Reactions involving ionic compounds.
As discussed earlier, ionically bonded compounds consist of large aggregations of cations and anions which pack together in crystal lattices in such a way that the electrostatic attractions between oppositely charged ions are maximised and repulsions between like charged ions minimised. When an ionic crystal is placed in water, in many cases the solid dissolves, releasing the component ions into the SOLVENT to form a SOLUTION. Such compounds are said to be SOLUBLE and the substance that dissolves is called the SOLUTE. A well known example of a soluble ionic compound is table salt or sodium chloride. The process of dissolving can be best represented by an equation which is slightly different from the formula equations used in Topic 5. Instead, an IONIC EQUATION is used to show the ions released into the solution as follows.

NaCl(s) → Na⁺ + Cl⁻

There is an upper limit to the amount of solute that can be dissolved in a solvent at a given temperature. When no more solute can be dissolved, the solution is said to be SATURATED. The maximum solubility of compounds increases with temperature because more energy is available at higher temperatures, allowing the ions to escape from the attractive forces in the crystal lattice.

Ionic Equations.
An ionic equation is able to show the physical states of all the reactants and products unambiguously by including any dissolved ionic species as ions. The same rules apply as for formula equations in that all species shown on the left must also be present on the right hand side of the equation. In addition, notice that the electrical charge present on both sides of the equation must also balance.

Thus the equation for another ionic solid, barium chloride, dissolving in water would be as follows

BaCl₂(s) → Ba²⁺ + Cl⁻ + Cl⁻

which is usually written as

BaCl₂(s) → Ba²⁺ + 2Cl⁻

Because the Cl⁻ ions are separate individual species, they are represented as 2Cl⁻ and not as Cl₂⁻, which would mean two Cl atoms bonded together and bearing a 2 negative charge, (Cl–Cl)²⁻.

Why do ionic solids dissolve in water?
When ions are released into water solution, they all experience attractions to water molecules which form spheres around them. The reason for this attractive force between water molecules and ions is the ability of the oxygen atoms in water molecules to attract the electrons in their O–H bonds to a greater extent than do the hydrogen atoms. The O atom is said to be more ELECTRONEGATIVE than the H atom. This results in a slight negative charge on the oxygen atom and a slight positive charge on each hydrogen atom. The O–H
bond is an example of a **POLAR BOND** and the water molecule, being angular in shape, has a non-symmetric distribution of charge and is a **POLAR MOLECULE**.

In the dissolution of ionic solids such as sodium chloride, the oxygen atoms of water molecules are attracted to the positive charge on cations (Na\(^+\) in this example) while its hydrogen atoms are attracted to the anions (Cl\(^-\) in this example). This results in considerable energy being released as individual ions become surrounded by the attracted water molecules as shown below.

Consequently, ions in water solution are said to be **AQUATED** and sometimes the suffix (aq) is used to emphasise this point. The large amount of energy released by the process of aquating ions may result in the crystal lattice breaking down.
Thus the two equations above might also be written as

\[
\text{NaCl}(s) \rightarrow \text{Na}^+(aq) + \text{Cl}^-(aq) \\
\text{BaCl}_2(s) \rightarrow \text{Ba}^{2+}(aq) + 2\text{Cl}^-(aq)
\]

Initially the (aq) suffix will be used here, but later it will be assumed that all ions in water solution are aquated and the (aq) suffix will be omitted. Some other examples of ionic equations for ionic solids dissolving follow.

\[
\text{K}_2\text{CO}_3(s) \rightarrow 2\text{K}^+(aq) + \text{CO}_3^{2-}(aq) \\
(\text{NH}_4)_3\text{PO}_4(s) \rightarrow 3\text{NH}_4^+(aq) + \text{PO}_4^{3-}(aq)
\]

Many text books use the (aq) symbolism (incorrectly) to indicate a solution of a substance in water by attaching (aq) to the formula of that solid. Clearly, there can be no such species as an aquated ionic substance because if it has dissolved, it is totally present as ions. Therefore equations showing species such as NaCl(aq) are most misleading and should be ignored.

Not all ionic solids will dissolve in water to a significant extent. For example, the amount of the ionic solid silver chloride, AgCl, which will dissolve in water is so small that it is classed as insoluble. Insoluble ionic compounds of common metals include three chlorides, about five sulfates, most carbonates, most phosphates and most sulfides. All nitrates are soluble in water.

**How does one know if a given ionic compound is soluble?**

A table classifying the solubilities of ionic compounds in water is given at the end of this Topic. This table lists the ionic compounds of the common cations as soluble, insoluble or slightly soluble. This information probably will be needed to be committed to memory at a later stage of your course, but for the present you should consult it if needed for the ensuing exercises. The contents of the table are more easily remembered in terms of the general situation for each anion. For example: all nitrates are soluble; most hydroxides are insoluble except for those of the first family members Na and K, plus Ba; most carbonates are insoluble except those of the first family members Na and K. It is also useful to remember that all of the compounds of the elements Na and K from the first family are soluble.

