Effect of Polarity on the Interaction Energies of some Organic Solvent (OS)-Water System and Formation of Donor-acceptor Complex: Quantum Mechanical MP4 Study

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Long-range or short-range van der Waals interaction in organic solvents (OS) is quite common; but, when an organic solvent or molecule directly interacts with a water molecule then it shows a significant change in the stability of such solvent-water system. Polar organic solvent/molecule shows strong hydrogen bonding interaction in aqueous phase whereas nonpolar solvent shows very weak interaction. In organic solvent-water complexes, solvent plays a very important role in the stability of such system, sometimes the solvent will interact with water molecule as proton donor or sometimes it will interact as proton acceptor mode. In gas phase, some common organic solvents such as aldehyde, ketone, carboxylic acid, cyclohexanol, amine, aniline, amide, *etc*. may interact with water molecules through different ways and we try to investigate the relative stability and actual mode of interaction of such organic solvent(OS)-water complexes by using computational method. Solvent polarity also plays an important role in such organic solvent (OS)-water complexes; therefore, the effect of polarity on the interacting counterpart need to study by changing the dielectric constant of the solvent system and thus it is necessary to analyze the effect of solvent on such OS-water systems. Quantum mechanical method, Møller-Plesset (MP) perturbation theory is an important and accurate tool for computing the interaction energy between the organic solvent-water complexes. The minute details of electron transfer pathway and donor-acceptor mechanism of such OS-water complexes have been studied by calculating single point energy and Natural Bond Orbital (NBO) analysis using MP4 methods.

Keywords: Organic solvent (OS), Hydrogen bonding, Water, MP4, NBO, *etc*.

INTRODUCTION

Past literature reveals that some organic solvents and water are miscible into a single liquid and have a transparent phase at constant temperature and pressure [1]. *Frank and Evans* proposed the so called "iceberg theory" in 1945 which discussed the entropic anomalies in several organic binary mixtures of water [2]. Further, the hydration of salt ions, the term "hydrophobic hydration" was developed to describe the "hydrate-like" or "ice-like" formations of water molecules [3-7]. As we know that the water molecule has a unique ability to form as multiple hydrogen bonds and it is supposed to be a reason for these abnormalities observed in binary mixtures [8]. The study of Raman spectra of water and organic solvents made possible the analysis of the hydrogen bond network (HBN) that water molecules develop with their neighboring organic solvent molecules [9]. By changing the ratio between the two components of binary mixtures of organic solvents and water we can simply change the solubilization and ionizations powers [10]. The mechanism of solvent action in a binary solvent can be readily achieved by studying solvatochromic indicators to analyze both solvent-probe and solvent-solvent interactions [11].

Acetic acid is a typical organic solvent which can act as both hydrogen bond donor as well as hydrogen bond acceptor [12]. It forms dimeric structures in gas phase and non-polar solvents owing to hydrogen bond interaction [13-

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14]. As water molecule has high dipole moment value and hence it has high ability to form hydrogen bonds, which has typical effects in solvation process. Many theoretical studies were done on micro-hydration of acetic acid, there have been on the hydrated clusters of acetic acids as well, *e.g.* micro hydration of Acetic acid dimer in chloroform, hydrogen bond in hydrates with acetic acid, *etc.* [15-17]. The intermolecular interactions between acetic acids and water molecules are also varying the numbers of H_2O molecules have been studied by using *ab initio* molecular dynamics simulations (AIMD) and other quantum chemical calculations [18]. Theoretical and experimental study reveals that in dilute solution of acetic acid, the more favorable molecular interaction between acetic acid and H_2O molecule is hydrogen bonding. The acetic acid molecule has both hydrophobic (methyl) and hydrophilic (carboxyl) sites, interactions involving the alkyl groups (-R) are rather weak while the carboxylic acid group (-COOH) is active to form hydrogen bond [19]. This feature of acetic acid leads to form a cyclic dimer with hydrogen bond in aqueous solution, an equilibrium between monomeric acetic acid and its cyclic dimer which was characterized by the dimerization constant (K_D) of the equilibrium [20-24]. On the other hand, some other studies have suggested that the favorable structure is acetic acid and water molecules rather than the acetic acid cyclic dimer, but in very dilute solutions, the interactions between the acetic acid and water molecules become more competitive [25-27].

Alcohol molecules show two types of solvent properties, self-association equilibria and the equilibria involving interactions with solutes. Hydrogen bond formation between two alcohol molecules lead to polarization of the hydroxyl groups which increase binding affinity for subsequent interactions with the unbound donor and acceptor sites. The overall solvation properties of alcohols depend on the speciation of different aggregates, the polarities of these species, and the polarities of the solutes. At low alcohol concentration, polar solutes are solvated by alcohol monomers and at higher alcohol concentration solutes are solvated by the more polar chain ends of linear polymers. Tertiary alcohols are marginally less polar solvents than primary alcohols, due to steric interactions that destabilize the formation of polymeric aggregates leading to lower concentrations of polar chain ends [28]. The aqueous

