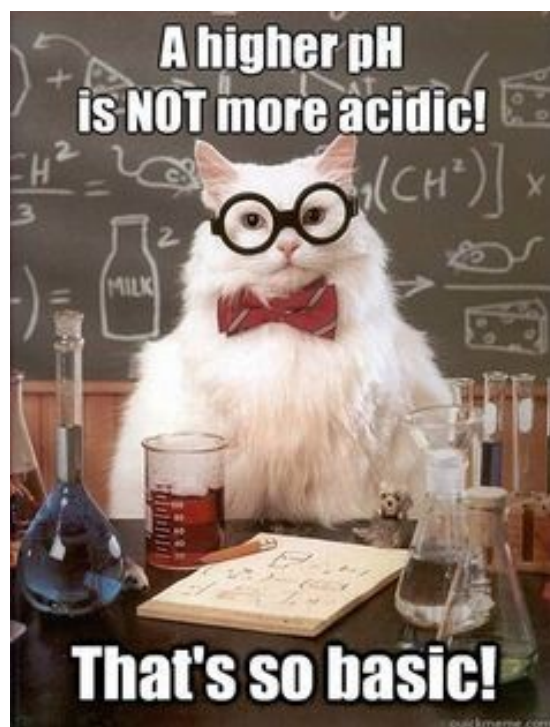
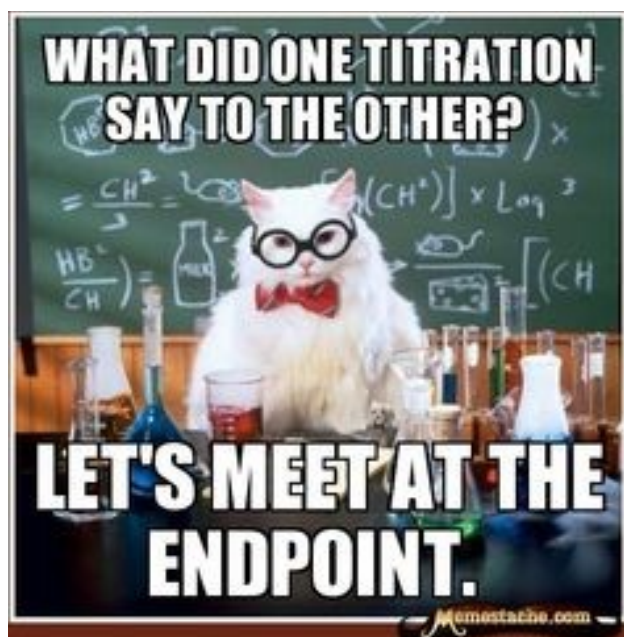




Kennedy
Baptist College

Acids and Bases



(Discover and save creative ideas n.d.)

| WEEK | TOPIC | UNIT CONTENT | ASSESSMENT |
|--------|-----------------|--|------------|
| 5 – 10 | Acids and Bases | <ul style="list-style-type: none"> models and theories are contested and refined or replaced when new evidence challenges them, or when a new model or theory has greater explanatory scope. Davy initially proposed that acids were substances that contained replaceable hydrogen (hydrogen that could be partly or totally replaced by metals) and bases were substances that reacted with acids to form salts and water. Arrhenius model, which includes only soluble acids and bases, identified acids as substances which produce hydrogen ions in solution and bases as substances which produce hydroxide ions in solution the Brønsted-Lowry model describes acid-base behaviour in terms of proton donors and proton acceptors. This approach includes a wider range of substances and can be more broadly applied. acids are substances that can act as proton (hydrogen ion) donors and can be classified as monoprotic or polyprotic, depending on the number of protons available for donation the relationship between acids and bases in equilibrium systems can be explained using the Brønsted-Lowry model and represented using chemical equations that illustrate the transfer of protons between conjugate acid-base pairs the strength of acids is explained by the degree of ionisation at equilibrium in aqueous solution which can be represented by chemical equations and acidity constants (K_a) the hydrolysis of salts of weak acids and weak bases can be represented using equations; the Brønsted-Lowry model can be applied to explain the acidic, basic and neutral nature of salts derived from bases and monoprotic and polyprotic acids | |

