

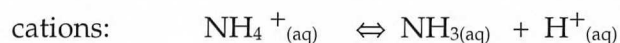
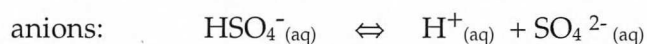
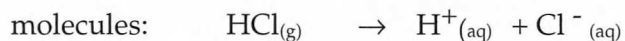
ACIDS and BASES

LESSON 1

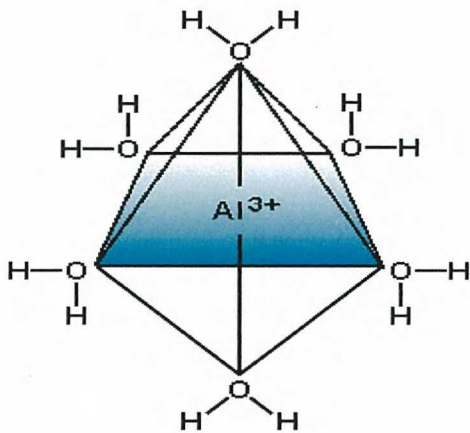
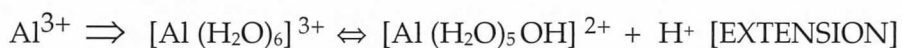
THEORIES OF ACIDS AND BASES

ARRHENIUS THEORY

- **ACIDS** are substances which produce H^+ ions in solution. Acids may be :

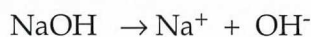


OR



- **BASES** are substances that produce OH^- in aqueous solution, in one of the following ways:

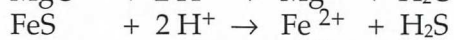
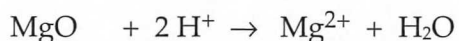
- (i) dissociation of ions of an ionic hydroxide :



- (ii) ionisation of some molecules and anions:



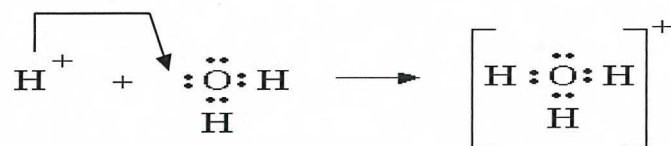
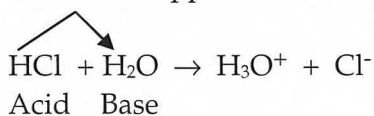
- (iii) insoluble ionic compounds that REACT with acids are also called bases:



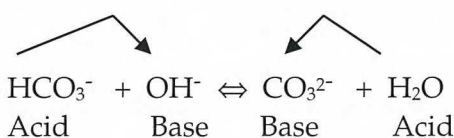
NB: Soluble bases are termed **ALKALIS**.

LOWRY-BRONSTED THEORY

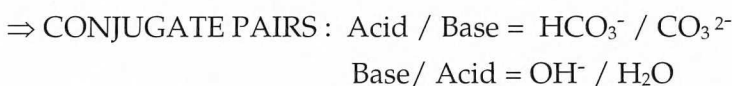
- An **ACID** is a **PROTON DONOR** (H^+). An acid-base reaction is a proton transfer reaction as opposed to a redox reaction which is an electron transfer reaction.



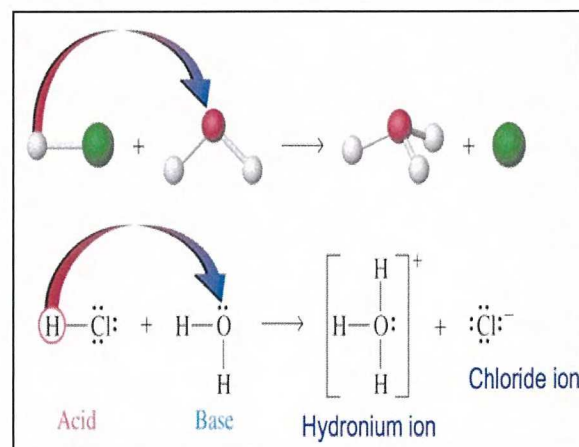
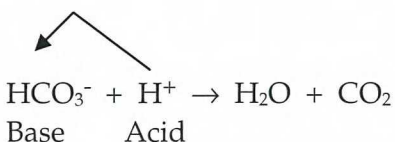
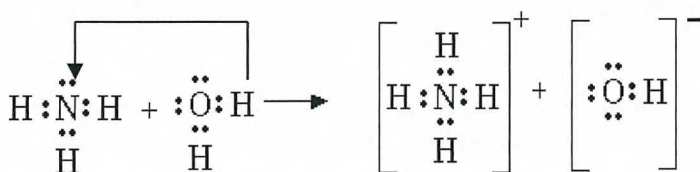
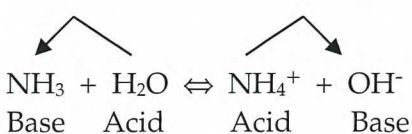
- Many proton transfer reactions can proceed in both the forward and reverse direction and are thus examples of "equilibrium" processes. When this occurs to any extent the products of the forward reaction become the reactants for the reverse reaction and can also be classified as acid or base.



