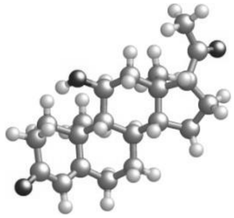


Chem!stry

Name: …………………… (   )

Class: ………………

Date: …… / …… / ……

Essential Notes: Acids, Bases and Salts

Names and Formulae of Some Common Acids, Bases and Salts:

• **Acids:**

Examples include: hydrochloric acid – HCl, nitric acid – HNO_3, sulfuric acid – H_2SO_4, phosphoric acid – H_3PO_4, ethanoic acid – CH_3COOH

• **Bases / Alkalis:**

Note: The term *alkali* is used to describe a base that is soluble in water, and has been dissolved in water so that it can be used as a solution. Not all bases are soluble in water, and so not all bases are alkalis. Examples include: copper(II) oxide – CuO, sodium hydroxide – $NaOH$, calcium hydroxide – $Ca(OH)_2$, ammonia – NH_3.

• **Salts:**

Examples include: sodium chloride – $NaCl$, copper(II) sulphate – $CuSO_4$, ammonium nitrate – NH_4NO_3

Definitions:

• **Acid:**

An acid is a chemical that dissolves in water to produce hydrogen ions, $H^+(aq)$, as the only positive ion:

$$HCl(g) \rightarrow H^+(aq) + Cl^-(aq)$$

$$HNO_3(l) \rightarrow H^+(aq) + NO_3^-(aq)$$

$$H_2SO_4(l) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$$

Note: Sulfuric acid is said to be dibasic because 1 mol of the acid produces 2 mol of hydrogen ions. The hydrogen ions that are produced do not exist on their own, but bond to water molecules to form hydroxonium ions, $H_3O^+(aq)$:

$$H^+(aq) + H_2O(l) \rightarrow H_3O^+(aq)$$

Therefore, the complete balanced chemical equation for an acid dissolving in water should be written:

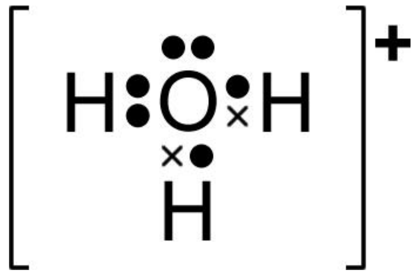
$$HNO_3(l) + H_2O(l) \rightarrow H_3O^+(aq) + NO_3^-(aq) \text{ (a monobasic acid)}$$

$$CH_3COOH(l) + H_2O(l) \rightarrow H_3O^+(aq) + CH_3COO^-(aq) \text{ (a monobasic acid)}$$

$$H_2SO_4(l) + 2H_2O(l) \rightarrow 2H_3O^+(aq) + SO_4^{2-}(aq) \text{ (a dibasic acid)}$$

$$H_3PO_4(l) + 3H_2O(l) \rightarrow 3H_3O^+(aq) + PO_4^{3-}(aq) \text{ (a tribasic acid)}$$

The dot and cross diagram of the hydroxonium ion is given below:



Note: The hydrogen that originated from the acid (shown on the left) is sharing two electrons (two dots) which *both* belong to the oxygen. This special type of covalent bond is known as a dative covalent bond.

An important fact to arise from this is that an acidic chemical will only exhibit its acidic properties when dissolved in water. Therefore, when writing balanced chemical equations, all acids must be in an aqueous state, (aq).