| | | | |
|--|--|---|----------------------------------|
| | | <ul style="list-style-type: none"> • water is a weak electrolyte; the self-ionisation of water is represented by $K_w = [H^+][OH^-]$ where $K_w = 1.0 \times 10^{-14}$ at 25 °C • K_w can be used to calculate the concentration of hydrogen ions or hydroxide ions in solutions of strong acids or bases • the pH scale is a logarithmic scale and the pH of a solution can be calculated from the concentration of hydrogen ions using the relationship $pH = -\log_{10} [H^+]$ • acid-base indicators are weak acids, or weak bases, in which the acidic form is a different colour from the basic form • buffer solutions are conjugate in nature and resist changes in pH when small amounts of strong acid or base are added to the solution; buffering capacity can be explained qualitatively; Le Châtelier's Principle can be applied to predict how buffers respond to the addition of hydrogen ions and hydroxide ions • volumetric analysis methods involving acid-base reactions rely on the identification of an equivalence point by measuring the associated change in pH, using appropriate acid-base indicators or pH meters, to reveal an observable end point • data obtained from acid-base titrations can be used to calculate the masses of substances and concentrations and volumes of solutions involved | <p>Practical 2</p> <p>Test 2</p> |
|--|--|---|----------------------------------|

Chemical and Physical Properties of Acids and Bases

Acidic solutions generally:

- conduct an electric current
- turn litmus red
- taste sour

Basic solutions generally:

- also conduct an electric current
- turn litmus blue
- taste bitter
- have a slippery feel

Give the net ionic equation and observations for the reactions below illustrating some of the **chemical properties** of acids and bases.

1. Acidic solutions dissolve most metals forming a salt and hydrogen gas. Au, Ag, Pt and Cu are unaffected by dilute acids.



(Eg: Some hydrochloric acid solution is poured onto a strip of nickel metal)

2. $\text{ACID} + \text{METAL HYDROXIDE} \rightarrow \text{SALT} + \text{H}_2\text{O}$

(Eg: Calcium hydroxide solution is added to a solution of nitric acid)

3. $\text{ACID} + \text{METAL OXIDE} \rightarrow \text{SALT} + \text{H}_2\text{O}$

(Eg: Some solid copper (II) oxide is added to a solution of sulphuric acid)



(Eg: Hydrochloric acid solution is added to nickel carbonate solid)



(Eg: Ethanoic acid solution is added to solid zinc hydrogencarbonate)



(Eg: Hydrochloric acid solution and sodium sulfite powder are mixed)



(Eg: A potassium hydroxide solution is mixed with an ammonium chloride solution)



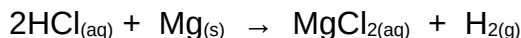
(Eg: Sulfur dioxide gas is bubbled through a potassium hydroxide solution)

Theories of Acids and Bases

Davy Theory

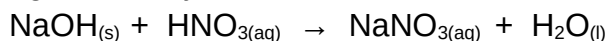
Acids: Have replaceable H (hydrogen could be partly or totally replaced by metals).

eg hydrochloric acid:



Bases: Reacted with acids to form salt and water.

eg sodium hydroxide:



Arrhenius Theory

Acids: Have H in their formula and produce hydrogen ions (H^+) when dissolved in water.

eg hydrochloric acid: $\text{HCl}_{(\text{g})} \rightarrow \text{H}^+_{(\text{aq})} + \text{Cl}^-_{(\text{aq})}$

eg ethanoic acid: $\text{CH}_3\text{COOH}_{(\text{s})} \rightarrow \text{H}^+_{(\text{aq})} + \text{CH}_3\text{COO}^-_{(\text{aq})}$

Bases: Have OH in their formula and produce hydroxide ions (OH^-) when dissolved in water.

eg sodium hydroxide: $\text{NaOH}_{(\text{s})} \rightarrow \text{Na}^+_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$

eg strontium hydroxide: $\text{Sr}(\text{OH})_{2(\text{s})} \rightarrow \text{Sr}^{2+}_{(\text{aq})} + 2\text{OH}^-_{(\text{aq})}$

Neutralisation: an acid plus a base produces a neutral solution (ie $\text{H}^+_{(\text{aq})} + \text{OH}^-_{(\text{aq})} \rightarrow \text{H}_2\text{O}_{(\text{l})}$)

Problems with the theory:

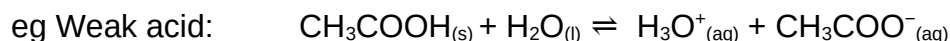
- Some bases produce OH^- ions in solution yet do not have OH in their formula (eg NH_3 and CO_3^{2-}).
- Restricted to aqueous solutions.
- Not all salts are neutral.
- Does not allow for the existence of hydronium ions (H_3O^+).

