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# Conjoint Beneficial Impacts of Lithium-Ion along with Aqueous Vitamin in Driven Rechargeable Batteries and also in Neurotransmitter Drugs

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ABSTRACT

Lithium ions are "energy capsules". The electrochemical calculations of the aqueous rechargeable lithium-ion battery has been widely investigated in efforts to design a green and safe technology that can provide a highly explicit capacity, high efficiency and long life for high power applications such as the smart grid and electric vehicle. It is assumed that the advantages of this battery will overcome the limitations of the rechargeable lithium-ion battery with organic electrolytes that encompass safety and create high production cost issues. This assessment focuses on the opportunities of the aqueous rechargeable lithium-ion battery compared to the conventional rechargeable lithium-ion battery with organic-based electrolytes. Previously reported studies are briefly summarized, together with the presentation of new findings based on the conductivity, morphology, electrochemical performance and cycling stability consequences. The factors that influence the electrochemical performance, the challenges and potential of the aqueous rechargeable lithium-ion battery are highlighted with emphasis on maintaining the excellent battery performance. Moreover current studies, tip to the possibility that long-term treatment with ascorbic acid based lithium drug might prove useful for reducing brain damage and motivating new neural growth in neurodegenerative diseases.

## 1. Introduction

Since years the specific biochemical mechanism of lithium deed in stabilizing mood is known. Presently, the majority of lithium ions utilized in aqueous rechargeable designs of batteries can be widely used in various kinds of portable electric devices because of their high performance with much more power, high output voltage and also can find applications in the development of electrolyte solutions [1, 2]. Moreover the interconnection of lithium salts with vitamin increases diverse impacts in biological systems [3, 4].

In the present work molecular dynamics calculations from the experimental determination of solvation phenomena for the ions of lithium halides in water-ascorbic acid solvents have been discussed. The results have been interpreted in terms of ion-solvent interactions and structural modifications in the solvent media [5, 6]. The electrical conductivity of electrolytes in mixed solvent solutions mainly depends upon the concentration of the electrolyte and also upon the viscosity of the solvent. The use of mixed solvents in high energy batteries has also extended the area in the field of mixed solvent systems [7, 8]. Hence, in studying the behavior of lithium salts in different solvent systems will help in the production of more useful and "cost-effective batteries" [9].

Vitamin is very suitable due to the excellent solubility characteristics. Vitamin plays an important role for the synthesis of several important neurotransmitters. So, in our recent work due to incorporation of lithium ion in aqueous ascorbic acid solution, enhancement of neurotransmission property occurs. The new battery is similar to many commercially-available lithium-ion batteries with one important difference – it has the probability to use butenolide (2-Furanone) from vitamin C (ascorbic acid) as the cathode, the part that stores the electricity that is released when connected to a device. While bio-derived battery parts have been created previously, this is the one that uses bio-derived long-chain molecules – for one of the electrodes, essentially allowing battery energy to be stored, instead of expensive, harder to process, and more environmentally-harmful metals such as cobalt. Modern batteries contain a positive terminal, a negative terminal, and an electrolyte solution. When a battery is connected to any electronic device that requires power, electrons flow from the anode– the negatively charged electrode of the device supplying current – to the device, then into the cathode. The ions also migrate through the electrolyte solution to balance the charge. When the device is connected to a charger, this process happens in reverse. The reaction in the anode creates electrons and the reaction in the cathode absorbs them when discharging. The result of this is electricity. The battery will continue to produce electricity until one or both of the electrodes run out of the substances necessary for these reactions to take place.

It is clear from the above discussion that electrochemical cells of lithium can be used to provide power, in which case they are called batteries; moreover lithium ions can use the reactants supplied continuously and thus can be called as fuel cells.

## 2. Experimental Methods

### *2.1 Materials*

All the lithium halides were purchased from Merck Company, Ascorbic acid used for this study were of analytical grade supplied by Sisco Research Laboratories (SRL). All these chemicals were used as supplied without any further purification. Lithium chloride (LiCl), lithium bromide (LiBr), and lithium iodide (LiI) with minimum mass fraction purities of (> 0.99) were dried in the electrical oven at about 110 °C for 24 h prior to use. Double distilled water was used for preparation of the solutions. Water has been used as the primary solvent obtained from series of distillations till its specific conductance attains the range of ($0.1x10^{-6}$ to $1.0x10^{-6}$ $\Omega^{-1}cm^{-1}$). The purity of water was further checked by measuring its density. Before use the chemicals were kept in a vacuum oven to eliminate any moisture.

### *2.2 Preparation of Solutions*

All the solutions were prepared by weight and conversions of molality to molarity were done by using the standard expression of an analytical balance (precision $1x10^{-5}$ g). All these solutions were prepared in 0.001 m, 0.003 m, 0.005 m ascorbic acid in water as solvent in the various temperatures of 298.15 K, 303.15 K, and 308.15 K (Scheme 1). Six concentrations ranging from 0.005 $molkg^{-1}$ to 0.08 $molkg^{-1}$ of LiCl, LiBr, and LiI were prepared by taking different compositions of ascorbic acid in water as solvent for all studies.

