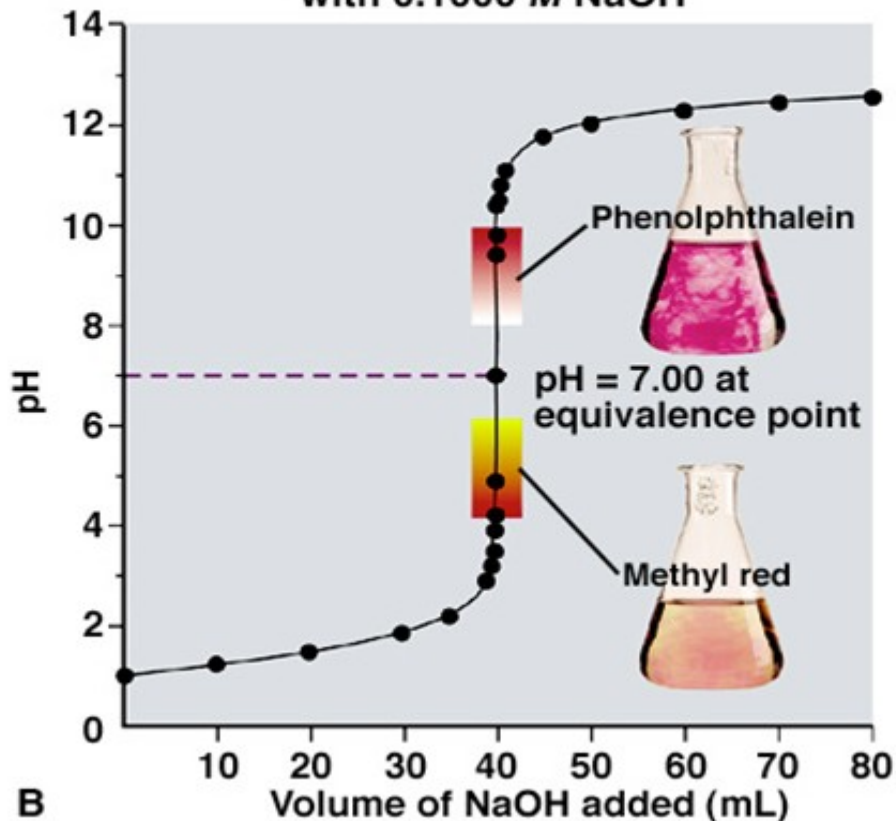


Strong Acid-Base Titration Curve

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

A

Titration of 40.00 mL of 0.1000 M HCl with 0.1000 M NaOH



B

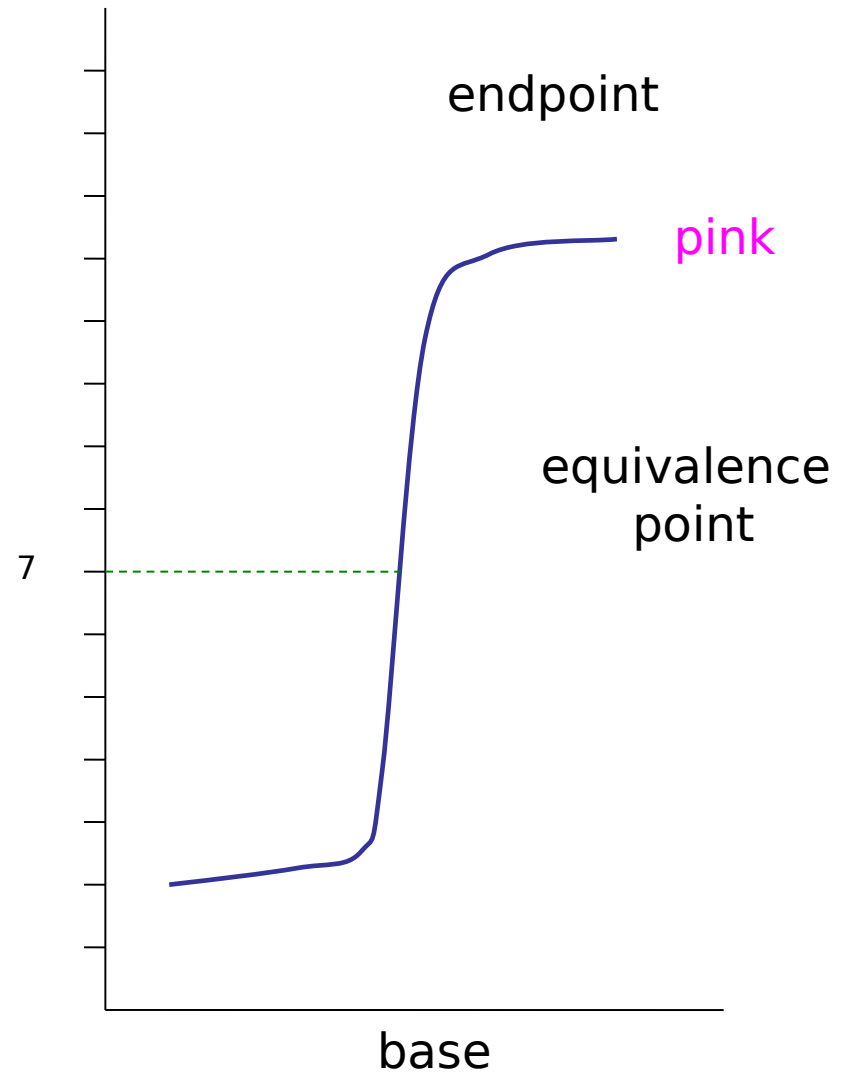
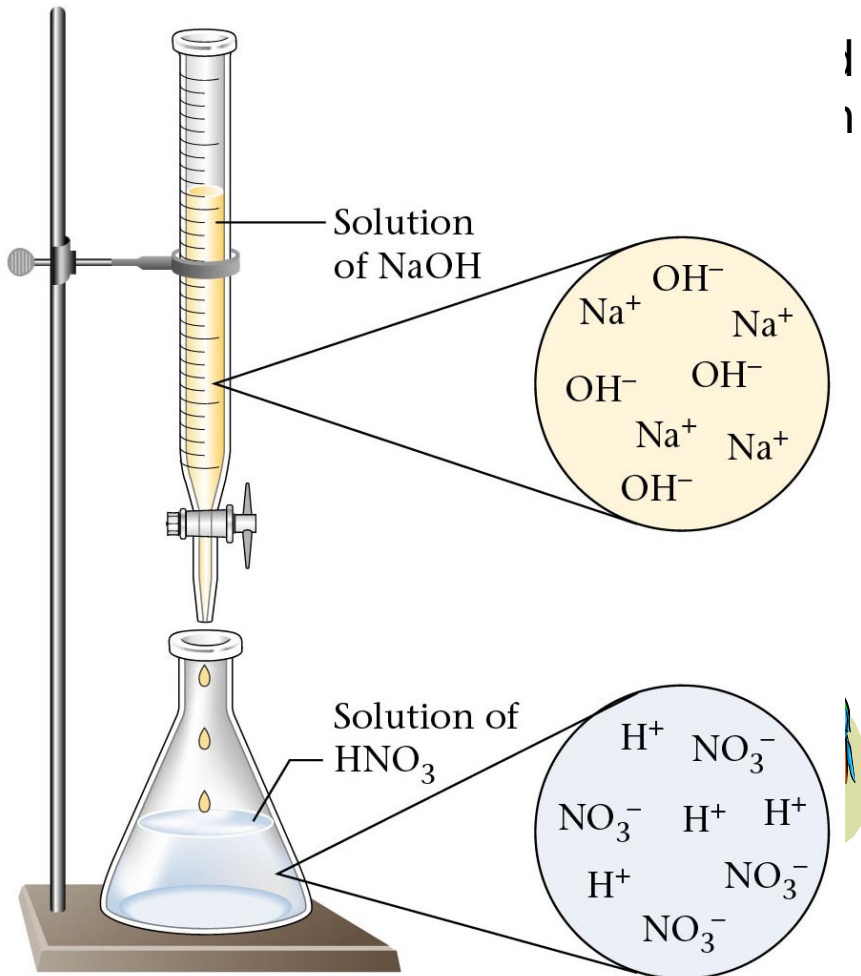
Features of the Strong Acid-Strong Base Titration Curve