<table>
<thead>
<tr>
<th>Check your understanding of this section:</th>
</tr>
</thead>
<tbody>
<tr>
<td>Why do ionic compounds when soluble, dissolve best in water?</td>
</tr>
<tr>
<td>Write an ionic equation for the dissolution of iron(III) chloride in water.</td>
</tr>
<tr>
<td>Account for the water molecule being polar.</td>
</tr>
<tr>
<td>What is the source of the energy required to break up an ionic crystal when it dissolves?</td>
</tr>
</tbody>
</table>
How can a soluble ionic compound be obtained back from a solution?
The aquated ions released into a solution when an ionic compound dissolves and its ionic bonds broken are free to move about throughout the solution. Ionic bonding is only present in the solid state. However, if the solution is boiled, the water is driven off as a gas (VOLATILE) but the ions remain in the solution (NON-VOLATILE). Volatile substances have sufficiently weak forces of attraction between their constituent entities to allow them to escape to the vapour phase when enough heat energy is supplied. In non-volatile substances, the attractive forces operating are much stronger and much higher temperatures would be required to vaporise them. When sufficient water has been evaporated the solution becomes saturated. The cations and anions are deprived of their surrounding water molecules so they can then recombine to form the ionically bonded solid again. In this way, any solution of an ionic compound can be evaporated sufficiently for crystals to form. The equation for the evaporation process would simply be the reverse of the equation for the dissolution, for example

\[
\text{Na}^+(aq) + \text{Cl}^-(aq) \rightarrow \text{NaCl}(s)
\]

To emphasise that the solution is being evaporated, it may be helpful to write "evap" or similar over the arrow.

Precipitation reactions.
Silver chloride is seen from the solubility table to be insoluble in water. Silver nitrate and sodium chloride are both soluble compounds and in water-solution would release their component ions as shown in the following ionic equations.

\[
\text{AgNO}_3(s) \rightarrow \text{Ag}^+(aq) + \text{NO}_3^-(aq)
\]

\[
\text{NaCl}(s) \rightarrow \text{Na}^+(aq) + \text{Cl}^-(aq)
\]

If a solution of silver nitrate were mixed with a solution of sodium chloride, the solution would momentarily be SUPER SATURATED with respect to AgCl and the Ag\(^+\) ions would react with the Cl\(^-\) ions to form a PRECIPITATE of the insoluble salt silver chloride, AgCl, as shown in the following ionic equation.

\[
\text{Ag}^+(aq) + \text{NO}_3^-(aq) + \text{Na}^+(aq) + \text{Cl}^-(aq) \rightarrow \text{AgCl}(s) + \text{NO}_3^-(aq) + \text{Na}^+(aq)
\]

Note that only the Ag\(^+\) and the Cl\(^-\) ions have reacted, leaving the Na\(^+\) and NO\(_3^-\) ions free in the solution because sodium nitrate is much more soluble than silver chloride. As these last two ions have not in fact entered into a reaction, they can be deleted from the equation in much the same way as common terms are cancelled from both sides of a mathematical equation. Such ions are called SPECTATOR IONS. Initially it may be helpful to write down all the ions which are being mixed together in order to establish whether any combination can form an insoluble salt, and then cancel out the spectator species. However, with practice you will be able to delete this step and write the final equation in one step. For this reaction it would be

\[
\text{Ag}^+(aq) + \text{Cl}^-(aq) \rightarrow \text{AgCl}(s)
\]

Being a solid, the silver chloride could be obtained by filtering the mixture. The precipitate would be retained in the filter paper and the solution containing the
spectator ions (called the **FILTRATE**) would pass through the filter paper. It would contain all of the Na\(^+\) and NO\(_3^-\) ions, provided exactly equal amounts of the silver nitrate and sodium chloride were in the solutions which were mixed originally. By evaporating the water, crystals of sodium nitrate could be isolated from the filtrate.

\[
\text{Na}^+(aq) + \text{NO}_3^-(aq) \xrightarrow{\text{evap}} \text{NaNO}_3(s)
\]

From a knowledge of the solubilities of ionic compounds, one can predict whether any combination of solutions of soluble compounds will lead to a precipitation reaction. The following examples show how to establish whether a precipitation reaction occurs when solutions are mixed.

**Example 1.**
*If a water solution of potassium chloride were added to a water solution of copper(II) nitrate, would any reaction occur and if so, what would be the product?*

To answer this, consider all possible combinations of the four ions which are to be mixed - K\(^+\), Cl\(^-\), Cu\(^{2+}\) and NO\(_3^-\). The formulas of the possible products are KNO\(_3\) and CuCl\(_2\). By consulting the solubility table given, it is seen that both of these compounds are soluble so there would be no reaction. If the solution formed by mixing this combination were to be evaporated, the resulting solid would be a mixture of all four possible compounds, KNO\(_3\), CuCl\(_2\), Cu(NO\(_3\))\(_2\) and KCl.

**Example 2.**
*Water solutions of lead(II) nitrate and sodium sulfate are mixed. What, if any, reaction occurs?*

The possible combinations of ions could produce the compounds of formulas PbSO\(_4\) and NaNO\(_3\). Of these two, lead(II) sulfate is insoluble and sodium nitrate is soluble. Therefore, a precipitate of PbSO\(_4\) would form according to the equation

\[
Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)
\]

In this example, only one of the possible products is insoluble (PbSO\(_4\)) but if both possible products were insoluble, then a mixture of both compounds would form.

**Example 3.**
*Water solutions of sodium carbonate and barium chloride are mixed. What, if any, reaction occurs?*

The possible combinations of ions could produce the compounds NaCl and BaCO\(_3\). Of these two, barium carbonate is insoluble while sodium chloride is soluble. Therefore, a precipitate of BaCO\(_3\) would form according to the equation

\[
Ba^{2+}(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s)
\]
Example 4.
Water solutions of iron(II) sulfate and potassium phosphate are mixed. What, if any, reaction occurs?

The possible combinations of ions could produce the compounds $K_2SO_4$ and $Fe_3(PO_4)_2$. Of these two, iron(II) phosphate is insoluble while potassium sulfate is soluble. Therefore, a precipitate of $Fe_3(PO_4)_2$ would form according to the equation

$$3Fe^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Fe_3(PO_4)_2(s)$$

Check your understanding of this section:
Given that the formation of an ionic solid from its component ions involves the release of energy, why is it necessary to supply heat to a solution of sodium chloride in order to reclaim the solid from solution?
What advantages are there in using an ionic equation to represent the precipitation of silver chloride from a mixture of sodium chloride and silver nitrate solutions?
In the previous reaction, which are the spectator ions?

