

Unit 4 - Acids, Bases, and Salts

Arrhenius Acids and Bases

- This theory of acids and bases was the first developed (over one hundred years ago), and is still relevant in many circumstances.

- Definitions:

1.) An _____ is a chemical species that releases H^+ (aq) (proton) in water.

A simpler way to say this is an **acid** is an ionic compound that starts with "H".

Ex. - H_2SO_4 , H_3PO_4 , HNO_3 , HCl

2.) A base is a chemical species which releases OH^- (aq) (hydroxide) in water.

Ex. -

3.) A salt is an ionic compound that does not contain hydrogen (H) or hydroxide (OH^-), or another way of saying this is it is an ionic compound that is not an acid or base!

Ex. -

- The most common chemical reaction involving acids and bases is an _____ reaction.

Ex. - $HCl + NaOH \rightarrow H_2O + NaCl$

Ex. 2 - $H_2SO_4 + KOH \rightarrow K_2SO_4 + H_2O$ (don't forget to balance)

- Properties of **Arrhenius acids:**

1.) acids react with bases.

2.)

3.) acids react with some metals (metals higher than hydrogen on reactivity) to form hydrogen gas.

(ex. - zinc, magnesium, potassium, iron, aluminium)

4.) acids turn litmus paper _____.

5.) acids **ARE** sour.

Common acids

1.) Sulphuric acid - H_2SO_4 **common names:** _____, _____

-

- Good electrolyte (conducts electricity)

- Used in manufacture of sulphates, in car batteries, to absorb water from solutions, and

_____.

2.) Hydrochloric acid - HCl **common names:** _____

- Good electrolyte
- Concentrated solutions have a choking odour
- Used to manufacture chlorides, to remove "lime scale" (calcium and magnesium carbonate), and cleaning metals of oxides.

3.) Nitric acid - _____

- Very reactive with metals
- Colours proteins (like skin!) yellow
- Used to manufacture _____ and _____.

- Properties of **Arrhenius bases:**

- 1.) bases react with acids
- 2.) bases form _____ in water
- 3.) bases feel slippery (if when spilled your hands are slippery lightly WASH THEM!)
- 4.) bases turn litmus paper _____.
- 5.) bases taste _____ (baking soda).

Common bases

1.) _____ - NaOH **common names:** _____ or lye

- Very corrosive (caustic) to animal and plant tissue
- Quickly " _____ " (removes water form air)
- Used in manufacture of sodium salts, in making soaps and cleaning products, and production of _____, _____, _____, and aluminium.

2.) Ammonia - _____

- Highly toxic and corrosive gas with pungent odour
- Soluble in water
- Used to manufacture nitric acid, fertilizers, explosives, and synthetic fibres.

- When acids are mixed with water the proton (H^+) released by the acid reacts with the water molecule. This creates what is known as a _____ (H_3O^+).
- Since all acid reactions take place in water the formerly mention proton (H^+) is actually always H_3O^+ .
- In the last chapter we understood the dissociation of HCl to be this:

We now know it should be written like this:

Bronsted-Lowry Acids and Bases

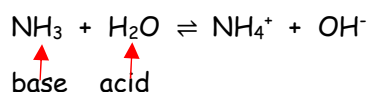
- The Bronsted-Lowry theory of acids and bases was developed to include Arrhenius acid/base theory but was needed due to the development in chemistry understanding that incorporated equilibrium.

- Definitions:

- 4.) An **acid** is a chemical species that _____ a proton (H^+) to another substance.
A simpler way to say this is an **acid** is a proton donor (H^+).
- 5.) A **base** is a chemical species that _____ a proton (H^+) from another substance.
A simpler way to say this is a **base** is a proton (H^+) acceptor.

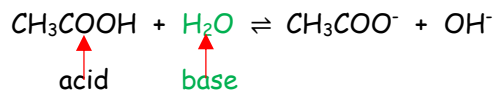
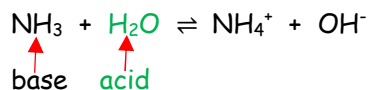
Ex. -

We can see the ammonium received and H^+ and the water lost an H^+ . So, overall we have



- Definitions for acids:

- 1.) Monoprotic acid -
- 2.) Diprotic acid -
- 3.) Triprotic acid -
- 4.) Polyprotic acid -
- 5.) Amphiprotic - a substance that can act like a **proton donor** (Bronsted-Lowry acid) or a **proton acceptor** (Bronsted-Lowry base). Ex. - H_2O , $H_2PO_4^-$, HS^- , HCO_3^{2-}



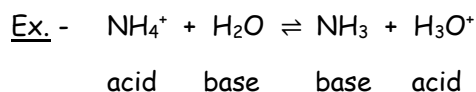
A substance is AMPHIPROTIC if

- 1.)
- 2.)