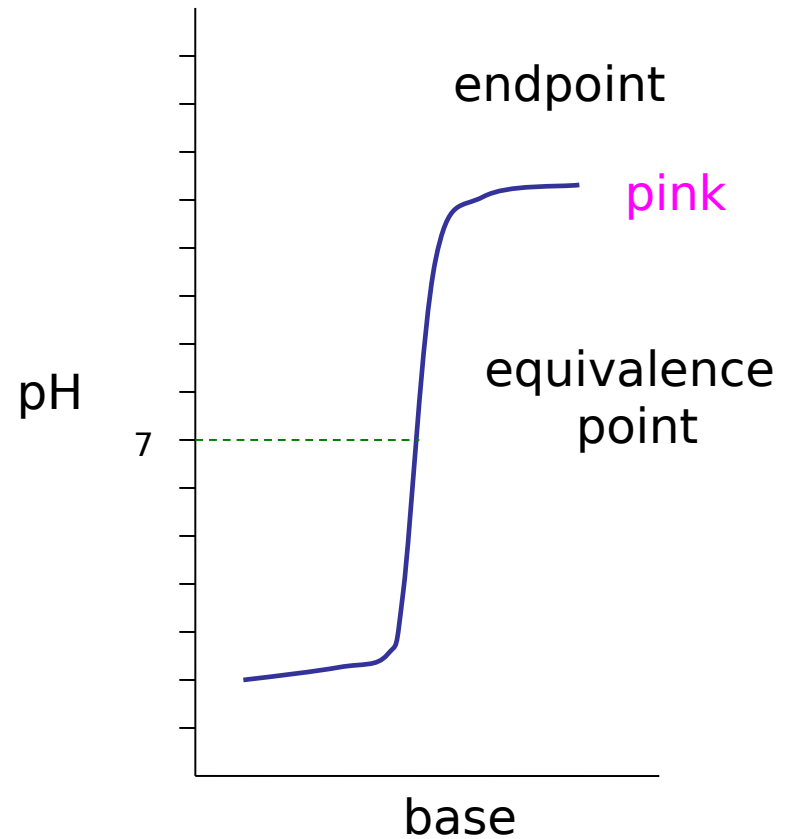
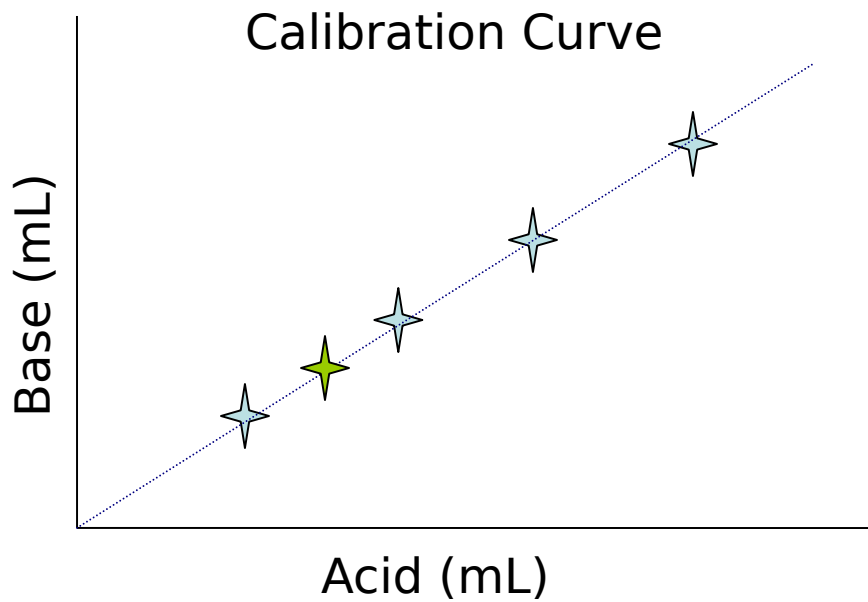
1. The pH starts out low, reflecting the high $[\text{H}_3\text{O}^+]$ of the strong acid and increases gradually as acid is neutralized by the added base.
2. Suddenly the pH rises steeply. This occurs in the immediate vicinity of the equivalence point. For this type of titration the pH is 7.0 at the equivalence point.
3. Beyond this steep portion, the pH increases slowly as more base is added.

