CHAPTER-II

GENERAL INTRODUCTION

(REVIEW OF THE EARLIER WORKS)

II.1. HOST-GUEST CHEMISTRY

Supramolecular host-guest chemistry emphases on the interaction between molecules rather than within molecules and on the study of intermolecular bond rather than covalent bond [1]. Supramolecular chemistry is often defined as "the chemistry beyond the molecule" or "the chemistry of the noncovalent bond". In host-guest chemistry host molecule encapsulate the guest molecule into the inner cavity of the host (**Figure II.1**). All the compounds exhibit supramolecular properties such as molecular recognition, self-assembly, self-organization, and kinetic and thermodynamic complementarity. A great deal of effort in supramolecular chemistry has been used in attempts to model or mimic biological processes [2]. Enzyme mimicking is one of the most attractive studies in recent years. The principles, perspectives and recent developments in the field of supramolecular chemistry have grown exponentially in the last few decades [3-10].
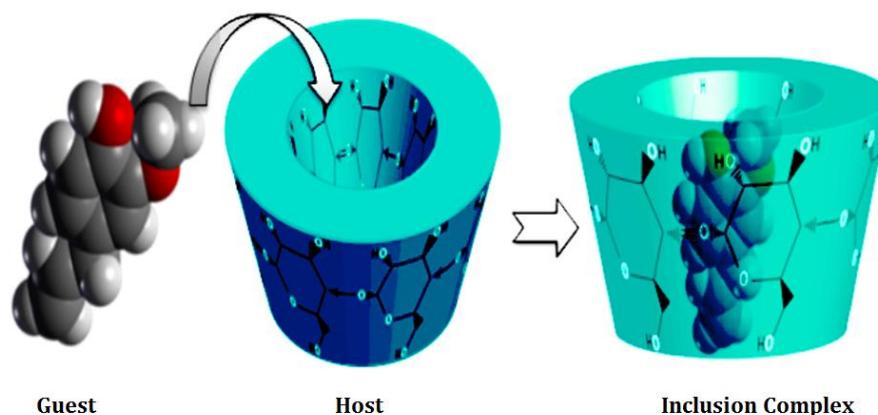

| Guest | Host | Inclusion Complex |

Figure II.1: Schematic representation for host-guest complexation by Cyclodextrin.

The emergence of supramolecular host-guest chemistry has had a profound effect on how efficiently chemists prepare structures using spontaneous secondary interactions such as hydrogen bonding, dipole-dipole, charge transfer, van der Wall's and

π-π stacking interactions (**Figure II.2**)[11-15]. The supramolecular field was and is the basis for most of the essential biochemical process of life. The diagrammatic outline from molecular to supramolecular chemistry is shown in **Figure II.3** [16]. The example of some supramolecular host systems are such as Cyclodextrins, Crown ethers, Calixarenes, Cucurbit etc. with large cavities capable of accommodating small guest molecules and ions [17]. Macrocyclic host molecules are of great importance in inclusion complexes as the cyclized and constrained conformation provide the advantages of molecular selectivity [18]. Here we have discussed mainly about the cyclodextrins and crown ethers (as host) in details.
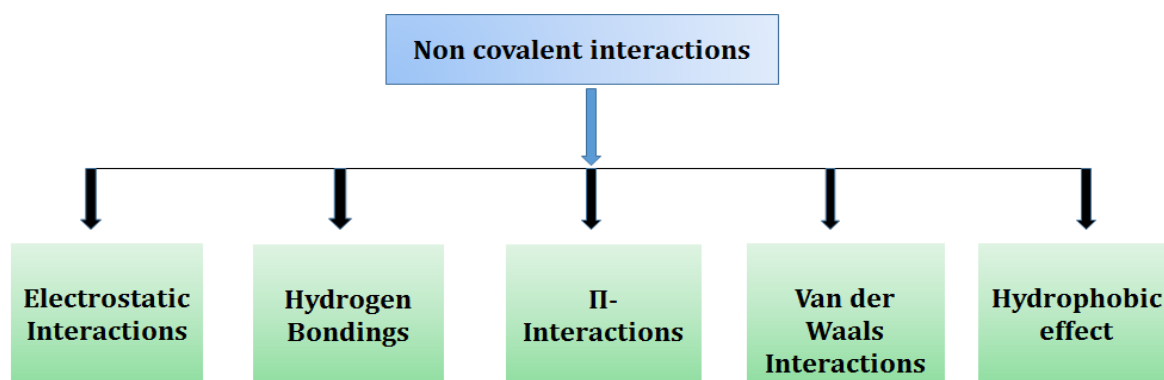

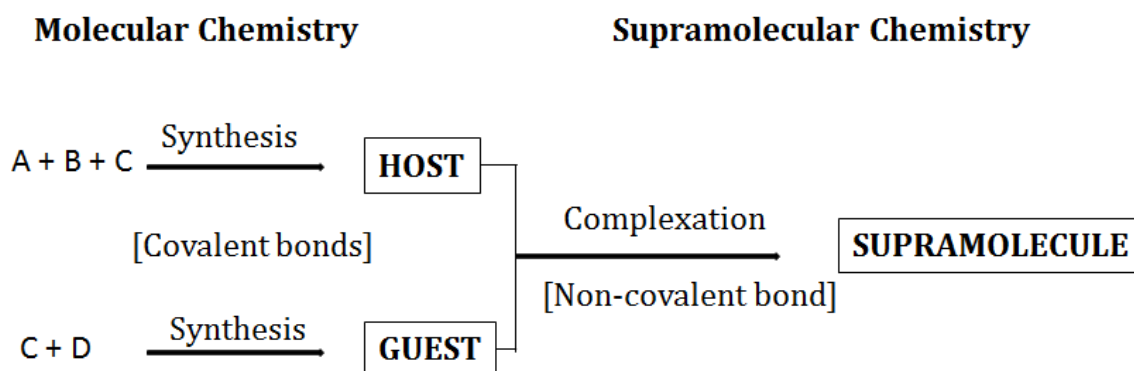**Figure II.2:** Schematic representation of various interactions involved in host-guest chemistry.

Figure II.3: From molecular to supramolecular chemistry.

II.2. CYCLODEXTRINS

Cyclodextrins (CDs) are useful molecular chelating agents. They possess a cage-like supramolecular structure, which is the same as the structures formed from cryptands, calixarenes, cyclophanes, spherands and crown ethers. Compared to all the supramolecular hosts mentioned above, cyclodextrins are most important. Through their inclusion complex forming ability, important properties of the complexed substrates can be modified significantly. As a result of molecular complexation phenomena CDs are widely used in many industrial products, technologies and analytical methods. The negligible cytotoxic effects of CDs are an important attribute in applications such as rug carrier, food and flavours, cosmetics, packing, textiles, separation processes, environment protection, fermentation and catalysis.

II.2.1 Discovery of Cyclodextrin

The oligosaccharide cyclodextrin were first isolated and discovered by Villers in 1891 by enzymatic cleavage of starch [19]. Schardinger confirms his result in 1904 by further identification of cyclic structure of glucose oligomers and enzymes responsible for the synthesis of cyclodextrins [20]. Because of his pioneering work the name Schardinger dextrins was often used for cyclodextrins in early literature. In the early literature, cyclodextrins are also sometimes called cycloamyloses or cyclomaltooligosaccharides. The next major contributor to cyclodextrin chemistry was Karl Freudenberg [21], who developed a method for obtaining of pure α-CD and β-CD and in the process also isolated another crystalline dextrin, which he named γ-CD. It was Freudenberg, who characterized these as cyclic structures composed of α- 1, 4 linked glucose unit [22]. However, the molecular weights of the most common CDs were not determined until much later [23]. The foundations of cyclodextrin chemistry were laid down in the first part of the 20th century. Cramer and French [24] were the first, who recognized and studied possible applications by forming complexes of CDs. Because of their research, the way was opened for the use of cyclodextrins as enzyme models. However, until 1970 only small amounts of cyclodextrins could be produced and high production costs prevented their wide spread usage. Today the cost of cyclodextrins is

reduced dramatically and as a result, the group of scientists interested in cyclodextrins is continually growing.

II.2.2 Structural features and Properties of cyclodextrins

Cyclodextrin (CD) are cyclic oligosaccharides consisting of glucopyranosyl units linked by α-(1,4) bonds. The widely used natural cyclodextrins are α-, β- and γ-cyclodextrin consisting of 6, 7 and 8 glucopyranose units, respectively (**Figure II.4**). Owing to lack of free rotation about the bonds connecting the glucopyranose units, the cyclodextrins are not perfectly cylindrical molecules but the torroidal or cone shaped with a hydrophobic cavity and a hydrophilic surface [**Figure II.5**] [25]. Based on this architecture the primary hydroxyl groups are located on the narrow side of the cone shape, while the secondary hydroxyl groups are located on the wider edge. The primary and secondary hydroxyls on the outside of the cyclodextrins make cyclodextrins water-soluble. Cyclodextrins are insoluble in most organic solvents. The hydrophobic cavity of the cyclodextrin is capable of trapping the hydrophobic parts present in the molecules [26-28]**.** From the crystal structure, it is shown that they have a truncated cone shape with a height of approximately 8Å and an inner cavity diameter varying between ≈5 (α-CD) and 8.30 nm (γ-CD). **Figure II.5** indicates the general structure of the native cyclodextrins. **Table II.1** indicate the common parameters reported by Szejtli [2]. Each D-glucose unit in the cyclodextrin structure contains five chiral carbon atoms and which make CD as a chiral macrocyle.
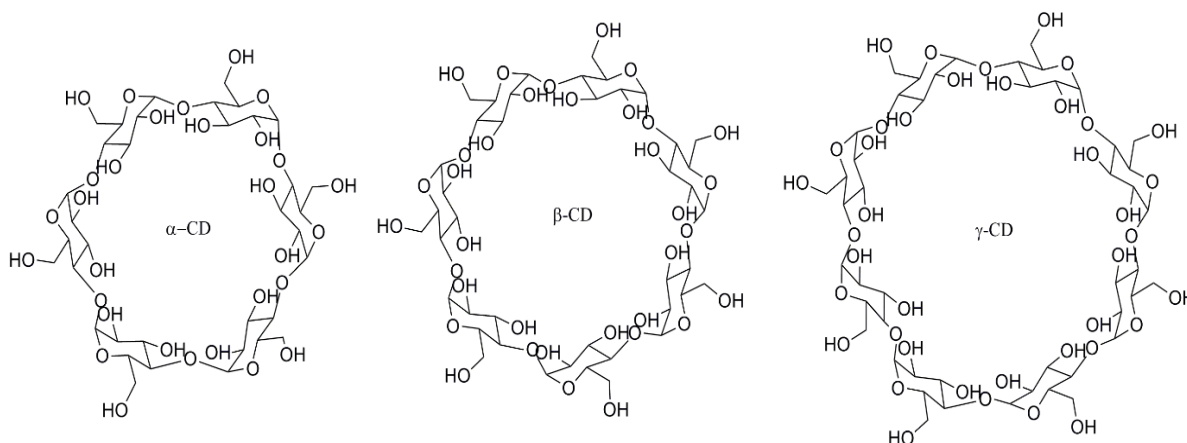

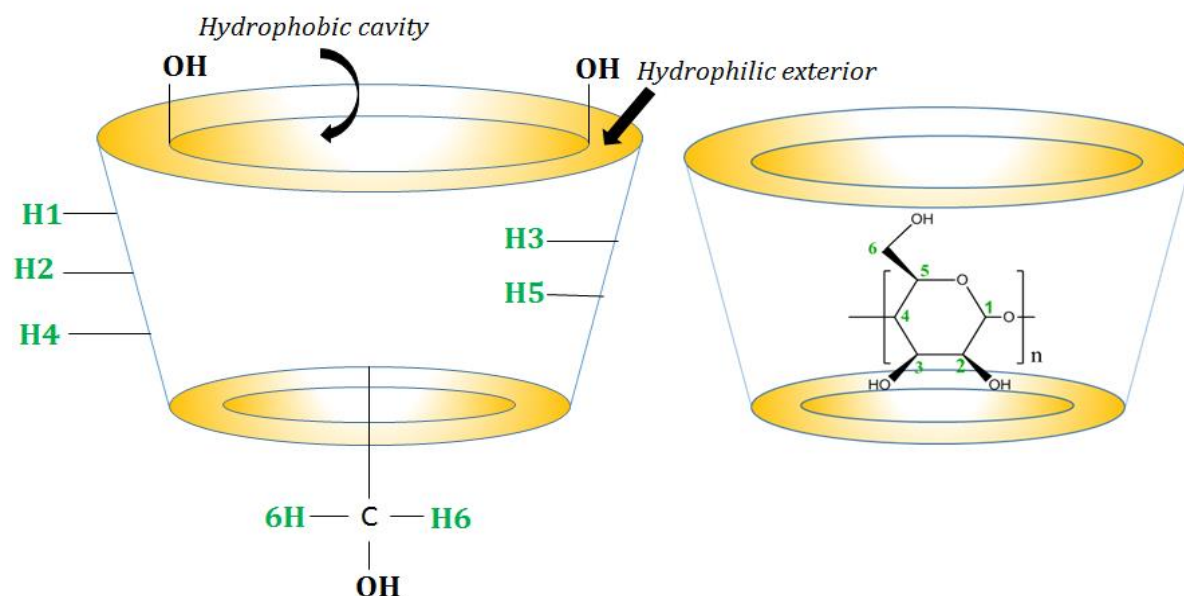**Figure II.4:** Structures of the α-, β- and γ-cyclodextrin

Figure II.5: General structure of cyclodextrin molecule with interior and exterior protons (n = 6, 7 for α-CD and β-CD respectively).

Table II.1: Charecteristics of Cyclodextrin.

Property	α-CD	β-CD	γ-CD
No. of glucose units	6	7	8
Empirical formula (anhydrous)	$C_{38}H_{60}O_{30}$	$C_{42}H_{70}O_{35}$	$C_{48}H_{80}O_{40}$
Mol. Wt. (anhydrous)	972	1135	1297
H_2O solubility (g/100 mL) at 25°C	14.5	1.85	23.2
Cavity diameter [Å]	4.7-5.3	6.0-6.5	7.5-8.3
Cavity length [Å]	7.9±0.1	7.9±0.1	7.9±0.1
Approx Cavity Volume ($Å^3$)	174	262	427

The low solubility of β-CD is attributed to the interruption by aggregated β-CD with its 7-fold symmetry units [29]. Szejtli [30] proposes that the intramolecular hydrogen bonds of the β-CD rim are responsible for its low solubility. Alkylation of β-CD hydroxyls leads to an increase in solubility and this phenomenon has constituted a motivation for carrying out such chemical modifications [27].

II.2.3 Inclusion Complex formation

Complex formation is a dimensional fit between the host cavity and the guest molecule [31]. The hydrophobic cavity of CD molecules arrange for a microenvironment into which appropriately sized non-polar moieties can enter to form inclusion complexes [32]. No covalent bonds are formed during the formation of the inclusion complex [33]. The main driving force of complex formation is the release of enthalpy-rich water molecules from the cavity resulting in a more stable lower energy state. The host-guest complexes can be prepared by various methods like co-precipitation method, paste method, dry mixing method, slurry method and analyzed by varieties of technique like FTIR, UV-visible and NMR spectroscopy, Powder XRD, DSC, TGA, Circular Dichroism spectroscopy, etc. [34]. The formation of inclusion complex of β-CD and a guest substance is accompanied by changes in their IR spectra as compared with the individual components [35, 36].

The binding of the guest molecules within the host CD is a dynamic equilibrium. Inclusion into CDs exerts a profound effect on the physicochemical properties of guest molecules giving rise to beneficial modification of guest molecules, which are not achieved otherwise [37]. The ability of a cyclodextrin to form an inclusion complex with a guest molecule is a function of two key factors. The first is steric and depends on the relative size of the CD to the guest molecule or certain key functional groups within the guest. The second critical factor is the thermodynamic interactions between the different components of the system (CD, guest and solvent). For a complex to form there must be a favorable net energetic driving force that pulls the guest into the CD.

II.2.4 Equilibrium

In an aqueous solution, the slightly apolar cyclodextrin cavity is occupied by water molecules which are energetically unfavored (polar-apolar interaction), and therefore can be readily substituted by appropriate 'guest molecules' which are less polar than water. One, two or three cyclodextrin molecules contain one or more entrapped 'guest' molecules. Most frequently the host: guest ratio is 1:1. The association of the host CD and guest (G) molecules, and the dissociation of the formed CD/guest complex are governed by a thermodynamic equilibrium,
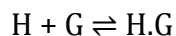
$$H + G \rightleftharpoons H.G$$

$$K = \frac{[H.G]}{[H][G]}$$

II.2.5 Higher order complexes

The complexation depends upon the cavity diameter of CD as well as the size of organic compound or guest molecule. Usually complexation occurs between one cyclodextrin and one guest (1:1 ratio) molecule. However, 1:2, 2:1, 2:2 and higher order complex equilibria also exist in the system [**Figure II.6**] [38-46].
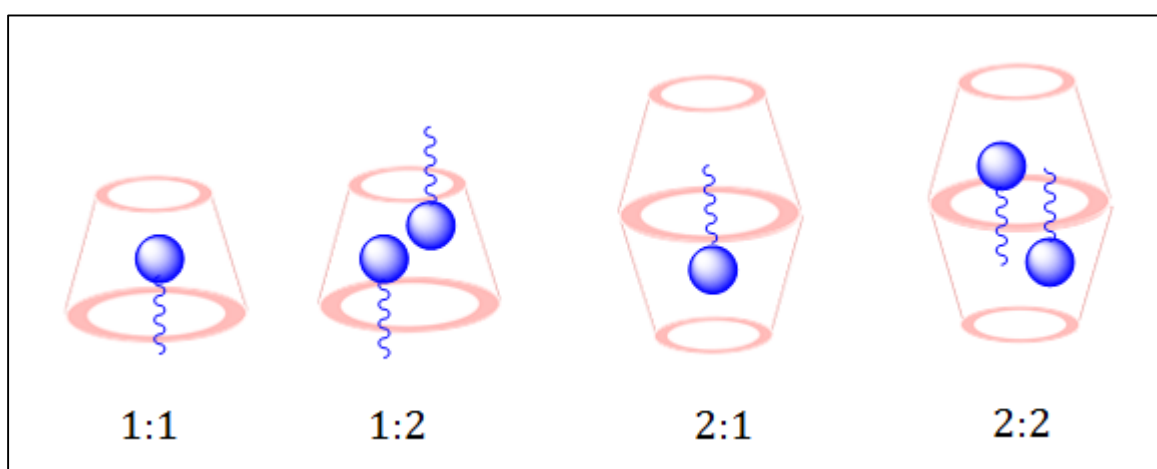
Figure II.6: Different stoichiometries of host-guest inclusion complexes.

