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Acids, Bases and Salts

- **THE THEORY of ACIDS and ALKALIS** and a few technical terms:
  - **Acids** are substances that **form hydrogen ions** ($H^+_{(aq)}$) when dissolved in water **eg** hydrochloric acid **H**Cl gives $\mathbf{H^+_{(aq)}}$ and $Cl^-_{(aq)}$ ions, sulphuric acid $\mathbf{H_2}SO_4$ gives $\mathbf{2H^+_{(aq)}}$ and SO_4^{2-} ions and nitric acid **H**NO_3 gives $\mathbf{H^+_{(aq)}}$ and $NO_3^-{}_{(aq)}$ ions.
  - **Alkalis** are substances that **form hydroxide ions** ($\mathbf{OH^-_{(aq)}}$) in water eg sodium hydroxide Na**OH** gives $Na^+_{(aq)}$ and $\mathbf{OH^-_{(aq)}}$ ions, calcium hydroxide Ca(**OH**)$_2$ gives $Ca^{2+}_{(aq)}$ and $\mathbf{2OH^-_{(aq)}}$ ions. Note: **an alkali is a base soluble in water**.
  - In **water**, there are trace quantities of H^+ and OH^- ions BUT they are of equal concentration and so water is neutral.
  - In **acid solutions** there are **more H^+ ions than OH^- ions**.
  - In **alkaline solution** there are **more OH^- ions than H^+ ions**.
  - Acids dissociate to different extents in aqueous solution. Acids that dissociate to a large extent are strong electrolytes and **strong acids**. In contrast, acids that dissociate only to a small extent are **weak acids** and weak electrolytes

    In a similar manner, bases can be strong or weak depending on the extent to which they dissociate and produce OH^- ions in solution. Most metal hydroxides are strong electrolytes and **strong bases**. Ammonia, NH_3, is a weak electrolyte and **weak base**.

$$NH_3(g) + H_2O(l) \rightleftarrows NH_4^+(aq) + OH^-(aq)$$

  - **BASES** eg oxides and hydroxides are substances that **react and neutralise acids to form salts and water**. **Bases which are soluble in water are called alkalis**.

Acids

Some common acids are listed below:

Name	Formula	Strong/Weak	Where is it found?
Hydrochloric acid	HCl	Strong	The stomach, in the lab.
Sulphuric acid	H_2SO_4	Strong	Acid rain, car batteries, the lab.
Nitric acid	HNO_3	Strong	Acid rain, in the lab.
Ethanoic (acetic) acid	CH_3COOH	Weak	Vinegar
Methanoic (formic) acid	HCOOH	Weak	Ant & nettle stings, descalers
Citric Acid	$C_6H_8O_7$	Weak	Citrus fruits

- Acids taste sour (e.g. vinegar, lemon juice).
- Acids are harmful to living cells.
- Aqueous solutions of all acids contain **hydrogen ions, H^+**.

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Bases

Most bases are oxides or hydroxides of metals. Not all bases fit into these categories, however (e.g. ammonia). Some examples of bases are shown below:

Name	Formula	Where is it found?
Sodium hydroxide (caustic soda)	NaOH	Oven cleaners, in the lab.
Calcium hydroxide	$Ca(OH)_2$	Soil lime, limewater
Magnesium oxide (magnesia)	MgO	Indigestion tablets
Calcium carbonate	$CaCO_3$	Limestone, soil lime
Sodium hydrogencarbonate (bicarbonate)	$NaHCO_3$	Baking powder
Ammonia	NH_3	Cleaning fluids, in the lab.

- Soluble bases are known as **alkalis**.
- Aqueous solutions of alkalis contain **hydroxide ions, OH^-**.

Reactions of Acids

- ***With metals***

Metals above copper in the reactivity series will react with acids, giving off hydrogen gas. The metal dissolves, forming a **salt**.

METAL + ACID → SALT + HYDROGEN

e.g. $Mg_{(s)} + H_2SO_{4(aq)} \rightarrow MgSO_{4(aq)} + H_{2(g)}$

This is why acids corrode metals, and must be stored in glass containers.

- ***With bases (metal oxides and hydroxides)***

The base dissolves in the acid and neutralises it. A salt is formed.

ACID + BASE → SALT + WATER

e.g. $H_2SO_{4(aq)} + CuO_{(s)} \rightarrow CuSO_{4(aq)} + H_2O_{(l)}$

- ***With metal carbonates***

With metal carbonates, much effervescence occurs when they react with acids, as carbon dioxide gas is released too.

ACID + CARBONATE → SALT + WATER + CARBON DIOXIDE

e.g. $2HCl_{(aq)} + CaCO_{3(s)} \rightarrow CaCl_{2(s)} + H_2O_{(l)} + CO_{2(g)}$

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Neutralisation

Acids are neutralised by bases. If you have indigestion (too much acid in the stomach), you may take a tablet containing a base (e.g. magnesia, MgO). A farmer may spread lime (calcium hydroxide, $Ca(OH)_2$) on fields to make the soil less acidic.

What happens during a neutralisation reaction?

ACID + ALKALI → SALT + WATER

e.g. hydrochloric acid + sodium hydroxide → sodium chloride + water

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

To understand why water is formed, we must consider what happens to the ions that are present in the reacting solutions:

$$H^+_{(aq)} + Cl^-_{(aq)} + Na^+_{(aq)} + OH^-_{(aq)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)} + H_2O_{(l)}$$
(from HCl) (from NaOH)

The H^+ ions from the acid combine with the OH^- ions from the alkali – we can show this by writing a simple **ionic equation**:

$$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$$

Water is formed. So if equal amounts of acid and alkali are mixed, the resulting solution will be neutral.

The Na^+ and Cl^- ions do not take part in the neutralisation reaction, they just remain in solution - they are **spectator ions**. It is unnecessary to include them in the ionic equation.

