8.1 – Theories of Acids and Bases

8.1.1 - Define acids and bases according to the Bronsted-Lowry and Lewis theories

There are two main theories that exist for classifying acids and bases: that of Bronsted and Lowry, and the Lewis theory. Both are used.

The **Bronsted-Lowry Theory** is that:

An **acid** is a substance that can donate a proton, or hydrogen ion

A **base** is a substance that can accept a proton, or hydrogen ion

$$\underset{acid}{HCl_{(g)}} + \underset{base}{H_2O_{(l)}} \leftrightarrow \underset{acid}{H_3O^+{}_{(aq)}} + \underset{base}{Cl^-{}_{(aq)}}$$

In the equation above, Cl^- is the **conjugate base** of HCl and H_3O^+ is the **conjugate acid** of H_2O. That is to say, it is what remains when the molecule has gained or lost a proton.

Interestingly, water (H_2O) is able to act as both an acid and a base in Bronsted-Lowry theory, called **amphiprotic**. Other examples include HCO_3^- and HCO_4^-. The substances are able to accept or donate a proton. Alternatively, a substance that can act as both an acid and a base is called **amphoteric**, such as Al_2O_3

The **Lewis Theory** is that:

An **acid** is an electron pair acceptor

A **base** is an electron pair donator

Therefore, all Lewis acids are deficient in an electron pair (such as H^+, BF_3 and SO_3), whilst a Lewis base must contain a non-bonding pair of electrons (such as OH^-, F^- and H_2O).

$$\begin{array}{ccccc} BF_3 & + & NH_3 & \rightarrow & BF_3NH_3 \\ acid & + & base & \rightarrow & Lewis\ acid-base\ complex \end{array}$$

BF_3 is a Lewis acid because it does not have any non-bonding pairs of electrons and only three bonding pairs.

8.1.2 - Deduce whether or not a species could act as a Bronsted-Lowry and/or a Lewis acid or base

Many acids or bases that do not fit the Bronsted-Lowry definition fit the Lewis definition, along with their reactions. Hence, substances like BF_3, SO_3, H^+ and SbF_5 are not Bronsted-Lowry acids, but are Lewis acids. Also, H_2O, OH^- and F^- are not Bronsted-Lowry bases, but are Lewis bases. Many other substances **can be classified as both**.

The reaction below could be classified as both:

$$H^+ + NH_3 \rightarrow NH_4{}^+$$

From the perspective of a **Bronsted-Lowry** acid-base reaction, the H^+ ion is being donated to the NH_3 molecule to form the new complex. As for the **Lewis** theory, the NH_3 molecule has a non-bonding pair of electrons, which are able to bond with the H^+ ion.

When classifying a reaction, the Lewis theory is usually only used if the Bronsted-Lowry theory does not apply.

8.1.3 - Deduce the formula of the conjugate acid (or base) of any Bronsted-Lowry base (or acid)

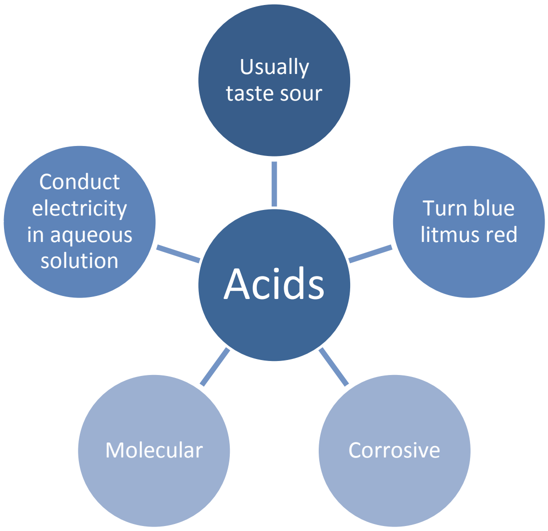
Remember that a **conjugate acid** is the complex formed after a Bronsted-Lowry base has received a proton. On the other hand, a **conjugate base** is the complex formed after a Bronsted-Lowry acid has donated a proton. Therefore, an acid will always have one more proton than its conjugate base, and vice versa.

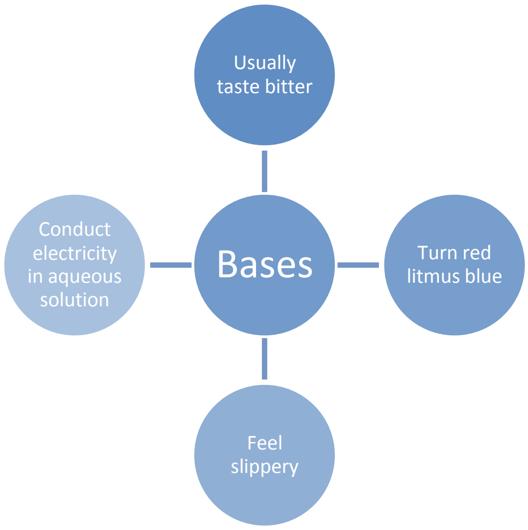
When you write the formula of the conjugate acid or base, you must always indicate where the proton is located. For example:

$$CH_3COOH \leftrightarrow CH_3COO^-$$

8.2 – Properties of Acids and Bases

8.2.1 - Outline the characteristic properties of acids and bases in aqueous solution





Note that an **alkali** is a base that dissolves in water.

