

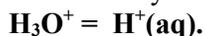
[A First Year blog on Acids and Bases: Part 1, 23rd March 2019](#)

(This is merely a *summary* of some of the things that you need to know. *Everything in this blog, and more, can be found in the First Year, Foundation Chemistry Book.*)

The Arrhenius definition of an acid

An acid is a substance that dissociates protons in Water, and in modern times¹ we now know that a proton cannot exist by itself in Water. It is such an aggressive species that when in Water it immediately piggy-backs onto a molecule of Water to form H_3O^+ : $\text{H}^+ + \text{H}_2\text{O} \longrightarrow \text{H}_3\text{O}^+$.

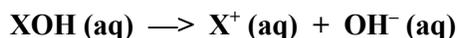
A proton cannot exist by itself in Water. There is literally a huge number of Water molecules even in a 10cc/10cm³ teaspoonful of Water (**334,000,000,000,000,000,000** or $3.342,796,149 \times 10^{23}$ molecules) – and each Oxygen atom in a molecule of Water has two unpaired electrons. A proton dissociated by an acid is therefore surrounded by a gigantic number of unpaired electrons and it just latches onto one of them (the nearest one) temporarily to form an Hydronium ion, H_3O^+ . In modern times we have stopped using the term H_3O^+ (although the Americans still use the term), and instead we now write H^+ (aq) viz. a proton attached to a molecule of Water and surrounded by many molecules of Water.



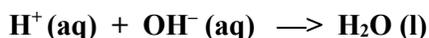
At GCSE level you learnt that an acid and a base gives you a salt plus Water, but in fact the ‘salt’ is just a ‘spectator’ ion and the real product of an acid-base reaction is **WATER**.



Under Arrhenius’ definition, a substance is not an acid until it dissociates its proton in Water, and a base is a substance that in Water dissociates its OH^- species (the hydroxide ion)



and it is the union of the proton and the hydroxide ion that forms a molecule of Water



NB “**H₂O (l)**” means a molecule of Water surrounded by other molecules of Water. Some bases cause the release of an OH^- species from a Water molecule – but wherever the OH^- species comes from, the effect will still be the same.

The Brønsted–Lowry definition of an acid

Under the Brønsted-Lowry definition, an acid is any substance that donates a proton, and a base is any substance that accepts a proton. We will then need to talk about conjugate acids and conjugate bases (and we shall do so shortly).

NB Water is **NOT** integral to the Brønsted- Lowry definition of an acid.

¹ Svante August Arrhenius won the Nobel prize for Chemistry in 1903, but at that time he did not know any of this stuff. It is thanks to his pioneering work that we now know so much about Acids.

The Lewis definition of an acid

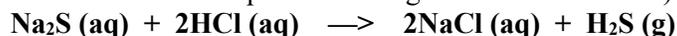
Under the Lewis definition, an acid is a substance that accepts a pair of electrons (it is an **electrophile**), and a base is a substance that donates a pair of electrons (it is a **nucleophile**). **The common laboratory acids are (alphabetically)**

<u>Name of acid</u>	<u>Formula</u>	<u>Resulting anion</u>	<u>Formula</u>
Carbonic acid	H ₂ CO ₃	Carbonate	(CO ₃) ²⁻
Hydrochloric acid	HCl	Chloride	(Cl) ⁻
Nitric acid	HNO ₃	Nitrate	(NO ₃) ⁻
Phosphoric acid	H ₃ PO ₄	Phosphate	(PO ₄) ³⁻
(the assumption above is that <i>three</i> protons have been dissociated)			
Sulphuric Acid	H ₂ SO ₄	Sulphate	(SO ₄) ²⁻
		Bisulphate or Hydrogen Sulphate =	(HSO ₄) ⁻

NB Please note the correlation between the charge on the anion and the number of H atoms in the acid. This is NOT a coincidence.

Using the Brønsted-Lowry definition of an Acid (instead of the Arrhenius model)

At GCSE Level, without knowing it we used the Arrhenius definition of an Acid-Base reaction i.e. the “neutralisation of an Acid and a Base to form a Salt plus Water” but you must now put that definition aside because it will no longer serve you well.² We are now going to learn about Acid-Base reactions in which sometimes no Water is formed! For example, under the Brønsted-Lowry definition of an Acid-Base reaction, an Acid-Base reaction occurs when HCl reacts with Na₂S – but no Water at all is formed in this reaction (instead we have a double displacement or gas evolution reaction)!