When an acid and a base are mixed in stoichiometric proportions, their acidic and basic properties disappear as the result of a **neutralization reaction**.

$$\underset{\text{Acid}}{HA(aq)} + \underset{\text{Base}}{MOH(aq)} \rightarrow \underset{\text{Water}}{H_2O(l)} + \underset{\text{A salt}}{MA(aq)}$$

Because the salts that form in neutralization reaction are generally strong electrolytes, we can write the neutralization reaction as an ionic equation.

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

When the spectator ions are removed, the net ionic equation is revealed.

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

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

$$\text{or} \quad H_3O^+(aq) + OH^-(aq) \rightarrow 2H_2O(l)$$

This net ionic equation is the same for the neutralization reaction of *any* strong acid and strong base.

For the reaction of a weak acid with a strong base, a similar neutralization occurs. Consider the neutralization of HF with KOH.

$HF(aq) + KOH(aq) \rightarrow KF(aq) + H_2O(l)$	molecular equation
$HF(aq) + \cancel{K^+(aq)} + OH^-(aq) \rightarrow \cancel{K^+(aq)} + F^-(aq) + H_2O(l)$	ionic equation
$HF(aq) + OH^-(aq) \rightarrow F^-(aq) + H_2O(l)$	net ionic equation

The weak acid HF is written as a molecular formula because its dissociation is incomplete.

Formation of H_30^+ion

The hydrogen ion $H^+(aq)$ does not exist as such in aqueous solutions. Hydrogen ions combine with water molecules to give a more stable species, the **hydronium ion H_3O^+**, as demonstrated in the following equation:

$$HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

Acids can contain different numbers of acidic hydrogens, and can yield different numbers of H_3O^+ ions in solution.

Sulfuric acid:

$$H_2SO_4(aq) + H_2O(l) \rightarrow HSO_4^-(aq) + H_3O^+(aq)$$

$$HSO_4^-(aq) + H_2O(l) \rightleftarrows SO_4^{2-}(aq) + H_3O^+(aq)$$

Phosphoric acid:

$$H_3PO_4(aq) + H_2O(l) \rightleftarrows H_2PO_4^-(aq) + H_3O^+(aq)$$

$$H_2PO_4^-(aq) + H_2O(l) \rightleftarrows HPO_4^{2-}(aq) + H_3O^+(aq)$$

$$HPO_4^{2-}(aq) + H_2O(l) \rightleftarrows PO_4^{3-}(aq) + H_3O^+(aq)$$

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Some important reactions of Bases (alkali = soluble base)

- **Neutralisation** with acids is dealt with above.
- **Ammonium salts are decomposed when mixed with a base** eg the alkali sodium hydroxide.
  - eg **sodium hydroxide + ammonium chloride ==> sodium chloride + water + ammonia**
  - **$NaOH + NH_4Cl ==> NaCl + H_2O + NH_3$**
  - The ammonia is readily detected by its pungent odour (strong smell) and by turning damp red litmus blue.
  - The ionic equation is: **$NH_4^+ + OH^- ==> H_2O + NH_3$**
  - This reaction can be used to prepare ammonia gas and as a simple chemical test for an ammonium salt.
- **Alkali's (soluble bases) are used to produce the insoluble hydroxide precipitates of many metal ions** from their soluble salt solutions.
  - eg sodium hydroxide + copper(II) sulphate ==> sodium sulphate + copper(II) hydroxide
  - **$2NaOH_{(aq)} + CuSO_{4(aq)} ==> Na_2SO_{4(aq)} + Cu(OH)_{2(s)}$** a blue precipitate
  - ionically: **$Cu^{2+}_{(aq)} + 2OH^-_{(aq)} ==> Cu(OH)_{2(s)}$**
  - This reaction can be used as a simple test to help identify certain metal ions.

ACIDIC, BASIC & AMPHOTERIC OXIDES

- **Oxygen combines with most other elements to form oxides of varying physical chemical character**.
  - On the left and middle of the Periodic Table are the **basic metal oxides** which react with acids to form salts eg Na_2O, MgO, CuO etc. These **metal oxides** tend to be ionic in bonding character with high melting points. The Group 1 Alkali Metals, and to a less extent, Group 2 oxides, dissolve in water to form alkali solutions. All of them react with , and neutralise acids to form salts.
  - As you move left to right the oxides become less basic and more acidic.
  - So on the right you have the **acidic oxides of the non-metals** CO_2, P_2O_5, SO_2, SO_3 etc. These tend to be covalent in bonding character with low melting/boiling points. Those of sulphur and phosphorus are very soluble in water to give acidic solutions which can be neutralised by alkalis to form salts.
  - **These oxides are another example of the change from metallic element to non-metallic element chemical behaviour from left to right across the Periodic Table**.
  - **BUT life is never that simple in chemistry!**:
    - Some oxides react with both acids and alkalis and are called **amphoteric oxides**. They are usually relatively insoluble and have little effect on indicators. An example is **aluminium oxide** dissolves in acids to form 'normal' aluminium salts like the chloride, sulphate and nitrate. However, it also dissolves in strong alkali's like sodium hydroxide solution to form 'aluminate' salts. This could be considered as 'intermediate' basic-acidic character in the Periodic Table.
    - **Some oxides are neutral**, tend to be of low solubility in water and have no effect on litmus, and do not react with acids or alkalis. eg CO

