Chemistry 30 IB
Acids & Bases

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Unfinished!

Contents

1 Theories

The following two equations mean the same thing.

H^+(aq) and H_3O^+(aq) are interchangable.

1.1 Arrhenius

$$HX(aq) \longrightarrow H^+(aq) + X^-(aq)$$

- doesn't specifically state water is present (aq)
- uses hydrogen ions, H^+(aq)
- cannot determine strong or weak

1.2 Brønsted-Lowry (aka. Modified Arrhenius)

$$HX(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + X^-(aq)$$
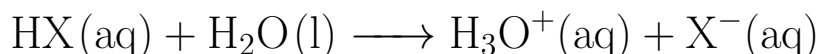
- specifically states water is present
- uses hydronium ions, H_3O^+(aq)
- can determine strong or weak

2 General Equations

- Chemical entities not present in the data booklet pages 8-9—such as K^+, Mg2+, Na^+, Cl^-—are spectators and do not need to be included in the equations.
- Chemical entities from the first six rows in the data booklet pages 8-9—such as ClO_4^-— do not need to be included in equations, as equations containing said entities ionize/dissociate completely, leaving the hydronium/hydroxide ions by themselves.

2.1 Ionization of Acids

Forming ions from molecular compounds.

2.1.1 Strong

$$HX(aq) + H_2O(l) \xrightarrow{> 99.9\%} H_3O^+(aq) + X^-(aq)$$
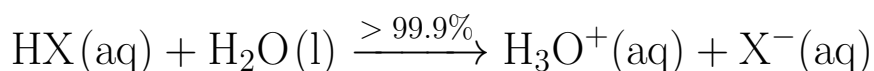
- ionize completely ($> 99.9\%$ of the reaction completes)
- irreversible ($\longrightarrow$)
- high K value ($K > 1$)

2.1.2 Weak

$$HX\,(aq) + H_2O\,(l) \overset{< 50\%}{\rightleftharpoons} H_3O^+(aq) + X^-(aq)$$

- do not ionize completely ($< 50\%$ of the reaction completes)
- reversible ($\rightleftharpoons$)
- ionize at equilibrium
- low K value ($K < 1$)

2.2 Dissociation of Bases

Separation of existing ions in solution.

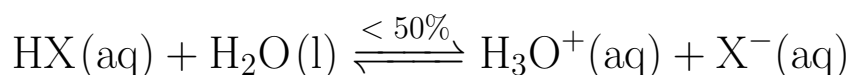### 2.2.1 Strong

$$M(OH)_n + H_2O\,(l) \xrightarrow{> 99.9\%} M^{n+}(aq) + nOH^-(aq)$$
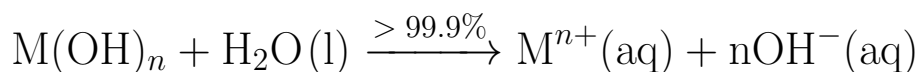
- M is a metal, $M(OH)_n$ is highly soluble
- dissociate quantitatively

2.2.2 Weak

$$X\,(aq) + H_2O\,(l) \overset{< 50\%}{\rightleftharpoons} HX^+(aq) + OH^-(aq)$$
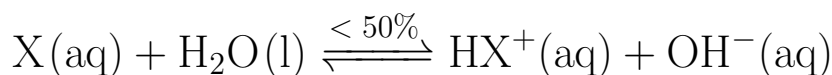
- dissociate at equilibrium

3 pH & pOH

3.1 pH/pOH from Concentration

$$pH = -\log\left[H_3O^+\right]$$

$$pOH = -\log\left[OH^-\right]$$

3.1.1 Shortcut

The absolute value of the exponent of a concentration is the pH/pOH, for hydronium and hydroxide concentrations respectively.

e.g. $1.00 \times 10^{-4}\,\mathrm{mol\,L^{-1}}$ of H_3O^+ has a $pH = 4$

3.2 Estimation

This isn't required to know.

This only applies if the base is equal to 1.00. If the base is not 1.00, you can predict ranges that the pH/pOH could be.

Of course, you can just calculate it, but this is a trick you can use to quickly figure out and compare pH/pOH from concentration.

- If [ ] = 1, then $\mathrm{pH} = |exponent|$
- If [ ] > 1, then $\mathrm{pH} < |exponent|$
- If [ ] < 1, then $\mathrm{pH} > |exponent|$

3.3 Concentration from pH/pOH

$$[\mathrm{H_3O^+}] = 10^{-\mathrm{pH}}$$

$$[\mathrm{OH^-}] = 10^{-\mathrm{pOH}}$$

3.4 Switching between pH and pOH

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

3.5 K_w

The equilibrium constant of water can be used to solve for hydrogen ion concentration or hydronium ion concentration when you have the other.

$$K_w = [\mathrm{H_3O^+}][\mathrm{OH^-}]$$

$$K_w = 1.00 \times 10^{-14}\,\mathrm{mol\,L^{-1}}$$

$$1.00 \times 10^{-14}\,\mathrm{mol\,L^{-1}} = [\mathrm{H_3O^+}][\mathrm{OH^-}]$$

3.6 Significant Digits

The significant digits of the concentration is the number of decimal pleaces of the pH/pOH.

