

## Physicochemical Properties and Structural Analysis of Some Organic Solvent (OS)-Water Systems: A Quantum Mechanical Study

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Molecular interaction in organic solvents and water causes a significant change in physicochemical properties of organic solvent (OS)-water systems, which may also affect the stability of such systems. Although optimized models of OS-water systems have been shown to have strong hydrogen bond formation, some changes may occur in bond distances and angles following a molecular interaction between solvents and H<sub>2</sub>O or H<sub>3</sub>O<sup>+</sup> ions. Organic solvents, such as aldehyde, ketone, carboxylic acid, amine, aniline, amide, etc., might have different interaction modes with water molecules, but the proton donor-acceptor nature of such systems greatly impacts their free energy, enthalpy, zero-point energy, and other thermodynamic parameters. The analysis of enthalpy changes ( $\Delta H^\circ$ ) for all OS-water systems showed negative values, which reflects their exothermic behavior; however, the Gibbs free energy change ( $\Delta G^\circ$ ) shows the feasibility of the formation of all types of OS-water systems. The hydrophobic or hydrophilic nature of such systems not only controls the stability of donor-acceptor complexes but also defines their acidic or basic nature. In all cases, optimized models provide some detailed information about the changes in the geometry of OS-water mixtures. Furthermore, the physicochemical properties of OS-H<sub>3</sub>O<sup>+</sup> and OS-H<sub>2</sub>O systems have been found to follow a similar trend. In this study, the standard free energy change ( $\Delta G^\circ$ ), enthalpy change ( $\Delta H^\circ$ ), zero-point energy change ( $\Delta ZPE$ ), the effect of protonation, electron charge density, and other parameters of OS-H<sub>2</sub>O and OS-H<sub>3</sub>O<sup>+</sup> systems were evaluated by quantum mechanical calculations.

**Keywords:** Organic solvents (OS), H<sub>2</sub>O, Free energy, Enthalpy, Physicochemical property

### INTRODUCTION

Computational methods have been widely used to analyze the structural and physicochemical properties of different organic molecules. The structural stability of molecular entities can be studied using density-functional theory (DFT) or second-order Møller-Plesset perturbation theory (MP2) calculations [1-2]. Moreover, some theoretical approaches, including Monte Carlo simulations, DFT calculations, time-resolved photoacoustic calorimetry, and isoperibol reaction-solution calorimetry has been used to investigate the effects of physicochemical properties, such as solvent effect and solvation enthalpies, on the energetics

of O-H bond in organic solvents [3-4]. Recent studies have revealed that the interaction between methanol and water can be considered as a prototype of a solute with both hydrophilic (hydroxyl) and hydrophobic (methyl) groups. In this interaction, the mobility of some water molecules, rather than those closest to the methane molecules, is slightly reduced due to the presence of hydrophobic solute [5].

The dissociation process of the dimeric structure and geometry parameters of some carboxylic acids and the interaction energies of acid dimers were extensively investigated [6-7]. The molecular interaction energy consists of three components, namely, electrostatic, dispersive, and repulsive components, which were used to study the origin of interactions in acetic acid clusters [8-9]. An unprotected NH<sub>2</sub>-group tends to bind with water

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molecules to form a strong  $O_w-H\dots N$  and a much weaker  $N-H\dots O_w$  H-bonds. Such H-bonding cooperatively stabilizes a cyclic structure formed by the presence of a third water molecule. Conversely, the unprotected  $NH_3^+$  group forms three strong and anti-cooperative  $N^+-H\dots O_w$  H-bonds with  $H_2O$  molecules [10-11].

The only difference in the chemical structure of these amino acids are found in the length of the carbon chain that connects the amine and carboxyl terminals. This makes a significant differences in the strength of intramolecular  $C=N-H$  hydrogen bond in unsolvated ions, which is particularly useful for the competition between solvation and hydrogen bonding interaction. While the solvation of the  $-NH_3^+$  group of amino acids weakens the  $C=O\dots H-N$  hydrogen bond, the solvation of the carboxyl  $-OH$  group results in the strengthening of such bonds. It is also observed that the addition of an  $H_2O$  molecule to the second solvation shell can directly influence the strength of the  $C=O\dots H-N$  hydrogen bonding interaction [12-13]. It is well known that the *cis*-conformation is energetically favorable over the *trans*-conformation in the gas phase. These two conformations become energetically competitive in solvents such as water, alcohol-water, and dimethyl sulfoxide (DMSO)-water [14]. The structures and spectroscopic properties of such organic molecules have been studied by applying principle-based electronic structure calculations [15-21].