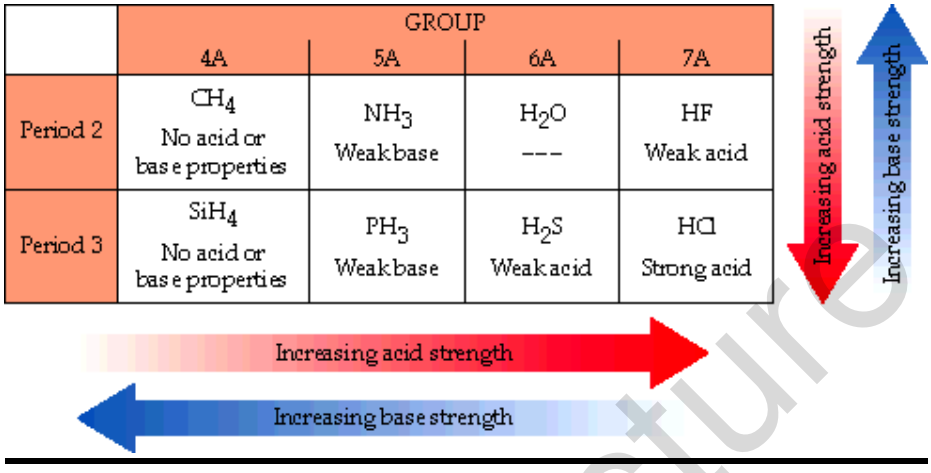
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carbon monoxide (note that CO_2 carbon dioxide is weakly acidic) and NO nitrogen monoxide (note that NO_2 nitrogen dioxide is strongly acidic in water). There is no way of simply predicting this kind of behaviour from periodic table patterns!



Salts

We have seen that when an acid reacts with a base, a salt is formed:

e.g. sulphuric acid + sodium hydroxide → **sodium sulphate** + water

Here sodium sulphate (Na_2SO_4) is the salt formed. Salts are ionic compounds.

The metal ion is provided by the base (in this case sodium ions, Na^+).
The non-metal ion is provided by the acid (in this case sulphate ions, SO_4^{2-}).

Note: Ammonia (NH_3) is an unusual base - it does not contain a metal. It forms **ammonium salts**, containing the ammonium ion, NH_4^+.

e.g. $NH_{3(aq)} + HNO_{3(aq)} \rightarrow NH_4NO_{3(aq)}$ (ammonium nitrate)

Soluble and Insoluble Salts

Many ionic salts are soluble in (cold) water. It is useful to know whether or not a salt is soluble in water, as this will influence your choice of method for making it. The table below gives a guide to the solubility of salts:

Soluble Salts
All common potassium, sodium and ammonium salts
All nitrates
All common ethanoates

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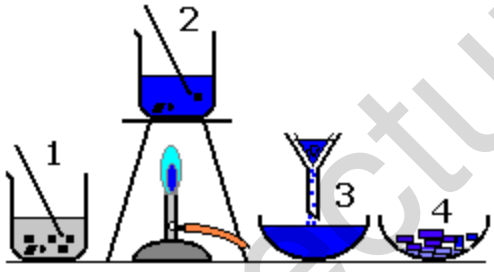
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All common chlorides, **except lead and silver chlorides**
All common sulphates, **except lead, barium and calcium sulphates**

Methods of making Salts which are water soluble

Soluble salts can be made in four different ways:

1) ACID + METAL → SALT + HYDROGEN
2) ACID + BASE → SALT + WATER
3) ACID + CARBONATE → SALT + WATER + CARBON DIOXIDE
4) ACID + ALKALI → SALT + WATER



Method 1 (Acid + Metal)
Not suitable for making salts of metals above magnesium, or below iron/tin in reactivity.

e.g. zinc + hydrochloric acid → zinc chloride + hydrogen

Apparatus used: (1) balance, measuring cylinder, beaker and glass stirring rod; **(2)** beaker/rod, bunsen burner, tripod and gauze; **(3)** filter funnel and filter paper, evaporating (crystallising) dish; **(4)** evaporating (crystallising) dish. **(ii)** A measuring cylinder is adequate for measuring the acid volume, you do not need the accuracy of a pipette or burette required in method (a).

* Add excess metal to (warm) acid. Wait until no more H_2 is evolved.
* Filter to remove excess metal.
* Heat the filtrate to evaporate off water until crystallisation starts.
* Set aside to cool slowly and crystallise fully.

- **Method 2 (Acid + Base)**

Useful for making salts of less reactive metals, e.g. lead, copper.

e.g. copper(II) oxide + sulphuric acid → copper(II) sulphate + water

* Add excess base to acid. Warm gently.
* Filter to remove excess base, then continue as in method 1…

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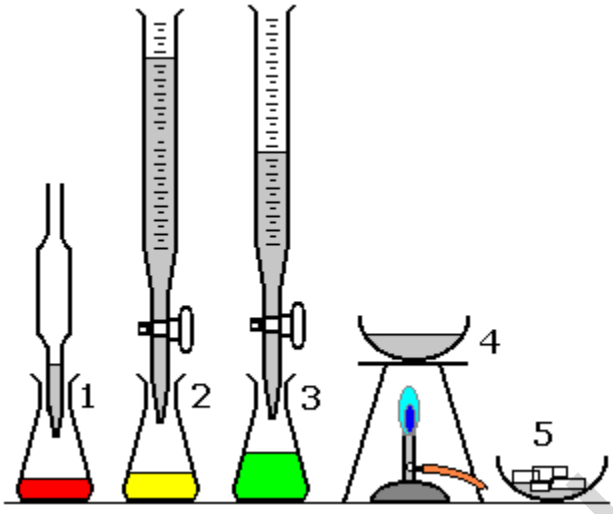
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- **Method 3 (Acid + Carbonate)**

Useful particularly for making salts of more reactive metals, e.g. calcium, sodium.

e.g. calcium carbonate + nitric acid → calcium nitrate + water + carbon dioxide

* Add excess metal carbonate to acid. Wait until no more CO_2 is evolved.
* Filter to remove excess carbonate, then continue as in method 1…



- **Method 4 (Acid + Alkali)**

This is useful for making salts of reactive metals, and ammonium salts. It is different from methods 1-3, as both reactants are in solution. This means neutralisation must be achieved, by adding exactly the right amount of acid to neutralise the alkali. This can be worked out by **titration**

e.g. sodium hydroxide + hydrochloric acid → sodium chloride + water

ammonia + sulphuric acid → ammonium sulphate

(1) A known volume of acid is pipetted into a conical flask and universal **indicator** added. The acid is titrated with the alkali in the burette

(2) until the indicator turns green.