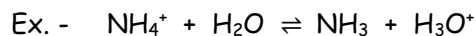
- 6.) Conjugate acid-base pair - also known as a conjugate pair is two chemicals which **differ by one proton**. Ex. - $NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$

- conjugate pairs

- 7.) Conjugate acid - the member of the above-mentioned pair that **contains the extra proton**.



- 8.) Conjugate base - the member of the above-mentioned pair that **lacks the extra proton**.

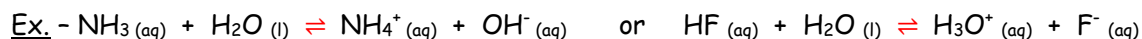


- Not all acid base reactions go to equilibrium.
 1. Reactions with **strong acids** and **bases** go to _____.
 2. Reactions with **weak acids** and **bases** go to _____.

- Strong acids/bases **fully ionize** (fully dissociate) in solution.



- Weak acids/bases **don't fully ionize** (fully dissociate) in solution.



***Note- **dilute** and **concentrated** refer to *molarity of solution* whereas **strong** and **weak** refer to *percent of ionization*. BE CAREFUL with language in chemistry!

Handout - [Relative Strengths of Acids and Bases](#)

- _____ is an expression describing how all strong acids (six acids on top of chart) are 100% dissociated in aqueous solution and are thus **equivalent** to solutions of $\text{H}_3\text{O}^+_{(aq)}$, and strong bases (two on bottom right) are also 100% dissociated in aqueous solutions rendering them equivalent to solutions of OH^- .
- Because of this condition of strong acids/bases, the central portion of the table from H_3O^+ down to H_2O shows the reactions which can actually occur. The top six and bottom two reactions are said to be levelled to produce H_3O^+ for acids and OH^- for bases.

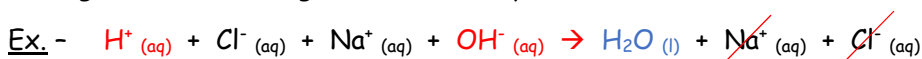
Equilibrium Constants for Water, Acids, and Bases

- Definitions:

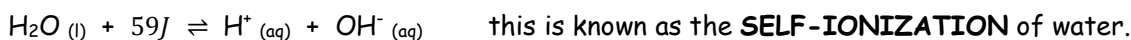
- 6.) A neutral solution is a solution when $[H_3O^+] = [OH^-]$
7.) An acid solution is a solution when $[H_3O^+] > [OH^-]$
8.) A base solution is a solution when $[H_3O^+] < [OH^-]$

Strong Acids and Bases

- Remember not all acid base reactions go to equilibrium. Reactions with **strong acids** and **bases** go to completion.
- When a strong acid and a strong base react, they neutralize and make water.



For a strong acid/base reaction we'd get this:



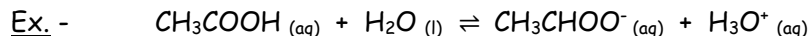
- An equilibrium expression can be written for this reaction: $K_w = [H^+][OH^-]$ (water is a liquid).
- The more common way to write this expression is $K_w = [H_3O^+][OH^-]$
- This K_w value has been measured at 1.00×10^{-14} at 25°C.

Ex. - What is the $[H_3O^+]$ and $[OH^-]$ in a 0.0010 M HCl (aq)?

Answer - *remember HCl is a strong acid, so it will completely dissociate* $[H_3O^+] = [HCl]$

Weak Acids and Bases

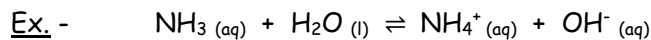
- Remember reactions with **weak acids** and **bases** go to _____.
- The ionization of weak acids produces the conjugate base and acid of the reactant acid and water.



The equilibrium expression for this reaction is $K_a = \frac{[CH_3COO^-][OH^-]}{[CH_3COOH]}$ $K_a = 1.76 \times 10^{-5}$

This value of K_a is called the _____

- The ionization of weak bases produces the conjugate acid and base of the reactant base and water.



The equilibrium expression for this reaction is $K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$ $K_b = 1.79 \times 10^{-5}$

This value of K_b is called the **Base Ionization Constant**

- The chart of Relative Strengths of Acids and Bases does not list K_b values. However, there is a relationship between K_a and K_b .

