Chemistry Notes: Acids and Bases

Arrhenius base is one that dissociates in water to form hydroxide ions.

Arrhenius acid is on that dissociates in water to form hydrogen ions (protons).

A Bronsted-Lowry acid is one that donates a hydrogen ion in a reaction.

A Bronsted-Lowry base is one that accepts a hydrogen ion in a reaction.

The act of ion donation cannot happen in isolation, there must first be a species that is capable of receiving an ion (playing the role of the acceptor). In the following equation:

$$HA + B \leftrightarrows A^{-} + BH^{+}$$

HA is the acid of the forward reaction because it donates a proton (H^+) to B thus forming BH^+. In the reverse reaction, BH^+ acts as the acid, and A^- acts as the base because the former loses a proton and the later gains one.

Conjugate Pairs are two species in the same reaction one of which is a base and the other of which is an acid. They differ by exactly on proton (H^+). In the reaction above,

HA and A^- is one conjugate pair, and BH^+ and B is the other. These are usually present in any Bronsted-Lowry acid-base reaction.

To form a conjugate acid, add on H^+ to the conjugate base's formula, to form a conjugate base, remove one H^+ ion from the conjugate acid. For example:

$H_2O \text{ and } H_3O^+ \text{are part of a conjugate pair.}$

When removing the hydrogen ion be careful to also subtract +1 from the charge.

Water can act as both an acid and a base depending on the reaction. Substances that can take the form of both acids and bases are called amphoteric or amphiprotic.

Bronsted-Lowry bases must have a lone pair of electrons to accept the donated H^+ ion. Using this rationale, Gilbert Lewis changed the definition of an acid and a base to:

A Lewis acid is an electron pair acceptor.

A Lewis base is an electron pair donator.

Lewis's definition no longer restricts the actions of an acid or base to the donation of accepting of an H^+ ion, and broadens it to any atom with an unfilled outer energy orbital.

Lewis Acid-Base reactions result in the formation of a covalent dative bond (because both electrons come from the base).

Many transition elements can act as Lewis bases because they often form ions with vacant orbitals and as such are able to accept a lone pair of electrons.

The Lewis Base is called the **ligand**. It is the lone pair donator. These ligands often surround the ion (Lewis acid) in a fixed number ratio. Dative bonds form, and a complex ion results. In the following case:

$$Cu^{2+} + 4H_2O \rightarrow [Cu(H_2O)_4]^{2+}$$

Copper (II) is the Lewis acid because it accepts a lone pair of electrons from each of the four water molecules that surround it. The water molecules are the ligands, and the result is a complex ion (remember that Lewis bases have lone pairs).

All Bronsted-Lowry acids are Lewis acids, but not all Lewis acids are Bronsted-Lowry acids. In general, Lewis acids are only acids that cannot release H^+ ions. Bronsted-Lowry acid-base reactions must involve H^+ transfer (these reactions can also be classified as Lewis acid-base reactions). However, if there is no H^+ transfer, it is not a Bronsted-Lowry acid-base reaction.

Metal Oxides and Hydroxides neutralize acids to produce water. **Alkalis** are bases that dissociate in water to form hydroxide ions.

Alkalis are bases, but not all bases are alkalis. Bases accept H^+ ions but not all bases dissociate in water to produce hydroxide ions.

Acid-Base Indicators change color in reversibly according to the concentration of hydrogen ions in a given solution.

Litmus is a dye derived from lichens which turns pink in the presence of an acid and blue in the presence of alkalis. Litmus is not useful for determining the strengths of acids or alkalis.

(Other indicators listed in table 16 of the IB data booklet)

Universal Indicator is formed by mixing together several indicators. Thus, the indicator changes many times across a range of different acids and alkalis. The wide range can thus be used to measure the concentration of hydrogen ions on the pH scale.

Acids react with metals, bases, and carbonates to forms salts. The term **salt** refers to a compound formed when hydrogen is **replaced** by a metal or other positive ion. A **parent acid** is the acid from which the hydrogen atom is taken and replaced (HCl is a common example of a parent of NaCl).