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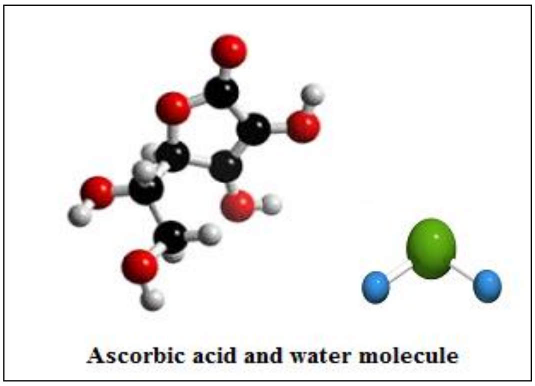


**Scheme 1** Three dimensional representation of solvent cage

### 2.3 Instrumentation

Density was measured by vibrating-tube digital Anton Paar Density-Meter (DMA 4500M) with a precision of ±0.00005 gcm$^{-3}$, was calibrated by double-distilled water and dry air. The temperature was automatically kept constant within ±0.01 K. It was calibrated by triply-distilled water and passing dry air. The viscosities were measured using a Brookfield DV-III Ultra Programmable Rheometer with fitted spindle size -42 fitted to a Brookfield digital bath TC-500. The viscosities were obtained using the following equation,

η= (100 / *RPM*) × TK× torque ×*SMC*

where *RPM*, *TK* (0.09373) and *SMC* (0.327) are the speed, viscometer torque constant and spindle multiplier constant, respectively. The instrument was calibrated against the standard viscosity samples supplied with the instrument, water and aqueous $CaCl_2$ solutions. Temperature of the solution was maintained to within ± 0.01 K using Brookfield Digital TC-500 temperature thermostat bath. The viscosities were measured with an accuracy of ±1%. Each measurement reported herein is an average of triplicate reading with a precision of 0.3%.

Refractive index was measured with the help of a digital refractometer Mettler Toledo. The light source was light emitting diode (LED), $\lambda$=589.3 nm. The refractometer was calibrated twice using distilled water and calibration was checked after every few measurements. The uncertainty of refractive index measurement was ±0.0002 units.

The conductance measurements were carried out in a Systronics-308 conductivity bridge of accuracy ±0.01%, using a dip-type immersion conductivity cell, CD-10 having a cell constant of approximately (0.1± 0.001) cm$^{-1}$ [10, 11] . Measurements were made in a thermostat water bath maintained at T = (298.15 ±0.01) K. The cell was calibrated by the method proposed by Lind et al. and cell constant was measured based on 0.01 M aqueous KCl solution. During the conductance measurements, cell constant was maintained within the range 1.10–1.12 cm$^{-1}$. The conductance data were reported at a frequency of 1 kHz and the accuracy was ±0.3%. During all the measurements, uncertainty of temperatures was ±0.01 K [12].

## 3. Results and Discussion

### 3.1 Density Calculation

The measured value of densities (ρ) of lithium halide in solvent mixtures of different mass fraction of ascorbic acid in water, as a function of concentration at 298.15 K, 303.15 K and 308.15 K are reported.

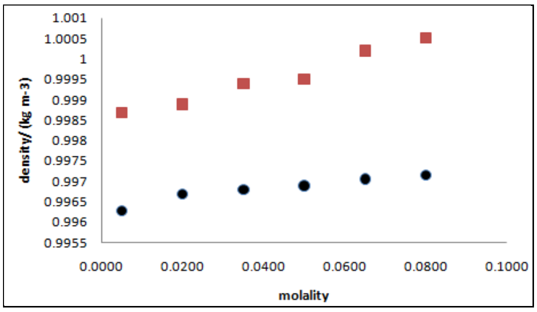


**Fig. 1** Comparison of measured densities for (ascorbic acid+ $H_2O$) binary system with literature at T = 298.15 K, this work;● with reference to the introduction of the salt ■ in the above mixture

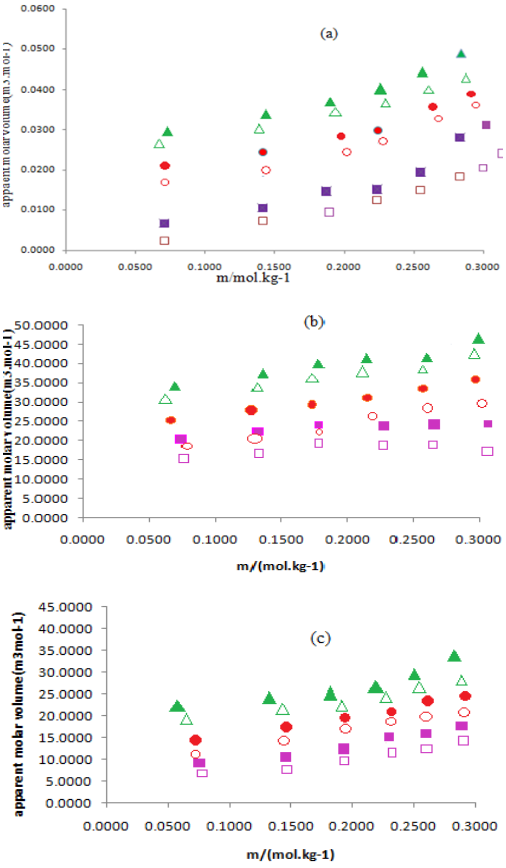
The measured values of density for binary (ascorbic acid + $H_2O$) system and ternary (ascorbic acid + LiCl + $H_2O$), (ascorbic acid +LiBr + $H_2O$), and (ascorbic acid + LiI + $H_2O$) mixtures at different concentrations of ascorbic acid and (LiCl, LiBr and LiI), at different temperatures are given. Fig. 1 shows the comparison of measured densities for (ascorbic acid + $H_2O$) binary system in this work and literature at T = 298.15 K, 303.15 K, 308.15 K. Both data show linear dependence to molality of ionic liquid with almost same intercepts but different slopes [13]. This discrepancy in slopes may come from differences in purities of used ascorbic acid and also little differences in densities due to pressures (Fig. 1).