Brønsted-Lowry Theory

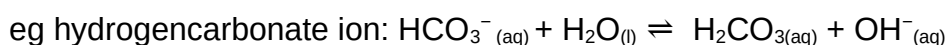
Acids are proton (H^+) donors.

eg Strong acid:

HCl is donating a proton and acting as an acid. The H_2O is accepting a proton and acting as a base.



Bases are proton (H^+) acceptors.



HCO_3^- (aq) is accepting a proton and acting as a base. The H_2O is donating a proton and acting as an acid.

Neutralisation: reaction between a proton donor and a proton acceptor.

eg carbonate ion plus water: CO_3^{2-} (aq) + H_2O (l) \rightleftharpoons HCO_3^- (aq) + OH^- (aq)

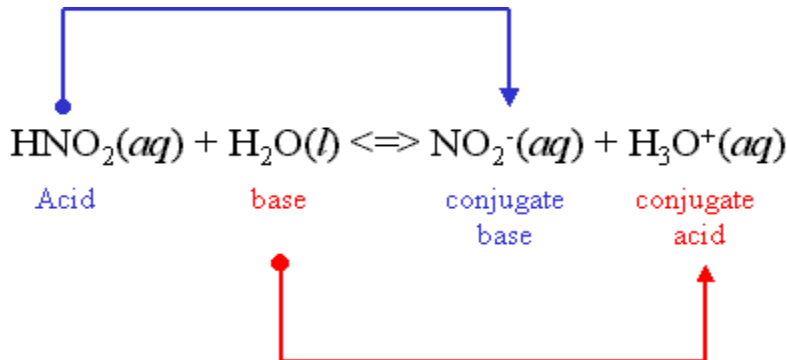
Conjugate acid- base pairs

Brønsted-Lowry theory recognises the reversibility of proton transfer reactions.

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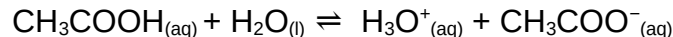
- This means both the forward and the reverse reactions are acid-base reactions.
eg NH_3 (g) + H_2O (l) \rightleftharpoons NH_4^+ (aq) + OH^- (aq) H_2O (l) + H_2O (l) \rightleftharpoons H_3O^+ (aq) + OH^- (aq)
- The acid and the base it forms (or the base and the acid it forms) are called a conjugate acid-base pair.
- In any conjugate acid- base pair the formula of the acid is always greater by the equivalent of H^+ .
- The stronger the acid the weaker its conjugate base.



(Blaber, 2000)

Acidity constant (K_a) and acid strength

- The larger the K_a the greater the tendency of the acid to donate a proton to water ie the stronger the acid is the greater the degree of its ionisation. (Strong acids fully ionise in water.)
- Eg the proton transfer equation for ethanoic acid is:



its K_a expression is: $K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]}$

The K_a value at 25°C is 1.8×10^{-5} . This indicates the reaction only proceeds to a very limited extent and only about 1% of ethanoic acid molecules are ionised.

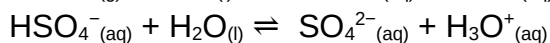
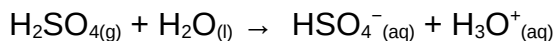
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- The K_a value also indicates the strength of the acid's conjugate base to accept a proton back from water so, if K_a is large (strong acid) its conjugate base is weak. As K_a is progressively smaller, the acid is progressively weaker and the conjugate base becomes progressively stronger. (**The stronger the acid is the weaker its conjugate-base.**)
- Acid-base reactions tend to occur in the direction in which the stronger acid and stronger base react to form a weaker acid and weaker base.

Eg

Polyprotic Acids

While weak monoprotic acids produce fewer hydrogen ions in solution than strong acids, they react to the same extent with bases, providing that the concentration is the same.

Explain (using equations) why one mole of H_2SO_4 yields between one and two moles of H_3O^+ ions in solution.



This process of becoming ions over multiple steps is called **successive ionisation**.

Polyprotic acids are acids which contain two or more acidic protons. Use equations to show the successive ionisation of phosphoric acid.