solution of formaldehyde at infinite dilution without hydration reaction was studied by different computational methods [29-30]. Aliphatic or aromatic amines are semi-volatile derivatives of ammonia which has molecular structure with a hydrophilic ($-NH_2$) and hydrophobic (-R/-Ar) group, it results in the amphiphilic character of alkyl/aryl-substituted ammonia [31]. Methylamine (MA) is an interesting solute found in the atmosphere which has an important physicochemical perspective since it combines one hydrophobic group (the methyl) connected to a more hydrophilic (the primary amine) one. Therefore, the study of mass accommodation of MA at the surface of liquid water together with a possible interpretation of the observed quantities in terms of the MA free energy of hydration and the molecular details of its solvation environment [32]. A lot of quantum mechanical investigation has been carried out for the hydrates of different amine isomers in the neutral and protonated forms by using DFT method. Frequency analysis of the amine-water complex reveals that the variation in the stretching frequency and intensity of the N–H and O–H bonds of the complex. This behavior is also explained by the application of the Natural Energy Decomposition Analysis (NEDA) and Natural Bond Orbitals (NBO) theory to the complexes and the dimeric form of amine-water pairs [33-34]. The $-NH_2$ group of an amine tends to originate with the water molecules a strong hydrogen bond *i.e.* O_w–H….N, and a much weaker hydrogen bond *i.e.* N–H….O_w. But, the presence of a water molecule leads to the formation of a cyclic structure stabilized by H-bonding. In contrast, the protonated amine group forms three strong and anticooperative N^+–H….O_w H-bonds with the water molecules but no water-water H-bonds are found to exist [34]. Moreover, DFT study on *p*-nitroaniline (PNA) reveals that the highest occupied molecular orbital (HOMO) electron density is mainly located on the $-NH_2$ group, whereas the lowest unoccupied molecular orbital (LUMO) electron density is on the $-NO_2$ group, indicating the donor and acceptor electronic transition. Due to its large difference in the ground state (6.1 D) and excited state (15 D) dipole moment, the absorption band is quite sensitive to the solvent polarity and other properties of the solvent medium [35-45]. Therefore, solvents with comparable polarity have been found to show significant synergistic behavior as a result of the formation

of stronger solvent-solvent hydrogen-bonded species [46].

In our current investigation, we are trying to analyze the van der Waals interaction or hydrogen bonding interaction exist between some simple organic solvents/molecules, such as aldehyde, ketone, amine, amide, carboxylic acid, *etc.,* with water molecule and also investigate the relative stability with the mode of interacting organic solvent (OS)-water complexes by using computational method. We also emphasize the effect of polarity on the interacting solvent counterpart by using different solvent systems and try to analyzes the solvent effect on such systems. The minute details of electron density shifting pathway and donor-acceptor mechanism of such organic solvent (OS)-water complexes have been studied by calculating single point energy and Natural Bond Orbital (NBO) analysis using MP4 methods.

METHODOLOGY

All the organic solvent (OS)-water models were constructed by using GaussView 5.0. We have studied the interaction energies between some selected organic solvents with water molecules by quantum mechanical and computational methods. Several OS-water complexes were constructed in such a way that they have different orientations with respect to each other and optimized these complexes in gas phase at room temperature. The OS-water complex structures were optimized by using Møller-Plesset (MP) perturbation methodology *i.e.* MP2 level of theory and double zeta basis set 6-311++G(d,p) at 298 K and 1atmospheric pressure. After optimization, the OS-water molecular complexes give constant intermolecular hydrogen bond distances and then interaction energies for all such complexes were calculated in gas phase and also with different solvent systems. All the single point energy for OS-water complexes and interaction energy calculations were computed by the MP4 level of theories with 6-311++G(d,p) basis set using Gaussian09 software package [47].

The organic solvent (OS)-water interaction energies (E_{IE}) are calculated by the following equation;

$$E_{IE} = E_{OS\text{-water}} - E_{OS} - E_{water}$$

Here, $E_{OS\text{-water}}$ = Energy for Organic Solvent-water complex, E_{OS} = Energy of Organic Solvent, E_{water} = Energy of water

Natural bond orbital (NBO) calculation is a very useful tool to study the interaction between organic solvent and water molecule of OS-water complexes. It was performed by using Gaussian09 to understand the electron charge delocalization in OS-water complexes during the interaction. Here, we also carried out the MP4 method for NBO analysis and the basis set was taken as MP4/6-311++G(d,p) POP = NBO. The NBO analysis was carried out for all possible hydrogen bonding interactions (O---H) in OS-water complexes. This interaction is taking place between "filled" Lewis-type NBOs (donor) and "empty" non-Lewis NBOs (acceptor) and calculating their energetic importance by perturbation theory. Since these interactions lead to donation of occupancy from the localized NBOs of the idealized Lewis structure into the empty non-Lewis orbitals, they are referred to as "delocalization" corrections to the zeroth-order natural Lewis structure. For each donor NBO (*i*) and acceptor NBO (*j*), the stabilization energy E(2) associated with delocalization ("2e-stabilization") *i* to *j* is estimated as;

$$E(2) = \Delta E_{ij} = q_i \frac{F(i,j)^2}{\varepsilon_j - \varepsilon_i}$$

Where; q_i is the donor orbital occupancy, e_i and e_j are diagonal elements (orbital energies), and $F(i,j)$ is the off-diagonal NBO Fock matrix element [48-50].

