

## Acids

(A First Year Blog.)

I cannot stress strongly enough how **DANGEROUS** Acids and Alkalis are. PLEASE treat them with the utmost caution. You are at the beginning of your lives. If you are blinded and have to spend the rest of your life sightless, I promise you, you will **NOT** like it one little bit.

In some textbooks you will see the statement “acids taste sour”. That refers to things like vinegar in your kitchen at home. Please do **NOT** put ANYTHING that you may find in a Chemistry laboratory into your mouth. That would be sheer utter madness!

Acids, Alkalis and all Laboratory chemicals are

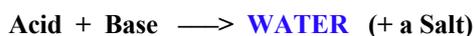
**MASSIVELY DANGEROUS ENTITIES.**

Please treat them with the utmost caution. Always wear your safety goggles/gloves/and lab coat during your experiments.

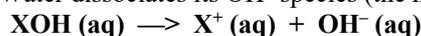
**IT IS BETTER TO BE ULTRA-SAFE  
THAN TO BE SORRY  
FOR THE REST OF YOUR LIFE.**

### A) The Arrhenius definition of an acid

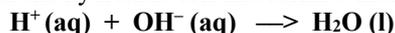
- An acid is a substance that dissociates protons in Water, and in modern times<sup>1</sup> we now know that a proton cannot exist by itself in Water. It is such an aggressive species that when in Water it immediately piggy-backs onto a molecule of Water to form H<sub>3</sub>O<sup>+</sup>:  $H^+ + H_2O \longrightarrow H_3O^+$
- **A proton cannot exist by itself in Water.** Depending on the volume of the Water, there is literally a huge number of Water molecules even in a 10cc/10cm<sup>3</sup> teaspoonful of Water (cf. page 6) – 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 temporarily to form an Hydronium ion, H<sub>3</sub>O<sup>+</sup>. In modern times we have stopped using the term H<sub>3</sub>O<sup>+</sup>, and instead we write H<sup>+</sup> (aq).  
 $H_3O^+ = H^+(aq).$
- At GCSE level you learnt that an acid and a base gives you a salt plus Water, but elsewhere I have you about ‘spectator ions’ and we then learnt that the ‘salt’ was just a ‘spectator’ and that 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, therefore a base must be a substance that in Water dissociates its OH<sup>-</sup> species (the hydroxide ion)



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



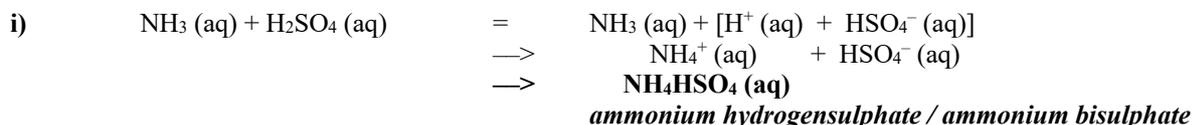
NB “H<sub>2</sub>O (l)” means a molecule of Water surrounded by other molecules of Water.

### B) The Brønsted–Lowry definition of an acid

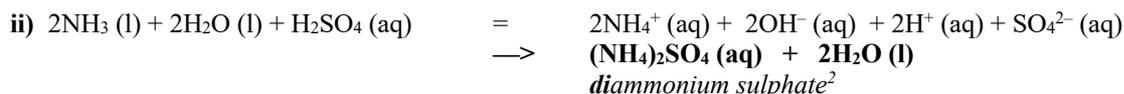
- Under the Brønsted-Lowry definition, an acid is any substance that donates a proton, and a base is any substance that accepts a proton. We will then need to talk about conjugate acids and conjugate bases (and we shall do so shortly). NB Water is **NOT** integral to the Brønsted–Lowry definition of an acid.

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

- I am sure that you will remember that the Nitrogen atom in Ammonia (NH<sub>3</sub>) 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.

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

<sup>2</sup> NH<sub>4</sub><sup>+</sup> has a charge of “+1”, whereas SO<sub>4</sub><sup>2-</sup> has a charge of “-2”, therefore Ammonium Sulphate HAS TO BE **Diammonium Sulphate**.

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

- At GCSE Level, without knowing it you 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.<sup>3</sup> 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 Na<sub>2</sub>S – but no Water at all is formed in this 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!
- 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<sup>+</sup> 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<sup>-</sup> is an anion such as Cl<sup>-</sup> or SO<sub>4</sub><sup>2-</sup> or NO<sub>3</sub><sup>-</sup>, 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<sup>+</sup>** and **A<sup>-</sup>**.