I should warn you that you are about to learn new ways of looking at some very familiar substances, and at first you may find it a bit difficult to look at familiar substances in a new light. For example, you have always been taught that Water is a neutral substance – and yet I am just about to tell you that (according to one way of defining Acids and Bases) Water can be both an Acid and/or a Base but clearly not both at the same time (except in Physics where Quantum theory allows all sorts of weird things to happen)!

There are different ways of defining Acids and Bases (e.g. using Arrhenius’ definition, or by using the Lewis definition) – but one of the ways of looking at Acid-Base reactions is the Brønsted-Lowry way whereby

An acid is a proton DONOR, and a Base is a proton ACCEPTOR.

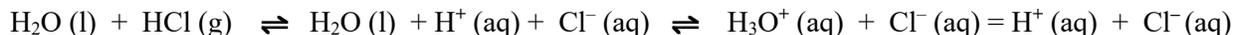
Any reaction where a proton (i.e. an H⁺ species) is donated to a substance that contains an unbonded (or a “lone”) pair of electrons that can receive the donated proton can be considered to be an acid-base reaction (and please remember that for a bond to form there has to be a pair of i.e. **TWO** bonding electrons).

If A⁻ is an anion such as Cl⁻ or SO₄²⁻ or NO₃⁻, then for any acid HA, HA must be able to break down into ions and thus dissociate a proton (or else it will **NOT** be able to act as an acid). Therefore, **HA** must be able to dissociate into H⁺ and A⁻.

Please note therefore that under both the Arrhenius and the Brønsted-Lowry definition, an acid does NOT behave as an acid until it actually dissociates its proton!

² That is not quite what Arrhenius said, but that is how it is taught at GCSE Level.

For an acid to *behave like an acid*, there has to be something to accept the proton. Therefore HCl gas (which in gaseous form is a molecular and not an ionic substance) **does NOT act like an acid when dissolved in Benzene** (because Benzene has no lone pair of electrons to form a dative bond with a proton and thus accept that proton, therefore *the protons are not dissociated in Benzene*) – but HCl **does behave like an acid when it is dissolved in Water**, because in a Water molecule the Oxygen atom does have at least one unbonded pair of electrons that can form a dative bond with the proton (to form H_3O^+) viz.



NB A proton can never exist in Water just as a proton. The charge density on a proton is so great that it immediately bonds to one of the two lone pairs of electrons in a Water molecule (*in any Water molecule*) and thus in reality H^+ never has an independent existence. H^+ in Water always exists as the Hydronium ion H_3O^+ .

However, in an acid-base reaction the Water molecule in the Hydronium ion does nothing. It plays no part in the reaction at all other than to carry the H^+ ion (the proton) to wherever it goes, and in modern times chemists therefore leave it out of acid-base reaction equations.

$\text{H}^+(\text{aq})$ is therefore the modern way of writing H_3O^+ i.e. protons piggy-backing/hitching a ride on a Water molecule.

NB The Oxygen atom in a Water molecule possesses two lone pairs of electrons, one of which accepts a proton, therefore **under the Brønsted-Lowry definition of a Base, Water here is a Base!**

OK, let me show you how (under our new definition) Water can act both as an acid and as a base (but clearly a Water molecule cannot act simultaneously as both a base and an acid).

Water acting as an ACID (Here Water has donated a proton therefore it has acted like an acid)

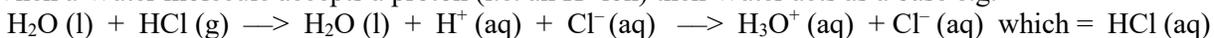
When a Water molecule donates a proton (i.e. an H^+ ion) to another species, then Water acts as an acid e.g.



The Nitrogen atom in an ammonia molecule has a lone (i.e. an unbonded) pair of electrons and this creates a dative bond with the H^+ proton to form a positively charged AMMONIUM ion, and this $\text{NH}_4^+(\text{aq})$ species attracts the negatively charged OH^- species and together they form Ammonium Hydroxide.

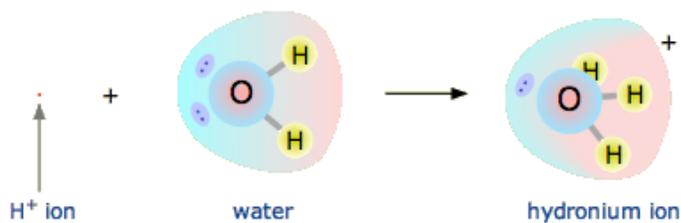
Water acting as a BASE (when it accepts a proton)

When a Water molecule accepts a proton (i.e. an H^+ ion) then Water acts as a base e.g.