For instance,

$$[H_3O^+] = 2.3 \times 10^{-6} \frac{mol}{L} \qquad pH = 5.64$$

3.7 Square

This may help you remember when to use what.

$$K_W = [H_3O^+][OH^-] = 1.0 \times 10^{-14} \frac{mol^2}{L^2} \quad (p.3)$$
$$(1.0 \times 10^{-7})(1.0 \times 10^{-7})$$

$$[H_3O^+] \rightleftharpoons [OH^-]$$

$$pH = -\log[H_3O^+] \qquad [H_3O^+] = 10^{-pH}$$

$$[OH^-] = 10^{-pOH} \qquad pOH = -\log[OH^-]$$

$$pH \longleftrightarrow pOH$$

$$pH + pOH = 14.00$$
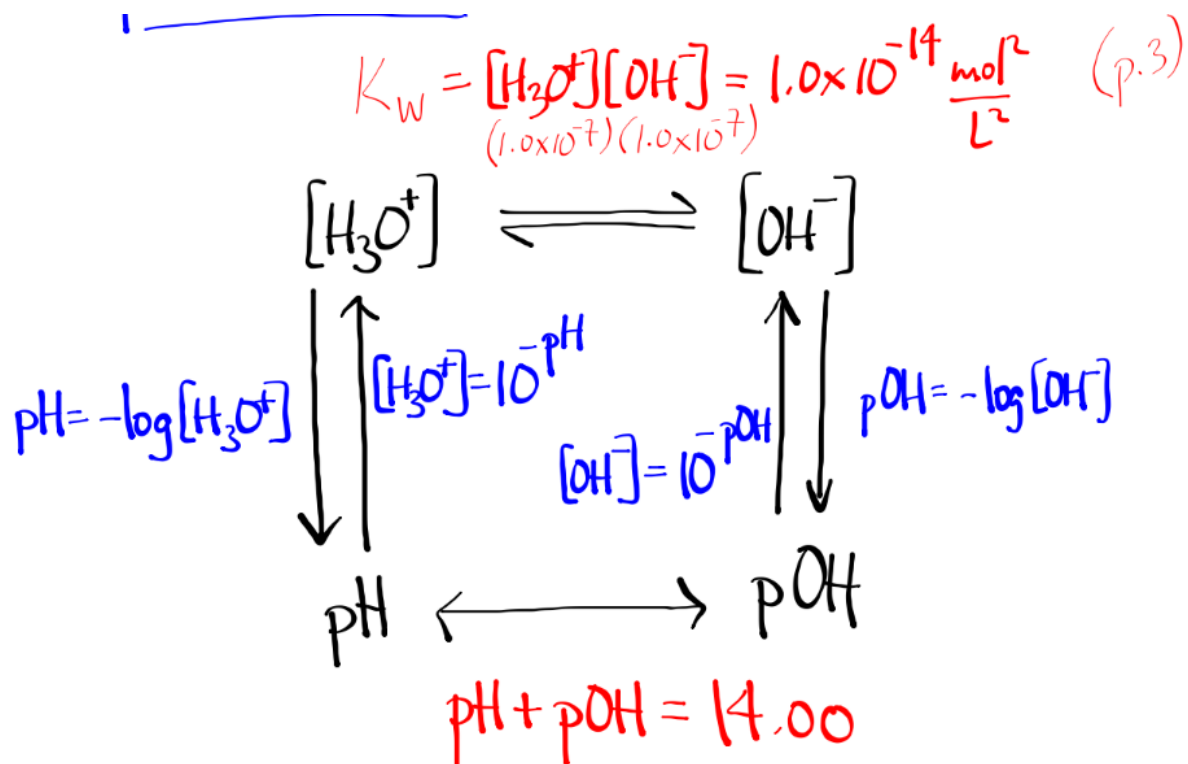
3.8 Example

Determine the pH of a 1.23 mol/L NaOH solution.

$$NaOH_{(aq)} \xrightarrow{>99.9\%} Na^+_{(aq)} + OH^-_{(aq)}$$

1.23 M   1.23 M
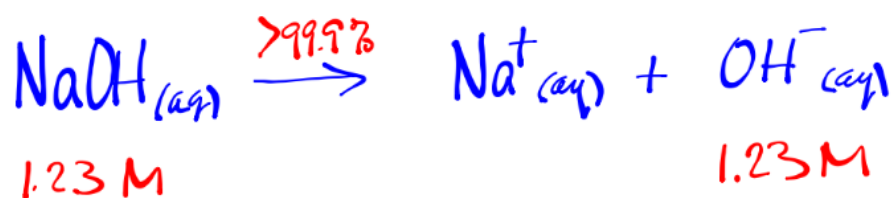
(Quantitative 1:1 dissociation of strong base).

$$pOH = -\log\left(1.23 \frac{mol}{L}\right)$$
$$= -0.090$$

$$pH = 14.00 - ^-0.090$$
$$= 14.09$$

$$[H_3O^+] = \frac{1.0 \times 10^{-14} \frac{mol^2}{L^2}}{1.23 \frac{mol}{L}}$$
$$= 8.1 \times 10^{-15} \frac{mol}{L}$$

$$pH = -\log(8.1 \times 10^{-15})$$
$$= 14.09$$

4 Acid/Base Reaction

Also known as a neutralization reaction.

The complete or partial transfer of protons (H^+) from the strongest acid to the strongest base. The acids donate protons, the bases accept protons.

ure.

ACID + BASE $\xrightarrow[\text{partial}]{\text{complete}}$ CONJUGATE ACID + CONJUGATE BASE
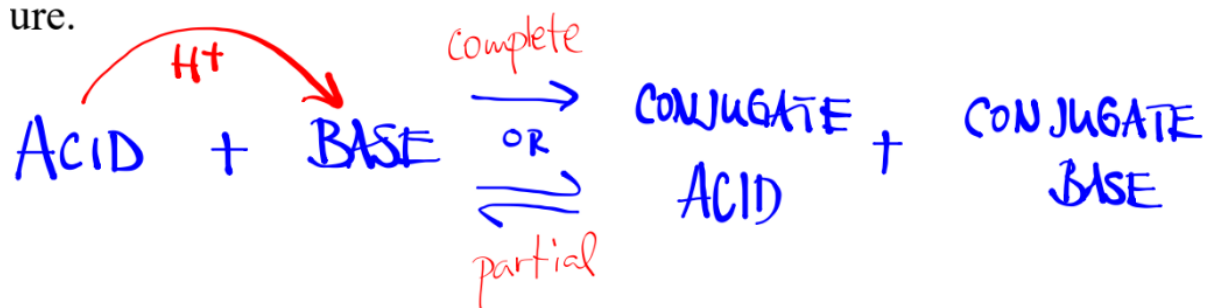
7

4.1 Net Reaction

At this level, we typically only see the net reaction.

For your information, this is what leads to the net reaction.