The computation of pKa values determines the proton solvation energies in water, which need an estimate of proton solvation energy to match the computed and measured pKa values. Kalidas *et al.* used a new quantum mechanical approximation to obtain free energies of transfer of the proton between water and some organic solvents; these energies of transfer, together with the absolute proton solvation energy in water, yielded proton solvation energies for the solvents [22-27]. In some studies, the proton solvation energy in methanol (MeOH), acetonitrile (MeCN), water, and DMSO was obtained from computational models and pKa values. Proton solvation energies in different solvents were also determined by comparing calculated pKa values using a combination of electrostatic energy and quantum chemical computations. The pKa values of MeCN, DMSO, and MeOH in water correlated well with the corresponding proton solvation energies in these liquids,

indicating that the solvated proton was attached to a single solvent molecule. One of the most challenging tasks in physical chemistry is to determine the absolute free energies of proton solvation [ $DG_{\text{sol}}(H1)$ ] in different solvents. Particularly, the solvation of protons in water is difficult since they are solvated dynamically by different water clusters [28-33]. Thermodynamic perturbation theory was used to evaluate the differences between the solvation Gibbs energies of phenol and phenoxy radicals in acetonitrile and benzene [34-37].

In this work, the relative changes in geometry and physicochemical properties, such as enthalpy change, Gibbs free energy change, of some organic molecules, including aldehyde, ketone, amine, amide, and carboxylic acid-water systems, were analyzed. In addition, the effect of protonation on the organic solvent (OS) was explored in the aqueous phase using a quantum mechanical approach. Such OS-water analyses are very essential and useful to study the physicochemical properties of crosslinked chitosan films, amphiphilic drug-additive systems, and amino acid-based deep eutectic solvents [38-40].

## METHOD

All organic solvent (OS)-water models were constructed using GaussView 5.0. The interaction energies between the selected organic solvents and water molecules were studied by quantum mechanical and computational methods. Several OS-water systems were constructed in such a way to have different orientations with respect to each other and optimized in the gas phase at room temperature. The structures of OS-water systems were optimized by the MP2 level of theory and a double-zeta basis set 6-311++G(d,p) at 298 K and 1 atmospheric pressure. After the optimization of OS-water molecular systems, some changes were observed in the intermolecular hydrogen bond distances and bond angles in the gas phase. All optimized parameters and frequencies of OS-water systems were computed by the MP2 level of theory using the Gaussian09 software package with a 6-311++G(d,p) basis set [41].

Different energies of OS-water systems were carried out using the following equation:

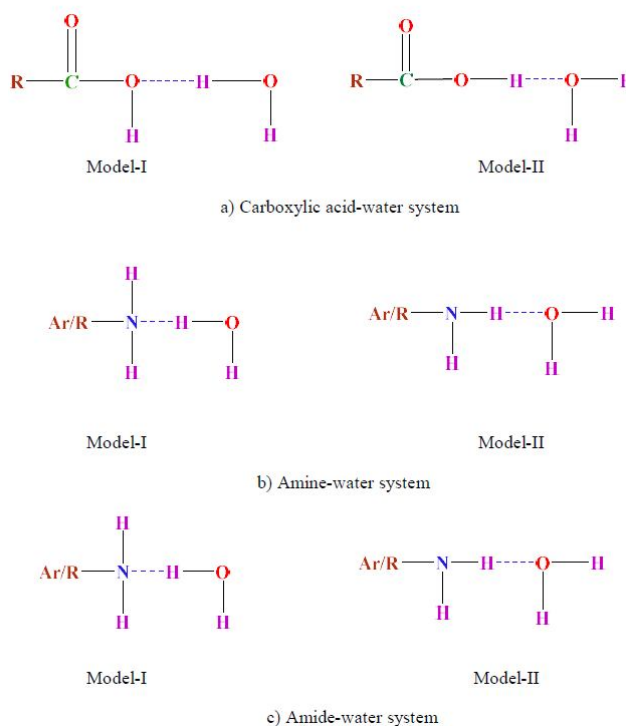
$$E_{IE} = E_{\text{OS-water}} - E_{\text{OS}} - E_{\text{water}}$$

In the above equation,  $E_{IE}$  = Interaction energy,  $E_{OS-water}$  = Energy for OS-water system,  $E_{OS}$  = Energy of OS, and  $E_{water}$  = Energy of water.

