Grade 12 Chemistry

Chemical Systems & Equilibrium

Class 13

Acid-Base Properties of Salts

- When an acid neutralizes a base, an ionic compound called a salt is formed
- **Salt** – ionic compound consisting of a cation and anion
- **Hydrolysis** – the reaction of a salt with water

Strong Acid + Strong Base → Neutral salt

Strong Acid + Weak Base → Acidic salt

Weak Acid + Strong Base → Basic salt

Neutral Salt Solutions

$NaOH + HCl \leftrightarrows \mathbf{NaCl} + H_2O \leftrightarrows Na^+ + Cl^-$

- The Na^+ and the Cl^- do not react with water to produce H_3O^+ or OH^-
  - These ions become hydrated (surrounded by water molecules)
- The following do not react appreciably with water to produce H_3O^+ or OH^-
  - Cations from strong bases (Group 1 and 2 except Be)
  - Anions from strong acids (Cl^-, Br^-, I^-, NO_3^-, ClO_4^-)

Acidic Salt Solutions

$NH_3 + HCl \leftrightarrows \mathbf{NH_4Cl} \leftrightarrows NH_4^+ + Cl^-$

- NH_4^+ is a stronger acid than water (it's the conjugate acid of NH_3, a weak base) and will react with water to form H_3O^+

$$NH_4^+ + H_2O \leftrightarrows NH_3 + H_3O^+$$

- Cl^- will not react with water
- Presence of H_3O^+ will yield an acidic salt solution

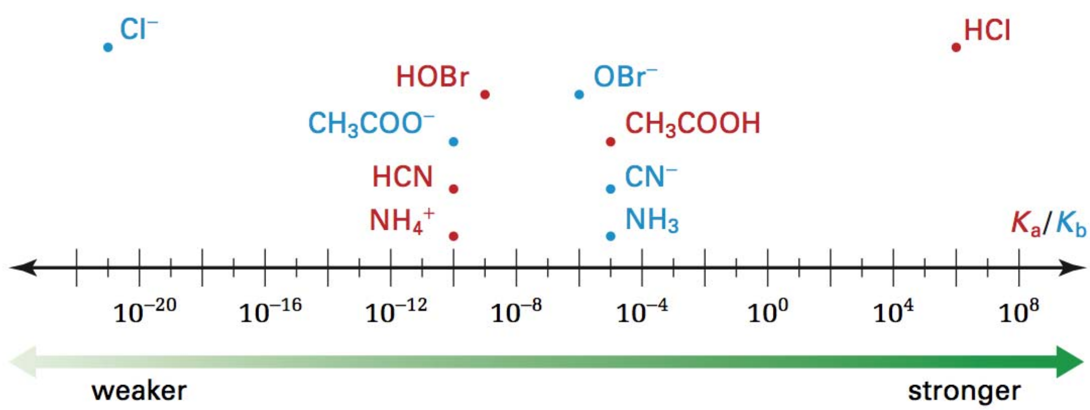
Basic Salt Solutions

$$NaOH + HCN \leftrightarrows \mathbf{NaCN} + H_2O \leftrightarrows Na^+ + CN^-$$

- Na^+ will not react with water
- CN^- is a stronger base than water (it's the conjugate base of HCN, a weak acid) and will react with water to form OH^-

$$CN^- + H_2O \leftrightarrows HCN + OH^-$$

- Presence of OH^- will yield a basic salt solution



Salts of Weak Acids and Bases

$$NH_4CN \leftrightarrows NH_4^+ + CN^-$$

- NH_4^+ is a stronger acid than water (it's the conjugate acid of a weak base)

$$NH_4^+ + H_2O \leftrightarrows NH_3 + H_3O^+$$

- CN^- is a stronger base than water (it's the conjugate base of a weak acid)

$$CN^- + H_2O \leftrightarrows HCN + OH^-$$

- Who wins?