Three main reaction types by which acids form salts:

$$acid + metal \rightarrow salt + hydrogen\ (2HCl + Zn \rightarrow ZnCl_2 + H_2)$$

The net ionic equation for this reaction is:

$$2H^+ + Zn \rightarrow Zn^{2+} + H_2\ (chlorine\ was\ a\ spectator\ ion)$$

There is a big range of reactivity of metals in this reaction. Less reactive metals are more resistant to corrosion. The common nitric acid (HNO_3) does react with metals but generally does not produce hydrogen gas.

$$acid + base \rightarrow salt + water\ (HCl + NaOH \rightarrow NaCl + H_2O)$$

These are also known as **neutralization reactions**. All neutralization reactions are represented by the common net ionic equation:

$$H^+ + OH^- \rightarrow H_2O$$

This reaction is used to calculate the exact concentration of an acid or an alkali when the other is known (**titration**). This involves reacting together a carefully measure volume of one of the solutions and adding the other solution gradually until the **equivalence point** is reached. This point is reached when the two solutions neutralize each other. The entire process is known as **standardization**.

$$acid + carbonate \rightarrow salt + water + carbon\ dioxide\ (2HCl + CaCO_3 \rightarrow CaCl_2 + H_2O + CO_2)$$

The net ionic reaction is:

$$H^+ + CO_3{}^{2-} \rightarrow H_2O + CO_2$$

These reactions involve gas being given off through the process of **effervescence** (visible production of bubbles).

Strong and Weak Acids and Bases

Reactions of acids and bases depend on the fact that they dissociate in solution. Acids produce hydrogen ions and alkalis/bases produce hydroxide ions. Their aqueous solutions exist as equilibrium mixtures containing **both the un-dissociated acids and the ions**. The position of this equilibrium is what defines the strength of the acid or base.

In the equation

$$HA \leftrightarrows H^+ + A^-$$

If the equilibrium lies to the right, the HA acid is considered a strong acid. If it lies to the left, it is considered a weak acid. This is basically measure of how readily the acid dissociates. The more it dissociates, the strong the acid. **Strength has nothing to do with concentration of the solution, it is the concentration of the ions that matters**.

It is possible for an acid to be strong but present in a dilute solution. It is also possible for an acid to be weak and present in a concentrated solution.

Weak acids and bases are more common that strong acids and bases, thus it is easy to recognize strong and weak acids and bases.

Strong Acids	Strong Bases	
HCl (hydrochloric acid)	LiOH (lithium oxide)	$Ba(OH)_2$ (barium hydroxide)
HNO_3 (nitric acid)	NaOH (sodium hydroxide)	
H_2SO_4 (sulfuric acid)	KOH (potassium hydroxide)	

Weak Acid

CH_3COOH and other organic acids (ethanoic acid)

H_2CO_3 (carbonic acid)

H_3PO_4 (phosphoric acid)

Weak Bases

NH_3 (ammonia)

$C_2H_5NH_2$ and other amines (ethylamine)

Strong acids and bases dissociate almost completely in solution; weak acids dissociate only partially in solution.

Strong acids and bases contain a higher concentration of ions than weak acids and bases do. Note that these **comparisons area only valid** when the **solutions are of the same concentration and temperature**.

The following properties depend on the concentration of ions in the solution:

- Electrical Conductivity:

Electrical conductivity depends on the concentration of mobile ions. **Strong acids and bases** will have **higher conductivity** than weak acids and bases.

- Rate of Reaction:

The higher the concentration of H^+ ions, the faster the reactions occur. Therefore **reactions happen at a faster rate with strong acids**.

- pH:

A **higher the hydrogen ion concentration** yields a **lower** the **pH value**.

The pH scale:

$$pH = -\log[H^+] \; therefore,$$

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

pH numbers are usually positive and have no units.

Normal range is 0-15, corresponding to the hydrogen ion concentration of

1 mol dm^{-3} to 10^{-15} mol dm^{-3}

The numbers are inversely related, higher pH = lower hydrogen ion concentration and vice versa.

A change of on pH represents a 10 fold change in hydrogen ion concentration.

The pH scale is logarithmic and therefore compresses a wide range of hydrogen ion concentration values into a small range of numbers.

The relationship between the concentration of hydroxide and hydrogen ion concentration is inverse in aqueous solutions. Solutions with high hydrogen ion concentration have lower hydroxide ion concentration, and vice versa.

In acidic solutions $[H^+] > [OH^-]$ and the $pH < 7$

In alkaline solutions $[H^+] < [OH^-]$ and $pH > 7$

In neutral solutions $[H^+] = [OH^-]$ and $pH = 7$

… at 25 degrees Celsius.

Measuring pH:

- Universal indicator on paper or solution. Substance tested gives characteristic color. Narrower range indicators give more accurate results.

- pH meter that reads hydrogen ion concentration through special electrode. Must be calibrated using buffer solution and standardized for temperature.