Mr SGs Acids & Bases Notes

Properties of Acids and Bases

-Acids and bases are common classes of chemicals, with a variety of important uses, ranging from household cooking ang cleaning products to chemicals with important industrial and biological applications

-While several theories have been proposed to classify acids and bases, they have traditionally been defined by their properties

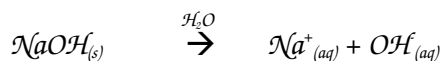| Properties of acids	Properties of Bases
-Taste sour	-Taste bitter
-Are corrosive	-Often corrosive
-Many acid solutions conduct electricity	-Many solutions of bases conduct electricity
-React with reactive metals to form H_2	-Do not react with most metals
-React with bases to form a salt and water	-React with acids to form salts and water
-React with carbonates and hydrogencarbonates to form H_2O and CO_2	-Feel soapy, slippery (react with fats/oils to produce soaps)
-Turn blue litmus red	-Turn red litmus blue

Arrhenius theory of acids & bases

-The Swedish Chemist Arrhenius proposed one of the earliest definitions of acids and bases

-Arrhenius acids have H in their formula and release H^+ ions in solution (e.g. HCl, H_2SO_4)
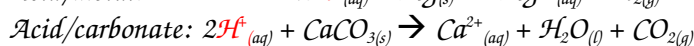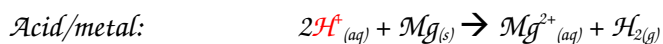
$$HCl_{(g)} \xrightarrow{H_2O} H^+_{(aq)} + Cl^-_{(aq)}$$

-Arrhenius bases have OH in their formula and release OH^- ions in solution (e.g. $NaOH$, KOH)

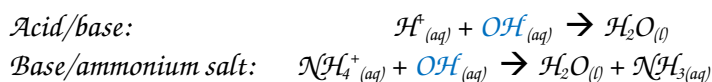$$NaOH_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + OH^-_{(aq)}$$

-Under the Arrhenius theory, all of the properties and reactions of acids can be explained by the behaviour of H^+ ions in solution
-This can be observed in the ionic equations of the reactions of strong acids, where the anion of the acid is typically a spectator ion, and the hydrogen ion is the species that reacts with the metal/carbonate/base

Acid/metal: $\quad 2H^+_{(aq)} + Mg_{(s)} \rightarrow Mg^{2+}_{(aq)} + H_{2(g)}$
Acid/carbonate: $2H^+_{(aq)} + CaCO_{3(s)} \rightarrow Ca^{2+}_{(aq)} + H_2O_{(l)} + CO_{2(g)}$

-Likewise, the Arrhenius theory suggests that the properties and reactions of bases can be explained by the behaviour of OH ions in solution

-This can be observed in the ionic equations of the reactions of strong bases, where the cation of the base is typically a spectator ion, and the hydroxide ion is the species that reacts with the acid/ammonium salt

Acid/base: $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

Base/ammonium salt: $NH_4^+{}_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)} + NH_{3(aq)}$

-While the Arrhenius theory explains most of the properties and reactions of acids and bases, it does have some limitations (particularly in relation to weak bases)

Arrhenius' theory explains:	Limitations of the Arrhenius Theory:
-Neutralisation reactions	-Not all bases contain OH (e.g. NH_3)
-Conductivity of solutions of acids	-Only applies to solutions in water
-Reactions of acids with metals, carbonates, and hydrogen carbonates	-Not all salts are neutral
	-H^+ ion quickly forms H_3O^+ (hydronium) ion in water

-Other acid/base theories have been developed to address these limitations of the Arrhenius theory (not assessed in Yr 11)

Theory	Brønsted-Lowry	Lewis
Acid definition	Proton donor	Electron pair acceptor
Example	$HCl + H_2O \rightarrow Cl^- + H_3O^+$	$2H^+ + 2e^- \rightarrow H_2$
Base definition	Proton acceptor	Electron pair donor
Example	$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$	$OH^- + H^+ \rightarrow H_2O$ (OH supplies both e^- for covalent bond)
Advantages (vs Arrhenius)	-Explain weak acids (e.g. NH_3/HCO_3^-, etc.) -Can use H_3O^+ (instead of H^+) to explain acidic properties -Applies to all reactions, not just those that occur in aqueous solution	