Acids.
Apart from the cations of metals and the polyatomic cation $NH_4^+$, another cation frequently encountered in reactions is the HYDROGEN ION, $H^+$, which is also sometimes called the oxonium ion or the hydronium ion. Recall from Topic 2 that the $H$ atom consists of just one proton for the nucleus, surrounded by a single orbiting electron. If this electron were removed, the $H^+$ ion would be formed and it would be just a free proton. A proton is extremely small and the charge density on it would therefore be very concentrated. Consequently, the $H^+$ ion does not have an independent existence in solution. Instead it associates with water molecules by joining on to them using one of the lone pairs of electrons on the O atom, and is more correctly represented as $H^+(aq)$, or frequently as $H_3O^+$. These are all equally acceptable ways of representing the hydrogen ion. An $H_3O^+$ ion could be represented as in the following diagram.

![Hydrogen Ion Diagram](image)

Hydrogen ions are supplied in water solution by compounds called ACIDS. Examples of commonly encountered acids include

nitric acid, $HNO_3$, which contains $H^+(aq)$ and $NO_3^-(aq)$ ions
sulfuric acid, $H_2SO_4$, which contains $H^+(aq)$, $HSO_4^-(aq)$ and $SO_4^{2-}(aq)$ ions
hydrochloric acid, which is a water solution of the gas hydrogen chloride, $HCl(g)$, which totally ionizes to $H^+(aq)$ and $Cl^-(aq)$ ions in the water.
Acids such as these provide $H^+$ ions as well as the anion from the acid in solution. Note that the anions associated with the $H^+$ ion in these acids have already been encountered in earlier topics. The chloride ion, $Cl^-$, is a component of many binary compounds that were described in Topic 3 and the nitrate ion ($NO_3^-$) and sulfate ions ($SO_4^{2-}$), both polyatomic anions, were described in Topic 4. Many other polyatomic anions occur as part of an acid including the carbonate ion ($CO_3^{2-}$) in carbonic acid, $H_2CO_3$, and the phosphate ion ($PO_4^{3-}$) in phosphoric acid ($H_3PO_4$). Other properties of acids are dealt with in more detail in Topic 13, but here we concentrate on the types of reaction into which acids can enter.

**Reactions involving acids.**

(a) Precipitation reactions.

Both the $H^+$ ions and the anions from acids can participate in reactions. The reactions of the anions from acids are exactly the same as for the same anions supplied by a soluble salt containing the anion. Thus addition of hydrochloric acid to silver nitrate solution produces silver chloride precipitate just like the previous example when solutions of silver nitrate and sodium chloride were mixed. Remaining in solution would be the spectator ions, $H^+$ and $NO_3^-$, which is a solution of nitric acid. The ionic equation for the reaction is

$$H^+(aq) + Cl^-(aq) + Ag^+(aq) + NO_3^-(aq) \rightarrow AgCl(s) + H^+(aq) + NO_3^-(aq)$$

When the spectator ions are deleted, it reduces to

$$Cl^-(aq) + Ag^+(aq) \rightarrow AgCl(s)$$

which is identical to the ionic equation given previously for the reaction between silver nitrate solution and sodium chloride solution. Hence as far as the precipitation reaction is concerned, there is no difference between the source of the chloride ion being a soluble chloride salt or the acid, hydrochloric acid. From this example, it can be seen that any soluble silver salt and any soluble chloride or hydrochloric acid, would react when mixed together in a water solution to precipitate silver chloride. Only the spectator ions will be different.

One advantage of writing ionic equations is that only the actual species reacting are shown.

Another example is the precipitation of the insoluble salt lead(II) sulfate by mixing solutions of the soluble salts lead(II) nitrate and sodium sulfate. The ionic equation for this reaction shown without spectator ions is

$$Pb^{2+} + SO_4^{2-} \rightarrow PbSO_4(s)$$

The spectator ions are nitrate and sodium ions which remain in solution.

Using sulfuric acid as the source of sulfate ions when lead(II) nitrate is mixed with sulfuric acid, the reaction equation is exactly the same as above,

$$Pb^{2+} + SO_4^{2-} \rightarrow PbSO_4(s)$$

The spectator ions in this case are $H^+$ and $NO_3^-$ instead of $Na^+$ and $NO_3^-$. 
(b) Reactions involving the $H^+$ ion.
In the precipitation reactions involving acids discussed above, it is the anion from the acid which is reacting and $H^+$ is the spectator ion. However, there are many reactions of acids where the important ion is $H^+$ while the anion from the acid is merely the spectator ion.

All acids supply hydrogen ions in solution. This ion can participate in a number of common reaction types, recognition of which makes the writing of chemical equations much simpler. These reaction types are:

1. Acids with reactive metals.
Many metals react with acids to form the cation of the metal and produce hydrogen gas, $H_2$, from the $H^+$ ions. These metals are sometimes called "reactive metals", as distinct from the "coinage metals" such as silver, gold, copper, and platinum which are inert to most acids. In particular, hydrochloric acid and sulfuric acid behave in this way with reactive metals, and in the process, will leave the chloride or sulfate ions in the solution. Evaporation of the water from the resulting solution will produce the ionic compound of the metal and the acid used. Ionic compounds are frequently called salts which is a general term and, as pointed out earlier, is not restricted to common table salt, NaCl.

For example, hydrochloric acid added to magnesium metal produces hydrogen which is evolved as a gas and leaves in solution magnesium ions, $Mg^{2+}$ and chloride ions, $Cl^-$.

