

*pH, pOH, BRONSTED –  
LOWRY AND LEWIS CONCEPT  
OF ACIDS AND BASES*

# pH and its scales

## **pH Scale:**

When water molecule dissociates and attains ionic equilibrium the concentration of both  $[H^+]$  and  $[OH^-]$  ions will be equal. Increase of either  $H^+$  or  $OH^-$  would result in the imbalance of the equation. As such this phenomenon is used to measure the acidic or alkaline condition of the solution. This is referred as pH scale.

## **Definition of pH:**

The pH is a quantity characterizing the activity (or concentration) of hydrogen ions and is numerically equal to the negative common logarithm of the activity or concentration expressed in g-ions/l.

The pH scale of acidity defines pH as the negative common logarithm of the concentration of  $H^+$

$$pH = - \log [H^+]$$

# Example

Dissociation of water:



- ⦿ In non neutral media, the activities  $a_{\text{H}^+}$  and  $a_{\text{OH}^-}$  are not equal to each other.
- ⦿ They are inversely proportional to each other. Thus on adding the acid to water we increase the concentration of hydrogen ions and therefore augment the value of  $a_{\text{H}^+}$  but this will accelerate the opposite direction of reaction.
- ⦿ Some of the added hydrogen ions bind an equivalent amount of  $\text{OH}^-$  to form  $\text{H}_2\text{O}$  and a reduction in  $a_{\text{OH}^-}$  occurs.
- ⦿ Equilibrium is established again when the product of the ion activities once more acquires the value it had prior to the addition of the acid.
- ⦿ Hence any increase in hydrogen ion concentration causes a corresponding decrease in hydroxyl ion concentration and vice versa.

# pOH

## Definition:

pOH is defined as the negative common logarithm of the concentration of hydroxide ion ( $\text{OH}^-$ )

$$\text{pOH} = -\log [\text{OH}^-]$$

**In the pOH scale, 7 is neutral, less than 7 is basic and greater than 7 is acidic.**

Relationship between pH and pOH,

$$\text{pH} + \text{pOH} = 14$$

This formula allows to readily convert the values of pH and pOH.

Because  $K_w$  is constant, the product of  $[\text{H}^+]$  and  $[\text{OH}^-]$  is always equal to  $10^{-14}$ .

## pKa and pKb

The pH and pOH as values are used to describe the acidic or basic strength of solutions.

Whereas pKa ( $-\log K_a$ ) and pKb ( $-\log K_b$ ) are used as a measure of acidity and basicity, respectively, for standard acids and bases.

In **Strong acid, pKa value is less than zero**. Because it almost completely dissociates in water, which gives an aqueous solution a relatively low pH.

In **Weak acid, pKa value is greater than zero**. Because they only partially dissociate in water, to make solutions with larger pH's than those strong acids produce at the same concentration.

Similarly,

In **strong base, pKb value is less than zero**

In **weak base, pKb value is greater than zero**.

Note: The prefix "p" in front of a symbol means "take the negative log".

# Lewis concept of acids and bases

Lewis concept mainly depends on lone pair or electron pair.

- According to this concept an acid is any species which is capable of accepting a pair of electrons. Lewis an acid may also be a species without any proton.

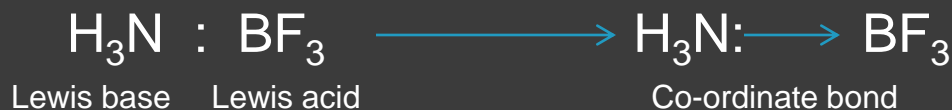
Eg-  $\text{BF}_3$ ,  $\text{AlCl}_3$ ,  $\text{ZnCl}_2$ ,  $\text{FeCl}_3$ ,  $\text{SO}_3$  etc.

- a base is a species having lone pair of electrons for donation.

- Types of lewis acid:**

1. Compounds whose central atom has an incomplete octet. In other words Lewis acids of this type are the electro deficient molecules.

Eg.  $\text{BF}_3$ ,  $\text{AlCl}_3$ ,  $\text{GaCl}_3$  etc



Example of Lewis acid:

1. Central atom has incomplete octate

E.-  $\text{BF}_3$ ,  $\text{BCl}_3$ ,  $\text{AlCl}_3$ ,  $\text{MgCl}_2$ ,  $\text{LiCl}$



Boron has 6 electrons, but only three places are occupied.  
That's why it has incomplete octate and known as lewis acid.