1. **Molecular equation:**
$$HCl(aq) + NaOH(aq) \rightleftharpoons H_2O(l) + NaCl(aq)$$

2. **Ionic Equation:**
$$H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightleftharpoons H_2O(l) + Na^+(aq) + Cl^-(aq)$$
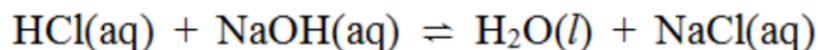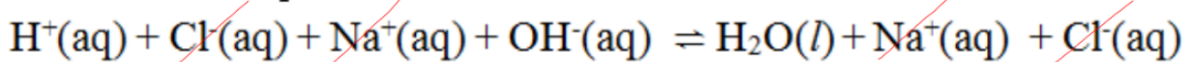
3. **Net Equation (Reaction):**
$$H^+(aq) + OH^-(aq) \rightleftharpoons H_2O(l)$$
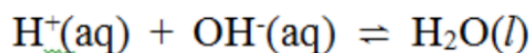
* ionize/dissociate all strong or highly soluble substances

remove all spectators for Brønsted-Lowry reactions.

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4.2 Example Net Reactions

a. $HNO_2(aq)$ and ~~K~~OH(aq)
b. ~~HCl~~(aq) and ~~Na~~OH(aq)
c. $CH_3COOH(aq)$ and ~~K~~OH(aq)
d. ~~HClO_4~~(aq) and ~~Mg(OH)_2~~(aq)

a.
$$HNO_{2(aq)} + OH^-_{(aq)} \xrightarrow{>99.9\%} NO^-_{2(aq)} + H_2O_{(l)}$$
 A   B   CB   CA

b.
$$H_3O^+_{(aq)} + OH^-_{(aq)} \xrightarrow{>99.9\%} H_2O_{(l)} + H_2O_{(l)}$$
 A   B   CB   CA

c.
$$CH_3COOH_{(aq)} + OH^-_{(aq)} \xrightarrow{>99.9\%} CH_3COO^-_{(aq)} + H_2O_{(l)}$$
 A   B   CB   CA

d.
$$H_3O^+_{(aq)} + OH^-_{(aq)} \xrightarrow{>99.9\%} H_2O_{(l)} + H_2O_{(l)}$$
 A   B   CB   CA
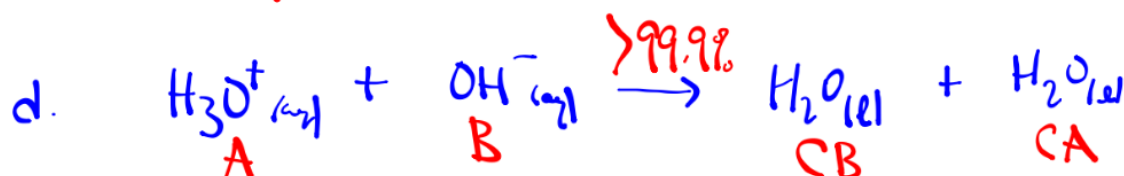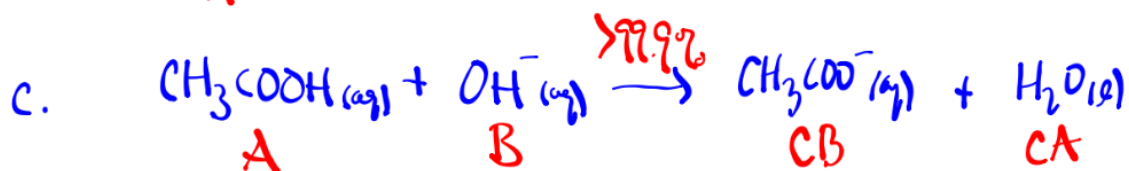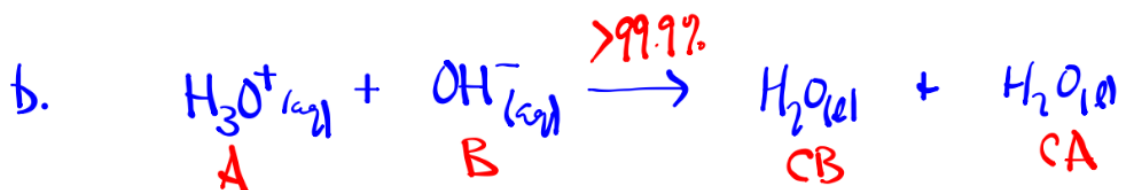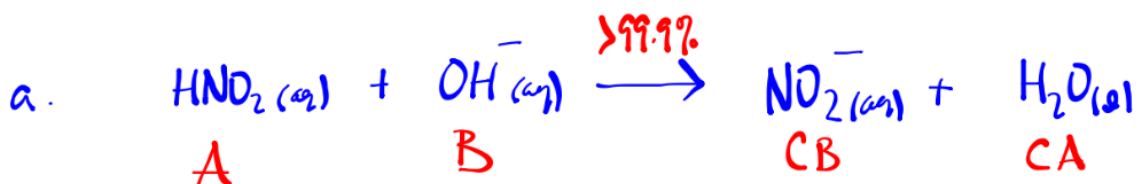
4.3 Conjugate

- The conjugate base (CB) is what an acid becomes on the product side.
- The conjugate acid (CA) is what a base becomes on the product side.