Titration

indicator changes color
to indicate pH change

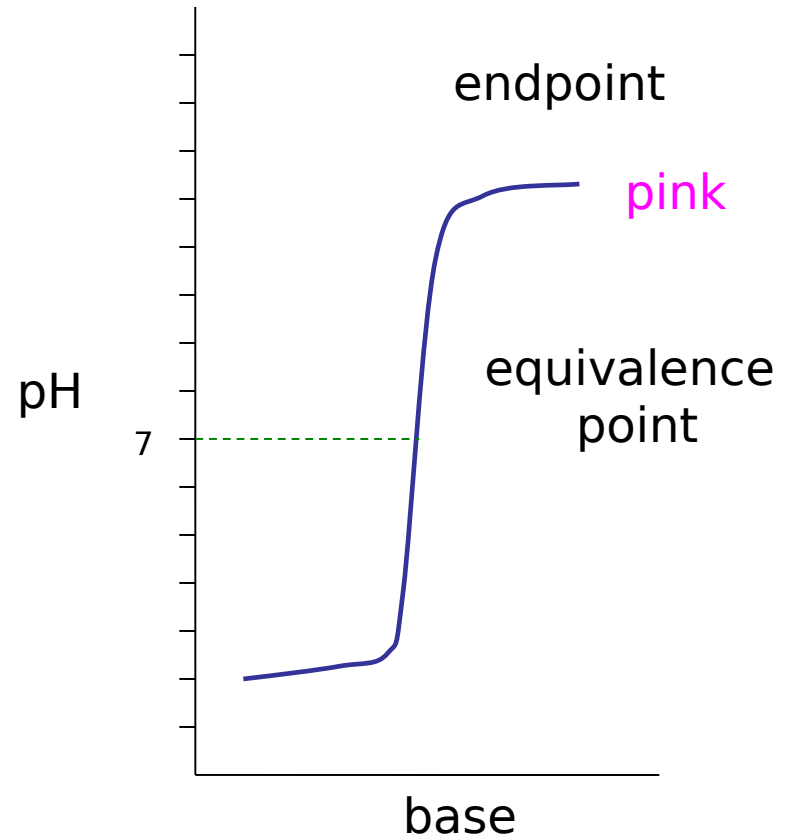
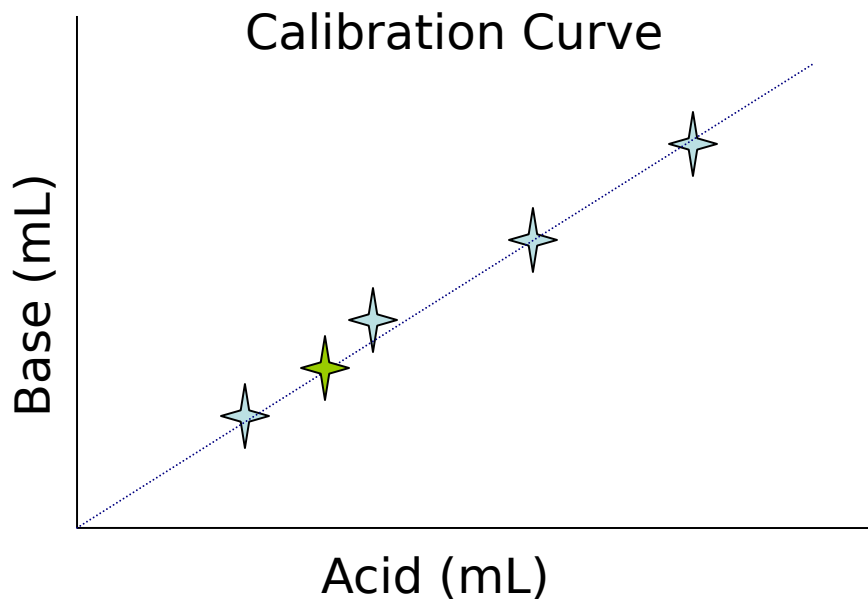
e.g.





indicator changes color to indicate pH change
 e.g. phenolphthalein is colorless in acid
 and pink in basic solution

