

# **ACIDITY AND BASICITY IN CHEMISTRY**

SAMPLE

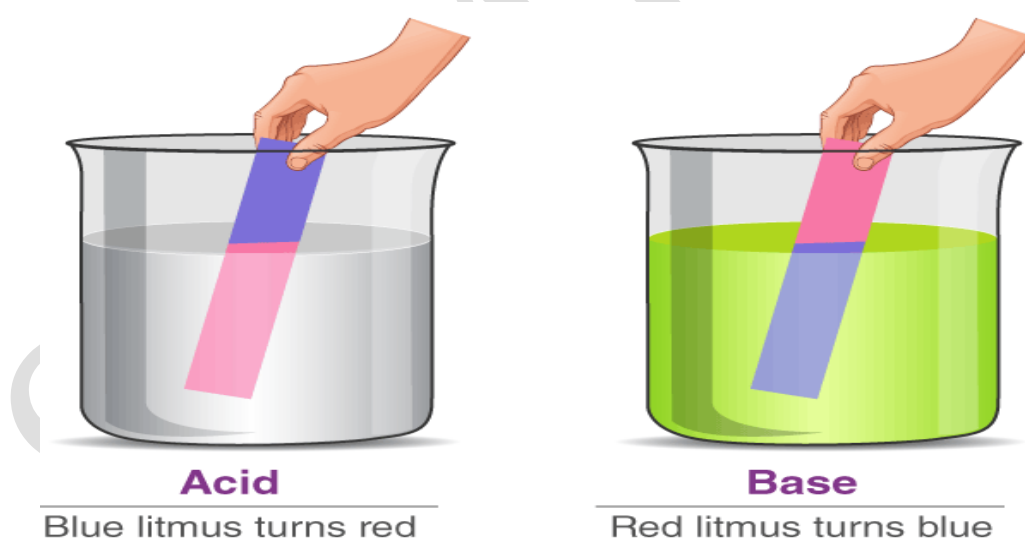
# CHAPTER 1: BASICS OF ACID-BASE CHEMISTRY

## 1.1. Introduction

Acid, base and salt are old notions that advanced chemistry has absorbed and developed. Bases and acids take part in a critical part in the lives of the humans and in the environment, through our body to the seas and minerals of the Earth. We have had bases & acids if we have ever sipped the juice of a lemon or used soap to wash our hands. Based on qualities like flavour and pH, researchers divide substances into bases (also known as alkali), acids, or neutral. Citric acid is a substance found in citrus fruits. Toothpaste is an ineffective starting point. Toothpaste preserves dental enamel by neutralizing the acids in meals (Ayers et al., 2015).

The bitter taste of acidic substances is typically a good indicator. An acid is a molecule that may give an  $H^+$  ion as residual thermodynamically favourable after releasing  $H^+$ . Blue litmus can become red when uncovered to acids.

Bases, contrarily, have a sour flavour and a slick consistency. The word alkali represents a base that may be soluble in water. Salts are produced if such compounds reacted chemically with acids. Bases are reported to change red litmus to blue (Shindy, 2012).



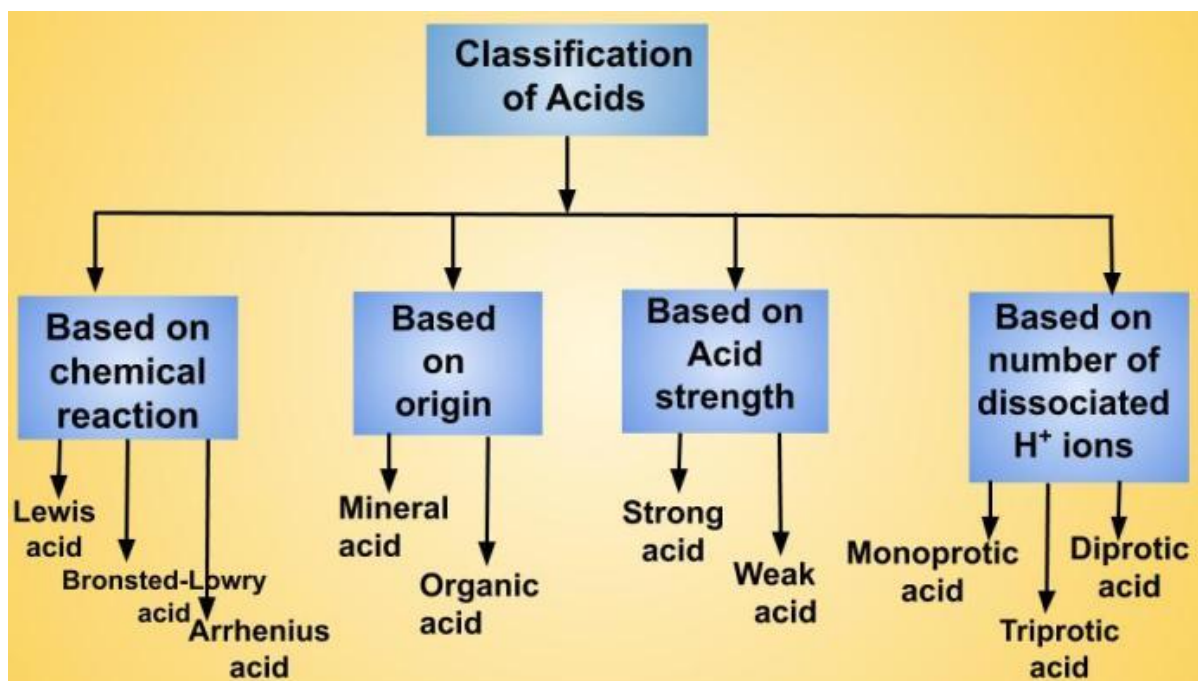
*Figure 1.1. Litmus Test of Bases & Acids*

*Source: <https://byjus.com/chemistry/acids-and-bases/>*

## 1.2. Acids

The word acid was first utilized in the 17th century, and this word was derived through the Latin word ac-, which means "pointed," like the taste of vinegar. Acids are acknowledged as a unique class of chemicals with the following characteristics in their aqueous solutions (Chen & Neibling, 2014):

- i. The ability to convert litmus from blue to red;
- ii. A distinct bitter flavour;
- iii. React with *bases* to produce water & sodium chloride;
- iv. React with some metals to form hydrogen gas.