Conjugates differ from their original by one proton. (H^+)

This is because, say a compound loses a proton from left to right, then it is an acid. Read that in reverse, right to left, and it gains a proton, making it a base.

$$\overset{\angle 50\%}{\underset{\underset{\displaystyle \text{CA}}{}}{}}$$

$$\underset{\text{B}}{\text{H}_2\text{O}(l)} + \underset{\text{A}}{\text{HNO}_2(aq)} \rightleftharpoons \underset{\text{CA}}{\text{H}_3\text{O}^+(aq)} + \underset{\text{CB}}{\text{NO}_2^-(aq)}$$

$$\underset{\text{A}}{\text{H}_2\text{O}(l)} + \underset{\text{B}}{\text{S}^{2-}(aq)} \overset{<50\%}{\rightleftharpoons} \underset{\text{CB}}{\text{OH}^-(aq)} + \underset{\text{CA}}{\text{HS}^-(aq)}$$
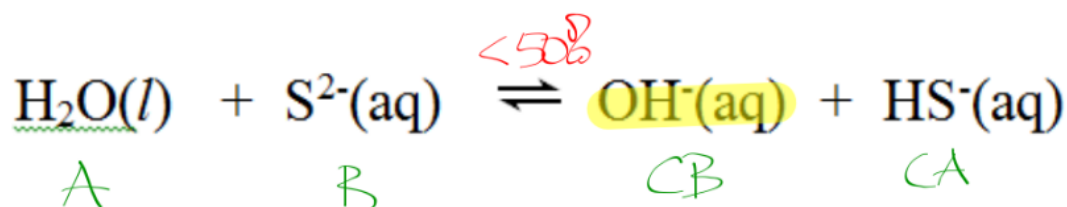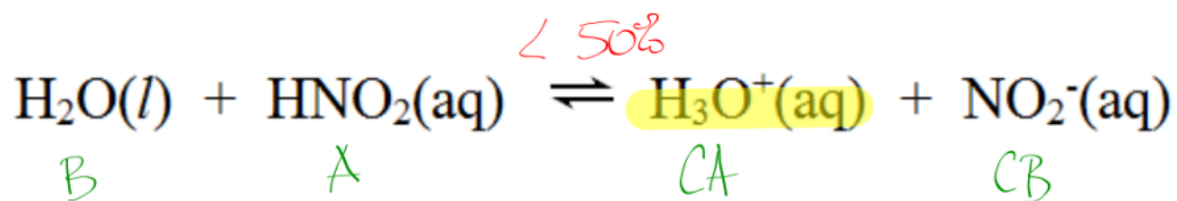
A conjugate pair would be either,

- an acid and its conjugate base
- a base and its conjugate acid

4.4   Polyprotic Acids

Acids that can donate a proton more than once.

They do it one at a time, yielding a different K_a value every time.

- $\text{H}_2\text{CO}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{HCO}_3^-(aq)$ $K_a = 4.4 \times 10^{-7}$
- $\text{HCO}_3^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{CO}_3^{2-}(aq)$ $K_a = 4.7 \times 10^{-11}$

For calculating the pH of a polyprotic acid, the K_a of the first ionization should be used for calculations.

The conjugate base of polyprotic acids are amphiprotic.

4.5   Amphiprotic

Sometimes called amphoteric. Water is amphiprotic.

Substances that are both a base and an acid. Whether they behave as an acid or a base depends on the other substance in the acid/base reaction.

It will do the opposite of the other substance, so if it were an acid, the amphiprotic will behave as a base, and vice versa.

Amphiprotics also often are a negative ion with hydrogen—HX⁻—such as $\text{HSO}_4^-(aq)$.

If both substances in an acid/base reaction can behave as an acid, the stronger acid will behave as an acid, the other as a base.

4.6 Favourability

4.6.1 Hydronium or Hydroxide Present

If H_3O^+ or OH^- are present, the reaction will favour the side they are not on.

If either one of these two ions are,

- a reactant, then products are favored. $\xrightarrow{> 99.9\%}$
- a product, then reactants are favored. $\underset{< 50\%}{\rightleftharpoons}$

4.6.2 No Hydronium/Hydroxide

On page 8-9 in your data booklet,

- if the acid is above the base, then products are favored. $\xrightarrow{> 99.9\%}$
- if the base is above the acid, then reactants are favored. $\underset{< 50\%}{\rightleftharpoons}$

The higher on the table, the stronger the acid. If the acid is stronger, it'll more likely donate protons, increasing the concentration of the opposing side of the reaction.

5 (AB) Relative Strength

The following is, oddly enough, non-IB.

- K_a = acid equilibrium constant
  how much $H_3O^+(aq)$ is produced when an acid interacts with water
- K_b = base equilibrium constant
  how much $OH^-(aq)$ is produced when an base interacts with water

Equilibrium constants are calculated the same as usual. Product of concentrations of products divided by product of concentrations of reactants, not including liquid water.

5.1 K_a of Common Substances

The equilibrium constant of many common acids is provided on page 8-9 of the Alberta data booklet, on the right-most column.

But what if you want K_b?

5.2 Converting between acid/base equilibrium constants

$$K_w = K_a \times K_b$$

This means you can calculate K_b as long as you have K_a, since K_w is a known constant.

The K_a of a base on the table is actually the conjugate acid. Doesn't really change anything—they are on the same line, after all—but good to know.

$$K_b = \frac{1.0 \times 10^{-14}\ \text{mol L}^{-1}}{K_a}$$

If the $K_b > K_a$, than the base is a stronger base than the conjugate is an acid. This doesn't always mean its a base—its just a stronger base.

5.2.1 Example

The target base is ammonia, NH_3. The conjugate acid of ammonia, ammonium $NH_4{}^+$, has a K_a listed on the table—$5.6 \times 10^{-10}\ \text{mol L}^{-1}$.

Therefore,

$$K_b = \frac{1.0 \times 10^{-14}\ \text{mol L}^{-1}}{5.6 \times 10^{-10}\ \text{mol L}^{-1}}$$

$$K_b = 1.7 \times 10^{-5}\ \text{mol L}^{-1}$$

Ammonia NH_3 is a stronger base than ammonium $NH_4{}^+$ is an acid.

6 pH/pOH and Ion Concentration of Weak Acids/Bases

6.1 Strong

Strong acids and bases ionize/dissociate completely, respectively. The concentrations of the acid/base is equal to the concentration of hydronium/hydroxide ions. They ionize/dissociate quantitatively.

$$[\text{acid}] = [\text{H}_3\text{O}^+] \qquad\qquad [\text{base}] = [\text{OH}^-]$$

6.2 Weak

If you want the full explanation of where these formulas are derived from, read the actual notes—page 38. The following are the shortcuts you truly only need.