**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!**

- 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<sub>3</sub>O<sup>+</sup>) viz.  
$$\text{H}_2\text{O}(\text{l}) + \text{HCl}(\text{g}) \rightleftharpoons \text{H}_2\text{O}(\text{l}) + \text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{Cl}^-(\text{aq}) = \text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq})$$
- **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<sup>+</sup> never has an independent existence. **H<sup>+</sup> in Water always exists as the Hydronium ion H<sub>3</sub>O<sup>+</sup>.**
- 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<sup>+</sup> ion (the proton) to wherever it goes, and in modern times chemists have therefore “demoted” its importance in acid-base reaction equations.
- **H<sup>+</sup>(aq) is therefore the modern way of writing H<sub>3</sub>O<sup>+</sup> 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).

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<sup>3</sup> That is not quite what Arrhenius said, but that is how it is taught at GCSE Level. We will talk about Arrhenius in greater detail next year in the Chapter on Acids.

### C) The Lewis definition of an acid

- 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 <sub>2</sub> CO <sub>3</sub>	Carbonate	(CO <sub>3</sub> ) <sup>2-</sup>
Hydrochloric acid	HCl	Chloride	(Cl) <sup>-</sup>
Nitric acid	HNO <sub>3</sub>	Nitrate	(NO <sub>3</sub> ) <sup>-</sup>
Phosphoric acid	H <sub>3</sub> PO <sub>4</sub>	Phosphate	(PO <sub>4</sub> ) <sup>3-</sup>
(the assumption above is that <i>three</i> protons have been dissociated)			
Sulphuric Acid	H <sub>2</sub> SO <sub>4</sub>	Sulphate	(SO <sub>4</sub> ) <sup>2-</sup>

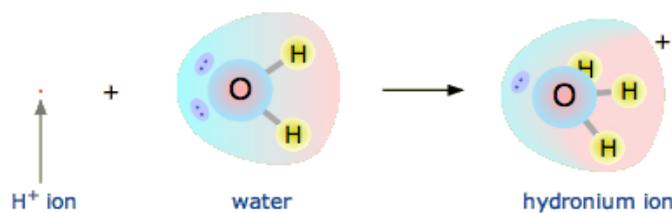
**NB Please note the correlation between the charge on the anion and the number of H atoms in the acid. This is NOT a coincidence.**

#### 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<sup>+</sup> ion) to another species, then Water acts as an acid e.g.  
$$\text{NH}_3(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_3(\text{g}) + \text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightleftharpoons \text{NH}_4^+\text{OH}^-(\text{aq})$$
- 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<sup>+</sup> proton to form a positively charged AMMONIUM ion, and this NH<sub>4</sub><sup>+</sup> (aq) species attracts the negatively charged OH<sup>-</sup> 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<sup>+</sup> ion) then Water acts as a base e.g.  
$$\text{H}_2\text{O}(\text{l}) + \text{HCl}(\text{g}) \longrightarrow \text{H}_2\text{O}(\text{l}) + \text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \longrightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{Cl}^-(\text{aq})$$
 which = HCl(aq)  
NB The symbol "H<sub>3</sub>O<sup>+</sup> (aq)" can be read as "a molecule of Water that has accepted a proton (i.e. an H<sup>+</sup> species) and the resulting positively charged H<sub>3</sub>O<sup>+</sup> species is *surrounded by Water molecules*". Equally, "H<sub>3</sub>O<sup>+</sup> (aq)" and "Cl<sup>-</sup> (aq)" can also be read as "H<sub>3</sub>O<sup>+</sup>" and "Cl<sup>-</sup>" 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<sup>+</sup> proton to become an H<sub>3</sub>O<sup>+</sup> species.<sup>4</sup>



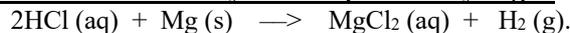
The diagram is from Prof Stephen Lower, Simon Fraser University. NB The physical space occupied by a proton is TINY and that is what Prof Lower was trying to indicate by his tiny red dot.  
[https://chem.libretexts.org/Textbook\\_Maps/Physical\\_and\\_Theoretical\\_Chemistry\\_Textbook\\_Maps/Supplemental\\_Modules\\_\(Physical\\_and\\_Theoretical\\_Chemistry\)/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)

<sup>4</sup> 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. This feature was described by Theodor von Grothuss 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 'Grothuss Mechanism'.