- The **acid** on the left hand side is *transformed* into a **base** on the right hand side of the reaction and vice versa. This acid and its corresponding base are referred to as a **CONJUGATE ACID/BASE PAIR**. The **base** on the left hand side and its corresponding acid on the right would form a **CONJUGATE BASE/ACID PAIR**.



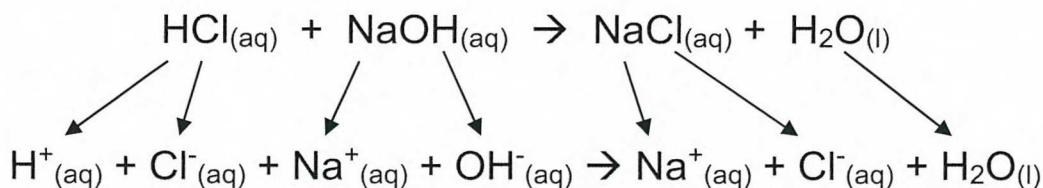
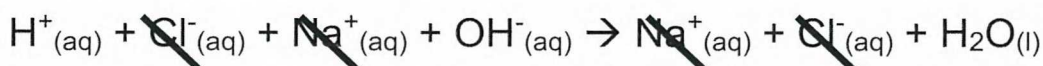
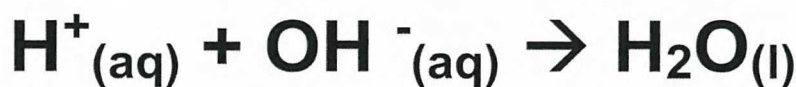
- A **BASE** is a **PROTON ACCEPTOR**.



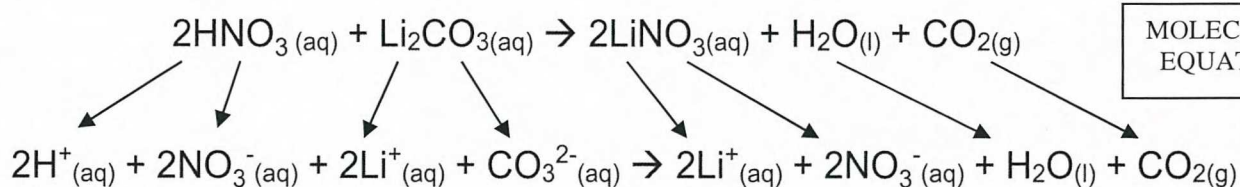
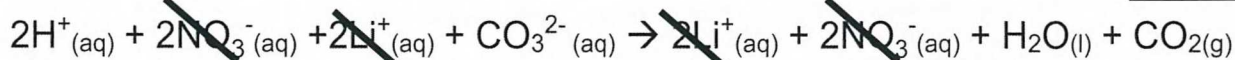
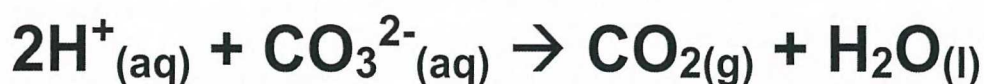
- Some substances such as HCO_3^- can be seen to act as **BOTH** acid and base in different circumstances. These substances are said to be "**AMPHOTERIC**" or "**AMPHIPROTIC**".

PROPERTIES OF ACIDS

1. They have a sour taste.
2. They conduct electricity; some are strong electrolytes and some are weak electrolytes.
3. They turn universal pink and blue litmus red.
4. They react with bases - **NEUTRALISATION** Reaction:

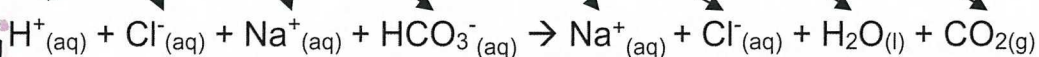
MOLECULAR
EQUATIONTRUE
SPECIESIONIC
EQUATION

5. They react with **CARBONATES**, $\text{CO}_2(\text{g})$ is given off:

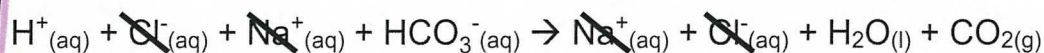
MOLECULAR
EQUATIONTRUE
SPECIESIONIC
EQUATION

6. They react with **BICARBONATES**, CO_2 (g) is given off :

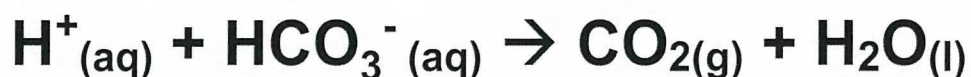
MOLECULAR
EQUATION



TRUE
SPECIES

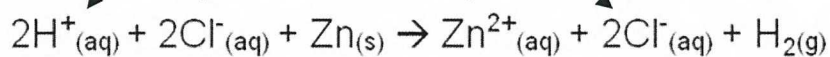


IONIC
EQUATION

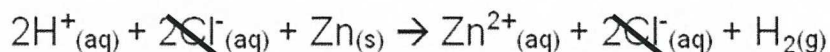


7. They react with *reactive* **METALS** to release H_2 gas:

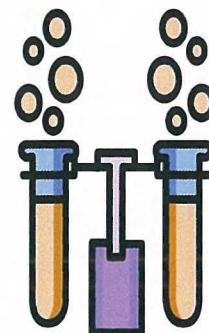
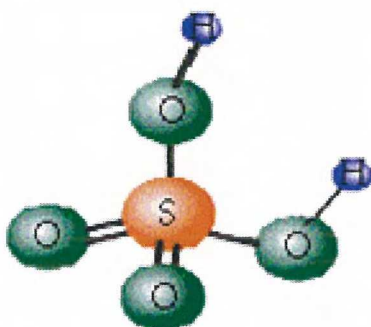
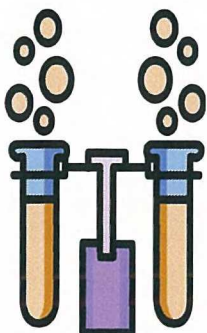
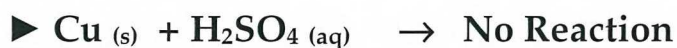
MOLECULAR
EQUATION



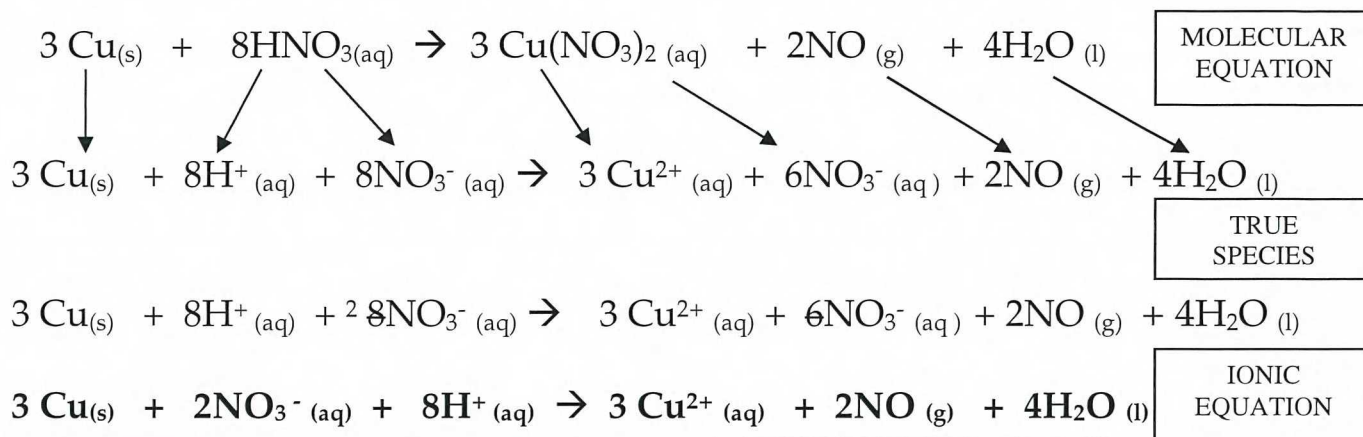
TRUE
SPECIES



IONIC
EQUATION



8. The **OXIDISING** acids react with metals to release NO(g) , $\text{*NO}_2\text{(g)}$ or $\text{*SO}_2\text{(g)}$:



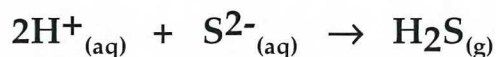
* Extension



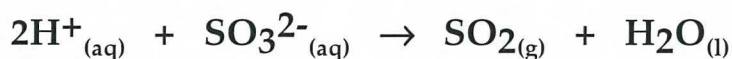
- Acids react with ETHANOATES (CH_3COO^-) to produce a SALT and ETHANOIC ACID giving off a "Vinegar" odour.



- Acids react with SULFIDES (S^{2-}) to produce a SALT and HYDROGEN SULFIDE (H_2S) which is a foul smelling gas known as "Rotten Egg Gas".



- Acids react with SULFITES (SO_3^{2-}) to produce a SALT and SULFUR DIOXIDE (SO_2) gas which is an industrial pollutant and contributing factor to "Acid Rain".