Formula equation: 

$$Mg + 2HCl \rightarrow MgCl_2 + H_2$$

Ionic equation: 

$$Mg(s) + 2H^+(aq) + 2Cl^-(aq) \rightarrow Mg^{2+}(aq) + H_2(g) + 2Cl^-(aq)$$

or, deleting the $Cl^-$ spectator ions,

$$Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$$

Evaporation of this solution results in isolation of the salt, magnesium chloride.

$$Mg^{2+}(aq) + 2Cl^-(aq) \rightarrow MgCl_2(s)$$

2. Acids with oxides of metals.
All ionic oxide compounds, regardless of whether they are water soluble, react with acids to release the metal ion from the oxide into the solution and also to form water.
As an example, consider the equation for the reaction of hydrochloric acid with calcium oxide.

Formula equation: 

$$CaO + 2HCl \rightarrow CaCl_2 + H_2O$$

Ionic equation: 

$$CaO(s) + 2H^+(aq) + 2Cl^-(aq) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq) + H_2O(l)$$

If the water were evaporated from the solution, the $Ca^{2+}$ and $Cl^-$ ions would crystallise to form the solid compound, calcium chloride.
Note that in the equation above, the chloride ion is a spectator ion. This ion should be omitted, to give for the correct equation

\[ \text{CaO(s)} + 2\text{H}^+(\text{aq}) \rightarrow \text{Ca}^{2+}(\text{aq}) + \text{H}_2\text{O(l)} \]

**an acid + an oxide forms a salt and water.**

### 3. Acids with hydroxides of metals.

In the same way as the \( \text{H}^+ \) ions of acids react with metal oxides to form a salt and water, so too can acids react with hydroxide compounds of metals, regardless of whether they are water soluble, to also form a salt and water. As in the case of oxides, if the salt formed is soluble in water, then the resulting solution must be evaporated in order to isolate the solid compound.

For example, nitric acid reacts with copper(II) hydroxide to form water and copper(II) nitrate in solution.

**Formula equation:** \( \text{Cu(OH)}_2 + 2\text{HNO}_3 \rightarrow \text{Cu(NO}_3)_2 + 2\text{H}_2\text{O} \)

**Ionic:** \( \text{Cu(OH)}_2(\text{s}) + 2\text{H}^+(\text{aq}) + 2\text{NO}_3^-(\text{aq}) \rightarrow 2\text{H}_2\text{O(l)} + \text{Cu}^{2+}(\text{aq}) + 2\text{NO}_3^-(\text{aq}) \)

Again, the \( \text{NO}_3^- \) (aq) ion is a spectator ion and should be deleted from both sides of the equation, as the \( \text{H}^+ \) (aq) is the only component of the acid which reacts with the \( \text{Cu(OH)}_2(\text{s}) \).

\[ \text{Cu(OH)}_2(\text{s}) + 2\text{H}^+(\text{aq}) \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{H}_2\text{O(l)} \]

Evaporation of this solution would produce the salt, \( \text{Cu(NO}_3)_2 \), according to the equation

\[ \text{Cu}^{2+}(\text{aq}) + 2\text{NO}_3^-(\text{aq}) \rightarrow \text{Cu(NO}_3)_2(\text{s}) \]

If the hydroxide reacting with the acid is soluble in water and a solution of that hydroxide is specified, then the cation of the hydroxide is also a spectator ion and should be deleted along with the anion from the acid. For example, hydrochloric acid reacting with a solution of sodium hydroxide would be represented by the ionic equation

\[ \text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O(l)} \]

The salt \( \text{NaCl(s)} \) could be obtained by evaporating the solution.

**an acid + a hydroxide forms a salt and water**

or, combining this with the previous section,

**an acid + an oxide or hydroxide forms a salt and water.**
4. Acids with carbonates.
Another common reaction of an acid leads to the evolution of a gas when the acid is mixed with the metal salts of certain anions. This type of reaction is typified by the well-known case of acids reacting with carbonates to produce carbon dioxide gas, water and a salt. Again, the reaction occurs regardless of whether the compound reacting with the acid is soluble or insoluble in water. For example, sulfuric acid reacts with solid magnesium carbonate to produce carbon dioxide gas and magnesium sulfate in solution as follows:

Formula equation: \( \text{MgCO}_3 + \text{H}_2\text{SO}_4 \rightarrow \text{MgSO}_4 + \text{CO}_2 + \text{H}_2\text{O} \)

Ionic: \( \text{MgCO}_3(\text{s}) + 2\text{H}^+(\text{aq}) \rightarrow \text{Mg}^{2+}(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \)

Note that the spectator ion, \( \text{SO}_4^{2-}(\text{aq}) \) has been deleted from both sides of the ionic equation.

If the carbonate compound used were soluble and was already in solution, then the ionic equation would not include the metal ion because it would be a spectator ion, as in the following example of adding nitric acid to a solution of sodium carbonate:

\[
2\text{Na}^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) + 2\text{H}^+(\text{aq}) + 2\text{NO}_3^-(\text{aq}) \quad \rightarrow \\
2\text{Na}^+(\text{aq}) + 2\text{NO}_3^-(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})
\]

which simply becomes, after deleting spectator ions,

\[
\text{CO}_3^{2-}(\text{aq}) + 2\text{H}^+(\text{aq}) \quad \rightarrow \quad \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})
\]

**Summary of reactions of acids.**

<table>
<thead>
<tr>
<th>An acid reacts with</th>
<th>To produce</th>
</tr>
</thead>
<tbody>
<tr>
<td>a reactive metal</td>
<td>hydrogen gas + the salt of the metal and the acid</td>
</tr>
<tr>
<td>an oxide or hydroxide of a metal</td>
<td>water + the salt of the metal and the acid</td>
</tr>
<tr>
<td>a carbonate of a metal</td>
<td>carbon dioxide gas + the salt of the metal and the acid</td>
</tr>
</tbody>
</table>

**Check your understanding of this section:**
By what criterion would you decide that a species is an acid?
What is the expected reaction when an acid is mixed with a reactive metal?
What happens when an acid is placed with an insoluble oxide or hydroxide?
When an acid is placed on a compound such as calcium carbonate, what is observed?
Objectives of this Topic.