II.2.6 Binding constants

Equilibrium constants for molecular association (binding constants/association constants/formation constants) have been measured using a variety of chemical approaches, including UV-visible, fluorescence and NMR spectroscopy and gas and liquid chromatography [47]. Benesi-Hildebrand method [48] is often used to calculate the binding constant values for 1:1 association. For UV studies, B.H equation is

$$\frac{1}{A - A_0} = \frac{1}{\Delta\varepsilon[G]K_a} \cdot \frac{1}{[CD]} + \frac{1}{\Delta\varepsilon[G]}$$

and for Fluorescence studies, the equation is

$$\frac{1}{I - I_0} = \frac{1}{[I' - I_0]K_a} \cdot \frac{1}{[CD]} + \frac{1}{I' - I_0}$$

II.2.7 Applications of cyclodextrins

Upon inclusion in cyclodextrins and modified CDs, the guest molecules undergo beneficial modifications. Cyclodextrins find versatile applications in innumerable fields ranging from agriculture to pharmaceuticals.

Cyclodextrins are extensively used in separations due to their ability to discriminate between positional isomers, functional groups, homologues and enatiomers. Cyclodextrins are used as chiral stationary phases bonded to solid support or as mobile phase additives in high performance liquid chromatography (HPLC) and in capillary electrophoresis for the separation of chiral compounds [49].

Cyclodextrins are potential candidates to play the role of drug delivery vehicles [50]. The majority of pharmaceutical active agents do not have sufficient solubility in water. The solubility of such drugs enhances by the formation of inclusion complex with cyclodextrin hence used in drug delivery system [51]. Drug solubility and dissolution, bioavailability, safety and stability are the main features in pharmaceutical applications of CDs [52]. Cyclodextrins stabilize penicillin from amylase degradation through complex formation [53]. The CDs also increase drug efficacy by sustained release of the drug [54]. The CD complexation enables enhanced absorption in different delivery routes such as oral, ocular, nasal, dermal, retinal etc. [55]

Cyclodextrins do not damage the microbial cells or enzymes. Cyclodextrins act as excellent 'enzyme models'. They catalyze biomimetic reactions of enzymes. Since CDs bind substrates, they are considered as artificial enzymes. They show substrate specificity and stereospecificity [56]. Quinine-β-CD complex provides an excellent model for mimicking enzyme-substrate interactions [57].

CDs are used in the food industry to protect active ingredients against light, heat and fermentation. Long-lasting perfumes are possible by slow release of fragrance through CD complexation. Unpleasant odors and bitter taste can be masked through complexation with CDs [58]. CD inclusion complex containing oily antimicrobial and

volatile agents are coated on a water absorbing sheet with a natural resin binder used for wrapping fresh products [59]. CDs find applications in waste water treatment. Also highly toxic materials from industrial effluents, organic pollutants and heavy metals from soil are solubilised in cyclodextrins and hence are removed [59]. Polymers in paints interact with each other resulting in thickening of paints and increase the coefficient of viscosity. β-cyclodextrin formulations prevent this undesirable effect [60]. A supramolecular chemosensor for aromatic compounds has been devised using β – cyclodextrin [61].

II.3. MACROCYCLIC POLYETHERS

Crown ethers, as originally defined are those compounds with multiple ether oxygen atoms incorporated in a monocyclic backbone. The term *"crown"* was used because the cavity shape of the macrocycle resembled a crown [62]. Crown ethers have proved to be unique cyclic molecules for molecular recognition of suitable substrates by hydrogen bonds, ionic interactions, cation−π or π−π interactions. Cation−π interactions are well described by a schematic drawing such as **Figure II.7** showing a K^+ ion interacting with the negatively charged π-electron cloud of benzene. The study of interactions involved in complexation of different cations with crown ethers in mixtures of solvents is important for a better understanding of the mechanism of biological transport, molecular recognition, and other analytical applications. **Figure II.8** shows some of the more common crown ethers.
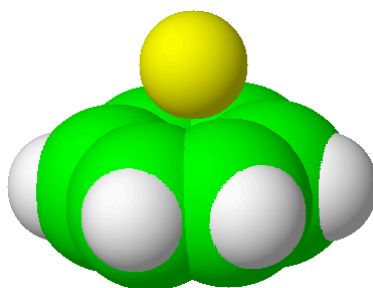

Fig. II.7: Schematic drawing of cation−π interactions showing the contact of K^+ ion and benzene.
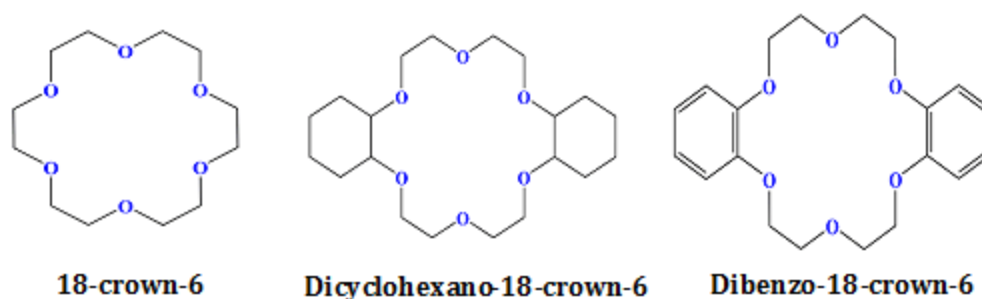
Figure II.8: Several example of crown ethers.

We chose 18-crown-6, Dicyclohexano-18-crown-6 (DC18C6) and Dibenzo-18-crown-6 (DB18C6) three crown ethers in our study. DB18C6 and DC18C6 were used to find out the effect of cyclohexyl group (electron donating group) and benzene group (electron withdrawing group) on the association constant.

Crown ethers are well known for their binding strengths and selectivities towards alkali and alkaline earth metal cations [63]. Their complexation with organic ammonium [64,65], arenediazonium [66-68], guanidinium [69-72], tropylium [73,74] and pyridinium [75] cations has also been studied.

II.3.1 Discovery of Crown Ethers

In 1987, Charles J. Pedersen, Jean-Marie Lehn and Donald J. Cram were jointly awarded the Nobel Prize in chemistry for their contributions to the development and use of supramolecular chemistry [76]. Pedersen serendipitously discovered dibenzo 18-crown-6 as a byproduct of a reaction in which he had sought to obtain bis[2-(*o*-hydroxyphenoxy) ethyl] ether, **1**. (**Figure II.9**). While trying to purify his product, he isolated a small amount of a white crystalline solid. Further analysis showed no alcohol groups to be present, and more interestingly, its UV spectrum changed significantly in the presence of sodium hydroxide. The unknown byproduct was identified as dibenzo-18-crown-6, **2**. Although this was not the first cyclic polyether to be reported in the literature, Pedersen was the first to recognize the unique properties of this type of compound, thereby opening up a whole new field of host-guest chemistry [77].
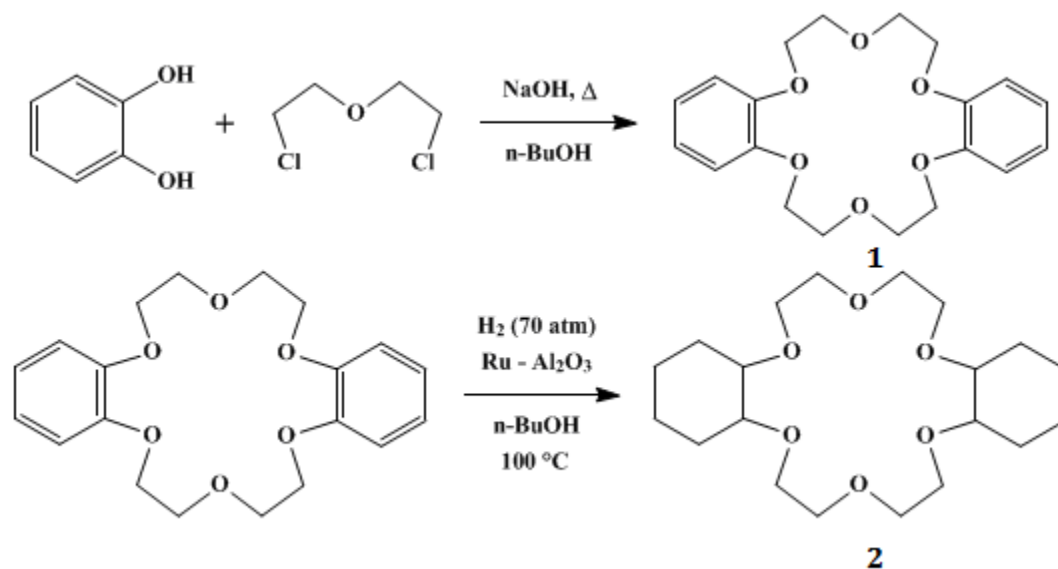
Figure II.9: Pedersen's reaction[78]

Actually long before Pedersen discovered crown ether by accident, there were others like Luttringhaus (Luttringhaus, 1937) who had preceded Pedersen in synthesizing macrocyclic polyethers. However, those researchers did not understand the unique cation-ligating properties of the cyclic polyethers. So Pedersen is widely regarded as the father of these important compounds. Since Pedersen's discovery in 1967, many scientists and research groups have synthesized crown ethers and have investigated the properties of many more examples of this class of compounds that serve as hosts for cations [79], anions [80, 81] and some neutral molecules [81].

II.3.2 Structure of Crown Ethers

Cyclic polyethers such as the one discovered by Pedersen consist of polyethylene glycol derivatives in a cyclic array wherein the heteroatom, most commonly oxygen, is directed toward the center of the cavity. The hydrocarbon backbone is located on the periphery of the ring, thereby rendering the cyclic polyether 'hydrophobic' on the outside and 'hydrophilic' on the inside. The heteroatoms, through their lone electron pairs are able to complex cationic guest molecules. Pedersen envisioned that the cyclic polyethers "crown" the cations; hence, the origin of the common name for these types of compounds: *crown ethers*.

II.3.3 Factors Governing Complexation and Selectivity of Crown Ethers

The factors responsible for the stability of cation-crown complexes can be classified as (a) relative size of a cation and crown compound cavity, (b) cation charge, (c) cation type, (d) counter anion, (e) number of donor atoms, (f) type of donor atom, (g) electron density of crown cavity, (h) crown substituents, (i) crown ether-ring flexibility, and (j) physical properties of a solvent. The major contributing factor towards binding and selectivity is attributed to size compatibility of the cation and the crown ether cavity. The cations that match the cavity size of the macrocycle are located in its center and optimize the interactions with the oxygen heteroatoms.

Functional groups also have been incorporated into host compounds in an effort to modify their complexation properties. Benzo substituent groups reduce the binding strength and selectivity of coronands and cryptands [82]. There are a number of reasons thought to be responsible for this:

(i) The O-O distance is smaller between the two oxygens joined by the benzene ring than those connected by an ethylene unit, thereby resulting in reduced cavity size.

(ii) The benzene ring also is electron withdrawing, thereby reducing the electron density and basicity of the adjacent oxygen atoms.

(iii) Benzo groups provide a rigid component that reduces the overall flexibility of the crown ether.

II.3.4 Applications of Crown ethers

Crown ethers have appreciable binding strengths and selectivities toward alkali and alkaline earth metal cations [77]. These special properties make crown ethers the first synthetic compounds that mimic many of the naturally occurring cyclic antibiotics. Due to the importance of alkali and alkaline earth metals (sodium, potassium, magnesium, calcium) in biological systems, in high –power batteries (lithium) and in isotope chemistry and radiochemistry, crown ethers are important ligands in the study of the chemistry of these metal ions [83].

Crown ethers are used in a wide range of areas such as analyses; separations, recovery or removal of specific species, ion selective electrodes, biological mimics and

reaction catalysts. Derivatives of crown ethers used as potential powerful anti tumor agents which is a very important step in fighting deadly diseases such as AIDS.

The studies of crown ethers and their derivatives have led to important advances in the area of molecular recognition and to the emergence of new concepts such as supramolecular chemistry [83]. The rapid development and the importance of molecular recognition as applied to macrocyclic compounds can be seen in the awarding of the Nobel Prize in chemistry in 1987 to three of its pioneers, namely Pedersen, Cram and Lehn.

II.4. METHODS OF DETECTING THE INCLUSION PROCESS

Structural characterization is of particular significance for supramolecular host-guest complexes, which are the basis of most CD applications in medicine, catalysis or in food chemistry, separation and sensor technology. NMR spectroscopy has become the most important method for structural elucidation of inclusion complexes. There are a few alternatives to NMR in the study of inclusion complexes such as fluorescence, UV-visible spectroscopy studies play a major role in measuring complexation energetics. IR, MASS methods are used in characterizing solid inclusion complexes.

II.4.1 NMR Spectroscopy

NMR spectroscopy has proved to be an efficient technique for the determination of the interactions between macrocyclic hosts and organic guests [84-89]. The study of cyclodextrin complexes by NMR was initiated by Demarco and Thakkar, [90] who observed 1H chemical shift variations of CD protons (H3 and H5) in the presence of various substrates and inferred that inclusion in the CD cavity had taken place. The spectacular advance of NMR techniques during the last few years has led to a much more detailed structural elucidation of inclusion complexes of macrocyclic hosts. NMR spectroscopy has been used from the very beginning to identify the hydrogen-bonding network in CDs. Inoue [91] and Schneider *et al.* [92] have reviewed NMR studies of the CD complexes. A typical structural inference is that if only H3 undergoes a shift in the presence of the substrate, then the cavity penetration is shallow, whereas if H5 also shifts,

the penetration is deep [27]. Structural information about the crown ether complexes can also be obtained from NMR spectra [75,93,94].

II.4.2 FT-IR Spectra of solid inclusion complexes

The solid inclusion complex formation is analyzed by FT-IR spectroscopy. FT-IR spectrum is used to confirm the formation of the solid inclusion complex by considering the deviation of peak shape position and intensity [88, 95-98].

II.4.3 Optical Spectroscopy

II.4.3.1 UV-visible spectroscopy

UV–vis spectroscopy is a convenient and widely used method for the study of binding phenomena [48,99-102]. Complexation between the macrocyclic hosts such as cyclodextrin and crown ethers and either completely or partially included guest molecules results in a number of interesting spectroscopic effects. Spectrophotomeric determinations of inclusion complexes rely on the difference in the absorptivities of free and complexed substrate. Since the host molecule may in principle modify both photochemical and photophysical properties of their guest molecules, both absorbance and fluorescence signals can be used to gain insight into the characteristics of the complexes formed [103]. On adding host to the substrate solution in a suitable solvent, there is a hike in the absorbance in most cases, followed by a red/blue shift in absorbance maximum in a few cases.

II.4.3.2 Fluorescence spectroscopy

The supramolecular interaction of guest molecules with CDs can be investigated by spectrofluorimetry [104-108]. With an increase in the CDs (α-CD and β-CD) concentration, however, the fluorescence intensity of the guest molecules enhanced accompanied by a slight hypsochromic shift of the emission peak. These findings indicates the formation of host-guest inclusion complexes. Cramer *et al.* first noticed that 8-(phenylamino)-lnaphthalenesulfonate exhibited a more intense emission in the presence of β-cyclodextrin and presumed a 1:1 complex was formed. Since then other

authors have reported enhancement of fluorescence of benzene and its derivatives in cyclodextrin [109]. The development of fluorescence methods for quantitative analysis is particularly important due to their sensitivity, selectivity and instrumental simplicity.

II.4.4 Surface Tension

Surface tension (γ) is another valuable parameter which also suggests the formation of inclusion complex of a guest molecule into cavity of host molecule. The addition of CD to pure water does not show any considerable change to the surface tension of water which is an indication that both cyclodextrins are almost surface inactive compounds. With the successive addition of aqueous cyclodextrin solution the surface tension values substantially increased for the surface active guest molecules probably due to the removal of surface active guest molecules from the surface of the solution, i.e., the hydrophobic part enter the hydrophobic cavity of the cyclodextrin forming the host guest inclusion complexes [110]. If we get a single break point in the surface tension curve and after that point if the γ value becomes approximately steady which confirms the formation of a 1 : 1 inclusion complex. More break points in the surface tension curve is the indication of the formation of inclusion complexes with complex stoichiometry such as 1 : 2, 2 : 1, 2 : 2 etc. [111,112].

II.4.5 Conductance Study

The conductivity (k) study not only confirms the formation of a host–guest inclusion complex but also gives the stoichiometry of the assembly [113,114]. With the successive addition of cyclodextrin in the solution of guest molecule, the conductivity of the guest molecules decreases on a regular basis. This type of observation is in good agreement with the formation of inclusion complexes. The insertion of the guest molecule inside the cavity of the CD molecule decreases the number of free guest molecule, resulting in the reduction in conductivity of the solution. The curves having a noticeable break suggests the formation of host-guest inclusion with a stoichiometry of 1 : 1.

II.5. SIGNIFICANT MOLECULES

Significant molecules are those molecules which are very much imperative and essential to our environment systems and which can be used in medical applications. Such molecules are ionic liquids, vitamins, drug molecules and amino acids etc. Here we have discussed about ionic liquids and drug molecules.

II.5.1 Drug molecules

➢ *Drugs: Definition*

By definition, drugs are chemical substances that affect/alter the physiology when taken into a living system. They can be natural or synthetic. Chemically, they are low molecular mass structures. When a drug is therapeutically active and is used for the diagnosis, treatment or prevention of a disease, it is called medicine (legal drugs). They target the macromolecules inside the body and generate a biological response. Most of them interrupt the nervous system (especially brain) for the generation of a proper biological response. However, they can be toxic in higher doses called lethal dose.

➢ *Classification of Drugs*

Classification of drugs can be done on the basis of different criteria. Some of the criteria for the classification of drugs are given below.

❖ *On the basis of pharmacological effect:*

How a drug or medicine affects or influences the cells of an organism is referred to as the pharmacological effect. Drugs have different pharmacological effects on an organism; say, for example, an analgesic reduces the pain while an anti-inflammatory drug reduces the inflammation of the body. Thus, drugs can be classified based on the pharmacological effect.

❖ *On the basis of drug action:*

Different drugs act differently i.e., each drug has its own way of generating response called drug action. Drug action is more specified according to how it generates a response. For example, there are lots of medicines to treat hypertension but each type of drug has different drug action. All the hypertension medicines reduce the blood pressure but in a different pathway.

❖ *On the basis of chemical structure:*

This is a common classification of drugs. Generally, drugs which have same drug action and pharmacological effect have a basic skeletal structure and a minute variation in the branching. This is why some drugs have more potential than the other. For example, all sulphonamides have the same skeletal structure.

❖ *On the basis of molecular targets:*

Drugs target the macromolecules inside the body to generate a biological response. Such macromolecules are called target molecules or drug targets. Drugs which have the same mechanism of action will have the same target. This basis for classification of drugs is more helpful during clinical trials.

➢ *Discovery of drug*

Drug discovery is the process of finding new drugs. It has evolved from early serendipitous discovery from natural sources, such as morphine from poppy seeds, to today's industrial-scale screening projects [115,116]. Modern drug discovery starts with the identification of a biological target that can be modulated to induce the desired therapeutic effect. To search for potential drugs, compounds are tested for their ability to modulate the target.

