RAFFLES INSTITUTION RAFFLES PROGRAMME - YEAR THREE CHEMISTRY

ACIDS AND BASES

===

ACID: - A substance that produces hydrogen H+ ions in an aqueous solution / dissociates in water to produce hydrogen H+ ions (Arrhenius Model)

- A substance that donates H+ ions (Bronsted-Lowry Model)

BASE: - A substance that produces hydroxide OH- ions in an aqueous solution / dissociates in water to produce hydroxide OH- ions (Arrhenius Model)

- A substance that accepts H+ ions (Bronsted-Lowry Model)

Chemistry of Acids (ABCMS)

1) Acid - Base Reaction (Neutralisation):

Acid + Base → Salt + Water H_2O

2) Acid - Carbonate Reaction:

Acid + Carbonate → Salt + Water H_2O + Carbon Dioxide CO_2

3) Acid - Reactive Metal Reaction:

Acid + Reactive Metal → Salt + Hydrogen Gas H_2

4) Acid - Sulfite SO_3^{2-} Salt Reaction:

Acid + Sulfite Salt → Salt + Water H_2O + Sulfur Dioxide SO_2

! Reminder ! - ide indicates anion of pure element. Think chloride, iodide, and sulfide.

- ate indicates polyatomic ion with more oxygen. Think sulfate SO_4^{2-}, nitrate NO_3^-.
- ite indicates polyatomic ion with less oxygen. Think sulfite SO_3^{2-}, nitrite NO_2^-.

Chemistry of Bases (BAAA)

1) Acid - Base Reaction (Neutralisation):

Acid + Base → Salt + Water H_2O

2) Base - Ammonium NH_4^+ Salt Reaction:

Base + Ammonium Salt → Salt + Water H_2O + Ammonia Gas NH_3

3) Alkali - Aqueous Salt:

Alkali + Aqueous Salt → Salt + Insoluble Hydroxide

Generally, reaction between any 2 aqueous/soluble reactants produces an aqueous solution and an insoluble salt as a precipitate, and this is called a precipitation. Specifically, if one of the aqueous reactants is basic in nature, then it is an alkaline.

Most alkalis are hydroxides and not oxides, hence in this specific precipitation reaction with an alkaline, the precipitate produced is a hydroxide.

Strength, Concentration and Basicity of Acids

Strength: The extent to which an acid dissociates/ionizes in water to produce hydrogen H+ ions.

Strong Acids dissociate/ionize in water completely/fully to produce hydrogen H+ ions.

(Hydrochloric Acid HCl, Sulfuric Acid, H2SO4, Nitric Acid HNO3)

Weak Acids dissociate/ionize in water partially to produce hydrogen H+ ions.

(Phosphoric Acid H3PO4, Ethanoic Acid CH3COOH, Carbonic Acid H2CO3)

Note that for the reaction of a weak acid with water, a reversible arrow is used.

Concentration: the number of molecules or ions in a given volume of substance (moldm-3 / M)

Concentrated Acid Solutions have more acid molecules for a given volume (Bigger moldm-3 / M).

Dilute Acid Solutions have less acid molecules for a given volume (Bigger moldm-3 / M).

Basicity: The number of H atoms in a molecule that are able to ionize to form Hydrogen H+ ions.

Monobasic Acids provide 1 H+ ion per acid molecule.

(Hydrochloric Acid HCl, Ethanoic Acid CH3COOH, Nitric Acid HNO3)

Dibasic Acids provide 2 H+ ions per acid molecule.

(Sulfuric Acid H2SO4, Carbonic Acid H2CO3)

Tribasic Acids provide 3 H+ ions per acid molecule.

(Phosphoric Acid H3PO4)

Note that non-monobasic acids (i.e. dibasic or tribasic acids) have more possible salts as different number of H+ ions are replaced. (eg. H2SO3 + Na+ → NaHSO3 or Na2SO3)

Strength, Concentration (of the acid) and Basicity are completely unrelated concepts, they do not affect each other at all. The only similarity they have is that they each can affect the concentration of H+ ions in the acid and hence the pH of the acid.

Types of Oxides (BAM NAN)

Generally, basic, amphoteric and neutral oxides are insoluble in water while acidic oxides are soluble in water.

BAM - Metal oxides are either Basic or Amphoteric.

Basic Oxides react with acids and are mostly insoluble in water.

(Sodium Oxide Na_2O, Calcium Oxide CaO, Copper (II) Oxide CuO)

Amphoteric Oxides react with both acids and bases and are all insoluble in water.

(ZAP - Zinc Oxide ZnO, Aluminum Oxide Al_2O_3, Lead (II) Oxide PbO, Gallium Oxide Ga_2O_3 not in syllabus)

NAN - Non-metal oxides are either Acidic or Neutral.

Acidic Oxides react with bases and are mostly soluble in water.

(Carbon Dioxide CO_2, Sulfur Dioxide SO_2, Nitrogen Dioxide NO_2)

Neutral Oxides do not react with both acids and bases and are mostly insoluble in water.

(Carbon Monoxide CO, Nitric Oxide NO, Water H_2O)

To remember, acidic oxides have higher oxidation states while neutral oxides have lower oxidation states.

pH: Power of Hydrogen

pH of a solution is a quantitative measure of the extent of alkalinity or acidity of a solution.

pH = -lg[H+]   the negative log of concentration of H+ ions in moldm-3

[H+] = 10^{-pH}

Acidic solutions: [H+] > [OH-]

Alkaline solutions: [H+] < [OH-]

Neutral solutions: [H+] = [OH-]

(The neutral pH is not always 7; it fluctuates with temperature. However at rtp it is 7)

Acid Rain

Sources: Sulfur Dioxide Gas (SO_2) and Oxides of Nitrogen (NO_x) react with water and oxygen to form strong acids.