Strong & Weak Acids & Bases

Strength vs Concentration

-Strength refers to electrolyte strength, the degree to which a substance ionises in solution (for strong acids) or dissociates in solution (strong bases)

-Concentration refers to the amount of solute per unit of solution (a concentrated solution has a high concentration as measured in $molL^{-1}$)

-Strength is a property of a particular acid/base, while solutions of any acid/base can be prepared in a range of concentrations

-Both the strength of an acid/base, and its concentration will affect the concentration of H^+ /OH^- ions in solution and hence the acidic/basic properties of the solution

Strong and weak acids

-Strong acids are those where almost all of the acid molecules ionise, e.g. the reaction below will proceed almost to completion

$$HA \rightarrow H^+ + A^-$$
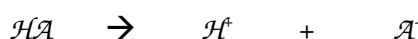
-Strong acids include hydrochloric (HCl), nitric (HNO_3) and sulfuric (H_2SO_4) acids

$$HCl \rightarrow H^+ + Cl^-$$
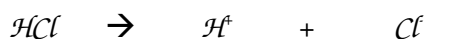
-Weak acids are those where only some of the acid molecules ionise, e.g. for the reaction below, both HA and A^- will be present to a significant extent at equilibrium

$$HA \leftrightarrow H^+ + A^-$$
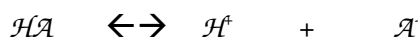
-Weak acids include ethanoic (CH_3COOH), phosphoric (H_3PO_4) carbonic (H_2CO_3), hydrofluoric (HF) and citric ($C_6H_8O_7$) acids (all acids except HCl/HI/H_2SO_4/HNO_3)
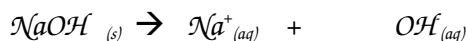
$$HF \leftrightarrow H^+ + F^-$$

-Most non-metal oxides can also be considered weakly acidic, as they react with water to form acidic hydroxyl containing compounds
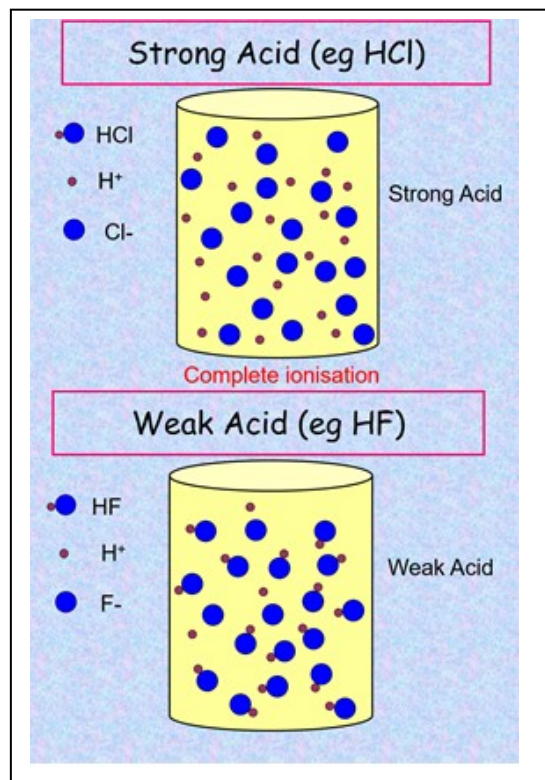
$$CO_2 + H_2O \rightarrow H_2CO_3$$

$$H_2CO_3 \leftrightarrow HCO_3^- + H^+$$
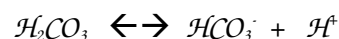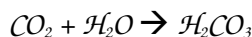
Strong and weak bases

-Strong bases are those that completely dissociate in water to produce hydroxide ions:
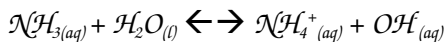
$$NaOH_{(s)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

-Strong bases include the Group I and II metal hydroxides, as well as most (but not all) metal oxides

-Weak bases are those that produce hydroxide ions in solution by removing an H^+ from a water molecule

-This is best explained by the Brønsted-Lowry theory, which defines bases as proton (H^+) acceptors

$$NH_{3(aq)} + H_2O_{(l)} \longleftrightarrow NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$$

-Weak bases include ammonia and sodium carbonate (the carbonate ion can act as a proton acceptor)