RESULT AND DISCUSSION

Interaction of Organic Solvent-water

Depending on the molecular structure and other parameters of organic solvents (OS), there might have more than one mode of hydrogen bonding interaction with water. In our investigation, we have taken organic solvents having –OH, –COOH, –CONH$_2$, and –NH$_2$ groups present in alcohol, carboxylic acid, amide, amine, *etc.* which generally shows two types of interactions with water molecules in OS-water complex. The first model depicts the interaction between O or N atom of –OH, –COOH, –CONH$_2$, and –NH$_2$ groups of solvent molecules with H atom of water molecule, whereas the second model depicts the interaction
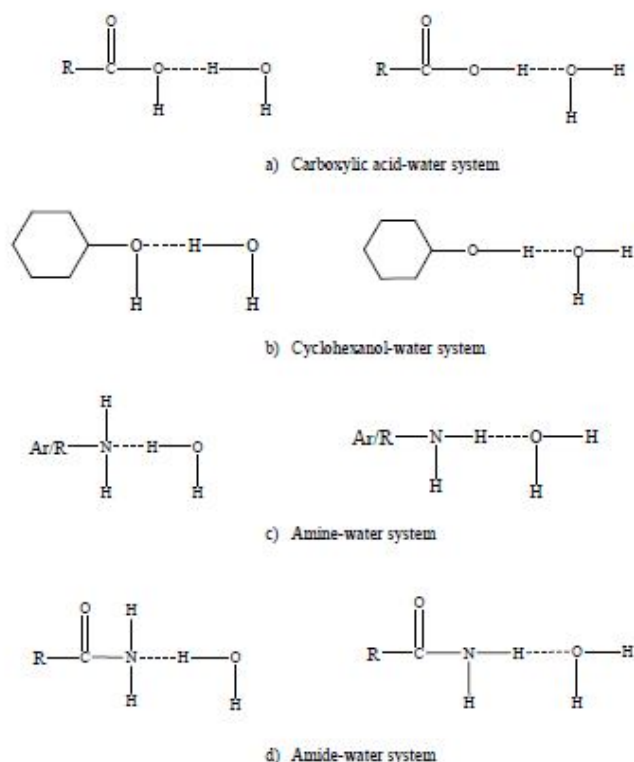
Fig. 1. Proton acceptor (Model-I) and proton donor (Model-II) mode of interaction for organic solvent-water complexes

Fig. 2. Optimized models of different organic molecules-water complexes.

between H atom of –OH, –COOH, –CONH$_2$, and –NH$_2$ groups of solvent molecule with O atom of water molecule as shown in Fig. 1.

In other words, model-I of OS-water complexes, the oxygen atom of –OH, –COOH, –CONH$_2$, or –NH$_2$ group in solvent is acting as *proton acceptor* as RCOHO---H$_2$O, RHO---H$_2$O, RCOH$_2$N---H$_2$O, and RH$_2$N---H$_2$O. However, model-II is proton donor and these are RCOOH---OH$_2$, ROH---OH$_2$, RCONH$_2$---OH$_2$, and RNH$_2$---OH$_2$, these models interact through the H atom of solvents with O atom of water molecule. Herein, we have investigated the interactions between some organic solvents like formic acid, acetic acid, propanoic acid, chloroacetic acid, cyclohexanol, amine, amide, *etc.* with water molecules as a pair of interacting molecular complexes. The optimized models for all such OS-water models are shown in Fig. 2.

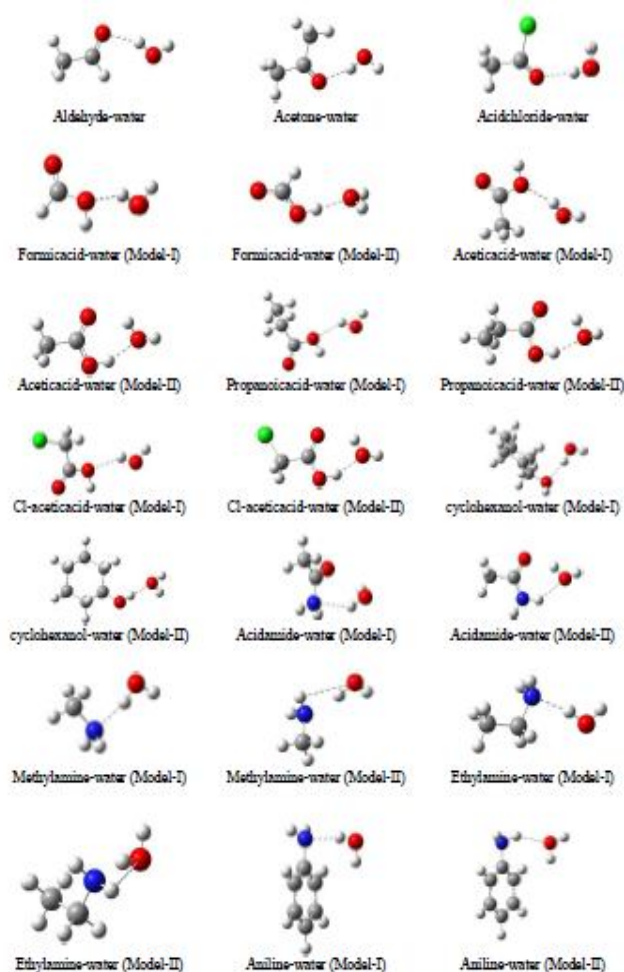At first, we computed the interaction energy of some selected organic solvent molecules and water complexes in

gas phase to compare their interaction energy values with different solvent systems. From these calculations, we have found that the interaction energy for model-II (proton donor) for all of the carboxylic acid-water complexes results in more negative value (-9.36 to 11.08 kcal mol^{-1}) than that of model-I (-2.55 to -6.91 kcal mol^{-1}) as the proton donor mode is theoretically found more favorable than that of proton acceptor mode (Table 1). Contrary to that, in cyclohexanol-water complex the proton acceptor mode, *i.e.* model-I results in more negative interaction value *i.e.* -7.48 kcal mol^{-1}, than that of the interaction energy value (-6.26 kcal mol^{-1}) for proton donor mode *i.e.* model-II as shown in Table 1. Interestingly, cyclohexanol-water shows the reverse trend than that of the studied carboxylic acid-

water complexes (Table 1). It happens due to the molecular conformation of proton acceptor mode (model-I) of cyclohexanol-water complex, which is sterically found more favored as compared to the conformation of proton donor model (model-II). The interaction energy of proton donor and acceptor modes of such systems also explain why the carboxylic acid shows more acidic in nature than that of cyclohexanol in aqueous phase.