*Figure 1.2. Categorization of acids*

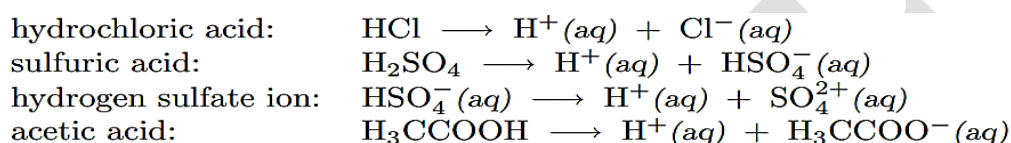
Source: <https://dewwool.com/10-types-of-acids/>

The initial chemical concept of acid was incorrect: in 1787, Antoine Lavoisier classified recognized acids like a distinct group of "complex substances" as a portion of his comprehensive taxonomy of compounds. He hypothesized that one's unique characteristics stemmed from the existence of a basic component that embodied the "acidity" concept, which he titled oxygen, after the Greek word for "acid former." Lavoisier was given this term to a novel gas element identified only a few years before by Joseph Priestly as the key ingredient which enables burning. Because several products of combustion (oxides) produce acidic solutions, and also most acids contain oxygen, Lavoisier's error is reasonable (McQuilton et al., 2012).

Humphrey Davy demonstrated in 1811 this muriatic (hydrochloric) acid (which Lavoisier was considered an element) doesn't really comprise oxygen and that only reassured a few that chlorine had been oxygen comprising compound rather than an element. However a few oxygen-free acids were identified by 1830, the hydrogen hypothesis of acids had not been broadly acknowledged till about 1840. Through this point, the misnomer oxygen had become too well-known to be altered (Brooks et al., 2005).

The importance of knowing acids (and also salts & bases) had to wait until Michael Faraday discovered the solutions of salt (called electrolytes) produce electricity in the mid-19th century. It suggests the presence of charged particles capable of migrating in response to an electric field. Such particles were given the name ions ("wanderers") by Faraday. The qualities we associated with acids are attributable to an overabundance of hydrogen ions in the solution, according to later investigations on electrolyte solution (Johnson, 2003).

Svante Arrhenius (1859 to 1927), a Swedish chemist, developed the 1st beneficial theory of acids in 1890: an acidic substance has at least 1 hydrogen atom which may dissociate, or ionize, once mixed with water, generating an anion & a hydrated hydrogen ion:



To be considered an "Arrhenius acid," it should include hydrogen. Many compounds, on the other hand, don't really comprise hydrogen natively, however when mixed with water, they nonetheless produce hydrogen ions; the hydrogen ions are produced by the water on their own, as a result of the interaction with the material. To provide a very relevant operational concept of an acid, the preceding is provided (Titov, 2016):

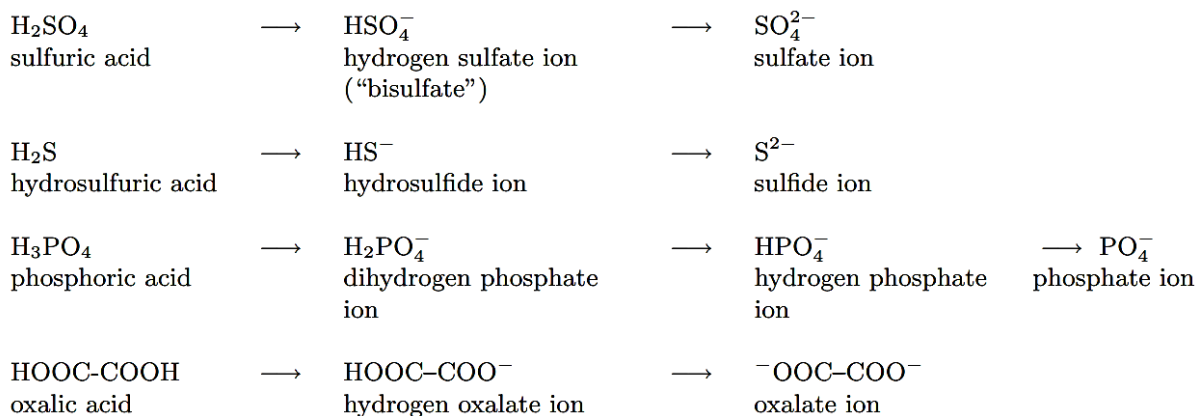
*When a chemical is mixed with water, it produces an oversupply of hydrogen ions, which is known as an acid.*

When it comes to hydrogen in acids, there will be 3 critical aspects to remember:

Although all Arrhenius acids comprise hydrogen atoms, not over all the atoms of hydrogen in a substance are vulnerable to detachment, hence the  $-\text{CH}_3$  hydrogens in acetic acid are classified as "non-acidic."

- One of the most significant aspects of understanding chemistry is the ability to anticipate whether hydrogen atoms in a material would be capable to dissolve at what time. The hydrogen's that do dissolve may dissolve to varying degrees of intensity. Strong acids, like hydrochloric acid and nitric acid, are efficiently dissolved to a hundred per cent degree in the solution they are in. A limited percentage of mostly organic acids, like acetic acid, is dissolved in many solutions; hence, mostly organic acids are weak acids. Fluoric acid and HCN are instances of inorganic acids that are weak (Malhotra, 2018).
- Sulphuric acid & phosphoric acid are two popular instances of polyprotic acids, acids that have a large number of ionizable hydrogen atoms are recognized as polyprotic acids. Ampholytes

are intermediary types of protons that can both absorb and shed protons, like the  $\text{HPO}_4^{2-}$  molecule.