- In this case, you need to know the K_a and K_b values:
  - $K_a > K_b$: solution will be acidic
  - $K_a = K_b$: solution will be neutral
  - $K_a < K_b$: solution will be basic

$$NH_4^+ + H_2O \leftrightarrows NH_3 + H_3O^+ \quad K_a = 5.6 \times 10^{-10}$$

$$CN^- + H_2O \leftrightarrows HCN + OH^- \quad K_b = 1.6 \times 10^{-5}$$

$K_a < K_b$, therefore the solution will be basic

Checkpoint

Classify each of the following salt solutions as acidic, basic, or neutral.

a) Na_3PO_4

b) NH_4NO_3

c) $NaCl$

d) NH_4HCO_3

e) NH_4F $\quad NH_4^+$ $K_a = 5.6 \times 10^{-10}$

$\quad\quad\quad F^-$ $K_b = 1.5 \times 10^{-11}$

pK_a and pK_b

- Measures of the strengths of acids and bases

Acids

$$HA + H_2O \leftrightarrows H_3O^+ + A^-$$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

$$pK_a = -\log K_a$$

- The smaller the value of pK_a, the stronger the acid

Bases

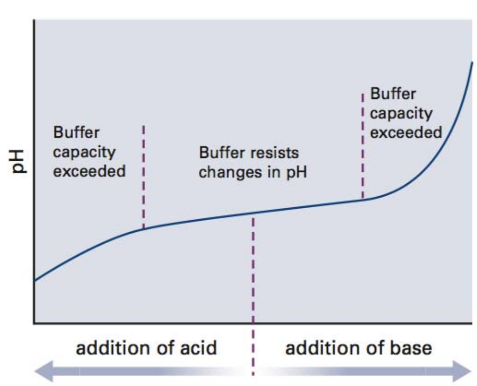
$$B + H_2O \leftrightarrows BH^+ + OH^-$$

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

$$pK_b = -\log K_b$$

- The smaller the value of pK_b, the stronger the base

Buffer Solutions



- A solution that contains a weak acid/conjugate base mixture or a weak base/conjugate acid mixture
- Buffer solutions resist changes in pH when a moderate amount of an acid or base is added

Blood Buffers

- Arterial blood pH is 7.4
- If blood pH drops to 7.0 (acidosis) or rises above 7.5 (alkalosis), life-threatening problems develop
  - Acidosis: CNS depression, disorientation, comatose
  - Alkalosis: Spasms, convulsions
- Blood contains blood buffers in the form of hydrogen carbonate ions

- Three mechanisms maintain blood pH:

1) Acid-Base Buffer System

When H^+ is added: $H^+ + HCO_3^- \rightarrow H_2CO_3$

When OH^- is added: $OH^- + H_2CO_3 \rightarrow HCO_3^- + H_2O$

– Although this can buffer drastic changes, this buffer can be overwhelmed

2) Respiratory Centre

Too Basic: $H_2O + CO_2 \rightarrow H_2CO_3 \rightarrow HCO_3^- + H^+$

Too Acidic: $H_2O + CO_2 \leftarrow H_2CO_3 \leftarrow HCO_3^- + H^+$

3) Kidney control excretion and absorption of H^+ and HCO_3^- at the nephron

- Buffer solutions can be made in two different ways:
  1. Weak Acid and one of its salts
     - Ex: CH_3COOH and CH_3COONa
  2. Weak Base and one of its salts
     - Ex: NH_3 and NH_4Cl
- Acids and bases are removed by the weak acid/base and its salts
- **Buffer Capacity** – the amount of acid or base that can be added before considerable change occurs to the pH

Henderson-Hasselbalch Approximation

- Allows us to approximate the pH of a buffer solution

For a weak acid: $HA + H_2O \leftrightarrows H_3O^+ + A^-$

$$pH = pK_a + \log\frac{[A^-]}{[HA]}$$

- Since HA is a weak acid and does not dissociate completely, we can say that $[HA] \approx [HA]_i$ and $[A^-] \approx [A^-]_i$

