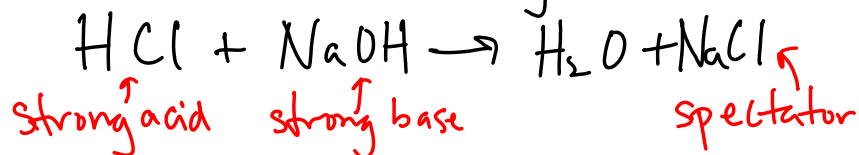


Consider the following neutralization reaction.



If the number of moles of H^+ equals the number of moles of OH^- , what is the pH of the solution? $\text{pH} = 7$. The solution is neutral.

What if there are more moles of H^+ than moles of OH^- ? $\text{pH} < 7$. The solution is acidic.

What if there are more moles of OH^- than moles of H^+ ? $\text{pH} > 7$. The solution is basic.

Eg. What is the pH that results when 25.0 mL of 0.250 M HCl is mixed with 35.0 mL of 0.200 M NaOH?



$$\text{moles HCl} = (0.0250 \text{ L})(0.250 \text{ mol/L}) = 0.00625 \text{ mol HCl}$$

$$\text{moles NaOH} = (0.0350 \text{ L})(0.200 \text{ mol/L}) = 0.00700 \text{ mol NaOH}$$

$$0.00700 - 0.00625 = 0.00075 \text{ moles of excess NaOH}$$

$$[\text{NaOH}] = [\text{OH}^-] = \frac{0.00075 \text{ mol}}{(0.025 \text{ L}) + (0.035 \text{ L})} = 0.013 \text{ M}$$

$$\text{pOH} = -\log[\text{OH}^-] = -\log(0.013) = 1.89$$

$$\text{pH} = 14 - \text{pOH} = 14 - 1.89 = \boxed{12.11}$$

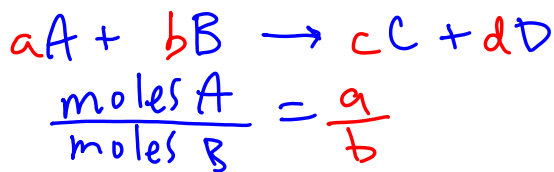
Eg. How many moles of HCl must be added to 40.0 mL of 0.180 M NaOH to produce a solution that has a $\text{pH} = 7$?

When $\text{pH} = 7$ moles of H^+ = moles of OH^-

$$\text{moles OH}^- = (0.180 \text{ M})(0.0400 \text{ L}) = 0.00720 \text{ moles H}^+$$

TITRATION: quantitative chemical analysis used to determine the unknown concentration of a known reactant.

Equivalence Point: When the ratio of moles present in the system exactly equals the mole ratio required by the stoichiometry of the reaction



So, if you know the mole ratio (a/b) and the moles of A in the system, you can calculate the moles of the other reactant, B.

That's the essential theory behind a titration.

In order to perform a titration, you must start with a solution with a precisely known concentration. This is called the standard solution or standardized solution.

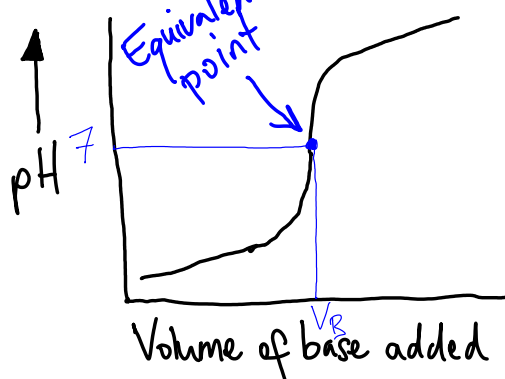
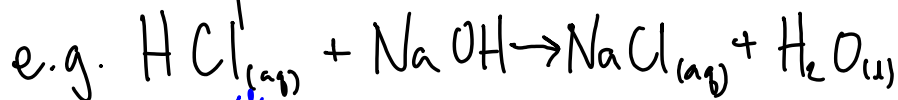
Standardized solutions are prepared by:

- a) Using a substance which can be obtained in a pure and stable form (ie. does not absorb water or carbon dioxide in the air) and which has a known molar mass so that it can be used to prepare a solution of known concentration (**PRIMARY STANDARD**).
e.g. Potassium hydrogen phthalate $\text{KHC}_8\text{H}_4\text{O}_4$ can be used to prepare Oxalic acid dihydrate $\text{H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$