### 3.2 Apparent Molar Volume Studies

The apparent molar volume of LiCl, LiBr and LiI in (0.001, 0.003, 0.005) m aqueous ascorbic acid solution have been calculated from density data (Figs. 2a, b and c) by using Eq.(1),

$$\phi_v=M_2/d-(d-d^0)/mdd^0 \qquad (1)$$

where $d^0$ is the density of solvent, d is the density of solution, m the molality of solution and $M_2$ the molecular weight of electrolyte.



At T = 298.15 K: green△,LiCl; red○, LiBr; purple □,LiI.

At T = 318.15 K: green▲,LiCl; red●,LiBr; purple■,LiI

**Fig. 2.** Variation tendency of apparent molar volumes,of Ascorbic acid in aqueous LiCl, LiBr and LiI solutions (a)$w_{salt}$ = 0.001(m) (b) $w_{salt}$ = 0.003(m) (c) wsalt =0.005(m)

The densities of various solutions of lithium chloride, lithium bromide and lithium iodide in (0.001, 0.003, 0.005) m aqueous ascorbic acid obey Root's equation and justify the use of Masson's Eq. (2) for the estimation of the limiting apparent molar volume.

$$\phi_v=\phi_v^0+S_v^*\sqrt{C} \qquad (2)$$

where $\phi_v^0$ and $S_V^*$ are calculated from the intercept and slope from the extrapolation of the plots of $\phi_v$ versus √C shown. The values of limiting apparent molar volume and slopes $S_V^*$ are recorded in Figs. 3a, b and c. The slope $S_V^*$ in Masson's equation may be attributed to be as a measure of ion–ion or solute–solute interactions. Low and positive values of $S_V^*$ for LiCl, LiBr and LiI accounts for weak ion–ion interactions [14, 15]. There is a decrease in inter ionic interactions with increase in temperature for lithium chloride, lithium bromide, lithium iodide in (0.001, 0.003, 0.005) m aqueous ascorbic acid solutions which may be due to more solvation of electrolytic ions with rise in temperature [16].

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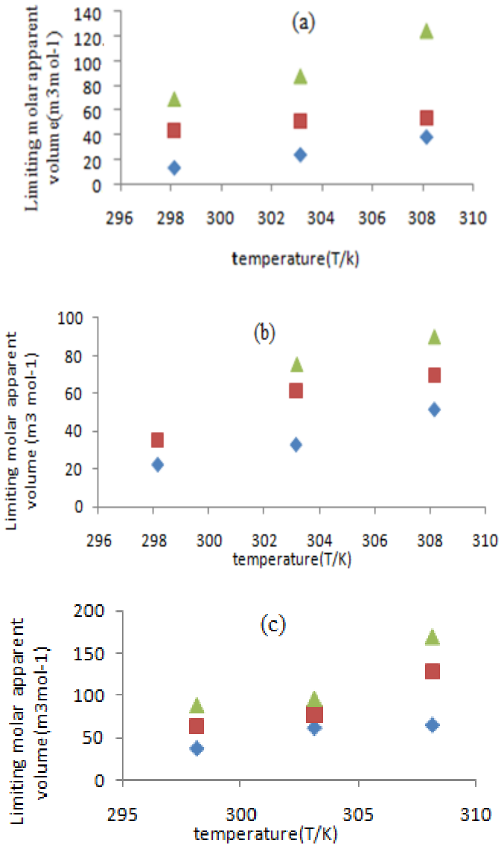
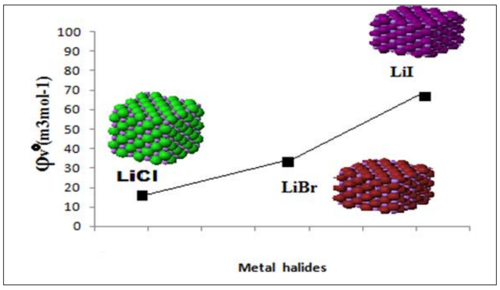
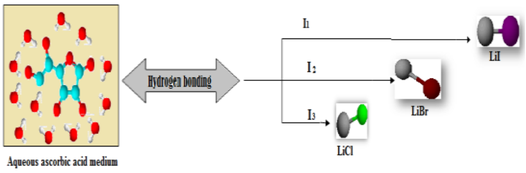