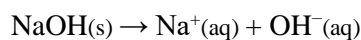


### 1.3. Bases

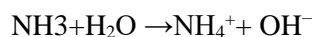
The term "*base*" has often been connected with a group of substances its aqueous solutions have the following characteristics (Bruen & Bruen, 2010):

- a nasty flavour;
- a "*soapy*" texture if applied to the skin;
- The capability to return litmus to its natural blue colour after it's been stained red through acids;
- The ability to produce salts when reacting with acids. The base is identical to the term "alkali." Since old times, woody ashes were the usual resource of the stronger base named as koh; the basic name stems from the similar Latin term kalium (potash), which is also the source of the representation for potassium(K).

Whenever a base is mixed with water, it produces hydroxide ions, similar to when an acid releases hydrogen ions into solution (Pankaj et al., 2014):

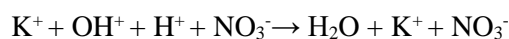


Since it includes  $\text{OH}^-$  ions, sodium-hydroxide (NaOH) is an Arrhenius base. Certain chemicals, on either hand, that don't possess hydroxide ions but may make them by reacting with water are categorized as bases. Hydrogen compounds & metal oxides of some non-metals are 2 types of these compounds:

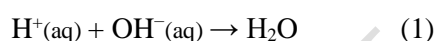


## 1.4. Neutralization

Bases & acids combine to form two items: an ionic & water molecule called salt. It is referred to as a neutralising reaction (Sykes & Matza, 1957).

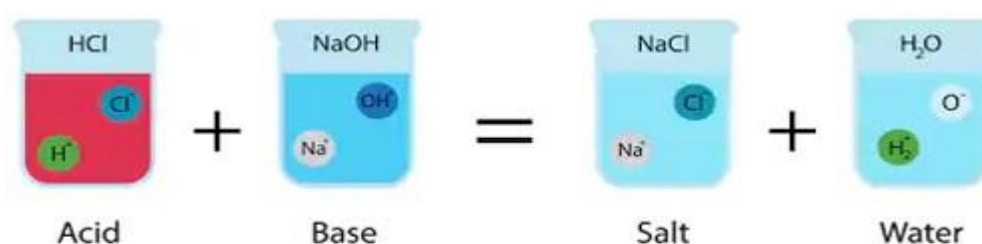


Both such reactions are exothermic; while each employs various bases & acids, these both produce a similar quantity of heat (57.7 kilojoules) for one mole of hydrogen ion neutralised. It indicates that all neutralisation reactions are equivalent to the following one (Maruna & Copes, 2005):



In a neutralisation process, the "salt" created is merely the cation & anion that already existed. Upon evaporating the water, the salt may be retrieved like a solid.

### Acid - base reactions

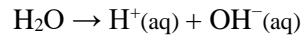


*Figure 1.3. A typical reaction of acid and base*

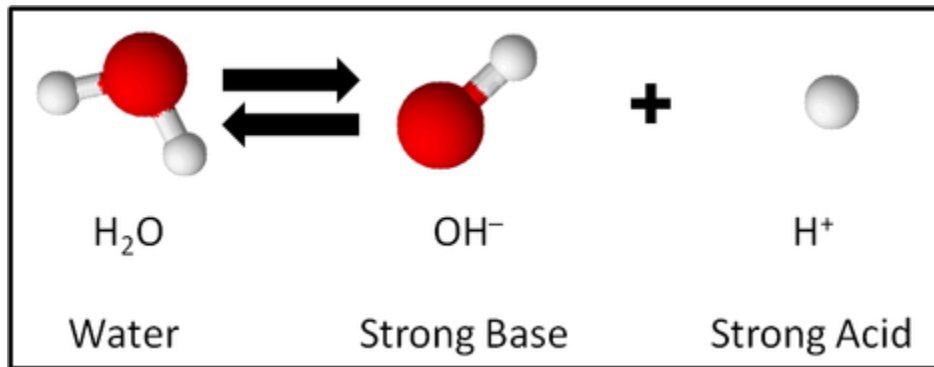
Source: <https://www.assignmentpoint.com/science/chemistry/an-acid.html>

## 1.5. Dissociation of water

Acids' capability to interact with bases is determined by the hydrogen ion's desire to mix with hydroxide ions to produce water. Due to the magnitude of such a propensity, the reaction in Equation 1 is almost completed. Furthermore, no reaction is fully completed; at equilibrium (while no additional net change in the quantities of substances occurs), there would always be the concentration of a very small reactant in the mixture. One more technique of putting it is that every reaction is reversible to some extent. It suggests that the reaction:



would occur to a little amount in distilled water (Feibelman, 2002). Experimental observations verify it: even the thoroughly distilled water prepared by chemists conducts electricity only minimally. Based upon the electrical conductivity, the optimum concentrations of both the hydrogen ion & hydroxide ions at 25°C are approximately precisely  $1.00 \times 10^{-7}$ . This results in the dissociation of around one  $\text{H}_2\text{O}$  molecule per fifty million (Sweeton et al., 1974).



**Figure 1.4. water dissociation reaction**

Source: <https://www.quora.com/What-is-the-pH-of-a-neutral-solution-Why-is-that-the-neutral-point>

Water's degree of separation is so minute that you may question why it is stated whatsoever. The explanation this is necessary is that the constant of equilibrium

$$[\text{H}^+][\text{OH}^-] = 10^{-7} \times 10^{-7} = K_w = 10^{-14}$$

must be defined using the concentrations of hydrogen ion and hydroxide ion in distilled water.