1. Titration of a base with an acid solution

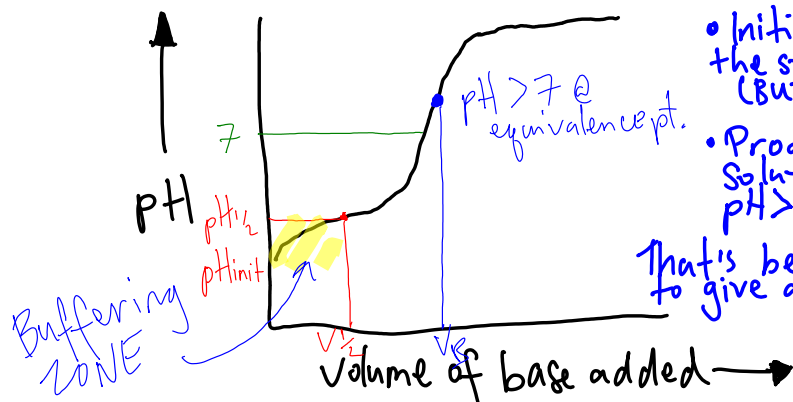
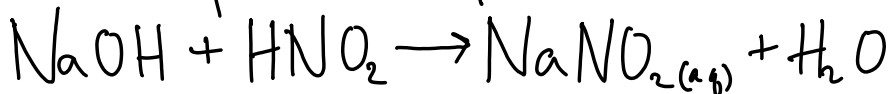
3 TYPES OF TITRATIONS

① Titration of a STRONG ACID with a STRONG BASE



- The pH rises almost vertically around the value of V_B .
- V_B is the volume of the base required to reach the equivalence point.
- SA-SB titrations produce a NEUTRAL salt solution so $\text{pH} = 7$ at equivalence point.
- Choose indicator which changes color around $\text{pH} = 7$, such as phenol red or neutral red.

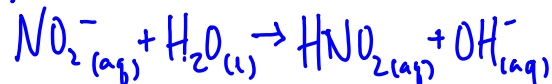
2. Titration of a WEAK ACID with a STRONG BASE



- Initial upswing in the pH at the start of the titration (BUFFERING ZONE).

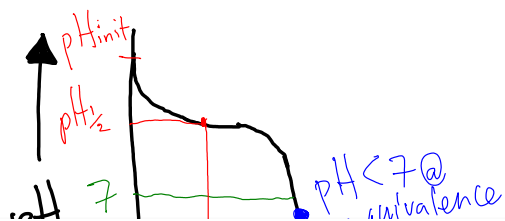
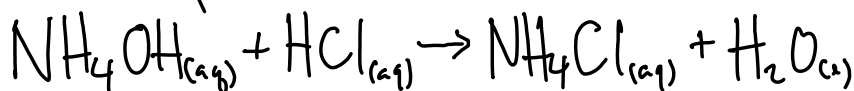
- Produces a BASIC salt solution. As such, the $\text{pH} > 7$ at the equivalence pt.

That's because NaNO_2 hydrolyzes to give a base.

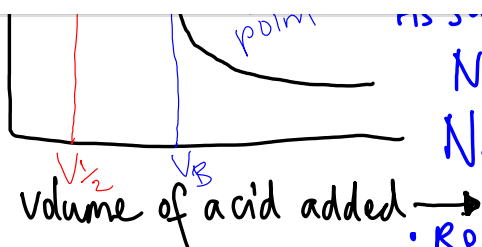


- Requires an indicator that changes colour at $\text{pH} > 7$. Such as phenolphthalein.

3. Titration of a WEAK BASE with a STRONG ACID



- Similar to WA-SB, except that the curve is flipped upside down.
- Graph gives pH values, not pOH. Therefore, pH's must be converted to pOH.
- Produces an ACIDIC solution.



NH_4Cl hydrolyzes to produce acid
 $\text{NH}_4^+(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{NH}_3(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$

• Requires an indicator which changes colour at $\text{pH} < 7$ such as methyl red.

STRONG ACID + STRONG BASE \rightarrow neutral ($\text{pH} = 7$)

weak acid + STRONG BASE \rightarrow basic ($\text{pH} > 7$)

STRONG ACID + weak base \rightarrow acidic ($\text{pH} < 7$)

PRACTICE

- 13.45 mL of 0.200 M NaOH is required to titrate 25.0 mL of a solution which is known to have HCl. Calculate the original [HCl]. (3 marks)
- 13.45 mL of 0.200 M HCl is required to titrate 25.0 mL of a solution which is known to have $\text{Ba}(\text{OH})_2$. Calculate the original [$\text{Ba}(\text{OH})_2$]. (3 marks)
- 13.45 mL of 0.200 M $\text{Sr}(\text{OH})_2$ is required to titrate 25.0 mL of a solution which is known to have HNO_3 . Calculate the original [HNO_3]. (3 marks)
- What volume of 0.100 M NaOH would be required to titrate 35.0 mL of a 0.231 M solution of $\text{H}_2\text{C}_2\text{O}_4$. (3 marks)