**Fig. 3** Variation tendency of apparent molar volumes at infinite dilution for ternary [(LiX + Ascorbic acid + $H_2O$) where X = Cl, Br, I] system at different temperature obtained from the plots: (a)*LiCl* shows $\varphi_v^0$ for 0.001(m) concentration as green ; 0.003(m) concentration as brown ; 0.005(m) concentration as blue(b) *LiBr* shows $\varphi_v^0$ for 0.001(m) concentration as green ; 0.003(m) concentration as brown ; 0.005(m) concentration as blue(c) *LiI* shows $\varphi_v^0$ for 0.001(m) concentration as green ; 0.003(m) concentration as brown ; 0.005(m) concentration as blue

The $\phi_v^o$ is a measure of ion–solvent interactions (Scheme 2). The positive values of $\phi_v^o$ for LiCl, LiBr and LiI shows positive ion–solvent interactions [17, 18]. The $\phi_v^o$ values for LiCl, LiBr and LiI increases with increase in temperature and it may be due to more solvation of electrolytic ions as a result of decrease in hydrogen bonding between solvent molecules with rise in temperature, thus making more free solvent molecules available for solvation of electrolytic ions shown in (Scheme 3), where $I_1 > I_2 > I_3$ [19, 20].



**Scheme 2** Array of limiting apparent molar volumes at a definite concentration



**Scheme 3** Array of molecular interactions

### *3.3 Temperature Dependent Limiting Apparent Molar Volume*

The variation of $\phi V^0$ with the temperature of the lithium halides in different solvents with the temperature can be expressed by the general polynomial equation as follows,

$$\phi V^0 = a_0 T + a_1 T + a_2 T^2 \qquad (3)$$

where $a_0$, $a_1$, $a_2$ are the empirical coefficients depending on the solute, mass fraction ($w_1$) of the co-solute, and T is the temperature range under study in Kelvin. The values of these coefficients of the above equation for the lithium salts in vitamin C and water are reported at 298.15 K, 303.15 K and 308.15 K are shown in Table 1.

**Table 1** Values of empirical coefficients ($a_0$, $a_1$, and $a_2$) of Eq. (8) for lithium halides in co-solute (ascorbic acid, water) at 298.15 K, 303.15 K, 308.15 K

| Concentration | Electrolyte | $a_0 \cdot 106/m^3 \cdot mol^{-1}$ | $a_1 \cdot 106/m^3 \cdot mol^{-1} \cdot K^{-1}$ | $a_2 \cdot 106/m^3 \cdot mol^{-1} \cdot K^{-2}$ |
|---|---|---|---|---|
| 0.001 | LiCl | 265.7 | 0.648 | -0.005 |
| 0.003 | | 290.6 | -0.633 | -0.002 |
| 0.005 | | 529.6 | -0.503 | -0.001 |
| 0.001 | LiBr | 251.6 | 0.854 | -0.006 |
| 0.003 | | 282.4 | 0.852 | -0.004 |
| 0.005 | | 319.4 | 0.755 | -0.002 |
| 0.001 | LiI | 178.0 | 2.008 | -0.007 |
| 0.003 | | 256.2 | 0.912 | -0.004 |
| 0.005 | | 397.4 | 0.380 | -0.001 |

The temperature dependence of $\phi v^o$ for LiCl, LiBr and LiI can be expressed as:

$$\phi v^o = 265.7 + 0.648T - 0.005T^2 \qquad (4)$$

for: (LiCl in 0.001 m aqueous ascorbic acid system)

$$\phi v_o = 290.6 - 0.633T - 0.002T^2 \qquad (5)$$

for : ( LiCl in 0.003 m aqueous ascorbic acid system)

$$\phi v_o = 529.6 - 0.503T - 0.001T^2 \qquad (6)$$

for : ( LiCl in 0.005 m aqueous ascorbic acid system)

$$\phi v_o = 251.6 + 0.854\,T - 0.006T^2 \qquad (7)$$

for : ( LiBr in 0.001 m aqueous ascorbic acid system)

$$\phi v_o = 282.4 + 0.852\,T - 0.004T^2 \qquad (8)$$

for: ( LiBr in 0.003 m aqueous ascorbic acid system)

$$\phi v_o = 319.4 + 0.755\,T - 0.002T^2 \qquad (9)$$

for: (LiBr in 0.005 m aqueous ascorbic acid system)

$$\phi v_o = 178.0 + 2.008\,T - 0.007T^2 \qquad (10)$$

for: ( LiI in 0.001 m aqueous ascorbic acid system)

$$\phi v_o = 256.2 + 0.912T - 0.004T \qquad (11)$$

for: ( LiI in 0.003 m aqueous ascorbic acid system)

$$\phi v_o = 397.4 + 0.380T - 0.001T^2 \qquad (12)$$

for: ( LiI in 0.005 m aqueous ascorbic acid system)

The limiting apparent molar expansibilities, $\phi E^0$ can be obtained by the following equation,

$$(\phi E^0) = (\delta \varphi v^0 / \delta T)_p = a_1 + 2a_2 T \qquad (13)$$

The limiting apparent molar expansibilities, $\phi E^0$, change in magnitude with the change of temperature. The values of $\phi E^0$ for different solutions of the studied sample at 298.15, 303.15, and 308.15 K are reported in Table 2. The table reveals that $\phi E^0$ is positive for lithium halides in all the concentrations of studied solvents and studied temperatures [21, 22]. This fact can ascribed to the absence of "caging or packing effect" for the lithium salts in solutions and their behavior is just like common electrolytes [23, 24].