NB The symbol " $\text{H}_3\text{O}^+(\text{aq})$ " can be read as "molecules of Water that have accepted protons (i.e. H^+ species) and the resulting positively charged H_3O^+ species are *surrounded by Water molecules*". Equally, " $\text{H}_3\text{O}^+(\text{aq})$ " and " $\text{Cl}^-(\text{aq})$ " can also be read as " H_3O^+ " and " Cl^- " species surrounded by Water molecules.

The Oxygen atom in a Water molecule has **two** unbonded (or lone) pairs of electrons on it, and it can use either one of them to accept an H^+ species (a proton) to become an H_3O^+ (Hydronium or Hydroxonium) species.³



[https://chem.libretexts.org/Textbook_Maps/Physical_and_Theoretical_Chemistry_Textbook_Maps/Supplemental_Modules_\(Physical_and_Theoretical_Chemistry\)/Acids_and_Bases/Acids_and_Bases_in_Aqueous_Solutions/The_Hydronium_Ion](https://chem.libretexts.org/Textbook_Maps/Physical_and_Theoretical_Chemistry_Textbook_Maps/Supplemental_Modules_(Physical_and_Theoretical_Chemistry)/Acids_and_Bases/Acids_and_Bases_in_Aqueous_Solutions/The_Hydronium_Ion) Prof Stephen Lower, Simon Fraser University.
NB The physical space occupied by a proton is TINY!

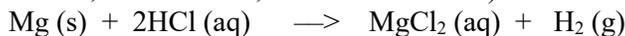
In fact, a proton does not stay bonded to any one molecule of Water, but instead it hops around from one Water molecule to another **many** times per second. It is this which allows Water to conduct electricity rapidly⁴.

Typical reactions of an Acid

A new definition for an Acid (the Brønsted-Lowry definition of an Acid, instead of the Arrhenius definition) does not mean that the reactions that we have observed all through the years of learning Chemistry are suddenly going to alter e.g.

Almost any Acid will still react with almost any Metal to produce Hydrogen

(There are only a very few metals, such as Gold, that will not do so.)



What happens is that $2HCl(aq) \longrightarrow 2H^+(aq) + 2Cl^-(aq)$, and then each proton or H^+ ion will attack the Magnesium metal (*which consists of positively charged Mg^{2+} ions set in a sea of delocalised electrons*) and steal an electron, and the proton will then become an atom/a Free Radical of Hydrogen – and these atoms/Free Radicals of Hydrogen join up to become molecules of Hydrogen (H_2), while each Mg^{2+} ion combines with $2Cl^-$ ions to form $MgCl_2$. *Would I be insulting you if I were to ask you to draw the dot-and-cross diagram for $MgCl_2$?* You will then see what is happening diagrammatically.

An acid will still react with a Carbonate to produce Water and Carbon Dioxide e.g.



and again, it is the dissociation of HCl into H^+ ions and Cl^- ions that allows the reaction to proceed (with Ca^{2+} bonding with $2Cl^-$ to become $CaCl_2$).

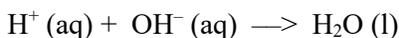
³ In fact, a proton does not stay bonded to any one molecule of Water, but instead it hops around from one Water molecule to another **many** times per second. It is this which allows Water to conduct electricity rapidly.

⁴ This feature was described by Theodor von Grotthuss in 1806 in his paper “Mémoire sur la décomposition de l’eau et des corps, qu’elle tient en dissolution a’ l’aide de l’électricité galvanique” and it subsequently became known as the ‘Grotthuss Mechanism’.

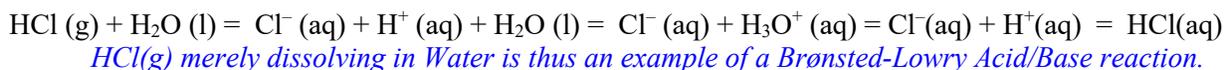
An acid and most Alkalis⁵ will still react together to form Water plus a Salt e.g.



and once again it is the dissociation of a proton from the acid that allows the reaction to proceed. We can show this by removing the spectator ions, whereupon we get our old friend

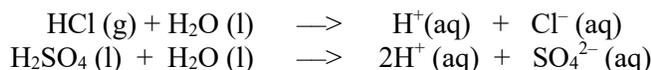


As a result of the new definition of an acid, we are now prepared to regard certain reactions as Acid-Base reactions which we would previously **not** have thought of as Acid-Base reactions. For example, when Hydrogen Chloride gas dissolves in Water to form Hydrochloric Acid, then as we have already seen, according to the Brønsted-Lowry theory or definition of an Acid-Base reaction, this **is** an Acid-Base reaction **because a molecule of Water has accepted a proton from HCl!** We can see this by writing out the reaction viz.