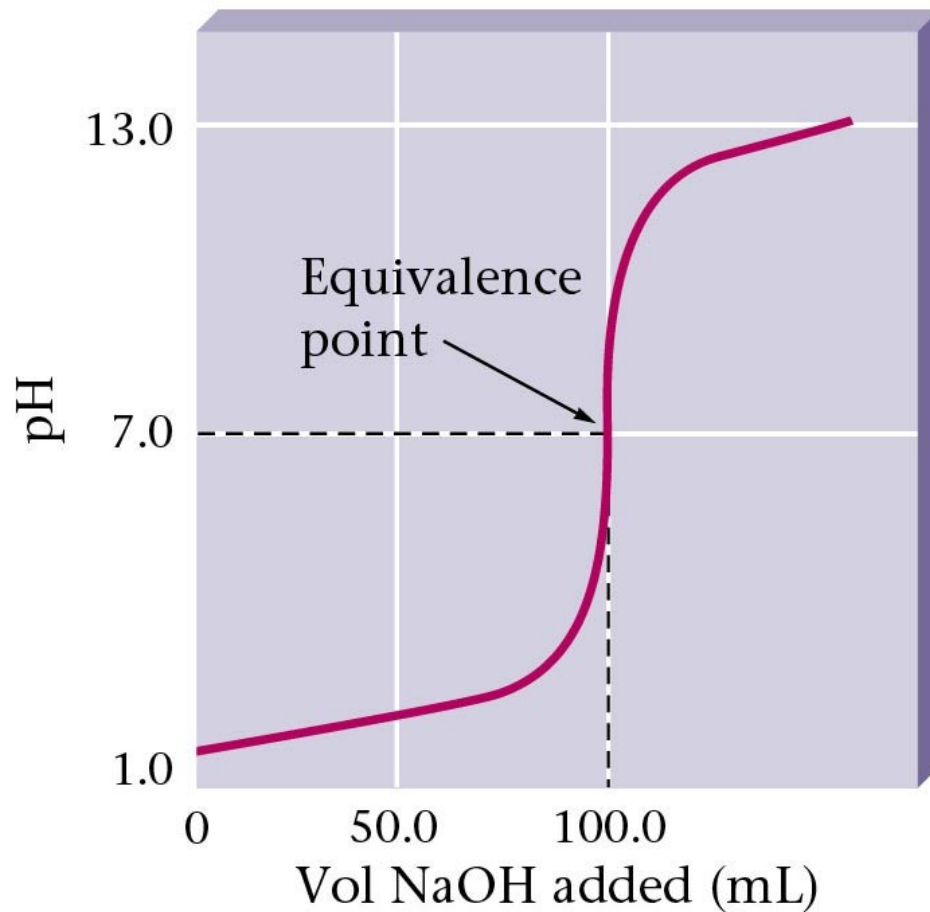
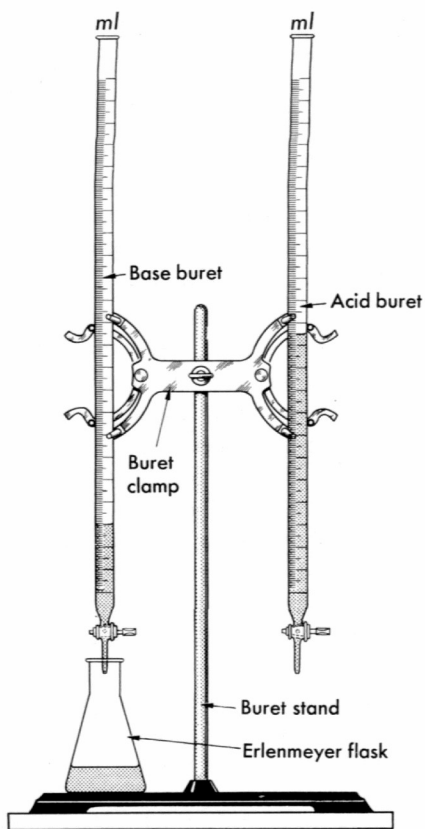




indicator changes color to indicate pH change
e.g. phenolphthalein is colorless in acid
and pink in basic solution



Titration Curve

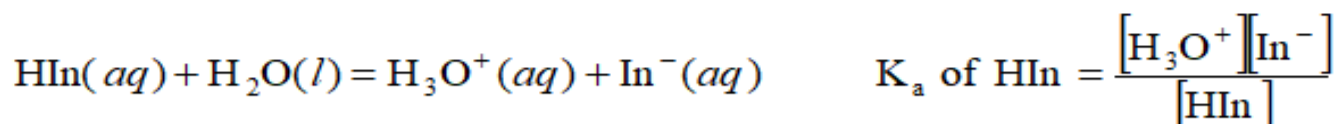


Some common indicators

	pH range for color change									
	0	2	4	6	8	10	12	14		
Methyl violet	Yellow		Violet							
Thymol blue	Red		Yellow		Yellow		Blue			
Methyl orange		Red		Yellow						
Methyl red			Red		Yellow					
Bromthymol blue				Yellow		Blue				
Phenolphthalein					Colorless		Pink			
Alizarin yellow R						Yellow		Red		

Acid-Base Indicators and the Measurement of pH

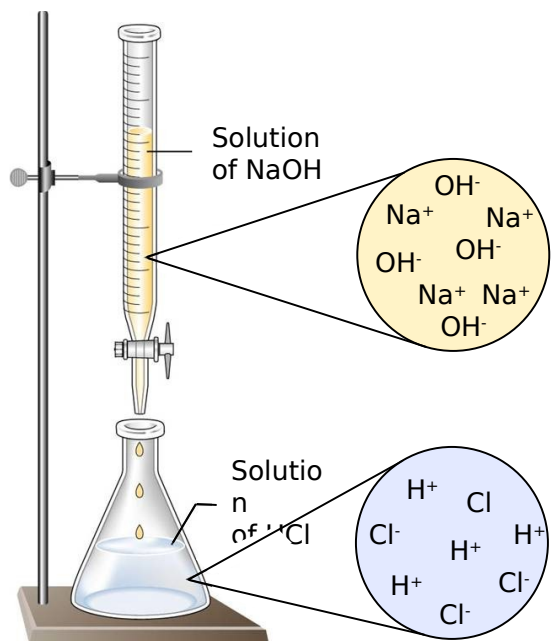
- Definition: A weak organic acid, HIn that has a different color than its conjugate base, In⁻, with the color change occurring over a specific and relatively narrow pH range.
- Typically, one or both forms are intensely colored, so only a tiny amount of indicator is needed, far too little to perturb the pH of the solution.
- Since the indicator molecule is a weak acid, the ratio of the two forms governed by the [H₃O⁺] of the test solution:



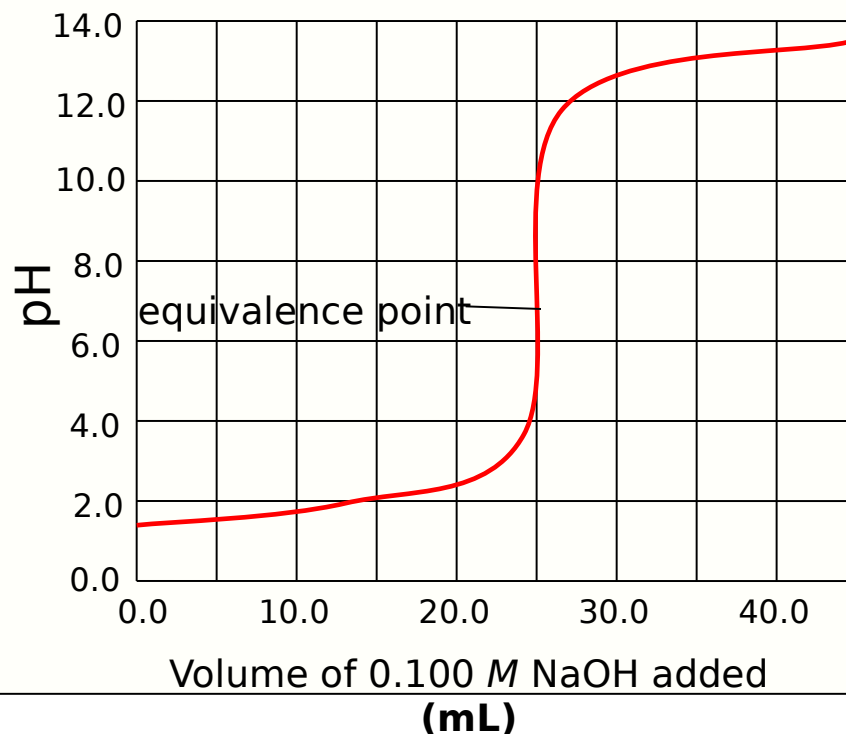
Therefore :