Reactions of Acids

Reaction with a **metal** to form a <u>salt</u> and <u>hydrogen gas</u>. Exceptions are Cu, Hg and Ag.

$$2HCl_{(aq)} + Mg_{(s)} \rightarrow MgCl_{2\,(aq)} + H_{2\,(g)}$$

Reaction with a **metal hydroxide** to form a <u>salt</u> and <u>water</u>.

$$2HCl_{(aq)} + Mg(OH)_{2\,(s)} \rightarrow MgCl_{2\,(aq)} + 2H_2O_{(l)}$$

Reaction with a **metal oxide** to form a <u>salt</u> and <u>water</u>

$$2HCl_{(aq)} + MgO_{(s)} \rightarrow MgCl_{2\,(aq)} + H_2O_{(l)}$$

Reaction with **metal carbonate** to produce a <u>salt</u>, <u>water</u> and <u>carbon dioxide</u>

$$2HCl_{(aq)} + MgCO_{3\,(s)} \rightarrow MgCl_{2\,(aq)} + H_2O_{(l)} + CO_{2\,(g)}$$

Reaction with a **metal hydrogen carbonate** to form a <u>salt</u>, <u>water</u> and <u>carbon dioxide</u>

$$2HCl_{(aq)} + Mg(HCO_3)_{2\,(s)} \rightarrow MgCl_{2\,(aq)} + 2H_2O_{(l)} + 2CO_{2\,(g)}$$

Reactions of Bases

Reaction of an **alkali** with an **acid** to form a <u>salt</u> and <u>water</u>

$$NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

Reaction of a **metal oxide** with an **acid** to form a <u>salt</u> and <u>water</u>

$$MgO_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2\,(aq)} + H_2O_{(l)}$$

Reaction of a **metal hydrogen carbonate** with an **acid** to product a <u>salt</u>, <u>water</u> and <u>carbon dioxide</u>

$$Na_2CO_{3\,(s)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + H_2O_{(l)} + CO_{2\,(g)}$$

Reaction of a **hydrogen carbonate** with an **acid** to form a <u>salt</u>, <u>water</u> and <u>carbon dioxide</u>

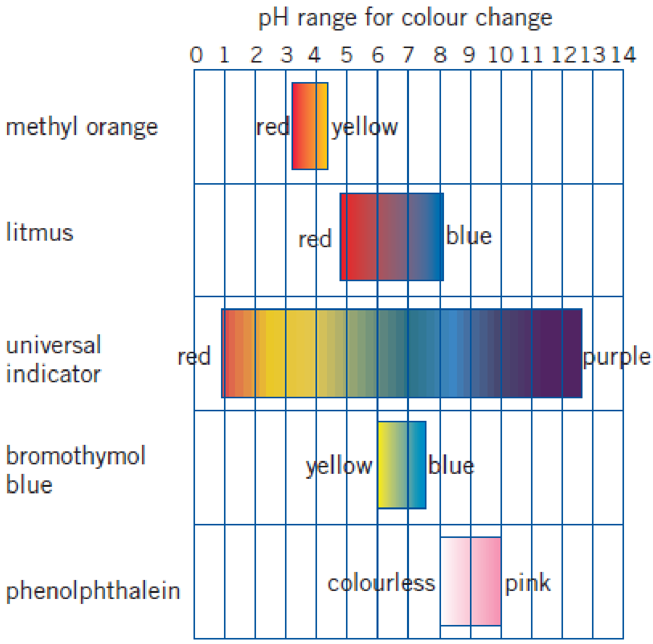
$$NaHCO_{3\,(s)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)} + CO_{2\,(g)}$$

Reaction of **ammonia** with an **acid** to produce an <u>ammonium salt</u>

$$NH_{3\,(aq)} + HCl_{(aq)} \rightarrow NH_4Cl_{(aq)}$$

Response to Indicators

Indicators change colour depending on the pH. Acids will cause the pigment for show a different colour from bases.