Technologies such as high-throughput screening (HTS) and combinatorial chemistry facilitate screening of libraries of millions of compounds. However, even though screening capacity reaches millions of compounds and continues to grow, the number of possible molecules that could be synthesized and tested is infinitely larger. This notion

has important consequences; apparently, we can only test a tiny sample of what is virtually available.
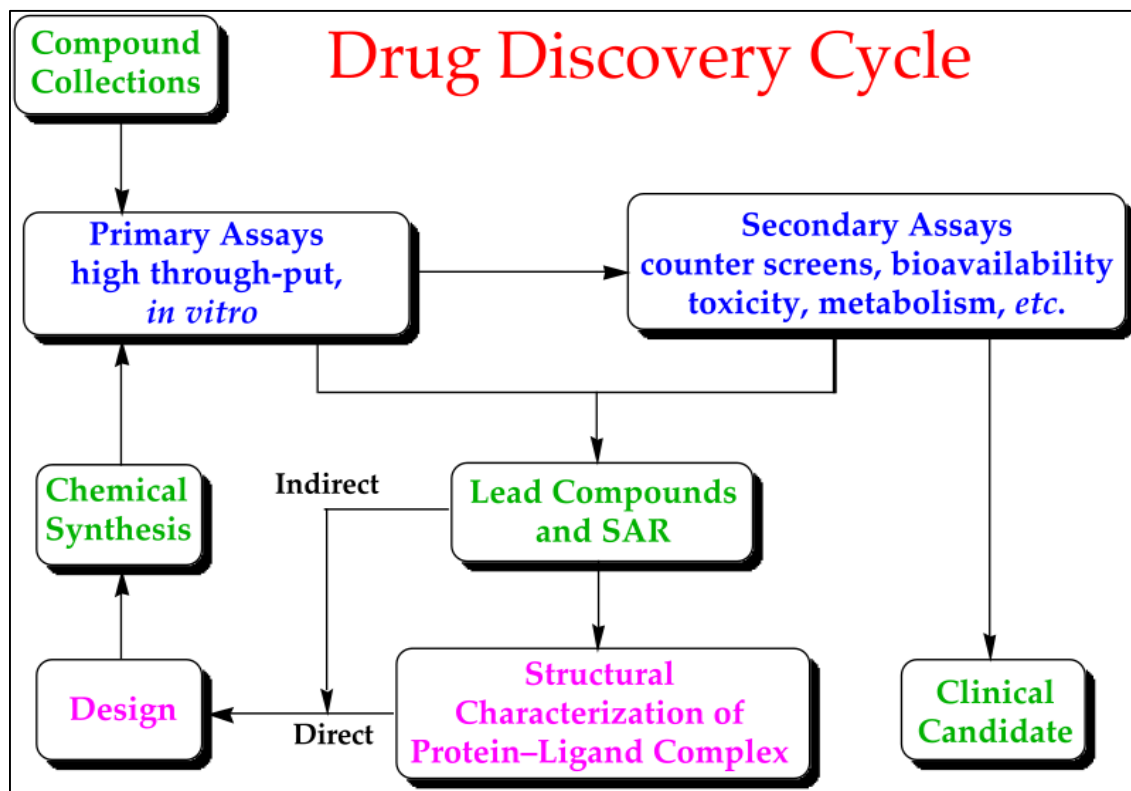
It includes pre-clinical research (microorganisms/animals) and clinical trials (on humans) and may include the step of obtaining regulatory approval to market the drug.

The development of every drug begins with the search for a target on which the drug can act. To find one, scientists need to have very precise knowledge of the biochemical processes that take place in the body and how these are changed by a disease. They must be able to bind well to the target protein; i.e. they must fit into the target like a key into a lock.

❖ **Some Examples of Drug Molecules**

Lipitor from mevastatin (cholesterol-lowering)

Doxorubicin (anti-cancer)

Cyclosporine (immunosuppressant)

Azithromycin from erythromycin (anti-bacterial)

Prograf (immunosuppressant)

Amoxicillin from penicillin (anti-bacterial)
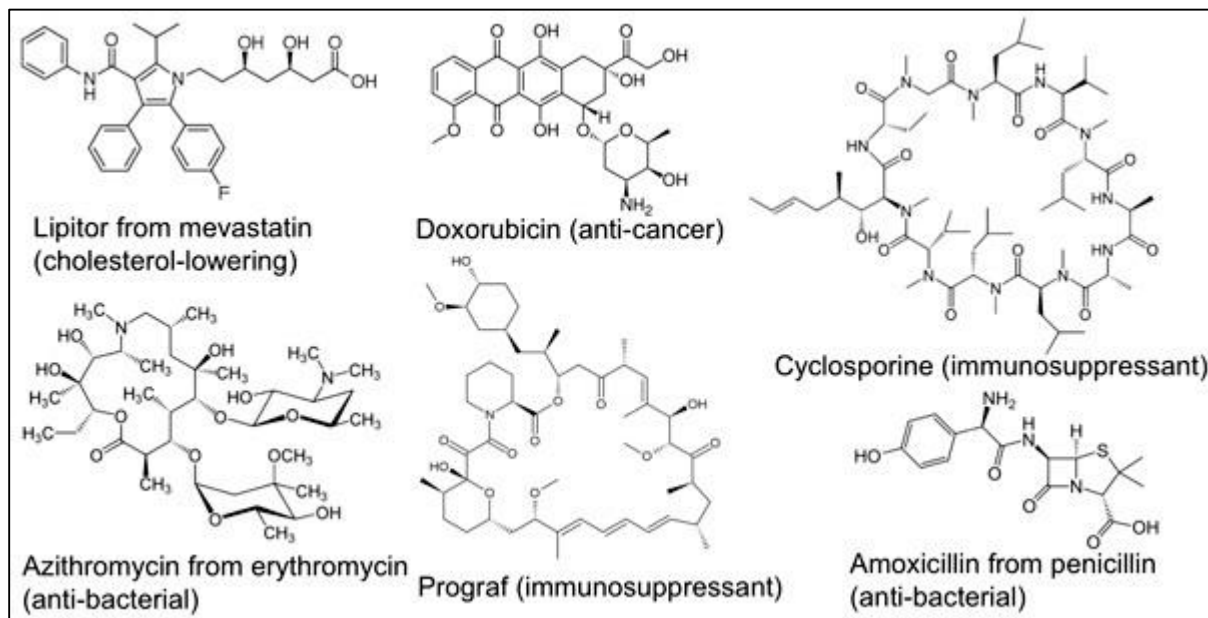
II.5.2 Ionic Liquids

In the last decades a new class of compounds came into the focus of many research groups around the world: ionic liquids (IL). The count of publications with the topic "*ionic liquids*" grew steadily over the last ten years. But what are ionic liquids and why are they so interesting? The following section gives a short overview over this wide field. A much more exhaustive overview about the possible applications and properties of ionic liquids can be found in the recent book "*Ionic Liquids in Syntheses*", edited by Peter Wasserscheid and Tom Welton [117]. The commonly accepted definition of ionic liquids is that they are "*ionic materials that are liquid below 373 K*". Ionic liquids have gained a lot of attention as emerging environmentally benign solvents. They can replace conventional organic solvents in several applications due to their unique features. Ionic liquids are salts with melting points below 373 K. They consist of an organic cation combined with an organic or inorganic anion. Ionic liquids show, in general, a very interesting set of properties to be used for different applications in chemical industry. The melting points of these organic salts are frequently found below 150 °C [118] and occasionally as low as –96 °C. Some ionic liquids are stable up to 500 K. At room temperature they have no

measurable vapor pressure due to their ionic nature. They normally have high solvency power for polar and non-polar compounds. Billions of ionic liquids can be designed and synthesized by selecting different ion pair combinations, which enable them to possess specific properties. Furthermore, the ability to tune the solvent properties of the ionic liquids is one of their outstanding features, which makes them unique solvents for various reactions and separations. Moreover ionic liquids are almost nonflammable, highly thermally and (electro) chemically stable and present a large liquid range.

Interest in ionic liquids has now been grown dramatically in the scientific community (both in academia and industry) with over 8000 papers having been published in the last decade [119].This growth can be observed in Figure **II.10**. (number of publications per year) and **Figure II.11**. (number of patents per year), where the number of publications and patents are increasing exponentially. There are about one million (10^6) simple ionic liquids that can be easily prepared in the laboratory by combination of different cations and anions and this total are just for simple primary systems. If there are one million possible simple IL systems, then there are one billion (1012) possible binary combinations of these, and one trillion (10^{18}) ternary possible IL systems that can be prepared from the combination of anions, cations, and substituents. Hence in this field many mores to discover.
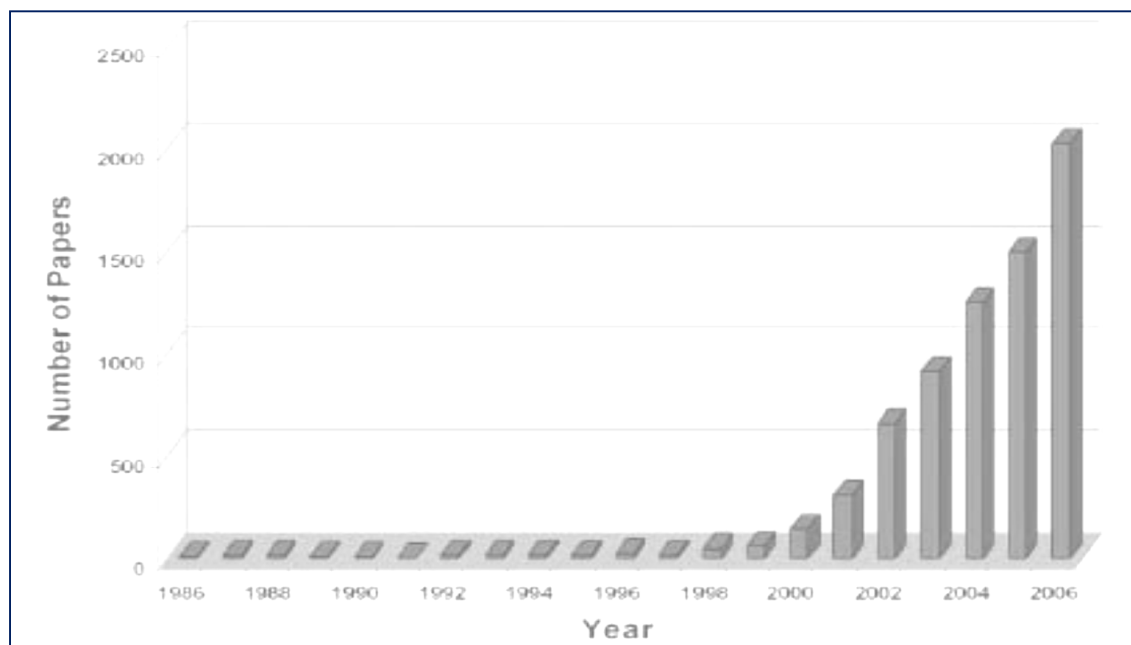

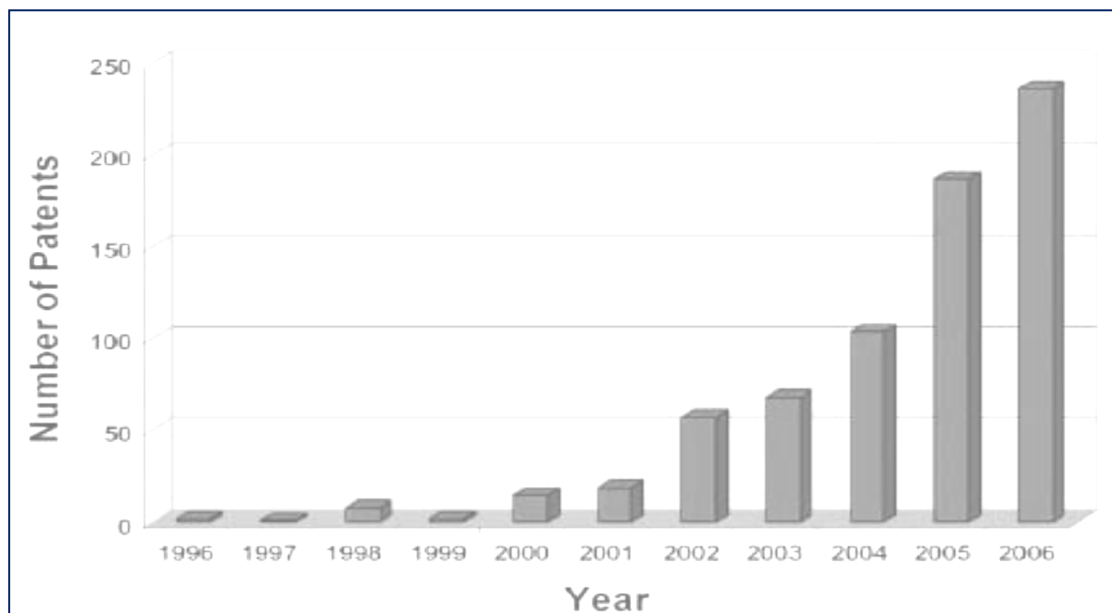**Figure II.10:** Growth rate of ionic liquid publications, 1986-2006.

Figure II.11: Annual growth of ionic liquid patents, 1996-2006.

Until 1998, the number of entries with the terms "*ionic liquid*" or "*ionic liquids*" in the Chemical Abstracts was below or around twenty per year, but this number is increasing from 45 per year in 1999 to 1255 per year in 2005 and 1717 in 2006 with their different use in various field.

II.6. SOLVATION BEHAVIOUR

The term "Solvation Behaviour" which is connected with dissolved state of matter has drawn the interest of physical chemists for quite a long time. This chemistry in fact forms the origin of careful investigations to understand the physicochemical properties of solution and to know the mechanistic paths of solvation of different solutes like ionic liquids and drug molecules [120-125]. In fact, lot of chemistry relates to and is conducted in liquid solutions and involves ionic and non-ionic species. The vast bodies of multi-component aqueous salt solutions of oceans, nature of physiological fluids, dissolution of proteins, nutrients, enzymes, sugars, etc., in body fluids, oxygen in blood, various industrial processes, use of different solvents in the fields of chemical, analytical, electrochemical, food, pharmaceuticals, ecological and photochemical chemistry all are

the general subject of interest of solution chemists. In earlier experiments the chemists were attracted [122-129] about the role of different solutes on various physicochemical properties and processes in aqueous, non-aqueous and mixed solvents. The role of solvents was assumed to offer an inert and homogeneous medium for the chemical reactions. But later, it has been duly recognized that solvent molecules play a significant role in dictating various aspects of physico-chemical properties, equilibrium, and kinetic behavior of reactions in solution phase.
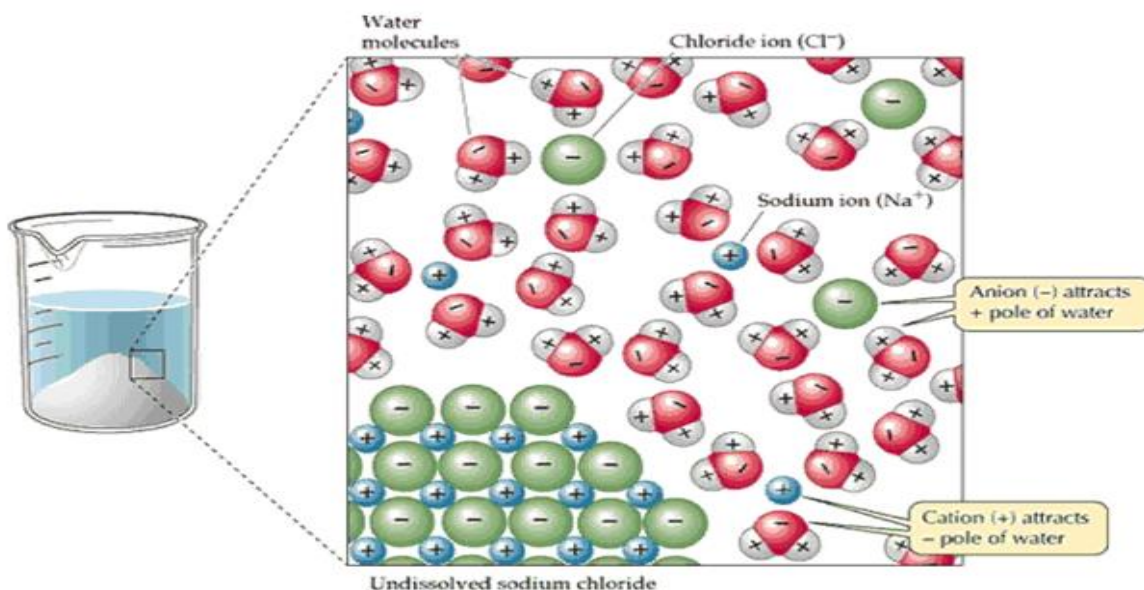
Solvation behaviour have been studied in a wide range of aqueous, non-aqueous, electrolytes and mixed solvents based on the various properties viz. lowering of vapor pressure by solutes [129], order of ionic mobilities [125,129,130], viscosity B-coefficients [129-132], partial molar volume of ionic and non-ionic solutes, dielectric decrement produced by ions [133], solubility of substances [134], effect of ionic and non-ionic solutes on solvent spectra [133,135,136], etc. Several investigations [122,124,130,135,137,138-144] were carried out by the use of spectroscopic methods such as UV-VIS, IR, NMR and considering thermodynamic, kinetic, transport and electrochemical properties at infinite dilution. All these studies provide an idea about the solute/ion-solvent interactions [129, 130,145]. Last few decades have witnessed an exponential growth on the fundamental research involving solute-solute, solute-solvent, and solvent-solvent interactions in aqueous and aquo-organic mixtures in particular but

in pure or mixed organic solvents the research was not established to that extent. In this context, much attention has been given also to determine the various thermodynamic parameters such as enthalpy, entropy and Gibbs free energy change of solutions [146-149]. The purposes of such studies were to gain the ideas about various mechanisms of some significant molecules during solvation. The exploration of molecular interaction existing in solution is always an interest to the chemists. Molecular interaction can be studied in solution phase by studying it's thermodynamic and transport properties. Results of such proposed studies will have enormous fundamental and physical consequences in the land of solution chemistry, biophysical chemistry, and pharmaceutical sciences.

Solvation (specifically, hydration) is important for many biological structures and processes. For instance, protein folding occurs spontaneously because of a favourable change in the interactions between the protein and the surrounding water molecules. Proteins are stabilized by 5-10 kcal/mol when folded due to a combination of solvent and hydrogen bonding effects. Minimizing the number of hydrophobic side-chains exposed to water by burying them in the centre of a folded protein is a driving force related to solvation. Solvation can also drive host-guest complexation. Host molecules having hydrophobic cavity can encapsulates a hydrophobic guest within its cavity. These interactions can be used in applications such as drug delivery, such that a hydrophobic drug molecule can be delivered in a biological system without needing to covalently modify the drug in order to solubilize it. Polarity of the solvent affects the binding constants for host-guest complexes.

