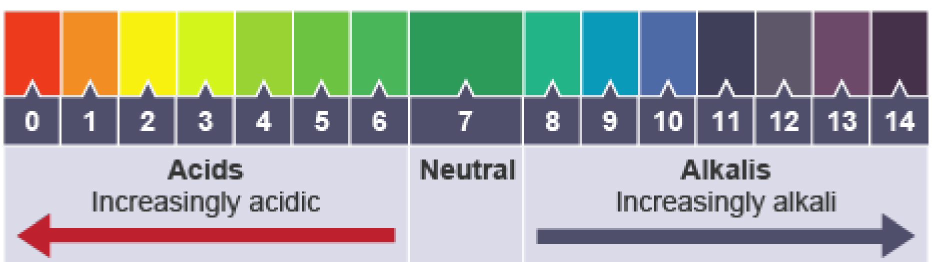
Acids and Bases

- Acids exist naturally in foods and plants
- Examples of acids include: hydrochloric acid, ethanoic acid, acetic acid, sulfuric acid, nitric acid etc.
- All acids form hydrogen ions, H^+, when dissolved in water. It is the hydrogen ions that make a solution acidic
- Alkalis are the opposite of acids when they react
- Examples of alkalis include: sodium hydroxide, calcium hydroxide and ammonia ‘
- Alkalis form hydroxide ions, OH^-, when they dissolve in water
- Substances that are neither acidic nor alkaline are neutral
- Pure water is neutral

The pH Scale

- The pH scale is used to show whether a solution is acidic, alkaline or neutral
- The pH scale runs from 0 to 14
- Acids have a pH below 7
- Alkalis have a pH above 7
- A solution with a lower pH has a higher concentration of hydrogen ions
- The higher the pH, the more alkaline the solution is
- A higher pH means a higher concentration of hydroxide ions
- We can measure pH using universal indicator paper or a pH electrode connected to a pH meter

Using Universal Indicator



- An acid-alkali indicator is a chemical which changes colour when we add an acid or an alkali
- Universal indicator is a mixture of indicators which shows a range of colours depending on the pH

The litmus test

- Litmus is an indicator that has two colours, red and blue
- If a solution is acidic it turns blue litmus red
- If a solution is alkaline it turns red litmus blue
- If a solution is acidic red litmus paper remains red
- If a solution is alkaline blue litmus paper remains blue
- The simplest definition of an acid states that an acid is a substance that dissolves in water to form hydrogen ions

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

- We can show that water plays an important part by first dissolving hydrogen chloride gas in a solvent called methylbenzene
- The solution does not turn blue litmus red
- This shows that the solution is not acidic
- When we shake this solution with water the litmus does turn red
- This shows that an acid has been formed once water has been added

<u>Chemical properties of acids</u>

Reaction of acids and metals

- Many metals react with dilute acids to form a salt and hydrogen
- However, nitric acid does not always give off hydrogen with metals unless it is very dilute

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

- The reaction of an acid with a metal is an example of a redox reaction. The metal loses electrons to form metal ions. The metal is oxidised. Hydrogen ions gain electrons to form hydrogen gas. The hydrogen ions are reduced

Reaction of acids with metal oxides

- Acids react with many metal hydroxides to form a salt and water

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

Reaction of acids with metal hydroxides or aqueous ammonia

- Acids react with many metal hydroxides to form a salt and water

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

- This is an example of a neutralisation reaction

Reaction of acids with carbonates

- Carbonates react with acids to form a salt, water and carbon dioxide. The carbon dioxide bubbles off as a gas

$$CuCO_3(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O\ (l) + CO_2(g)$$

Hydrogencarbonates react in a similar way:

$$NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g)$$

- Sodium hydrogencarbonate is used in baking powder together with powdered tartaric acid
- In the presence of moisture, carbon dioxide is given off which makes the cake rise