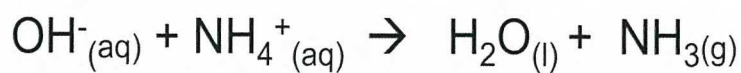
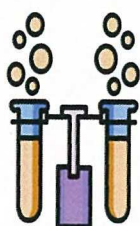
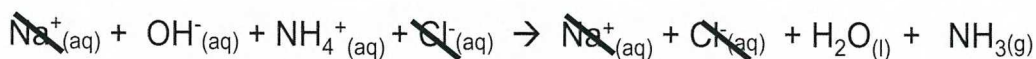
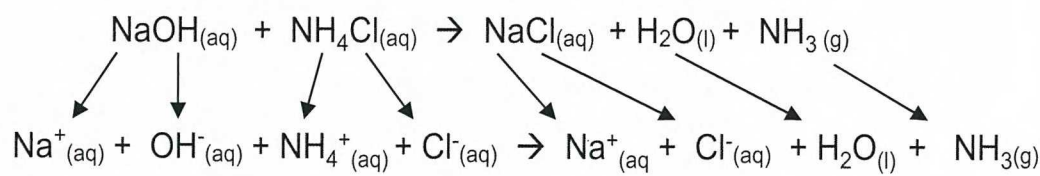


PROPERTIES OF BASES

1. They have a slippery feel.
2. They conduct electricity.
3. They react with acids - NEUTRALISATION.
4. They cause universal to turn purple and red litmus to turn blue.



5. They react with AMMONIUM SALTS to produce WATER and AMMONIA.

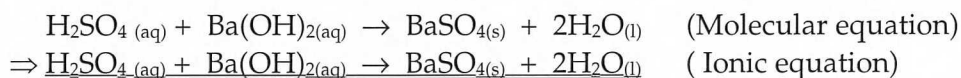
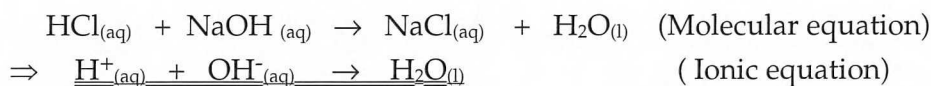
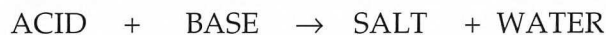
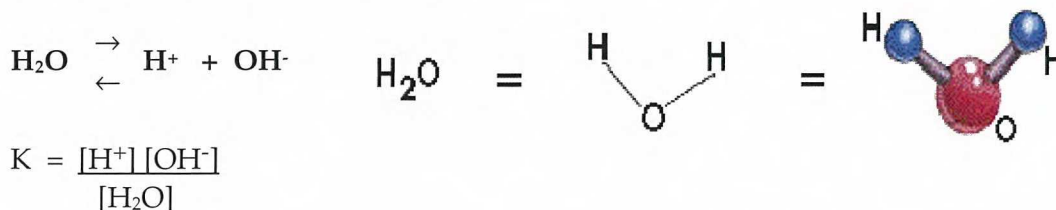


IONIC EQUATION



NEUTRALISATION

- Any time a basic substance reacts with an acid reducing its acidity we say that it has been neutralised. The most common examples are when an acid reacts with a hydroxide base :

AUTOIONISATION OF WATER

$[\text{H}_2\text{O}] = 55.5 \text{ M}$ (constant) \Rightarrow Taken as "1" and cancelled!

$$\therefore K_w = [\text{H}^+][\text{OH}^-] = (10^{-7})(10^{-7}) = \underline{10^{-14}} \quad (\text{at } 25^\circ\text{C})$$

NB: In ALL aqueous solutions K_w will always be 10^{-14} irrespective of what is dissolved in the water!

Calculations :

- (i) Calculate $[\text{H}^+]$ and $[\text{OH}^-]$ in 0.01 M HCl

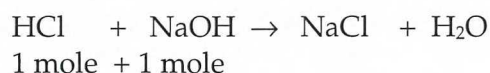
$$[\text{H}^+] = 10^{-2} \text{ M}$$

$$[\text{H}^+][\text{OH}^-] = 10^{-14}$$

$$\therefore [\text{OH}^-] = \frac{10^{-14}}{10^{-2}} = \underline{10^{-12} \text{ M}}$$



- (b) Calculate $[\text{H}^+]$ and $[\text{OH}^-]$ in a mixture of 100 mL of 0.1 M HCl and 50 mL of 0.1 M NaOH. Volumes are **ADDITIVE**.



$$n(\text{HCl})_{\text{in } 100 \text{ mL } 0.1 \text{ M}} = CV = 0.1 \times 0.1 = 0.01 \text{ mol}$$

$$n(\text{NaOH})_{\text{in } 50 \text{ mL } 0.1 \text{ M}} = CV = 0.1 \times 0.05 = 0.005 \text{ mol}$$

$$\text{If all HCl is used} \Rightarrow n(\text{NaOH}) = n(\text{HCl}) \\ = \underline{0.01 \text{ mol}}$$