Our works is concerned mainly with solvation of solutes (ionic liquids) and their thermodynamic aspects. The term 'solvation' can be defined as the more or less specific interaction between a solute molecule and one or more solvent molecule with a force intermediate between weak physical interaction and strong covalent bonding [151]. As we know solutes are generally two types; one is electrolytes and another is nonelectrolytes. Electrolytes when dissolved in suitable solvent, called the ionizing medium, produce ions. The basic difference between electrolytes and non-electrolytes then lies in the ability of the former to generate ions, which are distinguished by their integral charge, and the intense electric field associated with it. In fact, solute-solvent

interactions guide all the properties of solutions. Solvation phenomena have been studied in a wide range of aqueous, non-aqueous and mixed solvents from various properties. Remarkable advancements have been made by a number of researchers providing answer to a wide range of relevant problems, concerning solute-solvent interactions from both experimental and theoretical stand points also.
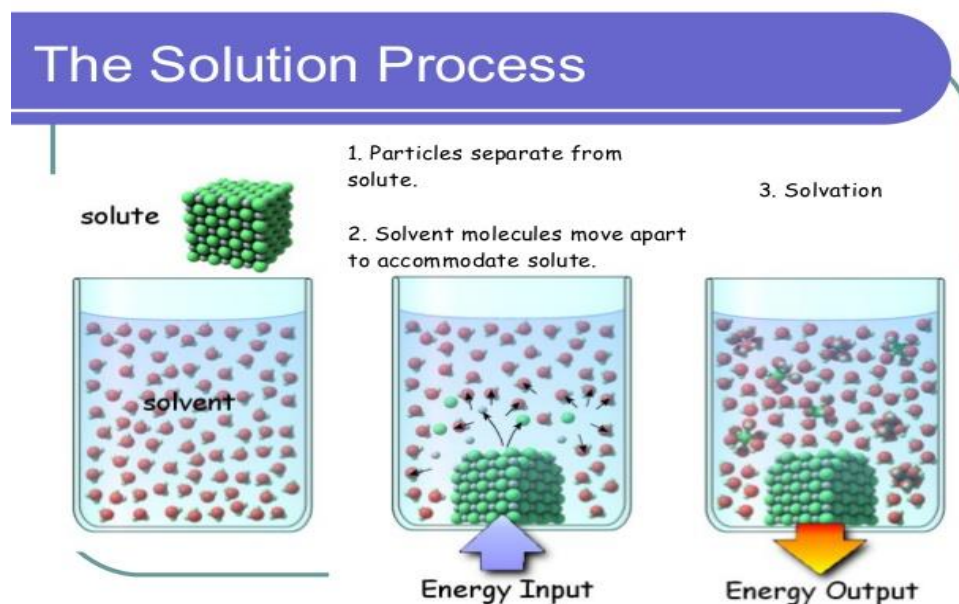
II.6.1 Interactions in Solution Systems

Three types of interactions are mainly observed the solution systems:

  (a) *Solvent–solvent interactions:* energy required to break weak bonds between solvent molecules.

  (b) *Solute–solute interactions:* energy required to break intermolecular bonds between the solute molecules.

  (c) *Solute–solvent interactions:ΔH* is negative since bonds are formed between them.
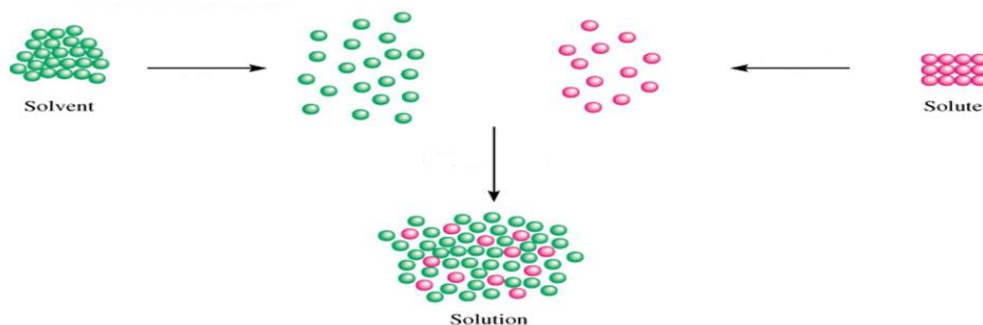
For liquid systems, the macroscopic properties are usually quite well known, whereas the microscopic structure is often much less studied. The liquid phase is characterized by local order and long-range disorder, and to study processes in liquids, it is therefore valuable to use methods that probe the local surrounding of the constituent particles. The same is also true for solvation processes: a local probe is important to obtain insight into the physical and chemical processes going on.

II.6.2 Investigation on Different Kind of Interactions

When salt is dissolved in water, the ions of the salt dissociate from each other and associate with the dipole of the water molecules. This result in a solution is called '*solvation*'.
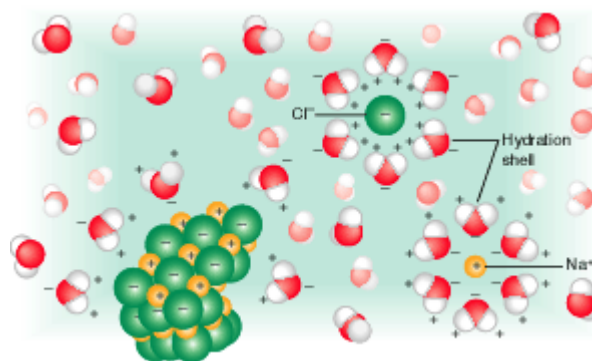
This means that the forces of interaction (attraction or repulsion) depending on whether like or unlike charges are closer together. On average, dipoles in a liquid orient themselves to form attractive interactions with their neighbours, but thermal motion makes some instantaneous configurations unfavourable.

The investigation on thermodynamic, transport and optical properties of different electrolytes in various solvents would thus offer an important step in this direction. Naturally, in the development of theories, dealing with electrolyte solutions, much attention has been devoted to 'ion-solvent interactions' which are the controlling forces in infinitely dilute solutions where ion-ion interactions are absent. The contributions due to cations and anions in the solute-solvent interactions can be obtained by separating these functions into ionic contributions. Thus ion-solvent interactions play a crucial role to understand the thermodynamic and physicochemical properties of solutions.

One of the causes for the intricacies in solution chemistry is that the structure of the solvent molecule is not known with certainty. The introduction of a solute also modifies the solvent structure to an uncertain magnitude whereas the solute molecule is also modified and the solute-solute, solute-solvent and solvent-solvent interactions become prominent, though the isolated picture of any of the forces is still not known completely to the solution chemist.

The ion-solvent interactions can be studied from the physicochemical, and thermodynamic point of view, where the changes of free energy, enthalpy and entropy, etc. associated with a particular reaction. Qualitatively and quantitatively evaluated various physicochemical parameters, from which concluded regarding the factors associated with the ion-solvent interactions occurred in the studied solutions.

Similarly, the ion-solvent interactions can be studied using solvational approaches involving the studies of different properties such as, density, viscosity, ultrasonic speed, refractive index and conductance of electrolytes and various derived parameters, factors associated with ionic solvation.

We shall particularly dwell upon the different aspects of these physicochemical, thermodynamic, transports, acoustic and optical properties as the present research work is warmly allied to the studies of ion-ion, ion-solvent and solvent-solvent interactions.

II.6.2.1 Ion-Solvent Interaction

Ion-solvation is a phenomenon of primary interest in many milieu of chemistry because solvated ions are omnipresent on Earth. Hydrated ions occur in aqueous solution in many chemical and biological systems [152]. Solvated ions appear in high

concentrations in living organisms, where their presence or absence can fundamentally alter the functions of life. Ions solvation in organic solvents, mixtures with water and other organic solvents are awfully common [153]. The exchange of solvent molecules around ions in solutions is fundamental clue to the understanding of the reactivity of ions in solution [154]. Solvated ions also play a key role in electrochemical applications, where for instance the conductivity of electrolytes depends on ion-solvent interactions [155].The significance of ion-solvent interactions was realized after extensive studies in aqueous, non-aqueous and mixed solvents [156-165].

Most chemical processes of individual and biological importance occur in solution. The role of solvent is so great that million fold rate changes take place in some reactions simply by changing the reaction medium. As water is the most abundant solvent in nature and its major importance to chemistry, biology, agriculture, geology, etc., water has been extensively used in kinetic and equilibrium studies. But still our knowledge of molecular interactions in water is extremely limited. Moreover, the uniqueness of water as a solvent has been questioned [166,167] and it has been realized that the studies of other solvent media like non aqueous and mixed solvents would be of great help in understanding different molecular interactions. The organic solvents have been classified on the basis of dielectric constants, organic group types, acid base properties, or association through hydrogen bonding donor-acceptor properties [168,169] hard and soft acid-base principles [170] etc. As a result, the different solvents show a wide change in properties, ultimately influencing their physicochemical, thermodynamic, transport and acoustic properties qualitatively and quantitatively, in presence of electrolytes and non-electrolytes in these solvents.

II.6.2.2 Ion-Ion Interaction

Ion-solvent interactions are only a part of the story of an ion related to its environment. The surrounding of an ion sees only other ions, no solvent molecules. The mutual interactions between these ions represent the essential part '*ion-ion interactions*'. The properties of solution is affected by the degree of ion-ion interactions. Solution chemistry will become understandable only when solute-solute/ion-ion, solute-solvent/ ion-solvent and solvent-solvent interactions are elucidated. And thus the work on

solvation of ionic liquid is related to the studies of solute-solute/ion-ion, solute-solvent/ion-solvent and solvent-solvent interactions in some liquid systems.

II.6.3 Density

The physicochemical properties of liquid mixtures have attracted much attention from both theoretical and engineering applications points of view. Many engineering applications require quantitative data on the density of liquid systems. They also provide information about the nature and molecular interactions between electrolyte or non-electrolyte and liquid components.

The volumetric property '*Density*' is a function of weight, volume and mole fraction and excess volumes of mixing. One of the well-recognized approaches to the study of molecular interactions in fluids is the use of physicochemical, thermodynamic methods. Physicochemical properties are generally convenient parameters for interpreting solute-solvent/ion-solvent and ion-ion/solute-solute interactions occurring in the solution phase. Fundamental properties such as enthalpy, entropy and Gibbs energy represent the macroscopic state of the system as an average of numerous microscopic states at a given temperature and pressure. Molecular phenomena is generally difficult to understand in terms of density value hence higher derivatives of these properties employed. The volumetric information may be of immense importance in this regard. Various concepts regarding molecular interaction in solutions is electrostriction [171], hydrophobic hydration [172], micellization [173], and co-sphere overlap during ion-solvent/solute-solvent interactions [174] have been derived and interpreted from the partial molar volume data for electrolytes and non-electrolytes.

❖ *Apparent and Partial Molar Volumes*

Density data can be used for the calculation of molar volume of a pure substance. In complex multi-component systems such as solutions, it is easier to describe a system in terms of the intrinsic or molal properties rather than the extensive properties. Any extensive property of a system can be calculated as the sum of the respective partial molal properties if all components have a known concentration. Although the additive

definition of partial molal properties is convenient, direct measurement of these solution properties are difficult, because interactions with other species contribute to partial molal properties. When dealing with solutions it is more common to measure the apparent moral quantities, ϕ_Y, which can be defined as the change in property, Y, due to a known amount of solute in a known amount of solvent, assuming the contribution by the solvent is the same as that of the pure species. In other words, all changes in the state properties can be attributed to the presence of the solute, even if these changes contribute to a change in partial molal property of the solvent. The apparent molal property of any solute, $\phi_{2,Y}$, can be defined as

$$\phi_{2,Y} = \frac{Y - n_1 \overline{\phi}_Y^0}{n_2}$$

(II. 1a)

Where, n_1 and n_2 represent the mole of the solvent and solute respectively, in the system. $\overline{\phi}_Y^0$, denotes the partial molal property of the pure solvent. Because the solvent is assumed to contribute a definite, constant quantity for all solute concentrations at fixed temperature and pressure, the partial derivative of the extensive property with respect to the number of moles of solute can be defined in terms of the apparent molal quantity:

$$\left(\frac{\partial Y}{\partial n_2} \right)_{n_1, T, P} = \left(\frac{\partial \phi_{2,Y}}{\partial n_2} \right)_{n_1} + \phi_{2,Y}$$

(II. 1b)

Eqs (II.1a) and (II.1b) shows that if the apparent molal property and its derivative with respect to moles of solute is known, the partial molal property can be calculated. In other words, if any appropriate Eq. for the apparent molal property with respect to any concentration scale, which is in good agreement with experimental data is known, the apparent molal property at infinite dilution can also be found out. Theoretically, as the concentration approaches zero, the apparent molal property approaches the partial molal property of the solute at infinite dilution, because by definition the solvent is already assumed to be in its pure form. If the apparent molal property is assumed to reflect the apparent molal property of the solute only and not the solute-solvent complex, then the apparent molal property at infinite dilution, $\phi_{2,Y}^\infty$, would be equal to the standard state molal property of the solute, as defined by Henry's law. Ignoring the previous

equality, equation of state developed for standard state partial molal variables have been used successfully to describe partial molal quantities at infinite dilution.
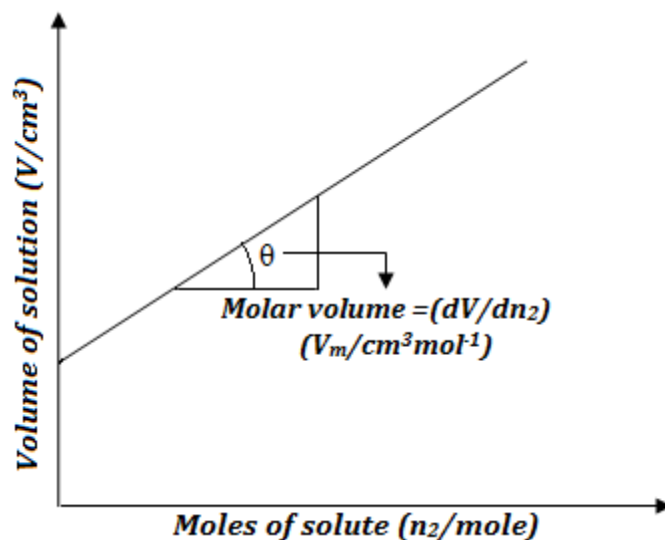
Figure II.12: A diagram for the explanation of molal volume.

The easiest way to explain this is in terms of the molal volume, V_m, shown in the **Figure II.12**, where the volume increases with respect to the amount of solute added. A dissolved solute has its own property, referred to as partial molal property.

Therefore for a two component system, where one component is the solvent and the other is solute, the total volume of the system can be represented as the sum of the partial molal volumes of the solvent, $\overline{V}_1$, and the solute, $\overline{V}_2$:

$$V = n_1 \overline{V}_1 + n_2 \overline{V}_2 \tag{II. 1c}$$

Dividing Eq. (II. 1c) by $n_1 + n_2$, the molal volume of the solution is obtained as:

$$V_m = x_1 \overline{V}_1 + x_2 \overline{V}_2 \tag{II. 1d}$$

where, x_1 and x_2 represent the mole fraction of the solvent and the solute, respectively. The partial molal property of a solute is defined as the change in the total property of the system with respect to the change in the number of moles of solute added, with all other variables (T, P, and the amount of the solvent) are held constant. An alternative, widely used property of the solute is the apparent molal property. The apparent molal volume

is the volume that should be attributed to the solute in solution if it is assumed that the solvent contributes the exact volume it would if it was in its pure state.

Under this assumption, the apparent molal volume of the solute, $\phi_{2,V}$, as defined by Harned and Owen, is the difference between the total volume (or the total molal volume) and the partial molal volume of the pure solvent ($\overline{V}_1^0$) divided by the number of moles (or the mole fraction) of the solute present:

$$\phi_{2,V} = \frac{V - n_1 \overline{V}_1^0}{n_2} \tag{II. 1e}$$

$$\phi_{2,V} = \frac{V_m - x_1 \overline{V}_1^0}{x_2} \tag{II. 1f}$$

In experimentation, $\overline{V}_1^0$ is generally considered to be constant over the range of solute concentration at constant temperature and pressure. Hence, $\phi_{2,V}$, can be easily calculated using Eq. (II. 1e) or (II. 1f) when the total volume or molal volume is known. Eq. (II. 1e) can be modified in order to find out the apparent molal volume of a solute using density of the solution containing the solute and the density of the pure solvent, ρ and ρ_0, respectively. Assuming there is 1 kilogram (kg) of solvent:

$$n_1 = \frac{1}{M_1}, \qquad \text{therefore,} \qquad V = \frac{\overline{V}_1^0}{M_1} + m_2 \cdot \phi_{2,V} \tag{II. 1g}$$

where, M_1 is the molal mass of the solvent.

Since, $\rho_1 = \frac{\overline{V}_1^0}{M_1}$, Eq. (II. 1g) becomes:

$$V = \frac{1}{\rho_1} + m_2 \cdot \phi_{2,V} \tag{II. 1h}$$

where, m_2 is the molality of the solute (which is equivalent to n_1 if 1kg of solvent is present). The entire mass of the solution will be composed of the mass of the solvent (1kg) and the mass of the solute ($m_2 \cdot M_2$). Since volume is the ratio of mass to density, the equation for V becomes:

$$\frac{1 + m_2 \cdot M_2}{\rho_1} = \frac{1}{\rho_1} + m_2 \cdot \phi_{2,V} \tag{II. 1i}$$

By rearranging this Eq. (II. 1h) and solving for $\phi_{2,v}$, an equation for apparent molal volume for 1 kg of the solute is obtained:

$$\phi_{2,v} = \frac{\rho_2 - \rho_1}{m_2 \rho_1 \rho_2} + \frac{M_2}{\rho_2} \tag{II. 1j}$$

However, the volume contributed to a solvent by the addition of one mole of an ion is difficult to determine. This is so because, upon entry into the solvent, the ions change the volume of the solution due to a breakup of the solvent structure near the ions and the compression of the solvent under the influence of the ion's electric field, *i.e.*, *electrostriction*. Electrostriction is a general phenomenon and whenever there are electric fields of the order of 10^9-10^{10} V·m^{-1}, the compression of ions and molecules is likely to be significant. The effective volume of an ion in solution, the partial molar volume, can be determined from a directly obtainable quantity apparent molar volume (ϕ_V). Thus the apparent molar volumes of the solutes can be calculated by using the following relation [175];

$$\phi_V = \frac{M}{\rho_0} - \frac{1000(\rho - \rho_0)}{c\rho_0} \tag{II.1k}$$

or

$$\phi_V = \frac{M}{\rho} - \frac{1000(\rho - \rho_0)}{m\rho\rho_0} \tag{II.1l}$$

Where M is the molar mass of the solute, c is the molarity, m is the molality of the solution; ρ_0 and ρ are the densities of the solvent and the solution respectively. The partial molar volumes, ϕ_{2v} can be obtained from the equation [176];

$$\phi_{2v} = \phi_v + \frac{(1000 - c\phi_v)}{2000 + c^{3/2}\left(\frac{\partial \phi_v}{\partial \sqrt{c}}\right)} c^{1/2}\left(\frac{\partial \phi_v}{\partial \sqrt{c}}\right) \tag{II. 2}$$

The extrapolation of the apparent molar volume of electrolyte to infinite dilution and the expression of the concentration dependence of the apparent molar volume have been made by four major equations over a period of years – the Masson equation [177], the

Redlich-Meyer equation [178], the Owen-Brinkley equation [179], and the Pitzer equation [180]. According to Masson

$$\phi_V = \phi_V^0 + S_V^* \sqrt{c} \tag{II. 3a}$$

$$\phi_V = \phi_V^0 + S_V^* \sqrt{m} \tag{II. 3b}$$

Where, ϕ_V^0 is the apparent molar volume (equal to the partial molar volume) at infinite dilution and S_V^* the experimental slope. The majority of ϕ_V data in water [181] and nearly all ϕ_V data in non-aqueous [182-186] solvents have been extrapolated to infinite dilution through the use of equation (II. 3a) or (II.3b).