Ionic equations for strong and weak bases

-When writing ionic equations, strong electrolytes in the aqueous state are written as their component ions (e.g. $HCl_{(aq)}$ would be written as $H^+{}_{(aq)}$ and $Cl^-{}_{(aq)}$)

-Strong acids and bases are strong electrolytes, so they are always written as their component ions when in the aqueous state, e.g. the ionic equation for the reaction between sodium hydroxide and hydrochloric acid would be written as:

$$H^+{}_{(aq)} \quad + \quad OH^-{}_{(aq)} \rightarrow \quad H_2O_{(l)}$$
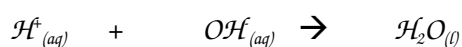
-Weak acids and bases are weak electrolytes that only partially ionise, so they are always written as their intact molecules when in the aqueous state, e.g. the ionic equation for the reaction between ammonia and ethanoic acid would be written as:

$$CH_3COOH_{(aq)} \quad + \quad NH_{3(aq)} \rightarrow \quad NH_4CH_3COO_{(aq)}$$
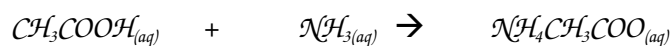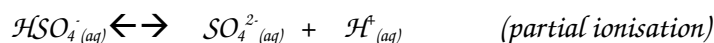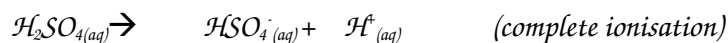
Polyprotic acids

-Acids that can donate one proton (H^+) per molecule (e.g. HCl, HNO_3) are monoprotic
(mono = 1, protic = protons (H^+) to donate)

-Acids that can donate more than one proton per molecule (e.g. H_2SO_4) are polyprotic (diprotic in the case of H_2SO_4)

-Only H atoms that can be given up in solution are counted when determining if an acid is polyprotic, e.g. CH_3COOH is monoprotic, as it can only donate the hydrogen in the COOH group)
-A 1 $mol L^{-1}$ solution of a polyprotic acid will have a higher H_3O^+ concentration than a 1 $mol L^{-1}$ solution of a monoprotic acid (assuming the first ionisation occurs to the same degree)

-Each successive ionisation occurs to a lesser degree than the previous ionisation, e.g.:

$$H_2SO_{4(aq)} \rightarrow \quad HSO_4^-{}_{(aq)} + \quad H^+{}_{(aq)} \qquad \text{(complete ionisation)}$$

$$HSO_4^-{}_{(aq)} \longleftrightarrow \quad SO_4^{2-}{}_{(aq)} + \quad H^+{}_{(aq)} \qquad \text{(partial ionisation)}$$

-A solution of sulfuric acid contains the following species in order of decreasing concentration
$$[H_2O] > [H^+] > [HSO_4^-] > [SO_4^{2-}] > [H_2SO_4] \text{ (negligible)}$$
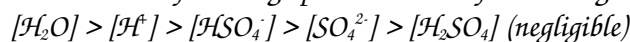
-For a weak diprotic acid with the formula H_2A, a solution of the acid will contain the following species in order of decreasing concentration

$$[H_2O] > [H_2A] > [H^+] > [HA^-] > [A^{2-}]$$
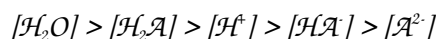
Reactions of Acids and Bases

-As acids and bases react in a predictable manner with many classes of chemicals, general equations can be used to predict

the products of these reaction types

-In the equations below, acids include non-metal oxides and bases include most metal oxides and hydroxides

-The salts produced in these reactions contain the cation from the species reacting with the acid and the anion from the acid

General equation	Example
acid + metal → salt + hydrogen	$2HCl_{(aq)} + Mg_{(s)} \rightarrow MgCl_{2(aq)}$
acid+ base → salt + water	$H_2SO_{4(aq)} + 2NaOH_{(aq)} \rightarrow Na_2SO_{4(aq)} + 2H_2O_{(l)}$
acid + metal (hydrogen) carbonate → salt + water + carbon dioxide	$2HNO_{3(aq)} + CaCO_{3(s)} \rightarrow Ca(NO_3)_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$
base + ammonium salt → salt + water + ammonia	$LiOH_{(aq)} + NH_4Cl_{(aq)} \rightarrow LiCl_{(aq)} + H_2O_{(l)} + NH_{3(g)}$
acid + metal sulfite → salt + water + sulfur dioxide (E)	$2HI_{(aq)} + K_2SO_{3(aq)} \rightarrow 2KI_{(aq)} + H_2O_{(l)} + SO_{2(g)}$