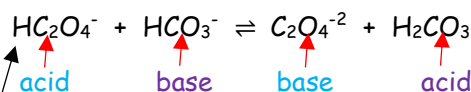
For a **conjugate pair** =

Ex. 1 - To find the K_a value for HSO_3^- look down the **left-hand** side of the table (acid side) and find HSO_3^- .

Ex. 2 - To find the K_b value for HSO_3^- look down the **right-hand** side of the table (base side) and find H_2SO_3^- .

Relative Strengths of Acids and Bases

- If solutions of HC_2O_4^- and HCO_3^- are mixed we need to be wary. Both are amphiprotic!
- HC_2O_4^- is higher on the chart meaning it is the stronger acid. This forces the HCO_3^- to be the base. The following equilibrium is set up:



- Notice how there are competing acids in equilibrium, HC_2O_4^- and H_2CO_3 , each of which can donate a proton. Which direction is the predominant reaction????
- Compare the acids on the chart. **Higher will drive reaction away from itself (forward direction, products favoured).**

- This can be shown mathematically as well.

$$K_{eq} = \frac{K_a(\text{reactant acid})}{K_a(\text{product acid})}$$

$$K_{eq} = \frac{6.4 \times 10^{-5}}{4.3 \times 10^{-7}} = \frac{[\text{products}]}{[\text{reactants}]}$$

$$K_{eq} = 1.5 \times 10^2$$

$K_{eq} > 1$ means products favoured.

Summary →

So $\frac{[\text{products}]}{[\text{reactants}]} = K_{eq} = \frac{K_a(\text{reactant acid})}{K_a(\text{product acid})}$

Ex. - What is the predominant reaction between dihydrogen citrate ion and hydrogen sulphate ion?

Answer -

pH and pOH

- Definitions:

$$1.) pH = -\log[H_3O^+]$$

$$2.) pOH = -\log[OH^-]$$

- pH is a term used to describe the strength of acids and pOH is used for bases. The "p" of pH comes from the German for "power" or "potency" and the "H" of pH is of course for the proton.
- This naming makes sense then as the pH scale is a logarithmic scale of the [protons] or [hydroxide] depending on whether you are talking about acids or bases respectively.
- Logarithm or simply logs are the amounts of **base 10** needed to give the value you seek. More simply it's how many 10's multiplied together give you your requested value.

Ex. - log of 1000? $10 \times 10 \times 10 = 1000$ so . . . the log of 1000 is 3 (3 tens make 1000)

- The reverse of addition is subtraction. The reverse of a log is an antilog.
- So, if $\log(1000) = 3$ then $\text{antilog}(3) = 1000$

****Reminder**** when bases are the same in math you can add exponents.

Ex. - $\log(10^x \times 10^y) = \log(10^x) + \log(10^y)$

- How do we convert a $[H_3O^+]$ and $[OH^-]$ to pH and pOH?

Ex. 1 - $[H_3O^+] = 3.94 \times 10^{-4} M$, what is the pH?

Ex. 2 - $[OH^-] = 9.51 \times 10^{-12} M$, what is the pOH?

*****Note***** In pH/pOH calculations only the values **after the decimal** are accurate!!! Be careful!!!!!!

- How do we convert a pH and pOH to $[H_3O^+]$ and $[OH^-]$?

(Hint - do the reverse math!) $[H_3O^+] = \text{antilog}(-pH)$

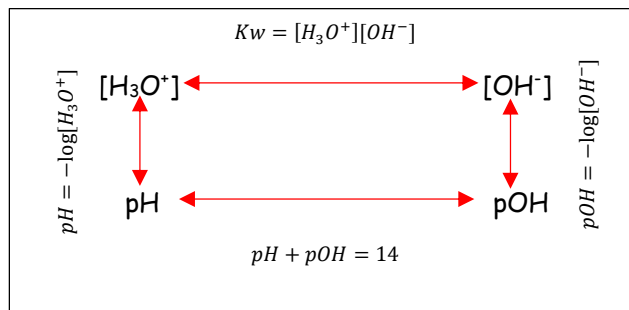
Ex. 1 - If pH is 3.405, then what is the $[H_3O^+]$?

Ex. 2 - If pOH is 11.022, then what is the $[OH^-]$?

- There clearly must be a connection between pH and pOH as well as $[H_3O^+]$ and $[OH^-]$. Yes!! But first . . .

$pH + pOH = 14$

- Any of the four corner values below can be calculated given any of the other values using the relationship.



Ex. 1 - pH is 11.59, what is pOH ?

Ex. 2 - pH is 4.59, what is $[OH^-]$?