During the past few years it has been emphasized by different workers that SV* is not the sole criterion for determining the structure-making or -breaking nature of any solute.

The structure making/ breaking capacity of LiCl, LiBr and LiI may be interpreted with the help of Hepler's reasoning shown in Table 2. Hepler developed a technique of examining the sign of $(\delta\phi E^0/\delta T)p$ for the solute in terms of long-range structure-making and -breaking capacity of the solute in the solvent systems using the general thermodynamic expression,

$$(\delta\phi E^0/\delta T)_p = (\delta 2\phi v^0/\delta T^2)_p = 2a_2 \qquad (14)$$

Eq. (14) can also be represented as,

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$$(\partial \bar{C}p^0/\partial P)_T = -T(\partial^2 \phi^o v/\partial T^2)_p \qquad (15)$$

where $\bar{C}p^0$ is the partial molar heat capacity at infinite dilution. From Eq. (15), it is clear that structure making electrolytes should have a positive value of $(\partial^2\phi v^o/\partial T^2)_p$ and structure breaking electrolytes should have negative value of $(\partial_2\phi v_o/\partial T^2)_p$. For LiCl, LiBr and LiI the sign of $(\partial^2\phi v^o/\partial T^2)_p$ has been found to be negative which suggests that LiCl, LiBr and LiI acts as structure-breaker in 0.001, 0.003, 0.005 m aqueous ascorbic acid solution.

**Table 2** Limiting apparent molal expansibilities ( $\varphi_E$) for lithium salts in co-solute (ascorbic acid and water) at 298.15 K to 308.15 K respectively

| Electrolyte | Temperature(T)(K) | $\phi E^o \cdot 10^6$ /($m^3mol^{-1}K^{-1}$) | $2a_2$ |
|---|---|---|---|
| LiCl (0.001) | 298.15 K | 2.432 | -0.014 |
| | 303.15 K | 2.382 | |
| | 308.15 K | 2.333 | |
| LiCl (0.003) | 298.15 K | 1.865 | -0.004 |
| | 303.15 K | 1.845 | |
| | 308.15 K | 1.825 | |
| LiCl (0.005) | 298.15 K | 1.119 | -0.016 |
| | 303.15 K | 1.109 | |
| | 308.15 K | 1.099 | |
| LiBr(0.001) | 298.15 K | 2.843 | -0.028 |
| | 303.15 K | 2.783 | |
| | 308.15 K | 2.723 | |
| LiBr(0.003) | 298.15 K | 1.613 | -0.036 |
| | 303.15 K | 1.573 | |
| | 308.15 K | 1.533 | |
| LiBr(0.005) | 298.15 K | 0.477 | -0.04 |
| | 303.15 K | 0.457 | |
| | 308.15 K | 0.437 | |
| LiI (0.001) | 298.15 K | 2.306 | -0.054 |
| | 303.15 K | 2.230 | |
| | 308.15 K | 2.166 | |
| LiI (0.003) | 298.15 K | 1.553 | -0.036 |
| | 303.15 K | 1.513 | |
| | 308.15 K | 1.473 | |
| LiI (0.005) | 298.15 K | 0.236 | -0.016 |
| | 303.15 K | 0.226 | |
| | 308.15 K | 0.216 | |

### *3.4 Viscosity Studies*

Another transport property of solution is viscosity has been studied for comparison and conformation of the solvation of the electrolyte in the chosen solvents. The viscosity data has been analyzed using Jones-Dole equation.

$$(\eta_r - 1)/\sqrt{c} = A + B\sqrt{c} \qquad (16)$$

where $(\eta/\eta_0) = \eta_r$, η and $\eta_0$ are relative viscosities, viscosities of solution and solvent respectively, c is the molar concentration and A and B are constants. The values of η, $(\eta_r - 1)/\sqrt{c}$ are shown. The values of A and B have been determined from the intercept and slope of linear plots of $(\eta_r - 1)/\sqrt{c}$ versus $\sqrt{c}$. The values of A and B of different solutions are recorded. Parameter A of Jones–Dole equation represents the contribution from ion–ion interactions. The values of A, shows that ion–ion interactions for LiCl, LiBr and LiI decreases with increase in temperature, which may be due to more solvation of ions.

The B parameter which measures the structure making/breaking capacity of an electrolyte in a solution also contain a contribution from structural effects and is responsible for solute–solvent interactions in a solvent. It has been emphasized by a number of workers that dB/dT is more important criteria for determining solute–solvent interactions. Viscosity study of a number of electrolytes has shown that structure-maker will have negative dB/dT and structure breaker will have positive dB/dT. The temperature effect on B coefficient for LiCl, LiBr and LiI shows a positive sign of dB/dT, showing thereby that LiCl, LiBr and LiI behaves as structure-breaker in (0.001, 0.003, 0.005) m aqueous ascorbic acid solution. The effect of temperature on the viscosity is given where A and B are viscosity coefficient indicates the ion–ion and ion–solvent interaction, respectively.