The temperature dependence of ϕ_V^0 in various solvents can be expressed by the general equation as follows:

$$\phi_V^0 = a_0 + a_1 T + a_2 T^2 + \cdots\cdots \tag{II. 4}$$

where a_0, a_1, a_2 are the coefficients of a particular electrolyte and T is the temperature in Kelvin.

The limiting apparent molar expansibilities (ϕ_E^0) can be obtained by the following equation:

$$\phi_E^0 = \left(\delta\phi_V^0 / \delta T \right)_P = a_1 + 2a_2 T \tag{II. 5}$$

The limiting apparent molar expansibilities (ϕ_E^0) change in magnitude with the change of temperature. During the past few years it has been emphasized by a number of workers that S_V^* is not the sole criterion for determining the structure-making or breaking tendency of any solute. Helper [187] developed a technique of examining the sign of $\left(\delta\phi_E^0 / \delta T \right)_P$ for the solute in terms of long–range structure-making and breaking capacity of the electrolytes in the mixed solvent systems. The general thermodynamic expression used is as follows:

$$\left(\delta\phi_E^0 / \delta T \right)_P = \left(\delta^2 \phi_V^0 / \delta T^2 \right)_P = 2a_2 \tag{II. 6}$$

If the sign of $\left(\delta\phi_E^0/\delta T\right)_P$ is positive or small negative the electrolyte is act as structure maker and when the sign of $\left(\delta\phi_E^0/\delta T\right)_P$ is negative, it is a structure breaker. Redlich and Meyer [178] have shown that an equation (II. 3a) or (II. 3b) cannot be any more than a limiting law where for a given solvent and temperature, the slope S_v^* should depend only upon the valence type. They suggested the equation:

$$\phi_v = \phi_v^0 + S_v\sqrt{c} + b_v c \tag{II. 7}$$

$$\text{Where } S_v = Kw^{3/2} \tag{II. 8}$$

S_V is the theoretical slope, based on molar concentration, including the valence factor where

$$w = 0.5\sum_i^j Y_i Z_i^2 \tag{II. 9}$$

$$\text{And } K = N^2 e^2 \left(\frac{8\pi}{1000 \in^3 RT}\right)^{1/2} \left[\left(\frac{\partial \ln \in}{\partial p}\right)_T - \frac{\beta}{3}\right] \tag{II. 10}$$

In equation (II. 10), K is the compressibility of the solvent and the other terms have their usual significance.

❖ *Limiting Ionic Partial Molar Volumes*

The individual partial ionic volumes provide information relevant to the general question of the structure near the ion, i.e., its solvation. The calculation of the ionic limiting partial molar volumes in organic solvents is, however, a difficult one. At present, however, most of the existing ionic limiting partial molar volumes in organic solvents were obtained by the application of methods originally developed for aqueous solutions to non-aqueous electrolyte solutions. In the last few years, the method suggested by Conway *et al.*[188] has been used more frequently. The authors used the method to determine the limiting partial molar volumes of the anion for a series of homologous tetra-alkyl ammonium chlorides, bromides and iodides in aqueous solution. They plotted the limiting partial molar volume $\phi_{v\,R_4NX}^0$, for a series of these salts with a halide ion in

common as a function of the formula weight of the cation, M_{R4N+} and obtained straight-lines for each series. Therefore, they suggested the following equation:

$$\phi^0_{v\,R_4NX} = bM_{R_4N^+} + \phi^0_{v\,X^-} \tag{II. 11}$$

The extrapolation to zero cationic formula weight gave the limiting partial molar volumes of the halide ions $\phi^0_{V\,X^-}$. Uosaki *et al.* [189] used this method for the separation of some literature values and of their own $\phi^0_{V\,R_4NX}$ values into ionic contributions in organic electrolyte solutions.

II.6.4 Viscosity

As fundamental and significant property of liquids is ***viscosity***, provide a lot of information on the structures and molecular interactions in liquid systems. Viscosity and volume are different types of properties of one liquid, and there is a certain relationship between them. So by measuring and studying them together, relatively more realistic and comprehensive information could be expected to be gained. Viscosity, one of the most important transport properties is used for the determination of ion-solvent interactions and studied extensively [190,191]. It is not a thermodynamic quantity, but of an electrolytic solution along with the physicochemical property, the partial molar volume $\phi^0_{v,2}$, gives a lot of information and insight regarding ion-solvent interactions and the nature of structures in the electrolytic solutions.

❖ *Viscosity of Pure Liquids and Liquid Mixtures*

Since the molecular motion in liquids is controlled by the influence of the neighboring molecules, the transport of momentum in liquids takes place, in sharp contrast with gases at ordinary pressures, not by the actual movement of molecules but by the intense influence of intermolecular force fields. It is this aspect of the mechanism of momentum transfer which forms the basis of the procedures for predicting the variations in the viscosity of liquids and liquid systems.

❖ *Viscosity of Electrolytic Solutions*

The viscosity relationships of electrolytic solutions are highly complicated. Because ion-ion and ion-solvent interactions are occurring in the solution and separation of the related forces is a difficult task. But, from careful analysis, vivid and valid conclusions can be drawn regarding the structure and the nature of the solvation of the particular system. As viscosity is a measure of the friction between adjacent, relatively moving parallel planes of the liquid, anything that increases or decreases the interaction between the planes will raise or lower the friction and thus, increase or decrease the viscosity. If large spheres are placed in the liquid, the planes will be keyed together in increasing the viscosity. Similarly, increase in the average degree of hydrogen bonding between the planes will increase the friction between the planes, thereby viscosity. An ion with a large rigid co-sphere for a structure-promoting ion will behave as a rigid sphere placed in the liquid and increase the inter-planar friction. Similarly, an ion increasing the degree of hydrogen bonding or the degree of correlation among the adjacent solvent molecules will increase the viscosity. Conversely, ions destroying correlation would decrease the viscosity. In 1905, Grüneisen [192] performed the first systematic measurement of viscosities of a number of electrolytic solutions over a wide range of concentrations. He noted non-linearity and negative curvature in the viscosity concentration curves irrespective of low or high concentrations. In 1929, Jones and Dole [193] suggested an empirical equation quantitatively correlating the relative viscosities of the electrolytes with molar concentrations (c):

$$\frac{\eta}{\eta_o} = \eta_r = 1 + A\sqrt{c} + Bc \tag{II. 12}$$

The above equation can be rearranged as:

$$\frac{\eta_r - 1}{\sqrt{c}} = A + B\sqrt{c} \tag{II. 13}$$

Where A and B are constants specific to ion-ion and ion-solvent interactions. The equation is applicable equally to aqueous and non-aqueous solvent systems where there is no ionic association and has been used extensively. The term $A\sqrt{c}$, originally ascribed to Grüneisen effect, arose from the long-range coulombic forces between the ions. The significance of the term had since then been realized due to the development Debye-

Hückel Theory [194] of inter-ionic attractions in 1923. The A -coefficient depends on the ion-ion interactions and can be calculated from interionic attraction theory [195-197] and is given by the Falkenhagen Vernon [197] equation:

$$A_{Theo} = \frac{0.2577 \Lambda_o}{\eta_o (\varepsilon T)^{0.5} \lambda_+^o \lambda_-^o} \left[1 - 0.6863 \left(\frac{\lambda_+^o \lambda_-^o}{\Lambda_o} \right)^2 \right] \quad (\text{II. 14})$$

Where the symbols have their usual significance. In very accurate work on aqueous solutions [131, 198], A-coefficient has been obtained by fitting η_r to equation (II. 13) and compared with the values calculated from equation (II. 14), the agreement was excellent. The accuracy achieved with partially aqueous solutions was however poorer [199]. A-coefficient suggesting that should be calculated from conductivity measurements. Crudden *etal*. [200] suggested that if association of the ions occurs to form an ion pair, the viscosity should be changed and thus analyzed by the equation:

$$\frac{\eta_r - 1 - A\sqrt{\alpha c}}{\alpha c} = B_i + B_P \left(\frac{1-\alpha}{\alpha} \right) \quad (\text{II. 15})$$

Where A, B_i and B_P are characteristic constants and α is the degree of dissociation of ion-pair. Thus, a plot of $\dfrac{\eta_r - 1 - A\sqrt{\alpha c}}{\alpha c}$ vs $\left(\dfrac{1-\alpha}{\alpha} \right)$, when extrapolated to $\left(\dfrac{1-\alpha}{\alpha} \right) = 0$ give the intercept B_i. However, for the most of the electrolytic solutions both aqueous and non-aqueous, the equation (II. 13) is valid up to 0.1 (M) [201,202] within experimental errors. At higher concentrations the extended Jones-Dole equation (II. 16), involving an additional coefficient D, originally used by Kaminsky [203], has been used by several workers [204,205] and is given below:

$$\frac{\eta}{\eta_o} = \eta_r = 1 + A\sqrt{c} + Bc + Dc^2 \quad (\text{II. 16})$$

The coefficient D cannot be evaluated properly and the significance of the constant is also not always meaningful and therefore, equation (II. 13) is used by the most of the workers.

The plots of $\dfrac{(\eta / \eta_o - 1)}{\sqrt{c}}$ against $\sqrt{c}$ for the electrolytes should give the value of A-coefficient. But sometimes, the values come out to be negative or considerably scatter

and also deviation from linearity occur [206-211]. Thus, instead of determining *A*-coefficient from the plots or by the least square method, the *A*-coefficient are generally calculated using Falkenhagen-Vernon equation (II. 14). *A*-coefficient should be zero for non-electrolytes. According to Jones and Dole, the *A*-coefficient probably represents the stiffening effect on the solution of the electric forces between the ions, which tend to maintain a space-lattice structure [193]. The sign of the *B*-coefficient may be either positive or negative which depends on the ions and the solvent. The *B*–coefficients are obtained as slopes of the straight lines using the least square method and intercepts equal to the *A*-coefficient.

The factors influencing *B*-coefficients are [208,209]:

(a) The destruction of the three dimensional structure of solvent molecules (i.e., structure breaking effect or depolymeriation effect) decreases η values.

(b) The solvent having high molal volume and low dielectric constant yield high *B*-values.

(c) Reduced *B*-values are obtained when the primary solution of ions is sterically hindered in high molal volume solvents or if either ion of a binary electrolyte cannot be specifically solvated.

❖ *Viscosities at Higher Concentration*

It had been found that the viscosity at high concentrations (1 M to saturation) can be represented by the empirical formula suggested by Andrade:

$$\eta = A \exp^{\frac{b}{T}} \tag{II. 17}$$

The several alternative formulations have been proposed for representing the results of viscosity measurements in the high concentration range [209-215] and the equation suggested by Angell [216,217] based on an extension of the free volume theory of transport phenomena in liquids and fused salts to ionic solutions is particularly noteworthy. The equation is:

$$\frac{1}{\eta} = A \exp\left[-\frac{K_1}{N_o - N}\right] \tag{II. 18}$$

Where N represents the concentration of the salt in eqv·litre^{-1}, A and K_1 are constants supposed to be independent of the salt composition and N_o is the hypothetical concentration at which the system becomes glass.

❖ *Division of B-Coefficient into Ionic Values*

The viscosity B-coefficients have been determined by a large number of workers in aqueous, mixed and non-aqueous solvents [218-248]. However, the B-coefficients as determined experimentally using the Jones-Dole equation, does not give any impression regarding ion-solvent interactions unless there is some way to identify the separate contribution of cations and anions to the total solute-solvent interaction. The division of B-values into ionic components is quite arbitrary and based on some assumptions, the validity of which may be questioned. The following methods have been used for the division of B-values in the ionic components:

[1]. Cox and Wolfenden [249] carried out the division on the assumption that B_{ion} values of Li$^+$ and IO$_3^-$ in LiIO$_3$ are proportional to the ionic volumes which are proportional to the third power of the ionic mobilities. The method of Gurney [250] and also of Kaminsky [203] is based on:

$$B_{K^+} = B_{Cl^-} \text{ (in water)} \qquad \text{(II. 19)}$$

The argument in favor of this assignment is based on the fact that the B-coefficients for KCl is very small and that the motilities' of K$^+$ and Cl$^-$ are very similar over the temperature range 288.15–318.15K. The assignment is supported from other thermodynamic properties. Nightingle [251], however preferred RbCl or CsCl to KCl from mobility considerations.

[2]. The method suggested by Desnoyers and Perron [204 is based on the assumption that the Et$_4$N$^+$ion in water is probably closest to be neither structure breaker not a structure maker. Thus, they suggest that it is possible to apply with a high degree of accuracy of the Einstein's equation [252],

$$B = 0.0025\overline{V}_o \qquad \text{(II. 20)}$$

and by having an accurate value of the partial molar volume of the ion, V_o, it is possible to calculate the value of 0.359 for $B_{Et_4N^+}$ in water at 298.15 K. Recently, Sacco *et al.* proposed the "*reference electrolytic*" method for the division of *B*-values.

Thus, for tetraphenylphosphonium tetraphenylborate in water, we have:

$$B_{BPh_4^-} = B_{PPh_4^+} = B_{BPh_4PPh_4} / 2 \tag{II. 21}$$

$B_{BPh_4PPh_4}$ (Scarcely soluble in water) has been obtained by the following method:

$$B_{BPh_4PPh_4} = B_{NaBPh_4} + B_{PPh_4Br} - B_{NaBr} \tag{II. 22}$$

The values obtained are in good agreement with those obtained by other methods. The criteria adopted for the separation of *B*-coefficients in non-aqueous solvents differ from those generally used in water. However, the methods are based on the equality of equivalent conductances of counter ions at infinite dilutions.

(a) Criss and Mastroianni assumed $B_{K^+} = B_{Cl^-}$ in ethanol based on equal mobilities of ions [253]. They also adopted $B^{25}_{Me_4N^+} = 0.25$ as the initial value for acetonitrile solutions.

(b) For acetonitrile solutions, Tuan and Fuoss [254] proposed the equality, as they thought that these ions have similar mobilities. However, according to Springer *et al.* [255] $\lambda^o_{25}(Bu_4N^+) = 61.4$ and $\lambda^o_{25}(Ph_4B^-) = 58.3$ in acetonitrile.

$$B_{Bu_4N^+} = B_{Ph_4B^-} \tag{II. 23}$$

(c) Gopal and Rastogi [219] resolved the *B*-coefficient in N-methyl propionamide solutions assuming that $B_{Et_4N^+} = B_{I^-}$ at all temperatures.

(d) In dimethyl sulphoxide, the division of *B*-coefficients were carried out by Yao and Bennion [207] assuming:

$$B_{[(i-pe)_3Bu_4N^+]} = B_{Ph_4B^-} = \frac{1}{2}B_{[(i-pe)_3BuNPh_4B]} \tag{II. 24}$$

at all temperatures.

It is apparent that almost all these methods are based on certain approximations and anomalous results may arise unless proper mathematical theory is developed to calculate B-values.

❖ *Thermodynamics of Viscous Flow*

Assuming viscous flow as a rate process, the viscosity (η) can be represented from Eyring's [256] approaches as:

$$\eta = Ae^{\frac{E_{vis}}{RT}} = \left(\frac{hN_A}{V}\right)e^{\frac{\Delta G^*}{RT}} = \left(\frac{hN_A}{V}\right)e^{\left(\frac{\Delta H^*}{RT} - \frac{\Delta S}{R}\right)}$$ (II. 25)

Where E_{vis} = the experimental entropy of activation determined from a plot of ln η against 1/T. ΔG^*, ΔH^* and ΔS^* are the free energy, enthalpy and entropy of activation, respectively.

Nightingale and Benck [257] dealt in the problem in a different way and calculated the thermodynamics of viscous flow of salts in aqueous solution with the help of the Jones-Dole equation (neglecting the $A\,c$ term). Thus, we have:

$$R\left[\frac{d\ln\eta}{d\left(\frac{1}{T}\right)}\right] = r\left[\frac{d\ln\eta_o}{d\left(\frac{1}{T}\right)}\right] + \frac{R}{1+Bc}\frac{d(1+Bc)}{d\left(\frac{1}{T}\right)}$$ (II. 26)

$$\Delta E^{\neq}_{\eta(So\ln)} = \Delta E^{\neq}_{\eta(Solv)} + \Delta E^{\neq}_V$$ (II. 27)

$\Delta E^{\neq}_V$ can be interpreted as the increase or decrease of the activation energies for viscous flow of the pure solvents due to the presence of ions, *i.e.*, the effective influence of the ions upon the viscous flow of the solvent molecules. Feakins *et al.* [258] have suggested an alternative formulation based on the transition state treatment of the relative viscosity of electrolytic solution. They suggested the following expression:

$$B = \frac{(\phi^0_{V,2} - \phi^0_{V,1})}{1000} + \phi^0_{V,2}\frac{(\Delta\mu_2^{0\neq} - \Delta\mu_1^{0\neq})}{1000RT}$$ (II. 28)

Where $\phi^0_{v,1}$ and $\phi^0_{v,2}$ are the partial molar volumes of the solvent and solute respectively and $\Delta\mu_2^{o\neq}$ is the contribution per mole of solute to the free energy of activation for viscous

flow of solution. $\Delta\mu_1^{o\neq}$ is the free energy of activation for viscous flow per mole of the solvent which is given by:

$$\Delta\mu_1^{o\neq} = \Delta G_1^{o\neq} = RT\ln(\eta_0\phi_{v,1}^0/hN_A) \tag{II. 29}$$

Further, if B is known at various temperatures, we can calculate the entropy and enthalpy of activation of viscous flow respectively from the following equations as given below:

$$\frac{d(\Delta\mu_2^{o\neq})}{dT} = -\Delta S_2^{o\neq} \tag{II. 30}$$

$$\Delta H_2^{o\neq} = \Delta\mu_2^{o\neq} + T\Delta S_2^{o\neq} \tag{II. 31}$$

❖ *Effects of Shape and Size*

Stokes and Mills have dealt in the aspect of shape and size extensively. The ions in solution can be regarded to be rigid spheres suspended in continuum. The hydrodynamic treatment presented by Einstein [252] leads to the equation:

$$\frac{\eta}{\eta_0} = 1 + 2.5\phi \tag{II. 32}$$

Where ϕ is the volume fraction occupied by the particles. Modifications of the equation have been proposed by (i) Sinha [259] on the basis of departures from spherical shape and (ii) Vand on the basis of dependence of the flow patterns around the neighboring particles at higher concentrations. However, considering the different aspects of the problem, spherical shapes have been assumed for electrolytes having hydrated ions of large effective size (particularly polyvalent monatomic cations). Thus, we have from equation (II. 32):

$$2.5\phi = A\sqrt{c} + Bc \tag{II. 33}$$

Since $A\sqrt{c}$ term can be neglected in comparison with Bc and $\phi = c\phi_{v,1}^0$ where $\phi_{v,1}^0$ is the partial molar volume of the ion, we get:

$$2.5\phi_{v,1}^0 = B \tag{II. 34}$$

In the ideal case, the B-coefficient is a linear function of partial molar volume of the solute, $\phi_{v,1}^0$ with slope to 2.5. Thus, $B_\pm$ can be equated to:

$$B_\pm = 2.5\phi_\pm^0 = \frac{2.5 \times 4}{3}\frac{\left(\pi R_\pm^3 N\right)}{1000} \tag{II. 35}$$

assuming that the ions behave like rigid spheres with an effective radii, $R_\pm$ moving in a continuum. $R_\pm$, calculated using the equation (II. 35) should be close to crystallographic radii or corrected Stoke's radii if the ions are scarcely solvated and behave as spherical entities. But, in general, $R_\pm$ values of the ions are higher than the crystallographic radii indicating appreciable solvation.