(3). The **volume of alkali needed for neutralisation is then noted**, this is called the **endpoint**. (1-3) are **repeated** with both known volumes mixed together BUT **without the contaminating indicator**.

(4) The solution is transferred to an evaporating dish and **heated to partially evaporate the water**.

(5) The solution is left to cool to complete the crystallisation.

(6) The residual liquid can be decanted away and the **crystals can be carefully collected and dried** by 'dabbing' with a filter paper OR the crystals can be collected by filtration (below) and dried (as above).

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Making Insoluble Salts

Insoluble salts cannot be prepared by acid-base reactions in the same way as soluble salts.

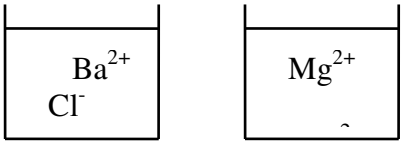
Insoluble salts are prepared by **precipitation**. This involves mixing solutions of two soluble salts that between them contain the ions that make up the insoluble salt. Here is an example:

barium chloride + magnesium **sulphate** → **barium sulphate** + magnesium chloride
(soluble salt) (soluble salt) (insoluble salt) (soluble salt)

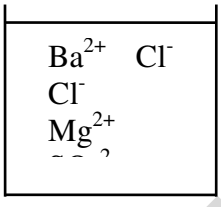
$$\mathbf{Ba}Cl_{2(aq)} + Mg\mathbf{SO_4}_{(aq)} \rightarrow \mathbf{BaSO_4}_{(s)} + MgCl_{2(aq)}$$

When the two solutions are mixed, a white solid **precipitate** of barium sulphate is formed.

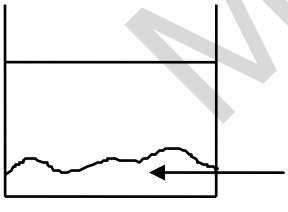
Let us consider what happens to the ions involved in the reaction:



1. At the start, the ions of each soluble salt move about freely in solution, in their separate containers.



2. After mixing, the ions of the two solutions are free to collide with each other, so new combinations of ions are possible.



3. If one of the new combinations of ions is an insoluble salt, it will be **precipitated** from the solution. The other ions simply remain in solution - they are **spectator ions**, and play no part in the reaction.

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The **ionic equation** for the reaction is:

$$Ba^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow BaSO_{4(s)}$$

Notice this equation does not include the spectator ions, only the ions that combine to form the precipitate.

Once the reaction is complete, the precipitate can be filtered off, washed with distilled water and dried.

This method can also be used to prepare many insoluble metal hydroxides and carbonates.

All common hydroxides/carbonates are insoluble except sodium, potassium and ammonium hydroxides/carbonates.

e.g. iron(III) chloride + sodium hydroxide → iron(III) hydroxide + sodium chloride
(rust brown ppt.)

$$FeCl_{3(aq)} + 3NaOH_{(aq)} \rightarrow Fe(OH)_{3(s)} + 3NaCl_{(aq)}$$

The ionic equation for this reaction, which shows the formation of the precipitate whilst missing out the spectator ions is as follows:

$$Fe^{3+}_{(aq)} + 3OH^-_{(aq)} \rightarrow Fe(OH)_{3(s)}$$

Types of Salts

Normal Salts:

Normal salts are formed when all the replaceable hydrogen ions in the acid have been completely replaced by metallic ions.

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

$$H_2SO_{4(aq)} + ZnO_{(aq)} \rightarrow ZnSO_{4(aq)} + H_2O_{(l)}$$

Normal salts are neutral to litmus paper.

Acid salts:

Acid salts are formed when replaceable hydrogen ions in acids are only partially replaced by a metal. Acid salts are produced only by acids containing more then one replaceable hydrogen ion. Therefore an acid with two replaceable ions e.g. H_2SO_4 will form only one

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acid salt, while acid with three replaceable hydrogen ions e.g. H_3PO_4 will form two different acid salts.

$H_2SO_{4(aq)} + KOH_{(aq)} \rightarrow KHSO_{4(aq)} + H_2O_{(l)}$

$H_3PO_{4(aq)} + NaOH \rightarrow NaH_2PO_{4(aq)} + H_2O_{(l)}$

$H_3PO_{4(aq)} + 2NaOH_{(aq)} \rightarrow Na_2HPO_{4(aq)} + 2H_2O_{(l)}$

An acid salt will turn blue litmus red. In the presence of excess metallic ions an acid salt will be converted into a normal salt as its replaceable hydrogen ions become replaced.

Basic Salts:

Basic salts contain the hydroxide ion, OH-. They are formed when there is insufficient supply of acid for the complete neutralization of the base. A basic salt will turn red litmus blue and will react with excess acid to form normal salt.

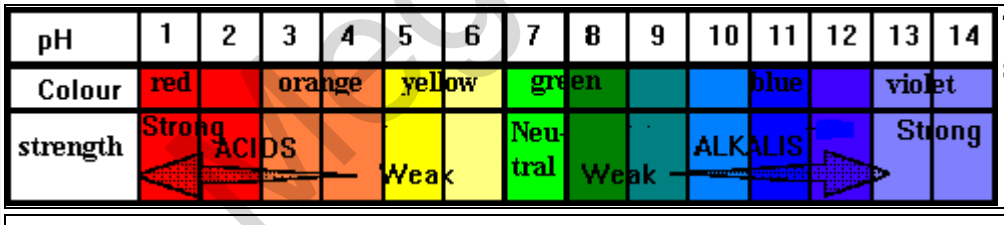
$Zn(OH)_{2(s)} + HCl_{(aq)} \rightarrow Zn(OH)Cl_{(aq)} + H_2O_{(l)}$

$Zn(OH)Cl_{(aq)} + HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_2O_{(l)}$