$$\frac{[\text{HIn}]}{[\text{In}^-]} = \frac{[\text{H}_3\text{O}^+]}{[K_a]}$$

Acid-Base Titrations



Titration of a Strong Acid With a Strong Base



1. Adding NaOH from the buret to hydrochloric acid in the flask, a strong acid. In the beginning the pH increases very slowly.

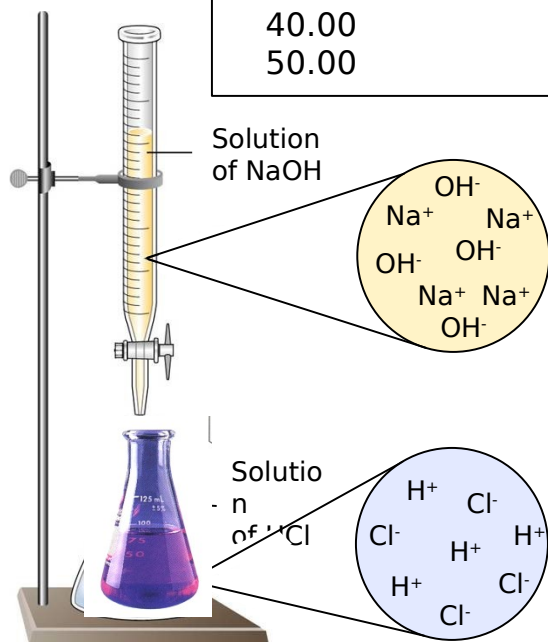
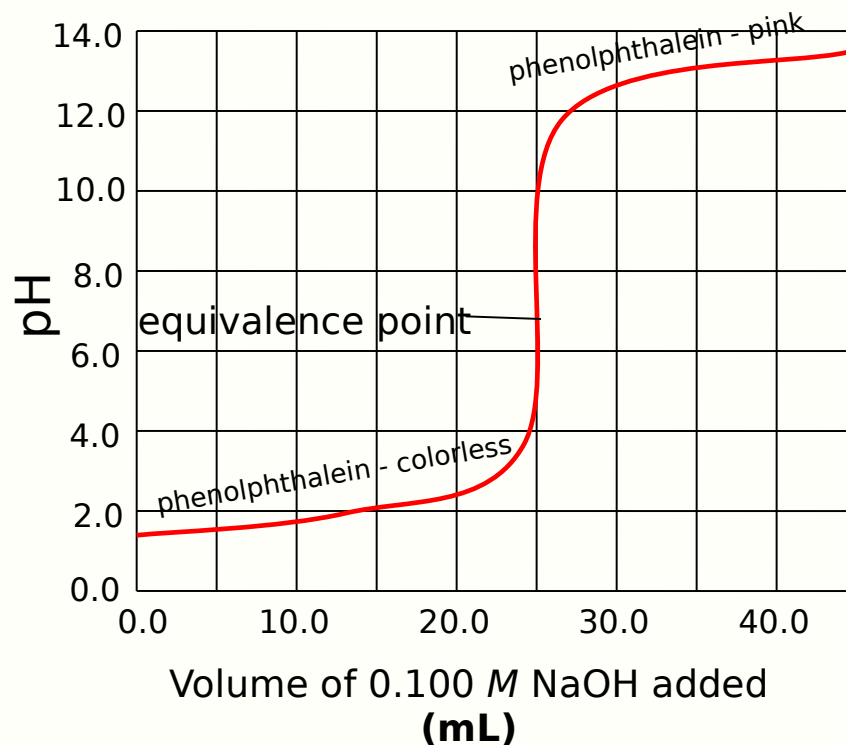
2. Adding additional NaOH is added. pH rises as the equivalence point is approached.

3. Additional NaOH is added. pH increases and then levels off as

Titration Data

NaOH added (mL)	pH
0.00	1.00
10.00	1.37
20.00	1.95
22.00	2.19
24.00	2.70
25.00	7.00
26.00	11.30
28.00	11.75
30.00	11.96
40.00	12.36
50.00	12.52

