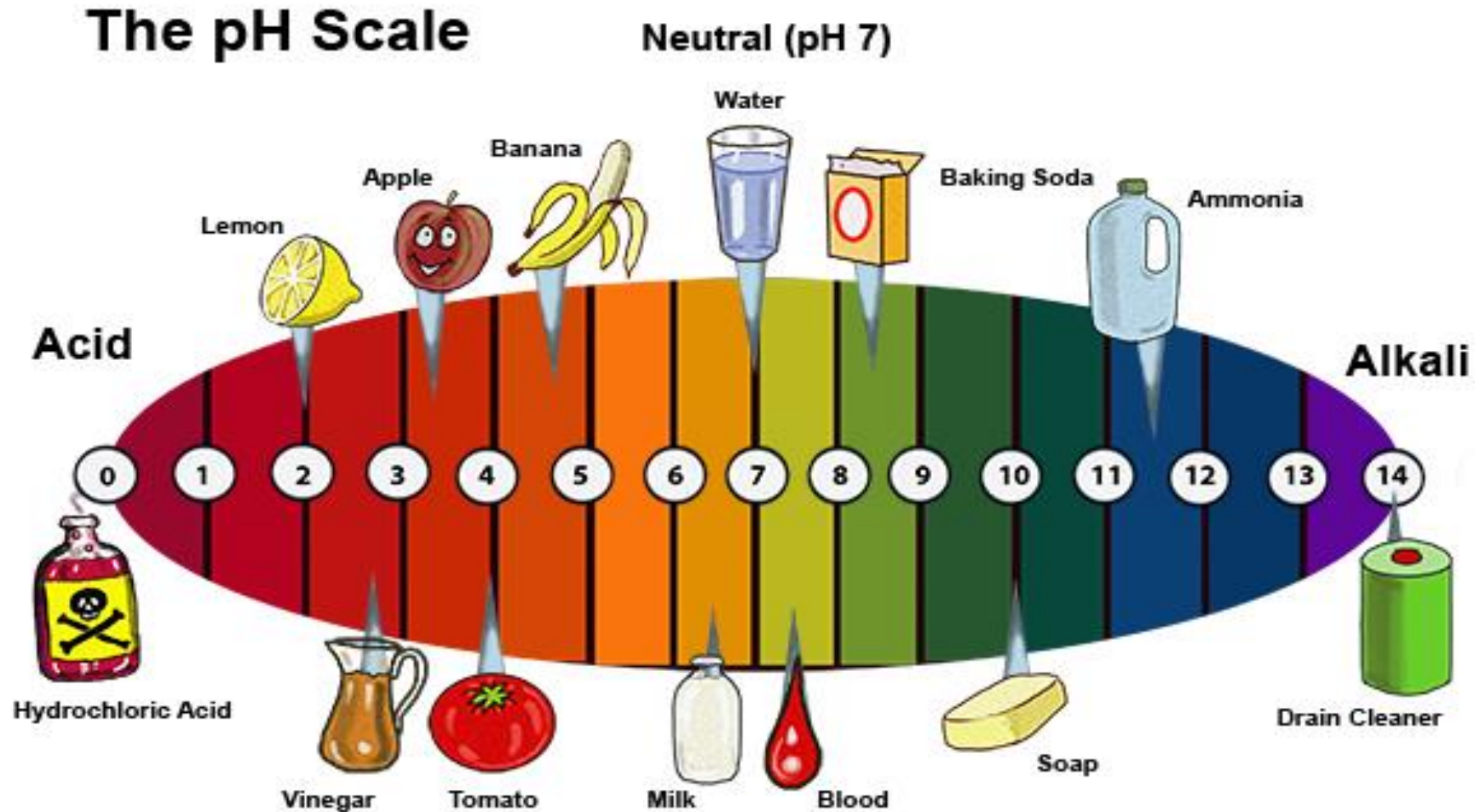


Acids And Bases



Optimist



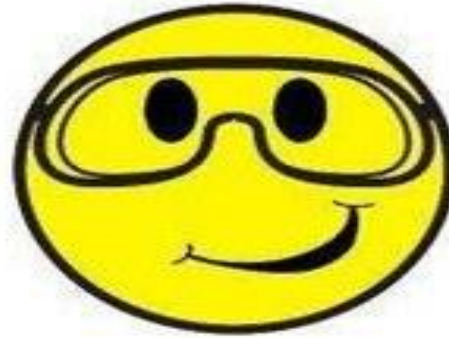
**The Glass
is Half
Full**

Pessimist

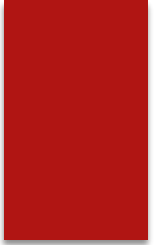


**The Glass
is Half
Empty**

Chemist



**The Glass
Contains
50% $\text{H}_2\text{O}(\text{l})$
39% $\text{N}_2(\text{g})$
10.5% $\text{O}_2(\text{g})$
.44% $\text{Ar}(\text{g})$
.06% $\text{CO}_2(\text{g})$**

- 
- There are three concepts of acids and bases in current use.
 - Each has its own peculiar advantages
 - These are-
 - Arrhenius concept
 - Bronsted-Lowry concept
 - Lewis concept

Svante Arrhenius (1859-1927)

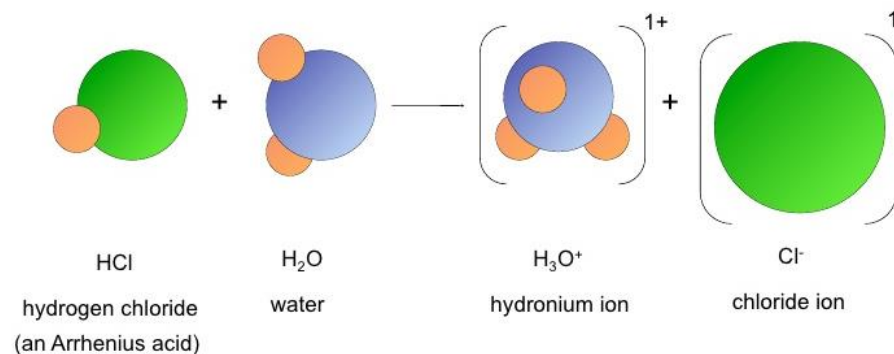
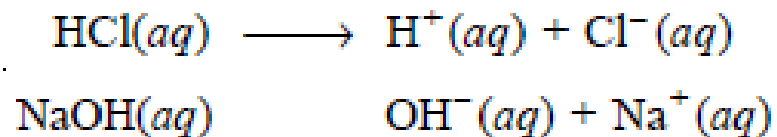
Swedish scientist



- The Nobel Prize in Chemistry 1903
- "in recognition of the extraordinary services he has rendered to the advancement of chemistry by his electrolytic theory of dissociation."

Arrhenius Concept Of Acids And Bases

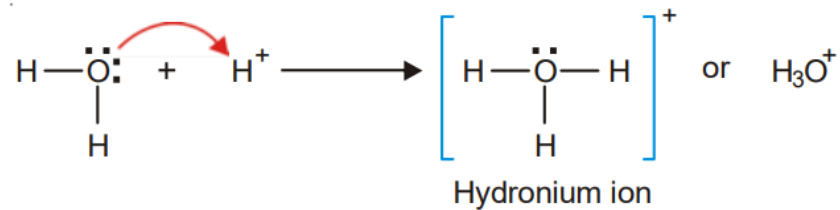
- Savante Arrhenius (1884) proposed his concept of acids and bases.
- According to this concept, an acid is a compound that releases H^+ ions in water; and a base is a compound that releases OH^- ions in water.
- For example, HCl is an Arrhenius acid and NaOH is an Arrhenius base.



Limitations of Arrhenius concept

Free H^+ and OH^- ions do not exist in water:

- The H^+ and OH^- ions produced by acids and bases respectively do not exist in water in the free state.
- They are associated with water molecules to form complex ions through hydrogen bonding.
- Thus the H^+ ion forms a hydronium ion



- Similarly, OH^- ion forms the complex H_3O_2^-

Limitations of Arrhenius concept

Limited to water only:

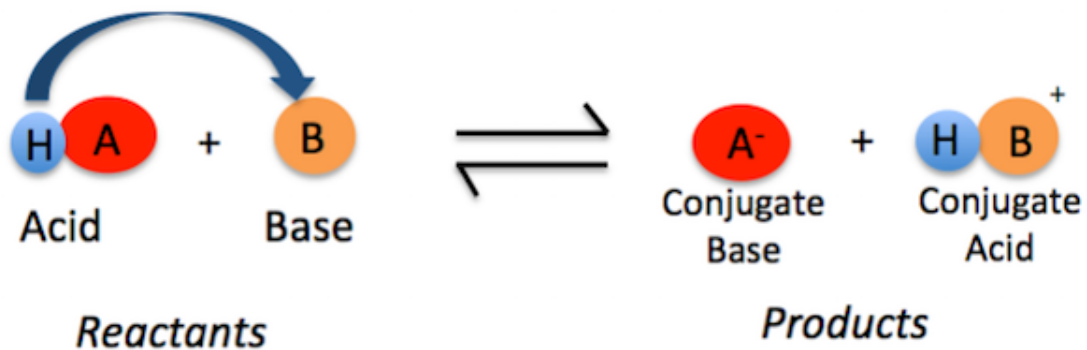
- It is only applicable in aqueous solutions.
- Arrhenius defined acids and bases as compounds producing H^+ and OH^- ions in water only.
- But a truly general concept of acids and bases should be appropriate to other solvents as well.

Some bases do not contain OH^- :

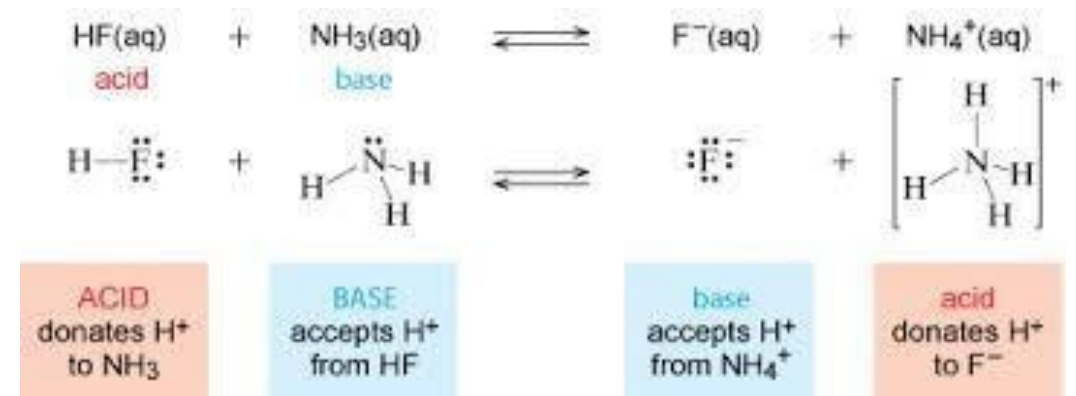
- Some bases do not contain OH^- Arrhenius base is one that produces OH^- ions in water.
- Yet there are compounds like ammonia (NH_3) and calcium oxide (CaO) that are bases but contain no OH^- ions in their original formulation.