As the electrolytes investigated here are found to be somewhat associated from conductivity measurements, the viscosity data have been analyzed by the following equation.

$$\eta_r = 1 + A(\alpha c) + B\alpha c + B/(1-\alpha) \qquad (17)$$

where A, B, and B/ are the characteristic constants and α is the degree of dissociation of the ion pair. Eq.(2) can be rearranged to give.

$$[\eta_r - 1 - A(\alpha c)^{1/2}]/\alpha c = B + B/(1-\alpha/\alpha) \qquad (18)$$

For the evaluation of B from Eq.(3), a method suggested earlier by us (in Eq.(1)) was followed. The viscosity data have also been analyzed on the basis of the transition state treatment of the relative viscosity of electrolyte solutions [25, 26]. Relative viscosity ($\eta_r$) (a synonym of "viscosity ratio") is the ratio of the viscosity of a solution (η) to the viscosity of the solvent used ($\eta_0$), shown.

### *3.5 Refractive Index Calculation*

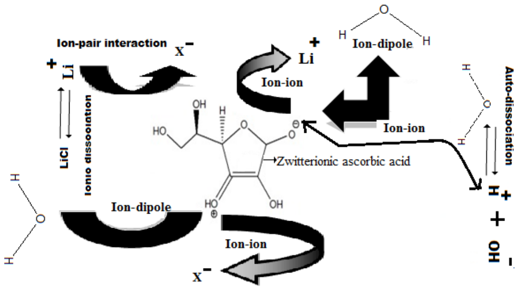
The molar refraction, $R_M$ can be evaluated from the Lorentz-Lorenz relation,

$$R_M = \{(nD^2 - 1)/(nD^2 + 2)\}(M/\rho) \qquad (19)$$

where $R_M$, $n_D$, M and ρ are the molar refraction, the refractive index, the molar mass and the density of solution respectively. The limiting molar refraction ($R^0_M$) estimated from the following equation,

$$R_M = R^0_M + R_S\sqrt{m} (20)$$

The refractive index measurement is also a convenient method for investigating the interaction occurring in the electrolytic solution. Stated more briefly by Deetlefs et al. that the refractive index of a compound describes its ability to refract light as it moves from one medium to another and thus, the higher the refractive index of a compound, the more the light is refracted. The values of refractive index $n_D$, molar refraction $R_M$, limiting molar refraction $R^0_M$ are reported in Table 3 respectively and Fig. 4. The refractive index of a substance is higher when its molecules are more tightly packed or in general when the compound is denser. The refractive index is directly proportional to molecular polarizability, scrutiny of Table 3 reveals the $n_D$ and $R_M$ values increases with an increasing concentration of mass fraction LiI into the solution of the used solvents, suggesting that in the ascorbic acid and water medium it is more tightly packed and more solvated. This is also in best agreement with the results obtained from apparent molar volume and viscosity B-coefficient and viscosity parameters discussed above [27, 28].



**Scheme 4** Different possible ionic group interactions in aqueous solutions of ascorbic acid in presence of LiX [X = Cl, Br, I]

**Table 3** Refractive index ($n_D$) , molar refraction ($R_M$), limiting molar refraction ($R^0_M$) of LiCl, LiBr, LiI in aqueous solution of ascorbic acid

| Solute | Solvent medium | $n_D$ | RM | $R^0_M$ |
|---|---|---|---|---|
| LiCl | Ascorbic Acid + Water | 1.3769 | 9.705 | 9.685 |
| | | 1.3776 | 9.715 | |
| | | 1.3781 | 9.723 | |
| | | 1.3787 | 9.733 | |
| | | 1.3792 | 9.744 | |
| | | 1.3797 | 9.754 | |
| LiBr | | 1.3775 | 19.824 | 19.98 |
| | | 1.3781 | 19.801 | |
| | | 1.3786 | 19.728 | |
| | | 1.3791 | 19.623 | |
| | | 1.3796 | 19.550 | |
| | | 1.3801 | 19.502 | |
| LiI | | 1.3779 | 30.705 | 30.65 |
| | | 1.3785 | 30.722 | |
| | | 1.3791 | 30.751 | |
| | | 1.3796 | 30.774 | |
| | | 1.3802 | 30.804 | |
| | | 1.3807 | 30.828 | |

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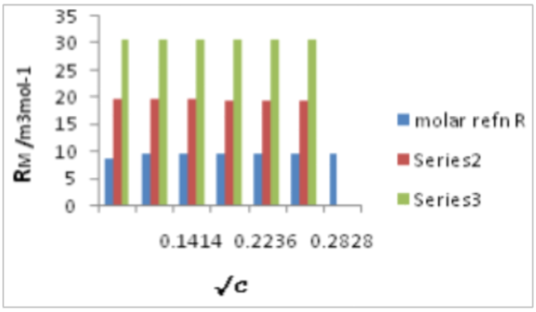


**Fig. 4** Plot of $R_M$ versus √c for alkali salt formers in ascorbic acid solutions: for LiCl (blue), LiBr (brown), LiI (green)