Moreover, in gas phase, the computed interaction energy for OS-water complex and its stability is also dependent on the intermolecular hydrogen bond distances between two entities of such complexes. It is well established that shorter is the hydrogen bond in OS-water complexes stronger is the interaction and hence it results in the more negative interaction energy values which can easily be reflected in Tables 1 and 2.

But, due to some other factors such as electron withdrawing effect, electron releasing effect, electronegativity, resonance, *etc.* of an organic molecule may also affect on OS-water hydrogen bonding distances. Therefore, the interaction energy values of such OS-water systems and their respective hydrogen bond distances might not always follow a similar trend.

Thus, the computed interaction energy trend of proton donor mode (model-II) for different carboxylic acid-water complexes in gas phase is found as;

Formic acid < Acetic acid ~ Propanoic acid < Chloroacetic acid

Among all carboxylic acid-water complexes, the formic acid-water complex shows the least negative interaction energy value *i.e.* -9.36 kcal mol^{-1}, which results in the less stable conformation exists for the formic acid-water complex than that of other acids. It happens because of the highly acidic nature of formic acid (calculated pK_a = 3.75) and hence the formic acid-water molecular complex becomes unstable. On the other hand, the interaction energy value for chloroacetic acid-water complex results from the most negative value *i.e.* -11.08 kcal mol^{-1}, it happens because of the presence of highly electronegative or electron withdrawing –Cl group attached at the acetate ion, which enhances the interaction energy of this pair. But, interestingly the interaction energy value for acetic acid and propanoic acid-water is found as an average of ~ -10.45 kcal mol^{-1}, due to the steric and electron-donating

Table 1. Computed Interaction Energies (kcal mol^{-1}) of Organic Solvent-water Complexes in Gas Phase at 298 K and 1 Atmospheric Pressure (MP4/6-311++G(d,p))

Solvent-water complexes	Interaction energies (kcal mol^{-1})	
Acetaldehyde-water	-5.82	
Acetone-water	-6.62	
Acetyl chloride-water	-4.02	
	Model-I	Model-II
Ethanamide-water	-3.99	-9.86
Methenamine-water	-7.94	-7.27
Ethanamine-water	-7.97	-7.74
Aniline-water	-7.04	-5.27
Formic acid-water	-6.91	-9.36
Acetic acid-water	-4.89	-10.44
Cl-Acetic acid-water	-2.55	-11.08
Propanoic acid-water	-3.62	-10.46
Cyclohexanol-water	-7.48	-6.26

Table 2. Computed Equilibrium Hydrogen Bond Distances for Optimized Organic Solvent-water Complexes in Gas Phase (MP2/6-311++G(d,p))

OS-Water complexes	OS-Water H-bond distances (Å)	
Acetaldehyde-water	1.971	
Acetone-water	1.942	
Acetyl chloride-water	2.113	
	Model-I	Model-II
Ethanamide-water	2.406	1.904
Methenamine-water	1.918	1.923
Ethanamine-water	1.920	1.932
Aniline-water	2.063	2.169
Formic acid-water	1.887	1.817
Acetic acid-water	2.042	1.806
Cl-Acetic acid-water	2.154	1.771
Propanoic acid-water	2.120	1.806
Cyclohexanol-water	1.906	1.958

effect caused by the –CH$_3$ and –C$_2$H$_5$ group present in these acids.

Moreover, for some $-NH_2$ containing molecules such as amine or aniline we have found a different picture of the stability and interaction energy for amine-water complexes. Although, both the proton acceptor and proton donor modes are possible in all such amine-water complexes but depending on their molecular structures, steric factor, and other parameters they show a selective mode of interaction with water molecules in gas phase. Therefore, the computed interaction energy of methenamine, ethanamine, and aniline-water complexes, proton acceptor mode *i.e.* model-I, results in more negative interaction energy values -7.94, -7.97 and -7.04 kcal mol^{-1}, respectively than that of the proton donor mode (model-II) *i.e.* -7.27, -7.74 and -5.27 kcal mol^{-1}. For these OS-water complexes, it is clearly observed that the proton acceptor mode (model-I) of interaction is theoretically found more favored than that of proton donor mode as shown in Table 1. Interestingly, aliphatic amines such as methenamine and ethanamine-water complexes result more negative interaction energy values for model-I, because the lone pair of electrons on N-atom is always intensified by alkyl group on it and that makes it more basic character in nature. But, in aromatic amine, *i.e.* aniline-water complex gives less negative interaction energy (-7.04 kcal mol^{-1}) than that of aliphatic amine-water complexes, it happens because the lone pair electrons on $-NH_2$ group of aniline undergoes delocalization which makes it less basic in nature. Contrary to that, for ethanamide-water complex the proton donor mode (model-II) results in much more negative interaction energy -9.86 kcal mol^{-1} than that of the proton acceptor mode (model-I) *i.e.* 3.99 kcal mol^{-1}, it happens due to its mild acidic nature of $-CONH_2$ functional group (Table 1). The computed interaction energy trend of proton acceptor mode (model-I) for different amine-water complexes in gas phase is found as;