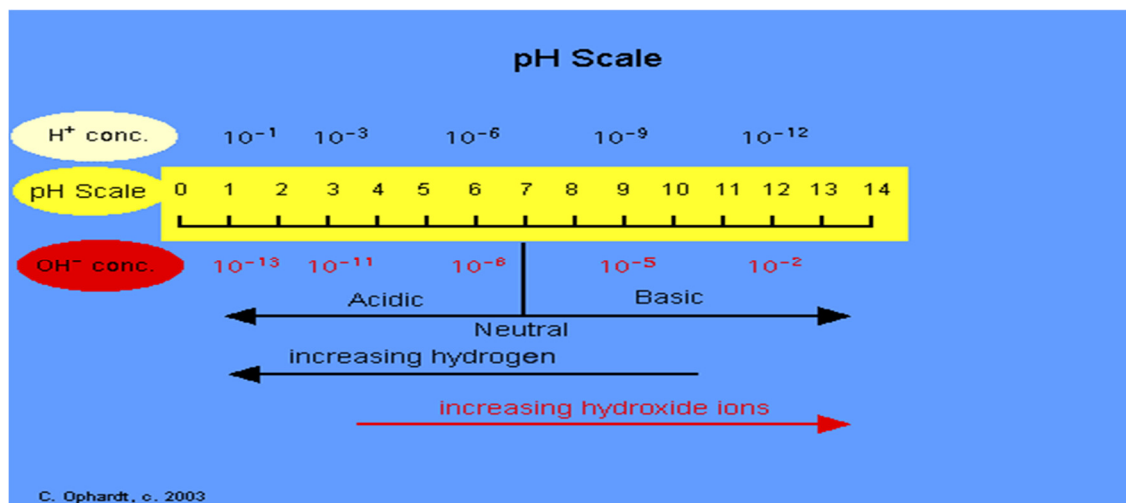
Note

$$pK_w = -\log(K_w) = -\log(1.00 \times 10^{-14}) \quad \text{so} \quad pK_w = 14.000, \text{ at } 25^\circ\text{C}$$

$$pK_w = pH + pOH$$

The pH Scale

- The levels of $[H_3O^+]$ and $[OH^-]$ are very small in solution. The pH scale is a logarithmic scale that is designed to be an easier way of conveying the $[H_3O^+]$ and $[OH^-]$.



- Notice that neutral (neither acid or base) is in the middle at $7 = pH = pOH$.
- pH scale is dependent on $K_w = K_{eq}$. As such this scale **IS** temperature dependent. If you raise the temperature the $[H_3O^+]$ will increase. This is a decrease (more acidic) on the pH scale.

Mixing Strong Acids and Bases

- If you mix 30.0 mL 0.12 M HCl with 40.0 mL 0.10 NaOH what will be the resulting pH?

1.) Equations -

2.) Dilution Calculation -

3.) Calculate excess -

4.) Calculate pH -

- Ex. - How many moles of HCl (g) must be added to 40.0 mL 0.180 M NaOH to make a solution with a $pH = 12.500$.

Answer - this is essentially the same question in **reverse!**

1.) Equations -

2.) Calculate pOH -

3.) Calculate excess -

4.) Moles of acid -

Hydrolysis

- Hydrolysis comes from the words "hydro" (water) "lysis" (to cut). Hydrolysis reactions typically are the "cutting out" of water or the removal of a water from a chemical.
- In Chemistry 12 we will be looking at the **hydrolysis of salts**. This is a reaction between water and the cation or anion of a salt to produce an acidic or basic solution.

***Note

In this section we will be focused on the reaction between ions and water and **NOT** the reaction that **may occur between the ions themselves**.

- **The conjugates of strong acids/bases are ALWAYS spectator ions.**

Ex. - $\text{NaOH} \rightarrow \text{Na}^+ + \text{OH}^-$ The Na^+ cannot reform NaOH (reverse reaction) as NaOH is a strong base and fully ionizes.

- To solve problems involving hydrolysis follow the following steps:

1.) Dissociate the salt $\text{NaNO}_2 \rightarrow \text{Na}^+ + \text{NO}_2^-$

2.) Discard the spectator ions $\text{Na}^+ + \text{NO}_2^-$

- All cations from alkali and alkaline earth $\rightarrow \text{Li}^+, \text{Na}^+, \text{K}^+, \text{Rb}^+, \text{Cs}^+, \text{Be}^{+2}, \text{Mg}^{+2}, \text{Ca}^{+2}, \text{Sr}^{+2}, \text{Ba}^{+2}$

- Anions of the top 6 rows on the K_a chart (strong acids) $\rightarrow \text{ClO}_4^-, \text{Cl}^-, \text{Br}^-, \text{I}^-, \text{NO}_3^-, \text{SO}_4^-$