<u>Neutralisation</u>

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

$$\cancel{Na^+}(aq) + OH^-(aq) + H^+(aq) + \cancel{Cl^-}(aq) \rightarrow \cancel{Na^+}(aq) + \cancel{Cl^-}(aq) + H_2O(l)$$

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

- Any combination of an acid and alkali neutralisation reaction will produce the above ionic equation

- The reasons for this being: every acid has hydrogen ions dissolved in water and every alkali has hydroxide ions dissolved in water
- So these two ions combine to make water

Acids, bases and protons

- A hydrogen ion is formed by the removal of a single electron from a hydrogen atom
- Nearly all hydrogen atoms have a single proton and no neutrons in their nuclei
- So a hydrogen ion is nothing more than a proton
- We define acids and bases more generally by seeing what happens to hydrogen ions when acids react with bases
- An acid is a proton donor
- A base is a proton acceptor

Strong and weak acids

- A strong acid is completely ionised when it dissolves in water. There are no molecules of the acid present
- Weak acids are generally organic acids such as citric acid, ethanoic acid
- A weak acid is only partly ionised when it dissolves in water
- There are lots of acid molecules present but very few ions
- The electric conductivity of a solution depends on the number of ions present, the more ions, the greater the conductivity
- A strong acid conducts better that a weak acid of the same concentration
- This is because there is a greater concentration if ions in the strong acid compared to the weak acid
- A stronger acid has a lower pH than a weak acid of the same concentration
- This is because there is a greater concentration of ions in the strong acid as compared with the weak acid
- A strong acid reacts faster than a weak acid
- This is also because there is a greater concentration of hydrogen ions in the strong acid compared with the weak acid

Strong and weak bases

- All hydroxides of alkali metals are strong bases. Strong bases are completely ionised in water
- Ammonia is a weak base because it is only partly ionised in water
- Compared with strong bases, weak bases have:
    i. A lower electrical conductivity
    ii. A less alkaline (lower) pH
    iii. A slower rate of reaction

Oxides

- Oxides are compounds of metals or non-metals with oxygen

Basic oxides

- Most metal oxides are basic oxides
- Many basic oxides are formed by the direct combination of a metal with oxygen
- Basic oxides react with acids to form a salt and water

$$CaCO\ (s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l)$$

- Basix oxides do not react with alkalis
- Many basic oxides do not react with water
- But those from Group I and many from Group II in the Periodic table form a metal hydroxide
- An alkaline solution is formed which turns red litmus blue

$$BaO(s) + H_2O(l) \rightarrow Ba(OH)_2$$

Acidic oxides

- Most non-metal oxides are acidic oxides. Many are formed by direct reaction with oxygen
- Acidic oxides react with alkalis to form a salt and water

$$CO_2(g) + 2NaOH(aq) \rightarrow Na_2CO_3(aq) + H_2O(l)$$

- Some acidic oxides react with bases such as metal oxides when heated

$$SiO_2 + CaO \rightarrow CaSiO_3$$

- Many acidic oxides react with water to form acidic solutions

$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$$

Neutral oxides

- Neutral oxides do not react with acids or bases
- Examples of neutral oxides are: nitrogen (I) oxide, N_2O nitrogen oxide, NO, and carbon monoxide CO.
- Most are lower oxides of non-metals
- Example: carbon monoxide is a neutral oxide but carbon dioxide is an acidic oxide

Amphoteric oxides

- Amphoteric oxides have both acidic and basic oxides
- The oxides of aluminium and zinc are examples
- They form salts when they react with acids
- They react with alkalis to form complex salts

$$ZnO(s) + 2HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + H_2O(l)$$

$$ZnO(s) + 2NaOH(aq) \rightarrow Na_2ZnO_2(aq) + H_2O(l)$$

$$Al_2O_3(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2O(l)$$

$$Al_2O_3(s) + 2NaOH(aq) \rightarrow 2NaAlO_2(aq) + H_2O(l)$$

- Zincates and aluminates have the ending –ate