## Typical reactions of an Acid

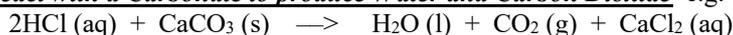
- Merely because we have changed the definition of (or learnt a new definition for) an Acid, it does not mean that the reactions that we have observed all through the years of learning Chemistry are suddenly going to alter e.g.

- Almost any Acid will still react with almost any Metal to produce Hydrogen e.g.



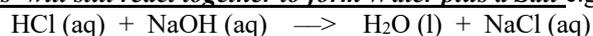
What happens is that  $2\text{HCl (aq)} \longrightarrow 2\text{H}^+ \text{ (aq)} + 2\text{Cl}^- \text{ (aq)}$ , and then each proton or  $\text{H}^+$  ion will attack the Magnesium metal (*which consists of positively charged  $\text{Mg}^{2+}$  ions set in a sea of delocalised electrons*) and steal an electron thus becoming an atom/a Free Radical of Hydrogen – and these atoms/Free Radicals of Hydrogen join up to become molecules of Hydrogen ( $\text{H}_2$ ), while each  $\text{Mg}^{2+}$  ion combines with  $2\text{Cl}^-$  ions to form  $\text{MgCl}_2$ . *Would I be insulting you if I were to ask you to draw the dot-and-cross diagram for  $\text{MgCl}_2$ ?*

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

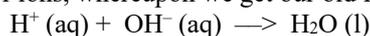


and again, it is the dissociation of  $\text{HCl}$  into  $\text{H}^+$  ions and  $\text{Cl}^-$  ions that allows the reaction to proceed (with  $\text{Ca}^{2+}$  bonding with  $2\text{Cl}^-$  to become  $\text{CaCl}_2$ ).

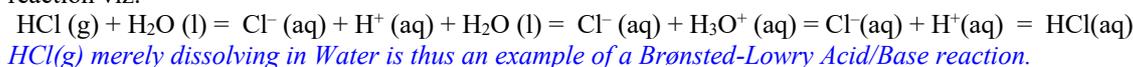
- An acid and most Alkalis<sup>5</sup> will still react together to form Water plus a Salt e.g.



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  $\text{HCl}$ !** We can see this by writing out the reaction viz.



## The difference between a Strong Acid and a Weak Acid

- A **STRONG** acid<sup>6</sup> (e.g.  $\text{HCl}$  or  $\text{H}_2\text{SO}_4$ ) is one that dissociates (usually in Water) almost completely into its cations ( $\text{H}^+$ ) and its anions  $\text{Cl}^-$  (in the case of Hydrochloric Acid), or  $\text{SO}_4^{2-}$  (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.<sup>7</sup> The position of dynamic equilibrium lies *very far* to the Right.

- In contrast, a **weak acid** is one where **very little dissociation occurs**. Thus e.g. in an aquated solution of Ethanoic Acid  $\text{CH}_3\text{COOH (aq)} \rightleftharpoons \text{CH}_3\text{COO}^- \text{ (aq)} + \text{H}^+ \text{ (aq)}$  the equilibrium lies so far to the LEFT that there are only a few  $\text{H}^+ \text{ (aq)}$  species in the solution.<sup>8</sup>

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<sup>5</sup> I do not want to go into the complexities of HF.

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

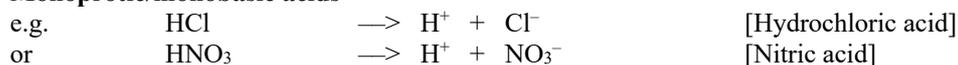
<sup>7</sup> Technically  $\text{HCl (aq)} \longrightarrow \text{H}^+ \text{ (aq)} + \text{Cl}^- \text{ (aq)}$  should be written as  $\text{HCl (aq)} \rightleftharpoons \text{H}^+ \text{ (aq)} + \text{Cl}^- \text{ (aq)}$ , but since the tendency to reverse the reaction is so small, it is acceptable 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 sometimes can be less than zero). In  $\text{HCl (aq)} \longrightarrow \text{H}^+ \text{ (aq)} + \text{Cl}^- \text{ (aq)}$ , even though the equilibrium lies so far to the right that there are virtually no  $\text{HCl (aq)}$  species at all in an aquated solution of Hydrochloric acid, nevertheless there may still be *some*  $\text{HCl (aq)}$  species in the solution therefore

$\text{HCl (aq)} \rightleftharpoons \text{H}^+ \text{ (aq)} + \text{Cl}^- \text{ (aq)}$  is technically a more accurate statement than  $\text{HCl (aq)} \longrightarrow \text{H}^+ \text{ (aq)} + \text{Cl}^- \text{ (aq)}$ .