2. Simple cation

Eg.-  $\text{H}^+$ ,  $\text{Na}^+$ ,  $\text{Li}^+$ ,  $\text{Mg}^{2+}$ ,  $\text{Al}^{3+}$

But conjugated cation like  $\text{NH}_4^+$ ,  $\text{PH}_4^+$  are not in this category

⦿ 2. Compounds in which the central atom has empty d-orbital



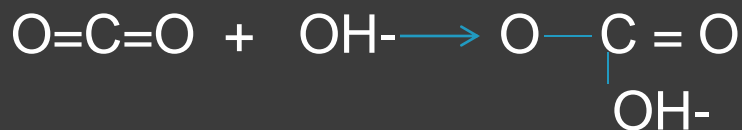
Si = 14 electron



Like this,  $\text{SF}_4$ ,  $\text{PF}_3$ ,  $\text{PF}_5$ ,  $\text{SeF}_4$  etc

3. Central atom formed multiple bond with dissimilar atom

Eg.-  $\text{CO}_2$ ,  $\text{SO}_2$



$\text{HCO}_3^-$  (Bicarbonate)

### Lewis bases

- ⦿ All the simple negative ions such as  $\text{Cl}^-$ ,  $\text{F}^-$  etc.
- ⦿ Molecules with one or two unshared pair of electrons ex:  $\text{H}_2\text{O}$ ,  $\text{NH}_3$ , Pyridine etc.
- ⦿ Multiple bonded compounds which form coordination compounds with transition metals Eg.  $\text{CO}$ ,  $\text{NO}$ , Ethylene, Acetylene.



## ***Bronsted - Lowry Concept for acid and bases***

- Bronsted - Lowry defined **an acid, as a substance** which **gives up proton** (proton donor)
  - and **a base as a substance** which **accepts proton** (Proton acceptor).
  - Example: The species B accepts a proton and thus behaves as a base while the species  $BH^+$  gives up a proton and thus acts as an acid
  - Such related pairs of an acid and a base is said to **conjugate to one another**
  - In the above example,  
    B is the conjugate base of the acid  $BH^+$  and  $BH^+$  is the conjugate acid of the base B.
- On the whole we can say that
- every base has its conjugate acid** and **every acid has a conjugate base**.
- Hence, **strong acid has a weak conjugate base** and **a weak acid has a strong conjugate base**

# Examples of conjugate acids and bases

Acid	Base		Conjugate acid	Conjugate base
$\text{H}_2\text{SO}_4$	$+\text{H}_2\text{O}$	$\longrightarrow$	$\text{H}_3\text{O}^+$	$\text{HSO}_4^-$
$\text{H}_2\text{O}$	$+\text{NH}_3$	$\longrightarrow$	$\text{NH}_4^+$	$\text{OH}^-$
$\text{HCl}$	$+\text{OH}^-$	$\longrightarrow$	$\text{H}_2\text{O}$	$\text{Cl}^-$
$\text{H}_3\text{O}^+$	$+\text{OH}^-$	$\longrightarrow$	$\text{H}_2\text{O}$	$\text{H}_2\text{O}$
$\text{NH}_3^+$	$+\text{H}_2\text{O}$	$\longrightarrow$	$\text{H}_3\text{O}^+$	$\text{NH}_3$

- Thus according to Bronsted -Lowry concept neutralization is a process in which a proton is transferred from an acid to a base.
- A substance acts as an acid only when another substance capable of accepting a proton (i.e Base) is present.
- For eg.  $\text{HCl}$  or  $\text{CH}_3\text{COOH}$  solution in benzene is not acidic because benzene is not in a position to take up protons.
- On the other hand  $\text{HCl}$  or  $\text{CH}_3\text{COOH}$  solutions in water are definitely acidic because  $\text{HCl}$  or  $\text{CH}_3\text{COOH}$  can ionize in water.

Water the most common solvent can act both as an acid as well as a base because it can give up a proton as well as can receive a proton. Such compounds as water can function both as an acid and base is are known as amphoteric or amphiprotic. Its dual behaviour may be represented by the following equation