$Mg(OH)_{2(s)} + HNO_{3(aq)} \rightarrow Mg(OH)NO_{3(aq)} + H_2O_{(l)}$

$Mg(OH)NO_{3(aq)} + HNO_{3(aq)} \rightarrow Mg(NO_3)_{2(aq)} + H_2O_{(l)}$

The pH Scale - Acids and Alkalis



The colours of solutions with universal indicator

-
- The **pH scale** is a measure of the relative acidity or alkalinity of a solution.
- To **find the pH of a solution an indicator is used** like Universal Indicator. An indicator is a substance or mixture of substances that when added to the solution gives **different colours depending on the pH** of the solution. **Universal indicator is a very handy indicator for showing whether the solution is acid, neutral or alkaline** and gives the pH to the nearest pH unit.
- **Water is a neutral liquid** with a pH of 7 (green). When a substance dissolves in water it forms an aqueous (aq) solution that may be acidic, neutral or alkaline.
- **Acidic solutions have a pH of less than 7**, and the lower the number, the stronger the acid is. The colour can range from orange-yellow (pH 3-6) for partially ionised **weak acids** like ethanoic acid (vinegar) to carbonated water. **Strong acids** like hydrochloric, sulphuric and nitric are fully ionised and give a pH 1 or less! and a **red colour with universal**

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indicator and litmus paper.

- **Neutral solutions have a pH of 7**. These are quite often solutions of salts, which are themselves formed from **neutralising** acids and bases.
- **The 'opposite' of an acid is called a base**. **Some bases are soluble** in water to give alkaline solutions - **these are known as alkalis**.
- **Alkaline solutions have a pH of over 7** and the higher the pH the stronger is the alkali. **Weak alkalis (soluble bases)** like ammonia give a pH of 10-11 but **strong alkalis (soluble bases)** like sodium hydroxide give a pH of 13-14. **They give blue/purple colour with universal indicator or litmus paper**.
- **NEUTRALISATION** usually involves **mixing an acid (pH <7)** with a **base or alkali (pH > 7)** which react to form a **neutral salt solution of pH 7**.

INDICATORS.

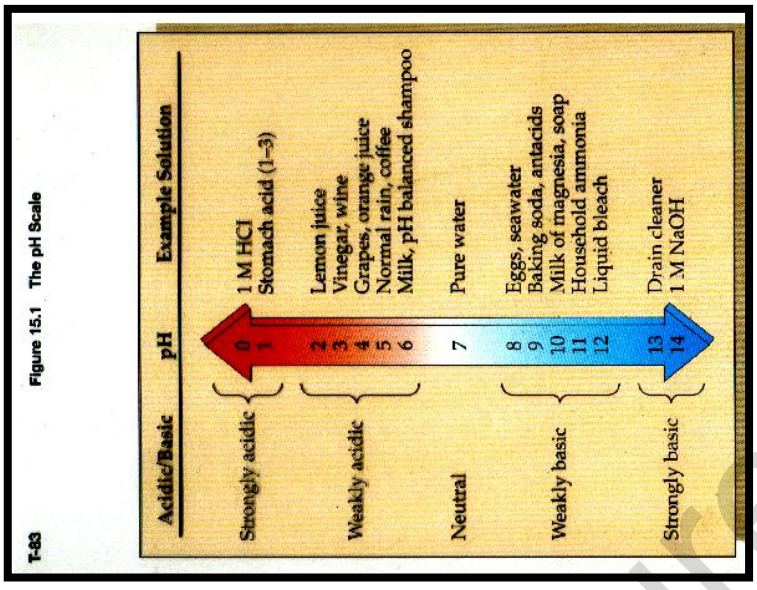
Indicators are the substances that have different colors in acidic and in alkaline solution. Some important indicators are given below

S.No.	Indicator	Color in strongly acidic solution	pH at which color changes	Color in strongly alkaline solution
1.	methyl orange	red	4	yellow
2.	bromothymol blue	yellow	7	blue
3.	phenolphthalein	colorless	9	red
4.	screened methyl orange	red	4	green

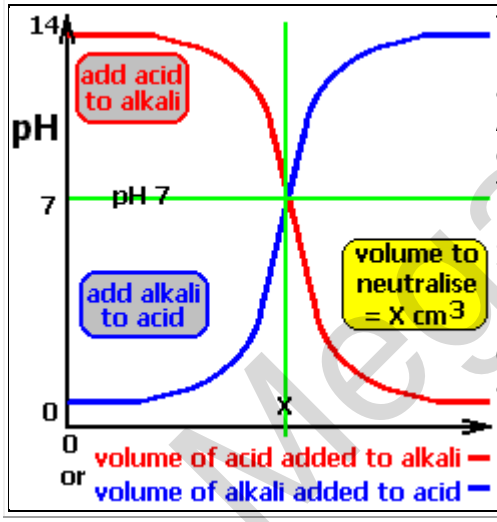
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What pH changes go on in a neutralisation reaction?



The graphs show how the pH changes when an alkali (soluble base) and an acid neutralise each other. If you add acid to an alkali (ind. = blue), the pH starts at 13 and falls little at first. As you get near to the neutralisation point at pH 7 (ind. = green), the change becomes more dramatic. With excess acid the pH falls and then levels out to about pH1 (ind. = red).

If you add alkali to an acid (ind. = red), the pH starts at about 1 and rises little at first. As you get near to the neutralisation point at pH 7 (ind. = green), the change becomes more dramatic. With excess alkali the pH rises and then levels out to about 13 (ind. = blue/violet).