The solvation number (n_b) can be easily calculated by comparing the Jones-Dole equation with the Einstein's equation:

$$B_\pm = \frac{2.5}{1000\left(\phi_i + n_b\phi_s\right)} \tag{II. 36}$$

Where ϕ_i is the molar volume of the base ion and ϕ_s, the molar volume of the solvent. The equation (II. 36) has been used by a number of workers to study the nature of solvation and solvation number.

❖ *Viscosity of Non-Electrolytic Solutions*

The equations of Vand [260], Thomas [261], and Moulik [262-264] proposed mainly to account for the viscosity of the concentrated solutions of bigger spherical particles have been also found to correlate the mixture viscosities of the usual nonelectrolytes. These equations are:

Vand equation: $$\ln \eta_r = \frac{\alpha}{1-Q} = \frac{2.5V_h c}{1-QV_h c} \tag{II. 37}$$

Thomas equation: $$\eta_r = 1 + 2.5V_h + 10.05cV_h^2 c \tag{II. 38}$$

Moulik equation: $$\eta^2 = I + Mc^2 \tag{II. 39}$$

Where η_r is the relative viscosity, a is constant depending on axial ratios of the particles, Q is the interaction constant, V_h is the molar volume of the solute including rigidly held

solvent molecules due to hydration, c is the molar concentration of the solutes; I and M are constants. The viscosity equation proposed by Eyring and coworkers for pure liquids on the basis of pure significant liquid structures theory, can be extended to predict the viscosity of mixed liquids also. The final expression for the liquid mixtures takes the following form:

$$\eta_m = \frac{6N_Ah}{\sqrt{2}r_m(V_m - V_{Sm})}\left[\sum_i^n\left\{1-\exp\left(\frac{-\theta_i}{T}\right)\right\}^{-x_i}\right]\exp\left[\frac{a_mE_{Sm}V_{Sm}}{RT(V_m - V_{Sm})}\right]$$
$$+\frac{V_m - V_{Sm}}{V_m}\left[\sum_i^n\frac{2}{3d_i^2}\left(\frac{m_ikT}{\pi^3}\right)^{\frac{1}{2}}x_i\right]$$

$$(\text{II. 40})$$

Where n is 2 for binary and 3 for ternary liquid mixtures. The mixture parameters, r_m, E_{Sm}, V_m, V_{Sm} and a_m were calculated from the corresponding pure component parameters by using the following relations:

$$r_m = \sum_i^n x_i^2 r_i + \sum_{i\neq j}2x_ix_jx_{ij} \qquad (\text{II. 41})$$

$$E_{Sm} = \sum_i^n x_i^2 E_{Si} + \sum_{i\neq j}2x_ix_jE_{Sij} \qquad (\text{II. 42})$$

$$V_m = \sum_i^n x_iV_i \qquad V_{Sm} = \sum_i^n x_iV_{Si} \qquad a_m = \sum_i^n x_ia_i \qquad (\text{II. 43})$$

$$r_{ij} = (r_ir_j)^{\frac{1}{2}} \text{ and } E_{Sij} = (E_{Si}E_{Sj})^{\frac{1}{2}} \qquad (\text{II. 44})$$

$$\theta = \frac{h}{\kappa 2\pi}\left(\frac{b}{m}\right)^{\frac{1}{2}} \qquad (\text{II. 45})$$

$$b = 2Z\varepsilon\left[22.106\left(\frac{N_A\sigma^2}{V_s}\right)^4 - 10.559\left(\frac{N_A\sigma^3}{V_s}\right)^2\right]\frac{1}{\sqrt{2}\sigma^2}\left(\frac{N_A\sigma^3}{V_s}\right)^{\frac{2}{3}} \qquad (\text{II. 46})$$

Where σ and ε_r are Lennard-Jones potential parameters and the other symbols have their usual significance.

❖ Viscosity Deviation

Viscosity of liquid mixtures can also provide information for the elucidation of the fundamental behavior of liquid mixtures, aid in the correlation of mixture viscosities with those of pure components, and may provide a basis for the selection of physico-chemical methods of analysis. Quantitatively, as per the absolute reaction rates theory [265], the deviations in viscosities ($\Delta\eta$) from the ideal mixture values can be calculated as:

$$\Delta\eta = \eta - \sum_{i=1}^{j}(x_i\eta_i) \tag{47}$$

Where η is the dynamic viscosities of the mixture and $x_i\eta_i$ are the mole fraction and viscosity of ith component in the mixture, respectively.

❖ Gibbs Excess Energy of Activation for Viscous Flow

Quantitatively, the Gibbs excess energy of activation for viscous flow ΔG^* can be calculated as [266]:

$$\Delta G^E = RT\left[\ln\eta V - \sum_{i=1}^{j}(x_i\ln\eta_iV_i)\right] \tag{II. 48}$$

Where η and V are the viscosity and molar volume of the mixture; η_i and V_i are the viscosity and molar volume of ith pure component, respectively.

II.6.5 Conductance

One of the most precise and direct technique available to determine the extent of the dissociation constants of electrolytes in aqueous, mixed and non-aqueous solvents is the "*conductimetric method.*" Conductance data in conjunction with viscosity measurements, gives much information regarding ion-ion and ion-solvent interaction.

ethanol
No conductivity

KCl
High conductivity

acetic acid solution
Low conductivity
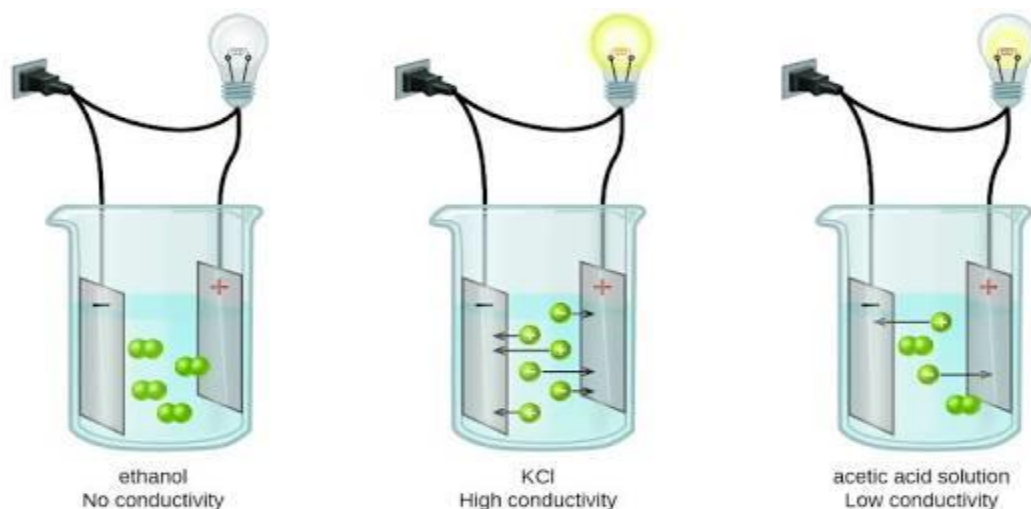
Dissolved Ions Conduct Electricity

The studies of conductance measurements were pursued vigorously during the last five decades, both theoretically and experimentally and a number of important theoretical equations have been derived. We shall dwell briefly on some of these aspects in relation to the studies in aqueous, non-aqueous, pure and mixed solvents. The successful application of the Debye-Hückel theory of interionic attraction was made by Onsager [267], to derive the Kohlrausch's equation representing the molar conductance of an electrolyte. For solutions of a single symmetrical electrolyte the equation is given by:

$$\Lambda = \Lambda_o - S\sqrt{c} \qquad (II. 49)$$

Where,

$$S = \alpha\Lambda_o + \beta \qquad (II. 50)$$

$$\alpha = \frac{(z^2)k}{3(2+\sqrt{2})\varepsilon_r kT\sqrt{c}} = \frac{82.406\times10^4\,z^3}{(\varepsilon_r T)^{\frac{3}{2}}} \qquad (II. 51a)$$

$$\beta = \frac{z^2 eFk}{3\pi\eta\sqrt{c}} = \frac{82.487 z^3}{\eta\sqrt{\varepsilon_r T}} \qquad (II. 51b)$$

The equation has no explanation for the short-range interactions and also of shape or size of the ions in solution. The ions were regarded as rigid charged spheres in an electrostatic and hydrodynamic continuum, i.e., the solvent [268]. In the subsequent years, Pitts (1953) [269] and Fuoss and Onsager (1957) [270] individually worked out the solution of the problem of electrolytic conductance accounting for both long-range

and short-range interactions. The conductance values at infinite dilution (Λ_o) are different for two different theory (Fuoss-Onsager theory and Pitt's theory) and the derivation of the Fuoss-Onsager equation was questioned [271,272]. Fuoss and Hsia [273,274] further modified the original Fuoss-Onsager equation who recalculated the relaxation field, retaining the terms which had previously been neglected.

The results of conductance theories can be expressed in a general form:

$$\Lambda = \frac{\Lambda_o - \alpha \Lambda_o \sqrt{c}}{(1+\kappa\alpha)} \left(\frac{1+\kappa\alpha}{\sqrt{2}} \right) - \frac{\beta\sqrt{c}}{(1+\kappa\alpha)} + G(\kappa\alpha) \tag{II. 52}$$

Where $G(\kappa\alpha)$ is a complicated function of the variable. The simplified form:

$$\Lambda = \Lambda_o - S\sqrt{c} + Ec\ln c + J_1 c + J_2 \sqrt[3]{c} \tag{II. 53}$$

However, it has been found that these equations have certain limitations, in some cases it fails to fit experimental data. Some of these results have been discussed elaborately by Fernandez-Prini [275]. Further correction of the equation (II. 53) was made by Fuoss and Accascin. They took into consideration the change in the viscosity of the solutions and assumed the validity of Walden's rule. The new equation becomes:

$$\Lambda = \Lambda_o - S\sqrt{c} + Ec\ln c + J_1 c + J_2 \sqrt[3]{c} - F\Lambda c \tag{II. 54}$$

Where,
$$Fc = \frac{4\pi R^3 N_A}{3} \tag{II. 55}$$

In most cases, however, J_2 is made zero but this leads to a systematic deviation of the experimental data from the theoretical equations. It has been observed that Pitt's equation gives better fit to the experimental data in aqueous solutions [276].

❖ *Ionic Association*

The equation (II. 54) successfully represents the behavior of completely dissociated electrolytes. The plot of Λ against $\sqrt{c}$ (limiting Onsager equation) is used to assign the dissociation or association of electrolytes. The electrolyte may be regarded as completely dissociated when positive deviation occurs $(\Lambda_{o\,exp} \rangle \Lambda_{o\,theo})$ but if negative deviation $(\Lambda_{o\,exp} \langle \Lambda_{o\,theo})$ or positive deviation from the Onsager limiting tangent

$(\alpha \Lambda_o + \beta)$ occurs, the electrolyte may be regarded to be associated. Here the electrostatic interactions are large so as to cause association between cations and anions. The difference in $\Lambda_{o\,exp}$ and $\Lambda_{o\,theo}$ would be considerable with increasing association [277].

Conductance measurements help us to determine the values of the ion-pair association constant, K_A for the process:

$$M^{z+} + A^{z-} = MA \tag{II. 56}$$

$$K_A = \frac{(1-\alpha)}{\alpha^2 c \gamma_\pm^2} \tag{II. 57}$$

$$\alpha = 1 - \alpha^2 K_A c \gamma_\pm^2 \tag{II. 58}$$

Where $\gamma_\pm$ is the mean activity coefficient of the free ions at concentration αc.

For strongly associated electrolytes, the constant K_A and Λ_o have been determined using Fuoss-Kraus equation [278] or Shedlovsky's equation [279].

$$\frac{T(z)}{\Lambda} = \frac{1}{\Lambda_o} + \frac{K_A}{\Lambda_o^2} \cdot \frac{c \gamma_\pm^2 \Lambda}{T(z)} \tag{II. 59}$$

Where $T(z) = F(z)$ (Fuoss-Kraus method) and $1/T(z)=S(z)$ (Shedlovsky's method).

$$F(z) = 1 - z(1 - z(1 - ...)^{-\frac{1}{2}})^{-\frac{1}{2}} \tag{II. 96a}$$

and $\qquad \frac{1}{T(z)} = S(z) = 1 + z + \frac{z^2}{2} + \frac{z^3}{8} + ... \tag{II. 96b}$

A plot of $T(z)/\Lambda$ against $c\gamma_\pm^2 \Lambda / T(z)$ should be a straight line having $1/\Lambda_o$ for its intercept and K_A / Λ_o^2 for its slope. The values of Λ_o and K_A obtained from equation (II. 95) show uncertainty when K_A is large.

The Fuoss-Hsia [273] conductance equation for associated electrolytes is given by:

$$\Lambda = \Lambda_o - S\sqrt{\alpha c} + E(\alpha c)\ln(\alpha c) + J_1(\alpha c) - J_2(\alpha c)^{\frac{3}{2}} - K_A \Lambda \gamma_\pm^2 (\alpha c) \tag{II. 60}$$

The equation was modified by Justice [280]. The conductance of symmetrical electrolytes in dilute solutions can be represented by the equations:

$$\Lambda = \alpha(\Lambda_o - S\sqrt{\alpha c} + E(\alpha c)\ln(\alpha c) + J_1 R(\alpha c) - J_2 R(\alpha c)^{\frac{3}{2}} \tag{II. 61}$$

$$\frac{(1-\alpha)}{\alpha^2 c\gamma_\pm^2} = K_A \tag{II. 62}$$

$$\ln \gamma_\pm = \frac{-k\sqrt{q}}{(1+kR\sqrt{\alpha c})} \tag{II. 63}$$

The conductance parameters are obtained from a least square treatment after setting, $R = q = \dfrac{e^2}{2\varepsilon kT}$ (Bjerrum's critical distance).

According to Justice the method of fixing the J-coefficient by setting, $R = q$ clearly permits a better value of K_A to be obtained. Since the equation (II. 61) is a series expansion truncated at the $c^{3/2}$ term, it would be preferable that the resulting errors be absorbed as must as possible by J_2 rather than by K_A, whose theoretical interest is greater as it contains the information concerning short-range cation-anion interaction. From the experimental values of the association constant K_A, one can use two methods in order to determine the distance of closest approach, 'a', of two free ions to form an ion-pair. The following equation has been proposed by Fuoss [281];

$$K_A = \frac{4\pi N_A\alpha^3}{3000} \exp\left(\frac{e^2}{\alpha\varepsilon kT}\right) \tag{II. 64}$$

In some cases, the magnitude of K_A was too small to permit a calculation of a. The distance parameter was finally determined from the more general equation due to Bjerrum [282].

$$K_A = \frac{4\pi N_A\alpha}{1000} \int\limits_{r=a}^{r=q} r^2 \exp\left(\frac{z^2 e^2}{r\varepsilon kT}\right) dr \tag{II. 65}$$

❖ *Ion Size Parameter and Ionic Association*

For plotting, equation (II. 54) can be rearranged to the 'Λ' function as:

$$\Lambda_1 = \Lambda + S\sqrt{c} - Ec\ln c = \Lambda_o + J_1 c + J_2 \sqrt[3]{c} = \Lambda_o + J_1 c \tag{II. 66}$$
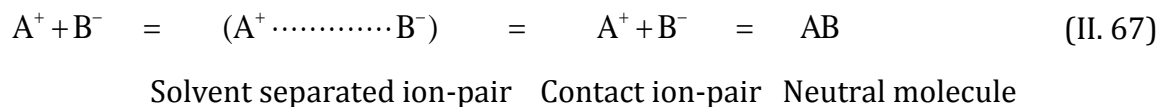
with J_2 term omitted.

Thus, a plot of Λ_o vs. c gives a straight line with Λ_o as intercept and J_1 as slope and 'a' values can be calculated from J_1 values. The 'a' values obtained by this method for

DMSO were much smaller than would be expected from sums of crystallographic radii. One of the reasons attributed to it is that ion-solvent interactions are not included in the continuum theory on which the conductance equations are based. The inclusion of dielectric saturation results in an increase in 'a' values (much in conformity with the crystallographic radii) of alkali metal salts (having ions of high surface charge density) in sulpholane. The viscosity correction leads to a larger value of 'a' [283] but the agreement is still poor. However, little of real physical significance may be attached to the distance of closest approach derived from J_1 [284]. Fuoss [285] in 1975 proposed a new conductance equation. Latter he subsequently put forward another conductance equation in 1978 replacing the old one as suggested by Fuoss and co-workers.

He classified the ions of electrolytic solutions in one of the three categories.

(i)   Ions finding an ion of opposite charge in the first shell of nearest neighbours (contact pairs) with $r_{ij} = a$. The nearest neighbours to a contact pair are the solvent molecules forming a cage around the pairs.

(ii)   Ions with overlapping Gurney's co-spheres (solvent separated pairs). For them $r_{ij} = a + ns$, where n is generally 1 but may be 2, 3 etc.; 's' is the diameter of sphere corresponding to the average volume (actual plus free) per solvent molecule.

(iii) Ions finding no other unpaired ion in a surrounding sphere of radius R, the diameter of the co-sphere (unpaired ions).

Thermal motions and interionic forces establish a steady state, represented by the following equilibria:

$$A^+ + B^- \quad = \quad (A^+ \cdots\cdots\cdots\cdots B^-) \quad = \quad A^+ + B^- \quad = \quad AB \qquad (II.\ 67)$$

Solvent separated ion-pair   Contact ion-pair   Neutral molecule

Contact pairs of ionogens may rearrange to neutral molecules $A^+B^- = AB$, e.g., H_3O^+ and CH_3COO^-. Let γ be the fraction of solute present as unpaired ($r > R$) ions. If $c\gamma$ is the concentration of unpaired ion and α is the fraction of paired ions($r \leq R$), then the concentration of unpaired ion and $c(1-\alpha)(1-\gamma)$ and that of contact pair is $\alpha c(1-\gamma)$.

The equation constants for Eq. (II. 67) are:

$$K_R = \frac{(1-\alpha)(1-\gamma)}{c\gamma^2 f^2}$$

(II. 68)

$$K_S = \frac{\alpha}{1-\alpha} = \exp\left(\frac{-E_S}{kT}\right) = e^{-\varepsilon}$$

(II. 69)

Where K_R refer to the formation and separation of solvent separated pairs by diffusion in and out of spheres of diameter R around cations and can be calculated by continuum theory; K_S is the constant describing the specific short-range ion-solvent and ion-ion interactions by which contact pairs form and dissociate. E_S is the difference in energy between a pair in the states $(r = R)$ and $(r = a)$; ε is E_S measured in units of kT.