6.2.1 Known Ion Concentration

The following should be used if you know the ion concentration or pH/pOH, since pH/pOH can be converted into ion concentration.

$$K_a = \frac{[\text{H}_3\text{O}^+]^2}{[\text{acid}] - [\text{H}_3\text{O}^+]} \qquad\qquad K_b = \frac{[\text{OH}^-]^2}{[\text{base}] - [\text{OH}-]}$$

6.2.2 Unknown Ion Concentration

If you are solving for ion concentration with the aforementioned formula, you will require the quadratic formula. To avoid this, the ion concentration in the denominator can be omitted, since it affects the calculation negligably.

$$K_a = \frac{[\text{H}_3\text{O}^+]^2}{[\text{acid}]} \qquad K_b = \frac{[\text{OH}^-]^2}{[\text{base}]}$$

Therefore, you can now find ion concentration of weak acids/bases.

$$[\text{H}_3\text{O}^+] = \sqrt{K_a \cdot [\text{acid}]} \qquad [\text{OH}^-] = \sqrt{K_b \cdot [\text{base}]}$$
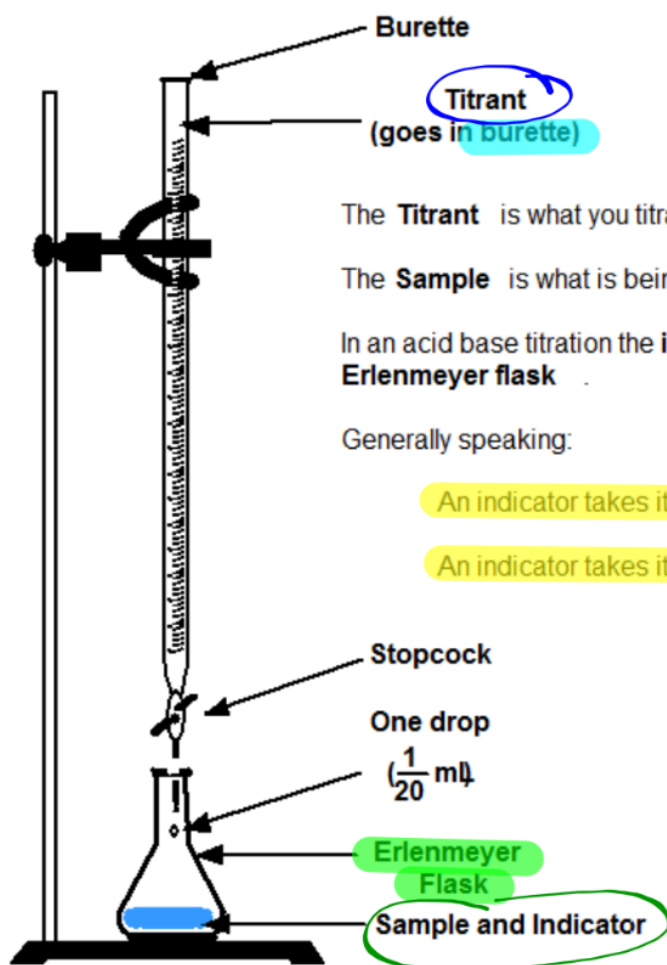
7   Acid Base Titrations

Titration is a method for measuring the concentration of a solution.

The **titrant** is the solution in the burette, typically just an acid or base.

The **sample** is the solution in the Erlenmeyer flask, typically composed of a solution and an indicator. The indicator can be used to determine the pH of the solution after titrating the titrant.

Indicators are given to you on page 10 of the data booklet. Each column has the acid variant on the left and base variant on the right. For instance, bromothymol blue is yellow below 6.0 pH, blue above 7.6 pH, and green in between (indicating the equivalence point).

Burette

Titrant
(goes in burette)

The **Titrant** is what you titrate with.

The **Sample** is what is being titrated.

In an acid base titration the **indicator** is added to the sample in the **Erlenmeyer flask** .

Generally speaking:

An indicator takes its **acid form** in an **acid sample** , and

An indicator takes its **base form** in a **base sample** .

Stopcock

One drop
$(\frac{1}{20}$ ml$)$

Erlenmeyer Flask

Sample and Indicator

bromothymol blue

HBb / Bb^{-}

(yellow) (blue)

acid   base

form   form

Endpoint color

(green)

8 Equivalence Point

- Adding base to an acid neutralizes it, $[H^+]$ decreases while $[OH^-]$ increases.
- Adding acid to an base neutralizes it, $[H^+]$ increases while $[OH^-]$ decreases.

Eventually, $[H^+] = [OH^-]$.

This is the equivalence point—when a proton is quantitatively exchanged completely. It shows up as the half way point in a steep jump/drop on neutralization graphs.

- The pH at the equivalence point is the pH of neutralization.
- The volume at the equivalence point is the volume of titrant needed to be titrated/added to neutralize sample.

8.1 Example Problem

It is just stoicheometry.

10.0 mL of an unknown concentration of NaOH is placed in a flask. A drop of indicator is added. A $2.00\,\mathrm{mol\,L^{-1}}$ solution of HCl is added until a colour change occurs. If it takes 25.0 mL of HCl to reach the equivalence point, what is the concentration of the NaOH?
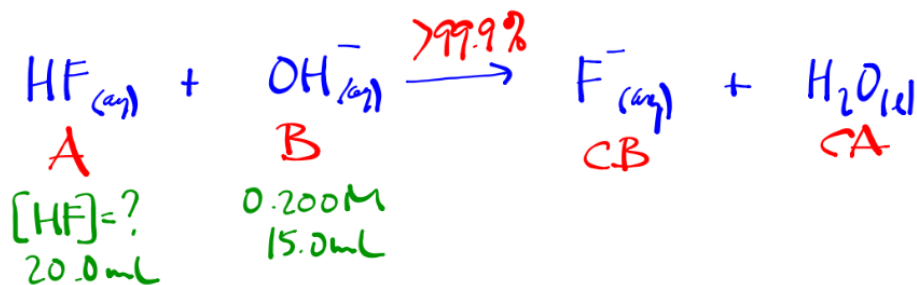
(Cl and Na are spectators.)