Bronsted-Lowry Concept

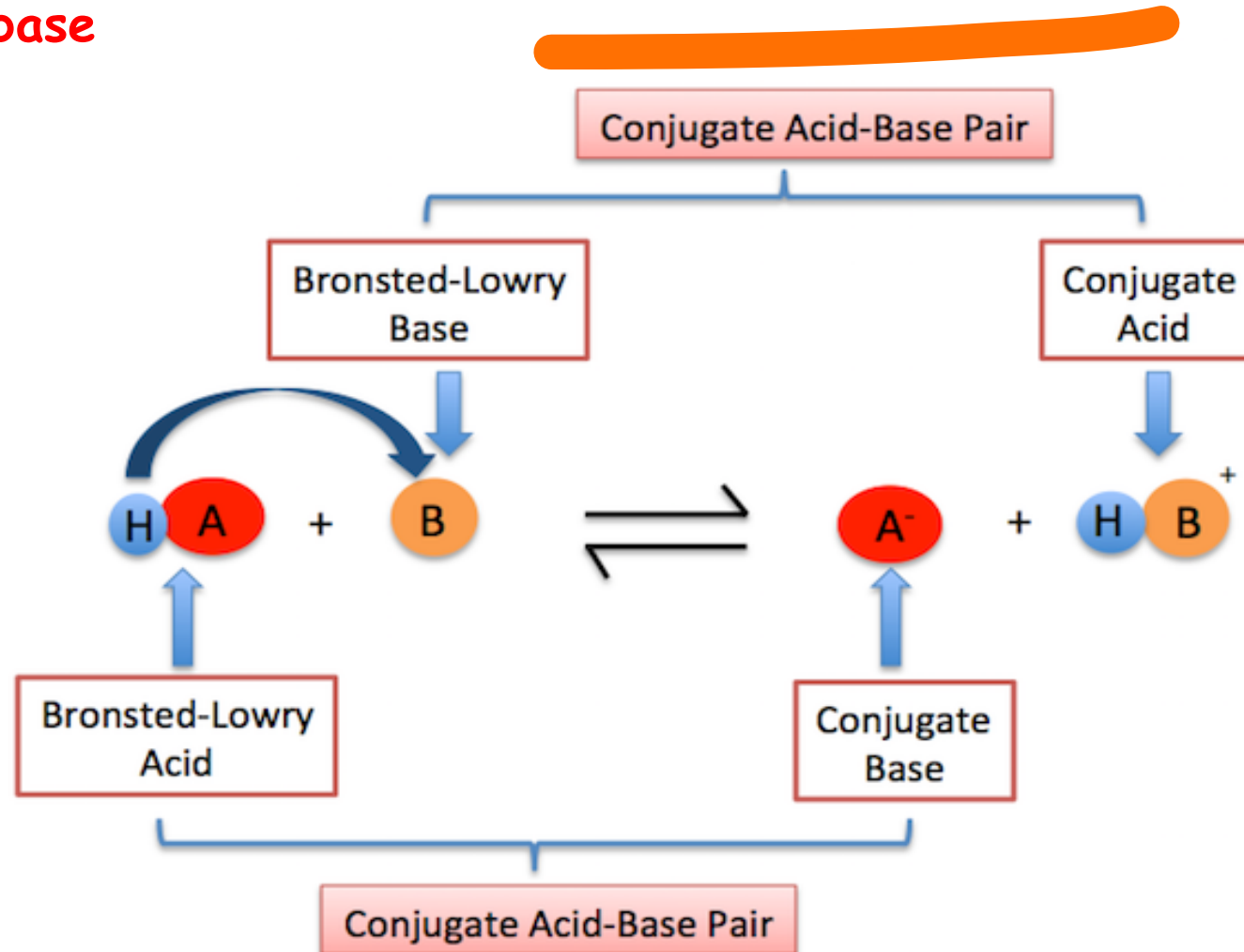
- In 1923 J.N. Bronsted and J.M. Lowry independently proposed a broader concept of acids and bases. According to this theory,
- an acid is any molecule or ion that can donate a proton (H^+)
- a base is any molecule or ion that can accept a proton
- For brevity we can say that an acid is a proton donor while a base is a proton acceptor.



Bronsted-Lowry Acid-Base Reaction

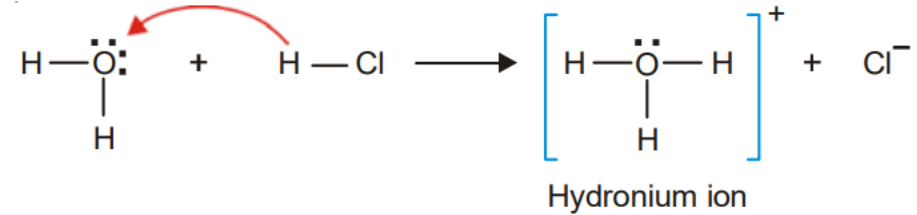


- An acid qualifying Bronsted-Lowry concept is termed a **Bronsted-Lowry acid** or **simply Bronsted acid**.
- A base qualifying Bronsted-Lowry concept is termed a **Bronsted-Lowry base** or **simply Bronsted base**.

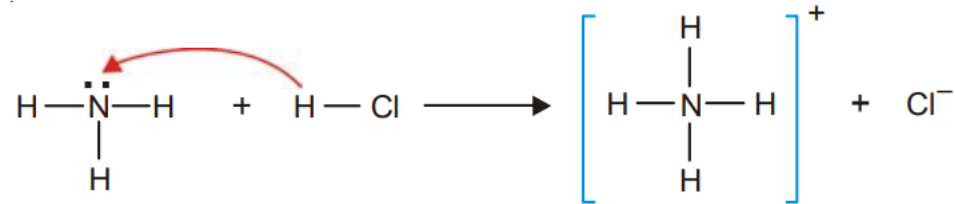


Examples of Bronsted acids and bases

- HCl gas and H₂O.



- HCl and Ammonia, NH₃.



- Calcium oxide and H₂O



Bronsted-Lowry concept is superior to Arrhenius concept

Much wider scope

- Arrhenius concept of acids and bases is restricted to the study of substances which can release H^+ or OH^- ions in water.
- Bronsted-Lowry concept embraces all molecules and ions that can donate a proton (acids) and those which can accept a proton (bases).

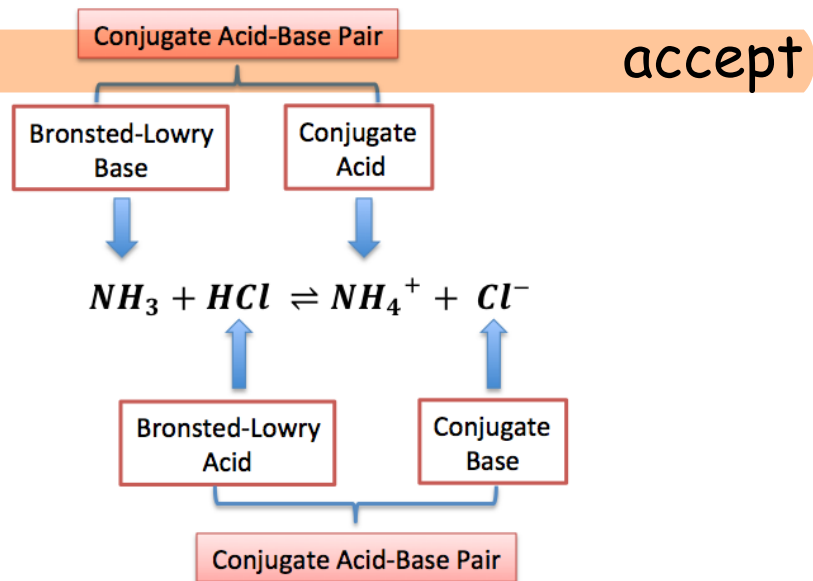
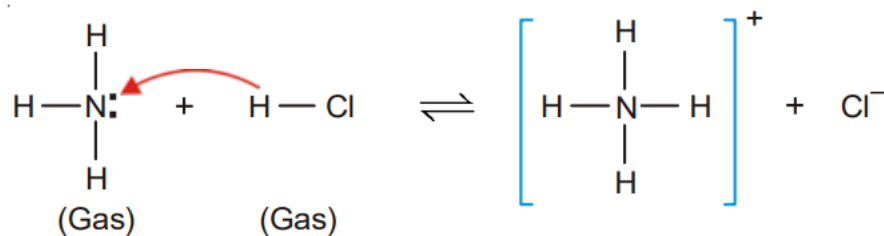
Not limited to aqueous solutions.

- The Bronsted-Lowry model is not limited to aqueous solutions as is the case with Arrhenius model.
- It can be extended even to the gas phase.
- For example, gaseous ammonia (a Bronsted base) can react with hydrogen chloride gas (a Bronsted acid) to give ammonium chloride.

Bronsted-Lowry concept is superior to Arrhenius concept

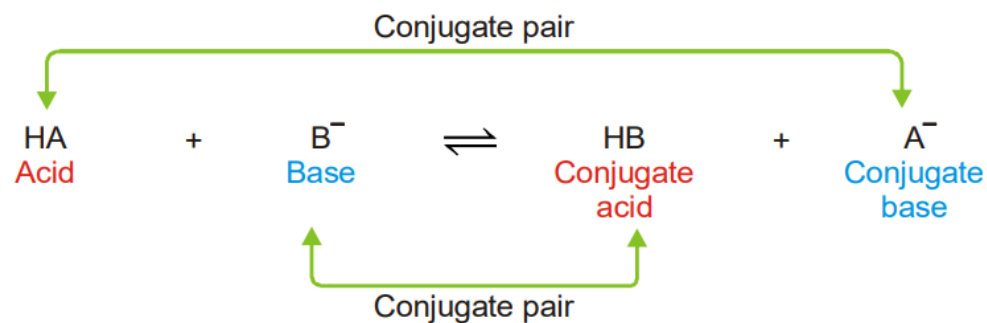
Release of OH⁻ not necessary to qualify as a base.