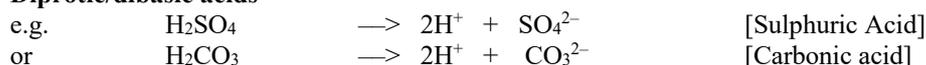
<sup>8</sup> Please could you remember that **most Carboxylic acids are weak acids** i.e. equilibrium is reached when they dissociate only a very small number of their protons – but, clearly, **if there is something to mop up those protons as they are dissociated, then even though a Carboxylic acid is 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!** That is the theoretical basis of buffer solutions.

- 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** (e.g. *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”.) We thus get

#### Monoprotic/monobasic acids



#### Diprotic/dibasic acids



#### Triprotic/tribasic acids



### The pH of pure Water

- The RMM or M<sub>r</sub> of Water is [(2 x 1) + (1 x 16.0) = ] 18 gmol<sup>-1</sup>, therefore  
18g of Water make up 1 mole of Water (cf. footnote<sup>9</sup>), therefore  
1000g of Water or 1dm<sup>3</sup> of Water has (1 mole x 1000g ÷ 18g =) 55.55555 moles of H<sub>2</sub>O molecules in it.
- By experimentation, scientists have established that at a temperature of 298K, 1 dm<sup>3</sup> of pure Water contains 1 x 10<sup>-7</sup> moles of H<sup>+</sup> ions and (since the stoichiometric ratio of H<sub>2</sub>O (l) ⇌ H<sup>+</sup> (aq) + OH<sup>-</sup> (aq) is 1 : 1 : 1) there are therefore also 1 x 10<sup>-7</sup> moles of OH<sup>-</sup> ions in 1 dm<sup>3</sup> of Water at 298K.
- [(1 x 10<sup>-7</sup>) ÷ 55.6] x 100 % = 0.000,000,18% of **Water** is thus ionised into H<sup>+</sup> and OH<sup>-</sup> 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<sup>-7</sup> moles of H<sup>+</sup> ions and 1 x 10<sup>-7</sup> moles of OH<sup>-</sup> ions in it at 298K, and it is “neutral” because the number of H<sup>+</sup> ions is exactly equal to the number of OH<sup>-</sup> ions, and these ions cancel out the effect of each other. **At a different temperature there would be a different number of H<sup>+</sup> and OH<sup>-</sup> 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.**

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<sup>9</sup> RMM of H<sub>2</sub>O ≈ (2 x 1) + 16 = 18 gmol<sup>-1</sup>.

The table below shows the effect of temperature on  $K_w$ . NB As the temperature rises, it just looks as though Water is becoming more acidic (the pH value is diminishing) – but, in reality, Water remains neutral at ANY temperature.

T (°C)	$K_w$ (mol <sup>2</sup> dm <sup>-6</sup> )	pH	pOH
0	$0.114 \times 10^{-14}$	7.47	7.47
10	$0.293 \times 10^{-14}$	7.27	7.27
20	$0.681 \times 10^{-14}$	7.08	7.08
25	$1.008 \times 10^{-14}$	7.00	7.00
30	$1.471 \times 10^{-14}$	6.92	6.92
40	$2.916 \times 10^{-14}$	6.77	6.77
50	$5.476 \times 10^{-14}$	6.63	6.63
100	$51.3 \times 10^{-14}$	6.14	6.14

Source: Chemistry LibreTexts

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 that are 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. 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.