$$H^+(aq) + OH^-(aq) \rightleftharpoons H_2O(l)$$
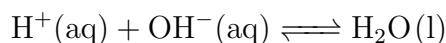
$$0.25\,\mathrm{L\ HCl} \times \frac{2.00\,\mathrm{mol\ HCl}}{1\,\mathrm{L\ HCl}} \times \frac{1\,\mathrm{mol\ NaOH}}{1\,\mathrm{mol\ HCl}} \times \frac{1}{0.0100\,\mathrm{L\ NaOH}}$$

$$= 5.00\,\mathrm{mol\,L^{-1}\ NaOH(aq)}$$

8.2 Example Problem 2

p.49

3. Find the concentration of an HF(aq) solution if 15.0 mL of 0.200 **titrant**
   mol/L KOH(aq) just neutralizes 20.0 mL of HF(aq)

   spectator   Sample

$$HF_{(aq)} + OH^-_{(aq)} \xrightarrow{>99.9\%} F^-_{(aq)} + H_2O_{(l)}$$

A   B   CB   CA

$[HF] = ?$ 0.200 M

20.0 mL   15.0 mL

$$\frac{0.200 \text{ mol } OH^-}{1 L \text{ } OH^-} \times 15.0 \text{ mL } OH^- \times \frac{1 \text{ mol } HF}{1 \text{ mol } OH^-} \times \frac{1}{20.0 \text{ mL } HF} = 0.150 \frac{\text{mol}}{L} HF$$

moles of given   mole ratio   required concentration

9 Neutralization Reactions Graphs

- A large jump/drop in the graph signifies an equivalence point
- The number of equivalence points = the number of protons transferred
- Proton exchange reactions that are at equilibrium do not result in large changes in pH

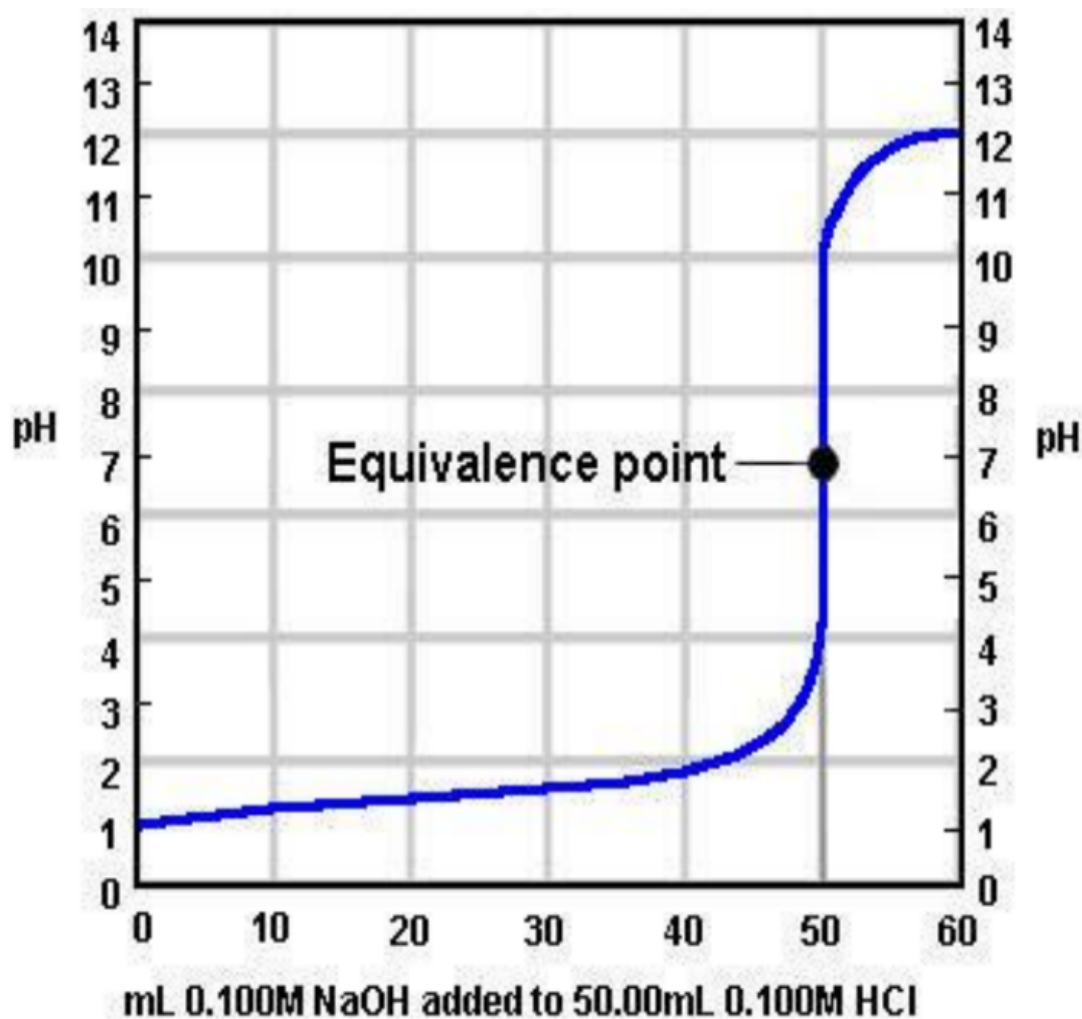## 9.1 Strong Acid & Strong Base

Figure 1: The line is both the acid and base pH, which are equal

Acidic sample—as pH starts low. Basic titrant—as pH ends high. A basic sample and acidic sample would be this graph reversed.

Equivalence point is always $= 7$.