- Arrhenius base is a substance that releases OH⁻ ions in water.
- On the other hand, Bronsted base is a substance that accepts a proton.
- Thus, liquid ammonia (NH₃) does not produce OH⁻ ions in water but it is a recognised base.
- But according to Bronsted-Lowry model, it qualifies as a base because it accepts a proton to form NH₄⁺ (an acid).



Conjugate Acid-Base pairs

- The acid (HA) and the conjugate base (A^-) that are related to each other by donating and accepting a single proton, are said to constitute a conjugate Acid-Base pair
- It may be noted that in any acid-base reaction, there are two conjugate acid-base pairs



Lewis concept of acids and bases

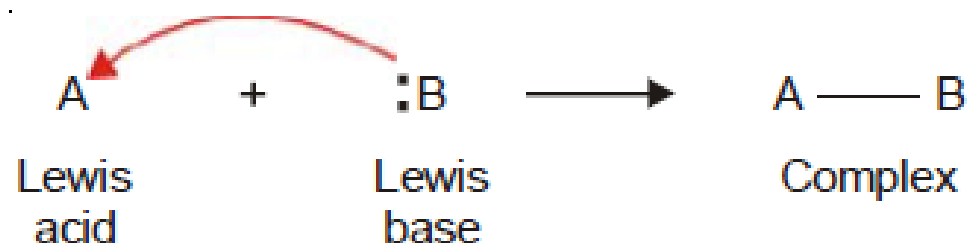
LEWIS CONCEPT OF ACIDS AND BASES

In the early 1930s, G.N. Lewis proposed even a more general model of acids and bases. According to Lewis theory,

an acid is an electron-pair acceptor

a base is an electron-pair donor

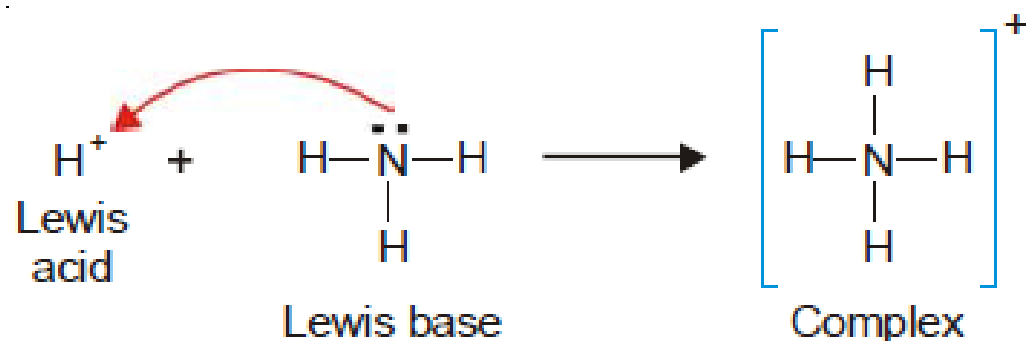
Lewis pictured an acid and base as sharing the electron pair provided by the base. This creates a covalent bond (or coordinate bond) between the Lewis acid and the Lewis base. The resulting combination is called a Complex. If the Lewis acid be denoted by A and the Lewis base by B, then the fundamental equation of the Lewis theory can be written as :



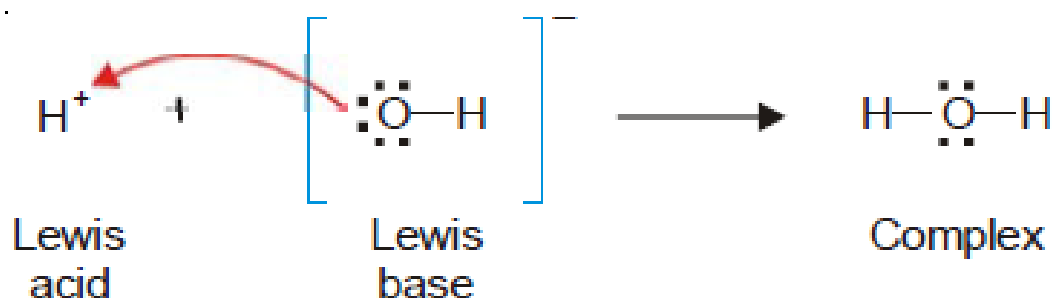
It may be noted that : (1) all cations or molecules short of an electron-pair act as Lewis acids; and (2) all anions or molecules having a lone electron-pair act as Lewis bases.

Examples of Lewis reactions

(1) Between H^+ and NH_3 . Proton (H^+) is a Lewis acid as it can accept an electron-pair. Ammonia molecule ($:\text{NH}_3$) has an electron-pair which it can donate and is a Lewis base. Thus the Lewis reaction between H^+ and NH_3 can be written as :



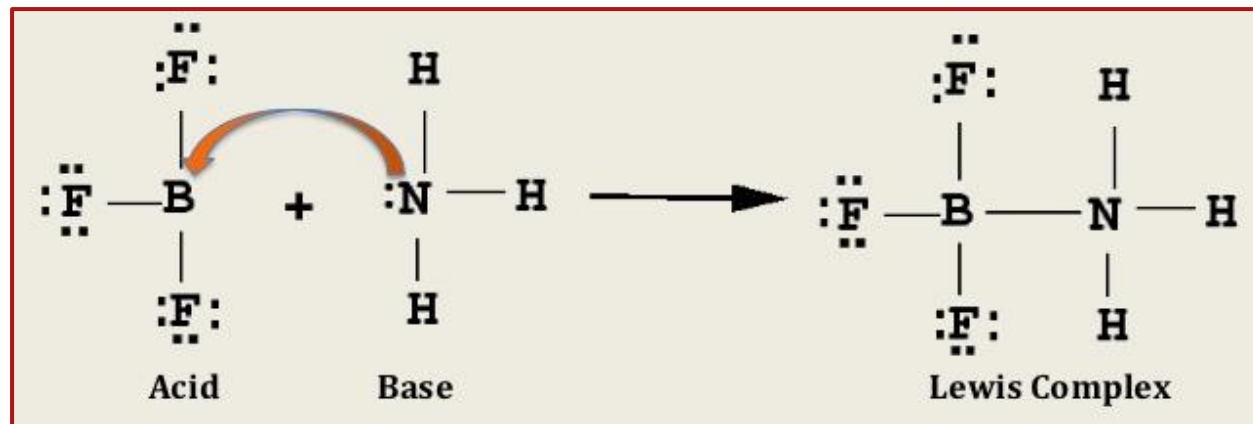
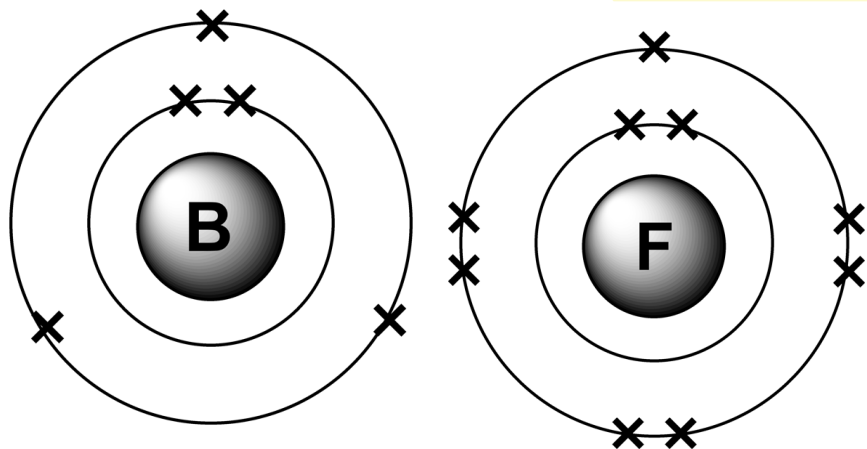
(2) Between H^+ and OH^- . A proton (H^+) is an electron-pair acceptor and, therefore, a Lewis acid. The OH^- is an electron-pair donor and hence a Lewis base. Thus Lewis reaction between H^+ and OH^- can be written as :



Superiority of Lewis model of acids and bases

The advantages of the Lewis acid-base model are :