• **Base / Alkali:**

Bases tend to be metal oxides and metal hydroxides. Ammonia, NH_3, is also basic. As opposed to producing hydrogen ions, which is the property of an acid, a base will accept hydrogen ions to form water as one of the products:

$$CuO(s) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + H_2O(l)$$

$$NaOH(s) + H^+(aq) \rightarrow Na^+(aq) + H_2O(l)$$

A base which is soluble in water is described as an alkali. Examples of alkalis include sodium hydroxide, NaOH, and ammonia, NH_3. Alkalis tend to dissolve in water to produce hydroxide ions, $OH^-(aq)$:

$$NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$$

$$NH_3(g) + H_2O(l) \rightarrow NH_4^+(aq) + OH^-(aq)$$

Hydroxide ions (from alkalis) react with hydrogen ions (from acids) to form water. This is known as a *neutralisation reaction*. The ionic equation is:

$$OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$$

• **Salt:**

Salts are ionic compounds, usually composed of a positive metal ion bonded to a negative non-metal ion. An exception would be a salt that contains a positive ammonium ion, NH_4^+, in place of the positive metal ion. Salts are often prepared by replacing the hydrogen ion(s) of an acid with a metal ion(s):

sodium chloride: $NaCl$ composed of Na^+ and Cl^-

copper(II) sulphate: $CuSO_4$ composed of Cu^{2+} and SO_4^{2-}

ammonium nitrate: NH_4NO_3 composed of NH_4^+ and NO_3^-

<u>Properties of Acids:</u>

- Acids turn blue litmus paper red.
- Acids turn universal indicator red / orange / yellow (depending upon the pH value of the acid).
- Acids have pH values <u>less than 7</u>. A pH of exactly 7 is neutral.

pH = $-\log_{10}[H^+]$

Where $[H^+]$ is the hydrogen ion concentration, mol/dm^3

Example of a pH calculation:

If $[H^+] = 0.001$, then pH = $-\log_{10} 0.001$

pH = $-(-3)$

pH = 3.00 (3 s.f.)

- An aqueous solution of an acid will conduct electricity because it contains mobile ions (charge carrying particles) which are free to move to the electrode of opposite charge. For example, an aqueous solution of sulfuric acid contains a mixture of mobile hydrogen ions, H^+, which will be attracted to the negative electrode (cathode) and mobile sulphate ions, SO_4^{2-}, which will be attracted to the positive electrode (anode).

- <u>Acids react with metals to produce a salt and hydrogen as the products</u>:

acid + metal → salt + hydrogen

nitric acid + magnesium → magnesium nitrate + hydrogen

$$2HNO_3(aq) + Mg(s) \rightarrow Mg(NO_3)_2(aq) + H_2(g)$$

Ionic equation: $2H^+(aq) + Mg(s) \rightarrow H_2(g) + Mg^{2+}(aq)$

This is a <u>redox reaction</u> because the metal is oxidised while the hydrogen is reduced. This reaction can also be considered as a <u>displacement reaction</u>. The hydrogen of the acid will only be displaced by a metal that is <u>more reactive than hydrogen</u> in the reactivity series, *i.e.* copper, gold and silver will <u>not</u> displace hydrogen from an acid (there will be <u>no observed reaction</u> when copper, gold or silver are added to an acid).

Note: Hydrogen gas extinguishes a burning splint with a “pop” sound.

- <u>Acids react with carbonates to produce a salt, water and carbon dioxide as the products</u>:

acid + carbonate → salt + water + carbon dioxide

sulfuric acid + sodium carbonate → sodium sulphate + water + carbon dioxide

$$H_2SO_4(aq) + Na_2CO_3(aq) \rightarrow Na_2SO_4(aq) + H_2O(l) + CO_2(g)$$

Ionic equation: $2H^+(aq) + CO_3^{2-}(aq) \rightarrow H_2O(l) + CO_2(g)$

The salt is formed when the hydrogen of the acid is replaced by the metal of the carbonate. This will work for any metal carbonate, including copper(II) carbonate and silver carbonate.
Note: Carbon dioxide gas forms a white precipitate when bubbled through lime water.

• Acids react with bases / alkalis to produce a salt and water as the products:

acid + base → salt + water

ethanoic acid + calcium hydroxide → calcium ethanoate + water

$$2CH_3COOH(aq) + Ca(OH)_2(aq) \rightarrow (CH_3COO)_2Ca(aq) + 2H_2O(l)$$

Ionic equation: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

The salt is formed when the hydrogen of the acid is replaced by the metal of the metal oxide / hydroxide. This will work for any metal oxide / hydroxide, including copper(II) oxide and silver hydroxide.