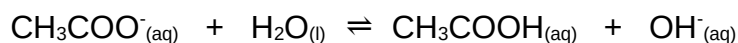
Acid-Base Properties of Salts

Eg 1

An ammonium chloride solution ($\text{NH}_4\text{Cl}_{(\text{aq})}$) is acidic. The ammonium ion (NH_4^+) is a weak acid in water as it acts as a proton donor and hydrolyses to produce hydronium ions (H_3O^+). The chloride ion (Cl^-) is neutral.

Eg2

A sodium ethanoate solution ($\text{NaCH}_3\text{COO}_{(\text{aq})}$) is basic. The ethanoate ion (CH_3COO^-) is a weak base in water as it acts as a proton acceptor and hydrolyses to produce hydroxide ions (OH^-). The sodium ion (Na^+) is neutral.



Summary

The table below classifies common ions according to whether they cause aqueous solutions to be acidic, basic or neutral.

| Type of Ion | Acidic | Basic | Neutral |
|-------------|--|---|--|
| Anions | Anions produced in first stage of ionisation of polyprotic acids: HSO_4^- , H_2PO_4^- | Anions from weak acids: F^- , CO_3^{2-} , HS^- , S^{2-} , HPO_4^{2-} , PO_4^{3-} , HCO_3^- , CH_3COO^- , SO_4^{2-} | Anions from strong acids: Cl^- , Br^- , I^- , NO_3^- |
| Cations | NH_4^+ , Fe^{3+} , Al^{3+} , Cr^{3+} Sn^{4+} | Not common | Li^+ , Na^+ , K^+ , Mg^{2+} , Ca^{2+} , Ba^{2+} |

Predict whether aqueous solutions of the following salts will be acidic, basic or neutral and justify your choice by writing suitable chemical equations.

- Ammonium nitrate
- Calcium hydrogencarbonate
- Potassium ethanoate
- Magnesium bromide
- Calcium hydrogensulfate

Self- Ionisation of Water, Kw and pH

At 25 °C, the equilibrium constant for the ionisation of water (K_w) equals 1.00×10^{-14} . It is given by:

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

In water or any solution where H_3O^+ and OH^- have the same concentration, the solution is said to be neutral. $[\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.00 \times 10^{-7} \text{ mol L}^{-1}$ at 25 °C.

From this, an acidic solution must have $[\text{H}_3\text{O}^+] > [\text{OH}^-]$ and a basic solution must have $[\text{H}_3\text{O}^+] < [\text{OH}^-]$. Also, since $[\text{H}_3\text{O}^+] \times [\text{OH}^-] = 1.00 \times 10^{-14}$ for all solutions at 25 °C, the concentration of one of these ions can be calculated given the concentration of the other.

Eg. Calculate the concentration of hydrogen ions in a 0.001 mol L^{-1} NaOH solution.

Eg. Calculate the concentration of hydroxide and hydrogen ions in a solution made by dissolving 2.34 g of $\text{Ba}(\text{OH})_2$ solid in 543 mL of water.

pH

Eg. Calculate the pH of the following solutions:

1. A 0.002 mol L^{-1} solution of HCl.
2. A 0.007 mol L^{-1} solution of H_2SO_4 (assuming complete ionisation)
3. A 0.005 mol L^{-1} solution of NaOH.
4. Calculate the final pH of a solution made up by adding 5.67 g of HCl solid to 23.5 mL of a $3.96 \times 10^{-1} \text{ mol L}^{-1}$ $\text{Ba}(\text{OH})_2$ solution (assume no change in volume).
5. Calculate $[\text{H}^+]$ of HNO_3 solution with a pH of 1.38.
6. Calculate $[\text{OH}^-]$ and $[\text{H}_3\text{O}^+]$ of a $\text{Mg}(\text{OH})_2$ solution with a pH of 12.68.

pH can be measured using an indicator solution (a specific weak acid that changes colour over a range of pH), indicator paper or a pH meter (which is much more precise, if used correctly). A pH meter measures the voltage difference across two electrodes placed in the solution.

- The self-ionisation of water is an **endothermic process**. An increase in temperature will favour the endothermic (forward) reaction which would result in an increased equilibrium constant. (eg at 40 °C, $K_w = 2.916 \times 10^{-14}$).
- Similarly, at lower temperatures, the exothermic (reverse) reaction is favoured and so the equilibrium constant is reduced (eg at 10 °C, $K_w = 2.93 \times 10^{-15}$).

It is important to remember that even though pure water at temperatures above 40 °C has a pH less than 7, it is not acidic (this is because a neutral pH of 7 only applies at 25 °C). Pure water at, whatever temperature, is always neutral.