The zero-point energy of a molecule or molecular system represents the lowest possible ground state energy of a quantum mechanical physical system. All quantum mechanical systems undergo fluctuations even in their ground state and have their associated zero-point energy due to their wave-like nature. The uncertainty principle requires every physical system to have a zero-point energy greater than the minimum of its classical potential well; accordingly, motion is possible even at absolute zero. Therefore, the lowest-energy state (the ground state) of a system must have a distributed position and momentum that satisfies the uncertainty principle, which implies that the energy of a system must be greater than the minimum of the potential well. The zero-point energy of OS-water systems were computed by the MP2 level of theory with a 6-311++G(d,p) basis set. Also, the Gibbs free energy and enthalpy changes of different OS-water systems were calculated to investigate their physicochemical properties [42-43].

## RESULTS AND DISCUSSION

In this current investigation, we focused on some organic solvents with functional groups such as  $>C=O$ ,  $-COOH$ ,  $-CONH_2$ ,  $-NH_2$ , which generally form different types of interactions with  $H_2O$  molecule to form solvent-water systems [44-47]. Generally, there are two possible organic solvent-water modes of interaction; the first mode is formed between the O or N atom of  $-COOH$ ,  $-CONH_2$ , and  $-NH_2$  groups of solvent molecules with the H atom of water molecules whereas the other mode of interaction occur between the H atom of  $-COOH$ ,  $-CONH_2$ , and  $-NH_2$  groups of solvent molecules with the O atom of water molecules (Fig. 1). Herein, the first and second modes of interaction for the OS-water systems are referred to as the proton acceptor mode (model-I) and the proton donor mode (model-II), respectively. However, organic solvents with a  $>C=O$  functional group, such as acetaldehyde, acetone, and acetyl chloride, exhibits only one type of hydrogen-bonding interaction with the  $H_2O$  molecule, as shown in Fig. 2. The optimized models of the selected OS-water systems are

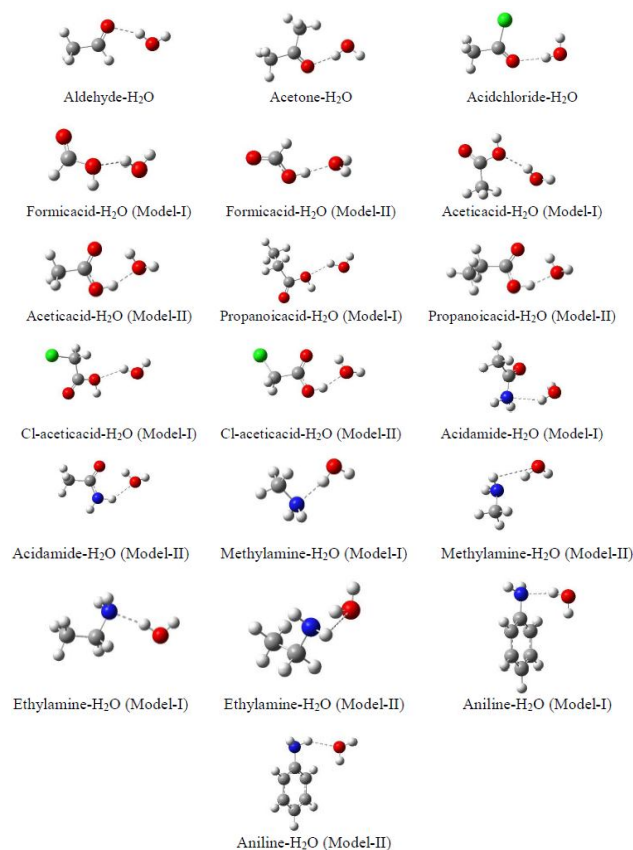


**Fig. 1.** Proton acceptor (Model-I) and proton donor (Model-II) modes of interaction for organic solvent-water systems.

depicted in Fig. 2, and a detailed analysis of structural changes which occurred in OS-water systems and their physicochemical properties are shown in Fig. 1.

### Analysis of Structural Changes in OS-Water Systems

It is well established that the bond length, or bond angle, of a molecule is constant and might change its structure when that molecule interacts with a new external entity. Therefore, a characteristic change was observed in the bond distance or bond angle of a functional group of the solvent molecule during its interaction with  $H_2O$  or  $H^+$  ions, which, in turn, led to the formation of a new molecular system. Computational results confirmed that some changes took place in the bond distance and bond angle of a solvent molecule after its interaction with water molecules and the  $H^+$  ion. The changes in the bond length and bond angle of OS or OS-water systems at the interaction sites can be



**Fig. 2.** Optimized models of different organic molecules- $H_2O$  mixtures.

computed from the optimized geometry. A comparative study of the bond length or bond angle of OS-water systems are shown in Table 1.

