ACIDS AND BASES

Topic-1
Lowry - Bronsted and Lewis theory of acids and bases with examples and applications

VERY SHORT ANSWER QUESTIONS

1. What is bronsted acid and base give one example?
Ans:

Strength of bronsted acids and bronsted bases will depend on the ability to donate or to accept proton.

Strong acids: which have more ability to donate protons.
Eg: $HClO_4$, HCl, H_2SO_4

Weak acids: which have fewer tendencies to donate proton.
Eg: HF, CH_3COOH, H_2CO_3, HCN

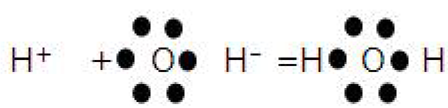
Strong bases: which have greater tendency to accept protons
Eg: OH^- CH_3COO^- CH_3^- etc.

Weak bases → which have fewer tendencies to accept the proton.
Eg: Cl^- ClO_4^- Br^- I^- NH_3…..etc

2. What is a Lewis acid and base?
Ans:

According to this concept, a base is defined as a substance which can furnish a pair of electrons to form a coordinate bond whereas an acid is a substance which can accept a pair of electrons. The acid is also known as electron acceptor or electrophile while the base is electron donor or nucleophile.

A simple example of an acid-base is the reaction of a proton with hydroxyl ion.



Acid **Base**

$H_3N: + BF_3 = H_3N \rightarrow BF_3$

Base **Acid**

SHORT ANSWER QUESTIONS

1. What is conjugate acid-base pair?

Ans:

Conjugate acid – base pairs:

The acid base pair which differs by a proton is called conjugated acid base pair.

Acid $-$ $H^{\oplus}$ = conjugate base

Base + $H^{\oplus}$ = conjugate acid

Eg :

Acid	Base
HCl	Cl^-
H_2SO_4	HSO_4^-
HSO_4^-	SO_4^{2-}
H_2S	HS^-
C_2H_2	$CH \equiv C^-$

2. Give two examples of Lewis acid and Lewis base?

Ans: Lewis theory of acids – bases:

Acid: which accepts electron pair.

All electrophiles are Lewis acids

Lewis acid must contain suitable vacant orbitals.

Eg: BF_3, BCl_3, $AlCl_3$, H^+, Ag^+ etc.

Base: which donates e^- pair.

All nucleophiles are Lewis bases

Lewis base must contain 1 or more lone pairs of electrons.

Eg: $\ddot{N}H_3, \dddot{P}H_3, R-\ddot{N}H_2$,

$\ddot{N}H_2-\ddot{N}H_2$, OH^-, Cl^-, $H_2\ddot{O}$

LONG ANSWER QUESTIONS

1. **Explain the concept of Bronsted acids and Bornsted bases. Illustrate the answer with examples?**

Ans:

Bronsted – Lowry theory of Acids and bases: This is also called protonic theory .

Acid: which donates proton

Eg: HCl, H_2SO_4, HNO_3 etc.

Base: which accepts proton

Eg: NH_3, PH_3, N_2H_4, $R-NH_2$ etc.

Bronsted acid could be a neutral molecule, anion, cation.

Neutral molecule $\rightarrow$ HCl, HNO_3, H_2SO_4 etc

Cations $\rightarrow$ H_3O^+, NH_4^+

Anions $\rightarrow$ HSO_4^-, HCO_3^-, $H_2PO_4^-$, $H(PO_4)^{2-}$, $H_2(PO_3)^-$ Bronsted base could be neutral, cation or anion

Neutral $\rightarrow$ NH_3 , PH_3, N_2H_4 etc.

anions $\rightarrow$ CNBrIOHetc.

Cations $\rightarrow$ $[Al(H_2O)_5OH)]^{2+}$, $[Fe(H_2O)_5OH$

- **<u>Merits of Bronsted – Lowery theory</u> :**

1) It could explain acid base behaviour in non – aqueous solvents also.
   Eg: liquid SO_2, liquid NH_3
2) It could explain the basic nature of substances like NH_3.
3) It could explain the acidic nature of HCl gas.
4) It is more generalized than Arrhenius theory.

- **<u>Limitations of Bronsted – Lowery theory</u> :**

1) This theory explains behaviour of acids only when there is a base (or) It explains behaviour of base only when there is a acid. (Acid base pairs must be present)
2) It fails to explain the acidic nature of electron deficient compound like BF_3, $AlCl_3$ etc.

2. **What are conjugate acids and conjugate bases? Give two examples**?

Ans:

<u>Conjugate acid – base pairs:</u> The acid base pair which differs by a proton is called conjugated acid base pair.

