

CONCEPT MAP

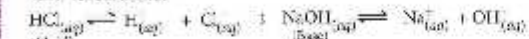
CONCEPTS OF ACIDS AND BASES

Number of concepts have been put forward to characterise acids and bases like Faraday considered them as electrolytes, Liebig termed acids as compounds that can replace its hydrogen with metals. Though the three most important are Arrhenius, Bronsted-Lowry and Lewis concepts.

Class XI

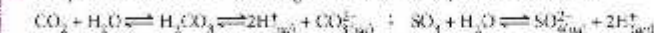
Arrhenius Concept

- A substance which furnishes H^+ ions when dissolved in water is an *acid* while the substance which furnishes OH^- ions is a *base*.

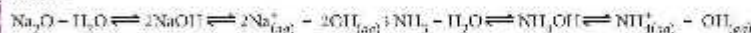


- Acids and bases that dissociate completely in aqueous solutions are strong acids (e.g., HCl, H_2SO_4) and strong bases (e.g., NaOH, KOH) while that dissociate to limited extent are weak acids (e.g., CH_3COOH , H_2CO_3) and weak bases (e.g., NH_4OH , $Al(OH)_3$).

- Many oxides of non-metals in water give H^+ ions thus, they are acidic in nature.



- Similarly, many oxides of metals and other compounds in water give OH^- ions thus, they are basic in nature.



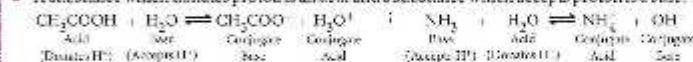
Svante Arrhenius

Limitations

- Applicable only to aqueous solutions.
- It cannot explain the acidic nature of certain salts e.g., $AlCl_3$, BF_3 , etc.

Bronsted-Lowry Concept

- A substance which donates proton is an *acid* and a substance which accepts proton is a *base*.



- Water is an *amphoteric* substance i.e., it can donate as well as accept the proton.

- If Bronsted acid is strong, then its conjugate base is weak and vice versa.

- The strength of acids (or bases) also depends upon the medium acting as base (or acid), e.g., HCl is neutral in benzene, since benzene does not accept proton.

- All Arrhenius acids are also Bronsted acids but all Arrhenius bases are not Bronsted bases.



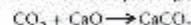
Johannes Bronsted



Thomas Martin Lowry

Limitations

- It cannot explain the acidic or basic behaviour of substances where no proton transfer takes place e.g.,



Lewis Concept

- A substance which accepts electron pair is an *acid* while the substance which donates electron pair is a *base*.

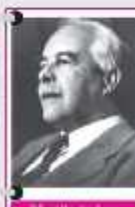


- Species that can act as Lewis acids are:

- Molecules with incomplete octet of central atom e.g., BF_3 , $AlCl_3$, $BeCl_2$, etc.
- Molecules with central atom having empty *d*-orbitals e.g., SiF_4 , $TiCl_4$, SnF_4 , $SnCl_4$, etc.
- All cations e.g., H^+ , Ag^+ , Cu^{2+} , etc.
- Molecules with a multiple bond between atoms of dissimilar electronegativity e.g., CO_2 , SO_2 , SO_3 , etc.

- Species that can act as Lewis bases are:

- Neutral molecules having at least one lone pair of electrons e.g., $\overset{-}{N}H_3$, R_3O^- , $\overset{-}{S}NH_2$, etc.
- All anions e.g., CN^- , F^- , OH^- , etc.



Gilbert Lewis

Limitations

- This concept is so general that any reaction with the formation of coordinate bond is considered to be acid-base reaction however, it is not always true.
- It explains that formation of coordinate bond is necessary for acid base reaction but no such bond is formed between HCl and NaOH which is an acid-base reaction.
- It does not predict the relative strength of acids and bases.
- It cannot explain the catalytic activity of acids which is due to H^+ ions. As for Lewis acid it is not necessary to give H^+ ions, many Lewis acids will not have this property.

CONCEPT MAP CLASS XI

Titrimetric / Volumetric Analysis

Titrimetric analysis has a wide applications in the various fields. It is used to determine the acid content of alcoholic beverages, calcium content in dairy products, total water hardness, vitamin C in fruit juices, etc.