The difference between a Strong Acid and a Weak Acid

A **STRONG** acid⁶ (e.g. HCl or H₂SO₄) is one that dissociates (usually in Water) almost completely into its cations (H⁺) and its anions Cl⁻ (in the case of Hydrochloric Acid), or SO₄²⁻ (in the case of Sulphuric Acid).



and these resulting ions show virtually no inclination to recombine, and in any case cannot recombine because they are surrounded by and separated by Water molecules.⁷ The position of dynamic equilibrium lies *so far* to the Right as to be considered to have dissociated to “Completion”.

In contrast, a **weak acid is one where very little dissociation occurs**. Thus e.g. in an aquated solution of Ethanoic Acid



the equilibrium lies so far to the LEFT that there are only a few H⁺ (aq) species in the solution.⁸

⁵ I do not want to go into the complexities of HF.

⁶ A strong acid has a pH value of about 0-3, while a weak acid has a pH value of about 3-6.

⁷ Technically $\text{HCl (aq)} \longrightarrow \text{H}^+ \text{ (aq)} + \text{Cl}^- \text{ (aq)}$ should be written as $\text{HCl (aq)} \rightleftharpoons \text{H}^+ \text{ (aq)} + \text{Cl}^- \text{ (aq)}$, but since the tendency to reverse the reaction is so small, it is acceptable to write the reaction as $\text{HCl (aq)} \longrightarrow \text{H}^+ \text{ (aq)} + \text{Cl}^- \text{ (aq)}$. Strong acids have a pH value of about 0-3 (and can sometimes be less than zero). In the case of $\text{HCl (aq)} \longrightarrow \text{H}^+ \text{ (aq)} + \text{Cl}^- \text{ (aq)}$, even though the equilibrium lies so far to the right that there are virtually no HCl (aq) species at all in an aquated solution of Hydrochloric acid, nevertheless there may still be *some* HCl species in the solution therefore $\text{HCl (aq)} \rightleftharpoons \text{H}^+ \text{ (aq)} + \text{Cl}^- \text{ (aq)}$ is technically a more accurate statement than $\text{HCl (aq)} \longrightarrow \text{H}^+ \text{ (aq)} + \text{Cl}^- \text{ (aq)}$.

⁸ Please could you remember that **most Carboxylic acids are weak acids** i.e. equilibrium is reached when they dissociate only a very small number of their protons – but, clearly, **if there is something to mop up those protons as they are dissociated, then even though a Carboxylic acid is a weak acid, then all the time that there is something mopping up its protons, it will keep on dissociating protons until all its protons have been dissociated!**

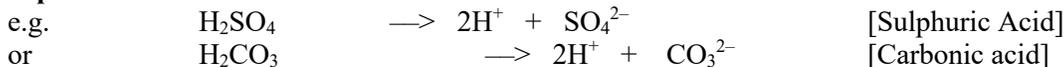
Please do not confuse the words **strong** and **weak** with the words **concentrated** and **diluted**! A **dilute** acid can be either a strong acid OR a weak acid **that has been diluted by Water** (*say 1 part of acid to 10 parts of Water or whatever*), whereas a **concentrated** acid has very little dilution at all.

An acid may dissociate some or all of its protons. If it dissociates one proton it is called monobasic or monoprotic, if it dissociates two protons it is called dibasic or diprotic, if it dissociates three protons it is called tribasic or triprotic,, and so on. (However, I prefer the term “.... protic” rather than “.....basic” because “protic” tells you how many protons are being dissociated from each molecule of the acid.) We thus get

Monoprotic/monobasic acids



Diprotic/dibasic acids



Triprotic/tribasic acids



The pH of pure Water

The RMM or M_r of Water is [(2 x 1) + (1 x 16.0) =] 18 gmol⁻¹, therefore
 18g of Water make up 1 mole of Water (cf. footnote⁹), therefore
 1000g of Water or 1dm³ of Water has (1 mole x 1000g ÷ 18g =) 55.55555 moles of H₂O molecules in it.
 (55.5 recurring is written “55.5”.)

