Acid-Base Titrations

Section 8.7

Homework
(Strong Acid-Strong Base)
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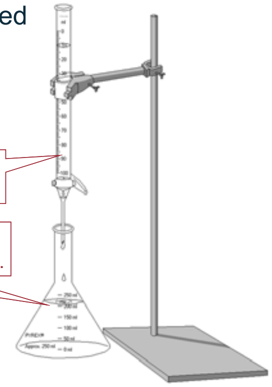
Worksheet

Titration is a quantitative technique used to find the concentration of a solution.

- involves the progressive addition of precise volumes of the titrant to the sample

The **titrant** is the solution in the burette; it has a KNOWN concentration.

The **sample** is the solution held in the flask; it has an UNKNOWN concentration.



The titrant is added to the sample until the point when all reactant in the sample is consumed. This is the **EQUIVALENCE POINT**.

The equivalence point is determined using stoichiometry (molar ratios).

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

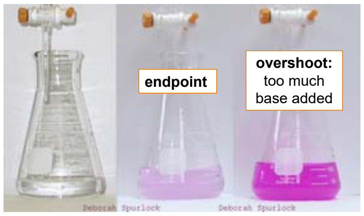
n = 0.150 mol n = 0.150 mol

$$2\,HCl_{(aq)} + Ca(OH)_{2\,(aq)} \rightarrow 2\,H_2O_{(l)} + CaCl_{2\,(aq)}$$

n = 0.150 mol n = 0.0750 mol

The most common type of titration is an acid-base titration.

- acid and base react in a neutralization reaction
- progress of reaction is monitored using a pH indicator
  - **ENDPOINT** – The point of the reaction when the indicator changes colour



For maximum experimental accuracy,

ENDPOINT = EQUIVALENCE POINT

Experimental	Theoretical
The visual point when the indicator changes colour	The stoichiometric point when $n_{acid} = n_{base}$

These are NOT inherently the same thing!
You need to be careful about picking an indicator that changes colour at the equivalence point.

Before a titration is performed, the titrant must be **standardized**.

- Standardization involves accurately determining the concentration of the titrant.
- It is accomplished by titrating the titrant with a stable, non-volatile acid or base, of whose concentration we can be certain.

Standardize the titrant.
(Find its concentration)

↓

With the titrant concentration now accurately known, use the titrant to **titrate the unknown sample**.

↓

Once the endpoint is reached, record the **volume of titrant.**

↓

Use stoichiometry to determine the **concentration of the unknown**.

Learning Checkpoint

- Titration involves reacting a precise volume of titrant with a sample of unknown concentration.
  - Use stoichiometry to find the unknown concentration.
- The endpoint of the reaction is a visual indicator that the reaction has reached its equivalence point.
  - They are not the same thing, though a properly selected pH indicator can ensure that the endpoint is achieved at the equivalence point.

Strong Acid-Strong Base Titrations

Qualitative Analysis

strong acid + strong base → salt + water

$$HNO_{3\,(aq)} + NaOH_{(aq)} \rightarrow NaNO_{3\,(aq)} + H_2O_{(l)}$$

$$HNO_{3\,(aq)} + NaOH_{(aq)} \rightarrow Na^+_{(aq)} + NO_3^-{}_{(aq)} + H_2O_{(l)}$$

Na^+ and NO_3^- do not have acidic/basic properties
∴ solution is neutral at the equivalence point.

Quantitative Analysis

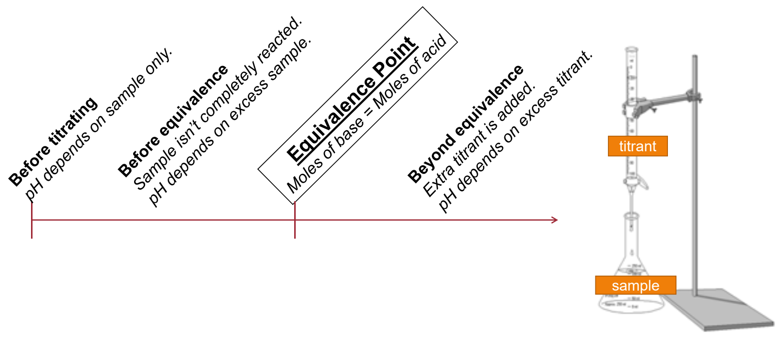
All titrations need to be analyzed in two steps:

1. As **stoichiometry** problems:
   - How many **MOLES** of acid/base are in the solution?
   - Which one is in excess, and how will that affect pH?
2. As **equilibrium** problems
   - In the case of **weak** acids or bases, what **CONCENTRATION** of acid/base will dissociate? This determines pH.