Aniline < Methenamine ~ Ethanamine

On the other hand, organic solvents like acetaldehyde, acetone, and acetyl chloride; show only one type of hydrogen bonding interaction with water molecules as depicted in Fig. 2, we also investigated the interactions of such OS-water complexes. In gas phase acetone-water complex, the acetone molecule interacts with H atom of H_2O molecule through $>C=O$ group *i.e.* the hydrogen bond exists in the fashion $>C=O---H-OH$. The two $-CH_3$ groups

of acetone molecule also enhances the electron charge density on the O atom of $>C=O$ group and computed Mulliken electron charge density of $>C=O$ group is -0.362, as a result of which there exist a strong interaction between acetone-water complex, it is clearly reflected in the interaction energy value *i.e.* -6.62 kcal mol^{-1}. Similarly, for aldehyde-water complex, the aldehyde molecule directly interacts with H atom of H_2O molecule through $>C=O$ group as observed in acetone-water complex; a single methyl ($-CH_3$) group of aldehyde molecule enhances electron charge density on the O atom of $>C=O$ group, but the hydrogen atom does not impact much on $>C=O$ functional group. The computed Mulliken electron charge density on O atom $>C=O$ functional group is found as -0.336, which results in lesser interaction in aldehyde-water complex *i.e.* -5.82 kcal mol^{-1} than that of acetone-water molecular complex. Similarly, for acid chloride-water complex the hydrogen bond exists in the fashion $>C=O---H-OH$. Here, one electron-donating $-CH_3$ group on $>C=O$ functional group enhances the electron charge density of acid chloride molecule whereas the strong electronegative $-Cl$ group reduces the electron charge density on the $>C=O$ functional group and the computed Mulliken electron charge density on O atom of $>C=O$ group is -0.213, as a result of which acid chloride-water complex shows least negative interaction energy value *i.e.* -4.02 kcal mol^{-1} than that of aldehyde-water and acetone-water complexes. The computed hydrogen bonding interaction energy trend for these complexes in gas phase is found as;

Acid chloride < Acetaldehyde < Ketone

Effect of Polarity on Organic Solvent(OS)-water Interaction

The polarity of solvent plays an important and crucial role in the interaction of OS-water complexes. Therefore, it is also important to study the interaction of an individual OS-water complex in different solvent systems. For our investigation, the interaction of OS-water complexes has been extensively studied in different solvent systems, *i.e.* from the least polar to higher polar solvents. The polarity of a solvent can be measured by the value of dielectric constants (K) of the solvent and thus the polarity trend of

Table 3. Dielectric Constants for Different Solvents at 298 K and 1 Atmospheric Pressure

Solvents	Dielectric constants (K)
Cyclohexane	2.02
Carbon tetrachloride	2.24
Toluene	2.38
Chloroform	4.81
Tetrahydrofuran (THF)	7.58
Dichloromethane (DCM)	8.93
Acetone	20.70
Methanol	32.70
Dimethyl sulfoxide (DMSO)	46.68
Water	80.10

the solvents is shown in Table 3. Herein, the sequence of the polarity of some selected solvents are shown below:

Cyclohexane < Carbon tetrachloride < Toluene < Chloroform < THF < DCM < Acetone < Methanol < DMSO < Water

The computed interaction energies for OS-water complexes reveal that as the dielectric constant (polarity) of the solvent increases, the interaction energy value of a

particular OS-water complex also changes to a great extent *i.e.* the polarity of a solvent is highly sensitive toward the interaction energy value of OS-water complexes. Current studies revealed that for all the selected OS-water complexes the interaction energy value is found to be more negative in cyclohexane medium, which is due to the low dielectric constant value *i.e.* 2.02 or least polar cyclohexane, whereas the interaction energy value is found to be less negative in highly polar water medium due to its high dielectric constant *i.e.* 80.10 (Tables 4a-c).

In less polar solvents, OS-water complexes result in effective interaction due to the stronger van der Waals interaction between them, whereas in highly polar solvent the interaction between OS-water complex reduces and it happens due to the very strong interaction between OS or water with the external solvent molecule and it can be revealed by the computed interaction energy values shown in Tables 4a-c.

As stated earlier, in the gas phase OS-water molecule interaction, similar proton acceptor and proton donor mode of binding were also investigated in the selected solvent phases and almost similar trend of interaction energy has been found as we have already observed in the gas phase interaction energy calculations. Among all the solvent systems, for carboxylic acid-water complexes, model II

Table 4a. Computed Interaction Energies (kcal mol^{-1}) of Organic Solvent-water Complexes in Different Solvent Systems at 298 K and 1 Atmospheric Pressure (Basis Set MP4/6-311++G(d,p))

Solvents	Interaction energies (kcal mol^{-1})		
	Acetaldehyde-water	Acetone-water	Acetyl chloride-water
Cyclohexane	-4.93	-5.59	-3.66
Carbon tetrachloride	-4.82	-5.47	-3.60
Toluene	-4.75	-5.39	-3.57
Chloroform	-4.15	-4.75	-3.25
THF	-3.88	-4.47	-3.10
DCM	-3.80	-4.38	-3.05
Acetone	-3.55	-4.13	-2.90
Methanol	-3.48	-4.05	-2.86
DMSO	-3.44	-4.01	-2.83
Water	-3.40	-3.97	-2.81

Table 4b. Computed Interaction Energies (kcal mol^{-1}) of Organic Solvent-water Complexes in Different Solvent Systems at 298 K and 1 Atmospheric Pressure (Basis Set MP4/6-311++G(d,p))