3.) Determine which side of the acid/base chart your ion is on. **base**

4.) Produce the hydrolysis reaction $\text{NO}_2^- + \text{H}_2\text{O} \rightleftharpoons \text{HNO}_2 + \text{OH}^-$

Ex. - NH_4Cl

Answer -

Ex. 2 - NaHC_2O_4

Answer -

Ex. 3 - NH_4NO_2

Answer -

Calculations Involving K_a and K_b

- Weak acids (HA) and bases (A^-) are out into water, they only partially ionize.
- The smaller the K_a the smaller the amount ionizing into H_3O^+ for acids and OH^- for bases.
- Since the acid/base does not ionize until contact with water, an old friend is best for solving. ("old friend" means ICE box!)
- When solving questions regarding weak acids/bases there are **three major pieces of information**:
 - 1.) $[HA]$ or $[Hb]$, depending on which type of solution (the original acid/base concentration).
 - 2.) K_a or K_b , depending again on the solution.
 - 3.) $[H_3O^+]$ / $[OH^-]$ or pH/POH, depending on solution.
- There are also **three types of problems** that you could be asked to solve:

Acids

- 1.) given $[HA]$ and K_a , find $[H_3O^+]$ (or pH)
- 2.) given $[HA]$ and $[H_3O^+]$ (or pH), find K_a
- 3.) given $[H_3O^+]$ (or pH) and K_a , find $[HA]$

Bases

- given $[A^-]$ and K_a (you calculate K_b), find $[OH^-]$ (or pOH)
- given $[A^-]$ and $[OH^-]$ (or pOH), find K_b
- given $[OH^-]$ (or pOH) and K_a (you calculate K_b), find $[A^-]$

- Ex. 1 - If $K_a = 1.8 \times 10^{-5}$ for CH_3COOH , what is the pH of a 0.500 M solution of CH_3COOH ?

Answer - 1.) List known and asked for to allow you to figure out the procedure for solving.

2.) Write acid ionization equation.



I				
C				
E				

3.) Write the K_a expression and **write your assumption**. You must write out your assumption.

4.) Solve the equation using the assumption.

Ex. 2 - If $pH = 1.70$ for a 0.100 M solution of an unknown weak acid, what is the K_a ?

Answer - 1.) List known and asked for to allow you to figure out the procedure for solving.

2.) Write acid ionization equation.

I				
C				
E				

3a.) Use pH to solve for $[H_3O^+]$.

b.) Write the K_a expression and **check your assumption**.

4.) Solve the equation.

Ex. 3 - How many grams of NH_4Cl will make 1.50 L of solution with a $pH = 4.75$?

Answer - 1.) List known and asked for to allow you to figure out the procedure for solving.

2.) Write acid ionization equation.

I				
C				
E				

3a.) Use pH to solve for $[H_3O^+]$.

b.) Write the K_a expression and **write your assumption**. You must write out your assumption.

4.) Solve the equation using the assumption.

Ex. 4 - What is the pH of a 0.10 M solution of NaCN?

Answer - 1.) List known and asked for to allow you to figure out the procedure for solving.

2.) Write base ionization equation.

I				
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3a.) Use K_a to solve for K_b .

b.) Write the K_b expression and **write your assumption**. You must write out your assumption.

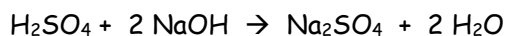
4.) Solve the equation using the assumption.

Acid/Base Titrations

- Review - titration is a process in which a **measured amount** of a solution is reacted with a **known volume** of another solution, in an effort to discern an unknown concentration, until a desired equivalence point (stoichiometric point) is reached.
- The point in a titration reaction when the acid/base neutralize each other is known as the "end point".
- The "titrant" is the chemical of known concentration that is being used to discover the unknown chemicals concentration. The equivalence point is the point when the moles of reactant chemicals are equal to the ratio of the balanced equation.
- Titration questions involve **5 main parts**. You will need to solve for the missing part in a given question.
 - 1.) concentration of acid
 - 2.) volume of acid
 - 3.) concentration of base
 - 4.) volume of base
 - 5.) acid/base mole ratio

- There are always **three parts** to the question:
 1. Calculate the moles of the chemical given.
 2. Use the acid/base mole ratio (hello, Chem 11!!) to calculate the moles of the second chemical.
 - 3.) Use the moles of the second substance to solve for the volume or concentration (whichever is needed).