Explanation

- pH increases slowly at first, as it initially takes a lot of acidic hydroxide ions to remove basic hydronium ions
- Once there are very little acidic ions, not a lot of basic ions are needed to increase the pH
- Therefore, the pH jumps very rapidly
- Since the acidic hydronium ion concentration is low, it will again take a lot of base to increase the hydroxide ion concentration
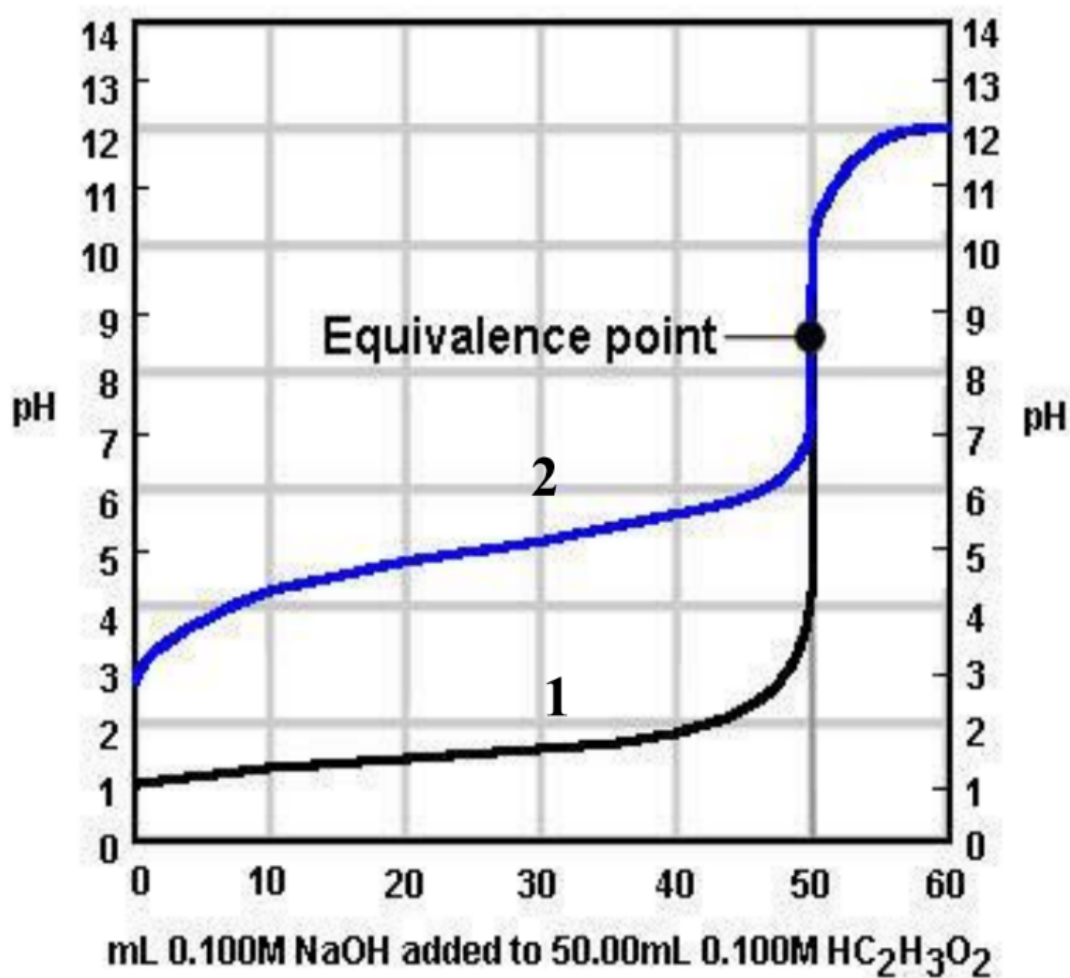
9.2   Strong Base & Weak Acid

Figure 2: Line 1 (black) is strong base / Line 2 (blue) is weak acid

Equivalence point is always > 7.

Weak Acid (Line 2) Differences

- Starts with a higher pH than the strong acid (line 1)
- Brief curve up at beginning (from 0-10 mL) before linear slope
- Region between brief curve and equivalence point (from 10-45 mL) is **buffer**
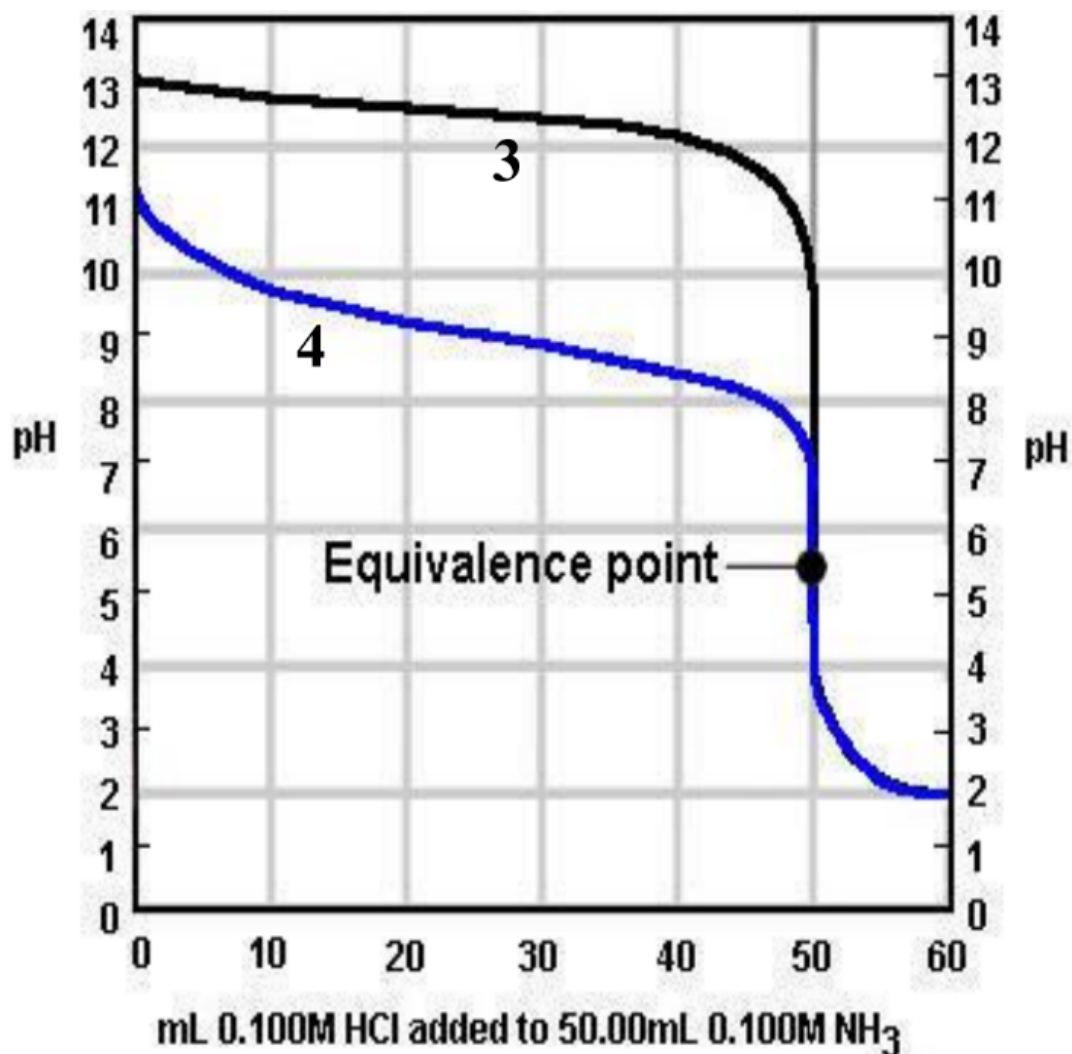
9.3 Strong Acid & Weak Base