Solvents	Interaction energies (kcal mol^{-1})							
	Aniline-water		Ethanamide-water		Methenamine-water		Ethanamine-water	
	Model-I	Model-II	Model-I	Model-II	Model-I	Model-II	Model-I	Model -II
Cyclohexane	-6.68	-4.13	-3.16	-8.04	-7.44	-6.83	-7.56	-7.33
Carbon tetrachloride	-6.62	-3.99	-3.05	-7.82	-7.37	-6.76	-7.50	-7.25
Toluene	-6.58	-3.90	-2.99	-7.70	-7.32	-6.72	-7.46	-7.19
Chloroform	-6.16	-3.18	-2.44	-6.59	-6.92	-6.30	-7.10	-6.73
THF	-5.94	-2.87	-2.19	-6.12	-6.73	-6.09	-6.92	-6.52
DCM	-5.81	-2.77	-2.06	-5.97	-6.58	-6.02	-6.78	-6.45
Acetone	-5.65	-2.48	-1.91	-5.54	-6.48	-5.81	-6.70	-6.24
Methanol	-5.57	-2.39	-1.84	-5.41	-6.43	-5.75	-6.64	-6.18
DMSO	-5.54	-2.35	-1.81	-5.35	-6.40	-5.71	-6.62	-6.15
Water	-5.50	-2.31	-1.78	-5.28	-6.37	-5.68	-6.59	-6.12

Table 4c. Computed Interaction Energies (kcal mol^{-1}) of Organic Solvent-water Complexes in Different Solvent Systems at 298 K and 1 Atmospheric Pressure (Basis Set MP4/6-311++G(d,p))

Solvents	Interaction Energies (kcal mol^{-1})									
	Formic acid-water		Acetic acid-water		Chloroacetic acid-water		Propanoic acid-water		Cyclohexanol-water	
	Model-I	Model -II	Model-I	Model-II	Model-I	Model-II	Model-I	Model-II	Model-I	Model-II
Cyclohexane	-3.85	-8.87	-3.97	-8.72	-2.01	-9.39	-2.87	-8.74	-6.55	-5.82
carbon tetrachloride	-3.75	-8.78	-3.86	-8.52	-1.95	-9.19	-2.78	-8.54	-6.44	-5.76
Toluene	-3.69	-8.72	-3.80	-8.40	-1.91	-9.06	-2.72	-8.42	-6.37	-5.72
Chloroform	-3.11	-8.15	-3.27	-7.38	-1.58	-8.02	-2.27	-7.39	-5.78	-5.33
THF	-2.84	-7.86	-3.06	-6.94	-1.45	-7.57	-2.08	-6.96	-5.52	-5.15
DCM	-2.75	-7.76	-2.99	-6.81	-1.41	-7.43	-2.02	-6.83	-5.44	-5.09
Acetone	-2.49	-7.47	-2.80	-6.42	-1.29	-7.02	-1.85	-6.44	-5.20	-4.91
Methanol	-2.41	-7.38	-2.75	-6.30	-1.25	-6.90	-1.80	-6.32	-5.13	-4.85
DMSO	-2.36	-7.33	-2.72	-6.24	-1.24	-6.83	-1.77	-6.26	-5.10	-4.82
Water	-2.32	-7.29	-2.69	-6.19	-1.22	-6.77	-1.75	-6.21	-5.06	-4.79

(proton donor) shows more negative interaction energy value than that of model I (proton acceptor) in cyclohexane medium, which is due to proton donating nature of the carboxylic acid. Again, in the case of cyclohexanol-water

complex, model I (proton acceptor) shows more negative interaction energy value (-6.55 kcal mol^{-1}) than that of model II (proton donor) *i.e.* -5.82 kcal mol^{-1} in cyclohexane medium because cyclohexanol has less proton donor tendency than carboxylic acids as shown in Table 4c. Similar results of interaction energy have been obtained in the case of amines and other OS-water complexes. Hence, for all such OS-water systems, the interaction energy values of an individual OS-water complex gradually decrease to less negative value on increasing the polarity of the solvent. Thus, the interaction energy value for OS-water complex is found more negative in the least polar cyclohexane medium, whereas the same OS-water complex shows less negative interaction energy value in highly polar solvent *i.e.* water (Tables 4a-c).

The overall trend of computed interaction energy or stability of OS-Water complexes in different solvent systems for proton acceptor or donor modes are shown below:

Formic acid < Acetic acid ~ Propanoic acid < Chloroacetic acid

Acid chloride < Acetaldehyde < Ketone
Aniline < Methenamine ~ Ethanamine

HOMO-LUMO Energy Analysis

The calculation of interaction energies for different OS-water complexes may also be justified by computing their HOMO-LUMO energy gap. The frontier orbitals of a molecule i.e., HOMO and LUMO are the most important orbitals in a molecule or molecular complex. The highest occupied energy of HOMO characterizes electron-donating ability of a molecule while the lowest unoccupied LUMO energy determines the ability to accept an electron. The computed energy of these orbitals determines the way how a molecule interacts with other species and it gives the information about stability or reactivity of specific regions of the molecule. Moreover, from the HOMO-LUMO energy of a molecular system, we could determine the chemical reactivity descriptors, chemical potential (μ), electronegativity (χ), hardness (η), softness (S), electrophilicity index (ω), *etc*. The HOMO-LUMO energies for OS-water complexes in gas phase were obtained by the MP4 level of theories and the values are incorporated in

Table 5. HOMO-LUMO Energy (ev) of Organic Solvent-Water Complexes in Gas Phase at 298 K and 1 Atmospheric Pressure

Solvent-water complexes	HOMO-LUMO energy (ev)	
Acetaldehyde-water	5.11	
Acetone-water	5.57	
Acetyl chloride-water	4.74	
	Model-I	Model-II
Ethanamide-water	6.51	6.58
Methenamine-water	7.88	7.82
Ethanamine-water	8.01	7.82
Aniline-water	4.15	3.92
Formic acid-water	6.54	6.55
Acetic acid-water	6.86	6.91
Cl-Acetic acid-water	6.89	6.96
Propanoic acid-water	6.62	6.66
Cyclohexanol-water	8.35	8.17

Table 5.