### *3.6 Challenges of Ionic Conductivity and Conductance Studies*

The concentrations and molar conductance's (Λ) of lithium halides in ascorbic acid and water at 298.15 K, 303.15 K, 308.15 K temperatures are given. The molar conductance (Λ) has been obtained from the specific conductance (κ) value using the following equation,

$$\Lambda = (1000\,\kappa) / c \qquad (21)$$

Linear conductance curve (Λ versus √c) were obtained for the electrolyte in aqueous ascorbic acid solution by extrapolation of √c to zero concentration, evaluated the starting limiting molar conductance for the electrolyte. Fig. 5a, b and c concludes at a particular concentration as the temperature is raised in the solute-solvent system then the conductance also rises steeply (i.e. the transport of the ions also rises steeply). The limiting molar conductance $\Lambda^0_m$ for LiCl, LiBr and LiI in (0.001, 0.003, 0.005) m aqueous ascorbic acid solutions were recorded, shows that limiting molar conductance increases with increase in temperature, it may be due to increase in ionic mobility with rise in temperature (Fig. 6). Since in the state of infinite dilution, the motion of an ion is limited solely by its interaction with surrounding solvent molecules, there are no other ions within a finite distance [29, 30]. Therefore evaluation of $\Lambda^0_m$ should give equally reliable information regarding ion–solvent interactions shown in Fig. 6. Greater value of $\Lambda^0_m$ may therefore be interpreted as a measure of greater ion–solvent interactions.

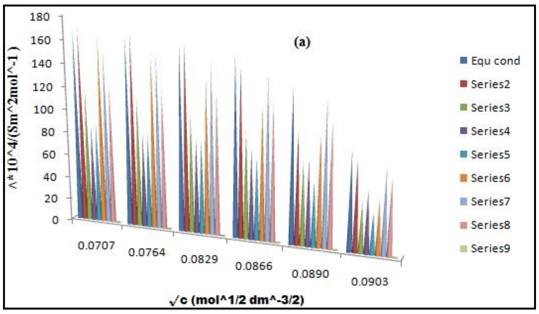


**Fig. 5(a)**Plot of molar conductance (Λ) and the square root of concentration (√c) of LiCl in $w_1$ = 0.003(m) shows series 1(■) at 308.15 K, series 2(■) at 303.15 K, series 3 (■) at 298.15 K, similar trends are also shown when $w_1$= 0.001(m) , $w_1$ =0.005(m) of ascorbic acid in water medium

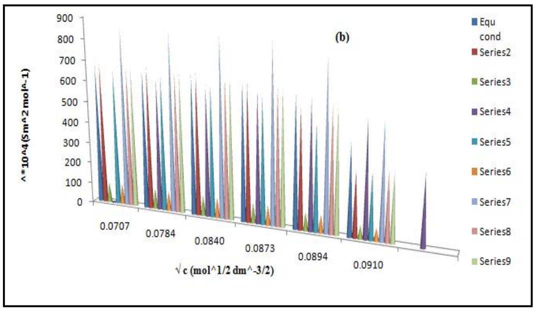


**Fig. 5(b)** Plot of molar conductance (Λ) and the square root of concentration (√c) of LiBr in $w_1$ = 0.005(m) shows series 1(■) at 308.15 K, series 2(■) at 303.15 K, series 3 (■) at 298.15 K, similar trends are also shown when $w_1$ = 0.001(m) , $w_1$ =0.003 (m) of ascorbic acid in water medium

The probable interaction is shown in Scheme 4. The order of solute–solvent interactions for LiCl, LiBr, LiI in (0.001, 0.003, 0.005) m aqueous ascorbic acid solutions follow as:

LiCl < LiBr < LiI

We know that hydration is inversely related with charge to radius ratio. Among all halide ions iodide has the largest ionic size than that of the other two concerned ions. That is why hydration is lowest in case of iodide ion and highest in case of chloride ion. Again ion – solvent interaction is inversely related to the extent of hydration. So, ion – solvent interaction is highest in case of iodide ion and is lowest in case of chloride ion (Scheme 5).

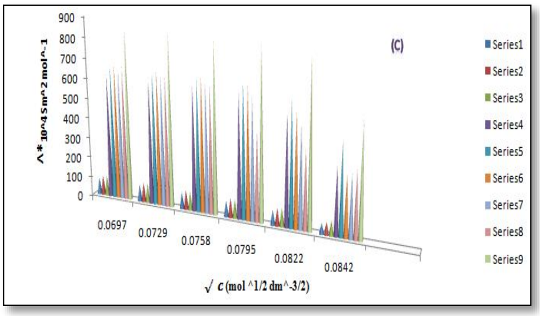


**Fig. 5(C)** Plot of molar conductance (Λ) and the square root of concentration (√c) of LiI in $w_1$ = 0.001(m) shows series 1(■) at 308.15 K, series 2(■) at 303.15 K, series 3 (■) at 298.15 K, similar trends are also shown when $w_1$ = 0.003(m), $w_1$ =0.005 (m) of ascorbic acid in water medium