Formulae of bases: oxides, hydroxides and carbonates	**Formulae of salts formed: soluble chlorides, sulphates and nitrates**	**The metal (or other ion) involved**
M_2O (oxide O^{2-}, soluble, alkali) **MOH** (hydroxide OH^-, soluble, alkali) **M_2CO_3** (carbonate CO_3^{2-}, soluble mild alkali)	**MCl** (chloride, Cl^-) **M_2SO_4** (sulphate, SO_4^{2-}) **MNO_3** (nitrate, NO_3^-)	**M =** Li, Na, K [usually Group 1], the ion is **M^+**

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$MHCO_3$ (hydrogencarbonate HCO_3^-, soluble, mild alkali)		
MO (oxide, often insoluble base) **$M(OH)_2$** (hydroxide, often insoluble, alkali if soluble) **MCO_3** (carbonate, often insoluble)	**MCl_2** (chloride) **MSO_4** (sulphate) **$M(NO_3)_2$** (nitrate)	**M =** Mg, Ca, Cu, Zn, Fe [often Group 2 or Transition], the ion is **M^{2+}**
Al_2O_3, $Al(OH)_3$ (insoluble bases, amphoteric)	**$AlCl_3$, $Al_2(SO_4)_3$, $Al(NO_3)_3$**	**Al**, aluminium from Group 3
the alkaline soluble base ammonia, **NH_3**, no stable oxide or hydroxide	**NH_4Cl, $(NH_4)_2SO_4$, NH_4NO_3**	the **ammonium** ion in the salts from ammonia, **NH_4^+**

FURTHER EXAMPLES of WORD & SYMBOL EQUATIONS for Salt Preparations

1a. copper(II) carbonate + sulphuric acid ==> copper(II) sulphate + water +CO2.

1b. $CuCO_{3(s)} + H_2SO_{4(aq)} ==> CuSO_{4(aq)} + H_2O_{(l)} + CO_{2(g)}$

2a. magnesium hydroxide + hydrochloric acid ==> magnesium chloride + water 2b. $Mg(OH)_{2(s)} + 2HCl_{(aq)} ==> MgCl_{2(aq)} + 2H_2O_{(l)}$

3a. zinc + sulphuric acid ==> zinc sulphate + hydrogen

3b. $Zn_{(s)} + H_2SO_{4(aq)} ==> ZnSO_{4(aq)} + H_{2(g)}$

4a. ammonia + nitric acid ==> ammonium nitrate

4b. $NH_{3(aq)} + HNO_{3(aq)} ==> NH_4NO_{3(aq)}$

5a. zinc oxide + hydrochloric acid ==> zinc chloride + water

5b. $ZnO_{(s)} + 2HCl_{(aq)} ==> ZnCl_{2(aq)} + H_2O_{(l)}$

6a. calcium carbonate + hydrochloric acid ==> calcium chloride + water + carbon dioxide

6b. $CaCO_{3(s)} + 2HCl_{(aq)} ==> CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

7a. sodium carbonate + hydrochloric acid ==> sodium chloride + water + carbon dioxide

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7b. $Na_2CO_{3(s)} + 2HCl_{(aq)} \Longrightarrow 2NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$

8a. sodium hydroxide + hydrochloric acid ==> sodium chloride + water

8b. $NaOH_{(aq)} + HCl_{(aq)} \Longrightarrow NaCl_{(aq)} + H_2O_{(l)}$

9a. sodium hydrogencarbonate + hydrochloric acid ==> sodium chloride + water + carbon dioxide

9b. $NaHCO_{3(s)} + HCl_{(aq)} \Longrightarrow NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$

IONIC EQUATIONS

Aqueous Reactions and Net Ionic Equations

The equations written up to this point have been molecular equations. All substances have been written using their full chemical formulas as if they were molecules. Because we now know that strong electrolytes dissociate in water to their component ions, it is more accurate to write an ionic equation in which all of the ionic species are shown.

- **In many reactions only certain ions change their 'chemical state' but other ions remain in exactly the same original physical and chemical state.**
- **The ions that do not change are called 'spectator ions'.**
- The ionic equation represents the 'actual' chemical change and omits the spectator ions.

To write a net ionic equation:

1. Write a balanced molecular equation.
2. Rewrite the equation showing the ions that form in solution when each soluble electrolyte dissociates into its component ions. Only dissolved strong electrolytes are written in ionic form.
3. Identify and cancel the spectator ions that occur unchanged on both sides of the equation.

- **Five types of examples are presented below.**

When reactions between ions occur, at least one kind of ion is removed from the "field of action". Simply put, its concentration decreases as the reaction proceeds.

There are three ways to remove ions:

1. Formation of an insoluble precipitate
2. Formation of a weakly ionized substance, and
3. Oxidation or reduction of an ion

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Let's examine each way individually:

1) Formation of An Insoluble Precipitate

$$Overall\ \text{Re}\ action:\ AgNO_3 + HCl \rightarrow AgCl\downarrow + HNO_3$$
$$Ionic\ \text{Re}\ action:\ Ag^+ + NO_3^- + H_3O^+ + Cl^- \rightarrow AgCl\downarrow + NO_3^- + H_3O^+$$
$$Net\ Ionic\ \text{Re}\ action:\ Ag^+ + Cl^- \rightarrow AgCl\downarrow$$

Excess chloride ion "drives" this reaction to the right.

2) Formation of A Weakly Ionized Substance

$$Overall\ \text{Re}\ action:\ NH_4Cl + NaOH \rightarrow NH_3 + H_2O + NaCl$$
$$Ionic\ \text{Re}\ action:\ NH_4^+ + Cl^- + Na^+ + OH^- \rightarrow NH_3 + H_2O + Na^+ + Cl^-$$
$$Net\ Ionic\ \text{Re}\ action:\ NH_4^+ + OH^- \rightarrow NH_3 + H_2O$$

3) Oxidation or Reduction of An Ion

$$S^{2-} \xrightarrow[removal]{oxidized\ for} S + 2e^-$$
$$Cu^{2+} \xrightarrow[removal]{oxidized\ for} Cu + 2e^-$$

Rules for Writing Ionic Equations

1) Ionic formulas are written for a strong electrolyte in solution e.g

$$Na^+ + Cl^-\ for\ NaCl$$
$$2Na^+ + SO_4^{2-}\ for Na_2SO_4$$
$$2H^+ + SO_4^{2-}\ for\ H_2SO_4$$
$$Na^+ + HCO_3^-\ for\ NaHCO_3$$
$$3Na^+ + PO_4^{3-}\ for\ Na_3PO_4$$