When you have completed this Topic, including the tutorial questions, you should have achieved the following goals:

1. Know the meaning of the terms solvent; solute; solution; saturated solution; super saturated solution; aquated ions; volatile; non-volatile; spectator ions; precipitate; filtrate; electronegative atom; polar bond; polar molecule; non-polar bond; non-polar molecule.
2. Understand the concept of polarity in covalent molecules, especially as it relates to the role of water as a solvent for ionic compounds.
3. Be able to write ionic equations for any dissolution process or any crystallisation of a salt.
4. Know how to find out if any given ionic compound is soluble.
5. Be able to write ionic equations for any precipitation of a salt.
6. By use of solubility tables, be able to predict whether any given combination of solutions of salts will lead to formation of a precipitate.
7. Recognise acids as a source of hydrogen ions.
8. Know the names and formulas for several common acids.
9. Recognise acids as being able to precipitate insoluble salts by supplying the relevant anion.
10. Know the reactions of acids with (i) reactive metals; (ii) oxides and hydroxides; (iii) carbonates.
11. Be able to write ionic equations for each of the above reaction types.
### SOLUBILITY TABLE OF SOME COMMON SALTS

<table>
<thead>
<tr>
<th>ANION</th>
<th>BEHAVIOUR WITH COMMON CATIONS</th>
<th>SLIGHTLY SOLUBLE</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>USUAL</td>
<td>EXCEPTIONAL</td>
</tr>
<tr>
<td>F⁻</td>
<td>soluble</td>
<td>except Mg²⁺, Ca²⁺, Sr²⁺, Ba²⁺, Mn²⁺, Pb²⁺</td>
</tr>
<tr>
<td>Cl⁻</td>
<td>soluble</td>
<td>except Ag⁺, Hg₂²⁺</td>
</tr>
<tr>
<td>Br⁻</td>
<td>soluble</td>
<td>except Ag⁺, Hg₂²⁺</td>
</tr>
<tr>
<td>I⁻</td>
<td>soluble</td>
<td>except Ag⁺, Hg₂⁺, Hg²⁺, Pb²⁺</td>
</tr>
<tr>
<td>SO₄²⁻</td>
<td>soluble</td>
<td>except Sr²⁺, Ba²⁺, Hg₂⁺, Pb²⁺</td>
</tr>
<tr>
<td>NO₃⁻</td>
<td>soluble</td>
<td></td>
</tr>
<tr>
<td>CH₃CO₂⁻</td>
<td>soluble</td>
<td></td>
</tr>
<tr>
<td>OH⁻</td>
<td>insoluble</td>
<td>except Na⁺, K⁺, Ba²⁺</td>
</tr>
<tr>
<td>SO₃²⁻</td>
<td>insoluble</td>
<td>except Na⁺, K⁺, NH₄⁺, Mg²⁺</td>
</tr>
<tr>
<td>PO₄³⁻</td>
<td>insoluble</td>
<td>except Na⁺, K⁺, NH₄⁺</td>
</tr>
<tr>
<td>CO₃²⁻</td>
<td>insoluble</td>
<td>except Na⁺, K⁺, NH₄⁺</td>
</tr>
<tr>
<td>C₂O₄²⁻</td>
<td>insoluble</td>
<td>except Na⁺, K⁺</td>
</tr>
<tr>
<td>S²⁻</td>
<td>insoluble</td>
<td>except Na⁺, K⁺, NH₄⁺, Mg²⁺, Ca²⁺, Sr²⁺, Ba²⁺</td>
</tr>
</tbody>
</table>

The ammonium ion, NH₄⁺, is not an element but is a polyatomic cation which also forms salts like any normal metal cation.
SUMMARY

Ionic compounds consist of very large aggregates of cations and anions packed into a crystal lattice so as to maximise the attractive forces between cations and anions and minimise the repulsive cation/cation and anion/anion forces. Despite the strength of ionic bonds, many ionic compounds can dissolve in solvents such as water to form a solution, a process best represented by an ionic equation. Ionic equations show the species involved in the actual form in which they are present, so ions in solution are shown as such. Ionic equations require not only that the mass should balance on both sides but also the charge. Otherwise, ionic equations follow the same rules as for formula equations.

The reason why ionic compounds may dissolve in water is that water is a polar molecule in which the oxygen atoms attract more than their share of the bonding electrons (said to be more electronegative) than the hydrogen atoms, leaving a slight negative charge on the O atoms and a slight positive charge on the H atoms. This uneven charge distribution allows water molecules to be attracted to both cations and anions in the ionic crystal, releasing much energy in the process. If enough energy is released through these attractions, the ions may leave the crystal lattice and go into solution as aquated ions, each surrounded by a sheath of water molecules. These aquated ions are free to move around throughout the solution. Once dissolved, ionic bonds cease to exist as this form of bonding is really only present in the solid state. Not all ionic compounds are soluble and solubility tables listing this data are available.

If the solvent is boiled off from a solution of an ionic compound, the ions, stripped of their surrounding water molecules, are non-volatile and remain to recombine as crystals of the solid compound.

If cations and anions from two different salts are mixed in the same solution and any combination of them corresponds to an insoluble salt, then that compound will be precipitated from the solution. Other cations and anions present will remain in solution and are termed spectator ions. The precipitated solid can be collected by filtration and the remaining solution (the filtrate) can be boiled down to collect any soluble compounds. This type of reaction is best represented by an ionic equation in which only the reacting ions are shown and the spectator ions are deleted as they are present on both sides of the reaction equation.

Acids are compounds that release the H+ ion in solution. This ion would strictly be just a proton with no electron, but this species would have too large a charge on too small a volume to be stable, so it associates with water molecules via a lone pair of electrons on the O atom and thus can also be represented as H+(aq) or H3O+. Apart from the anion from the acid in solution, reactions of the H+ ion are conveniently grouped into several classes.

Acids on reactive metals: form a salt and hydrogen gas.
Acids on oxides of metals: form a salt and water.
Acids on hydroxides of metals: form a salt and water.
Acids on a carbonate: form a salt and water and carbon dioxide gas.

In each case, the anion in the salt formed is the anion that belonged to the acid, for example, chloride from hydrochloric acid or nitrate from nitric acid. Each of these reaction types are conveniently represented by ionic equations which more clearly show the process which is taking place than overall formula equations.
TUTORIAL QUESTIONS - TOPIC 6.

1. **Reactions of acids with reactive metals.**
   
   (a) The tutor will show the following:
   hydrochloric acid is added to magnesium metal.
   Observations:
   
   (i) Write the formula equation
   
   (ii) Write the ionic equation

   (b) The tutor will show the following:
   hydrochloric acid is added to zinc metal.
   Observations:
   
   (i) Write the formula equation
   
   (ii) Write the ionic equation

   (c) The tutor will show the following:
   sulfuric acid is added to zinc metal.
   Observations:
   
   (i) Write the formula equation
   
   (ii) Write the ionic equation

2. **Reactions of acids with oxides and hydroxides of metals.**
   
   (a) The tutor will show the following:
   sulfuric acid is added to solid sodium hydroxide.
   Observations:
   
   (i) Write the formula equation
   
   (ii) Write the ionic equation
(b) The tutor will show the following:
hydrochloric acid is added to zinc oxide.

Observations:

(i) Write the formula equation

(ii) Write the ionic equation

3. **Reactions of acids with carbonates of metals.**
(a) The tutor will show the following:
hydrochloric acid is added to solid sodium carbonate.
Observations:

(i) Write the formula equation

(ii) Write the ionic equation

(b) The tutor will show the following:
sulfuric acid is added to copper(II) carbonate.
Observations:

(i) Write the formula equation

(ii) Write the ionic equation

(c) The tutor will show the following:
nitric acid is added to calcium carbonate.
Observations:

(i) Write the formula equation

(ii) Write the ionic equation
4. Precipitation reactions.

(a) The tutor will show the following solutions being mixed:
   a water solution of sodium sulfate and a water solution of barium chloride.
   Observations:
   (i) Write the formula equation
   (ii) Write the ionic equation

(b) The tutor will show the following solutions being mixed:
   a water solution of silver nitrate and a water solution of sodium chloride.
   Observations:
   (i) Write the formula equation
   (ii) Write the ionic equation

(c) The tutor will show the following solutions being mixed:
   a water solution of potassium iodide and a water solution of lead(II) nitrate.
   Observations:
   (i) Write the formula equation
   (ii) Write the ionic equation

(d) The tutor will show the following solutions being mixed:
   a water solution of potassium iodide and a water solution of sodium chloride.
   Observations:
   Explanation:

(e) The tutor will show the following solutions being mixed:
   a water solution of cobalt(II) nitrate and a water solution sodium carbonate.
   Observations:
   (i) Write the formula equation
(ii) Write the ionic equation

5. Write **ionic** equations for the reaction of magnesium carbonate with each of the acids - hydrochloric acid, nitric acid, sulfuric acid, hydrobromic acid, hydriodic acid. What do you notice about the equations? Explain your observation.

6. Explain the meaning of the following terms:
solvent; solution; precipitate; filtrate; formula equation; ionic equation; volatile substance.

7. Write equations for the dissolution of the following ionic compounds in water.

<table>
<thead>
<tr>
<th>Compound</th>
</tr>
</thead>
<tbody>
<tr>
<td>cobalt(II) sulfate</td>
</tr>
<tr>
<td>ammonium nitrate</td>
</tr>
<tr>
<td>iron(III) chloride</td>
</tr>
<tr>
<td>sodium phosphate</td>
</tr>
<tr>
<td>aluminium nitrate</td>
</tr>
<tr>
<td>potassium carbonate</td>
</tr>
<tr>
<td>iron(II) bromide</td>
</tr>
</tbody>
</table>
8. Give the missing formula or name for each entry in the following table.

<table>
<thead>
<tr>
<th>FORMULA</th>
<th>NAME</th>
</tr>
</thead>
<tbody>
<tr>
<td>CuBr₂</td>
<td>cobalt(II) chloride</td>
</tr>
<tr>
<td>BaSO₄</td>
<td>magnesium sulfate</td>
</tr>
<tr>
<td>NH₄I</td>
<td>lead(II) carbonate</td>
</tr>
<tr>
<td>Mn(NO₃)₂</td>
<td>strontium sulfite</td>
</tr>
<tr>
<td>AgF</td>
<td>aluminium nitrate</td>
</tr>
<tr>
<td>NaCl</td>
<td>silver oxalate</td>
</tr>
<tr>
<td>Fe₃(PO₄)₂</td>
<td>tin(II) acetate</td>
</tr>
<tr>
<td>FePO₄</td>
<td>iron(II) bromide</td>
</tr>
<tr>
<td>ZnCO₃</td>
<td>sodium nitrite</td>
</tr>
<tr>
<td>K₂CO₃</td>
<td>iron(III) bromide</td>
</tr>
<tr>
<td>Ca(CH₃CO₂)₂</td>
<td>nickel(II) iodide</td>
</tr>
<tr>
<td>Zn(HCO₃)₂</td>
<td>lithium carbonate</td>
</tr>
</tbody>
</table>
9. What is a "salt"? Which of the following compounds are salts and how did you decide in each case?

\[ \text{Na}_3\text{PO}_4, \text{SO}_3, \text{BaBr}_2, \text{NH}_3, \text{Mg(NO}_3\text{)_2}, \text{CO}_2, \text{H}_2\text{O}, \text{AgCl}, \text{CCl}_4, \text{H}_2\text{SO}_4, \text{H}_2\text{S}, \text{HNO}_3 \]

**ANSWERS TO TUTORIAL TOPIC 6**

When writing ionic equations, make sure that you have given the correct number of each ion on both sides of the equation for any compound that contains more than one of a given ion - all equations must balance in the same way as formula equations.

1. (a) hydrochloric acid is added to magnesium.

Observations: Metal dissolved with gas evolved - pop test shows hydrogen; A colourless solution resulted.

(i) Formula equation

\[ \text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2 \]

(ii) Ionic equation

\[ \text{Mg}(s) + 2\text{H}^+(aq) \rightarrow \text{Mg}^{2+}(aq) + \text{H}_2(g) \]

(b) hydrochloric acid is added to zinc.

Observations: The metal dissolved with a gas evolved - pop test shows hydrogen; formed a colourless solution.

(i) Formula equation

\[ \text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2 \]

(ii) Ionic equation

\[ \text{Zn}(s) + 2\text{H}^+(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{H}_2(g) \]

(c) sulfuric acid is added to zinc.

Observations: The metal dissolved and a gas evolved - pop test shows hydrogen; formed a colourless solution.

(i) Formula equation

\[ \text{Zn} + \text{H}_2\text{SO}_4 \rightarrow \text{ZnSO}_4 + \text{H}_2 \]
2. Reactions of acids with oxides and hydroxides of metals.

(a) Sulfuric acid is added to solid sodium hydroxide.

Observations: The solid dissolved to form a colourless solution; test tube warmed by the reaction.

(i) Formula equation
\[ 2\text{NaOH} + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O} \]

(ii) Ionic equation
\[ \text{NaOH}(s) + \text{H}^+(aq) \rightarrow \text{Na}^+(aq) + \text{H}_2\text{O}(l) \]

(b) Hydrochloric acid is added to zinc oxide.

Observations: The solid dissolved to form a colourless solution; test tube warmed by the reaction.

(i) Formula equation
\[ \text{ZnO} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2\text{O} \]

(ii) Ionic equation
\[ \text{ZnO}(s) + 2\text{H}^+(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{H}_2\text{O}(l) \]

3. Reactions of acids with carbonates of metals.

(a) Hydrochloric acid is added to solid sodium carbonate.

Observations: The solid dissolved to form a colourless solution and gas released - pop test negative.

(i) Formula equation
\[ \text{Na}_2\text{CO}_3 + 2\text{HCl} \rightarrow 2\text{NaCl} + \text{H}_2\text{O} + \text{CO}_2 \]

(ii) Ionic equation
\[ \text{Na}_2\text{CO}_3(s) + 2\text{H}^+(aq) \rightarrow 2\text{Na}^+(aq) + \text{H}_2\text{O}(l) + \text{CO}_2(g) \]
(b) sulfuric acid is added to copper(II) carbonate.

Observations: The solid dissolved to form a blue solution and gas released - pop test negative.

(i) Formula equation

$$\text{CuCO}_3 + \text{H}_2\text{SO}_4 \rightarrow \text{CuSO}_4 + \text{H}_2\text{O} + \text{CO}_2$$

(ii) Ionic equation

$$\text{CuCO}_3(\text{s}) + 2\text{H}^+(\text{aq}) \rightarrow \text{Cu}^{2+}(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$$

(c) nitric acid is added to calcium carbonate.

Observations: The solid dissolved to form a colourless solution and gas released - pop test negative.

(i) Formula equation

$$\text{CaCO}_3 + 2\text{HNO}_3 \rightarrow \text{Ca(NO}_3)_2 + \text{H}_2\text{O} + \text{CO}_2$$

(ii) Ionic equation

$$\text{CaCO}_3(\text{s}) + 2\text{H}^+(\text{aq}) \rightarrow \text{Ca}^{2+}(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$$

4. Precipitation reactions.

(a) a water solution of sodium sulfate and a water solution of barium chloride.

Observations: A white precipitate formed.

(i) Formula equation

$$\text{Na}_2\text{SO}_4 + \text{BaCl}_2 \rightarrow \text{BaSO}_4 + 2\text{NaCl}$$

(ii) Ionic equation

$$\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})$$

[Consider the formulas of the two possible compounds that might form, \(\text{BaSO}_4\) and \(\text{NaCl}\). Only barium sulfate is insoluble, so it will precipitate while sodium ions and chloride ions remain in solution.]

(b) a water solution of silver nitrate and a water solution of sodium chloride.

Observations: A white precipitate formed.
(i) Formula equation

$$\text{AgNO}_3 + \text{NaCl} \rightarrow \text{AgCl} + \text{NaNO}_3$$

(ii) Ionic equation

$$\text{Ag}^{+}(\text{aq}) + \text{Cl}^{-}(\text{aq}) \rightarrow \text{AgCl}(s)$$

[The combination of silver(I) ions and chloride ions results in an insoluble compound, silver(I) chloride, being precipitated while the sodium and nitrate ions remain in solution.]

(c) a water solution of potassium iodide and a water solution of lead(II) nitrate.

Observations: A bright yellow precipitate formed.

(i) Formula equation

$$2\text{KI} + \text{Pb(NO}_3)_2 \rightarrow 2\text{KNO}_3 + \text{PbI}_2$$

(ii) Ionic equation

$$2\text{I}^{-}(\text{aq}) + \text{Pb}^{2+}(\text{aq}) \rightarrow \text{PbI}_2(s)$$

[The potassium iodide provides K\(^{+}\)(aq) and I\(^{-}\)(aq) ions in the solution and Pb\(^{2+}\)(aq) and NO\(_3\)\(^{-}\)(aq) ions come from the lead(II) nitrate solution. The combination of Pb\(^{2+}\) and I\(^{-}\) ions results in the formation of lead(II) iodide precipitate. The potassium and nitrate ions remain in solution.]

(d) a water solution of potassium iodide and a water solution of sodium chloride.

Observations: No reaction observed. No temperature rise detected.

Explanation: The ions in the mixture of the two solutions are K\(^{+}\)(aq), I\(^{-}\)(aq), Na\(^{+}\)(aq) and Cl\(^{-}\)(aq). There is no combination of these cations and anions that form an insoluble compound.

(e) water solution of cobalt(II) nitrate and a water solution sodium carbonate.

Observations: A purple precipitate formed.

(i) Formula equation

$$\text{Co(NO}_3)_2 + \text{Na}_2\text{CO}_3 \rightarrow \text{CoCO}_3 + 2\text{NaNO}_3$$

(ii) Ionic equation

$$\text{Co}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{CoCO}_3(s)$$

[The combination of cobalt(II) ions and carbonate ions results in an insoluble
compound, cobalt(II) carbonate, being precipitated while the sodium and nitrate ions remain in solution.]

5. $\text{MgCO}_3(s) + 2\text{H}^+(aq) \rightarrow \text{Mg}^{2+}(aq) + \text{CO}_2(g) + \text{H}_2\text{O}(l)$

The equations are identical because the reaction does not involve the chloride, nitrate, sulfate, bromide or iodide ions which are merely spectator ions in each case.

6. **Solvent:** Substance in which another substance (solute) dissolves, the combination being termed a solution. Solvents are usually but not necessarily liquids.

**Solution:** A homogeneous combination of two or more substances which retain their separate identities.

**Precipitate:** A solid formed in a chemical reaction.

**Filtrate:** The liquid phase that remains after any solids previously present have been removed, for example by filtering.

**Formula equation:** A representation of a chemical reaction in which all the reactants and products are represented by their entire formulas.

**Ionic equation:** A representation of a chemical reaction in which any species which are present as ions are shown as such. Only the species actually reacting are shown and any other ions are called spectator ions because they play no part in the actual reaction.

**Volatile substance:** Element or compound that can be readily converted to the gaseous state.

7. $\text{CoSO}_4(s) \rightarrow \text{Co}^{2+}(aq) + \text{SO}_4^{2-}(aq)$

$\text{NH}_4\text{NO}_3(s) \rightarrow \text{NH}_4^+(aq) + \text{NO}_3^-(aq)$

$\text{FeCl}_3(s) \rightarrow \text{Fe}^{3+}(aq) + 3\text{Cl}^-(aq)$ (note $3\text{Cl}^-$ required to balance)

$\text{Na}_3\text{PO}_4(s) \rightarrow 3\text{Na}^+(aq) + \text{PO}_4^{3-}(aq)$ (note $3\text{Na}^+$ required to balance)

$\text{Al(NO}_3)_3(s) \rightarrow \text{Al}^{3+}(aq) + 3\text{NO}_3^-(aq)$ (note $3\text{NO}_3^-$ required to balance)

$\text{K}_2\text{CO}_3(s) \rightarrow 2\text{K}^+(aq) + \text{CO}_3^{2-}(aq)$ (note $2\text{K}^+$ required to balance)

$\text{FeBr}_2(s) \rightarrow \text{Fe}^{2+}(aq) + 2\text{Br}^-(aq)$ (note $2\text{Br}^-$ required to balance)
8. Formulas: \( \text{CoCl}_2; \text{MgSO}_4; \text{PbCO}_3; \text{SrSO}_3; \text{Al(NO}_3)_3; \text{Ag}_2\text{C}_2\text{O}_4; \text{Sn(CH}_3\text{CO}_2)_2; \text{FeBr}_2; \text{NaNO}_2; \text{FeBr}_3; \text{NiI}_2; \text{Li}_2\text{CO}_3. \)

Names: copper(II) bromide; barium sulfate; ammonium iodide; 
manganese(II) nitrate; silver(I) fluoride; sodium chloride; 
iron(II) phosphate; iron(III) phosphate; zinc carbonate; 
potassium carbonate; calcium acetate; zinc hydrogencarbonate; 
lithium hydroxide; copper(II) sulfate; lead(II) nitrate; 
lead(II) nitrite; lithium sulfite.

Note the use of Roman numerals in conjunction with those cations that can 
occur with more than one cationic charge, e.g. copper.

9. A salt is any ionic compound in which the cation is not \( \text{H}^+ \) or the anion is not \( \text{OH}^- \). The compounds \( \text{Na}_3\text{PO}_4, \text{BaBr}_2, \text{Mg(NO}_3)_2, \text{AgCl} \) are all salts 
because they meet these criteria. The other compounds are either covalently 
bonded compounds involving only non-metals (\( \text{SO}_3, \text{NH}_3, \text{CO}_2, \text{H}_2\text{O}, \text{CCl}_4, \text{H}_2\text{S} \)) or acids (\( \text{HNO}_3, \text{H}_2\text{SO}_4 \)).