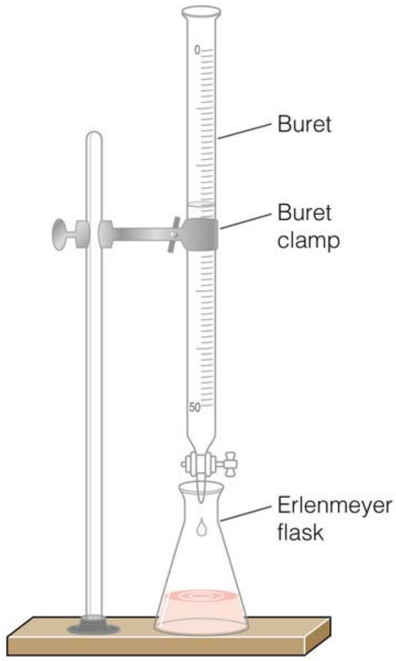
Checkpoint

What ratio of $[A^-]/[HA]$ will create an acetic acid buffer of pH 5.0? (K_a of acetic acid is 1.75×10^{-5})

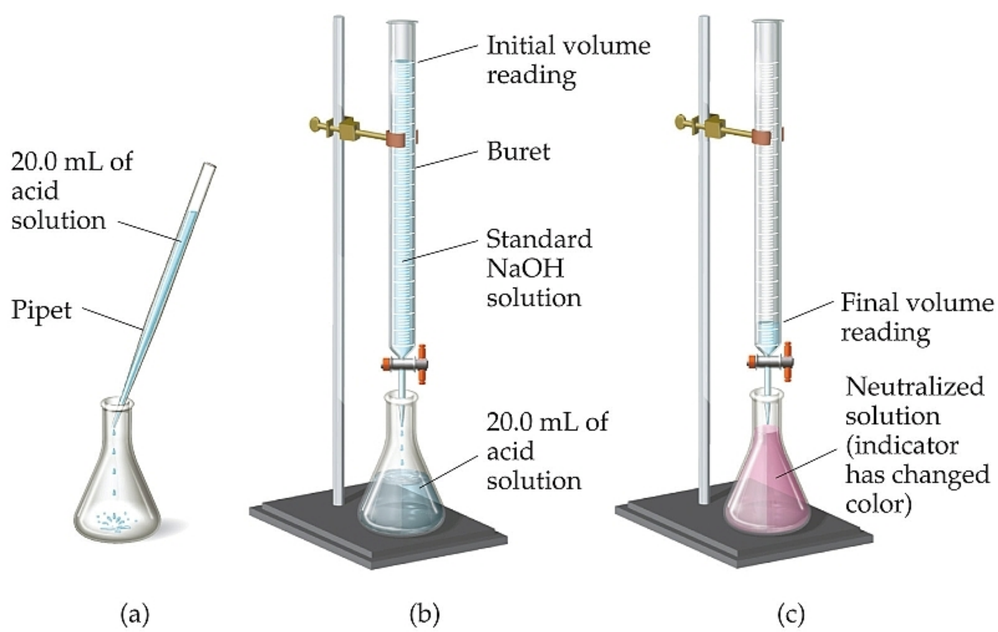
Checkpoint

Find $[H^+]$ in a solution 1.0M HNO_2 and 0.225M $NaNO_2$. The K_a for HNO_2 is 5.6×10^{-4}.