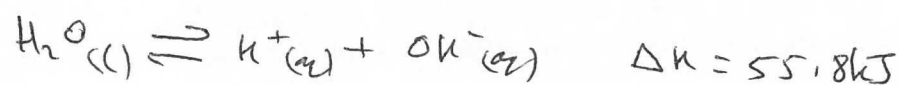
*We have only 0.005 mol of NaOH
 \therefore NaOH is the **LIMITING REACTANT!**

$$\begin{aligned} \therefore n(\text{HCl})_{\text{used}} &= n(\text{NaOH}) \\ &= \underline{0.005 \text{ mol}} \\ \therefore n(\text{HCl})_{\text{remaining}} &= \text{TOTAL} - n(\text{HCl})_{\text{used}} \\ &= 0.01 - 0.005 \\ &= \underline{0.005 \text{ mol}} \\ \therefore [\text{H}^+] = [\text{HCl}] &= \frac{n}{V} = \frac{0.005}{0.150} = \underline{0.0333 \text{ M}} \end{aligned}$$

$$\begin{aligned} \Rightarrow [\text{H}^+] &= 3.33 \times 10^{-2} \text{ M} \\ [\text{H}^+][\text{OH}^-] &= 10^{-14} \\ \therefore [\text{OH}^-] &= \frac{10^{-14}}{3.33 \times 10^{-2}} = \underline{3.00 \times 10^{-13} \text{ M}} \end{aligned}$$



autoionization



endothermic

$$\text{at } 0^\circ\text{C} \quad K_w = 0.144 \times 10^{-14}$$

$$100^\circ\text{C} \quad K_w = 55 \times 10^{-14}$$

increasing temp favors endothermic

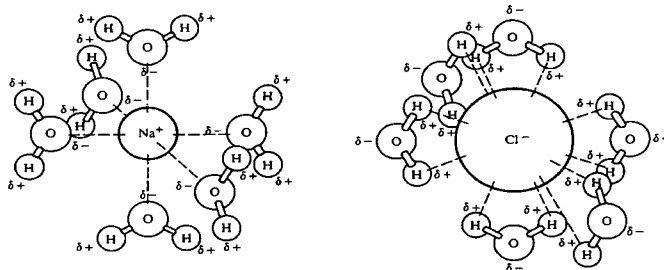
ACIDS and BASES

LESSON 2

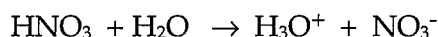
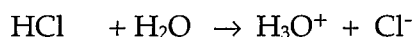
STRONG ACIDS

- Most strong acids are covalent molecular species and therefore are not composed of ions. Through the process of ***IONISATION** which involves their **REACTION** with **WATER** (another molecule) ions may be created through proton transfer.

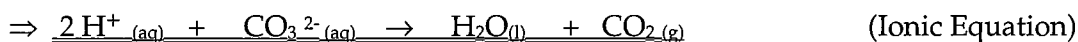
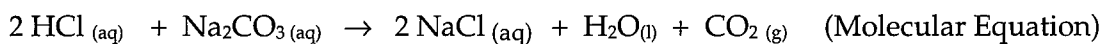
* Ionisation contrasts with the process of **DISSOCIATION** which involves *already existing ions* in an ionic compound simply separating as water surrounds and **HYDRATES** each type of ion :



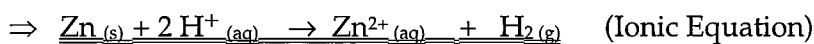
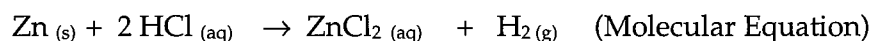
- A **STRONG ACID** is one that **IONISES FULLY** in aqueous solution. There are no molecules left unionised. eg: H_2SO_4 , HCl or HNO_3 .



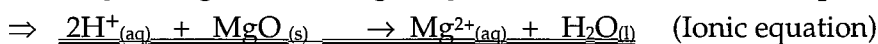
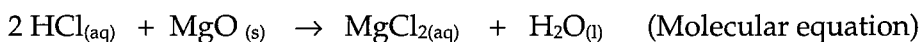
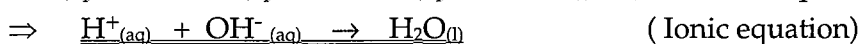
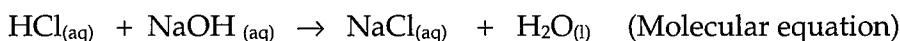
REACTIONS WITH CARBONATES:



REACTIONS WITH METALS:



REACTIONS WITH BASES: (Metal Hydroxides and oxides)

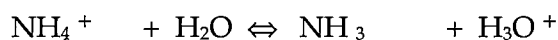
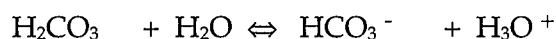