SO_2 produced in power stations and motor vehicles; oxides of nitrogen produced in car engines.

($2SO_2 + 2H_2O + O_2 \rightarrow 2H_2SO_4$ $4NO_2 + 2H_2O + O_2 \rightarrow 4HNO_3$)

Effects: These acids damage lungs, causing bronchitis. They also damage metal bridges and stone buildings. They also lower pH of water bodies, killing aquatic life.

Solutions: To minimise SO_2 pollution, remove S from fossil fuels before burning, or remove SO_2 from waste gases using Calcium Carbonate ($CaCO_3 + SO_2 \rightarrow CaSO_3 + CO_2$). To minimise NO_x pollution, fit cars with catalytic converters which utilize catalysts that are only required in small amounts and can be reused many times. ($2NO + 2CO \rightarrow N_2 + 2CO_2$)

For acidic soils or lakes, try to use $CaCO_3$, because it is insoluble and excess can be cleared easily. Also, it is not alkaline and will not increase pH beyond 7 if excess is added.

Titration

A titration is used to find the concentration of a reactant by reacting a known volume of it with another reactant of known volume and concentration.
The equivalence point is when the acid and base in the titration are in the right proportions to neutralize each other.

An indicator is used during titration to observe when the right proportions of acid and base are present. The indicator changes colour gradually over a range of pH values. Hence for the indicator to be used, the pH range where the indicator changes colour must fall within sharp drop of pH at one single volume. (Recall that pH is the negative log H+ ions, so a single drop of reactant can change pH drastically.)
The endpoint is when the indicator changes colour, and most of the time the equivalence point and endpoint are the same, hence these two are often used interchangeably. It is still important to know the difference between them though.

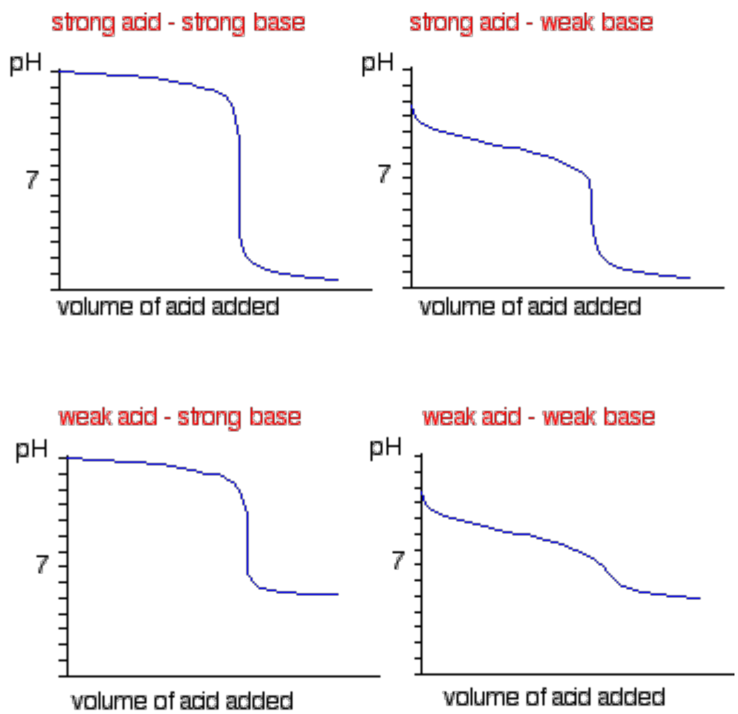
Back titration (the titration we had for our practical test) is a type of titration where two reactions are present; the titration being the second reaction. Use the set and measured values to calculate for the second reaction, and then use the calculated value from the second redaction to find the required value in the first reaction. What is important here is to actually identify and understand what is going on here; once you identify the two reactions and what you need to calculate, follow the usual steps for a titration.

For titration curves, there are several things to take note of:

1) What is being titrated into what. This will affect your starting and ending pH. (strong/weak acid/base)

2) The relative strength of the two reactants. This will affect the pH of the equivalence point, as well as whether a region of gradual pH change is present in the graph.

3) The mol ratio of the reaction. This will affect what volume the equivalence point is present at.



Precipitation: The formation of a solid/insoluble product when two aqueous solutions are reacted. As stated before, this is the general case. A more specific case is the Alkali – Aqueous salt reaction under Chemistry of Bases

For experimental type questions, to get the insoluble salt/precipitate, filter the mixture, then wash the residue with distilled water before pressing dry between sheets of filter paper.

Ionic Equations: Equations that show specifically which ions take part in the reaction.

Number of atoms and net charge on both sides must be balanced.

1) Write the balanced chemical reaction with state symbols.

2) Separate the free ions (in the aqueous solutions).

3) Cancel away corresponding spectator ions (ions which remain unchanged throughout the reaction, same ion and same state).

4) Write down the uncancelled formulas as the ionic equation.

Below are the general ionic equations for different reactions.

Neutralization: Ions of water	$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
Precipitation: Ions of precipitate	$Na^+(aq) + Cl^-(aq) \rightarrow NaCl(aq)$
Acid-Metal: H+ ions and Metal	$Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$

Acid-Carbonate: H+, Carbonate ions $2H^+(aq) + CO_3^{2-}(aq) \rightarrow H_2O(l) + CO_2(g)$

Displacement: Ions displacing/displaced $Zn(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Zn^{2+}(aq)$

However it is still good practice to do the 4 steps for ionic equations and only using the general ionic equations to check to be as safe as possible.