Acid Base Titrations



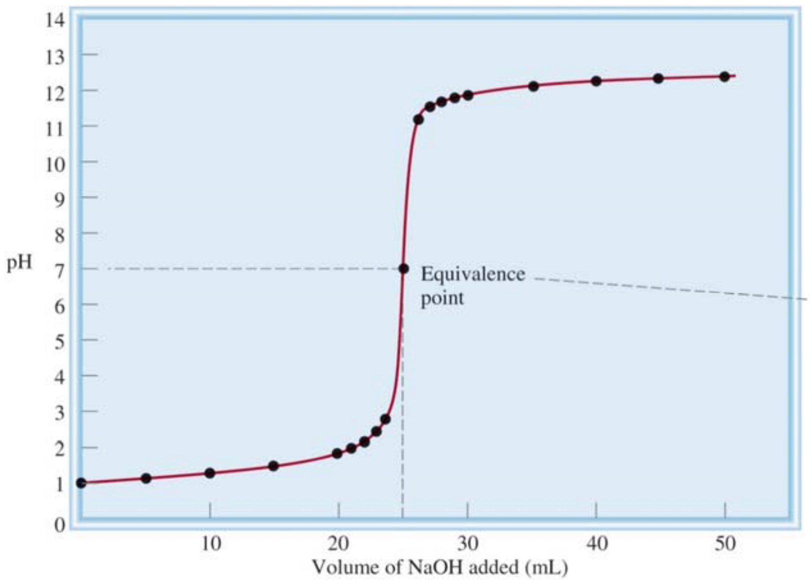
- Used to determine the concentration of one of the reactants
- Three types:
  1. Titrations of Strong Acid + Strong Base
  2. Titrations of Strong Acid + Weak Base
  3. Titrations of Weak Acid + Strong Base



- **Equivalence Point** – the point at which equimolar amounts of acid and base have reacted

Strong Acid + Strong Base Titration

$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$



Volume NaOH added (mL)	pH
0.0	1.00
5.0	1.18
10.0	1.37
15.0	1.60
20.0	1.95
22.0	2.20
24.0	2.69
25.0	7.00
26.0	11.29
28.0	11.75
30.0	11.96
35.0	12.22
40.0	12.36
45.0	12.46
50.0	12.52

Checkpoint

Calculate the pH in the titration of 0.100M NaOH to 25.0mL of 0.100M HCl after the addition of:

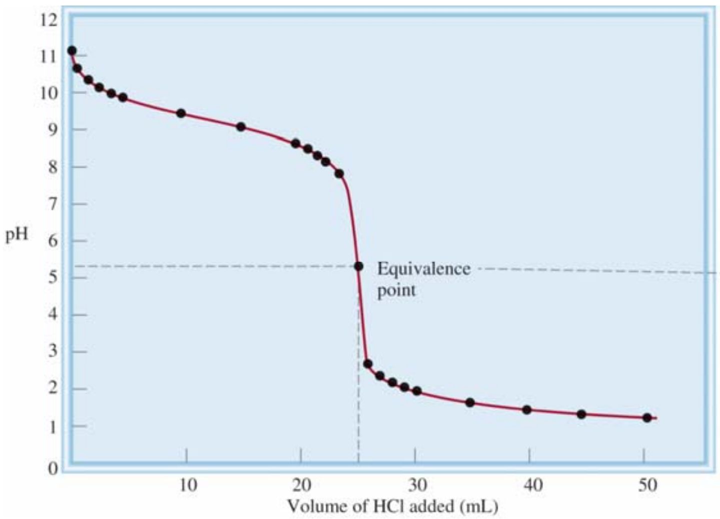
a) 10.0mL of NaOH

b) 25.0mL of NaOH

c) 35.0mL of NaOH

Strong Acid + Weak Base Titration

$$HCl(aq) + NH_3(aq) \rightarrow NH_4Cl(aq)$$



Volume HCl added (mL)	pH
0.0	11.13
5.0	9.86
10.0	9.44
15.0	9.08
20.0	8.66
22.0	8.39
24.0	7.88
25.0	5.28
26.0	2.70
28.0	2.22
30.0	2.00
35.0	1.70
40.0	1.52
45.0	1.40
50.0	1.30