Titration of a Strong Acid With a Strong Base

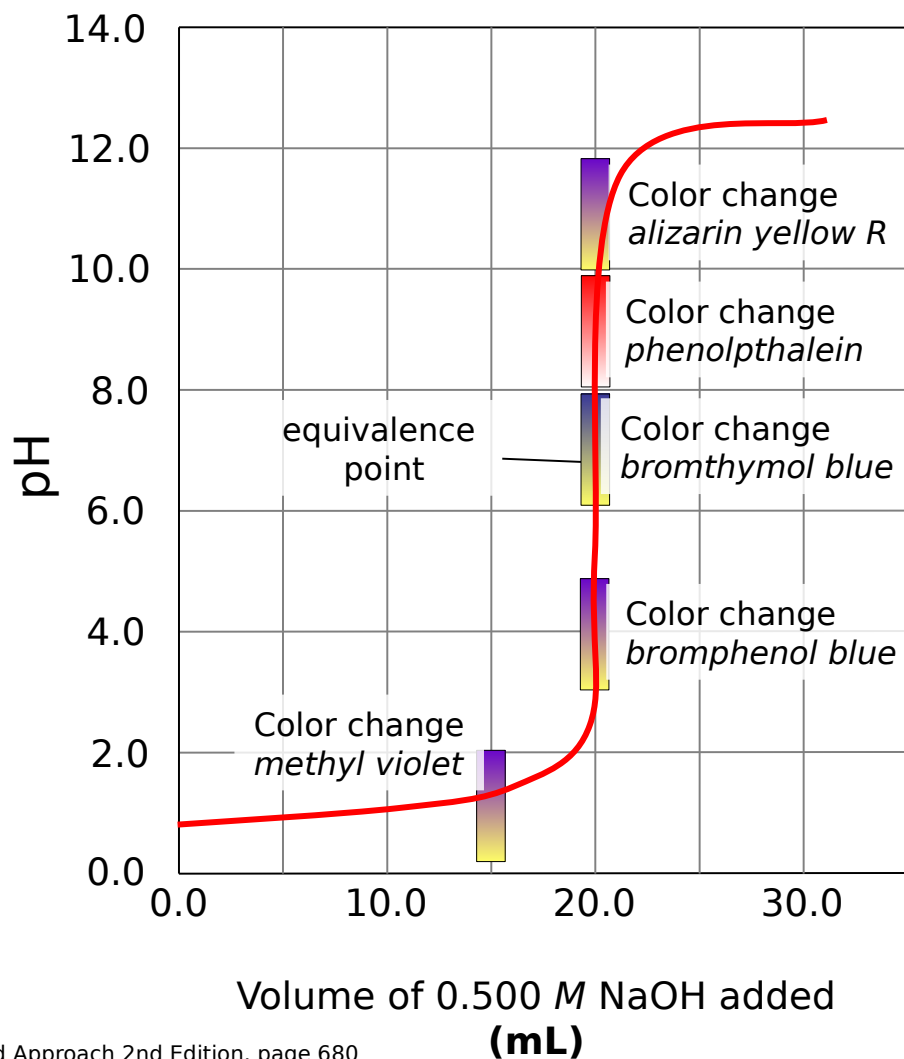


Yellow → Blue

Bromthymol blue is best indicator: pH change 6.0 - 7.0

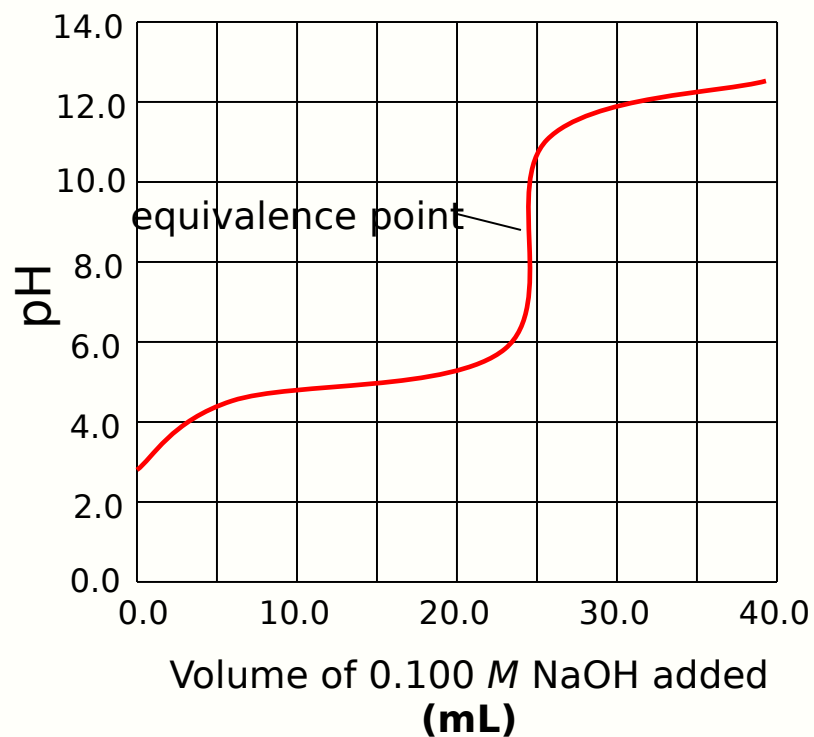
Titration of a Strong Acid With a Strong Base

(20.00 mL of 0.500 M HCl by 0.500 M NaOH)



Titration of a Weak Acid With a Strong Base

Titration of a Weak Acid With a Strong Base



Titration Data

NaOH added (mL)	pH
0.00	2.89
5.00	4.14
10.00	4.57
12.50	4.74
15.00	4.92
20.00	5.35
24.00	6.12
25.00	8.72
26.00	11.30
30.00	11.96
40.00	12.36

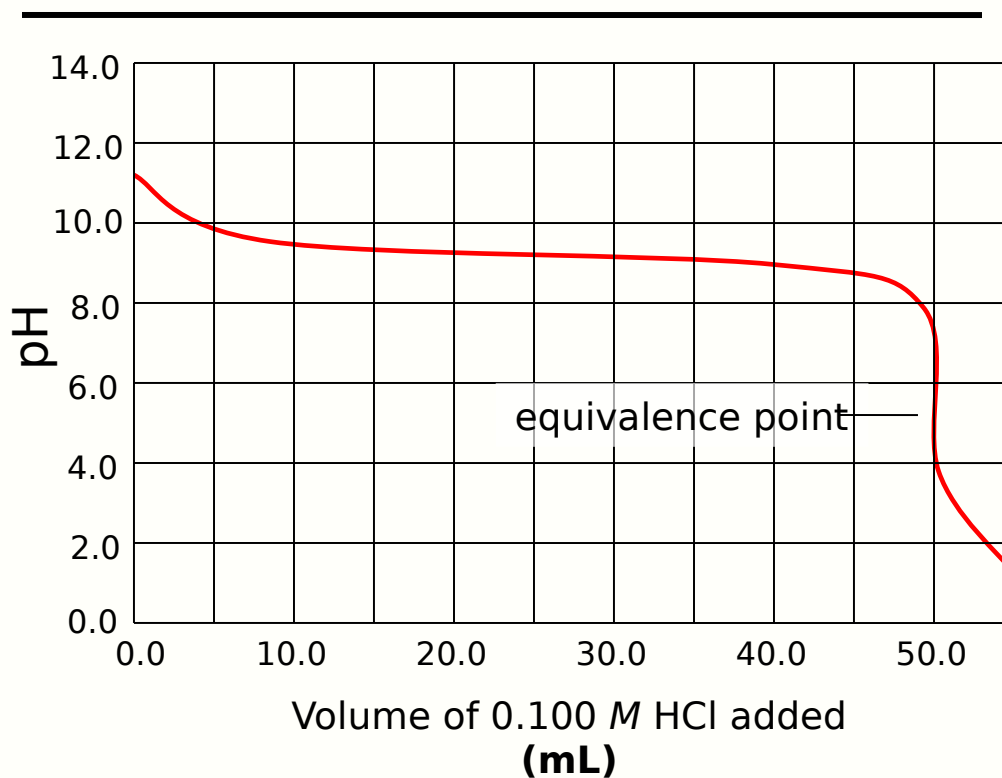
Phenolphthalein is best indicator: pH change 8.0 - 9.6