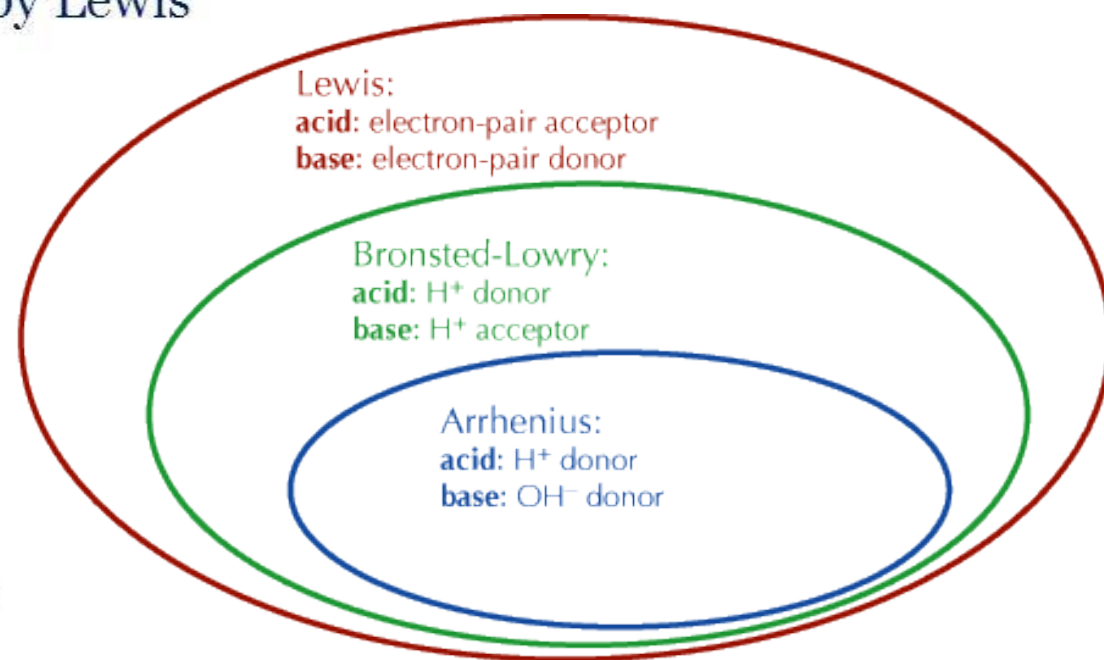
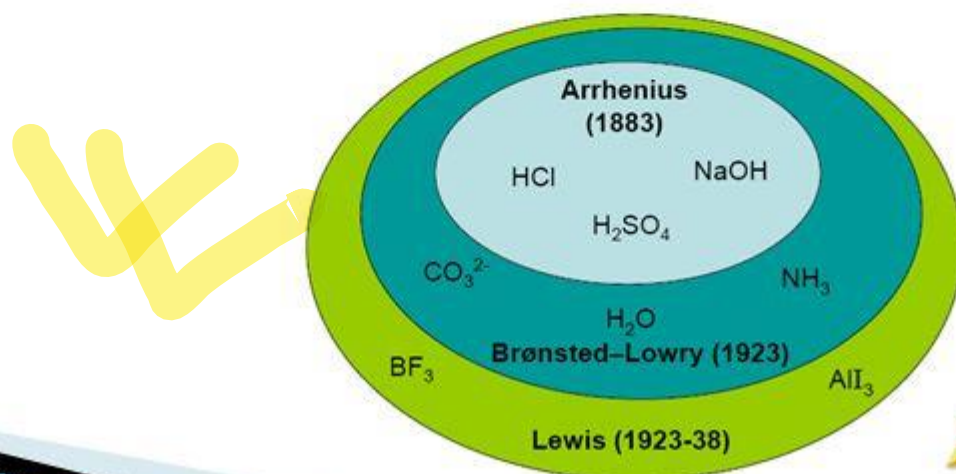
- (1) All the Bronsted-Lowry acid base reactions are covered by the Lewis model. It is so because the transfer or gain of a proton is accompanied by the loss or donation of an electron-pair in both types of reactions.
- (2) Many reactions which do not involve transfer of a proton are also covered by the Lewis theory. e.g.,



Comparison of Bronsted-Lowry and Lewis Theories of Acids and Bases

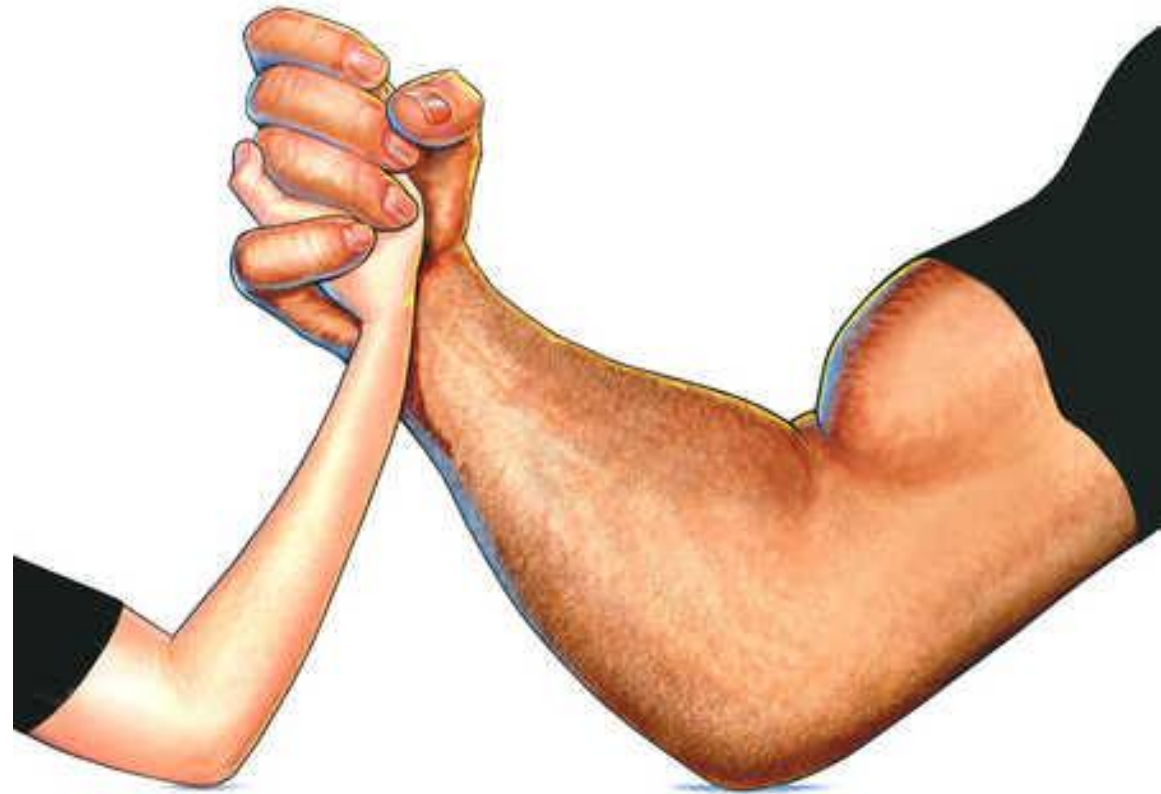
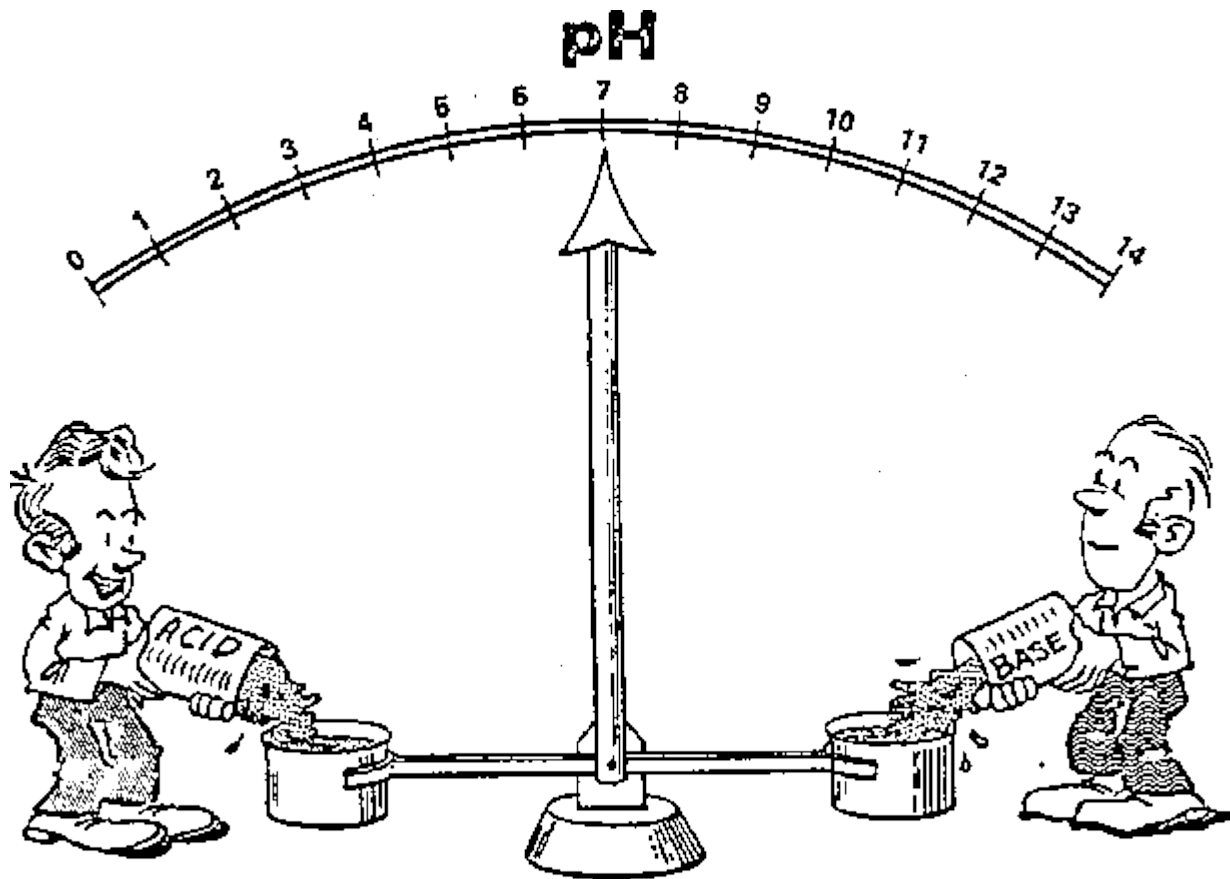
Theory	Definition of acid	Definition of base
Bronsted-Lowry	Proton donor	Proton acceptor
Lewis	Electron pair acceptor	Electron pair donor

Although all Bronsted-Lowry acids are Lewis acids, not all Lewis acids are Bronsted-Lowry acids, so the term Lewis acids are usually reserved for those species which can only be described by Lewis theory (that is those that do not release H^+).





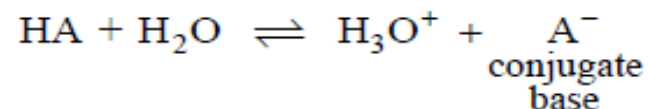
Relative Strength Of Acids And Bases



Relative strength of acids

- The strength of an acid is defined as the concentration of H^+ ions in its aqueous solution at a given temperature.
- The strength of an acid depends on its ability to transfer its proton (H^+) to a base to form its conjugate base.
- When a monoprotic acid (HA) dissolves in water, it transfers its proton to water (a Bronsted base) to form hydronium ion (H_3O^+) and a conjugate base.

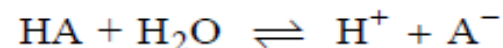
ACIDS AND BASES



For simplifying our discussion, we take



Thus we can write the equilibrium reaction (1) as



This equation represents the dissociation of the acid HA into H^+ ion and A^- ion.

