Chemistry 12

Acid-Base Equilibrium VII

Name:
Date:
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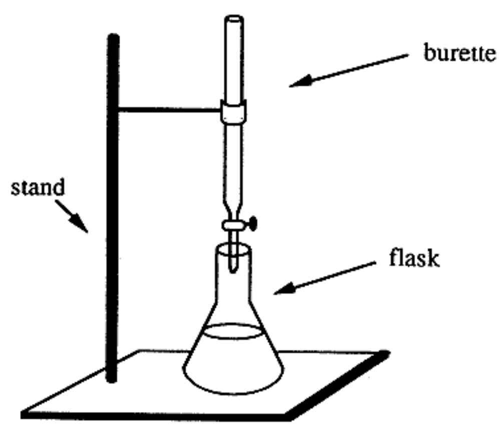
1. **Titrations**
2. **Indicators**

Titrations

Titration is a form of volumetric analysis where the number of moles of solute in a solution is determined by adding a sufficient volume of another solution of known concentration to just produce a complete reaction.

- The reaction is completed when the number of moles of H_3O^+ equals the total number of moles of OH^- - this is called the **equivalence point** or **stoichiometric point** of the titration.
- An **indicator** is used that will indicate when the equivalence point has been reached by changing colour at or very near the pH associated with the equivalence point.



We will be dealing with 3 scenarios:

1.

2.

3.

<u>**(Chemistry 11) Example 1:**</u> A student standardizing a solution of NaOH finds that 28.15 mL of that solution is required to neutralize 25.00 mL of a 0.1072 M standard solution of HCl. Calculate the [NaOH].

- Write the balanced equation.

- What is the concentration of NaOH?

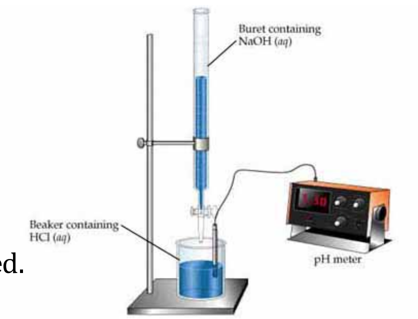
Example 2: A 20.0 mL sample of 0.450 M HNO_2 is titrated with a 0.500 M NaOH solution. What volume of base is added at exactly halfway to the equivalence point?

- Draw a sketch of the set-up:

- What is the balanced reaction?

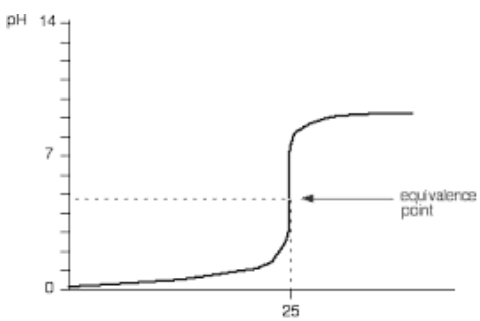
- What volume of NaOH was needed at equivalence point?

- What is halfway to equivalence point?



TITRATION CURVES:

- A graph that plots the pH of the solution vs. the volume of the titrant added.



As the titrant is added, the pH is gradually *(increasing/decreasing)* which means that *(acid/base)* is the titrant.

How many mL of titrant was required to attain equivalence point? __________

Looking at the equivalence point pH, do you think the reaction produced an acidic or basic salt?

What do you think the graph would look like if an acid was the titrant? Explain your answer.

1. Strong Acid & Strong Base

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

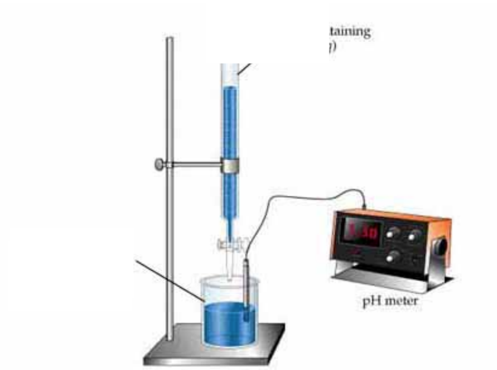
- Typical strong acid + strong base neutralization type of question.
  (1) Dilution
  (2) Neutralization
  (3) Acidic or basic?