Ex. 1 - 23.10 mL of 0.2055 M NaOH was added to 25.00 mL H₂SO₄ to reach the equivalence point. What is the [H₂SO₄] in the following reaction:



Answer - Solve for moles

Solve for second chemical

Solve for [second chemical]

Ex. 2 - 25.00 mL of 0.1255 M NaOH was added to 38.74 mL of 0.02700 H₄P₂O₇ to reach the equivalence point. How many of the polyprotic protons are removed from the H₄P₂O₇, and what is the balanced equation?

Answer - Solve for moles

Solve for second chemical

Solve for difference

Ex. 3 - An 3.4786 g impure sample of NaHSO₄ is added to 250.0 mL of water. 0.02500 L sample was titrated using 0.02677 L of 0.09974 M NaOH. What is the percent purity of the acid?

Answer - Solve for moles

Solve for second chemical

Solve for molarity

Solve for theoretical

Solve percent purity

Indicators

- Indicators are **weak** organic acids and bases that have different colours between the conjugate acid and base.
- The names of these indicators are long and complex. As such, they are often abbreviated by "HIn" as in H for acid with In for indicator.
- Ex. - **HIn** + H₂O ↔ **In**⁻ + H₃O⁺
- When the indicator is added to an acid the excess H₃O⁺ shifts the equilibrium left (yellow colour). What this tells us is the indicator is in its conjugate **acid form** when in **highly acidic solutions**.
- When the indicator is added to a base the [H₃O⁺] is very low so the equilibrium shifts right (red colour). What this tells us is the indicator is in its conjugate **base form** when in **highly basic solutions**.
- If a base is added to an acid solution then at some point the [HIn] = [In⁻]. At this point the yellow and red molecules will be equal and the colour seen is orange. This is called **end point** or **transition point**.
- So at the endpoint →
$$K_a = \frac{[H_3O^+][In^-]}{[HIn]} = [H_3O^+] \quad \text{since } [HIn] = [In^-]$$

- So at the endpoint →
$$pK_a = -\log K_a \quad K_a = [H_3O^+] \quad pK_a = pH$$

- Ex. - Methyl orange changes colour in a pH range from 3.2 – 4.4. What is the K_a of methyl orange?

Answer -

- Ex. - Alizarin yellow R changes colour from yellow to red at pH = 11.0. If Aliz⁻ ion is red, what colour is Alizarin Yellow R in 1.0 × 10⁻⁵ M NaOH?

Answer -

- A universal indicator is an indicator that changes a variety of colours depending on the pH.

More on Titrations

- To do a titration you need a _____, that is, a solution of known concentration. This standard solution is used to titrate unknown solutions.
- To make a standardized solution you need to use a chemical that is both stable (won't absorb water or react with air) and pure. This substance is commonly referred to as a _____.
- The most common acidic primary standards are:

1. potassium hydrogen phthalate, $\text{KHC}_8\text{H}_4\text{O}_4$ (monoprotic acid). **Phenolphthalein** is an excellent indicator
2. oxalic acid dihydrate, $\text{H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$ (diprotic acid). **Phenolphthalein** is the preferred indicator.

The most common basic primary standard is:

sodium bicarbonate, Na_2CO_3 (diprotic base). **Methyl orange** is a good choice for indicator.

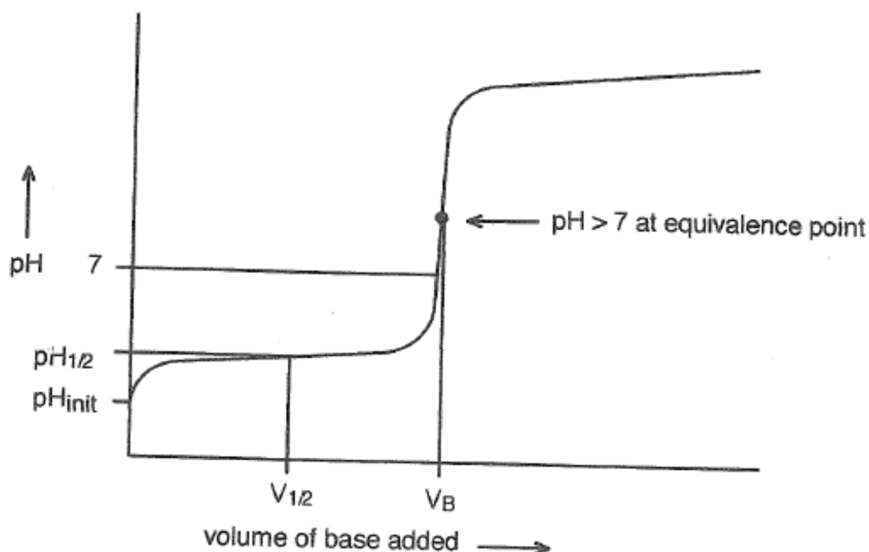
Solving Titrations

- This section will help you predict approximate $\text{p}K_a$ values to adequately choose indicators for titrations.
- **Strong Acid with a Strong Base** titrations ← understand that the acid is being titrated by the base!!