Note: When a dibasic or tribasic acid reacts, it is possible for the hydrogen ions of the acid to be replaced by metal ions one-at-a-time, *i.e.* the hydrogen ions of a dibasic or tribasic acid do not need to be replaced all-at-once.

• Example, phosphoric acid reacting with sodium hydroxide:

• Replace first H^+ with Na^+ to form monosodium phosphate:

$$H_3PO_4(aq) + NaOH(aq) \rightarrow NaH_2PO_4(aq) + H_2O(l)$$

• Replace second H^+ with Na^+ to form disodium phosphate:

$$NaH_2PO_4(aq) + NaOH(aq) \rightarrow Na_2HPO_4(aq) + H_2O(l)$$

• Replace third H^+ with Na^+ to form trisodium phosphate:

$$Na_2HPO_4(aq) + NaOH(aq) \rightarrow Na_3PO_4(aq) + H_2O(l)$$

• Overall:

$$H_3PO_4(aq) + 3NaOH(aq) \rightarrow Na_3PO_4(aq) + 3H_2O(l)$$

• Acids fall into two categories, strong acids and weak acids.

Strong acids include hydrochloric acid, nitric acid and sulfuric acid. These acids fully ionize or dissociate when dissolved in water. For example, if 100 molecules of nitric acid, HNO_3, are added to water, all of them will ionize to from 100 hydrogen ions, H^+, and 100 nitrate ions, NO_3^-. There will be no nitric acid molecules left intact:

$$HNO_3(l) \rightarrow H^+(aq) + NO_3^-(aq)$$

Weak acids tend to be organic acids (ones that contain carbon) such as ethanoic acid and citric acid. These acids only partially ionize or dissociate when dissolved in water. For example, if 100 molecules of ethanoic acid, CH_3COOH, are added to water, only 5 of them will ionize to from 5 hydrogen, H^+, ions and 5 ethanoate ions, CH_3COO^-. The remaining 95 ethanoic acid molecules will remain intact:

$$CH_3COOH(l) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$$

The $\rightleftharpoons$ symbol indicates that the reaction is reversible. This means that once the hydrogen ions and ethanoate ions have been produced, they can combine together again to reform molecules of ethanoic acid.

What is the consequence of this? 1 mole of nitric acid will dissolve in water to form 1 mole of hydrogen ions, but 1 mole of ethanoic acid will dissolve in water to form substantially less than 1 mole of hydrogen ions.

One other common acid that is considered to be a relatively weak acid is phosphoric acid – H_3PO_4. Phosphoric acid is often used in the food industry as a preservative, for example, in fizzy drinks.

Properties of Bases / Alkalis:

- Alkalis turn red litmus paper blue.
- Alkalis turn universal indicator blue / purple / violet (depending upon the pH value of the alkali).
- Alkalis have pH values greater than 7. A pH of exactly 7 is neutral.
- Alkalis react with acids to form a salt and water as the reaction products:

nitric acid + sodium hydroxide → sodium nitrate + water

$$HNO_3(aq) + NaOH(aq) \rightarrow NaNO_3(aq) + H_2O(l)$$

sulfuric acid + ammonium hydroxide → ammonium sulphate + water

$$H_2SO_4(aq) + 2NH_4OH(aq) \rightarrow (NH_4)_2SO_4(aq) + 2H_2O(l)$$

- Alkalis react with ammonium salts to form a salt, water and ammonia as the reaction products:

potassium hydroxide + ammonium bromide → potassium bromide + water + ammonia

$$KOH(aq) + NH_4Br(aq) \rightarrow KBr(aq) + H_2O(l) + NH_3(g)$$

The Nature of Metal and Non-metal Oxides:

Oxides of the chemical elements can be classified in four different ways:

- **Acidic:** Generally, the oxides of the non-metallic elements are acidic. Common examples are carbon dioxide, CO_2, and sulphur dioxide, SO_2. Carbon dioxide in the atmosphere dissolves in rain water to form carbonic acid, H_2CO_3, thus making rain water naturally slightly acidic:

$$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$$

Sulphur dioxide is an atmospheric pollutant produced by the combustion of fossil fuels, such as coal, oil and petrol. It dissolves in rain water to form sulphurous acid, H_2SO_3, (acid rain) which is harmful to plants, animals and corrodes buildings:

$$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$$

• **Basic:** Generally, the oxides of metallic elements are basic. Common examples include copper(II) oxide, CuO, and calcium oxide, CaO. Both copper(II) oxide and calcium oxide will react with acids to form a salt and water:

calcium oxide + nitric acid → calcium nitrate + water

$$CaO(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2O(l)$$

• **Amphoteric:** These oxides can exhibit both acidic and basic properties, depending upon the chemical environment that they are placed in. For example, if an amphoteric oxide is added to an acidic solution, it will behave as a base. Alternately, if an amphoteric oxide is added to an alkaline solution, it will behave as an acid! Examples of amphoteric oxides include aluminium oxide, Al_2O_3, lead(II) oxide, PbO, and zinc oxide, ZnO.

• **Neutral:** Neutral oxides have a pH value of 7 when dissolved in pure water. Examples include carbon monoxide, CO, dinitrogen monoxide N_2O, and of course water, H_2O! Note: The neutral oxides all tend to be the *monoxides* of non-metallic elements.

The Properties and Preparation of Salts:

• **Solubility:**

The table below summarises the solubilities of common salts:

Soluble in Water	Insoluble in Water
All salts of Group I metals, *e.g.* Na^+ and K^+.	All carbonates **except** those of the Group 1 metals and ammonium carbonate, $(NH_4)_2CO_3$.
All ammonium salts, NH_4^+.	All hydroxides and oxides **except** those of the Group 1 metals, ammonium hydroxide, NH_4OH, and barium hydroxide $Ba(OH)_2$.
All nitrates, NO_3^-.	All phosphates **except** those of the Group 1 metals and ammonium phosphate, $(NH_4)_3PO_4$.
All chlorides, Cl^- **except** silver chloride, $AgCl$, and lead(II) chloride, $PbCl_2$.	
All sulphates, SO_4^{2-}, **except** barium sulfate, $BaSO_4$, calcium sulfate, $CaSO_4$, and lead(II) sulfate, $PbSO_4$.	
All hydrogencarbonates, HCO_3^-.	

For example:

magnesium chloride, $MgCl_2$, is <u>soluble</u>, but silver chloride, $AgCl$, is <u>insoluble</u>.

sodium sulphate, Na_2SO_4, is <u>soluble</u>, but barium sulphate, $BaSO_4$, is <u>insoluble</u>.

potassium carbonate, K_2CO_3, is <u>soluble</u>, but magnesium carbonate, $MgCO_3$, is <u>insoluble</u>.

sodium hydroxide, NaOH, is <u>soluble</u>, but copper(II) hydroxide, $Cu(OH)_2$, is <u>insoluble</u>.

potassium phosphate, K_3PO_4, is <u>soluble</u>, but calcium phosphate, $Ca_3(PO_4)_2$, is <u>insoluble</u>.

• **Preparation:** There are three possible methods:

1) <u>Reaction of an acidic solution with an alkaline solution to form a soluble salt – titration method</u>

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

• The acidic solution is pipetted into a flask and the alkaline solution is poured into a burette.

• The acidic and alkaline solutions are then titrated against each other using a suitable indicator.

• When the acid and alkali have exactly neutralised each other, as shown by the colour change of the indicator, the burette reading is taken.

• The titration is then repeated <u>without the indicator</u>, but using exactly the same volumes of acid and alkali that were used in the original titration.

• The resulting solution is heated until it becomes saturated, and the salt then left to crystallise from solution at room temperature.