2. Weak Acid & Strong Base

$$CH_3COOH_{(aq)} + NaOH_{(aq)} \rightarrow \mathbf{NaCH_3COO_{(aq)}} + H_2O_{(l)}$$

50.0 mL of 0.100 M acetic acid is titrated with 0.150 M NaOH.

What do you think the pH curve will look like?



Because the pH meter is measuring **from the beaker**, we always have to think about ________________!

We will be calculating (and comparing) the pH at 4 points on the curve:

1.
2.
3.
4.

Calculate the pH of the solution produced in the reaction flask at the following points:

1) The pH of the solution of acetic acid when no NaOH is yet added.

- This is a weak acid calculation.

2) When 10.0 mL of 0.150 M NaOH has been added.

- Acts very similarly to a buffer as the added hydroxide ions reacts with acetic acid to produce acetate ions.
- What are the diluted concentrations of reactant acid and base before the reaction (initial concentrations)?

- Create an ICE table.

- CH_3COOH and CH_3COO^- creates an acid buffer.

- Calculate the pH.

3) At the equivalence point when _____ mL of NaOH has been added.

- What volume of NaOH is needed at equivalence point?

- What are the diluted concentrations of reactant acid and base before the reaction (initial concentrations)?

- Create an ICE table.

- The anion of the dissociated salt, $NaCH_3COO$, is the conjugate base of a weak acid and is thus capable of accepting protons from water in a hydrolysis reaction.
- What is the hydrolysis reaction?

- Create an ICE table.

- Calculate pH of the solution resulting from the anionic hydrolysis of the acetate ion.

4) Beyond the equivalence point when 60.0 mL of NaOH added

Practice:

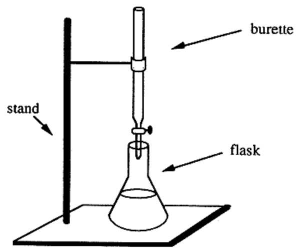
A 20.0 mL sample of 0.450 M HNO_2 is titrated with a 0.500 M NaOH solution. What will the pH be in the reaction flask at the following points:

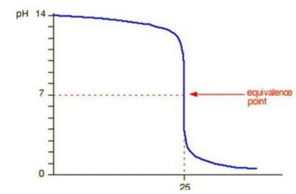
a) 2.0 mL before exactly halfway to the equivalence point?

b) At equivalence point?

3. Strong Acid & Weak Base

$NH_{3\,(aq)} + HCl_{(aq)} \rightarrow \mathbf{NH_4^+}_{\mathbf{(aq)}} + Cl^-_{(aq)}$





100.0 mL of 0.050 M NH_3 is titrated with 0.10 M HCl.

Calculate the pH of the solution produced in the reaction flask at the following points:

1) Before any HCl is added.

2) At the midpoint of the titration.

3) At the equivalence point when _______ mL of HCl has been added.

4) When 60.0 mL of HCl has been added.

Practice:
Calculate the pH of the solution produced in the reaction flask when 13.00 mL of 0.100 M $HClO_4$ has been added to 25.00 mL of 0.100 M $NaNO_2$. (This is just beyond halfway to the equivalence point.)

Indicators

We measure pH using either an acid-base indicator or a pH meter. Acid-base indicators are weak organic acids whose conjugate pairs display different and normally intense colours.

Acid-base indicators are complex organic molecules and refer to them as simply “HIn.”

$$\mathbf{HIn + H_2O \leftrightharpoons In^- + H_3O^+}$$

K_a =

<u>Example:</u> Consider methyl red indicator.

pH:

Colour:

- The pH value at which the indicator exhibits a colour change should be close to the pH at equivalence point. (__________ → __________ ← __________).
- When the colour changes, (reached ____________________ point) it is an indication that the titration has reached equivalence point.

pH at equivalence point	$[H_3O^+]$	Indicator	Colour exhibited
		Alizarin yellow	
6.8			
			Light pink
		Thymol blue	Green
2.0			

When 1.0 M NH_3 is titrated with 1.0 M HCl, the most suitable indicator is:

A. methyl violet.
B. indigo carmine.
C. phenolphthalein.
D. bromcresol green.

The most appropriate indicator for the titration of 0.50 M CH_3COOH with 0.50 M NaOH is

A. methyl violet.
B. indigo carmine.
C. phenolphthalein.
D. bromcresol green.

<u>Hebden Workbook Pg. 162 #108-116</u>