Constants of equilibrium and their computation are discussed further in-depth in a subsequent chapter. For the time being, all that is required is that you understand the ensure that all aspects (Hundt et al., 2014):

At twenty-five degrees Celsius, the multiplication of the hydroxide ion and hydrogen ion concentrations in an aqueous medium of the solution is always  $1.00 \times 10^{-14}$ .

$$\text{In other terms, } [\text{H}^+][\text{OH}^-] = 1.00 \times 10^{-14} \text{ (2)}$$

It is referred to as the multiplication of the ion of the water, and it relates to all aqueous mediums, not only distilled water. That has far-reaching implications since it indicates that even if the concentration of hydrogen ion is high, the concentration of hydroxide ion would be low, and conversely. It suggests that hydrogen ions exist in all water solutions, not only those that are acidic. As a result, the following critical concepts should be memorized (Zabolotskii et al., 1988):

neutral solution:  $[H^+] = [OH^-]$  ( $= 1.00 \times 10^{-7} M$  at  $25^\circ C$ )

alkaline solution:  $[H^+] < [OH^-]$

acidic solution:  $[H^+] > [OH^-]$

## 1.6. The pH scale

The probable values of  $[OH^-]$  and  $[H^+]$  in an aqueous medium of the solution range from around  $10^{-1}$  to  $10^{-15}$  magnitudes. As a result, it's easier to depict them on a compacted logarithmic scale. The *pH* scale is usually utilized to show the concentrations of hydrogen ions (Bjellqvist et al., 1993):

$$pH = -\log_{10} [H^+]$$

or on the contrary,

$$[H^+] = 10^{-pH}$$

you may describe also

$$pOH = -\log_{10} [OH^-]$$

and

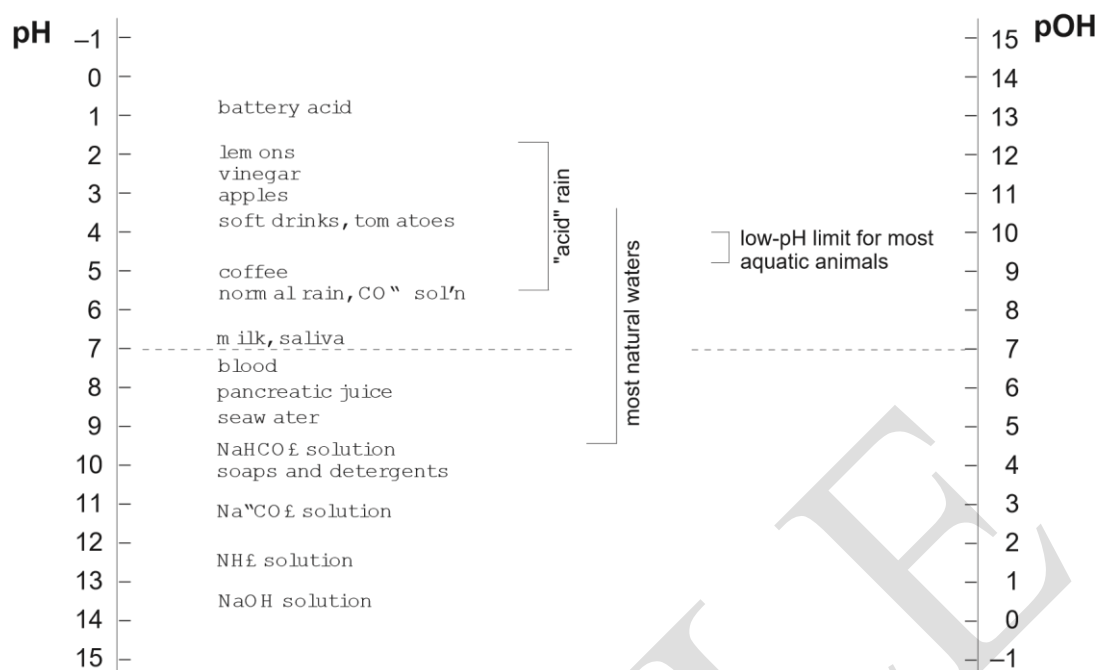
$$pK_w = -\log K_w$$

by Equation 2

$$pH + pOH = pK_w (= 14.0 \text{ in distilled water at twenty-five degrees celsius}) \quad (3)$$

At 25 degrees Celsius, the value of the pH of a neutral solution is 7.0; a higher pH suggests a basic solution, while a lower pH indicates an acidic medium of the solution. A solution having  $[H^+] = 1M$  has a pH of 0; a 0.00010M  $H^+$  solution has a pH of 4.0. A 0.00010M sodium hydroxide solution has a pOH of 4.0, and a pH of 10.0 as a result. You must have a thorough understanding of the pH scale and be able to transform  $[H^+]$  or  $[OH^-]$  to pH in both directions (Bosch et al., 1996).





**Figure 1.5. The pH scale**

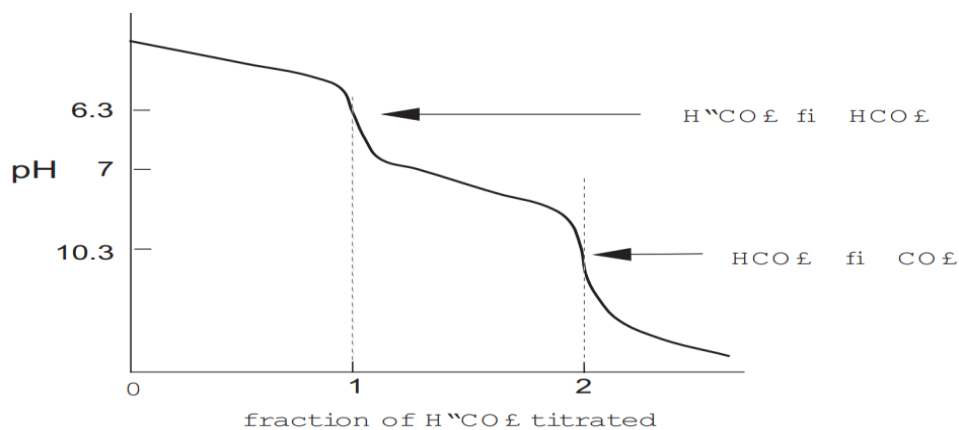
Source: <https://www.sciencenewsforstudents.org/article/scientists-say-ph>

As stated in further explanation on page 9, hydrogen ions don't exist independently in water, and hence the species denoted by "[H<sup>+</sup>]" is a more complex one. Additionally, if the overall ions concentration of all types in the solution surpasses approximately 0.001M, an important fraction would be linked into unbiased pairs like H<sup>+</sup>·Cl<sup>-</sup>, decreasing the concentration of "free" ion to the lower values we would then refer to as the efficient concentration. The efficient concentration of H<sup>+</sup> influences the level to which a solution is acidic, and that's what pH determination techniques measure. Consequently, the pH value is stated in terms of the H<sup>+</sup> concentration that is most efficient. You don't have to worry about the details right now, but it's something to keep in mind later when doing acid-base balance calculations (Bjellqvist et al., 1994).