The higher values of HOMO and LUMO energy (ev) indicate that the OS-water complex is chemically stable, while a small HOMO-LUMO energy gap represents the small excitation energies to the manifold of excited states. Therefore, we can tally the computed interaction energies of the above studied OS-water complexes with their HOMO-LUMO energy to justify the calculated interaction energy and it is found that all those complexes which have more negative interaction energy values also results from the higher HOMO-LUMO energy as shown in Tables 1 and 5. Although, we could not compute the accurate HOMO-LUMO energy for all such molecular complexes, still in most cases higher HOMO-LUMO energy results the stable OS-water complex formation with more negative interaction energy value. The computed HOMO-LUMO energy gap for acid chloride, aldehyde, and acetone-water complexes are found as 4.74, 5.11, and 5.57 ev, respectively, which clearly justifies their trend in stability and the interaction energy values also follow almost similar trend. Similarly, for carboxylic acid-water systems, the computed HOMO-LUMO energy for proton donor mode (model-II) is found within a range of ~ 6.55-6.96 ev, which is also higher than

that of proton acceptor mode (model-I) *i.e.* ~ 6.54-6.89 ev. Moreover, amongst all carboxylic acid-water complexes chloroacetic acid-water system results in the maximum HOMO-LUMO energy value *i.e.* 6.96 ev (model-II) and it also gives more negative interaction energy value *i.e.* -11.08 kcal mol^{-1}, which results in the stable complex formation as compared to the other systems. On the other hand, the formic acid-water system shows the least HOMO-LUMO energy *i.e.* 6.55 ev (model-II) which also gives the least negative interaction energy *i.e.* -9.36 kcal mol^{-1}, it clearly shows the least stable complex formation with water due to the strong acidic nature of formic acid. From the above HOMO-LUMO energy calculation, we have observed almost similar trend as we observed in interaction energy calculation. Thus, similar stability trend is observed in cyclohexanol-water, amine-water, and other OS-water complexes.

Natural Bond Orbital (NBO) Analysis

From the electronic stabilization energy E (2) values in NBO analysis, we could have a brief idea about the molecular interaction taking place between the interacting counterparts within a molecular complex. Delocalization of electron charge density between occupied Lewis type (bonding or lone pair) NBO orbitals and formally unoccupied (antibonding or Rydberg) non-Lewis NBO orbitals correspond to a stabilizing donor-acceptor interaction. All the OS-water complexes show there exist either van der Waals or hydrogen bonding interaction between organic solvent and water molecule, amongst these molecular interactions some are strong hydrogen bonds and few are weak. There observe two possibilities of shifting of electron charge density between the entities of a molecular complex, *i.e.* the shifting of electron is taking place either from organic solvent to H$_2$O molecule or from H$_2$O molecule to organic solvent, which would finally form a stable donor-acceptor complex. Such donor-acceptor electronic stabilization energy (E2) of the interacting sites for all the OS-water complexes are shown in Table 6. From the different interacting sites in OS-water complexes, it is clearly observed that the lone pairs (LP) of O or N (either from organic solvent or water) behaves as electron pair donor whereas the O–H antibonding (BD*) orbital behaves as electron pair acceptor. NBO analysis shows that higher

is the value of computed electronic stabilization energy E(2) in OS-water complexes results from favorable donor-acceptor complex formation and the possible mode of interaction is either through van der Waals interaction or hydrogen bonding (Table 6). Thus, the NBO analysis for the calculation of stabilization energy of OS-water complexes gives a brief idea about the donor-acceptor stabilization energy and therefore, we could predict the stability and the way of shifting of electron cloud density within the OS-water complexes. As analyzed in the interaction energy calculation of different OS-water complexes shown in Table 1, the proton donor mode(model-II) of different acidic systems like carboxylic acid-water complexes show effective hydrogen bonding interaction in gas phase than that of proton acceptor mode (model-I) which is also clearly reflected in the calculation of stabilization energy in NBO analysis (Table 6). For all the studied carboxylic acid-water molecular complexes it observes that the electron density is shifted from the water molecule (lone pair of O) to the O–H bond of the carboxylic acid as shown in Table 6.

The interaction energy and stabilization energy in NBO analysis for proton donor mode (model-II) follows almost similar trend as shown below:

Formic acid < Acetic acid ~ Propanoic acid < Chloroacetic acid

Whereas for another basic system like amine-water complexes the lone pair (LP) of N atom of amine (–NH$_2$) becomes electron donor and O–H bond (antibonding BD*) of water molecule is electron acceptor. But, for aromatic amine, the weak base aniline-water system the NBO stabilization energy is quite low *i.e.* 7.14 than that of aliphatic amine methenamine-water (14.76) and ethanamine-water (14.83) systems. Hence, the aliphatic amine-water complex may show strong molecular interaction than that of aromatic amine-water complex. The molecular structure, polarity, and steric effect of organic solvents may also change the NBO stabilization energy of OS-water complexes, and it can directly affect the stability of such hydrogen bonding interaction in such system. Similarly, for such basic systems the interaction energy and stabilization energy for donor-acceptor complex in NBO analysis follow almost similar trend as shown below:

Aniline < Methenamine ~ Ethanamine

On the other hand, amongst all the studied OS-water

Table 6. NBO Calculations for Organic Solvent-Water Complexes in Gas Phase (MP4/6-311++G(d,p) POP=NBO)