Titrimetric / Volumetric Analysis

Single Titration

- Acid-base titration:** The concentration of an acid is estimated by adding a solution of standard base. The equivalence point is detected by adding a few drops of suitable indicator to the acid solution.
- Redox titration:** The concentration of oxidant is estimated by adding reductant or vice versa. For example, Fe^{2+} ions can be estimated by titrating against acidified $KMnO_4$ solution.
- Precipitation titration:** Cations and anions combine to form a compound of very low solubility. $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$ (White ppt.)
- Complexometric titration:** Titrate combines with the titrant to form a complex salt. $CaSO_4 + 4NTi_2OH \rightarrow [Ca(NTi_2)_4]SO_4 + 4H_2O$

If a mixture of NaOH & NaHCO₃ is taken

- Equivalents of NaOH = equivalents of HCl required upto phenolphthalein end point.
- Equivalents of $NaHCO_3$ = equivalents of HCl required from phenolphthalein to methyl orange end point.

If a mixture of Na₂CO₃ & NaHCO₃ is taken

- 1/2 Equivalents of Na_2CO_3 ($n = 2$) = equivalents of HCl required upto phenolphthalein end point.
- 1/2 Equivalents of Na_2CO_3 ($n = 2$) + equivalents of $NaHCO_3$ ($n = 1$) initially = Equivalents of HCl required from phenolphthalein end point to methyl orange end point.

Back Titration

Substance A of known concentration is titrated with reagent B (V_1 litre) of known concentration. Equivalents of B reacted with A = $N_1 V_1$. In another flask, reagent C of unknown concentration is reacted with A and excess of A is titrated with B (V_2 litre). Equivalents of B reacted with excess of A = $N_2 V_2$. Equivalents of pure C = $N_1 V_1 - N_2 V_2$. Moles of pure C = $\frac{N_1 V_1 - N_2 V_2}{n}$

Here, *n*-factor
For acids, *n* = basicity of acid
For bases, *n* = acidity of base
In case of redox reactions:
n-factor = change in oxidation number per molecule
= total number of electrons transferred per molecule
Mass of pure C = $\frac{N_1 V_1 - N_2 V_2}{n} \times \text{Molar mass of C}$
% Purity of C = $\frac{N_1 V_1 - N_2 V_2}{N_1 V_1} \times \frac{\text{Molar mass of C}}{\text{Molar mass of C}} \times 100$
where, N_1 = Normality of B,
n = Mass of C.

In the given calculation, we have taken alkalimixture of NaOH and Na_2CO_3 but other alkalimixtures can also be taken.

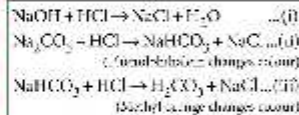
If a mixture of NaOH, Na₂CO₃ & NaHCO₃ is taken

- Equivalents of NaOH = 1/2 equivalents of Na_2CO_3 ($n = 2$) = Equivalents of HCl required upto phenolphthalein end point.
- 1/2 equivalents of Na_2CO_3 ($n = 2$) + equivalents of $NaHCO_3$ ($n = 1$) = Equivalents of HCl required from phenolphthalein end point to methyl orange end point.

Double Titration

The purpose of double titration is to determine the percentage composition of an alkali mixture or an acid mixture.
Process:
• Dissolve mixture of Na_2CO_3 and NaOH in water to make alkaline solution, add two indicators phenolphthalein and methyl orange.
• Now this solution is titrated against HCl.

Reactions Taking Place



Calculations

- At methyl orange end point:**
- Moles of HCl consumed by $NaHCO_3$ = $M_1 V_2$ = Moles of $NaHCO_3$ formed from Na_2CO_3 = Moles of Na_2CO_3 in the mixture.
 - Mass of Na_2CO_3 in the mixture = $M_1 V_2 \times 106$
 - % of Na_2CO_3 in mixture = $\frac{M_1 V_2 \times 106}{M_1 V_1 + M_2 V_2 \times 40} \times 100$

- At phenolphthalein end point:**
- Moles of HCl used in eq. (1) = $M_1 V_1 - M_2 V_2$ = Moles of NaOH
 - Mass of NaOH = $(M_1 V_1 - M_2 V_2) \times 40$
 - % of NaOH = $\frac{(M_1 V_1 - M_2 V_2) \times 40}{M_1 V_1 + M_2 V_2 \times 40} \times 100$
- where, V_1 = volume of HCl used in phenolphthalein changes colour; V_2 = volume of HCl required from phenolphthalein colour change to methyl orange colour change
n = Mass of Na_2CO_3 , NaOH mixture