## 1.7. Titration

Because acids & bases rapidly react, that is fairly simple to determine the quantity of acid in a solution experimentally by estimating how many moles of the base is necessary to neutralize it. This procedure is known as titration, and you must be aware of it in your Lab research (Thordarson, 2011).

We may titrate an acid in the presence of a base or a base in the presence of an acid. The drug being titrated is the chemical whose concentration we are evaluating; the chemical we are injecting in a specified titrant amount. The goal is to continue adding titrant till the solution is completely neutralised; at this stage, the number of titrant moles injected indicates the basic (or acid) concentration in the solution getting titrated (Gaudette et al., 1974).



**Figure 1.6.  $\text{Na}_2\text{CO}_3$  titration curve in the presence of HCl**

Source: <https://www.quora.com/What-are-reasons-for-getting-two-different-concentration-values-of-Na2CO3-when-it-was-titrated-with-HCL-using-Phenolphthalein-and-Methyl-Orange>

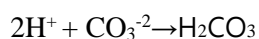
### 1.7.1. Titration curves

Titration progress may be tracked by drawing the pH of the solution as the amount of titrant administered. Figure 1.6 shows 2 identical curves, first for a powerful acid (HCl) and the 2<sup>nd</sup> for a moderate acid, acetic acid (HAc). Notice how the pH varies extremely slowly until the acid is virtually neutralised when looking at the HCl curve first. Only one more drop of sodium hydroxide solution, as seen in the larger image at the top of Fig., would cause the pH to rise to a greater value—nearly as much as the pure sodium hydroxide solution (Tanford, 1963).

Comparing and contrasting the hydrochloric acid and HAc curves. The pH spike around the neutralisation point is much less steep for a weak acid. It's also worth noting that the pH of the solution at the neutralisation point exceeds seven. Such 2 properties of a weak acid's titration curve are critical to understanding (Tanford & Kirkwood, 1957).

There would be a pH spike for every proton titrated whereas if acid or base is polyprotic. The titration of a sodium carbonate solution with HCl is shown in Figure. 1.6. The concentrations of both solutions are the same. Because the  $\text{CO}_3^{2-}$  ion is a base, the pH of the solution begins to rise.  $\text{CO}_3^{2-}$  is converted to  $\text{HCO}_3^-$  and then to carbonic acid,  $\text{H}_2\text{CO}_3$  when protons are injected (Asuero & Kirkwood, 1957).

The vertical portions of the titration curve correspond to quantities of Hydrochloric acid that are equivalent to, as well as twice, the real volume of  $\text{CO}_3^{2-}$  solution (for example, identical numbers of moles due to equal concentrations of both solutions). The fact that one mole of carbonate ion takes two moles of hydrochloric acid to neutralize is demonstrated by the second equal point:



### 1.7.2. Finding the equivalence point: indicators

Whenever a sufficient amount of base is introduced to entirely react with the hydrogens in a monoprotic acid, the equilibrium is achieved. Titration having a stronger alkali & a stronger acid result in a pH of seven at the point of equilibrium. If the strength of the acid is weaker, then the pH would be more than seven; hence, the "neutralised" solution would not be pH neutral. Every titratable hydrogen in a polyprotic acid would have an equivalency point; they generally appear at pH levels four to five units separately (Mocchiutti & Zanuttini, 2007).

The secret to a good titration is determining the point of equivalency. The simplest technique to determine the equivalency point is to utilize an indicator dye; an indicator dye is a material its colour is pH-sensitive. However, one indicator that's also frequently seen in the lab is phenolphthalein; it is colourless in the solution of acidic medium however rapidly changes vividly red in alkaline solution. When acid is being titrated, only a small number of drops of phenolphthalein should be added to the solution before starting the titration. A local red colour occurs as the titrant is introduced, however soon disappears as the solution is agitated. As the equilibrium is near, the colour fades more gradually; the secret is to cease adding base until a solitary drop consequences in a persistently pink colour solution (Checchetti & Lanzo, 2015).

Various indicators exhibit varying degrees of colour change at certain pH levels. Because the pH of the equilibrium fluctuates as the concentration of the acid is titrated, one attempts to match the indicator to the specific acid. The titration of polyprotic acids may be accomplished via utilising an appropriate mix of numerous indicators.

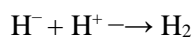
## 1.8. Classification of acids and bases

You would have noted that not each chemical containing hydrogen atoms is acidic;  $\text{NH}_3$  is a base, for instance. Likewise, certain molecules with the  $-\text{OH}$  group are alkaline, while others are acidic. Recognizing which chemicals would display acidic and alkaline characteristics in an aqueous medium is a critical component of knowing chemistry. Luckily, the majority of typical acids and alkalies fit into a limited amount of very well-known groups, making this task quite straightforward (Pearson, 1966).

### 1.8.1. Hydrides as acids and bases

Technically, the term hydride refers to ionic compounds composed of electropositive metals that make up the hydride ion;  $\text{H}^-$ . The term is often used in a wider meaning to refer to any binary compound  $\text{XH}_n$ , where X stands for any element.