Now,

$$(1-\alpha) = \frac{1}{1+K_S}$$

(II. 70)

and the conductometric pairing constant is given by:

$$K_A = \frac{(1-\alpha)}{c\gamma^2 f^2} = \frac{K_R}{1-\alpha} = K_R(1+K_S)$$

(II. 71)

The equation determines the concentration, $c\gamma$ of active ions that produce long-range interionic effects.

The various patterns can be reproduced by theoretical fractions in the form:

$$\Lambda = p\left[\Lambda_o\left(\frac{1+\Delta X}{X}\right) + \Delta\Lambda_e\right] = p\left[\Lambda_o(1+R_X) + E_L\right]$$

(II. 72)

Which is a three parameter equation $\Lambda = \Lambda(c, \Lambda_o, R, E_S), \Delta X / X$ (the relaxation field) and $\Delta\Lambda_e$ (the electrophoretic counter current) are long range effects due to electrostatic interionic forces and p is the fraction of Gurney co-sphere.

The parameters K_R (or E_S) is a catch-all for all short range effects:

$$p = 1 - \alpha(1-\gamma)$$

(II. 73)

In case of ionogens or for ionophores in solvents of low dielectric constant, α is very near to unity $(-E_S/kT) >> 1$ and the equation becomes:

$$\Lambda = \gamma\left[\Lambda_o\left(\frac{1+\Delta X}{X}\right)\right] + \Delta\Lambda_e$$

(II. 74)

The equilibrium constant for the effective reaction, $A^+ + B^- + AB$, is then

$$K_A = \frac{(1-\gamma)}{c\gamma^2 f^2} \approx K_R K_S \qquad (\text{II. 75})$$

as $K_S \gg 1$. The parameters and the variables are related by the set of equations:

$$\gamma = 1 - \frac{K_R c\gamma^2 f^2}{(1-\alpha)} \qquad (\text{II. 76})$$

$$K_R = \left(\frac{4\pi N_A R^3}{3000}\right) \exp\left(\frac{\beta}{R}\right) \qquad (\text{II. 77})$$

$$-\ln f = \frac{\beta\kappa}{2(1+\kappa R)}, \quad \beta = \frac{e^2}{\varepsilon\kappa T} \qquad (\text{II. 78})$$

$$\kappa^2 = 8\pi\beta\gamma\eta = \frac{\pi\beta N_A \gamma c}{125} \qquad (\text{II. 79})$$

$$-\varepsilon = \ln\left[\frac{\alpha}{1-\alpha}\right] \qquad (\text{II. 80})$$

The details of the calculations are presented in the 1978 paper [285]. The shortcomings of the previous equations have been rectified in the present equation that is also more general than the previous equations and can be used for higher concentrations (0.1 N in aqueous solutions).

❖ *Extension of Fuoss Conductance Equation*

As Fuoss 1978 conductance equation contained a boundary condition error [286,287], Fuoss introduced a slight modification to his model [288,289]. According to him, the ion pairs (ion approaching with their Gurney co-sphere) are divided into two categories- contact pairs (with no contribution to conductance) and solvent separated ion pairs (which can only contribute to the net transfer of charge). To rectify the boundary errors contained in Fuoss 1978 conductance equation, Lee-Wheaton [290(a)] in the same year described in the derivation of a new conductance equation, based on the Gurney co-sphere model and henceforth the new equation is referred to as the Lee-Wheaton equation [290(b)].The conductance data were analyzed by means of the Lee-Wheaton conductance equation [291] in the form:

$$\Lambda = \alpha_i \left[\begin{array}{l} \Lambda_o \left\{ 1 + C_1\beta\kappa + C_2(\beta\kappa)^2 + C_3(\beta\kappa)^3 \right\} \\ -\dfrac{\rho\kappa}{1+\kappa R} \left\{ 1 + C_4\beta\kappa + C_5(\beta\kappa)^2 + \dfrac{\kappa R}{12} \right\} \end{array} \right] \qquad (II.\,81)$$

The mass action law association [292] is

$$K_{A,c} = \frac{(1-\alpha_i)\gamma_A}{\alpha_i^2 c_i \gamma_\pm^2} \qquad (II.\,82)$$

and the equation for the mean ionic activity coefficient:

$$\gamma_\pm = \exp\left[-\frac{q\kappa}{1+\kappa R} \right] \qquad (II.\,83)$$

Where C_1 to C_5 are least square fitting coefficients as described by Pethybridge and Taba [293], Λ_o is the limiting molar conductivity, $K_{A,c}$, is the association constant, α_i is the dissociation degree, q is the Bjerrum parameter and γ the activity coefficient and $\beta = 2q$. The distance parameter R is the least distance that two free ions can approach before they merge into ion pair. The Debye parameter κ, the Bjerrum parameter q and ρ are defined by the expressions [293]:

$$\kappa = 16000\pi N_A q c_i \alpha_i \qquad (II.\,84)$$

$$q = \frac{e^2}{8\varepsilon_o \varepsilon_r \kappa T} \qquad (II.\,85)$$

$$\rho = \frac{F\,e}{299.79 \times 3\pi\eta} \qquad (II.\,86)$$

Where the symbols have their usual significance [294].

The equation (II. 78) was resolved by an iterative procedure. For a definite R value the initial value of Λ_o and $K_{A,c}$, were obtained by the Kraus-Bray method [295]. The parameter Λ_o and $K_{A,c}$, were made to approach gradually their best values by a sequence of alternating linearization and least squares optimizations by the Gauss-Siedel method [296] until satisfying the criterion for convergence. The best value of a parameter is the one when equation (II. 78) is best fitted to the experimental data corresponding to

minimum standard deviation (σ_A) for a sequence of predetermined R value and standard deviation (σ_A) was calculated by the following equation:

$$\sigma_A^2 = \sum_{i=1}^{n} \frac{[\Lambda_{i(calc)} - \Lambda_{i(obs)}]^2}{n - m} \qquad \text{(II. 87)}$$

Where n is the number of experimental points and m is the number of fitting parameters. The conductance data were analyzed by fixing the distance of closest approach R with two parameter fit ($m=2$). For the electrolytes with no significant minima observed in the σ_A versus R curves, the R values were arbitrarily preset at the centre to centre distance of solvent-separated pair:

$$R = a + d \qquad \text{(II. 88)}$$

where $a = r_c^+ + r_c^-$, i.e., the sum of the crystallographic radii of the cation and anion and d is the average distance corresponding to the side of a cell occupied by a solvent molecule. The definitions of d and related terms are described in the literature [297]. R was generally varied by a step 0.1 Å and the iterative process was continued with equation (78).

❖ *Limiting Ionic Conductance*

The limiting ionic conductance of an electrolyte can be easily determined from the theoretical equations and experimental observations. At infinite dilutions, the motion of an ion is limited solely by the interactions with the surroundings solvent molecules as the ions are infinitely apart. Under these conditions, the validity of Kohlrausch's law of independent migration of ions is almost axiomatic. Thus:

$$\Lambda_0 = \lambda_o^+ + \lambda_o^- \qquad \text{(II. 89a)}$$

At present, limiting ionic conductance is the only function which can be divided into ionic components using experimentally determined transport number of ions, i.e.

$$\lambda_o^+ = t_+ \Lambda_0 \quad \text{and} \quad \lambda_o^- = t_- \Lambda_0 \qquad \text{(II. 89b)}$$

Thus, from accurate value of λ_0 of ions it is possible to separate the contributions due to cations and anions in the solute-solvent interactions. However, accurate transference number determinations are limited to few solvents only.

In the absence of experimentally measured transference numbers it would be useful to develop indirect methods to obtain the ionic limiting equivalent conductances in solvents for which experimental transference numbers are not yet available.

The method has been summarized by Krumgalz [298] and some important points are mentioned as follows:

(i)   Walden equation [299]

$$(\lambda_o^\pm)_{\text{water}}^{25} \cdot \eta_{\text{o,water}} = (\lambda_o^\pm)_{\text{acetone}}^{25} \cdot \eta_{\text{o,acetone}} \tag{II. 90}$$

(ii) $(\lambda_{\text{o,pic}} \cdot \eta_o) = 0.267, \quad \lambda_{\text{o,Et}_4\text{N}^+} \cdot \eta_o = 0.269\,[299,300] \tag{II.91}$

        based on $\Lambda_{\text{o,Et}_4\text{N}_{\text{pic}}} = 0.563$

Walden considered the products to be independent of temperature and solvent. However, the $\Lambda_{\text{o,Et}_4\text{N}_{\text{pic}}}$ values used by Walden were found to differ considerably from the data of subsequent more precise studies and the values of (ii) are considerably different for different solvents.

(iii) $\lambda_o^{25}(\text{Bu}_4\text{N}^+) = \lambda_o^{25}(\text{Ph}_4\text{B}^-) \tag{II. 92}$

The equality holds good in nitrobenzene and in mixture with CCl₄ but not realized in methanol, acetonitrile and nitromethane.

(iv) $\lambda_o^{25}(\text{Bu}_4\text{N}^+) = \lambda_o^{25}(\text{Bu}_4\text{B}^-)\,[301] \tag{II. 93}$

The method appears to be sound as the negative charge on boron in the Bu₄B⁻ ion is completely shielded by four inert butyl groups as in the Bu₄N⁺ ion while this phenomenon was not observed in case of Ph₄B⁻.

(v)   The equation suggested by Gill [302] is:

$$\lambda_o^{25}(R_4N^+) = \frac{zF^2}{6\pi N_A\eta_o[r_i - (0.0103\varepsilon_o + r_y)]} \tag{II. 94}$$

Where Z and r_i are the charge and crystallographic radius of proper ion, respectively: η_o and ε_o are solvent viscosity and dielectric constant of the medium, respectively; $r_y =$ adjustable parameter taken equal to 0.85 Å and 1.13 Å for dipolar non-associated solvents and for hydrogen bonded and other associated solvents respectively.

However, large discrepancies were observed between the experimental and calculated values [298(a)]. In a paper [298(b)], Krumgalz examined the Gill's approach more critically using conductance data in many solvents and found the method reliable in three solvents e.g. butan-1-ol, acetonitrile and nitromethane.

$$\text{(vi)} \qquad \lambda_o^{25}\left[(i-Am)_3 \, BuN^+\right] = \lambda_o^{25}\left(Ph_4B^-\right) \qquad\qquad \text{(II. 95)}$$

It has been found from transference number measurements that the $\lambda_o^{25}\left[(i-Am)_3 \, BuN^+\right]$ and $\lambda_o^{25}\left(Ph_4B^-\right)$ values differ from one another by 1%.

$$\text{(vii)} \qquad \lambda_o^{25}\left(Ph_4B^-\right) = 1.01\lambda_o^{25}\left[(i-Am)_4 \, B^-\right] \qquad\qquad \text{(II. 96)}$$

Krumgalz suggested a method for determining the limiting ion conductance in organic solvents. Large tetraalkyl (aryl) onium ions are not solvated in organic solvents as there is extremely weak electrostatic interactions between solvent molecules and the large ions having low surface charge density and this phenomenon can be utilized as a suitable model for distributing Λ_o values into ionic components for non-aqueous electrolytic solutions.

Considering the motion of solvated ion in an electrostatic field as a whole, it is possible to calculate the radius of the moving particle by the Stokes equation:

$$r_S = \frac{|z| F^2}{A\pi\eta_o\lambda_o^{\pm}} \qquad\qquad \text{(II. 97)}$$

Where A is a coefficient varying from 6 (in the case of perfect sticking) to 4 (in case of perfect slipping). Since the r_S values, the real dimension of the non-solvated tetraalkyl (aryl) onium ions must be constant, we have:

$$\lambda_o^{\pm}\eta_0 = \text{constant} \qquad\qquad \text{(II. 98)}$$

This relation has been verified using $\lambda_o^{\pm}$ values determined with precise transference numbers. The relationship can be well utilized to determine $\lambda_o^{\pm}$ of ions in other organic solvents from the determined Λ_o values

❖ *Solvation*

Several types of interactions exist between the ions in solutions. Due to this types of interactions solvent molecules orient themselves towards the ion. The number of solvent molecules that are involved in the solvation of the ion is called *solvation number*. If the solvent is water, this is called *hydration number*. Solvation region can be classified as primary and secondary solvation regions. Here we are concerned with the primary solvation region. The primary solvation number is defined as the number of solvent molecules which surrender their own translational freedom and remain with the ion, tightly bound, as it moves around, or the number of solvent molecules which are aligned in the force field of the ion.

If the limiting conductance of the ion *i* of charge Z_i is known, the effective radius of the solvated ion can be determined from Stokes' law. The volume of the solvation shell is given by the equation.

$$V_S = \left(\frac{4\pi}{3}\right)\left(r_S^3 - r_C^3\right) \qquad (II.\ 99)$$

Where r_c is the crystallographic radius of the ion. The solvation number n_s would then be obtained from

$$n_S = \frac{V_S}{V_0} \qquad (II.\ 100)$$

Assuming Stokes' relation to hold well, the ionic solvated volume can be obtained, because of the packing effects [303], from

$$V_S^o = 4.35 r_S^3 \qquad (II.\ 101)$$

Where V_s^o is expressed in mol·lit^{-1} and r_s in angstroms. However, this method is not applicable to ions of medium size though a number of empirical and theoretical corrections [304-307] have been suggested in order to apply it to most of the ions.

❖ *Stokes' Law and Walden's Rule*

According to Stokes' law the limiting ionic Walden product (the product of the limiting ionic conductance and solvent viscosity) for any singly charged, spherical ion is as function only of the ionic radius and thus, under normal conditions, is constant. For of a spherical ion having radius R_i with movements in a solvent of dielectric field, the limiting conductances λ_o^i can be written, according to Stokes' hydrodynamics, as

$$\lambda_o^i = \frac{|z_i e| eF}{6\pi \eta_o R_i} = \frac{0.819|z_i|}{\eta_o R_i} \qquad \text{(II. 102)}$$

Where η_o = macroscopic viscosity by the solvent in poise, R_i is in angstroms. If the radius R_i is assumed to be the same in every organic solvent, as would be the case, in case of bulky organic ions, we get:

$$\lambda_o^i \eta_o = \frac{0.819 z_i}{R_i} = cons\,tant \qquad \text{(II. 103)}$$

This is known as the Walden rule [308]. The effective radii obtained using this equation can be used to estimate the solvation numbers. However, Stokes' radii failed to give the effective size of the solvated ions for small ions.

Robinson and Stokes [309], Nightingale [251] and others [310-312] have suggested a method of correcting the radii. The tetraalkylammonium ions were assumed to be not solvated and by plotting the Stokes' radii against the crystal radii of those large ions, a calibration curve was obtained for each solvent. However, the experimental results indicate that the method is incorrect as the method is based on the wrong assumption of the invariance of Walden's product with temperature. The idea of microscopic viscosity [313] was invoked without much success [314,315] but it has been found that:

$$\lambda_o^i \eta_0 = \text{constant} \qquad \text{(II. 104)}$$

where i is usually 0.7 for alkali metal or halide ions and $i = 1$ for the large ions [316,317] Gill [302] has pointed out the inapplicability of the Zwanzig theory [318] of dielectric friction for some ions in non-aqueous and mixed solvents and has proposed an empirical modification of Stokes' Law accounting for the dielectric friction effect quantitatively and predicts actual solvated radii of ions in solution. This equation can be written as:

$$r_i = \frac{|z|F^2}{6\pi N_A \eta_o \lambda_o^i} + 0.0103D + r_y \tag{II. 105}$$

Where r_i is the actual solvated radius of the ion in solution and r_y is an empirical constant dependent on the nature of the solvent [302,318].

The dependence of Walden product on the dielectric constant led Fuoss to consider the effect of the electrostatic forces on the hydrodynamics of the system. Considering the excess frictional resistance caused by the dielectric relaxation in the solvent caused by ionic motion, Fuoss proposed the relation:

$$\lambda_o^i = \frac{Fe|z_i|}{6\pi R_\infty}\left(\frac{1+A}{\varepsilon R_\infty^2}\right) \tag{II. 106}$$

or,

$$R_i = R_\infty + \frac{R}{\infty} \tag{II. 107}$$

Where R_∞ is the hydrodynamic radius of the ion in a hypothetical medium of dielectric constant where all electrostatic forces vanish and A is an empirical constant.

Boyd [305] gave the expression:

$$\lambda_o^i = \frac{Fe|z_i|}{6\pi\eta_o r_i\left[\left(1+\frac{2}{27}\pi\eta_o\right)\cdot\left(\frac{z_i^2 e^2 \tau}{r_i^4 \varepsilon_o}\right)\right]} \tag{II. 108}$$

By considering the effect of dielectric relaxation in ionic motion; τ is the Debye relaxation time for the solvent dipoles. Zwanzig [306] treated the ion as a rigid sphere of radius r_i moving with a steady state viscosity, V_i through a viscous incompressible dielectric continuum. The conductance equation suggested by Zwanzig is:

$$\lambda_o^i = \frac{z_i^2 eF}{\left[A_V \pi\eta_o r_i + A_D\left\{\frac{z_i^2 e^2(\varepsilon_r^o - \varepsilon_r^\infty)\tau}{\varepsilon_r^o(2\varepsilon_r^o + 1)r_i^3}\right\}\right]} \tag{II. 109}$$

Where ε_r^o and ε_r^∞ are the static and limiting high frequency (optical) dielectric constants. $A_V = 6$ and $A_D = 3/8$ for perfect sticking and $A_V = 4$ and $A_V = 3/4$ for perfect slipping. It has been found that Born's and Zwanzig's equations are very similar and both may be written in the form:

$$\lambda_o^i = \frac{Ar_i^3}{r_i^4 + B} \qquad \text{(II. 110)}$$

The theory predicts [319] that λ_o^i passes through a maximum of $27^{\frac{1}{4}} A / 4B^{\frac{1}{4}}$ at $r_i = (3B)^{1/4}$. The phenomenon of maximum conductance is well known. The relationship holds good to a reasonable extent for cations in aprotic solvents but fails in case of anions. The conductance, however, falls off rather more rapidly than predicted with increasing radius. For comparison with results in different solvents, the equation (II. 109) can be rearranged as [320]:

$$\frac{z_i^2 eF}{\lambda_o^i \eta_o} = \frac{A_V \pi r_i + A_D z_i^2}{r_i^3} \cdot \frac{e^2 (\varepsilon_r^o - \varepsilon_r^\infty)}{\varepsilon_r^o (2\varepsilon_r^o + 1)} \left(\frac{\tau}{\eta_o} \right) \qquad \text{(II. 111)}$$

$$L^* = A_V \pi r_i + \frac{A_D z_i^2}{r_i^3 P^*} \qquad \text{(II. 112)}$$

In order to test Zwanzig's theory, the equation (II. 111) was applied for Me_4N^+ and Et_4N^+ in pure aprotic solvents like methanol, ethanol, acetonitrile, butanol and pentanol [319-324]. Plots of L^* against the solvent function P^* were found to be straight line. It is noted that relaxation effect is not the predominant factor affecting ionic mobility and these mobility differences could be explained quantitatively if the microscopic properties of the solvent, dipole moment and free electron pairs were considered the predominant factors in the deviation from the Stokes' law.