2) Molecular formulas are written for:

a) Elements, gases, solids and non-electrolytes, e.g.:

$$Cl_2 \quad SO_2 \quad Cu \quad HCl(g)$$
$$CH_3CSNH_2 \quad CCl_4 \quad C_2H_5OH$$

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b) Weak electrolytes in solution, e.g.:

$$H_2O \quad HOAc \quad NH_3$$

c) Solid strong electrolytes or precipitates, e.g.:

$$\underline{AgCl} \text{ or } AgCl\downarrow$$
$$\underline{BaSO_4} \text{ or } BaSO_4\downarrow$$
$$\underline{CaCO_3} \text{ or } CaCO_3\downarrow$$

Writing Ionic Equations

When writing these equations, do so to answer the following three (3) questions:

1. What kind of reaction is it? Double decomposition? Redox?
2. What are the possible products of the reaction?
3. Are any of the possible products or reactants insoluble or weakly ionized?

Double Decomposition Reactions (Precipitate and Weak Electrolyte Reactions)

Examples

a) $KCl + NaNO_3 \rightarrow$ NR

Even the products would be soluble and, hence, no reaction occurs.

Since the hydrogen ion and the nitrate ion are spectators, the net ionic reaction is the result.

- **(1) Acid-base reactions: Acids can be defined as proton donors. A base can be defined as a proton acceptor.**
  - eg any acid-alkali neutralisation involves the hydroxide ion is (base) and this accepts a proton from an acid.
    - **$HCl_{(aq)} + NaOH_{(aq)} ==> NaCl_{(aq)} + H_2O_{(l)}$** which can be re-written as
    - **$H^+Cl^-_{(aq)} + Na^+OH^-_{(aq)} ==> Na^+Cl^-_{(aq)} + H_2O_{(l)}$**
    - **$H^+_{(aq)} + OH^-_{(aq)} ==> H_2O_{(l)}$**
    - the **spectator ions** are **Cl^- and Na^+**
- **(2) Insoluble salt formation: An insoluble salt is made by mixing two solutions of soluble compounds to form the insoluble compound in a process called 'precipitation'.**
  - **(a) Silver chloride** is made by mixing solutions of solutions of silver nitrate and sodium chloride.
    - silver nitrate + sodium chloride ==> silver chloride + sodium nitrate
    - **$Ag^+NO_3^-_{(aq)} + Na^+Cl^-_{(aq)} ==> AgCl_{(s)} + Na^+NO_3^-_{(aq)}$**
    - the ionic equation is: **$Ag^+_{(aq)} + Cl^-_{(aq)} ==> AgCl_{(s)}$**
    - the **spectator ions** are **NO_3^-** and **Na^+**

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(b) Silver nitrate and hydrochloric acid -- precipitate formation

$$Overall\ \text{Re}action: AgNO_3 + HCl \rightarrow AgCl + HNO_3$$
$$Ionic\ \text{Re}action: Ag^+ + NO_3^- + H^+ + Cl^- \rightarrow AgCl\downarrow + H^+ + NO_3^-$$
$$Net\ Ionic\ \text{Re}action: Ag^+ + Cl^- \rightarrow \underline{AgCl}$$

- **Lead(II) iodide** can be made by mixing lead(II) nitrate solution with potassium iodide solution.
  - lead(II) nitrate + potassium iodide ==> lead(II) iodide + potassium nitrate
  - **$Pb(NO_3)_{2(aq)} + 2KI_{(aq)} ==> PbI_{2(s)} + 2KNO_{3(aq)}$**
  - the ionic equation is: **$Pbg^{2+}_{(aq)} + 2I^-_{(aq)} ==> PbI_{2(s)}$**
  - the **spectator ions** are **NO_3^-** and **K^+**
- **(c) Calcium carbonate** forms on eg mixing calcium chloride and sodium carbonate solutions
  - calcium chloride + sodium carbonate ==> calcium carbonate + sodium chloride
  - **$CaCl_{2(aq)} + Na_2CO_{3(aq)} ==> CaCO_{3(s)} + 2NaCl_{(aq)}$**
  - ionically: **$Ca^{2+}_{(aq)} + CO_3^{2-}_{(aq)} ==> CaCO_{3(s)}$**
  - the **spectator ions** are **Cl^-** and **Na^+**
- **(d) Barium sulphate** forms on mixing eg barium chloride and dilute sulphuric acid
  - barium chloride + sulphuric acid ==> barium sulphate + hydrochloric acid
  - **$BaCl_{2(aq)} + H_2SO_{4(aq)} ==> BaSO_{4(s)} + 2HCl_{(aq)}$**
  - ionically: **$Ba^{2+}_{(aq)} + SO_4^{2-}_{(aq)} ==> BaSO_{4(s)}$**
  - the **spectator ions** are **CO_3^{2-}** and **H^+**