WEAK ACIDS

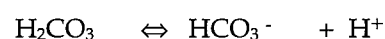
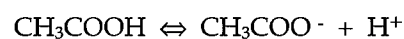
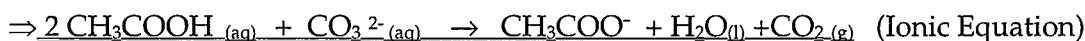
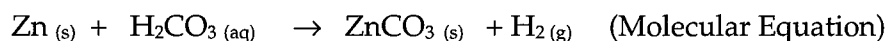
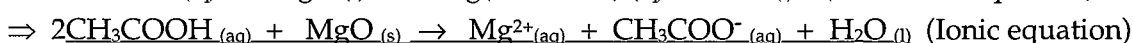
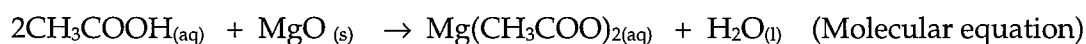
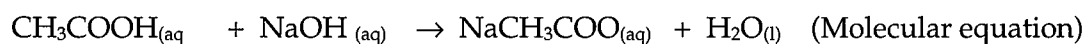
• A **WEAK ACID** is one that **DOES NOT IONISE FULLY** in aqueous solution. The ionisation reaction is an equilibrium process where both the forward and reverse reactions are occurring all the time. There will be a certain amount of both the reactants and the products in the solution at any one time.



• In general, in weak acid solutions, the major chemical species present (apart from water) is the unionised acid molecule. For this reason the equilibrium reaction are said to "favour" the left hand side. When writing ionic equations involving weak acids we always write the **FULL MOLECULAR SPECIES** rather than simplifying them to H^+ .



* These equations are commonly abbreviated:

REACTIONS WITH CARBONATES:REACTIONS WITH METALS:REACTIONS WITH BASES: (Metal Hydroxides and oxides)

Name of Acid	Acid	K _a
Sulfuric acid	H ₂ SO ₄	large
Hydrochloric acid	HCl	large
Nitric acid	HNO ₃	large
Hydronium ion	H ₃ O ⁺	55.5
Hydrogen sulfate ion	HSO ₄ ⁻	1.2 × 10 ⁻²
Phosphoric acid	H ₃ PO ₄	7.5 × 10 ⁻³
Hexaaquairon(III) ion	Fe(H ₂ O) ₆ ³⁺	6.3 × 10 ⁻³
Hydrofluoric acid	HF	7.4 × 10 ⁻⁴
Formic acid	HCO ₂ H	1.8 × 10 ⁻⁴
Benzoic acid	C ₆ H ₅ CO ₂ H	6.3 × 10 ⁻⁵
Acetic acid	CH ₃ CO ₂ H	1.8 × 10 ⁻⁵
Hexaaquaaluminum ion	Al(H ₂ O) ₆ ³⁺	7.9 × 10 ⁻⁶
Carbonic acid	H ₂ CO ₃	4.2 × 10 ⁻⁷
Hydrogen sulfide	H ₂ S	1 × 10 ⁻⁷
Dihydrogen phosphate ion	H ₂ PO ₄ ⁻	6.2 × 10 ⁻⁸
Hypochlorous acid	HClO	3.5 × 10 ⁻⁸
Ammonium ion	NH ₄ ⁺	5.6 × 10 ⁻¹⁰
Hydrocyanic acid	HCN	4.0 × 10 ⁻¹⁰
Hexaaquairon(II) ion	Fe(H ₂ O) ₆ ²⁺	3.2 × 10 ⁻¹⁰
Hydrogen carbonate ion	HCO ₃ ⁻	4.8 × 10 ⁻¹¹
Hydrogen phosphate ion	HPO ₄ ²⁻	3.6 × 10 ⁻¹³
Water	H ₂ O	1.8 × 10 ⁻¹⁶
Hydrogen sulfide ion	HS ⁻	1 × 10 ⁻¹⁹

STRONG ACIDS



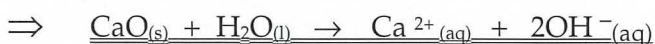
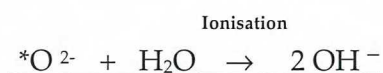
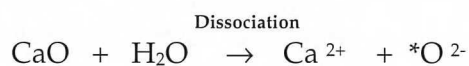
* A small number such as the K_a of Hydrocyanic acid tells us that this weak acid is indeed very weak. This means that there is only a very low concentration of ions in the solution at any one time. There is a greater concentration of unionized reactants.

STRONG BASES

- Many **STRONG BASES** are ionic hydroxides and are able to **DISSOCIATE FULLY** to produce hydroxide ions. It is rare for a covalent molecule to be basic, but if it were a strong base, it would ionise fully as would any strongly basic ion in solution.