Complete Essential Chemistry Set 5

Buffers

Eg Mixing 2 mol of ethanoic acid, CH_3COOH , with 2 mol sodium ethanoate NaCH_3COOH and making the solution volume up to 1.00 L. This solution contains a high concentration of weak acid, CH_3COOH , and its conjugate weak base, CH_3COO^- (from NaCH_3COOH).

The equilibrium set up is given as: $\text{CH}_3\text{COOH}_{(\text{s})} + \text{H}_2\text{O}_{(\text{l})} \rightleftharpoons \text{H}_3\text{O}^+_{(\text{aq})} + \text{CH}_3\text{COO}^-_{(\text{aq})}$

If acid is added, the system adjusts (according to Le Châtelier's Principle) by shifting the equilibrium to the left. In this case, the acetate ions will react with the excess H_3O^+ to keep the overall $[\text{H}_3\text{O}^+]$ relatively constant. Similarly, if a base is added, the system adjusts by shifting the equilibrium to the right to replace the H_3O^+ which reacted with the OH^- .

In both these cases, the acid or base can be consumed without a great alteration to the $[H_3O^+]$ in the original mixture. As a result, the pH of the mixture does not alter appreciably. Of course, buffer solutions have their limits and eventually if enough acid or base is added, the pH will change significantly.

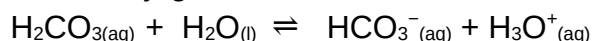
Explain buffering using Collision Theory:

The buffer capacity of a solution depends on two factors:

- The relative concentrations of the weak acid (base) and its conjugate base (acid). Equal concentrations of these two offers the best buffering capacity.
- The concentration of the weak acid (base) and its conjugate base (acid). The greater the concentrations (in general) the greater the buffering capacity.

Blood has a system of buffers present in it to maintain a fairly consistent pH range between 7.35 and 7.45. This is essential to maintain the body systems functioning.

One of the biffer systems to help blood maintain this narrow pH range is the **carbonic acid** and **hydrogen carbonate ion** buffer system. Carbonic acid is the weak acid and the hydrogen carbonate ion is the weak conjugate base.



An enzyme found in blood called carbonic anhydrase catalyses the formation of carbonic acid from carbon dioxide which ensures the equilibrium adjusts quickly to changes in $H_2CO_{3(aq)}$ or $CO_{2(aq)}$. The $H_2CO_3/HCO_3^-/CO_2$ buffer system is particularly useful because it allows the excess H_3O^+ to exit the body through the lungs as CO_2 .

Another blood buffer system is the dihydrogenphosphate and hydrogenphosphate, $H_2PO_4^-/HPO_4^{2-}$ buffer system.

Complete Essential Chemistry Set 6

Acid-Base Titrations

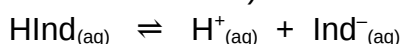
Acid- base titrations (also called volumetric analysis) are an analytical procedure often used to determine the concentration of a solution or the amount of a particular substance present. The concentration of one solution must be accurately known. This is called the **standard solution**. The purpose of the titration is to determine the concentration of the acid or base in the unknown solution.

Five main pieces of equipment are used in titrations:

- Pipette – used to accurately deliver a known volume (aliquot) of liquid.
It is rinsed with the liquid to be delivered from it prior to use.
 - Burette – used to accurately deliver a variable volume (titre) of liquid.
It is rinsed with the liquid to be delivered from it prior to use.
 - Volumetric flask – used to prepare an accurately known volume of solution.
It is rinsed with de-ionised water prior to use.
 - Analytical balance – used to weigh out an accurately known mass.
 - Conical flask – used to hold the solutions during the titration. It is rinsed with de-ionised water prior to use.
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Indicators are used to determine when stoichiometrically equivalent amounts of the two reactants have reacted. This point is called the equivalency point. The point at which the indicator changes colour is called the end point. To be a useful indicator, the end point should match the equivalency point.

For an indicator (which is a weak acid) in solution, the following reaction takes place:



The colour changes are due to the difference in the colours of HInd and Ind⁻.

A standard solution has an accurately known concentration. A primary standard solution is used to determine the concentration of a secondary standard solution. A primary standard solution should possess the following characteristics:

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Common primary standards used in acid-base titrations include **anhydrous sodium carbonate**, **oxalic acid ($\text{H}_2\text{C}_2\text{O}_4$)** and **potassium hydrogenphthalate ($\text{KHC}_8\text{H}_4\text{O}_4$)**.