- There is one more thing that I want/I need to tell you about strong acids viz. on page 6 I told you that pure Water dissociates a tiny proportion of its water molecules into  $H^+$  ions and  $OH^-$  ions<sup>10</sup>, and from an inspection of the Periodic Table we can calculate that
 

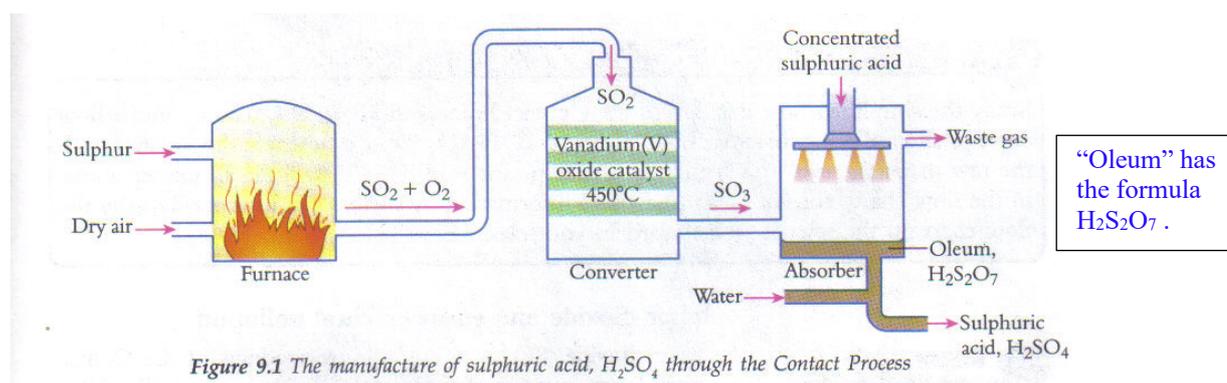
1 mole of Water	has a mass of	18.0g, therefore	
X moles	.....	1000.0g	(cf. footnote <sup>11</sup> )
where $X = (1 \text{ mol} \times 1000.0 \text{ g}) \div 18.0 \text{ g} = 55.5555\dots$ recurring = 55.5 moles.			
- The Acid Dissociation Constant,  $K_a = [H_2O]$  which from the above =  $55.5 \text{ mol dm}^{-3}$  therefore  $pK_a = -\log 55.5 = -1.745$ . We will need this next year when we examine Acids in greater depth.

<sup>10</sup> ... and in this sense Water can thus be thought of as an acid.

<sup>11</sup> The RMM of  $H_2O \approx (2 \times 1.00) + 16.0 = 18.0 \text{ g mol}^{-1}$ , but if we use Carbon-12 as the yardstick for the measurement of Mass, then the more accurate **isotopic** RAM of Hydrogen is 1.00794 atomic mass units, and the RAM for Oxygen would be 15.9994 amu and the RMM for Water would be 18.01528 amu (but I am not interested in such refinements for the moment), and I am happy to use here  $18.0 \text{ g mol}^{-1}$  for the RMM of Water instead of  $18.01528 \text{ g mol}^{-1}$ .

## The Contact Process for producing Sulphuric Acid

- Some British Exam Boards require a knowledge of the Contact Process for the manufacture of Sulphuric Acid,  $\text{H}_2\text{SO}_4$  – so let me say a few words on that subject.
  - Sulphuric Acid is a strong diprotic/ dibasic acid which is often used as an **oxidising agent**.
  - It is also a powerful **dehydrating agent** and is used to remove a molecule of water from many organic compounds.
  - It can be a **catalyst**, and it can also be
  - a **protonating agent**.
- Industrially, Sulphuric Acid is made by the Contact Process (below)

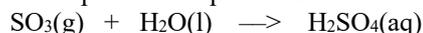


where Sulphur is roasted in a furnace to produce Sulphur Dioxide,  $\text{SO}_2(\text{g})$ , and then  $\text{SO}_2(\text{g})$  is oxidised into  $\text{SO}_3(\text{g})$  using  $\text{V}_2\text{O}_5$  as the catalyst

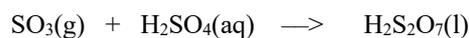


The process is **EXHO**thermic. It gives off heat.

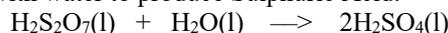
and then the  $\text{SO}_3$  is reacted with water to produce Sulphuric Acid.



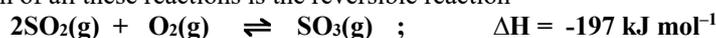
- Unfortunately, the reaction between  $\text{SO}_3(\text{g})$  and water is a very exothermic reaction and the reactants and products immediately boil and in consequence hardly any Sulphuric acid is produced.
- The resolution of this problem is to react/dissolve the  $\text{SO}_3(\text{g})$  not with water but with concentrated Sulphuric Acid to produce something called "Oleum",  $\text{H}_2\text{S}_2\text{O}_7(\text{l})$



and then the Oleum is reacted with water to produce Sulphuric Acid.