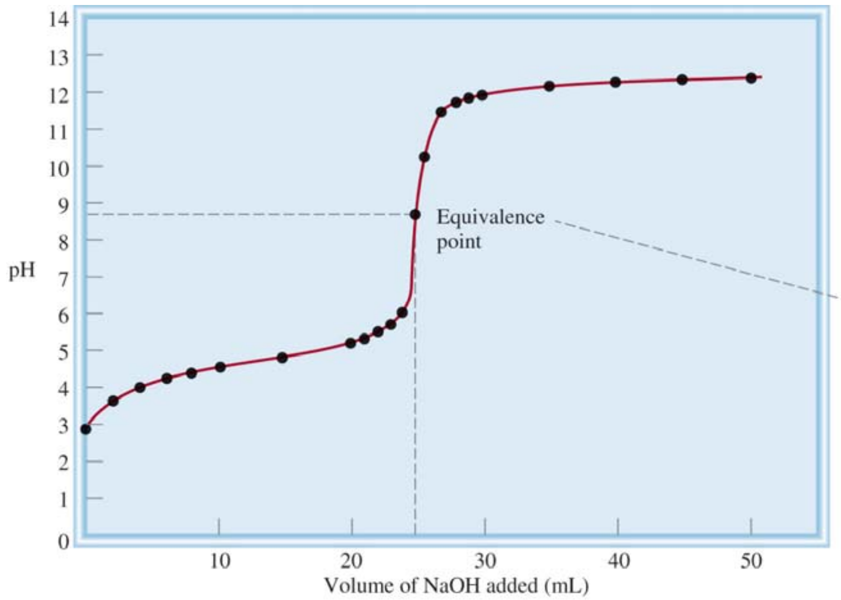
- *Due to volatility of the $NH_3(aq)$, it is safer to add HCl from a buret to the $NH_3(aq)$*

Checkpoint

Calculate the pH at equivalence point when 25.0mL of 0.100M NH_3 is titrated with 25.0mL of 0.100M HCl solution.

Weak Acid + Strong Base Titration

$CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$



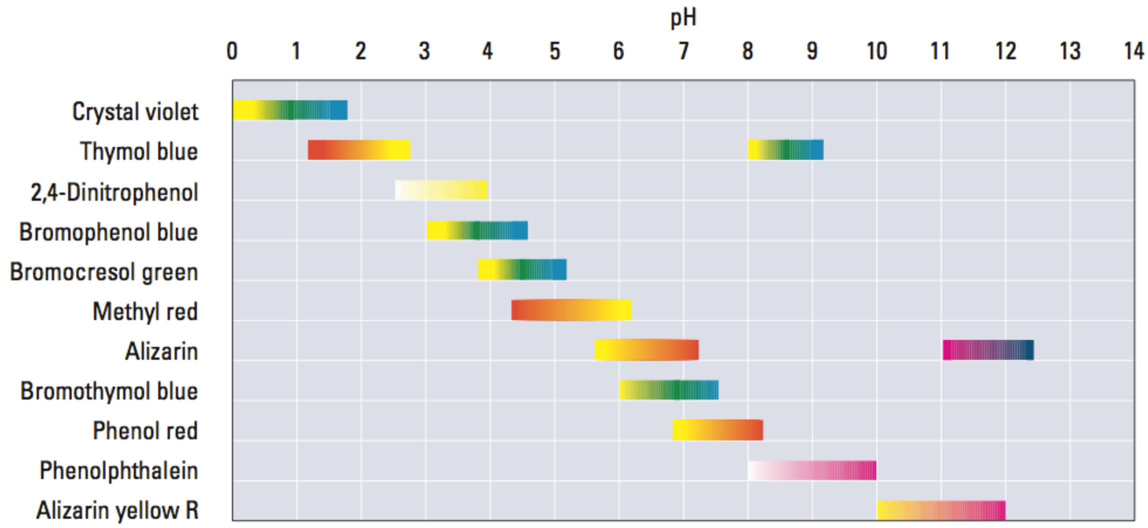
Volume NaOH added (mL)	pH
0.0	2.87
5.0	4.14
10.0	4.57
15.0	4.92
20.0	5.35
22.0	5.61
24.0	6.12
25.0	8.72
26.0	10.29
28.0	11.75
30.0	11.96
35.0	12.22
40.0	12.36
45.0	12.46
50.0	12.52