Solutions

1. mol NaOH: $0.200 \text{ M} \times 0.01345 \text{ L} = 0.00269 \text{ mol NaOH}$



mol HCl: $0.00269 \text{ mol NaOH} \times \frac{1 \text{ mol HCl}}{1 \text{ mol NaOH}} = 0.00269 \text{ mol HCl}$

$$[\text{HCl}] = \frac{0.00269 \text{ mol}}{0.0250 \text{ L}} = \boxed{0.108 \text{ M}}$$

$$2. \text{ mol HCl: } 0.200 \text{ M} \times 0.01345 \text{ L} = 0.00269 \text{ mol HCl}$$



$$\text{mol Ba}(\text{OH})_2: 0.00269 \text{ mol HCl} \times \frac{1 \text{ mol Ba}(\text{OH})_2}{2 \text{ mol HCl}} = 0.001345 \text{ mol Ba}(\text{OH})_2$$

$$[\text{Ba}(\text{OH})_2] = \frac{0.001345 \text{ mol}}{0.0250 \text{ L}} = \boxed{0.0538 \text{ M}}$$

$$3. \text{ mol Sr}(\text{OH})_2: 0.200 \text{ M} \times 0.01345 \text{ L} = 0.00269 \text{ mol Sr}(\text{OH})_2$$



$$\text{mol HNO}_3: 0.00269 \text{ mol Sr}(\text{OH})_2 \times \frac{2 \text{ mol HNO}_3}{1 \text{ mol Sr}(\text{OH})_2} = 0.00538 \text{ mol HNO}_3$$

$$[\text{HNO}_3] = \frac{0.00538 \text{ mol}}{0.0250 \text{ L}} = \boxed{0.215 \text{ M}}$$

$$4. \text{ mol H}_2\text{C}_2\text{O}_4 = 0.231 \text{ M} \times 0.0350 \text{ L} = 0.008085 \text{ mol H}_2\text{C}_2\text{O}_4$$



$$\text{mol NaOH} = 0.008085 \text{ mol H}_2\text{C}_2\text{O}_4 \times \frac{2 \text{ mol NaOH}}{1 \text{ mol H}_2\text{C}_2\text{O}_4} = 0.01617 \text{ mol NaOH}$$

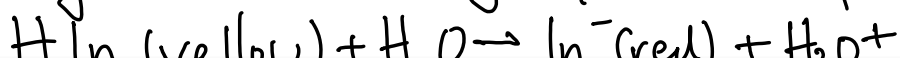
$$\text{Volume NaOH} = \frac{0.01617 \text{ mol}}{0.100 \text{ M}} = 0.162 \text{ L} \text{ or } \boxed{162 \text{ mL}}$$

Indicators

Indicators are used to determine the pH of a solution.

Indicator = weak organic acid or base that is different colours in its conjugate acid and base forms.

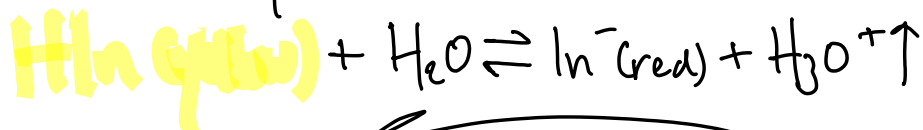
Imagine the following equilibrium of an indicator.



↑ Indicator's
acid form

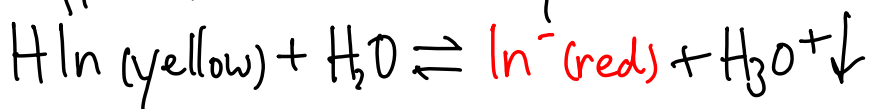
↑ Indicator's
base form

When an indicator is placed into an acid solution, the H_3O^+ causes a shift in the indicator's equilibrium.



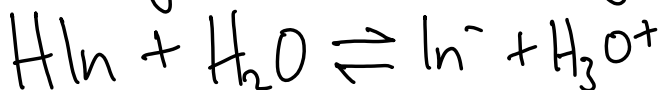
An indicator will be in its Conjugate acid (HIn) form in HIGHLY ACIDIC solutions.

When an indicator is placed into a basic solution, the OH^- reacts to decrease the $[H_3O^+]$ and causes a shift in the indicator's equilibrium.



An indicator will be in its conjugate base (In^-) form in HIGHLY BASIC solutions.

Transition Point: When the indicator is mid-way through its colour change and $[HIn] = [In^-]$



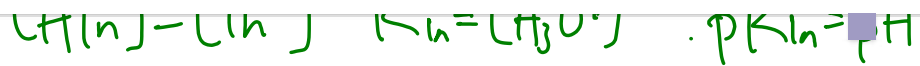
So, at the transition point,

$$K_{In} = \frac{[H_3O^+][In^-]}{[HIn]} = [H_3O^+]$$

$$\therefore -\log K_{In} = -\log [H_3O^+]$$

$$pK_{In} = pH \text{ @ transition}$$

At the transition point of any indicator, the following relationships exist:



The K_a of an indicator, K_{in} , can be calculated using the pH range over which the indicator changes colour.