- The critical reaction of all these reactions is the reversible reaction



There are 3 moles of gases on the left hand side of the equation and only one on the right hand side therefore **a high pressure would be used to increase the yield.**

- According to Le Chatelier's Principle, a reversible reaction that is in dynamic equilibrium system will seek to oppose any changes that are made to it, therefore if the temperature of an exothermic reaction is reduced, then the system will seek to raise the temperature by producing more product. A low temperature should thus be used to drive the position of dynamic equilibrium to the right and produce more  $\text{SO}_3(\text{g})$ . However, a low temperature **reduces the rate of the reaction** and thus it is *uneconomic* to use a low temperature because even though a large amount of the  $\text{SO}_2(\text{g})$  might be converted into  $\text{SO}_3(\text{g})$ , the rate of this conversion makes the process uneconomic. To convert 100% of  $\text{SO}_2(\text{g})$  into  $\text{SO}_3(\text{g})$  might be good Chemistry but it is useless **Economics** if the process were to produce just 1 tonne of  $\text{SO}_3(\text{g})$  per annum! Depending on the cost of the plant, an industrial plant might need to produce 250,000 (or whatever) tonnes of Sulphuric Acid per annum to be a viable one.

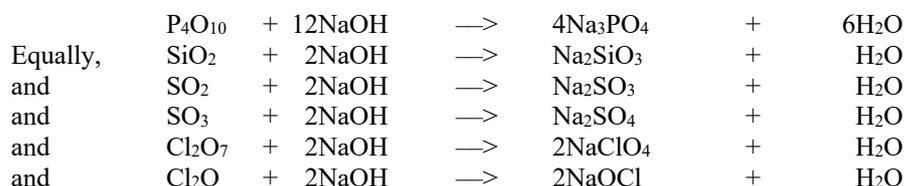
- **The temperature used in the Contact Process is thus 450°C** (because this gives a conversion factor of 97% whereas at 550°C the conversion rate drops off sharply to 86%).
- The pressure used is just over atmospheric pressure because even at this pressure the conversion rate is so high at 450°C that a high pressure is not needed. **(High pressures cost more money and reduce profitability and can be EXTREMELY dangerous.)**

*NB It is interesting to note that the crude oil from many countries contains Sulphur (about 2-3% on average), and this Sulphur is removed before the petrol is sold on forecourts to the motorist. The Sulphur that is removed from crude oil is normally made into Sulphuric Acid.*

- The industrial producers of Sulphuric acid say<sup>12</sup> that it is “a colourless, viscous liquid and it is hazardous and corrosive in nature. Sulphur is the major raw material used in the production of sulphuric acid, and around 90 percent of global sulphur production is used in the production of sulphuric acid. The other methods of sulphuric acid production include production from smelter gas and iron pyrites. Sulphuric acid finds its applications in a myriad of sectors and industries including Fertilizers, Chemicals, Agricultural, and others. Major applications of sulphuric acid include the production of chemicals, as an additive in the Pulp and Paper industry, in rubber vulcanising applications, and as battery acid in the Automotive sector”.
- Depending on how well the world is doing economically, something like 200-250 million tonnes of Sulphuric Acid are produced each year worldwide, and some 60% of that output goes into the making of fertilisers in some form or another (either as Ammonium Sulphate itself or in some other form).

### Postscript

- It will be of considerable value to you to know that **the Oxides of non-metals are acidic and that the oxides of most metals are basic** (i.e. they will react with water to produce OH<sup>-</sup> ions, and react with acids). Many bases will react with and dissolve in water, and those that do are called **alkalis**.
- **The oxides of non-metals are ACIDIC and the oxides of metals are on the whole BASIC.**
- The reactions of metal oxides with water tends to be fairly well covered at GCSE Level, but the reactions of **non-metal** oxides are less well-covered. For example, Phosphorous Pentoxide has the empirical formula P<sub>2</sub>O<sub>5</sub> and the molecular formula P<sub>4</sub>O<sub>10</sub>, and when it reacts with water it forms H<sub>3</sub>PO<sub>4</sub> and when it reacts with NaOH it forms Na<sub>3</sub>PO<sub>4</sub> and H<sub>2</sub>O viz.



OK, that is the end of this note.

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<sup>12</sup> The English in the Industrial handout was atrocious so I have corrected it.