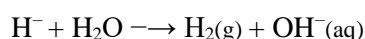
The hydride ion is an excellent proton acceptor:



*Table 1.1. Approximate acid strengths for some binary hydrogen compounds*

								CH <sub>4</sub>
10 <sup>-46</sup>	NH <sub>3</sub>	10 <sup>-35</sup>	H <sub>2</sub> O	10 <sup>-16</sup>	HF			
		PH <sub>3</sub>	10 <sup>-27</sup>	H <sub>2</sub> S	10 <sup>-7</sup>	HCl	10 <sup>7</sup>	
				H <sub>2</sub> Se	10 <sup>-4</sup>	HBr	10 <sup>9</sup>	
				H <sub>2</sub> Te	10 <sup>-3</sup>	HI	10 <sup>10</sup>	

The H<sup>-</sup> ion can't survive in an aqueous medium as it is a stronger base as compared to H<sub>2</sub>O and removes protons from H<sub>2</sub>O:



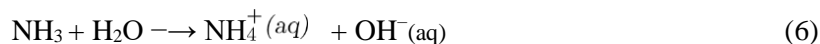
As a result, ionic hydrides are alkaline; molecules like NaH & CaH<sub>2</sub> dissolve in the water to release hydrogen gas, resulting in a basic solution (Balarew & Duhlev, 1984).

The covalent hydrides are acidic, except very weakly. Certain, such as H<sub>2</sub>O and NH<sub>3</sub>, have alkaline qualities as well; nevertheless, the alkaline property predominates in ammonia.

Non-metallic hydrides' acidity grows in proportion to the number of protons in the element to which they are linked. As the element M moves from the left side to the right side of the periodic table or down within such a group, the acids MH become more powerful (Table 1.4).

Efforts to describe such tendencies using a single metric, like M's electronegativity, are usually ineffective. The problem is that various factors determine the acid's strength, such as the strength of the M-H bond and the energy expended when the resultant ions become hydrated in solution. It's better to just study the rule at this point (Lu et al., 2020).

We normally conceive of NH<sub>3</sub> as a base instead of acid; it is amphoteric, like H<sub>2</sub>O, however, the basic qualities prevail in an aqueous medium. Just a proportion of the NH<sub>3</sub> molecules in water would take protons since ammonia is a weak alkali:

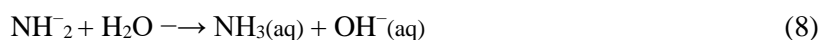


Ammonium hydroxide is the name given to an aqueous medium of NH<sub>3</sub>. That misunderstanding stems from the necessity to presuppose the presence of a material NH<sub>4</sub>OH that may ionize regarding Arrhenius' theory to produce the products of Equation (6) in pre-Bronsted periods. Though there's no bodily proof for the presence of NH<sub>4</sub>OH, the term appears to be permanently inscribed on reagent bottles in the chemistry lab (Carvalho, 1951).

NH<sub>3</sub> generates the amide ion NH<sub>2</sub><sup>-</sup> whenever it interacts like an acid.



Because  $\text{NH}_3$  is a weaker acid, the conjugate base it forms is a powerful acceptor of the proton. Because it has a stronger base as compared to hydroxide ion, and so on, similarly the  $\text{H}^-$  ion, can't reside in water:



When solid  $\text{NaNH}_2$ , is mixed with  $\text{H}_2\text{O}$ , it produces a basic solution with a powerful ammonia odour. As a result, ammonia's acidic character may only be manifested in a solvent apart from water.

### 1.8.2. Hydroxy compounds like acids and alkaline

The organic acids (covered separately further on) are the biggest classification of acids, with compounds bearing the group  $-\text{OH}$  being the biggest group of acids. Among the most prevalent bases are also found in  $\text{M}-\text{OH}$  complexes.

The final examination of the comparative strengths of the  $\text{M}-\text{O}$  &  $\text{O}-\text{H}$  bonds determines that a compound of the common class  $\text{M}-\text{O}-\text{H}$  would operate like an acid or a basic. If the  $\text{M}-\text{O}$  link is weak, the  $-\text{OH}$  component of the molecule would keep its uniqueness and behave like a hydroxide ion. The  $\text{MO}-$  portion of the molecule would stay intact if the  $\text{O}-\text{H}$  bond is weak, and the material would be acidic (Niemelä, 1990).

### 1.8.3. Metal-OH compounds

As a rule, if element  $\text{M}$  is a metal element, the  $\text{MOH}$  compound would be alkaline. It is worth noting that the strongly electropositive elements of Groups 1 & 2 are unique in that their solid  $\text{MOH}$  compounds are joining two structures of metal cations as well as the ions of  $\text{OH}^-$ ; therefore, those which solubilize quickly in water form strongly basic solutions;  $\text{NaOH}$  &  $\text{KOH}$ , on the other hand, are popular instances of strong alkalies. From the perspective of Brønsted, several "bases" are simply diverse reservoirs of the one powerful base hydroxide ion.

As one progresses through Group 2 of the periodic table, the  $\text{M}-\text{OH}$  compounds have lesser solubility; as a result, a saturated solution of  $\text{Ca}(\text{OH})_2$  is only moderately basic. It is so difficult to dissolve the  $\text{OH}^-$  of p-block metal elements as well as transition metals in the water that their solutions are not basic. Nonetheless, because such solids are soluble quickly in an acidic medium to generate salt and water, they are classified as basic in the formal sense (Xu et al., 2018).

### 1.8.4. $-\text{OH}$ compounds of the nonmetals

Because the electronegative central atom displaces the -ve charge from the hydroxylic oxygen atom in such compounds, which are generally called oxyacids, they have an acidic nature. The total result is that the oxygen becomes somewhat more +ve, allowing the hydrogen to leave the cell as  $\text{H}^+$  to leave more easily.

It is important to note that the existence of additional electron gaining groups on the central atom has a significant impact on the strength of an oxyacid compound. The oxygen atom that has formed a double bond is very significant. Excluding the halogen halides, all of the typical strong acids, such as  $\text{NO}_2(\text{OH})$ ,  $\text{SO}_2(\text{OH})$ , and  $\text{PO}(\text{OH})_3$ , include several of these oxygens. Generally, the strength of such acids is determined by the number of oxygen atoms present rather than by any other factor, therefore periodic patterns are not particularly significant. As it occurs, Cl is the only halogen by which all 4 oxyacids have been identified, and the  $K_a$  values for such a series demonstrate how significantly the Cl–O oxygen atoms influence the strength of the acid (Yang et al., 2007).