Since strong acids and bases completely dissociate, you don't really have to consider these systems as equilibrium problems.
You will see the equilibrium component when we do weak acid or weak base problems.

You will use mainly stoichiometry to analyze strong acid-strong base systems.

All titrations can be analyzed at multiple points:

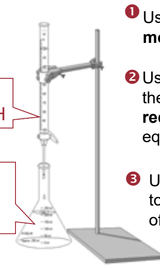


Example 1. $HNO_{3\,(aq)} + NaOH_{(aq)} \rightarrow NaNO_{3\,(aq)} + H_2O_{(l)}$

a) Use stoichiometry to calculate the volume of NaOH that will be required to react completely with the sample of HNO_3.

Titrant: 0.100 mol/L NaOH

Sample: 50.0 mL of 0.200 mol/L HNO_3



❶ Use n = c × v to find **moles of HNO_3.**

n = 0.200 mol/L x 50.0 mL
n = 10.0 mmol

❷ Use mole ratio to find the **amount of NaOH required** to reach the equivalence point.

$n_{NaOH} = 10.0 \text{ mmol } HNO_3 \times \frac{1 \text{ mol NaOH}}{1 \text{ mol } HNO_3}$
$n_{NaOH} = 10.0$ mmol NaOH

❸ Use the concentration to convert the amount of NaOH to a **volume**.

$$V_{NaOH} = \frac{10.0 \text{ mmol}}{0.100 \text{ mol/L}} = 100 \text{ mL} \quad (3\ SD)$$

The reaction will have reached its equivalence point once **100 mL** of base has been added.

b) Before the titrant is added, what is the pH of the solution?

❶ **Stoichiometry**: How many moles of acid are present?

No base has been added, therefore no acid has reacted yet. The solution contains the original amount of HNO_3.

❷ Consider all species in the solution. Which one controls the pH?

$H^+_{(aq)}$, $NO_3^-{}_{(aq)}$, $H_2O_{(l)}$

HNO_3 is a strong acid ∴ it controls the pH

❸ **Equilibrium**: What is the concentration of H^+?

HNO_3 is a strong acid. Therefore, 0.200 mol/L HNO_3 contains hydrogen ions at a concentration of **0.200 mol/L**.

❹ Calculate pH

pH = -log [H^+]
= - log (0.200)
pH = 0.699

Before any base is added, the HNO_3 solution has a pH of **0.699**.

c) Before the equivalence point: after 10.0 mL of the titrant is added, what is the pH of the solution?

❶ **Stoichiometry**: How many moles of acid are present?

Soon after the base is added to the sample, the OH^- will react with the H^+:

$$H^+ + OH^- \rightarrow H_2O$$

	H^+	OH^-	H_2O
Moles before reaction	50.0 mL x 0.200 mol/L = **10.0 mmol**	10.0 mL x 0.100 mol/L = **1.00 mmol**	
Moles after reaction	10.0 mmol – 1.00 mmol = **9.0 mmol**	1.00 mmol – 1.00 mmol = **0 mmol**	

There is more H^+ than OH^-. The OH^- will get used up, leaving excess H^+.

❷ Consider all species in the solution. Which one controls the pH?

$H^+_{(aq)}$, $NO_3^-{}_{(aq)}$, $Na^+_{(aq)}$, ~~$OH^-_{(aq)}$~~, $H_2O_{(l)}$

The excess leftover ion (H^+) will control the pH.

❸ **Equilibrium**: What is the concentration of H^+?

There is **9.0 mmol** of H^+ remaining.
Total volume of solution is
50.0 mL (acid) + 10.0 mL (base) = **60.0 mL**

$$\text{Therefore, } [H^+] = \frac{9.0 \text{ mmol}}{60.0 \text{ mL}} = \mathbf{0.15 \text{ mol/L}}$$

❹ Calculate pH

pH = -log [H^+]
= - log (0.15)
pH = 0.82

After 10.0 mL of titrant has been added, the pH of the solution rises to **0.82**

d) At the equivalence point: After 100.0mL of titrant is added to the sample, what is the pH?