2) <u>Reaction of acidic solution with an insoluble base or insoluble carbonate to form a soluble salt</u>

e.g. $H_2SO_4(aq) + CuO(s) \rightarrow CuSO_4(aq) + H_2O(l)$

• The acidic solution is pipetted into a flask.

• The insoluble base or insoluble carbonate is slowly added to the acid with continuous stirring.

• It is assumed that all of the acid has been neutralised and therefore only the soluble salt (reaction product) is dissolved in solution when no more of the insoluble base or insoluble carbonate is seen to react.

• The excess insoluble base or insoluble carbonate is removed by filtration, leaving a solution of the desired salt.

• The resulting solution is heated until it becomes saturated, and the salt then left to crystallise from solution at room temperature.

3) <u>Reaction of two soluble salts to form an insoluble salt – precipitation method</u>

e.g. $Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaNO_3(aq)$

- A solution of one soluble salt is added to a solution of a second soluble salt to form an insoluble precipitate of a third salt.
- The insoluble precipitate is removed from solution by filtration and rinsed with distilled water to remove any soluble impurities.
- The solid may be pressed between filter papers to aid in its drying.

Note: Never use an insoluble base or insoluble carbonate to form an insoluble salt. You will produce a mixture of two insoluble solids which will be very difficult or even impossible to separate!

Uses of Acids, Bases and Salts:

Sulfuric acid, H_2SO_4, is a very important acid that is manufactured on an industrial scale in many countries around the world. Sulfuric acid has the following uses:

- Sulfuric acid is used as the electrolyte in car batteries.
- Sulfuric acid is used in the manufacture of detergents that are used to remove grease and oil from clothing and cooking utensils.
- Sulfuric acid is used in the manufacture of fertilisers, such as ammonium sulfate, $(NH_4)_2SO_4$, that add nutrients to the soil for plants to grow.

Sulfur dioxide is an acidic oxide that has the following important uses:

- Sulfur dioxide is used to bleach wood pulp during the manufacture of paper. This is because the sulfur dioxide is a *reducing agent* that will reduce the coloured pigments in the wood pulp, turning them white.
- Sulfur dioxide is used as a preservative in certain foods, such as dried fruits and white wine. It acts as a preservative by killing bacteria and mould that would otherwise grow on the food.

The Industrial Manufacture of Ammonia:

Ammonia, NH_3, is a very important alkaline gas that is manufactured on an industrial scale in many countries around the world. Ammonia has the following uses:

- Ammonia is used as a raw material in the manufacture of nitric acid, HNO_3.
- Ammonia is used in the manufacture of fertilisers, such as ammonium nitrate, NH_4NO_3, that add nutrients to the soil for plants to grow.
- Ammonia is used in the manufacture of explosives, such as 2,4,6-trinitrotoluene (TNT), $C_7H_5N_3O_6$.
- Ammonia is used in the manufacture of plastics, such as nylon.

Ammonia is manufactured by reacting nitrogen gas and hydrogen gas directly together at a temperature of 450 °C, a pressure of 250 atm. and using an iron catalyst to increase the rate of the reaction:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

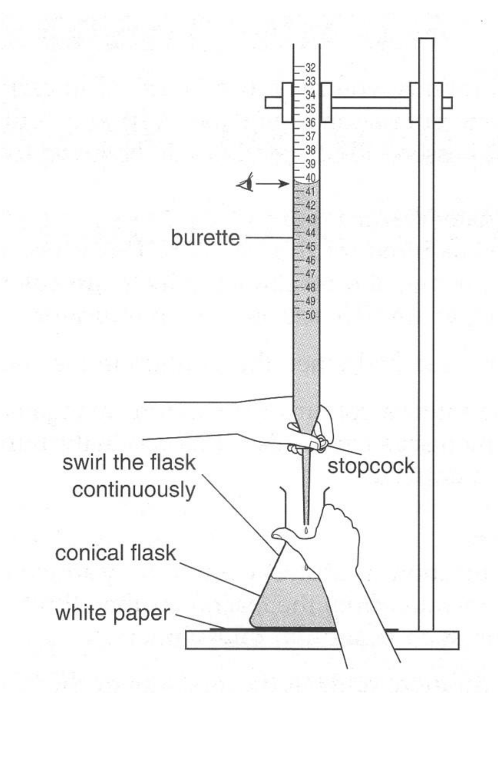
The reaction is reversible, meaning that nitrogen and hydrogen react together to form ammonia, and ammonia can decompose to form nitrogen and hydrogen.