Titration of a Weak Base With a Strong Acid

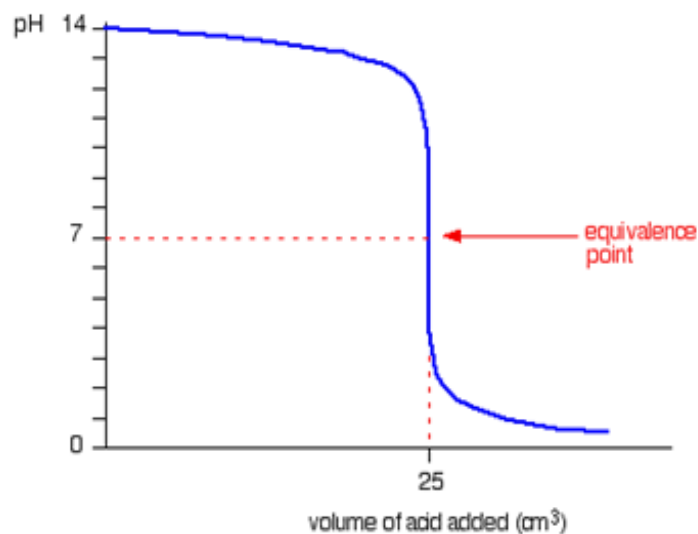
Titration Data

HCl added (mL)	pH
0.00	11.24
10.00	9.91
20.00	9.47
30.00	8.93
40.00	8.61
45.00	8.30
47.00	7.92
48.00	7.70
49.00	7.47
50.00	5.85
51.00	3.34

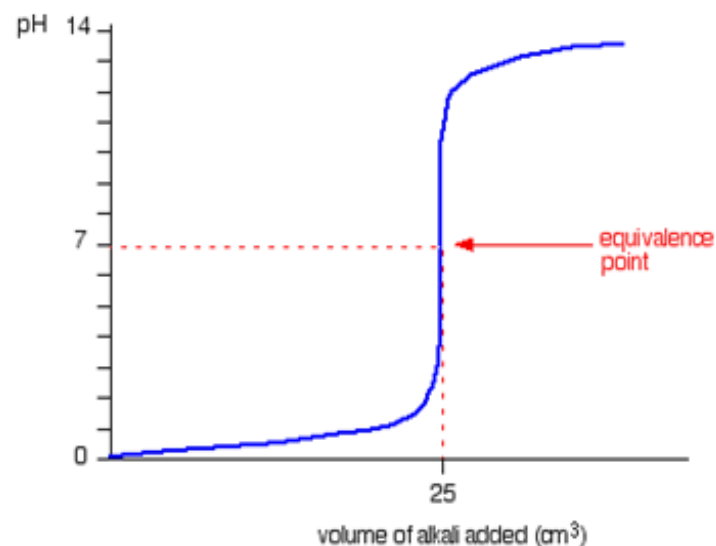
Titration of a Weak Base With a Strong Acid



Titration curves for strong acid vs. strong base



Running acid into the base



Running base into the acid

Sample Calculation: Strong Acid-Strong Base Titration Curve

Problem 24-1. Consider the titration of 40.0 mL of 0.100 M HCl with 0.100 M NaOH.

Region 1. Before the equivalence point, after adding 20.0 mL of 0.100 M NaOH. (Half way to the equivalence point.)

Initial moles of H_3O^+ =

Moles of OH^- added =

$$[\text{H}_3\text{O}^+] = \frac{\text{amount (mol) of } \text{H}_3\text{O}^+ \text{ remaining}}{\text{original volume of acid} + \text{volume of added base}}$$

Region 2. At the equivalence point, after adding 40.0 mL of 0.100 M NaOH.

Initial moles of $\text{H}_3\text{O}^+ = 0.0400 \text{ L} \times 0.100 \text{ M} = 0.00400 \text{ M } \text{H}_3\text{O}^+$

- Moles of OH^- added = $0.0400 \text{ L} \times 0.100 \text{ M} = 0.00400 \text{ mol } \text{OH}^-$

$$[\text{H}_3\text{O}^+] = \frac{\text{amount (mol) of } \text{H}_3\text{O}^+ \text{ remaining}}{\text{original volume of acid} + \text{volume of added base}}$$

Region 3. After the equivalence point, after adding 50.0 mL of 0.100 M NaOH. (Now calculate excess OH^-)

Total moles of $\text{OH}^- = 0.0500 \text{ L} \times 0.100 \text{ M} = 0.00500 \text{ mol OH}^-$

Moles of H_3O^+ consumed = $0.0400 \text{ L} \times 0.100 \text{ M} = 0.00400 \text{ mol}$

$$[\text{OH}^-] = \frac{\text{amount (mol) of OH}^- \text{ remaining}}{\text{original volume of acid} + \text{volume of added base}}$$

Q1. *What is the pH of a solution made by dissolving 2.5 g NaOH in 400 mL water*

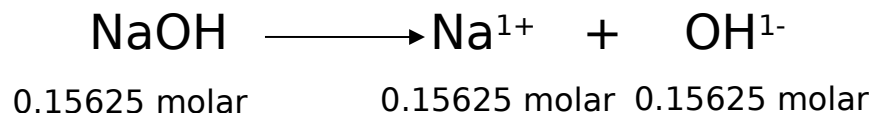
Determine number of moles of NaOH

$$x \text{ mol NaOH} = 2.5 \text{ g NaOH} \left(\frac{1 \text{ mol NaOH}}{40 \text{ g NaOH}} \right) = 0.0625 \text{ mol NaOH}$$