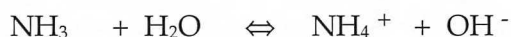


- All metal oxides are classified as strong bases due to the ionisation reaction of any aqueous oxide ions with water:



WEAK BASES

- A **WEAK BASE** is one that **DOES NOT IONISE FULLY** in aqueous solution. The ionisation reaction is an equilibrium process where both the forward and reverse reactions are occurring all the time. There will be a certain amount of both the reactants and the products in the solution at any one time.



Name of Base	Base	K_b
hydrogen sulfate ion	HSO_4^-	very small
chloride ion	Cl^-	very small
nitrate ion	NO_3^-	very small
water	H_2O	1.8×10^{-16}
sulfate ion	SO_4^{2-}	8.3×10^{-13}
dihydrogen phosphate ion	H_2PO_4^-	1.3×10^{-12}
pentaquahydroxoiron(III) ion	$\text{Fe}(\text{H}_2\text{O})_5\text{OH}^{2+}$	1.6×10^{-12}
fluoride ion	F^-	1.4×10^{-11}
formate ion	HCO_2^-	5.6×10^{-11}
benzoate ion	$\text{C}_6\text{H}_5\text{CO}_2^-$	1.6×10^{-10}
acetate ion	CH_3CO_2^-	5.6×10^{-10}
pentaquahydroxoaluminum ion	$\text{Al}(\text{H}_2\text{O})_5\text{OH}^{2+}$	1.3×10^{-9}
hydrogen carbonate ion	HCO_3^-	2.4×10^{-8}
hydrogen sulfide ion	HS^-	1×10^{-7}
hydrogen phosphate ion	HPO_4^{2-}	1.6×10^{-7}
hypochlorite ion	ClO^-	2.9×10^{-7}
ammonia	NH_3	1.8×10^{-5}
cyanide ion	CN^-	2.5×10^{-5}
pentaquahydroxoiron(II) ion	$\text{Fe}(\text{H}_2\text{O})_5\text{OH}^+$	3.1×10^{-5}
carbonate ion	CO_3^{2-}	2.1×10^{-4}
phosphate ion	PO_4^{3-}	2.8×10^{-2}
hydroxide ion	OH^-	55.5
sulfide ion	S^{2-}	1×10^5

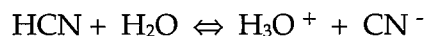


* A small number such as the K_b of sulfate ion tells us that this weak base is indeed very weak. This means that there is only a very low concentration of ions in the solution at any one time. There is a greater concentration of unionized reactants.

RELATIVE STRENGTHS OF WEAK ACIDS AND BASES (EXTENSION)

- To gain an idea of how weak a weak acid or base is, we compare the relative amounts of the *products* of ionisation with the remaining unionised *reactants*. We construct a fraction known as the "Equilibrium Constant" in which we divide the concentration of the products by the concentration of the reactants. If there are more than one reactant or product we multiply their concentrations together.

NB: The concentration of water is a constant and is assumed to be "1" so that it may be left out. In reality $[\text{H}_2\text{O}] \sim 55.5 \text{ M}$ (1 Litre of $\text{H}_2\text{O} = 1,000\text{g} = 1,000/18 = 55.5 \text{ mol L}^{-1}$)

WEAK ACID:

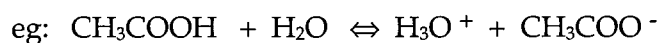
$$K = \frac{[\text{Products}]}{[\text{Reactants}]} = \frac{[\text{H}_3\text{O}^+][\text{CN}^-]}{[\text{HCN}]} = \underline{4.0 \times 10^{-10}}$$

NB: As this "K" value is for the ionisation of a weak acid we refer to it as K_a .

* We often abbreviate H_3O^+ to H^+ to simplify these expressions further!

* A small number such as the K_a of Cyanic acid tells us that this weak acid is indeed very weak as there is only a very low concentration of ions in the solution at any one time. There is a greater concentration of unionised reactants.

- The **BIGGER** the K_a the **STRONGER** the **WEAK ACID** and vice versa.



$$K_a = \frac{[\text{Products}]}{[\text{Reactants}]} = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} = \underline{1.8 \times 10^{-5}}$$

\therefore Ethanoic acid is a stronger weak acid than Cyanic acid!

**WEAK BASE:**

- The principal of "K" expressions can also be extended to weak bases to gain a measure of which is the stronger weak base.

- The **BIGGER** the K_b the **STRONGER** the **WEAK BASE** and vice versa.



$$K_b = \frac{[\text{Products}]}{[\text{Reactants}]} = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = \underline{1.8 \times 10^{-5}}$$



$$K_b = \frac{[\text{Products}]}{[\text{Reactants}]} = \frac{[\text{HSO}_4^-][\text{OH}^-]}{[\text{SO}_4^{2-}]} = \underline{8.3 \times 10^{-13}}$$

\therefore Ammonia is a far stronger weak base than sulphate ions!