Note about temperature: A *low* temperature gives a high yield of ammonia, but a low temperature also means that the reaction will be *slow*. A temperature of 450 °C is a compromise, offering a reasonably good yield of ammonia at an acceptable rate.

Note about pressure: A high pressure gives a high yield of ammonia, but high pressures are both hazardous and expensive to maintain. A pressure of 450 atm. is used as a compromise, offering a reasonably good yield of ammonia at an acceptable cost and level of safety.

pH Titration Curves:

• Introduction to titration:



• Titration, also known a volumetric analysis, is often used to determine the concentration of an acid or an alkali.

• One of the reagents is contained in a burette while a known volume of the other reagent is contained in a conical flask.

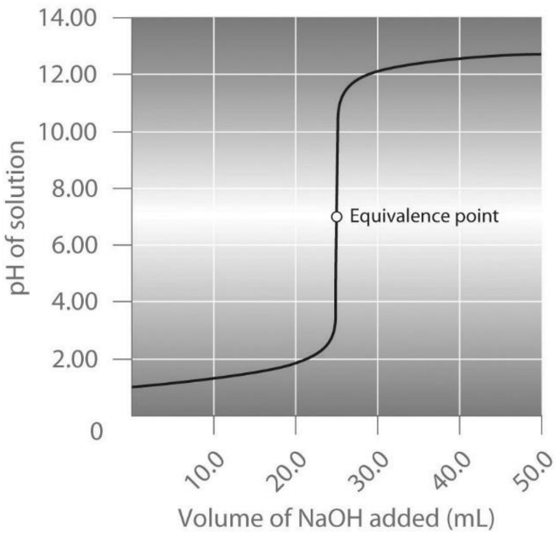
• The reagent in the burette is slowly added to the reagent in the conical flask. An indicator is used to tell when the two reagents have exactly neutralised each other.

• Once the volume of reagent added from the burette is known, the unknown concentration is calculated from:

$$c_{acid} \times V_{acid} = c_{alkali} \times V_{alkali}$$

c = concentration in mol/dm^3
v = volume in cm^3

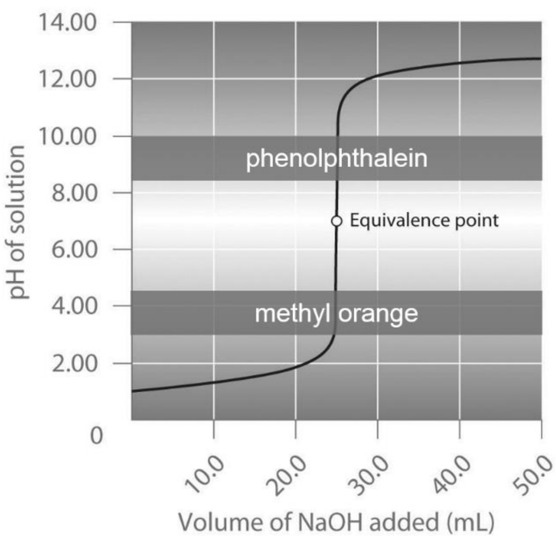
• pH curve obtained when a strong alkali is added to a strong acid:



• This is an example of a pH titration curve that would be obtained when a solution of a strong alkali, *e.g.* NaOH(aq) is added to a solution of a strong acid, *e.g.* HC*l*(aq).