Acid	Formula	$K_a$ Value
hypochlorous acid	$\text{ClOH}$	$10^{-7.2}$
chlorous acid	$\text{ClO}(\text{OH})$	0.011
chloric acid	$\text{ClO}_2(\text{OH})$	10
perchloric acid	$\text{ClO}_3(\text{OH})$	$\approx 10^{10}$

### 1.8.5. Organic acids

The carboxyl group  $-\text{CO}(\text{OH})$  is the organic acids' distinctive functional group. The carboxylic hydrogen atoms acidity is nearly completely owing to the withdrawal of electrons by the non-hydroxylic oxygen atom; if it weren't, we'd contain alcohol  $-\text{COH}$ , which has acidity still low as compared to the water.

This partial electron removal through a single atom might affect not just the atom next to it, as well as the atom next to it. As a result, the bonding environment of the carbon atom with that it is attached would impact the strength of a carboxylic acid. The inductive effect is a very significant concept in organic chemistry because it propagates partial electron removal to multiple nearby atoms. Contrasting the acetic acid strength and the increasingly more strongly substituted chloroacetic acids provides a nice illustration of the inductive effect induced by chlorine (a very powerful electronegative atom) (Jones, 1998):

Acid	Formula	Strength
acetic acid	$\text{CH}_3-\text{COOH}$	$1.8 \times 10^{-5}$
Monochloroacetic acid	$\text{ClCH}_2-\text{COOH}$	0.0014
dichloroacetic acid	$\text{Cl}_2\text{CH}-\text{COOH}$	0.055
trichloroacetic acid	$\text{Cl}_3-\text{COOH}$	0.63

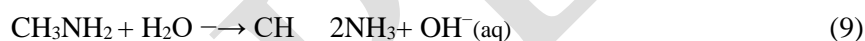
**Phenols:** The increased acidity of the  $-\text{OH}$  group as impacted through the 2nd atom of the oxygen which composes the  $-\text{COOH}$  group results in the carboxyl group's acidic nature. Because the benzene

ring has a comparable but lesser withdrawing of electron function, hydroxyl groups connected to benzene rings behave as acids too. Phenol, C<sub>6</sub>H<sub>5</sub>OH, is a very well instance of an acid. Phenolic acids are weaker in comparison to carboxylic acids (Ricke, 2003):

Acid	Formula	Strength
acetic acid	CH <sub>3</sub> -COOH	$1.8 \times 10^{-5}$
benzoic acid	C <sub>6</sub> H <sub>5</sub> -COOH	$6.3 \times 10^{-5}$
Phenol	C <sub>6</sub> H <sub>5</sub> OH	$1.1 \times 10^{-10}$

### 1.8.6. Organic bases and amines

Given that we contain just examined organic acids, maybe a brief mention of organic bases are in order. Because the -OH group is acidic instead of basic when attached to carbon, alcohols aren't analogues of inorganic hydroxy compounds. Amines are a very abundant class of organic alkalies. They are composed of the -NH<sub>2</sub> group coupled to a carbon atom. In water, amines form weak alkaline solutions:



Amines are the byproducts of microorganisms degrading nitrogenous organic compounds like proteins. They frequently smell strong like "rotten fish." This is not coincidental, since seafood includes a high concentration of nitrogen-containing chemicals that degrade rapidly. Methylamine, CH<sub>3</sub>NH<sub>2</sub>, is particularly suitable to make itself known to humans since it is a gas at room temperature. When lemon juice or another acidic material is added to fish, the methylamine is converted to the methylammonium ion CH<sub>3</sub>NH<sub>3</sub><sup>+</sup>. Since ions are non-volatile, they contain no fragrance (Greenberg, 1966).

### 1.8.7. Oxides as acids and bases

The separation of oxygen compounds into acidic & basic oxygen compounds is broadly analogous to the divide of -OH compounds. The compounds of oxygen of the Groups 1-2 highly electropositive metals have the O<sup>2-</sup> ion. Such ion is another example of a proton acceptor which is strong as compared to OH<sup>-</sup> and hence can't reside in water. As a result, ionic oxides tend to produce extremely basic solutions:



In other circumstances, like MgO, the solid is sufficiently insoluble that it causes a minimal change in pH when added to water. Furthermore, CaO, often referred to as *quicklime*, is enough soluble to create a highly alkaline solution when adequate heat is generated; the result is the somewhat dissolvable *slaked lime*, Ca(OH)<sub>2</sub> (Pearson & Songstad, 1967).

Transition metal-oxygen compounds are often insoluble solids with relatively complicated comprehensive lattices. While certain would soluble in acids, they are not acidic when dissolved in H<sub>2</sub>O.

### 1.8.8. Acid anhydrides

When the binary oxygen compounds of the non-metallic elements are introduced to water, they tend to form acidic solutions. Acid anhydrides are compounds like SO<sub>3</sub>, CO<sub>2</sub>, and P<sub>4</sub>O<sub>6</sub> that are occasionally used to describe a group of compounds (acids having no water).



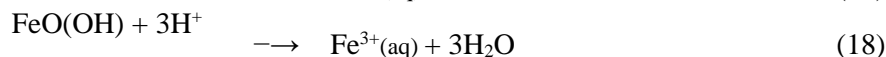
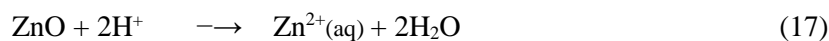
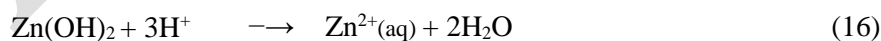
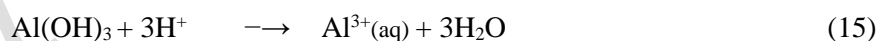
In certain circumstances, the reaction entails more than the incorporation of water's constituents. Consequently, NO<sub>2</sub>, which is utilised in the commercial manufacture of NH<sub>3</sub>, is not strictly an anhydride:



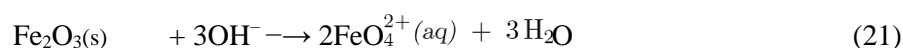
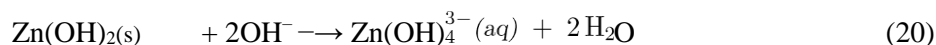
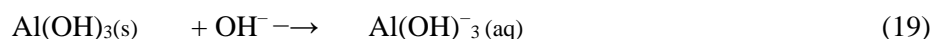
### 1.8.9. Amphoteric oxides and hydroxides

The hydroxides and oxides of Group 3 and above metals are typically very mildly basic, and the majority is amphoteric. The majority of such compounds are so water-insoluble that their acidity or alkalinity properties become apparent only when they react with strong acids or alkalis.