Applying the Law of Mass action to the acid dissociation equilibrium, we can write

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

Calculation of Relative strength of Weak acids from K_a

We have seen that for an acid in aqueous solution we have, $\text{HA} \rightleftharpoons \text{H}^+ + \text{A}^-$

and
$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

Let C moles per litre be the concentration of the acid and α its degree of dissociation. Then,

$$[\text{H}^+] = C\alpha$$

$$[\text{A}^-] = C\alpha$$

$$[\text{HA}] = C(1 - \alpha)$$

Substituting the values in the equilibrium expression we have

$$\begin{aligned} K_a &= \frac{C\alpha \times C\alpha}{C(1 - \alpha)} \\ &= \frac{C\alpha^2}{(1 - \alpha)} \end{aligned}$$

For weak acids $1 - \alpha \approx 1$. Therefore,

$$K_a = C\alpha^2$$

For two different acids, 1 and 2, let the degree of dissociation be α_1 and α_2 ; and the dissociation constants K_1 and K_2 . Then,

$$\text{for acid 1} \quad K_1 = C\alpha_1^2 \quad \dots(1)$$

$$\text{for acid 2} \quad K_2 = C\alpha_2^2 \quad \dots(2)$$

Dividing equation (1) by (2), we get,

$$\frac{\alpha_1}{\alpha_2} = \sqrt{\frac{K_1}{K_2}}$$

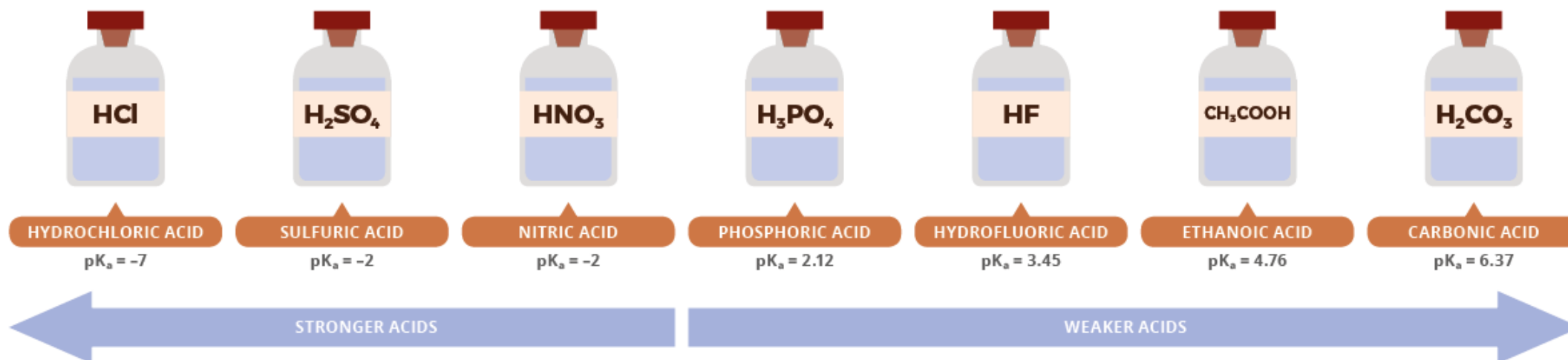
Since $[\text{H}^+]$ is a measure of acid strength and it depends on the degree of dissociation α , we can write

$$\frac{\text{Strength of acid 1}}{\text{Strength of acid 2}} = \sqrt{\frac{K_1}{K_2}}$$

Evidently, the ratio $\sqrt{K_1/K_2}$ would give us the relative strengths of the two acids.

A GUIDE TO ACIDS, ACID STRENGTH, AND CONCENTRATION

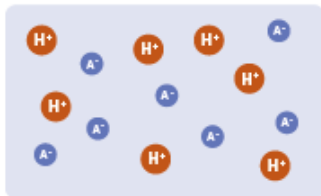
What's the difference between acid strength and concentration? And how does pH fit in with these? This graphic explains the basics.



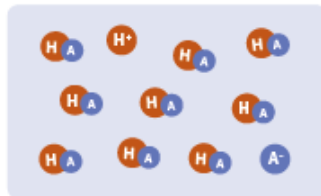
STRONG ACIDS VS. WEAK ACIDS



The H⁺ ion is transferred to a water molecule, forming H₃O⁺



STRONG ACID



WEAK ACID

Hydrogen ions Negative ions Acid molecules

Acids react with water when they are added to it, forming ions. The degree to which they do this is what determines whether they are strong or weak acids. Strong acids are essentially 100% ionised in solution. Weak acids ionise very little in solution.

ACIDS, K_a AND pK_a

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

$$\text{pK}_a = -\log_{10}[K_a]$$

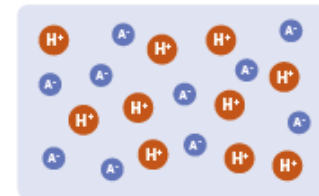
	K _a	pK _a
VERY STRONG ACID	>0.1	<1
FAIRLY STRONG ACID	10 ⁻³ –0.1	1–3
WEAK ACID	10 ⁻⁵ –10 ⁻³	3–5
VERY WEAK ACID	10 ⁻¹⁵ –10 ⁻⁵	5–15
EXTREMELY WEAK ACID	<10 ⁻¹⁵	>15

The acid dissociation constant, K_a, is a measure of the strength of an acid. The higher its value, the stronger the acid (i.e. the more readily it ionises in water). pK_a converts K_a number to a logarithmic scale that makes it easier to compare strengths of different acids.

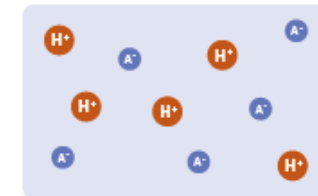
CONCENTRATION AND pH

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

A decrease of one on the pH scale represents a tenfold increase in H⁺ concentration.



CONCENTRATED ACID



DILUTE ACID

Hydrogen ions Negative ions

Concentration is distinct from strength. It refers to the amount of acid in a given solution. A concentrated acid contains a large amount of acid in a given volume; a dilute solution contains a small amount. The pH scale gauges the amount of hydrogen ions in solution.

Calculation of Relative strength of Weak acids from K_a

Do yourself:

- The dissociation constants of formic acid and acetic acid are 21.4×10^{-5} & 1.81×10^{-5} respectively. Find the relative strengths of the acids
- Two hypothetical acids HA and HB have the dissociation constants 1×10^{-3} and 1×10^{-5} respectively in water at 25°C . Calculate the strength of HA with respect to HB

Relative strength of bases

- The strength of a base is defined as the concentration of OH^- ions in its aqueous solution at a given temperature.

According to the Arrhenius concept, a base is a substance which produces OH^- ions in aqueous solution. The basic properties of such a substance are due to these hydroxyl ions. Let us consider a base BOH whose dissociation can be represented as



Applying the Law of Mass action to the above equilibrium we can write the equilibrium expression as

$$K_b = \frac{[\text{B}^+][\text{OH}^-]}{[\text{BOH}]} \quad \dots(2)$$

From the equilibrium expression (2), it is evident that the concentration of OH^- ions, $[\text{OH}^-]$, depends on the value of K_b . Therefore, the value of K_b for a certain base is a measure of its base strength. In the aqueous solution of a strong base, practically all the original base is dissociated and the value of K_b is large. In the case of a weak base, it is dissociated in aqueous solution to a very small extent and the value of K_b is also small.

Relative Strengths of Acid and Bases

As acid strength decreases, base strength increases;
the weaker the acid, the stronger its conjugate base.

As base strength decreases, acid strength increases;
the weaker the base, the stronger its conjugate acid.

Strong 100% ionized (H ⁺ completely donated to water)		Extremely weak (negligible H ⁺ acceptance from water)			
Conjugate acid		Conjugate base			
Strong	H ₂ SO ₄	HSO ₄ ⁻	Extremely weak		
	HBr	Br ⁻			
	HCl	Cl ⁻			
	HNO ₃	NO ₃ ⁻			
Acid strength increasing ↑	H ₃ O ⁺	H ₂ O	Base strength increasing ↓		
	H ₂ SO ₃ (sulfurous)	HSO ₃ ⁻			
	HSO ₄ ⁻	SO ₄ ²⁻			
	H ₃ PO ₄ (phosphoric)	H ₂ PO ₄ ⁻			
	HF (hydrofluoric)	F ⁻			
	HNO ₂ (nitrous)	NO ₂ ⁻			
	CH ₃ COOH (acetic)	CH ₃ COO ⁻			
	H ₂ CO ₃ (carbonic)	HCO ₃ ⁻			
	H ₂ S (hydrosulfuric)	HS ⁻			
	H ₂ PO ₄ ⁻	HPO ₄ ²⁻			
	NH ₄ ⁺	NH ₃			
	HCN	CN ⁻			
	HCO ₃ ⁻	CO ₃ ²⁻			
	H ₂ O	OH ⁻			
	Extremely weak	OH ⁻		O ²⁻	Strong
		H ₂		H ⁻	
		CH ₄		CH ₃ ⁻	
	Extremely weak (negligible H ⁺ donation to water)			100% reacted with H ⁺ from water	

The pH of Solutions

- A knowledge of the concentration of hydrogen ions (more specifically hydronium ions) is of the greatest importance in chemistry.
- Hydrogen ion concentrations are typically quite small numbers.
- Therefore, chemists report the hydrogen ion concentration of a solution in terms of pH.
- It is defined as "the negative of the base-10 logarithm (log) of the H⁺ concentration". Mathematically it may be expressed as

$$\text{pH} = -\log [\text{H}^+]$$

where $[\text{H}^+]$ is the concentration of hydrogen ions in moles per litre.

Alternative and more useful forms of pH definition are :

$$\text{pH} = \log \frac{1}{[\text{H}^+]}$$

and

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

The pH of Solutions

- The pH concept is very convenient for expressing hydrogen ion concentration.
- It was introduced by Sorensen in 1909.
- It is now used as a general way of expressing other quantities also

For example:

- Concentration of OH^- ions in aqueous solution of a base is expressed as
$$\text{pOH} = -\log [\text{OH}^-]$$
- Equilibrium constant for water is written as
$$\text{pK}_w = -\log [\text{K}_w]$$
- For any quantity X , we can write $\text{pX} = -\log X$
- The "p" in these expression means "- log of the quantity"

Auto-ionisation of water

- Water is an amphoteric substance.
- It can behave either as an acid or a base.
- One molecule of water transfers a proton to another molecule.
- There results a hydronium ion (H_3O^+) and a hydroxyl ion (OH^-).
- In this reaction one molecule of water acts as a Bronsted acid and the other as a Bronsted base.
- The above reaction in which water molecules interact to produce a hydronium ion and a OH^- ion is called auto-ionisation of water.