1. Typical shape of titration graph

- A strong acid and base will result in a **neutral solution**. As such, an indicator that changes pH around 7 ($\text{p}K_a = 7$) is best.

2. **Weak Acid with a Strong Base** titrations



- The weak acid and strong base generally result in basic solution.

** The value of K_a can be found using the following equation $K_a = [\text{H}_3\text{O}^+]_{1/2}$ where $[\text{H}_3\text{O}^+]_{1/2}$ is from $\text{pH}_{1/2}$

** The molarity of a weak acid **at equilibrium**, $[\text{HA}]$, can be found from the equation

$$[\text{HA}]_{\text{equ}} = \frac{[\text{H}_3\text{O}^+]^2}{K_a} \quad \text{where } [\text{H}_3\text{O}^+] \text{ is from } \text{pH}_{\text{init}}$$

Ex. - The following information was recorded from the titration of the weak acid furoic acid (C_4H_3O)COOH with NaOH.

Furoic acid volume	25.0 mL
NaOH volume at equivalence point	28.8 mL
Initial pH of acid	2.021
pH at halfway point (14.4 mL)	3.170

a.) Calculate the K_a of the acid.

Answer -

b.) Calculate the $[HA]_{init}$.

Answer -

c.) Calculate the $[NaOH]$.

Answer -

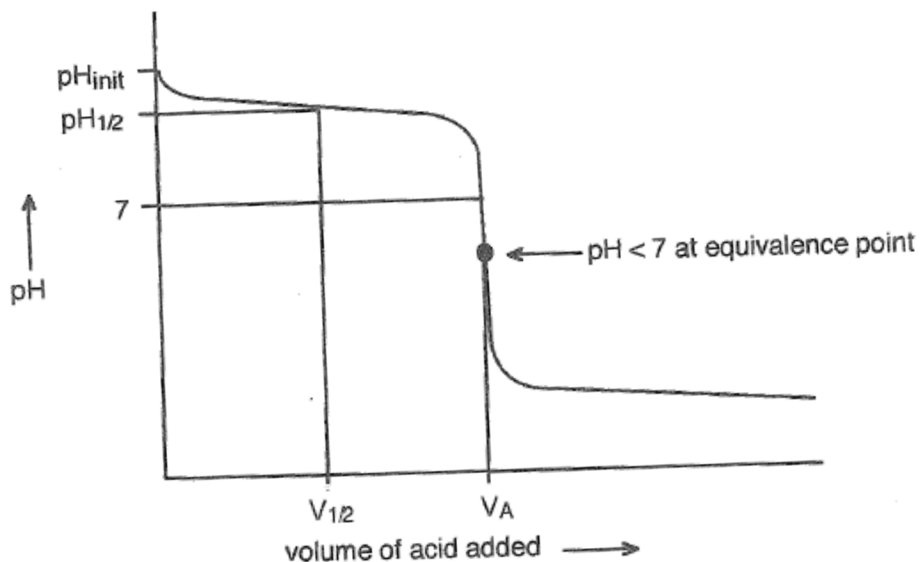
d.) Is the titrate solution acid, basic, or neutral at the equivalence point?

Answer -

e.) What is a good indicator for this titration?

Answer -

3. Weak Base with a Strong Acid titrations



- The weak base with a strong acid generally results in an acid solution.

*** Note - the above graph gives pH values BUT pOH values are required for all calculations. So, the first step will always be to convert to pOH values.

** The value of K_b can be found using the following equation $K_b = [OH^-]_{1/2}$ where $[OH^-]_{1/2}$ is from $pH_{1/2}$

** The molarity of a weak base **at equilibrium**, $[A^-]$, can be found from the equation

$$[A^-]_{equ} = \frac{[OH^-]^2}{K_b} \quad \text{where } [OH^-] \text{ is from } pOH_{init}$$

Ex. - The following information was recorded from the titration of the weak monoprotic base ethylamine with HCl.

ethylamine volume	25.00 mL
HCl volume at equivalence point	19.22 mL
Initial pH of base	11.855
pH at halfway point (9.60 mL)	10.807

a.) Calculate the K_b of the base.

Answer -

b.) Calculate the $[A^-]_{init}$.

Answer -

c.) Calculate the $[HCl]$.