8.3 – Strong and Weak Acids and Bases

8.3.1 - Distinguish between strong and weak acids and bases in terms of the extent of dissociation, reaction with water and electrical conductivity

Strong Acid/ Base

- Almost completely dissociate
- React rapidly with water
- Higher electrical conductivity

Weak Acid/ Base

- Only Partially dissociate
- React slowly with water
- Lower electrical conductivity

8.3.2 - State whether a given acid or base is strong or weak

Examples of **strong acids** include:

- Hydrochloric acid - HCl
- Nitric acid - HNO_3
- Sulfuric acid - H_2SO_4

Examples of **strong bases** include:

- Sodium hydroxide - $NaOH$
- Potassium hydroxide - KOH
- Barium hydroxide - $Ba(OH)_2$

Examples of **weak acids** include:

- Ethanoic acid - CH_3COOH
- Carbonic acid - H_2CO_3

Examples of weak bases include:

- Ammonia - NH_3
- Aminoethane - $C_2H_5NH_2$

In a reaction with water, strong acids and bases will completely dissociate, so the reaction is shown to go to completion:

$$HCl + H_2O \rightarrow H_3O^+ + Cl^-$$
$$KOH + H_2O \rightarrow K^+ + OH^-$$

On the other hand, weak acids and bases reacting with water are shown to be at equilibrium, with only partial dissociation.

$$CH_3COOH + H_2O \leftrightarrow CH_3COO^- + H_3O^+$$
$$NH_3 + H_2O \leftrightarrow {NH_4}^+ + OH^-$$

One mole of a **monoprotic acid**, such as HCl, produces one mole of H^+ ions. One mole of a **diprotic acid**, such as H_2SO_4, produced two moles of H^+ ions.

8.3.3 - Distinguish between strong and weak acids and bases, and determine the relative strengths of acids and bases, using experimental data

pH measurement

The pH of a strong acid will be lower than the pH of a weak acid. Likewise, the pH of a strong base will be higher than the pH of a weak base. This can be found using a pH meter or an indicator substance.

Conductivity measurement

A strong acid or base will have a higher reading on a conductivity meter than a weak acid or base, provided they are at equal concentrations, or **equimolar**, since they have more ions in the solution.

<u>Sidenote:</u>

Some exams might ask you about the cause of **acid rain**. You should be aware that the atmosphere contains non-metal oxides (i.e. CO_2), which form acids when they are dissolved in water. An example of this is carbonic acid, which is formed from CO_2 and water. However, rain is only defined as acidic if the pH is below 5. This is most common around polluted areas, where there is increased output of CO_2 and other such gases. These may also come from volcanoes.

Acid rain is harmful to aquatic species because the pH of their environment may change, plants because the pH of the soil changes, and buildings because the acid erodes metal and stone.

8.4 – The pH Scale

8.4.1 - Distinguish between aqueous solutions that are acidic, neutral or alkaline using the pH scale

The pH scale is used to give a scale of how acidic or alkaline a solution is based on the concentration of H^+ ions in solution. Any solution with a $pH < 7$ is acidic. A solution with a pH of 7 is neutral, and a solution with a $pH < 7$ is alkaline.

These concentrations will usually have negative exponents, such as 6.82 x 10^{-9} mol dm^{-3}. Using this as the comparison point would prove impractical, so the concentration is converted into another value according the equation:

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

This gives a positive value such as 9.2 or 3.7, which can be more easily used to compare solutions. The equation can be rearranged to give:

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

8.4.2 - Identify which of two or more aqueous solutions is more acidic or alkaline using pH values

When two solutions are compared, the one with the largest pH value is more alkaline, and the one with the smallest pH value is the most acidic. In order to determine pH, indicator substances or pH meters can be used.

Universal Indicator

If the indicator paper is used, the pH is determined based on the colour the paper turns when placed in the solution. Different indicators can be used to accurately determine the pH, although the problem arises that people interpret colours differently.

Alternatively, indicator solution can be dropped into the solution, allowing the entire solution to change colour. The pH is likewise determined by the new colour of the solution.

pH Meter

This reads the concentration of H^+ ions through an electrode, giving the pH with an accuracy of a few decimal points. This method is more objective than using indicator solutions and is more accurate.

8.4.3 - State that each change of one pH unit represents a 10-fold change in the hydrogen ion concentration

The concentration of H^+ ions in a solution and the pH of the solution have an inverse relationship. A change of one pH unit equates to a ten-fold change in $[H^+]$ because the scale is logarithmic. If the pH increases, the $[H^+]$ decreases, and vice versa.

8.4.4 - Deduce changes in $[H^+_{(aq)}]$ when the pH of a solution changes by more than one pH unit

If the pH increases, the $[H^+]$ decreases, and vice versa.

The pH of a solution changes from 2 to 6, deduce how the $[H^+]$ changes

pH = 2

$\therefore [H^+] = 10^{-2}$ mol dm^{-3}

pH = 6

$\therefore [H^+] = 10^{-6}$ mol dm^{-3}

So, the pH has changed by 10^{-4}, a decrease of 10000