- **(3) Redox reaction analysis:**
  - **(a) magnesium + iron(II) sulphate ==> magnesium sulphate + iron**
    - **$Mg_{(s)} + FeSO_{4(aq)} => MgSO_{4(aq)} + Fe_{(s)}$**
    - this is the 'ordinary molecular' **equation for a typical metal displacement reaction**, but this does not really show what happens in terms of atoms, ions and electrons, so we use **ionic equations** like the one shown below.
    - The sulphate ion **$SO_4^{2-}{}_{(aq)}$** is called a **spectator ion**, because it doesn't change in the reaction and can be omitted from the ionic equation. No electrons show up in the full equations because electrons lost by x = electrons gained by y!!
    - **$Mg_{(s)} + Fe^{2+}_{(aq)} ==> Mg^{2+}_{(aq)} + Fe_{(s)}$**
    - Mg oxidised by electron loss, Fe^{2+} reduced by electron gain
  - **(b) zinc + hydrochloric acid ==> zinc chloride + hydrogen**
    - **$Zn_{(s)} + 2HCl_{(aq)} => ZnCl_{2(aq)} + H_{2(g)}$**
    - the chloride ion Cl^- is the spectator ion
    - **$Zn_{(s)} + 2H^+_{(aq)} ==> Zn^{2+}_{(aq)} + H_{2(g)}$**
    - Zn oxidised by electron loss, H^+ reduced by electron gain
  - **(c) copper + silver nitrate ==> silver + copper(II) nitrate**
    - **$Cu_{(s)} + 2AgNO_{3(aq)} ==> 2Ag + Cu(NO_3)_{2(aq)}$**
    - the nitrate ion NO_3^- is the spectator ion
    - **$Cu_{(s)} + 2Ag^+_{(aq)} ==> 2Ag_{(s)} + Cu^{2+}_{(aq)}$**
    - Cu oxidised by electron loss, Ag^+ reduced by electron gain
  - **(d) halogen** (more reactive) **+ halide salt** (of less reactive halogen) **==> halide salt** (of more reactive halogen) **+ halogen** (less reactive)
    - $\mathbf{X}_{2(aq)} + 2K^+\mathbf{Y}_{(aq)} ==> 2K^+\mathbf{X}_{(aq)} + \mathbf{Y}_{2(aq)}$
    - $\mathbf{X}_{2(aq)} + 2\mathbf{Y}^-_{(aq)} ==> 2\mathbf{X}^-_{(aq)} + \mathbf{Y}_{2(aq)}$
    - the potassium ion K^+ is the spectator ion
    - halogen **X** is more reactive than halogen **Y**, F > Cl > Br > I)

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- **X** is the **oxidising agent** (electron acceptor, so is reduced)
- K**Y or Y$^-$** is the **reducing agent** (electron donor, so is oxidised)

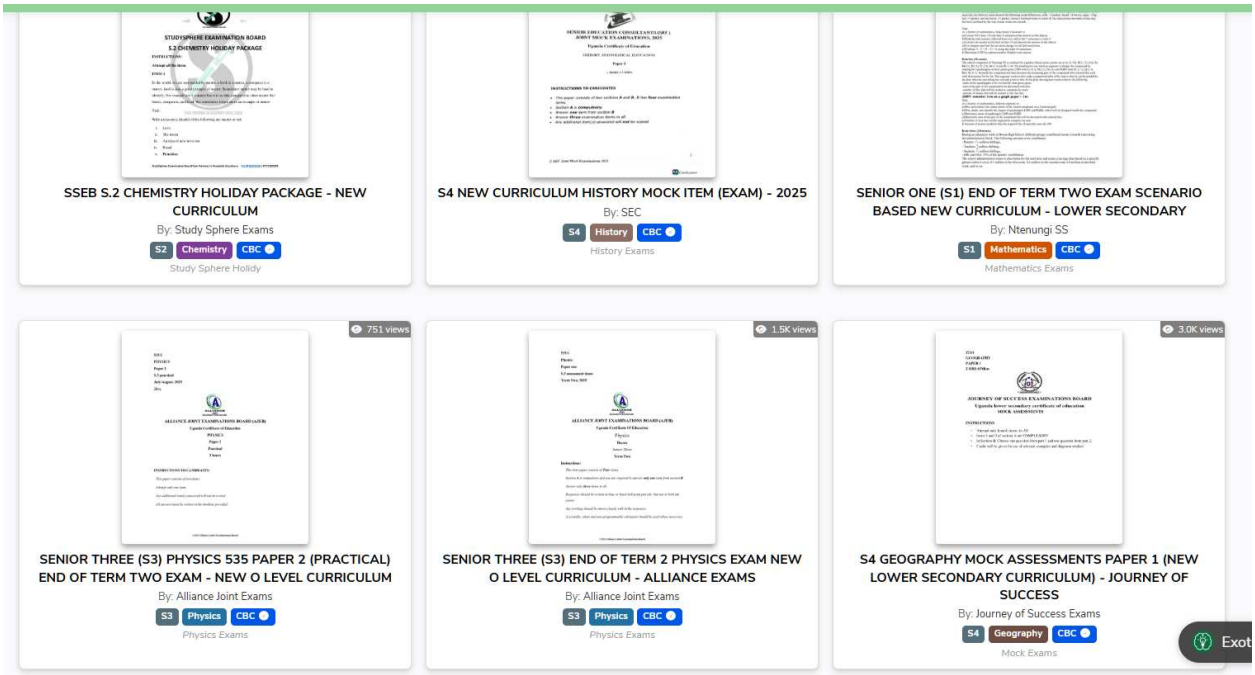
- **(4) Ion Exchange Resins:** Ion exchange polymer resin columns hold hydrogen ions or sodium ions. These can be replaced by calcium and magnesium ions when hard water passes down the column. The calcium or magnesium ions are held on the negatively charged resin. The freed hydrogen or sodium ions do not form a scum with soap.
  - eg **$2[resin]^-H^+_{(s)} + Ca^{2+}_{(aq)} ==> [resin]^-Ca^{2+}[resin]^-_{(s)} + 2H^+_{(aq)}$**
  - or **$2[resin]^-Na^+_{(s)} + Mg^{2+}_{(aq)} ==> [resin]^-Mg^{2+}[resin]^-_{(s)} + 2Na^+_{(aq)}$** etc.
- **(5) Scum formation with hard water:** On mixing hard water with soaps made from the sodium salts of fatty acids, insoluble calcium or magnesium salts of the soap are formed ...
  - **$CaSO_{4(aq)} + 2C_{17}H_{35}COONa_{(aq)} ==> (C_{17}H_{35}COO)_2Ca_{(s\ for\ scum!)} + Na_2SO_{4(aq)}$**
  - or more simply ionically: **$Ca^{2+}_{(aq)} + 2C_{17}H_{35}COO^-_{(aq)} ==> (C_{17}H_{35}COO^-)_2Ca^{2+}_{(s)}$**
  - the **spectator ions** are **SO_4^{2-}** and **Na^+**

DONE

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