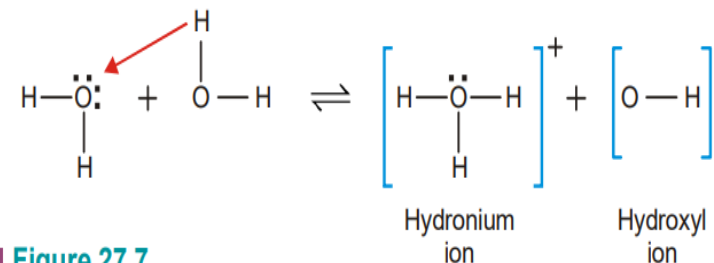


Figure 27.7
Auto-ionisation of water.

- It may be written as



Auto-ionisation of water

- That auto-ionisation actually occurs was proved by Friedrich Kohlrausch (1840-1910).
- He found that even the purest water conducts electricity to a very small extent which was due to the generation of H_3O^+ and OH^- ions in water by ionisation.
- Water auto-ionisation is most fundamental to our study of acids and bases.
- Obviously, H^+ ions are associated with water to give hydronium ions (H_3O^+).
- But for simplicity we shall generally write the dissociation equilibrium of water as $\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$

pH Scale

In order to express the hydrogen ion concentration or acidity of a solution, a pH scale was evolved. The pH is defined as

$$\text{pH} = -\log [\text{H}^+] \quad \text{or} \quad [\text{H}^+] = 10^{-\text{pH}}$$

The hydrogen ion concentration of different acidic solutions were determined experimentally. These were converted to pH values using the above relations. Then these pH values were computed on a scale taking water as the reference substance. **The scale on which pH values are computed is called the pH scale.**

Water dissociates to H^+ and OH^- ions to a very small degree so that we have the equilibrium.



We can write the equilibrium expression as

$$K = \frac{[\text{H}^+][\text{OH}^-]}{[\text{H}_2\text{O}]} \quad \dots(2)$$

Since water is so little dissociated, the concentration of undissociated molecules, $[\text{H}_2\text{O}]$, is presumed to be constant. Therefore from expression (2) we can write

$$[\text{H}^+][\text{OH}^-] = K [\text{H}_2\text{O}] = K_w = \text{a constant.} \quad \dots(3)$$

where K_w is called the **water dissociation constant or the water ionisation constant.**

When the concentrations of H^+ and OH^- ions in water are expressed in mole per litre, the value of

K_w found experimentally is 1.0×10^{-14} .

pH Scale

K_w found experimentally is 1.0×10^{-14} . From the equation (1) it is obvious that one molecule of water dissociates to give one H^+ ion and one OH^- ion. This means that the concentration of H^+ and OH^- ions in pure water is equal. Using the expression (3), we have

$$[H^+][OH^-] = \sqrt{K_w} = \sqrt{1.0 \times 10^{-14}}$$

or $[H^+] = [OH^-] = 1.0 \times 10^{-7} \text{ mol/l}$

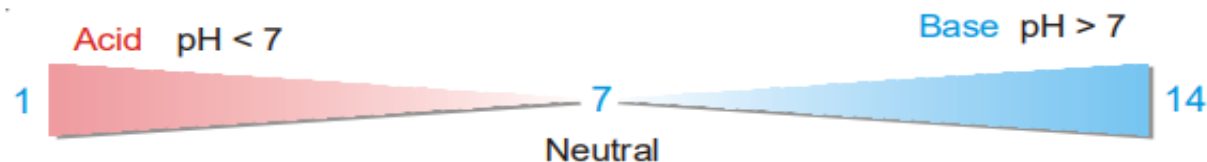
Thus the H^+ ion and OH^- ion concentrations in pure water are both $10^{-7} \text{ mol l}^{-1}$ at 25°C and it is said to be neutral. In acidic solution, however, the concentration of H^+ ions must be greater than $10^{-7} \text{ mol l}^{-1}$. Similarly in a basic solution, the concentration of OH^- ions must be greater than $10^{-7} \text{ mol l}^{-1}$. Thus we can state :

neutral solution $[H^+] = [OH^-]$

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

basic solution $[H^+] < [OH^-]$

Expressing the $[H^+]$ in terms of pH for the different solutions cited above, we get what we call the **pH scale**. On this scale (Fig. 27.9) the values range from 0 to 14. Since pH is defined as $-\log [H^+]$ and the hydrogen ion concentration of water is 10^{-7} , the pH of water is 7. All solutions having pH less than 7 are acidic and those with pH greater than 7 are basic.



As shown by the pH scale, pH decreases with the increase of $[H^+]$. The lower the pH, higher is the $[H^+]$ or acidity.

pH scale

Acid pH low + H^+ High

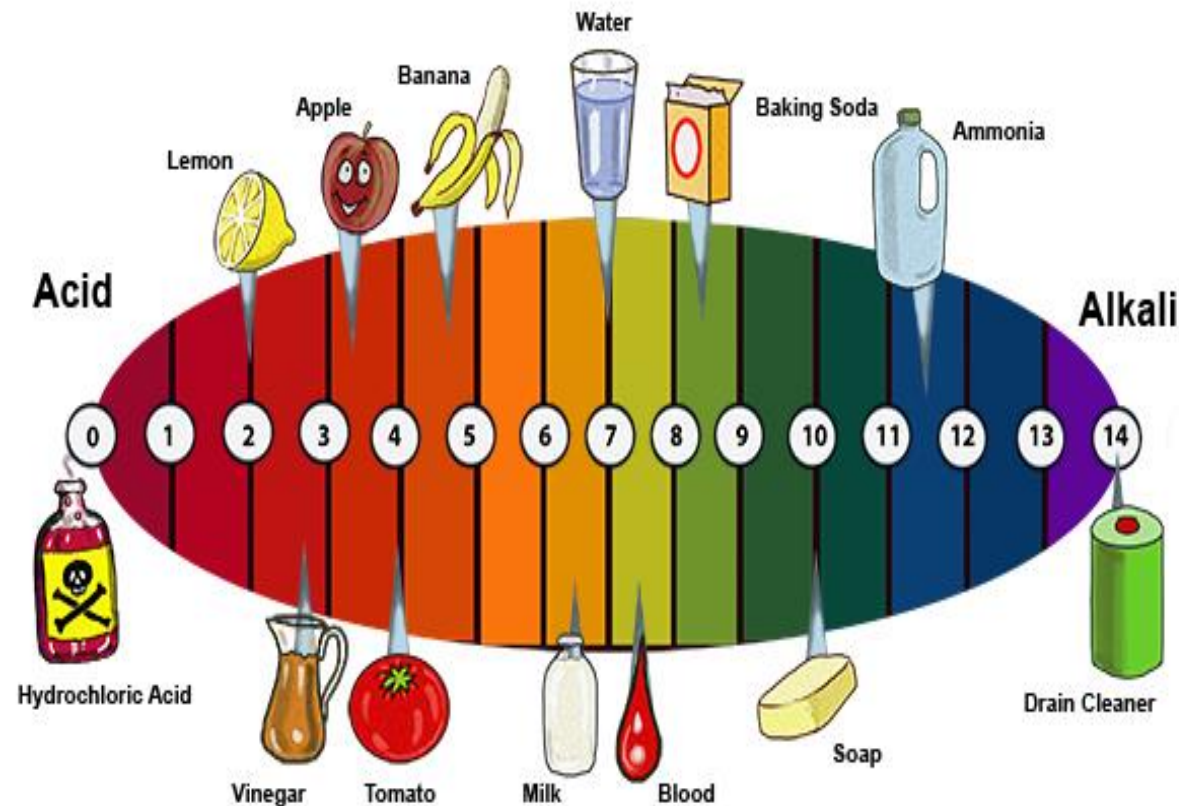
- The pH scale, (0 - 14), is the full set of pH numbers which indicate the concentration of H^+ and OH^- ions in water.

pH Scale Principle

- H^+ ion concentration and pH relate inversely
- OH^- ion concentration and pH relate directly.

- a. Increasing pH means the H^+ ions are decreasing.
- b. Decreasing pH means H^+ ions are increasing.
- c. Increasing pH means OH^- ions are.....?
- d. Decreasing pH means OH^- ions are.....?

The pH Scale



Relation between pH and pOH

Let us consider the log form of the expression

$$K_w = [\text{H}^+][\text{OH}^-]$$

That is, $\log K_w = \log [\text{H}^+] + \log [\text{OH}^-]$

or $-\log K_w = -\log [\text{H}^+] - \log [\text{OH}^-]$

Thus $\text{p}K_w = \text{pH} + \text{pOH}$

Since $K_w = 1.0 \times 10^{-14}$

$$\text{p}K_w = -\log (1.0 \times 10^{-14}) = 14.00$$

Hence, for any aqueous solution at 25°C, pH and pOH add up to 14.00. That is,

$$\text{pH} + \text{pOH} = 14.00$$

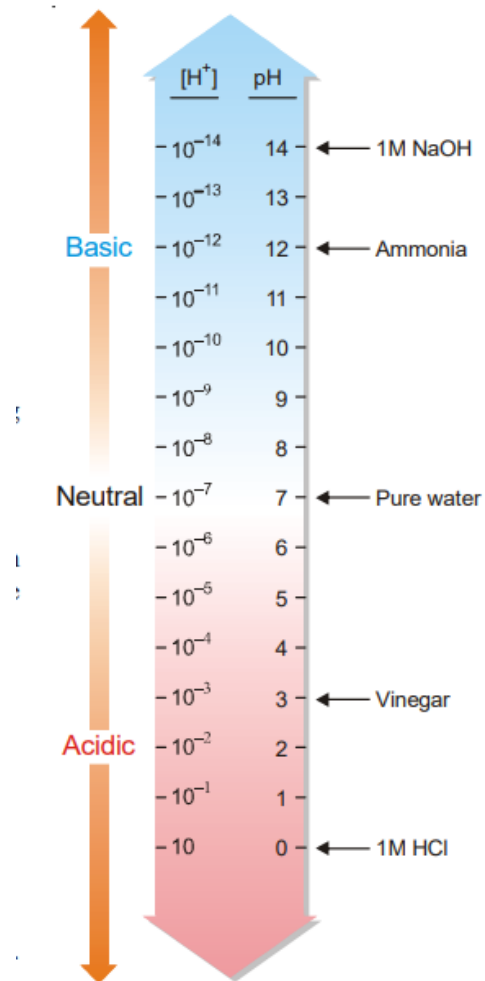


Figure 27.10
pH Scale and pH values of
some common substances.

pH related problems: Do yourself

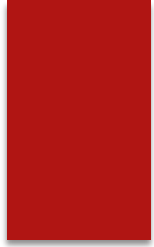
- The hydrogen ion concentration of a fruit juice is 3.3×10^{-2} M. What is the pH of the juice ? Is it acidic or basic ?
- If a solution has a pH of 7.41, determine its H^+ concentration.
- Determine the pH of 0.10 M NaOH solution
- Calculate the pH of a 0.020 M $Ba(OH)_2$ solution.
- The pH of a solution of HCl is 2. Find out the amount of acid present in a litre of the solution
- If a solution has a pH of 5.50 at $25^\circ C$, calculate its $[OH^-]$.

pKa & Henderson-Hasselbalch equation

- The pKa expresses, on a logarithmic scale, the relative strength of a weak acid or base

Henderson-Hasselbalch equation

- This equation shows how the pH could be estimated using the pKa.
- pKa is a measure of acid strength. It depends on the identity and chemical properties of the acid. pH is a measure of $[H^+]$ in a solution.
- For acids, the smaller the pKa, the more acidic the substance is (the more easily a proton is lost, thus the lower the pH).
- The stronger the acid, the lower its pKa; the stronger the base, the higher its pKa. The pKa can be determined experimentally; it is the pH at the midpoint of the titration curve for the acid or base.