Figure 3: Line 3 (black) is strong acid / Line 4 (blue) is weak base

Equivalence point is always < 7.

- Starts with a lower pH than the strong base (line 3)
- Brief curve down at beginning (from 0-10 mL) before linear slope
- Region between brief curve and equivalence point (from 10-45 mL) is **buffer**

9.4 Polyprotic Acid Titrations

mL 0.100M NaOH added to 25mL 0.100M H_2CO_3
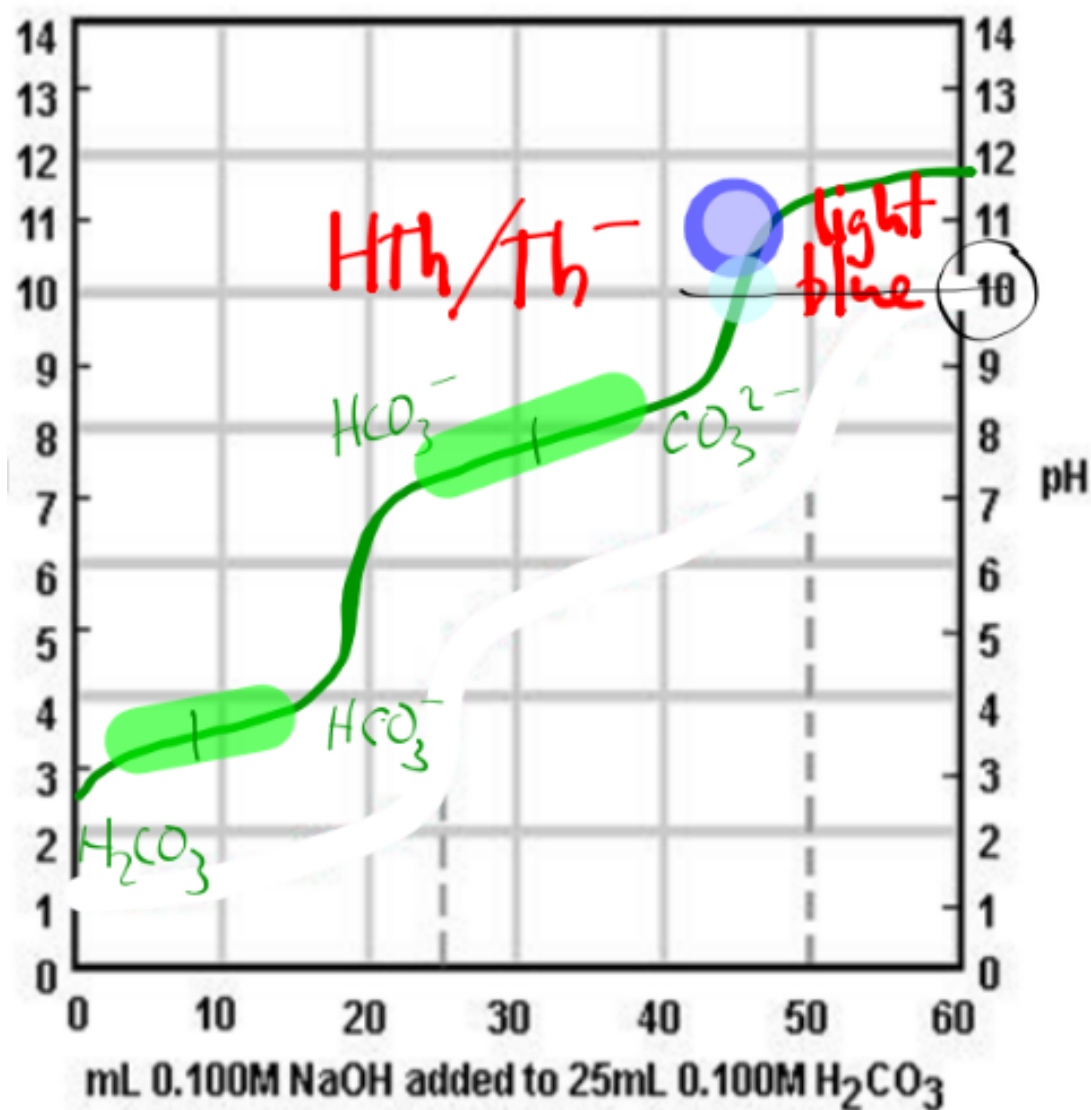
- Multiple jumps in line (multiple equivalence points)
- Flat regions between equivalence points (highlighted green) are buffers
- Last flat region is not a buffer
- Jumps after the first one are not as large