The Zwanzig's theory is found to be successful for bulky organic cations in aprotic media where solvation is seems to be minimum and the viscous friction predominates over that caused by dielectric relaxation. The theory breaks down whenever the dielectric relaxation term becomes large, i.e., for solvents of high P^* and for ions of small r_i. Like any continuum theory Zwanzig has the inherent weakness of its inability to account for the structural features [325], e.g.,

(i) It does not allow for any correlation in the orientation of the solvent molecules as the ion passes by and this may be the reason why the equation is not applicable to the hydrogen-bonded solvents [321].

(ii) The theory does not distinguish between positively and negatively charged ions and therefore, cannot explain why certain anions in dipolar aprotic media possess considerably higher molar concentrations than the fastest cations.

❖ *Thermodynamics of Ion-Pair Formation*

The standard Gibbs energy changes (ΔG^{o}) for the ion- association process can be calculated from the equation

$$\Delta G^{o} = -RT \ln K_{A} \qquad\qquad \text{(II. 113)}$$

The values of the standard enthalpy change, ΔH^{o} and the standard entropy change, ΔS^{o}, can be evaluated from the temperature dependence of values as follows;

$$\Delta H^{o} = -T^{2}\left[\frac{d(\Delta G^{o}/T)}{dT}\right]_{P} \qquad\qquad \text{(II. 114)}$$

$$\Delta S^{o} = -T^{2}\left[\frac{d(\Delta G^{o})}{dT}\right]_{P} \qquad\qquad \text{(II. 115)}$$

The values can be fitted with the help of a polynomial of the type:

$$\Delta G^{o} = c_{0} + c_{1}(298.15 - T) + c_{2}(298.15 - T)^{2} \qquad\qquad \text{(II. 116)}$$

And the coefficients of the fits can be compiled together with the σ % values of the fits. The standard values at 298.15 K are then:

$$\Delta G^{o}_{298.15} = c_{0} \qquad\qquad \text{(II. 117)}$$

$$\Delta S^{o}_{298.15} = c_{1} \qquad\qquad \text{(II. 118)}$$

$$\Delta H^{o}_{298.15} = c_{o} + 298.15 c_{1} \qquad\qquad \text{(II. 119)}$$

The standard entropy of ion-association of electrolytes depends on

(i)   the size and shape of the ions,

(ii)   charge density on the ions,

(iii)   electrostriction of the solvent molecules around the ions,

(iv)   penetration of the solvent molecules inside the space of the ions, and the influence of these factors are discussed later.

The non-columbic part of the Gibbs energy ΔG^o can also be calculated using the following equation:

$$\Delta G^o = N_A W_{\pm} \tag{II. 120}$$

$$K_A = \left(\frac{4\pi N_A}{1000}\right)\int_a^R r^2 \exp\left(\frac{2q}{r} - \frac{W_{\pm}}{kT}\right)dr \tag{II. 121}$$

Where the symbols have their usual significance. The quantity 2q/r is Columbic part of the interionic mean force potential and $W_{\pm}$ is its non-columbic part.

❖ *Triple-Ion Formations from Electrical Conductance*

While solutions of electrolytes in solvents of high and of intermediate dielectric constant have been studied extensively, similar solutions in solvents of very low dielectric constant have not been investigated systematically. We know only that such solutions generally are poor conductors and that the equivalent conductance falls rapidly with decreasing concentration. In addition to a number of isolated observations on the conductance of solutions in benzene [326] and several series of measurements relating to the conductance of complex compounds in various solvents at relatively high concentrations [327], the literature includes two important papers by Walden and his co-workers [328], who investigated the conductance of a variety of *salts* in benzene, ether, carbon tetrachloride and similar solvents. In solvents of somewhat higher dielectric constant, the conductance passes through a minimum at moderate concentration and thereafter increases.

The influence of the dielectric constant on conductance is satisfactorily accounted for by the interionic attraction theory in solvents of high dielectric constant, it is not known to what extent interionic forces are primarily concerned in solvents of low dielectric constant. Due to the deviation of the conductometric curves from linearity in case of low dielectric constant solvents, the conductance data have been analyzed by the classical Fuoss-Kraus theory of triple-ion formation [329,330] in the form

$$\Lambda \, g(c)\sqrt{c} = \frac{\Lambda_0}{\sqrt{K_p}} + \frac{\Lambda_0^T K_T}{\sqrt{K_p}}\left(1 - \frac{\Lambda}{\Lambda_0}\right)c \qquad (\text{II.122})$$
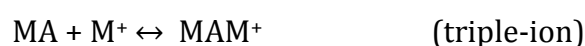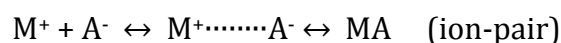
where $g(c)$ is a factor that lumps together all the intrinsic interaction terms and is defined by

$$g(c) = \frac{\exp\{-2.303\,\beta'\,(c\Lambda)^{0.5}/\Lambda_0^{0.5}\}}{\{1 - S(c\Lambda)^{0.5}/\Lambda_0^{1.5}\}(1 - \Lambda/\Lambda_0)^{0.5}} \qquad (\text{II.123})$$

$$\beta' = 1.8247 \times 10^6 / (\varepsilon_r T)^{1.5} \qquad (\text{II.124})$$

$$S = \alpha\,\Lambda_0 + \beta = \frac{0.8204 \times 10^6}{(\varepsilon_r T)^{1.5}}\Lambda_0 + \frac{82.501}{\eta(\varepsilon_r T)^{0.5}} \qquad (\text{II.125})$$

In all the above usable equations, the Λ_0 term signify the sum of the molar conductance of the simple ions at infinite dilution, the Λ_0^T is the sum of the conductance value of the two possible formation of triple-ions. Where the constants as K_P and K_T are implies that the ion-pair and triple-ion formation constants respectively and S is the limiting Onsager coefficient. If K_P is greater than K_T, indicates the electrolytes are exists as ion-pair with a major portion and as triple-ion with a minor portion. We know that the electrostatic ionic interactions are very large due to the higher force field effect, for very low relative permittivity solvents, i.e., $\varepsilon_r < 10$. Therefore the formed ion-pairs were attracted by the free movable cations or anions present in the solution medium; as the distance of the closest approach of the ions becomes minimum, these results in the formation of triple-ions, which acquires the charge of the respective ions, attracted from the solution bulk [330,331] i.e.;

$$M^+ + A^- \leftrightarrow M^+\cdots\cdots A^- \leftrightarrow MA \quad \text{(ion-pair)}$$
$$MA + M^+ \leftrightarrow MAM^+ \qquad\qquad \text{(triple-ion)}$$
$$MA + A^- \leftrightarrow MAA^- \qquad\qquad \text{(triple-ion)}$$

Where the symbols M+ and A- are implies for the cation and anions respectively. The effect of ternary association [332] thus clearly explained the non-linearity of the conductometric curve. According to the consequence of this ternary association, some formative non-conducting species MA, removed from solution and replaces by triple-ions which increase the conductance values evident by non-linearity observed in conductance curves.

Additionally, the ion-pair and triple-ion concentrations, C_P and C_T, respectively, at the minimum molar concentration of the salt solution have also been calculated using the K_P and K_T value by following set of equations [330]

$$\alpha = 1 / \left(K_P^{1/2} \cdot C^{1/2} \right) \tag{II.126}$$

$$\alpha_T = \left(K_T / K_P^{1/2} \right) C^{1/2} \tag{II.127}$$

$$C_P = C \left(1 - \alpha - 3\alpha_T \right) \tag{II.128}$$

$$C_T = \left(K_T / K_P^{1/2} \right) C^{3/2} \tag{II.129}$$

At this point, the fraction of ion-pairs (α) and triple-ions (α_T), present in the salt-solutions. From the appraisal of comparison of the C_P and C_T values, if C_P is higher with respect to C_T; indicates the major portion of ions are present as ion-pair even at high concentration and a small fraction exist as triple-ion.

Using the K_P values, the interionic distance parameter a_{IP} has been calculated with the aid of the Bjerrum's theory of ionic association [332] in the form

$$K_P = \frac{4\pi N_A}{1000} \left[\frac{e^2}{\varepsilon_r KT} \right]^3 Q(b) \tag{II.130}$$

$$Q(b) = \int_2^b y^{-4} \exp(y) \, dy \tag{II.131}$$

$$b = \frac{e^2}{a_{IP} \varepsilon_r KT} \tag{II.132}$$

The $Q(b)$ and b values have been calculated by the literature procedure [329].

The interionic distance a_{TI} for the triple ion can be calculated using the expressions [331]

$$K_T = \frac{2\pi N_A a_{IP}^3}{1000} I(b_3)$$ (II.133)

$$b_3 = \frac{e^2}{a_{TI}\varepsilon_r KT}$$ (II.134)

$I(b_3)$ is a double integral tabulated in the literature [330] for a range of values of b_3. Since $I(b_3)$ is a function of a_{TI}, the a_{TI} values have been calculated by an iterative computer program. These values also suggested the small fraction exist as triple-ion formation compared to the ion-pair.

❖ *Solvation Models (Some Recent Trends)*

The interactions between particles in chemistry have been based upon empirical laws- principally on Coulomb's law. This is also the basis of the attractive part of the potential energy used in the Schrödinger equation. Quantum mechanical approach for ion-water interactions was begun by Clementi in 1970. A quantum mechanical approach to salvation can provide information on the energy of the individual ion-water interactions provided it is relevant to solution chemistry, because it concerns potential energy rather than the entropic aspect of salvation. Another problem in quantum approach is the mobility of ions in solution affecting salvation number and coordination number. However, the Clementi calculations concerned stationary models and cannot have much to do with the dynamic salvation numbers. Covalent bond formation enters little into the aqueous calculations; however, with organic solvents the quantum mechanical approaches to bonding may be essential. The trend pointing to the future is thus the molecular dynamics technique. In molecular dynamic approach, a limited number of ions and molecules and Newtonian mechanics of movement of all particles in solution is concerned. The basis of such an approach is the knowledge of the intermolecular energy of interactions between a pair of particles. Computer simulation approaches may be useful in this regard and the last decade (1990-2000) witnessed some interesting trends in the development of solvation models and computer software. C.J. Cramer, D.G. Truhlar and co-workers from the University of Minnesota, U.S.A. constructed a series of solvation models (SM1-SM5 series) based on a collection of experimental free energy of solvation data, to predict and calculate the free energy of

solvation of a chemical compound [333-337]. These models are applicable to virtually any substance composed of H, C, N, O, F, P, S, Cl, Br and/or I. The only input data required are, molecular formula, geometry, refractive index, surface tension, Abraham's *a*(acidity parameter) and *b*(basicity parameter) values, and, in the latest models, the dielectric constants. The advantage of models like SM5 series is that they can be used to predict the free energy of self-solvation to better than 1 KCl/mole. These are especially useful when other methods are not available. One can also analyze factors like electrostatics, dispersion, hydrogen bonding, etc. using these tools. They are also relatively inexpensive and available in easy to use computer codes.

A. Galindo *et al.* [338,339] have developed Statistical Associating Fluid Theory for Variable Range (SAFT-VR) to model the thermodynamics and phase equilibrium of electrolytic aqueous solutions. The water molecules are modeled as hard spheres with four short-range attractive sites to account for the hydrogen-bond interactions. The electrolyte is modeled as two hard spheres of different diameter to describe the anion and cation. The Debye-Hückel and mean spherical approximations are used to describe the interactions. The relative permittivity becomes very close to unity, especially when the mean spherical approximation is used, indicating a good description of the solvent. E. Bosch *et al.* [340] of the University of Barcelona, Spain, have compared several "*Preferential Solvation Models*" specially for describing the polarity of dipolar hydrogen bond acceptor-cosolvent mixture.

II.6.6 Refractive Index

Optical data (refractive index, n_D) provide interesting information related to molecular interactions and structure of the solutions, as well as complementary data on practical procedures, such as concentration measurement or estimation of the extent of salvation of electrolytes/non-electrolytes in liquid systems.

The light bending property is a result of variation of the velocity with which light is transmitted. Refractive index (n_D) of liquid, changes not only with the wave length of light used but also with the temperature. Molar refractions are influenced by the arrangement of atoms in the molecule or by factors like unsaturation, ring closure etc. linear optical properties of liquids and liquid mixtures have been widely studied to obtain

information on their physical, chemical, and molecular properties. Fialkov et. al. [341,342] stated that the refractive index is an additive properties of pure components when composition is expressed in terms of volume fraction. Several researchers have estimated the refractivity of liquid systems using the well known mixting rules viz. Arago-Biot, Newton, Heller, Gladstone-Dale, Eyring-John, Eykman, Lorentz-Lorenz, Weiner and Oster relations [343-346]. These empirical approaches for calculating the excess properties attempt to explain the non-ideality in terms of specific and non-specific intermolecular interactions. Refractive index or refractivity is a property of intrinsic interest in the fields of pharmaceutical research such as formulation of eye preparations, in optoelectronic and photonic applications.

The ratio of the speed of light in a vacuum to the speed of light in another substance is defined as the index of refraction (n_D) for the substance.

$$Refractive\ Index\ (n_D) = \frac{Speed\ of\ light\ in\ vacuum}{Speed\ of\ light\ in\ solution\ systems}$$

Due to the change in speed of light its direction of travel also changes as it crosses a boundary from one medium into another, i.e., it is refracted. The relationship between light's speed in the two mediums (V_A and V_B), the angles of incidence ($sin\,\theta_A$) and refraction ($sin\,\theta_B$) and the refractive indexes of the two mediums (n_A and n_B) is shown below:

$$\frac{V_A}{V_B} = \frac{sin\,\theta_A}{sin\,\theta_B} = \frac{n_B}{n_A} \tag{II. 135}$$

It is possible to determine the refractive index of the sample quite accurately by measuring the angle of refraction, and knowing the index of refraction of the layer that is in contact with the sample, instead of measuring speed of light.

The refractive index of mixing can be correlated by the application of a composition-dependent polynomial equation. Molar refractivity, was obtained from the Lorentz-Lorenz relation [347,348] by using, n_D experimental data according to the following expression

$$R_M = \frac{(n_D^2 - 1)}{(n_D^2 + 2)} \left(\frac{M}{\rho} \right) \qquad \text{(II. 136)}$$

Where M is the mean molecular weight of the mixture and ρ is the mixture density. n_D can be expressed as the following:

$$n_D = \sqrt{\frac{(2A + 1)}{(1 - A)}} \qquad \text{(II. 137)}$$

Where A is given by:

$$A = [\{\frac{(n_1^2 - 1)}{(n_1^2 + 2)}(1/\rho_1)\} - \{\frac{(n_1^2 - 1)}{(n_1^2 + 2)}(w_2/\rho_1)\} + \{\frac{(n_2^2 - 1)}{(n_2^2 + 2)}(w_2/\rho_2)\}\rho] \qquad \text{(II. 138)}$$

Where n_1 and n_2 are the pure component refractive indices, w_j the weight fraction, ρ the mixture density, and ρ_1 and ρ_2 the pure component densities.

The molar refractivity deviation is calculated by the following expression:

$$\Delta R = R - \phi_1 R_1 - \phi_2 R_2 \qquad \text{(II. 139)}$$

Where ϕ_1 and ϕ_2 are volume fractions and R, R_1, and R_2 the molar refractivity of the mixture and of the pure components, respectively.

The deviations of refractive index were used for the correlation of the binary solvent mixtures:

$$\Delta n_D = n_D - x_1 n_{D1} - x_2 n_{D2} \qquad \text{(II. 140)}$$

Where Δn_D is the deviation of the refractive index for this binary system and n_D, n_{D1}, and n_{D2} are the refractive index of the binary mixture, refractive index of component-1, and refractive index of component-2, respectively, 'x' is the mole fractions.

The computed deviations of refractive indices of the binary mixtures are fitted using the following Redlich-Kister expression [349].

$$\Delta n_{Dew} = w_e w_w \sum_{P=0}^{S} B_p (w_e w_w)^p \qquad \text{(II. 141)}$$

Where B_p are the adjustable parameters obtained by a least squares fitting method, w is the mass fraction, and S is the number of terms in the polynomial.

The molar refractivity is isomorphic to a volume for which the ideal behavior may be expressed in terms of mole fraction: in this case smaller deviations occur but data are more scattered because of the higher sensitivity of the expression to rounding errors in the mole fraction. For the sake of completeness, both calculations of refractivity deviation function, molar refractivity deviation was fitted to a Redlich and Kister-type expression [349] and the adjustable parameters and the relevant standard deviation *ó* are calculated for the expression in terms of volume fractions and in terms of mole fractions, respectively.

II.6.7 FTIR Spectroscopy

The spectroscopic study has been established by the investigation of FTIR spectroscopy. The study has been taking into account to qualitative interpreting the molecular as well as ionic association of the electrolytes in the solutions. FTIR spectroscopy is one of the most appropriate optical properties which qualitatively interpreted the nature, mode, manner of the electrolytes and non-electrolytes in the solution system, eventually it also is able to give information about the configurational structure of the solute or solvents present in the solutions.

Infrared (IR) spectroscopy is is the absorption measurement of different IR frequencies by a sample positioned in the path of an IR beam. The chemical functional groups present in the sample can be analysed by IR spectroscopy. IR spectrometers can be applicable for wide range of sample types such as gases, liquids, and solids. Hence, IR spectroscopy is found to be an important analytical technique for structural elucidation and compound identification.

Infrared radiation spans a section of the electromagnetic spectrum having wave numbers from roughly 13,000 to 10 cm^{-1}, or wavelengths from 0.78 to 1000 μm. It is bound by the red end of the visible region at high frequencies and the microwave region at low frequencies.

IR absorption positions are generally presented as either wave numbers (v) or wavelengths (l). Wave number defines the number of waves per unit length i.e. wave numbers are directly proportional to frequency, as well as the energy of the IR absorption. The unit of the wave number is cm^{-1} (reciprocal centimeter) in modern IR

instruments that are linear in the cm^{-1} scale. In the contrast, wavelengths are inversely proportional to frequencies and their associated energy. At present, the recommended unit of wavelength is μm (micrometers), but μ (micron) is used in some older literature. Wave numbers and wavelengths can be interconverted using the following equation:

$$v\ (cm^{-1}) = \frac{1}{\lambda\ (\mu m)} \times 10^4 \qquad (II.142)$$

In the IR spectrum, wavelength or wavenumber taken as the x-axis and absorption intensity or percent transmittance as the y-axis.

Transmittance, T, is the ratio of radiant power transmitted by the sample (I) to the radiant power incident on the sample (I_0). Absorbance (A) is the logarithm to the base 10 of the reciprocal of the transmittance (T).

$$A = log_{10}(1/T) = -log_{10}(T) = -log_{10}\left(\frac{I}{I_0}\right) \qquad (II.143)$$

The transmittance spectra offer better contrast between intensities of strong and weak bands as transmittance ranges from 0 to 100% T whereas absorbance ranges from infinity to zero.