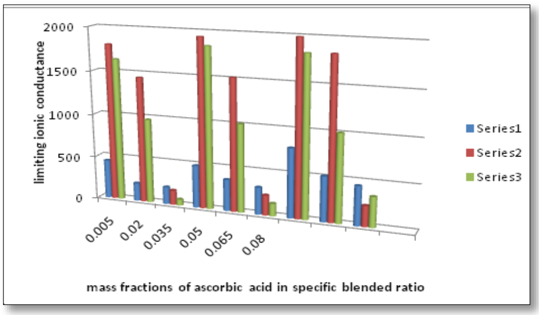
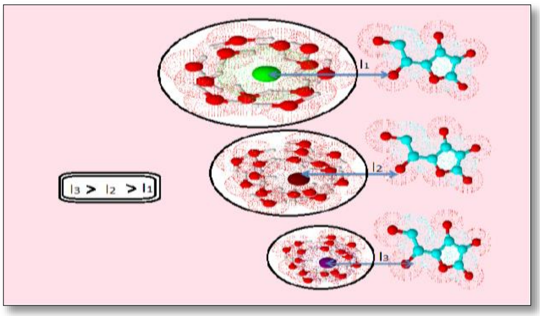


**Fig. 6** Plot of limiting molar conductance ($\lambda_o$) vs mass fraction (w) for LiCl, LiBr, LiI in definite 0.001(m), 0.003(m), 0.005(m) fractions at the variable ranges of temperatures of 298.15 K, 303.15 K, 308.15 K. Series1(■) denotes LiCl, Series2 (■) denotes LiBr, Series3 (■) denotes LiI in increasing trends of mass fractions from the left hand side of the graph



**Scheme 5** Ion-solvent interaction

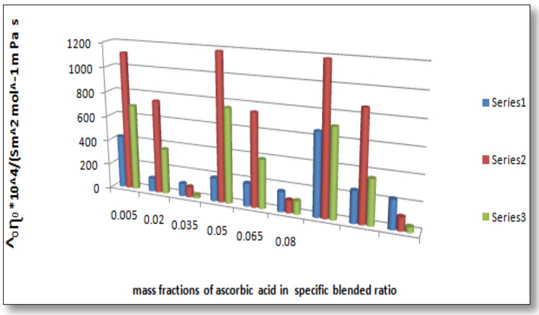


**Fig. 7** Plot of ionic Walden product ($\lambda_o\eta$) vs mass fraction (w) for LiCl, LiBr, LiI in definite 0.001(m) , 0.003(m), 0.005(m) fractions at the variable ranges of temperatures of 298.15 K, 303.15 K, 308.15 K. Series1(■) denotes LiCl , Series2 (■) denotes LiBr, Series3 (■ ) denotes LiI in increasing trends of mass fractions from the left hand side of the graph

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This may be explained on the basis of dissociation order of different metal halides as below:

LiI >LiBr > LiCl

The Walden product data ($\Lambda^0_m\eta^0$) have been recorded. The structure making/breaking nature of electrolyte has been determined from temperature coefficient of Walden product i.e. [$d(\Lambda^0_m\eta^0)/dT$]. The negative temperature coefficient of Walden product for LiCl, LiBr, LiI in (0.001, 0.003, 0.005) m aqueous ascorbic acid suggests that LiCl, LiBr, LiI behaves as structure-breaker in (0.001, 0.003, 0.005) m aqueous ascorbic acid which is depicted in Fig. 7.

### *3.7 Novelty and Characterization Following This Paper*

In this review discussion of a wide range of the representative cathode as 2-furanone and anode as lithium materials, starting from commercially available and currently used materials to promising novel materials that may be commercialized in the future. Fundamental properties, opportunities, challenges, and latest progress of anode and cathode material research have been discussed. It should be noted that even though flexible batteries have recently been of high interest, we will not cover this topic in this review, and as a substitute refer to other literature. An intercalation cathode is a solid host network, which can store guest ions. The guest ions can be inserted into and be removed from the host network reversibly. In a Li-ion battery, $Li^+$ is the guest ion and the host network compounds are aqueous ascorbic acid solution which may dissociate in solution and organize to form 2-furanone molecules. As new supplies and strategies are found, Li-ion batteries will no doubt have an ever greater impact on our lives in the years to come [31-33].

## 4. Conclusion

The Li-ion battery has essential advantages and decades of research which have developed it into the high energy density, high cycle life, high efficiency battery that it is today. So far research continues on new electrode materials to push the boundaries of cost, energy density, power density, cycle life, and safety. Various promising anode and cathode materials exist, but many suffer from limited electrical conductivity, slow Li transport, dissolution or other unfavorable interactions with electrolyte, low thermal stability, high volume expansion, mechanical brittleness. But these batteries would prove to be useful.

Studies of the transportation properties of different electrolytes in solvent media are of importance to obtain information on the behavior of ions in solution. Molecular interactions can be studied in the solution phase by studying it's thermodynamic and transport properties. These properties provide important information about the nature and strength of intermolecular forces operating among the components. The ion–ion interactions for LiCl, LiBr and LiI decreases with increase in temperature, which may be due to more solvation of ions (Schemes 3-5). The temperature effect on B coefficient for LiCl, LiBr and LiI shows a positive sign of dB/dT, showing thereby that LiCl, LiBr and LiI behaves a structure-breaker in (0.001, 0.003, 0.005) m aqueous ascorbic acid solution. This allows usage for a green battery with high capacity and high voltage. It also paves the way for cheaper consumer electronics.

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