Calculate the molarity of the solution

$$M = \frac{\text{mol}}{\text{L}} \Rightarrow \frac{0.0625 \text{ mol NaOH}}{0.4 \text{ L}} \quad [\text{Recall } 1000 \text{ mL} = 1 \text{ L}]$$

$$M_{\text{NaOH}} = 0.15625 \text{ molar}$$



$$\boxed{\text{pOH} = -\log [\text{OH}^-]}$$

or

$$\boxed{K_w = [\text{H}^+][\text{OH}^-]}$$

$$\text{pOH} = -\log [0.15625 \text{ M}]$$

$$\text{pOH} = 0.8$$

$$\text{pOH} + \text{pH} = 14$$

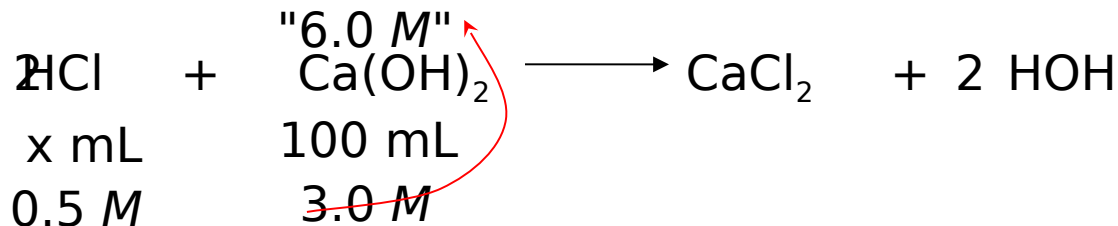
$$0.8 + \text{pH} = 14 \longrightarrow \text{pH} = 13.2 \longleftarrow \text{pH} = -\log [6.4 \times 10^{-14} \text{ M}]$$

$$1 \times 10^{-14} = [\text{H}^+] [0.15625 \text{ M}]$$

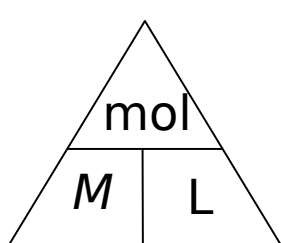
$$[\text{H}^+] = 6.4 \times 10^{-14} \text{ M}$$

$$\text{pH} = -\log [\text{H}^+]$$

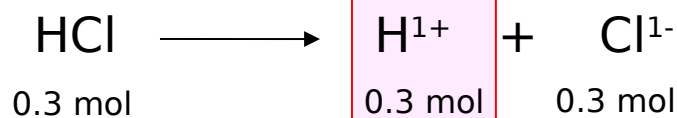
Q2 What volume of 0.5 M HCl is required to titrate 100 mL of 3.0 M Ca(OH)₂



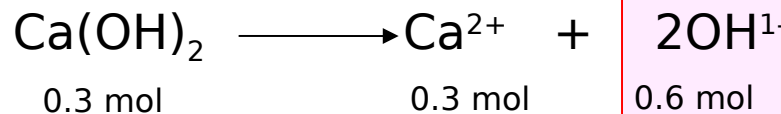
$$\begin{array}{lcl}
 M_1 V_1 = M_2 V_2 & & M_1 V_1 = M_2 V_2 \\
 (0.5 \text{ M}) (x \text{ mL}) = (3.0 \text{ M}) (100 \text{ mL}) & & (0.5 \text{ M}) (x \text{ mL}) = (6.0 \text{ M}) (100 \text{ mL}) \\
 x = 600 \text{ mL of } 0.5 \text{ M HCl} & & x = 1200 \text{ mL of } 0.5 \text{ M HCl}
 \end{array}$$



$$\begin{array}{l}
 \text{HCl} \\
 \text{mol}_{\text{HCl}} = M \times L \\
 \text{mol} = (0.5 \text{ M})(0.6 \text{ L}) \\
 \text{mol} = 0.3 \text{ mol HCl}
 \end{array}$$



$$\begin{array}{l}
 \text{Ca(OH)}_2 \\
 \text{mol}_{\text{Ca(OH)}_2} = M \times L \\
 \text{mol} = (3.0 \text{ M})(0.1 \text{ L}) \\
 \text{mol} = 0.3 \text{ mol Ca(OH)}_2
 \end{array}$$

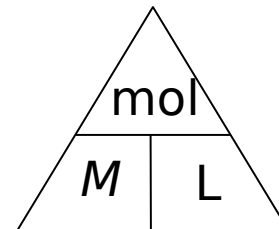


$$[\text{H}^+] = [\text{OH}^-]$$

Q3 10.0 grams vinegar titrated with 65.40 mL of 0.150 M NaOH

(acetic acid + water)

A) $\text{moles HC}_2\text{H}_3\text{O}_2 = \text{moles NaOH}$



NaOH

$$\text{mol}_{\text{NaOH}} = M \times L$$

therefore, you have ...

$$\text{mol} = (0.150 \text{ M})(0.0654 \text{ L})$$

$$0.00981 \text{ mol HC}_2\text{H}_3\text{O}_2 \quad \text{mol} = 0.00981 \text{ mol NaOH}$$

$$\text{B) } x \text{ g HC}_2\text{H}_3\text{O}_2 = 0.00981 \text{ mol HC}_2\text{H}_3\text{O}_2 \left(\frac{60 \text{ g HC}_3\text{H}_2\text{O}_2}{1 \text{ mol HC}_3\text{H}_2\text{O}_2} \right) = 0.59 \text{ g HC}_2\text{H}_3\text{O}_2$$

C) $\% = \left(\frac{\text{part}}{\text{whole}} \right) \times 100\%$

$$\% = \left(\frac{0.59 \text{ g acetic acid}}{10.0 \text{ g vinegar}} \right) \times 100\%$$

$$\% = 5.9 \% \text{ acetic acid}$$

Commercial vinegar is sold as 3 - 5 % acetic acid