When performing titrations, a practice run is always performed and at least three trials which should agree to within ± 0.3 mL (at least two of them). The practice run is never used in the calculating of the concentration.

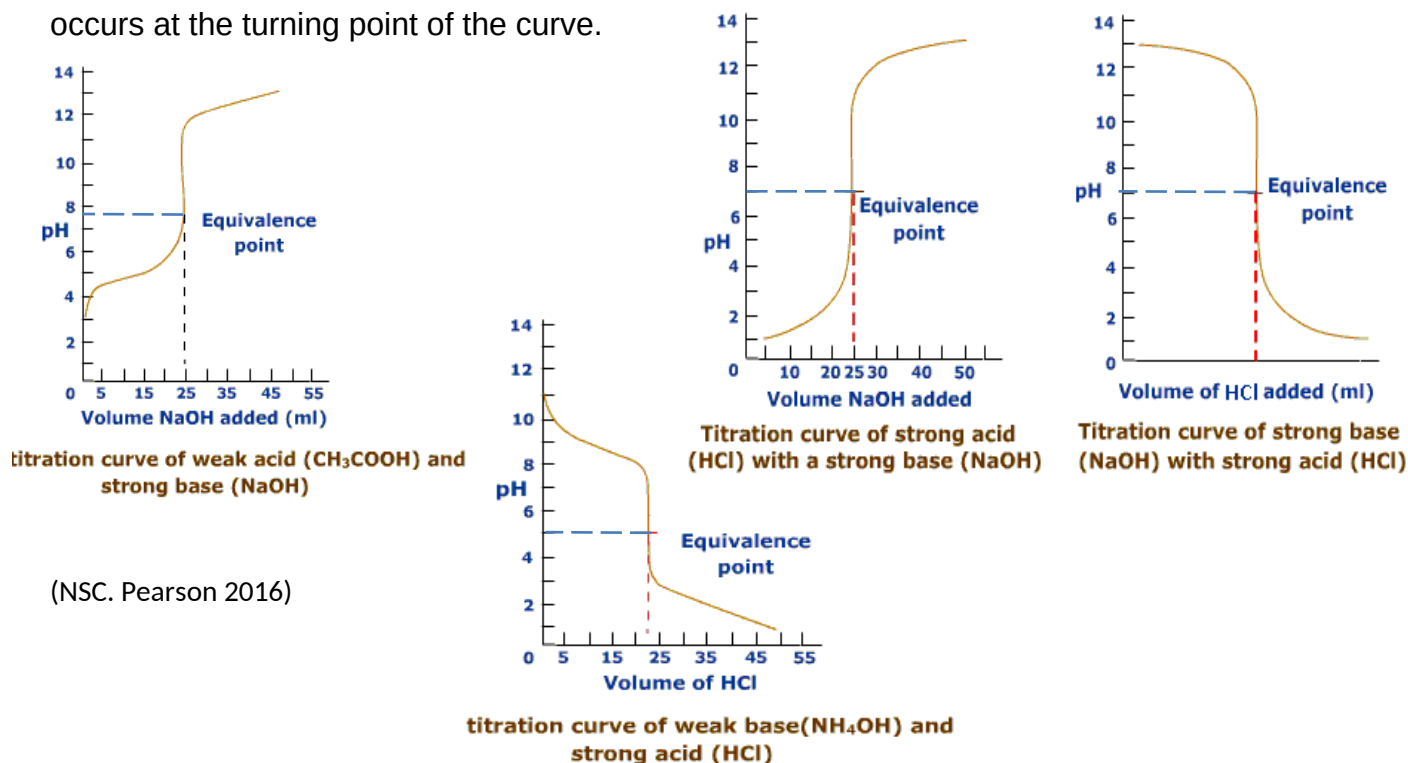
The Products of Neutralisation Reactions

Not all equivalency points of acid-base reactions will be neutral. Only the reaction between a strong acid and a strong base will produce a neutral solution. As such, the choice of indicator is very important. Below is a list of common indicators and their useful pH range.