Henderson-Hasselbalch equation:

- can be used to determine the **relationship between buffer pH and pKa**.
- is **important for understanding buffer action and acid-base balance** in the blood and tissues of vertebrates.
- simply a **useful way of restating the expression for the dissociation constant** of an acid.
- For the dissociation of a weak acid HA into H and A, the Henderson- Hasselbalch equation can be derived as follows:

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

First solve for $[H^+]$:

$$[H^+] = K_a \frac{[HA]}{[A^-]}$$

Then take the negative logarithm of both sides:

$$-\log [H^+] = -\log K_a - \log \frac{[HA]}{[A^-]}$$

Substitute pH for $-\log [H^+]$ and pK_a for $-\log K_a$:

$$pH = pK_a - \log \frac{[HA]}{[A^-]}$$

Now invert $-\log [HA]/[A^-]$, which involves changing its sign, to obtain the Henderson-Hasselbalch equation:

$$pH = pK_a + \log \frac{[A^-]}{[HA]} \quad \text{🧊 (2-9)}$$

Stated more generally,

$$pH = pK_a + \log \frac{[\text{proton acceptor}]}{[\text{proton donor}]}$$

- An **equivalence point** is when the moles of a standard solution (titrant) equal the moles of a solution of unknown concentration (analyte).
- At half the equivalence point: $pH = pK_a$

$$\text{pH} = \text{p}K_a + \log \frac{[\text{conjugate base}]}{[\text{acid}]}$$

Solving Problems Using the Henderson-Hasselbalch Equation

1. Calculate the $\text{p}K_a$ of lactic acid, given that when the concentration of lactic acid is 0.010 M and the concentration of lactate is 0.087 M, the pH is 4.80.

$$\text{pH} = \text{p}K_a + \log \frac{[\text{lactate}]}{[\text{lactic acid}]}$$

$$\text{p}K_a = \text{pH} - \log \frac{[\text{lactate}]}{[\text{lactic acid}]}$$

$$= 4.80 - \log \frac{0.087}{0.010} = 4.80 - \log 8.7$$

$$= 4.80 - 0.94 = 3.9 \quad (\text{answer})$$

2. Calculate the pH of a mixture of 0.10 M acetic acid and 0.20 M sodium acetate. The $\text{p}K_a$ of acetic acid is 4.76.

$$\text{pH} = \text{p}K_a + \log \frac{[\text{acetate}]}{[\text{acetic acid}]}$$

$$= 4.76 + \log \frac{0.20}{0.10} = 4.76 + 0.30$$

$$= 5.1 \quad (\text{answer})$$

3. Calculate the ratio of the concentrations of acetate and acetic acid required in a buffer system of pH 5.30.

$$\text{pH} = \text{p}K_a + \log \frac{[\text{acetate}]}{[\text{acetic acid}]}$$

$$\log \frac{[\text{acetate}]}{[\text{acetic acid}]} = \text{pH} - \text{p}K_a$$

$$= 5.30 - 4.76 = 0.54$$

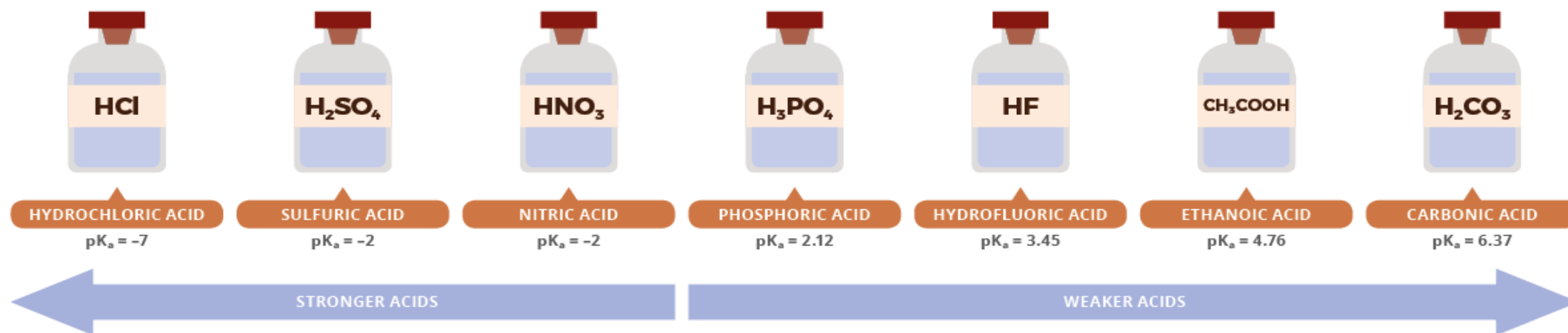
$$\frac{[\text{acetate}]}{[\text{acetic acid}]} = \text{antilog } 0.54 = 3.5 \quad (\text{answer})$$



To see the effect of pH on the degree of ionization of a weak acid, see the Living Graph for Equation 2–9.

A GUIDE TO ACIDS, ACID STRENGTH, AND CONCENTRATION

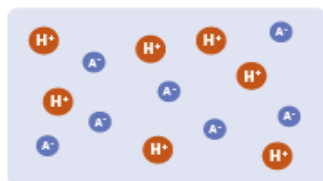
What's the difference between acid strength and concentration? And how does pH fit in with these? This graphic explains the basics.



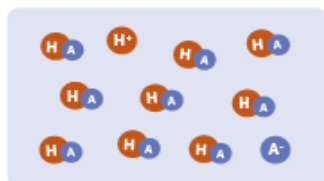
STRONG ACIDS VS. WEAK ACIDS



The H⁺ ion is transferred to a water molecule, forming H₃O⁺



STRONG ACID



WEAK ACID

Hydrogen ions Negative ions Acid molecules

Acids react with water when they are added to it, forming ions. The degree to which they do this is what determines whether they are strong or weak acids. Strong acids are essentially 100% ionised in solution. Weak acids ionise very little in solution.

ACIDS, K_a AND pK_a

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

$$\text{pK}_a = -\log_{10}[K_a]$$

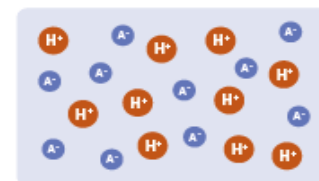
	K _a	pK _a
VERY STRONG ACID	>0.1	<1
FAIRLY STRONG ACID	10 ⁻³ –0.1	1–3
WEAK ACID	10 ⁻⁵ –10 ⁻³	3–5
VERY WEAK ACID	10 ⁻¹⁵ –10 ⁻⁵	5–15
EXTREMELY WEAK ACID	<10 ⁻¹⁵	>15

The acid dissociation constant, K_a, is a measure of the strength of an acid. The higher its value, the stronger the acid (i.e. the more readily it ionises in water). pK_a converts K_a number to a logarithmic scale that makes it easier to compare strengths of different acids.

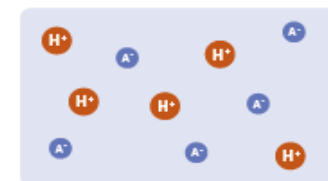
CONCENTRATION AND pH

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

A decrease of one on the pH scale represents a tenfold increase in H⁺ concentration.



CONCENTRATED ACID



DILUTE ACID

Hydrogen ions Negative ions

Concentration is distinct from strength. It refers to the amount of acid in a given solution. A concentrated acid contains a large amount of acid in a given volume; a dilute solution contains a small amount. The pH scale gauges the amount of hydrogen ions in solution.



Buffer solution

- ❖ A buffer solution is one which maintains its pH fairly constant even upon the addition of small amounts of acid or base.
- ❖ We can add a small amount of an acid or base to a buffer solution and the pH will change very little.
- ❖ a weak acid together with a salt of the same acid with a strong base. These are called Acid buffers e.g., $\text{CH}_3\text{COOH} + \text{CH}_3\text{COONa}$.
- ❖ a weak base and its salt with a strong acid. These are called Basic buffers. e.g., $\text{NH}_4\text{OH} + \text{NH}_4\text{Cl}$.
- ❖ **Acid Buffer= Weak acid + Its salt**
- ❖ **Base Buffer= Weak Base + Its salt**



Buffer Capacity

Buffering capacity is the number of grams of strong acid or alkali which is necessary for a change in pH of one unit of one liter of buffer solution.

On the other hand, the ability of the buffer to resist changes in pH when an acid or base is added.