Acid-Base Indicators

- The indicator is usually a weak, monoprotic organic acid or base that has two distinct colours when dissociated

$$\underset{\textbf{colour 1}}{HIn_{(aq)}} + H_2O_{(\ell)} \rightleftharpoons H_3O^+{}_{(aq)} + \underset{\textbf{colour 2}}{In^-{}_{(aq)}}$$

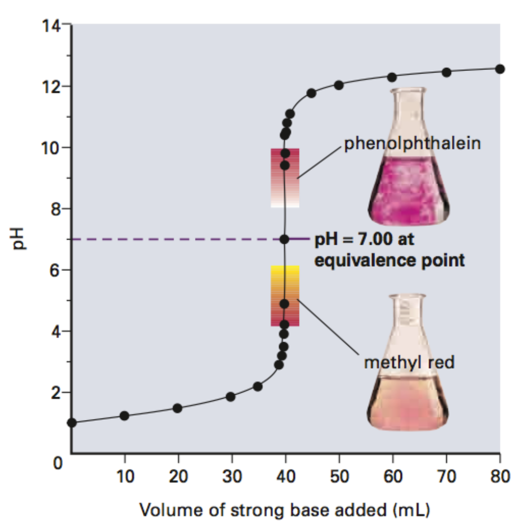
- End Point – occurs when the indicator changes colour
- By choosing the proper indicator, you can use the end point to determine the equivalence point



- Choose an indicator whose end point lies on the steep part of the titration curve

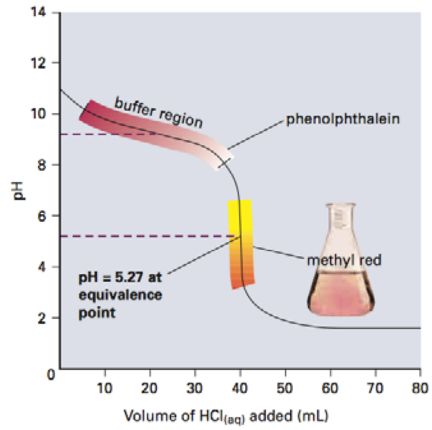
Table 8.4 Data on the Endpoints of Three Common Acid-Base Indicators

Indicator	Colour change at endpoint	Approximate range
bromocresol green	yellow to blue	3.8–5.2
methyl red	red to yellow	4.3–6.2
phenolphthalein	colourless to pink	8.2–10.0



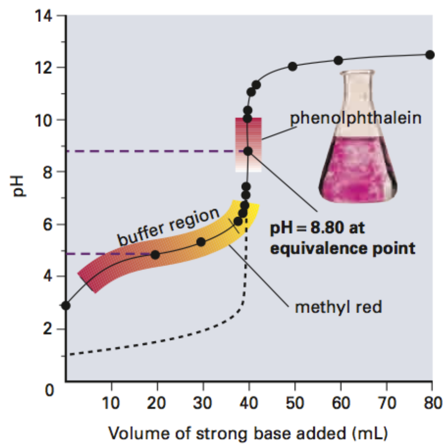
Scenario #1: Strong Acid with a Strong Base

- Use Phenolphthalein, Methyl Red



Scenario #2: Strong Acid with Weak Base

- Use Methyl Red or Bromocresol Green



Scenario #3: Weak Acid with Strong Base

- Use Phenolphthalein

Checkpoint

20ml of 0.20M NaOH(aq) is titrated against 0.20M HF (aq). Calculate the pH at equivalence and find an appropriate indicator.