Complexes	Models	Donor	Acceptor	Stabilization energy
Acetaldehyde-water		LP(2)O6(AL)	BD*(1)O8-H10 (W)	5.63
Acetone-water		LP(2)O6(AC)	BD*(1)O7-H9(W)	5.98
Acetyl chloride-water		LP(2)O6(ACL)	BD*(1)O7-H9(W)	1.72
Ethanamide	Model-I	LP(1)N4(EAD)	BD*(1)O1-H3(W)	1.18
	Model-II	LP(2)O4(EAD)	BD*(1)O6-H7(W)	7.69
		LP(2)O6(W)	BD*(1)N1-H2(EAD)	4.66
Methenamine-water	Model-I	LP 1)N4(MAN)	BD*(1)O1-H3(W)	14.76
	Model-II	LP(1)N1(MAN)	BD*(1)O4-H6(W)	14.46
Ethanamine-water	Model-I	LP(1)N4(EAN)	BD*(1)O1-H3(W)	14.83
	Model-II	LP(1)N1(EAN)	BD*(1)O4-H6(W)	13.79
Aniline-water	Model-I	LP(1)N4(AN)	BD*(1)O1-H3(W)	7.14
	Model-II	LP(2)O15(W)	BD*(1)N12-H14(AN)	2.67
Formic acid-water	Model-I	LP(1)O1(FA)	BD*(1)O3-H5(W)	1.67
	Model-II	LP(2)O3(W)	BD*(1)O1-H2(FA)	15.26
Acetic acid-water	Model-I	LP(1)O1(AA)	BD*(1)O3-H5(W)	2.37
	Model-II	LP(2)O3(W)	BD*(1)O1-H2(AA)	15.84
		LP(2)O11(AA)	BD*(1)O3-H5(W)	3.44
Cl-Acetic acid-water	Model-I	LP(1)O1(CAA)	BD*(1)O3-H5(W)	1.10
	Model-II	LP(2)O3(W)	BD*(1)O1-H2(CAA)	18.66
Propanoic acid-water	Model-I	LP(1)O1(PA)	BD*(1)O3-H5(W)	1.14
	Model-II	LP(2)O3(W)	BD*(1)O1-H2(PA)	15. 76
Cyclohexanol-water	Model-I	LP(2)O1(CH)	BD*(1)O3-H5(W)	7.91
	Model-II	LP(2)O3(W)	BD*(1)O1-H2(CH)	7.72

AL: Acetaldehyde, AC: Acetone, ACL: Acetyl chloride, EAD: Ethanamide, MAN: Methenamine, EAN: Ethanamine, AN: Aniline, FA: Formic acid, AA: Acetic acid, CAA: Cl-Acetic acid, PA: Propanoic acid, CH: Cyclohexanol, W: water and LP: Lone pair, BD*: Antibonding orbital.

complexes acetyl choride-water complex results very low NBO stabilization energy *i.e.* 1.72, which shows the formation of very weak donor-acceptor complex between the lone pair of O of acetyl choride with bonding pair of –OH of water. It happens because of the presence of highly electronegative –Cl atom in acetyl chloride, which may reduce the interaction in acetyl chloride-water complex. Again, the NBO stabilization energy for aldehyde-water (5.63) and acetone-water (5.98) complex shows that there exist strong molecular interaction with water molecule, *i.e.* the donor-acceptor complex is formed between the lone pair (LP) of O (aldehyde or ketone) as a donor to the antibonding orbital (BD*) of –OH group of H_2O molecule

as an acceptor, it happens due to the presence of electron donating –CH_3 group in aldehyde and ketone. The details of NBO stabilization energy and shifting of electron charge density to form a stable donor-acceptor complex within OS-water complexes are clearly displayed in Table 6.

As usual, we also found the similar trends of interaction energy and stabilization energy in NBO analysis for donor-acceptor complex is shown below:

Acid chloride < Acetaldehyde < Ketone

Therefore, the calculation of stabilization energy in NBO clearly justifies the interaction energy calculation as well as the formation of stable donor-acceptor complexes in OS-water system in gas and solvent phase.

CONCLUSIONS

In carboxylic acid-water complexes proton donor model (model II) shows more negative interaction energy value than that of proton acceptor (model I) for gas and all solvent phases which results in more favored systems are proton donor mode. But, interestingly in the case of cyclohexanol-water complex as the cyclohexanol has less proton donor tendency than carboxylic acids, hence the proton acceptor mode (model-I) shows more negative interaction energy value. The computed stabilization energy in NBO and HOMO-LUMO energy also justifies the calculated interaction energy as well as the formation of stable donor-acceptor complexes in OS-water system in gas and different solvent phase. But, for some basic and electron donation systems such as amine-water and aniline-water complexes proton acceptor mode (model-I) is favored as it results from more negative interaction energy values. Moreover, above studies reveals that in some OS-water systems where no such direct proton donor or acceptor mode exist, their electron-donating and electron accepting group controls the interaction energies of such systems. Thus, we have observed that more negative interaction energy in ketone with two electrons-donating $-CH_3$ group, whereas least negative interaction energy values observed in acid chloride with one electron withdrawing $-Cl$ group. Most importantly, the polarity of solvent plays major role in the interaction energy of OS-water complexes. For the above investigation, it has been revealed that the interaction of OS-water complexes has been changed in different solvent systems *i.e.* in the least polar and more polar solvents. The computed interaction energy value is found to be more negative in least polar cyclohexane medium and the interaction energy value is found to be less negative for highly polar water medium. Hence, for all OS-water complexes, the favored model is found in the solvents with low polarity.

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