There are only a very few H⁺ ions in Water, and by experimentation, scientists have established that at a temperature of 298K, 1 dm³ of pure Water contains 1 x 10⁻⁷ moles of H⁺ ions, and (since the stoichiometric ratio of H₂O (l) ⇌ H⁺ (aq) + OH⁻ (aq) is 1 : 1 : 1) then there are therefore also 1 x 10⁻⁷ moles of OH⁻ ions in 1 dm³ of Water at 298K.

As a percentage, [(1 x 10⁻⁷) ÷ 55.6 x 100] % = 0.000,000,18% of **Water** is thus ionised into H⁺ and OH⁻ ions – therefore a truly **TINY** amount of Water is ever ionised. We define acidic substances as those having a pH value of less than 7 at 298K, and those with pH values of more than 7 at 298K are called bases, and “neutral” Water has a pH value of 7 at 298K (**and only at 298K**).

Pure Water is defined as “neutral”, even though it has both 1 x 10⁻⁷ moles of H⁺ ions and 1 x 10⁻⁷ moles of OH⁻ ions in it at 298K, and it is “neutral” because the number of H⁺ ions is exactly equal to the number of OH⁻ ions, and these ions cancel out the effect of each other. **At a different temperature there would be a different number of H⁺ and OH⁻ ions in one mole of Water, and Water would thus have a different pH value at a different temperature – but it would still be neutral.**

⁹ RMM of H₂O ≈ (2 x 1) + 16 = 18 gmol⁻¹.

The table below (from Jim Clark writing for Chemistry LibreTexts) shows the effect of temperature on K_w .

T (°C)	K_w (mol ² dm ⁻⁶)	pH	pOH
0	0.114 x 10 ⁻¹⁴	7.47	7.47
10	0.293 x 10 ⁻¹⁴	7.27	7.27
20	0.681 x 10 ⁻¹⁴	7.08	7.08
25	1.008 x 10 ⁻¹⁴	7.00	7.00
30	1.471 x 10 ⁻¹⁴	6.92	6.92
40	2.916 x 10 ⁻¹⁴	6.77	6.77
50	5.476 x 10 ⁻¹⁴	6.63	6.63
100	51.3 x 10 ⁻¹⁴	6.14	6.14

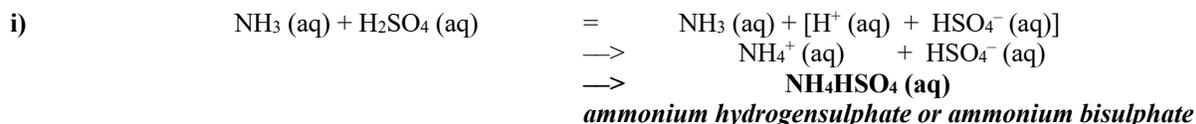
You can see that the pH of pure Water **decreases** as the temperature increases, i.e. **as the temperature RISES** the number of protons that are dissociated increases – but the Water is still NEUTRAL at all these temperatures. In this instance, a lower pH value does **not** mean that the Water has become more acidic. The number of protons dissociated is determined mathematically by K_w , and even though the number of dissociated protons increases as the temperature rises, the Water remains neutral throughout the changes in temperature.

NB Please do not worry about the meaning of K_w for the moment. I will write about K_w in a different blog. At this stage, all that I want you to do is to notice that *mathematically* the pH value of Water is temperature dependent, but that there is no alteration in the actual **acidity** of the Water. The Water remains neutral throughout.

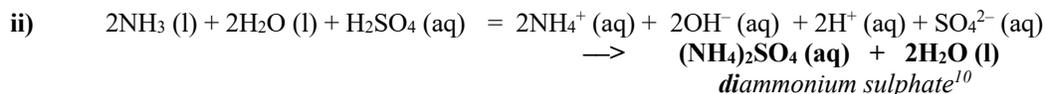
According to the Brønsted-Lowry definition, Ammonia is a Base

I am sure that you will remember that the Nitrogen atom in Ammonia (NH₃) has a lone pair of electrons on it, therefore it can form a dative bond with the proton of an Acid *and thus accept a proton from the Acid*. Therefore, according to the Brønsted-Lowry definition of a Base, Ammonia is a Base and the reaction of Ammonia with an Acid is an Acid-Base reaction.

For example, if Ammonia were reacted with Sulphuric Acid, then (depending on how much Water there was in the Sulphuric Acid) we would get either



or



and the *diammonium sulphate* (which is usually referred to as “Ammonium Sulphate”) that is formed plays a very important part in the production of fertilisers.

¹⁰ NH₄⁺ has a charge of “+1”, whereas SO₄²⁻ has a charge of “-2”, therefore Ammonium Sulphate HAS TO BE **Diammonium** Sulphate.