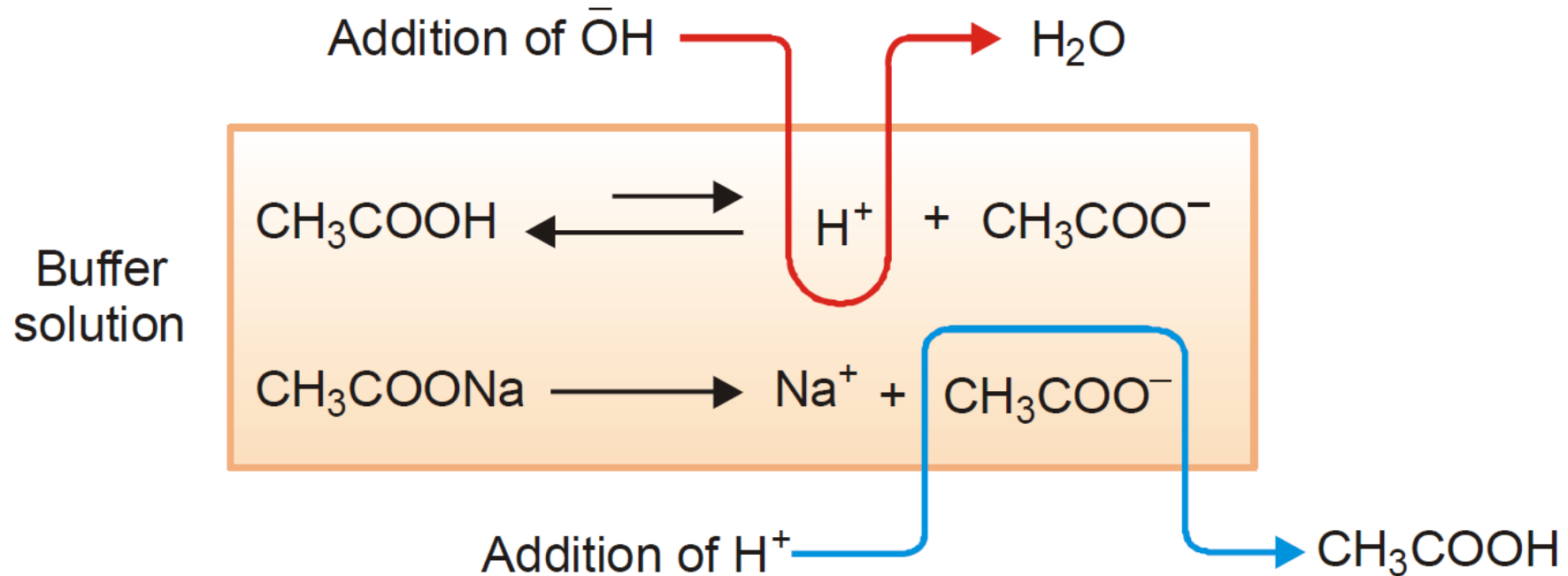


How a buffer operates?

For acid buffer

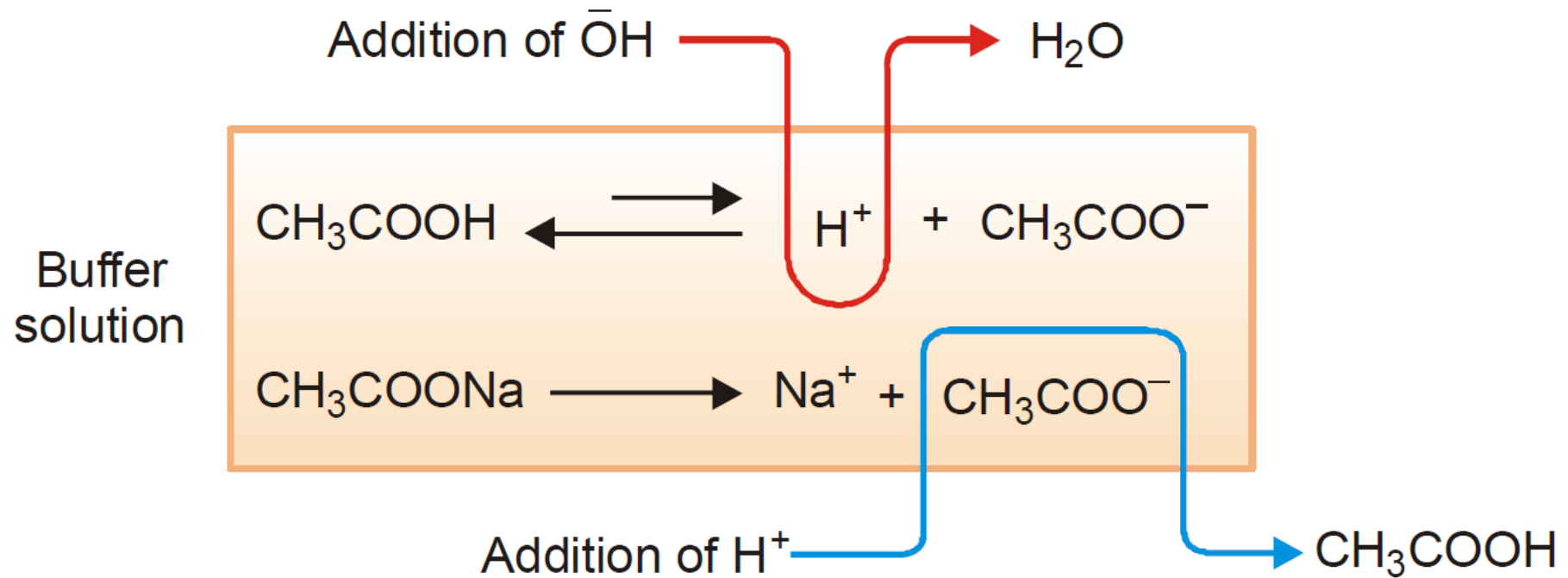
Addition of H^+

H^+	OH^-	Mechanism
1	0	Normal State
2	0	Addition of H^+ (transition state)
1	0	Formation of CH_3COOH



Addition of OH⁻

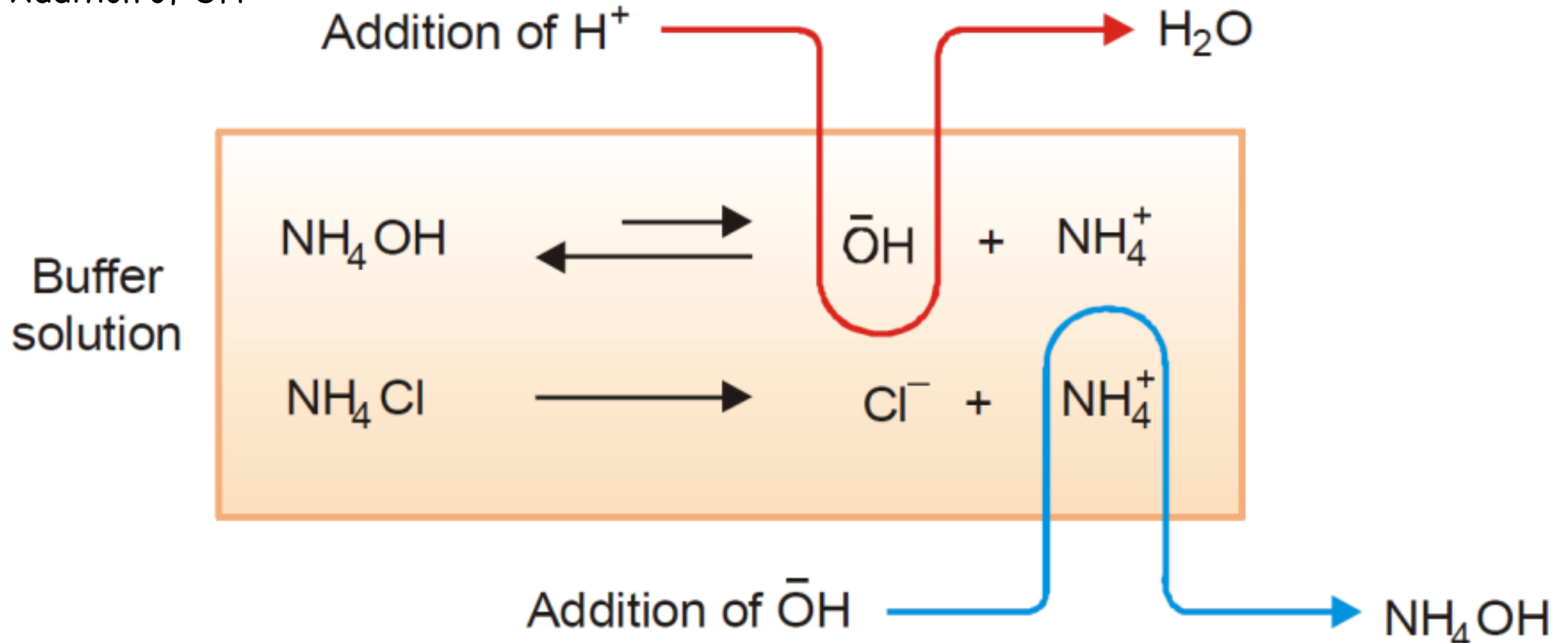
H ⁺	OH ⁻	Mechanism
1	0	Normal State
1	1	Addition of OH ⁻ (transition state)
0	0	Formation of H ₂ O (transition state)
1	0	Formation of H ⁺ & CH ₃ COO ⁻ from old CH ₃ COOH

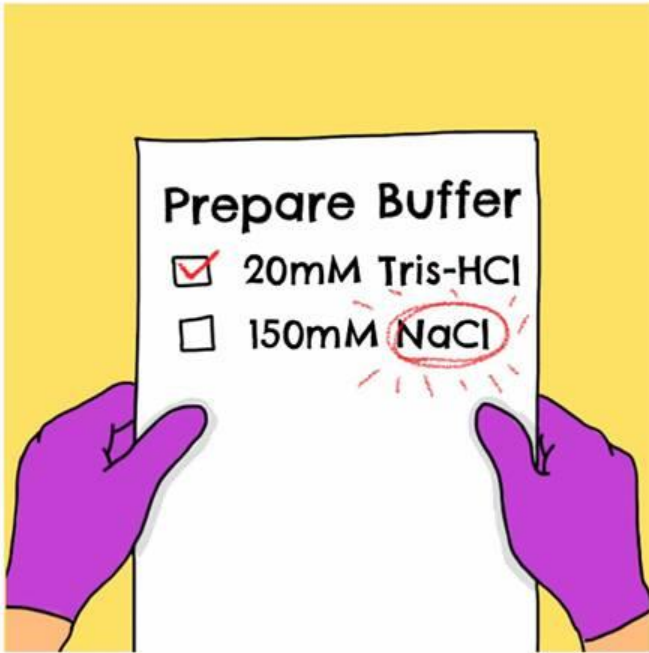


How a buffer operates?

For base buffer

1. Addition of H^+
2. Addition of OH^-





Thank You!