| Indicator | Colour in Acid | Colour in Base | pH Range |
|---------------------|----------------|----------------|-----------|
| Universal indicator | Red | Purple | 0.0-14.0 |
| Methyl orange* | Red | Yellow | 3.1-4.4 |
| Bromophenol blue* | Yellow | Purple | 3.0-4.6 |
| Methyl red | Red | Yellow | 4.4-6.2 |
| Litmus | Red | Blue | 4.5-8.3 |
| Bromothymol blue | Yellow | Blue | 6.0-7.6 |
| Phenolphthalein* | Colourless | Pink | 8.3-10.0 |
| Alizarine yellow R | Yellow | Red | 10.2-12.0 |

*denotes more commonly used (Although Universal Indicator is never used as an acid/base titration, you must be familiar with its colours in various pH ranges i.e. red 0-2, orange 3-4, yellow 5-6, green 7, blue 8-10, violet 11-14)

During a titration, the pH does not change in a linear fashion as shown by the **titration curves** below. Below are typical titration curves for various reactions. Note that the equivalence point occurs at the turning point of the curve.



(NSC. Pearson 2016)

Reactions Between a Weak Acid and a Strong Base

Write an equation for the reaction between acetic acid and sodium hydroxide in aqueous solution. Which salt is produced? Are solutions of this salt acidic, basic or neutral? Which indicator should be used?

Write an equation for the reaction between hypochlorous acid (HClO) and potassium hydroxide in aqueous solution. Which salt is produced? Are solutions of this salt acidic, basic or neutral? Which indicator should be used?

Summing up, the end point of titrations between a weak acid and a strong base in aqueous solution are basic and the indicator to be used is phenolphthalein.

Reactions Between a Strong Acid and a Weak Base

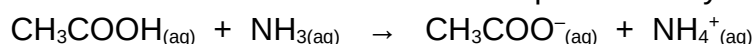
Which salt is formed when an aqueous solution of ammonia is added to a hydrochloric acid solution? Are solutions of this salt acidic, basic or neutral? Which indicator should be used?

Which salt is formed when an aqueous solution of sodium hydrogencarbonate is added to nitric acid? Are solutions of this salt acidic, basic or neutral? Which indicator should be used?

Summing up, the end point of titrations between a strong acid and a weak base in aqueous solution are acidic and the indicator to be used is methyl orange or bromophenol blue.

Reaction Between a Weak Acid and A Weak Base

The reaction between acetic acid and ammonia can be represented by the equation:



The pH of solutions of salts such as ammonium acetate depend on the relative strengths of the anion as a base and the cation as an acid. Ammonium acetate solutions are very close to neutral. As such, a pH meter is definitely the more preferred choice to monitor this type of acid-base titration as the pH change around the equivalence point is too gradual.

The titration calculation

Eg 1 20.00 mL aliquots of hydrochloric acid solution are titrated with $9.039 \times 10^{-2} \text{ mol L}^{-1}$ sodium carbonate solution. Several titrations were performed and the volumes of sodium carbonate solution used were 19.45 mL, 18.90 mL, 18.77 mL and 18.85 mL. What is the concentration of the hydrochloric acid solution?

In a **back titration**, a known volume of an excess reactant is added to the reactant being analysed. A titration is then carried out to determine the amount of excess reactant remaining after the reaction.

Eg 2 Magnesium oxide is not very soluble in water and is difficult to titrate directly. Its purity can be determined by use of the back titration method. 4.06 g of impure magnesium oxide was completely dissolved in 100.0 mL of hydrochloric acid, of concentration 2.00 mol L^{-1} (in excess). The excess acid required 19.70 mL of sodium hydroxide (0.200 mol L^{-1}) for neutralisation. Determine the percentage purity of the magnesium oxide in the original sample.

Sometimes the solution to be analysed is too concentrated for the standard solution and would require a titre too large for the burette. In this case the solution of high concentration is **diluted**.

The following equation will be used in these calculations: $c_1V_1 = c_2V_2$ where

c_1 = original undiluted solution concentration in mol L^{-1}

V_1 = original undiluted solution volume in litres
 c_2 = new diluted solution concentration in mol L^{-1}
 V_2 = new diluted solution volume in litres

Eg.3 A laboratory technician needs to accurately confirm the concentration of a batch of commercial hydrochloric acid. In order to do so a 15.00 mL sample is transferred to a 2.500 L volumetric flask and made up to the mark with distilled water. 20.00 mL aliquots require on average 16.65 mL of $9.08 \times 10^{-2} \text{ mol L}^{-1} \text{ NaOH}_{(\text{aq})}$ for equivalence. Determine the concentration of $\text{HCl}_{(\text{aq})}$ in the commercial solution.

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