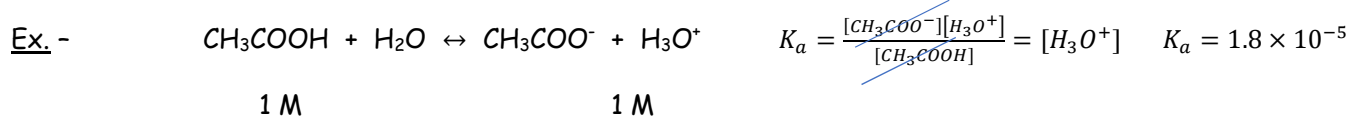
Answer -

d.) What is a good indicator for this titration?

Answer -

Buffers

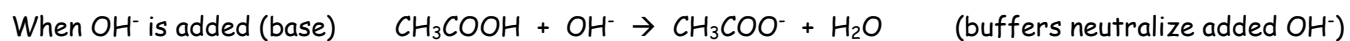
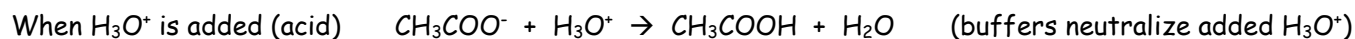
- A buffer is a solution consisting of a substantial amount of _____ and its conjugate _____.
- To make a buffer you need to mix both the acid and the base in water.



Conclusion

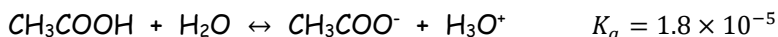
If **equal** amounts of weak acid and its conjugate base are mixed with water, the pH will equal the pK_a of the weak acid.

- Typically, when acids and bases are diluted the pH's are changed. However, if a buffer is diluted the pH _____.
- The purpose of creating a buffer solution is to **minimize** pH changes when an acid or base is added to the solution.
- Adding 10% more acid to a solution would typically result in a pH change of 6. In a typical buffer solution, the pH would **only** change by 0.08!
- This "buffering effect" is caused by the ratio _____, which doesn't change much when small amounts of acid or base are added.



- Acid buffer example

Mixing 1.0 mol CH₃COOH (acetic acid) and 1.0 mol CH₃COO⁻ (sodium acetate) in 1.0 L of water.



This solution "buffers" the pH around $pH = pK_a = -\log[1.8 \times 10^{-5}] \quad pH = 4.74$

- Base buffer example

Mixing 1.0 mol NH₃ (ammonia) and 1.0 mol NH₄NO₃ (ammonium nitrate) in 1.0 L of water.

Since NH₄NO₃ (s) → NH₄⁺ + NO₃⁻ and NO₃⁻ is a spectator, the following equilibrium is set up

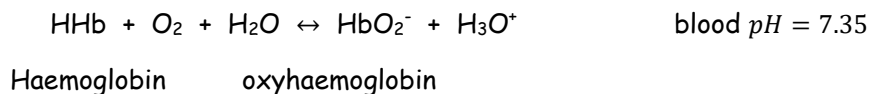


This solution "buffers" the pH around $pH = pK_a = -\log[5.6 \times 10^{-10}] \quad pH = 9.25$

- The importance of the above shows that the buffer solution needs to be set up near the pH you are trying to keep. Different buffers will buffer at different pH.
- End note → when a **weak acid or base** is titrated, a buffer solution will be established in the middle of the titration curve (pH_{1/2})!

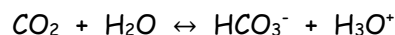
Where are buffers used?

1.) Haemoglobin of the blood.



- If you eat a tomato or lemon, without a buffer in your blood, the increased acid level would drive the above reaction to the left (Le Chatelier), decreasing the HbO_2^- level, resulting in passing out (death).

2.) Lakes naturally have a buffer system from the minerals in the surrounding rocks and soil.



- Business and automobiles produce SO_2 , SO_3 , NO_2 , NO_3 or more commonly known as SO_x and NO_x . These dissolve into precipitation and produce H_2SO_4 , H_2SO_3 , HNO_2 , HNO_3 (acids).

Metal and Non-metal Oxides

- As seen above, lakes become acidified from **covalent oxides (non-metal oxides) dissolve into water (precipitation) to form acids.**
Ex. - $\text{SO}_3 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_4$ (sulphuric acid)
Ex. - $\text{SO}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_3$ (sulphurous acid)
Ex. - $\text{CO}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{CO}_3$ (carbonic acid) } Need to know grade 12
- When **ionic metal oxides** are added to water they form basic solutions. For grade 12 we will only concern ourselves with the alkali and alkaline earth metal oxides.