• The <u>equivalence point</u> is the point in the reaction at which the amount (*moles*) of acid and amount (*moles*) of alkali are in the same ratio as the one given by the balanced chemical equation.

• Choice of indicator for a strong acid / strong alkali titration:



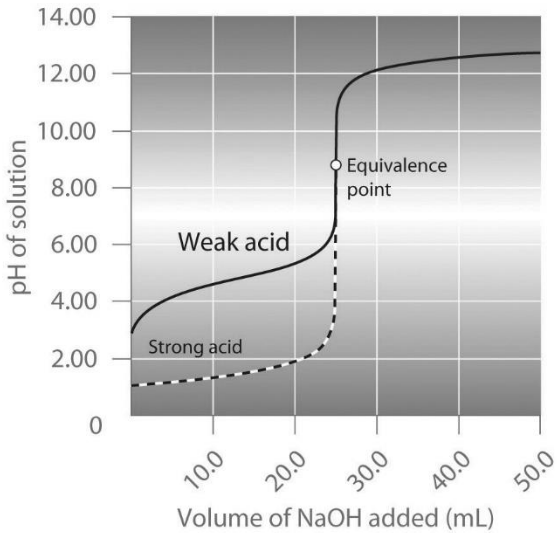
• The equivalence point for the addition of a strong alkali to a strong acid is at pH 7.00.

• A suitable indicator for this titration would ideally change colour at pH 7.00.

• Methyl orange changes colour over the range of 3.1 – 4.4. From the graph, it can be seen that methyl orange will change colour at the equivalence point, *i.e.* when 25.0 cm^3 of NaOH(aq) has been added, hence <u>methyl orange is a suitable indictor for this titration</u>.

• Phenolphthalein changes colour over the range of 8.3 – 10.0. From the graph, it can be seen that phenolphthalein will change colour at the equivalence point, *i.e.* when 25.0 cm^3 of NaOH(aq) has been added, hence <u>phenolphthalein is a suitable indictor for this titration</u>.

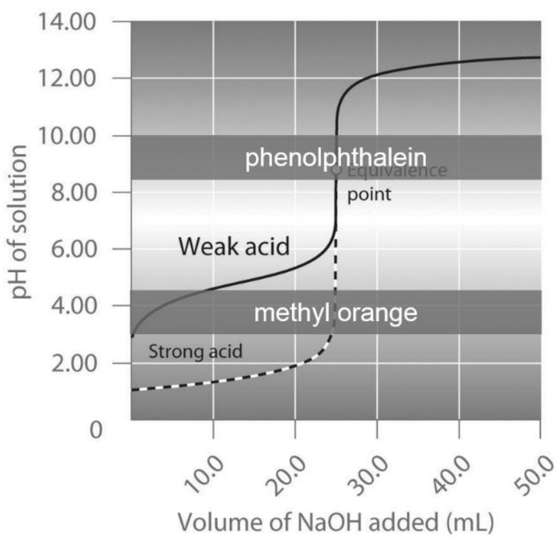
• pH curve obtained when a strong alkali is added to a weak acid:



• The equivalence point for the addition of a strong alkali to a weak acid is at pH 9.00.

• A suitable indicator would ideally change colour at pH 9.00.

• Choice of indicator for a weak acid / strong alkali titration:



• Methyl orange changes colour over the range of 3.1 – 4.4. From the graph, it can be seen that methyl orange will *not* change colour at the equivalence point of this titration. Methyl orange will change colour when only 1 – 10 cm^3 of NaOH(aq) have been added, but a suitable indicator needs to change colour when 25.0 cm^3 of NaOH(aq) have been added, hence methyl orange is *not* a suitable indictor for this titration.

• Phenolphthalein changes colour over the range of 8.3 – 10.0. From the graph, it can be seen that phenolphthalein will change colour at the equivalence point, *i.e.* when 25.0 cm^3 of NaOH(aq) has been added, hence phenolphthalein is a suitable indictor for this titration.

+ + + The End + + +