By and large, such compounds are more basic as compared to acidic; for example, the hydroxides and oxides of iron, aluminium, & zinc are all soluble in concentrated acid (Pearson, 1963):



Furthermore, in concentrated OH<sup>-</sup> solutions, such compounds create anionic species that are the oxide or hydroxide's conjugate bases:





These are referred to as the ferrate, aluminate, and zincate ion. Additional products are generated in which just a portion of the  $\text{-OH}$  groups of the parent  $\text{OH}^-$  are deprotonated, resulting in a sequence of such oxyanions for the majority of metals.

### 1.8.10.Salts

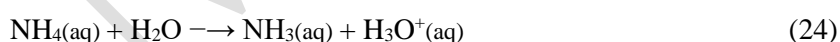
In distilled water, sodium chloride is completely neutral because the chloride ion, being the conjugate base of a very strong acid, is a very weak proton acceptor. When a specific amount of sodium acetate  $\text{CH}_3\text{COO}^- \text{Na}^+$  is solubilized in water, the solution turns basic. Because HF and  $\text{CH}_3\text{COOH}$  are both weak acids, their conjugate bases  $\text{F}^-_{(\text{aq})}$  and  $\text{CH}_3\text{COO}^-_{(\text{aq})}$  would be good proton acceptors (Pearson, 1987).

At the very least, they are powerful enough to behave as weak bases in  $\text{H}_2\text{O}$ :



Consequently, the rule (that you should be familiar with) states that a water solution of a salt of a weak acid would be basic. This activity is occasionally referred to as hydrolysis, a throwback to a time when the notion of proton transfer acid-base reactions gained acceptance.

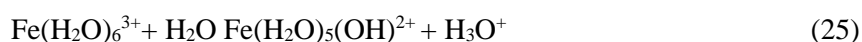
When the salt has the ammonium ion as its cation, hydrolysis would result in an acidic solution when no weak-acid anions are present to counteract the effect. Therefore, an ammonium sulphate solution would be acidic (Hundelshausen, 1971):



### 1.8.11.Metal cations

Iron(III) chloride is a bright orange solid that solubles readily in water to form a highly acidic solution. How is that possible? Because neither the  $\text{Cl}^-$  nor the  $\text{Fe}^{3+}$  ions possess protons, how are they able to give protons to  $\text{H}_2\text{O}$  to form  $\text{H}_3\text{O}^+$ ?

The protons originate from the water molecules in the metal cation's main hydration shell. These are the water molecules, typically 6 in number, which is most closely connected to the cation via ion-dipole attraction. If the cation has a charge of +2 or more, the electric field intensity near the hydration shell's edge is sufficient to encourage the release of a hydrogen ion from one of the  $\text{H}_2\text{O}$  molecules (Duhlev et al., 1991):

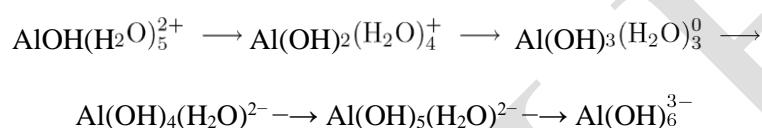


A solution of  $\text{FeCl}_3$  is a stronger acid as compared to an equimolar solution of acetic acid as a result of this reaction. A  $\text{FeCl}_2$  solution, on the other hand, would be a lot weaker acid; the +2 charge would be

much lesser helpful in reducing proton loss. Alkali metals and ions like  $\text{Ag}^+(\text{aq})$  have very little acidity; in general, the smaller and more strongly charged the cation, the more acidic it is.

Ion	$\text{In}^{3+}$	$\text{Bi}^{3+}$	$\text{Fe}^{3+}$	$\text{Sn}^{2+}$	$\text{Fe}^{2+}$	$\text{Cu}^{2+}$	$\text{Mg}^{2+}$
Acid Constant	0.6	.01	.007	$1\text{E}-4$	$5\text{E}-9$	$5\text{E}-9$	$1.6\text{E}-13$

A hydrated cation must be capable of losing more than one proton. An  $\text{Al}(\text{H}_2\text{O})_6^{3+}$  ion, for instance, must produce the following species in order:



Furthermore, when the charge declines from a high +ve to a -ve value, the elimination of protons becomes more difficult; the final 3 species have not been observed in solution. The main species in dilute aluminium chloride solutions are  $\text{Al}^{3+}$  &  $\text{AlOH}^{2+}$  (such as  $\text{Al}(\text{H}_2\text{O})_6^{3+}$  and  $\text{AlOH}(\text{H}_2\text{O})_5^{2+}$ ).

The issue is compounded in highly concentrated solutions by processes in which 2 or more hydroxylations polymerize form multi-centre complexes. In concentrated aluminium ion solutions, for instance, the 2 aluminium species stated above are mostly  $\text{Al}_4(\text{OH})_{10}^{2+}$  and  $\text{Al}_6(\text{OH})_{15}^{3+}$ . Likewise, the major ions in bismuth salt solutions are  $\text{BiOH}^{2+}$ ,  $\text{Bi}^{3+}$ , and  $\text{Bi}_6(\text{OH})_{12}^{6+}$ ;  $\text{Bi}(\text{OH})_2^+$  is not present (Brown & Duhlev, 1991).

SAMPLE