❶ **Stoichiometry**: How many moles of **base** are present?

	H^+	OH^-	H_2O
Moles before reaction	50.0 mL x 0.200 mol/L = **10.0 mmol**	100.0 mL x 0.100 mol/L = **10.0 mmol**	
Moles after reaction	10.0 mmol – 10.0 mmol = **0 mmol**	10.0 mmol – 10.0 mmol = **0 mmol**	

H+ and OH- are present in stoichiometric amounts.

❷ Consider all species in the solution. Which one controls the pH?

All the H^+ and OH^- have been consumed.

$NO_3^-{}_{(aq)}$, $Na^+{}_{(aq)}$, $H_2O_{(l)}$

Na^+ and NO_3^- don't affect pH.
The only H^+ ions are those produced by autoionization of water.

❸ **Equilibrium**:
What is the concentration of H^+?

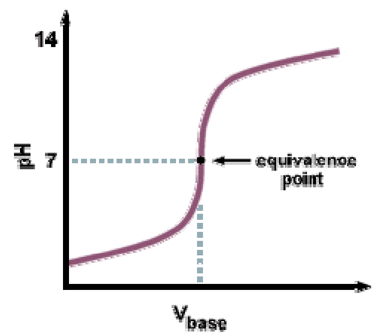
$[H^+] = 1.00 \times 10^{-7}$ mol/L (from autoionization)

❹ Calculate pH

pH = -log $[H^+]$
= - log (1.00×10^{-7})
pH = 7.00

At the equivalence point of a strong acid-strong base titration, the pH is exactly **7.00**

A typical pH curve for a strong acid-strong base titration



e) Beyond the equivalence point: After 150.0mL of titrant is added to the sample, what is the pH?

❶ **Stoichiometry**: How many moles of base are present?

	H^+	OH^-	H_2O
Moles before reaction	50.0 mL x 0.200 mol/L = **10.0 mmol**	150.0 mL x 0.100 mol/L = **15.0 mmol**	
Moles after reaction	10.0 mmol – 10.0 mmol = **0 mmol**	15.0 mmol – 10.0 mmol = **5.0 mmol**	

Now H^+ is limiting and OH^- is excess

❷ Consider all species in the solution. Which one controls the pH?

Past the equivalence point, there is no H^+ left to react with the additional OH-. Therefore, the major entities in solution are:

$OH^-{}_{(aq)}$, $Na^+{}_{(aq)}$, $NO_3^-{}_{(aq)}$, and $H_2O_{(l)}$

Since OH^- is a much stronger base than H_2O, it will control pH.

❸ **Equilibrium**:
What is the concentration of OH^-?

There is **5.0 mmol** of OH^- remaining.
Total volume of solution is
50.0 mL (acid) + 150.0 mL (base) = **200.0 mL**

Therefore, $[OH^-] = \frac{5.0 \text{ mmol}}{200.0 \text{ mL}} =$ **0.025 mol/L**

❹ Calculate pOH and then pH

pOH = -log $[OH^-]$
= - log (0.025)
pOH = 1.60

pH = 14.00 – 1.60 = **12.40**

When 150.0 mL of the titrant has been added, the pH rises to **12.40**.

Practice.

In a titration, a 25.00 mL sample of 0.350 mol/L hydrochloric acid is titrated with 0.500 mol/L sodium hydroxide.

a) What is the pH of the HCl solution BEFORE any titrant is added? *[0.456]*
b) Determine the amount of unreacted hydrogen ions after 10.00 mL of base has been added. *[3.75 mmol]*
c) Determine the pH after 10.00 mL of base has been added. *[0.971]*
d) What volume of base must be added to reach the equivalence point? *[17.5 mL]*

If you finish… work on Pg. 547 Practice #1, 2

Summary

- Titrations involving strong acids and strong bases involve mainly stoichiometric calculations.
- The equivalence point of a titration occurs when stoichiometric amounts of both reactants have been mixed.
  - ie, when molar ratios are present, and all reactant molecules from both solutions have been completely consumed
- At the equivalence point of a strong acid-strong base titration, the pH of the solution is exactly 7.00 (neutral).

Homework

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