Measuring Acidity

Self ionisation of water

-Water self-ionises to a small extent, according to the reversible reaction:

$$H_2O_{(l)} \longleftrightarrow H^+_{(aq)} + OH^-_{(aq)}$$
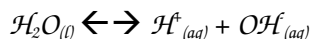
-In any aqueous solution, at 25°C:

$$[H^+][OH^-] = 10^{-14} \ mol^2 \ L^{-2}$$

-The higher the hydrogen ion concentration of a solution, the lower the hydroxide ion concentration and vice versa, e.g. if a solution had a hydrogen ion concentration of $10^{-6} \ mol \ L^{-1}$ it would have a hydroxide ion concentration of $10^{-8} \ mol \ L^{-1}$

-The hydrogen ion concentration of a solution can be related to the hydroxide ion concentration using the formulae

$$[H^+] = \frac{10^{-14}}{[OH^-]} \qquad and \qquad [OH^-] = \frac{10^{-14}}{[H^+]}$$

Acidic, basic, and neutral solutions

-The acidity or basicity of a solution is defined in relation to the relative concentration of hydrogen and hydroxide ions

Acidity	pH (at 25°C)	$[H^+]$ (at 25°C)	$[OH^-]$ (at 25°C)
Acidic	<7	$>10^{-7} \ mol \ L^{-1}$	$<10^{-7} \ mol \ L^{-1}$
Neutral	7	$10^{-7} \ mol \ L^{-1}$	$10^{-7} \ mol \ L^{-1}$
Basic	>7	$<10^{-7} \ mol \ L^{-1}$	$>10^{-7} \ mol \ L^{-1}$

The pH scale

-The acidity of a solution can be measured by measuring the concentration of hydrogen ions

-$[H^+]$ of solutions can vary wildly, from as low as 10^{-15} to as high as $10 \ mol L^{-1}$, so it is more convenient to measure acidity using a logarithmic scale (the pH scale)

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

e.g. a solution with a hydrogen ion concentration of $10^{-8} \ mol L^{-1}$ has a pH of 8.0

-At 25°C, acidic solutions have a pH of below 7 and basic solutions have a pH of above 7

-For pH values, only digits after the decimal point are significant e.g. a pH of 6.4 has one significant figure

-The pH formula can be rearranged to calculate the hydrogen ion concentration of a solution with a given pH

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

NOTE: pH calculations are NOT in the Year 11 Syllabus

Indicators

-The acidity or basicity of solutions can also be determined using indicators

-Indicators are weak acids that exist in equilibrium with their conjugate base according to:

$$HInd \longleftrightarrow Ind^- + H^+$$

-The acidic (Hind) and basic (Ind-) forms of indicators have different colours (they can either both be coloured, or one can be colourless)

-When the pH of a solution decreases, some of the extra hydrogen ions react with the basic form of the indicator increasing the concentration of the acidic form and changing the colour of the solution

-When the pH of a solution increases, the hydrogen ion concentration decreases, so there are less hydrogen ions to react with the basic form of the indicator, decreasing the concentration of the acidic form and changing the colour of the solution

-Different indicators will change colour at different pH values, and some polyprotic indicators may change colour twice at different pH values

-Universal indicator is a mixture of several indicators, allowing pH to be monitored over a wide range
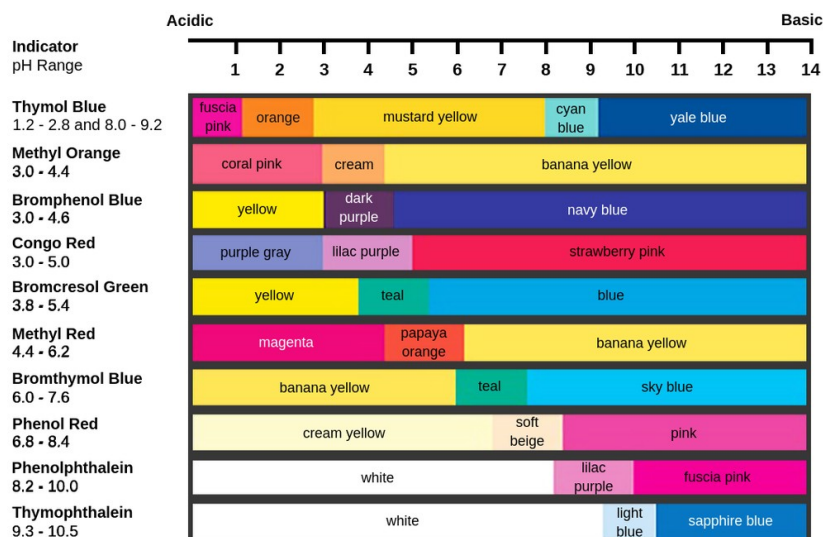
Acid Base Stoichiometry and Limiting Reagents