Acid̃ $H^{\oplus}$ = conjugate base

Base + $H^{\oplus}$ = conjugate acid

Eg :	**Acid**	**Base**
	HCl	Cl

H_2SO_4 HSO_4^-

HSO_4^- SO_4^{2-}

H_2S HS^-

In any conjugate acid base pair, if the acid is stronger, the base is weak and if base is stronger, the acid is weak.

Eg:

1. HCl is strong acid, Cl^- is weak conjugate base.
2. CH_3COOH is weak acid CH_3COO^- is strong base
3. NH_3 is weak base, NH_4^+ strong acid
4. $HClO_4$ is strongest acid, ClO_4^- weak base

Strength of bronsted acids and bronsted bases will depend on the ability to donate or to accept proton.

Strong acids: which have more ability to donate protons

Eg: $HClO_4$, HCl, H_2SO_4

Weak acids: which have fewer tendencies to donate proton.

Eg: HF, CH_3COOH, H_2CO_3, HCN

Among Hydracids HCN is the weakest acid

Strong bases: which have greater tendency to accept protons

Eg: OH^- CH_3COO^- CH_3^- etc.

Weak bases $\rightarrow$ which have fewer tendencies to accept the proton.

Eg: Cl^- ClO_4^- Br^- I^- NH_3…..etc.

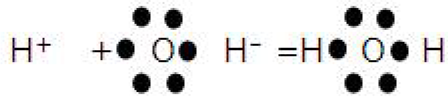
3. **Explain Lewis Acid-Base Theory?**

Ans:

<u>Lewis concept of Acids and bases :</u>

This concept was proposed by G.N. Lewis, in 1939. According to this concept, a base is defined as a substance which can furnish a pair of electrons to form a coordinate bond whereas an acid is a substance which can accept a pair of electrons. The acid is also known as electron acceptor or electrophile while the base is electron donor or nucleophile.

A simple example of an acid-base is the reaction of a proton with hydroxyl ion.



Acid Base

Some other examples are:

$H_3N: + BF_3 = H_3N \rightarrow BF_3$

Base Acid

$H^+ +: NH_3 = [H \leftarrow NH_3]^+$

Acid Base

$BF_3 + [F]^- = [F \rightarrow BF_3]^+$

Acid Base

Lewis concept is more general than the Bronsted Lowry concept.

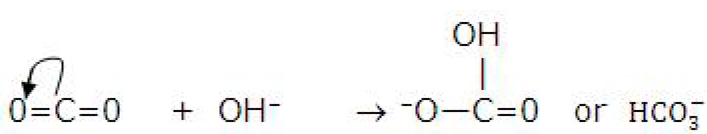
According to **Lewis concept**, the following species can act as **Lewis acids**.

(i) Molecules in which the central atom has incomplete octet: All compounds having central atom with less than 8 electrons are **Lewis acids**, e.g., BF_3, $BC1_3$, $A1C1_3$, $MgCl_2$. BeCL. etc.

(ii) Simple cations: All cations are expected to act as **Lewis acids** since they are deficient in electrons. However, cations such as Na^+, K^+, Ca^{2+}, etc., have a very little tendency to accept electrons, while the cations like H^+, Ag^+, etc., have greater tendency to accept electrons and, therefore, act as **Lewis acids**.

(iii) Molecules in which the central atom has empty d-orbitals: The central atom of the halides such as SiX_4, GeX_4, $TiCl_4$, SnX_4, PX_3, PF_5, SF_4, SeF_4, $TeCl_4$, etc., have vacant d-orbitals. These can, therefore, accept an electron pair and act as **Lewis acids**.

(iv) Molecules having a multiple bond between atoms of dissimilar electronegativity: Typical examples of molecules falling in this class of **Lewis acids** are $C0_2$, $S0_2$ and $S0_3$. Under the influence of attacking Lewis base, one -electron pair will be shifted towards the more negative atom.



Lewis acid Lewis base

The following species can act as **Lewis bases**.

(i) Neutral species having at least one lone pair of electrons: For example, ammonia, amines, alcohols, etc., act as **Lewis bases** because they contain a pair of electrons

$$:NH_3,\quad -\ddot{N}H_2,\quad R-\underset{\cdot\cdot}{\ddot{O}}-H$$

(ii) Negatively charged species or anions: For example, chloride, cyanide, hydroxide ions, etc., act as **Lewis bases**.

CN^-, Cl^-, OH^-

It may be noted that all Bronsted bases are also **Lewis bases** but all Bronsted acids are not **Lewis acids**.

Limitations:

Since the strength of the **Lewis acids and bases** is found to depend on the type of reaction, it is not possible to arrange them in any order of their relative strength.

The choice of which definition of acids and bases one wishes to use in a particular instance depends largely on the sort of chemistry that is studied. But Arrhenius concept is perfectly satisfactory and simplest for dealing with reactions in aqueous solutions. It explains satisfactorily the strength of acids and bases in aqueous solutions, neutralisation, salt hydrolysis, etc.