ACIDS and BASES

LESSON 3

DIPROTIC AND TRIPROTIC ACIDS

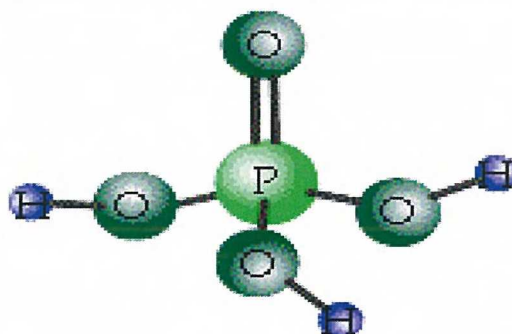
- An acid molecule or ion that has more than one hydrogen atom that is *ionisable* is termed a *multiprotic* acid. An acid with a single ionisable hydrogen is said to be **MONOPROTIC**, while one with *two* is **DIPROTIC** and in turn one with *three* is **TRIPROTIC**.
- The ionisation of multiprotic acids **always** occurs in *separate stages* where the strength of each subsequent ionisation always diminishes. The first stage ionisation is always stronger than the second and the second is always stronger than the third.



- * The reason this acid is termed **strong** is that the *1st ionisation* is **complete** even though the *2nd* is only partial. The second ionisation being, an equilibrium, will have an associated K_a value.
- The number of ionisable hydrogens *does not* give a measure of strength only the *degree of ionisation* at any one stage is important. Phosphoric acid (H_3PO_4) is a triprotic acid but is weaker than H_2SO_4 , a diprotic acid. All stages of the ionisation of Phosphoric acid are weak and have associated K_a values which decrease progressively.



$$K_{a1} > K_{a2} > K_{a3}$$

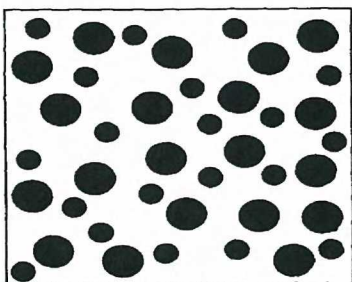


ACIDS and BASES

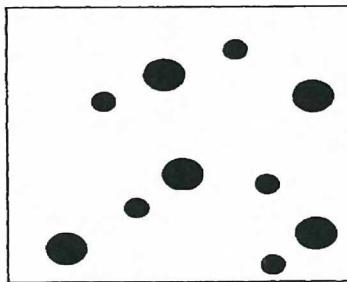
LESSON 4

CONCENTRATED AND DILUTE

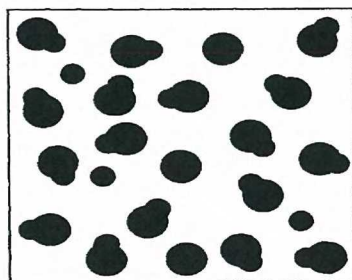
- The terms **concentrated** and **dilute** are often confused with the concept of **strong** and **weak** when referring to acids and bases.
- Concentrated and dilute describe solutions and refer to the amount of **SOLUTE** dissolved in a given **VOLUME**. A **CONCENTRATED** solution has a *large amount* of dissolved acid molecules (or basic species) *per unit of volume* irrespective of whether or not they have actually ionised.
- A **DILUTE** solution simply has only a small amount of solute acid or base per unit of volume.
- * The most dangerous acid or base solutions are those that are both **STRONG** and **CONCENTRATED**. eg: 8 M H_2SO_4 or 6 M NaOH .
- * The least dangerous acid or base solutions are those that are both **WEAK** and **DILUTE**. eg: 0.001 M CH_3COOH or 0.01 M NH_3 .



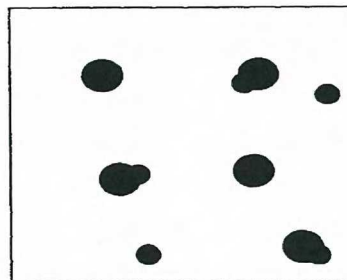
A concentrated solution of a strong acid.



A dilute solution of a strong acid.



A concentrated solution of a weak acid.



A dilute solution of a weak acid.



Key to diagrams:

	un-ionised acid HA(aq)
	hydrogen ions $\text{H}^+(\text{aq})$
	acid anions A^-

Water molecules are not shown



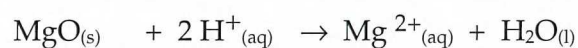
ACIDS and BASES

LESSON 5

ACID, BASE OXIDES AND HYDROXIDES

METAL OXIDES AND HYDROXIDES ⇒ BASIC

- Metal oxides and hydroxides produce aqueous OH^- . They do this by either direct **dissociation** (ionic hydroxides) or through dissociation and **ionisation** (oxide ions). *See Lesson 2, Page 3.
- **METAL OXIDES** and hydroxides are able to neutralise acids:



NON-METAL OXIDES AND HYDROXIDES ⇒ ACIDIC

- **NON-METAL OXIDES** and hydroxides produce aqueous H^+ . They do this through **ionisation** of the molecular species:

