Acids, Alkalis and Carbonates

Chemistry is the study of how atoms of the various different elements interact. Fortunately they do interact, otherwise there would be no chemistry and, amongst the many consequences, we wouldn’t be here to ask questions.

In more detail, the interaction is between the negatively charged electrons that surround the nucleus of every atom. As two atoms approach each other, the electrons of one feel additional electrical and magnetic forces due to the nearby presence of the other’s electrons. As a result of these additional forces, the electrons will move into a different configuration. If that configuration has less energy than the original one, the two atoms will tend to stay together, as you would have to add energy to separate them and return them to their original configurations. That, in essence, is the origin of the chemical bonds that hold atoms together to form molecules.

For simplicity, let us consider a molecule that contains just two atoms. In this molecule, the electron configuration of one atom will overlap with that of the other. In some sense, the atoms are sharing some of their electrons\(^1\). Only rarely will this sharing be precisely equal; the more usual case is for one atom to take a greater share than the other. The atom taking the greater share of the electron configuration will therefore end up having a slightly negative charge, with the other atom having an equal and opposite overall positive charge.

There are many cases where this unbalanced sharing is taken to the extreme and one atom takes complete control of one or more electrons that originally belonged to the other\(^2\). In such a case it is easiest to picture the molecule as made up of two charged atoms, or ‘ions’, with equal and opposite charges, held together by their mutual attraction. If such a substance is a solid (and, at room temperature, most of them are) its crystals are made up of a lattice of ions of alternating type and charge.

When you put such a solid in water, the surrounding water acts to weaken the attraction between the ions (it would take too long to explain how, but it happens – water is amazing stuff). If the attraction is weakened enough, ions at the surface drift away into the water and the solid dissolves. The solution consists of a mixture of the individual ions, some positive and some negative, drifting around independently in the surrounding water, a bit like the molecules of a gas. The substance is in a state where it is said to be ionised.

Incidentally, that’s where ‘de-ionised’ water – the stuff you use to top up your car battery, or fill up your steam iron – gets its name. The de-ionisation process is just the removal of all dissolved minerals (present as ions, in solution) that would otherwise disrupt the workings of the battery or clog up the steam iron with ‘scale’.

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\(^1\) Chemists describe this type of bonding as ‘covalent’.

\(^2\) ‘Electrovalent’, or ‘ionic’ bonding.
Acids and Alkalis

What makes something an acid or an alkali? To understand this, it’s best to start with pure (de-ionised!) water, which is neither!

Nor is it totally free of ions. In addition to weakening the forces of attraction between the atoms of other molecules, water has the same effect on itself (as I said earlier, it is amazing stuff). As a result, a very small proportion of the H$_2$O molecules of which it is composed break into a pair of ions:

$$\text{H}_2\text{O} \leftrightarrow \text{H}^+ + \text{OH}^-$$

An equilibrium is set up, with the rate of formation of the ions being exactly equal to the rate at which they recombine. At equilibrium, about one molecule in ten million is ionised, so the concentration of H$^+$ ions (and of OH$^-$ ions) is one in ten to the power 7 (or $10^{-7}$). The equilibrium process keeps the product of the two concentrations constant (at $10^{-14}$). If, for example, you could find a way to double the OH$^-$ concentration, the equilibrium point would shift to halve the H$^+$ concentration and keep the product constant. In any case, the number of ions present in pure water is sufficiently small that it is a very poor conductor of electricity.

Suppose we dissolve some hydrogen chloride gas in the water. The hydrogen chloride ionises completely in solution:

$$\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$$

You can tell that the number of ions has increased by the fact that the solution is a good conductor of electricity. As well as the Cl$^-$ ions, we have added an equal number of H$^+$ ions. This increases their concentration and so the concentration of OH$^-$ ions must fall, to preserve the balance. Now a solution of hydrogen chloride gas in water has acidic properties – it will, for example, dissolve magnesium metal, liberating hydrogen gas – and is known as hydrochloric acid.

Suppose, instead, we dissolve some sodium hydroxide in pure water. Again, the sodium hydroxide ionises completely in solution:

$$\text{NaOH} \rightarrow \text{Na}^+ + \text{OH}^-$$

and, as in the previous case, the solution will conduct electricity. In this case we have added OH$^-$ (and Na$^+$) ions, so the concentration of H$^+$ ions must fall. A solution of sodium hydroxide shows alkaline properties – it feels soapy on the skin (don’t try this at home unless the solution is extremely weak, and then wash your hands immediately, or you might lose some skin) and will neutralise acids.

The relationship between the ions and the acidity or alkalinity of the solution is not a coincidence. It turns out that an acid is a substance that creates an excess of H$^+$ ions (and a deficit of OH$^-$ ions) in water, and an alkali (or base, in the chemist’s jargon) is a substance that does the opposite. You can measure how acidic a solution is by determining the H$^+$ concentration. If the concentration turns out to be $10^{-7}$, the solution is neutral, if it is greater – say, by a factor of 100 (i.e. $10^{-5}$) – the solution is acidic and if the concentration is less – again, say, by a factor of 100 (i.e. $10^{-9}$) – the solution will be alkaline (or basic).
Incidentally, the numeric value of these powers, ignoring the minus sign, is what you measure when you determine the pH value: a neutral solution has a pH of 7, a smaller value (0 to 6) means the solution is acidic and a greater value (8 to 14) means that it is alkaline.

Neutralisation
I mentioned earlier that a base (or alkali) will neutralise an acid. The standard recipe, that I had to learn by heart in my school chemistry lessons, is:

\[
\text{an acid} + \text{a base} \rightarrow \text{a salt} + \text{water}
\]

Suppose we mix, in the right proportions, solutions of sodium hydroxide and hydrochloric acid. The ions that are initially present will be \(\text{Na}^+, \text{OH}^-, \text{H}^+\) and \(\text{Cl}^-\) with equal concentrations of each. However, both the \(\text{H}^+\) and \(\text{OH}^-\) will be present in far greater concentrations than the equilibrium values, so they will immediately combine to form neutral water molecules:

\[
\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}
\]

and all that will be left in the solution (apart from the tiny equilibrium amount of ionised water) is \(\text{Na}^+\) and \(\text{Cl}^-\) ions, in other words a solution of sodium chloride – common salt.

The overall reaction is:

\[
\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}
\]

which conforms with the above recipe.

The solution no longer contains excessive concentrations of either \(\text{H}^+\) or \(\text{OH}^-\) ions and so will be neither acid nor alkaline. The acid and the base have ‘neutralised’ each other and the solution will have a pH very close to 7.

Strong and weak
Some acids (such as hydrochloric acid) are said to be ‘strong’ and others (such as boric acid) are said to be ‘weak’. The difference is in how they ionise when dissolved in water. Strong acids are fully ionised in solution and weak acids are only partially ionised. A weak acid is weak because it doesn’t increase the \(\text{H}^+\) ion concentration as much as an equivalent amount of a strong acid does.

The same applies to bases. Sodium hydroxide is a strong base because it is fully ionised in solution, but an ammonia solution:

\[
\text{NH}_3 + \text{H}_2\text{O} \rightarrow \text{NH}_4\text{OH}
\]

is a weak base because only some of the ammonium hydroxide is ionised:

\[
\text{NH}_4\text{OH} \leftrightarrow \text{NH}_4^+ + \text{OH}^-
\]

What happens if you mix a strong acid like HCl with a weak base, such as ammonia solution? Well, the reaction will take place in the way described in the previous section:
HCl + NH₄OH \rightarrow NH₄Cl + H₂O

and we end up with water containing NH₄⁺ and Cl⁻ ions (and the tiny equilibrium amounts of H⁺ and OH⁻ ions) in other words a solution of ammonium chloride.

There is, however, one slight – but important – difference. Remember that ammonium hydroxide is a weak base, so some of those NH₄⁺ ions will recombine with some of the tiny amount of OH⁻ ions still present in a neutral solution:

\[ \text{NH}_4^+ + \text{OH}^- \rightleftharpoons \text{NH}_3\text{OH} \]

Reducing the concentration of OH⁻ ions means that the H⁺ concentration will be forced to increase, to maintain the product of H⁺ and OH⁻ concentrations at the equilibrium value of $10^{-14}$. The solution will therefore contain more H⁺ ions than are found in pure water, which means it will have a pH of less than 7 and must, therefore, be acidic.

In general, all salts formed from a strong acid and a weak base will form acid solutions. By a similar argument, salts formed from a weak acid and a strong base will, in solution, be alkaline.

**Carbonic Acid and the Carbonates**

Carbonic acid is formed when carbon dioxide dissolves in water:

\[ \text{H}_2\text{O} + \text{CO}_2 \rightleftharpoons \text{H}_2\text{CO}_3 \]

It is very unstable, readily decomposing back into water and carbon dioxide, and can exist only in dilute solution. It partially ionises, in two stages:

\[ \text{H}_2\text{CO}_3 \rightleftharpoons \text{H}^+ + \text{HCO}_3^- \text{ and then } \text{HCO}_3^- \rightleftharpoons \text{H}^+ + \text{CO}_3^{2-} \]

and is therefore a weak acid.

It forms two families of salts, the bicarbonates (or as chemists now prefer, ‘hydrogen carbonates’)

\[ \text{NaOH} + \text{H}_2\text{CO}_3 \rightarrow \text{NaHCO}_3 + \text{H}_2\text{O} \]

and the carbonates

\[ 2\text{NaOH} + \text{H}_2\text{CO}_3 \rightarrow \text{Na}_2\text{CO}_3 + 2\text{H}_2\text{O} \]

depending on whether one or both H⁺ ions are replaced.

The carbonates are always more stable than the bicarbonates, some of which are not known and several others only exist in solution, and all bicarbonates are easily decomposed (to the carbonate, liberating CO₂ and water) by moderate heat. That’s why baking soda (which is pure sodium bicarbonate) causes things to rise when you cook them.

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³ Bicarbonates were originally given this name because they contain twice as much ‘carbonate’ as the normal carbonate salt, e.g. for sodium carbonate CO₃²⁻/Na =1.3, for sodium bicarbonate CO₃²⁻/Na = 2.6.
As a salt of a strong base and a weak acid, sodium carbonate in solution is quite strongly alkaline, as are all the soluble carbonates. Barium and strontium carbonates would be alkaline in solution, except for the fact that both are virtually insoluble.

Bicarbonates, which are effectively made by only half-neutralising the carbonic acid, are less alkaline; sodium bicarbonate solution, for example, is close to neutral (a pH somewhere around 7 to 8).

Carbonates vary in their stability. Potassium and sodium carbonates can be strongly heated without decomposing. Magnesium carbonate is easily decomposed by moderate heat and calcium carbonate by strong heating. Strontium and barium carbonates are increasingly harder to decompose.

Many metals do not form simple carbonates. The so-called copper carbonate is really a basic carbonate – a more complex molecule that can appear in at least two different forms, with formulae \( \text{Cu}_2(\text{OH})_2\text{CO}_3 \) and \( \text{Cu}_3(\text{OH})_2(\text{CO}_3)_2 \). If you try to make ‘copper carbonate’ by adding sodium or potassium carbonate to a solution of copper sulphate (or copper chloride), you end up with an insoluble mixture of one or both of the basic carbonates with a variable amount of basic copper sulphate (or basic copper chloride).

All carbonates, whether soluble or not, are attacked by acids. For example, if you mix a solution of sodium carbonate with dilute hydrochloric acid, you get a mixture of \( \text{Na}^+ \), \( \text{CO}_3^{2-} \), \( \text{H}^+ \) and \( \text{Cl}^- \) ions. Since carbonic acid is weak, some of the will combine with the \( \text{CO}_3^{2-} \) ions to form undissociated carbonic acid:

\[
2\text{H}^+ + \text{CO}_3^{2-} \rightleftharpoons \text{H}_2\text{CO}_3
\]

which, being unstable, will largely decompose into water and carbon dioxide gas:

\[
\text{H}_2\text{CO}_3 \rightleftharpoons \text{H}_2\text{O} + \text{CO}_2
\]

The standard way of describing this reaction is the chemical equation:

\[
\text{Na}_2\text{CO}_3 + 2\text{HCl} \rightarrow 2\text{NaCl} + \text{H}_2\text{O} + \text{CO}_2
\]

This will happen with any acid that is stronger and/or more stable than carbonic acid – which is pretty much all of them, including boric acid.

If an acid exists as a solid substance, it can be mixed with a dry carbonate without reacting. However, if you add water, some of the acid will dissolve\(^4\) and start reacting, even if the carbonate is insoluble. Baking powder, which is a mixture of sodium bicarbonate and tartaric acid, shows this property; you can keep it for years in your kitchen cupboard but as soon as it is mixed with water in a dough, it begins to ‘work’.

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\(^4\) Acids must, by definition, dissolve in water: otherwise they could not affect the \( \text{H}^+ \) concentration.
Implications for Pyrotechnics

1) Aluminium is attacked vigorously by alkalis, liberating hydrogen. Try putting a small piece of aluminium foil in a strong solution of sodium carbonate\(^5\) and see what happens. It might take a few minutes for the surface oxide layer to be broken down, but eventually you will see bubbles of hydrogen start to form as the aluminium begins to dissolve. It is a fairly slow, but definite reaction. Don’t try this with powdered aluminium as the large surface area means that the reaction will be much faster, and could be quite violent. It follows that you should never include a soluble carbonate in a mixture that also contains aluminium. It is, however, OK to mix insoluble carbonates, and even sodium bicarbonate, with aluminium – several of Winokur’s glitter formulae do so.

2) If you are going to add water to a composition containing aluminium and a carbonate, you might want to think carefully about any attempt to stabilise it with boric acid. Boric acid will, in the presence of water, react with barium or strontium carbonate and puff up your star with bubbles of carbon dioxide. However, the reaction is very slow at room temperature because boric acid is both a very weak acid and not very soluble.

3) If you are going to make your own basic copper carbonate from a soluble copper salt and either sodium or potassium carbonate, it would be better to start from copper chloride rather than the more readily available copper sulphate\(^6\). That way, the impurity in the carbonate will be basic copper chloride, whose main effect will be the beneficial one of slightly increasing the available chlorine – assuming that you are aiming for a blue flame.

What’s in a Name?

The first alkaline substance commonly available in Europe was made by mixing wood ashes with water and boiling the filtered liquid in iron pots. It was called ‘potash’ and was a somewhat impure potassium carbonate. In the Arab world, the same substance was known as ‘al kali’, roughly meaning ‘the ash’. That, broadly speaking, is the origin of the name of potassium and the word ‘alkali’.

So why do we use the letter K as the chemical symbol for potassium? The letter P wasn’t available as, inconveniently, phosphorus had already been discovered and the letter was already taken. As was commonly the case in such circumstances, people looked to Latin and found the word ‘kalium’ – which was, of course, derived from the Arabic ‘al kali’.

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\(^5\) If you don’t have any sodium carbonate, you can make it by heating some sodium bicarbonate in an oven at 200 degrees for half an hour or so.

\(^6\) It isn’t too difficult to convert copper sulphate to copper chloride using only commonly available chemicals.