-Many of the reactions we study using stoichiometry are acid-base reactions

-The coefficients their chemical equations give us the molar ratios in which the reactants are consumed, and the products are produced

-When performing practical chemistry, this allows us to calculate the amount of product able to be produced for a given amount of reactants (or vice versa)

-The coefficients in the equation to the right tell us $\qquad$ $6HCl + 2Al \rightarrow 2AlCl_3 + 3H_2$
that 6 mol of HCl will react with 2 mol of Al to produce
2 mol of $AlCl_3$ and 3 mol of H_2

-In reality, it is practically impossible to combine reactants in stoichiometric amounts

-The reactant that is consumed first, limiting the degree to which the reaction proceeds, is known as the limiting reagent

-The limiting reagent must be used for any stoichiometric calculations, as it determines how much product will be formed and how much of the other reactants will be consumed

-If 2 mol of Al was combined with 12 mol of HCl, only 3 mol of H_2 would be produced rather than the 6 mol you would calculate based on the number of moles of HCl, as the reaction would stop once all of the Al had been consumed

-To calculate which is the limiting reagent for the reaction represented by the following equation where lower case letters are coefficients and uppercase letters are formulae:

$wW + xX \rightarrow yY + zZ$
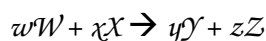
SR (stoichiometric ratio) = w/x AR (actual ratio) = $n(W)/n(X)$
If AR > SR, W is in excess and X is the limiting reagent
If AR < SR, X is in excess and W is the limiting reagent

-A summary of other acceptable methods for calculating LR is shown on p284-5 of CWA1

-The flow chart across the page details common approaches to solving stoichiometry problems, including those where a limiting reagent is present

While limiting reagents are not assessable in the Yr 11 course, we will study them extensively in Yr 12

REACTION STOICHIOMETRY

For the general equation: $wW + xX \longrightarrow yY + zZ$

$m_X(\text{impure})$

$m = \dfrac{\%(X)}{100} \times m_X(\text{impure})$

$V_{X(l)}, p_X$

$m = pV$

m_X

$n = \dfrac{m}{M}$

$c_X, V_{X(aq)}$

$n = cV$

n_X

$n = \dfrac{PV}{RT}$

$V_{X(g)}$

$n = \dfrac{V_{STP}}{22.71}$

$V(STP)_{X(g)}$

LR calc

n_{LR}

$SR = \dfrac{w}{x} \qquad AR = \dfrac{n_W}{n_X}$

If AR > SR, W = ER, X =LR

If AR < SR, X = ER, W =LR

(assuming X is LR)

$n_Y = \dfrac{y}{x} \times n_X$

n_Y

$m = nM$

m_Y

$\%(Y) = \dfrac{m_Y}{m_{total}} \times 100$

$\%(Y)_{(w/w)}$

$V_{Y(l)}$

$V = \dfrac{m}{p}$

$c = \dfrac{n}{V}$

c_Y

DILUTION

$c_1V_1 = c_2V_2$

c_Y

$V = \dfrac{n}{c}$

$V_{Y(aq)}$

$V_{Y(aq)}$

$V = \dfrac{nRT}{P}$

$V_{Y(g)}$

$V_{STP} = n \times 22.71$

$V(STP)_{Y(g)}$

If one or more reactant/product is an acid/base:

$pH = -\log_{10}[H^+] \qquad [H^+] = 10^{-pH} \qquad [H^+][OH^-] = 10^{-14} \text{ (at } 25^0C)$