Such a comparative study also provides some information about the stability and structural changes taking place in the OS-water systems. For example, the bond length of  $>C=O$  bond of the optimized aldehyde was found to be 1.215 Å whereas that of the aldehyde-water system slightly changed to 1.220 Å. On the other hand, the bond angle of  $H-C=O$  in the optimized aldehyde was found to be  $120.18^\circ$  whereas the bond angle of the optimized aldehyde-water system decreased to  $119.65^\circ$  (Table 2).

Such minor increases or decreases in the bond length or bond angles are considered normal in molecules and could have happened due to the hydrogen-bonding interaction in OS-water systems, as shown in Fig. 1 and 2. Two modes of

**Table 1.** Computed Bond Distances for Different Interaction Sites of OS and OS-Water Systems

Name	Interacting sites	Bond distances (Å)		
		Normal (OS)	OS-water system	
			Model-I	Model-II
Acetaldehyde	C=O	1.215	1.220	-
	C-H	1.110	1.106	-
Acetone	C=O	1.219	1.224	-
	C-C	1.515	1.512	-
Acetyl chloride	C=O	1.192	1.195	-
	C-Cl	1.795	1.788	-
Acetic acid	C=O	1.210	1.207	1.220
	C-OH	1.359	1.373	1.341
Formic acid	C=O	1.198	1.197	1.203
	C-OH	1.355	1.360	1.343
Propanoic acid	C=O	1.211	1.209	1.221
	C-OH	1.360	1.365	1.342
Chloroacetic acid	C=O	1.203	1.180	1.213
	C-OH	1.357	1.360	1.339
Methenamine	C-N	1.465	1.472	1.475
	N-H	1.014	1.015	1.016
Ethanamine	C-N	1.467	1.472	1.473
	N-H	1.016	1.017	1.015
Aniline	N-H	1.011	1.014	1.015
	N-Ar	1.406	1.416	1.402
Ethanamide	N-H	1.009	1.012	1.015
	CO-N	1.377	1.383	1.362

interaction, *i.e.*, proton acceptor and proton donor, were observed between carboxylic acid, amine, and amide molecules and water molecules. Hence, a change could be observed in the bond distance and bond angle of the interaction sites of these molecules. The bond lengths of the interaction sites of C-OH and C=O groups of different optimized carboxylic acids and carboxylic acid-water systems were found to be in the range of  $\sim 1.3$  Å- $1.2$  Å and  $\sim 1.2$ - $1.3$  Å, respectively (Table 1). Again, a variety of changes were observed in the bond angle of the interaction sites of carboxylic acids, *i.e.*,  $O=C-O/H$  and  $C-O-H$  for the

**Table 2.** Computed Bond Angles of Different Interaction Sites of OS and OS-Water Systems

Name	Interacting sites	Bond angle (°)		
		Normal (OS)	OS-water complex	
			Model-I	Model-II
Acetaldehyde	H-C=O	120.18	119.65	-
	O=C-C	121.97	121.17	-
Acetone	O=C-CH <sub>3</sub>	121.97	122.29	-
	O=C-Cl	120.61	120.39	-
Acetic acid	O=C-O	122.67	121.62	123.74
	C-O-H	105.88	105.62	107.60
Formic acid	O=C-H	124.05	124.64	123.47
	C-O-H	113.33	109.13	108.02
Propanoic acid	O=C-O	122.54	122.33	123.62
	C-O-H	105.80	106.61	107.65
Chloroacetic acid	O=C-O	124.17	123.64	125.10
	C-O-H	106.19	106.52	107.40
Methenamine	H-N-H	106.66	106.12	105.64
Ethanamine	H-N-H	106.70	105.99	106.14
Aniline	H-N-Ar	113.52	112.77	112.13
Ethanamide	H-N-H	116.91	114.62	118.94

optimized model (Table 2). Model-II, compared to model-I, resulted in a more stable structure for all carboxylic acid-water systems. This finding shows a characteristic change in the O=C-O/H and C-O-H bond angle of carboxylic acid-water systems in model-II, which is clearly reflected in Table 2. The bond angle for the optimized models of the carboxylic acid O=C-O/H site ranged from ~122 to 124°; the same bond angle in the carboxylic acid-water in model-II showed a dramatic change, ranging from ~123 to 125°. Similar changes were observed in the bond angle of the optimized models of carboxylic acids C-O-H site, which ranged from ~105 to 113°; the same bond angle in carboxylic acid-water in model-II reduced and ranged from ~107 to 108°. However, in amine-water systems, the proton acceptor mode interaction, *i.e.*, model-I, led to a more stable conformation. Interestingly, we did not observe any characteristic changes in the bond lengths (N-H) and bond

**Table 3.** Computed IR Stretching Frequencies for Different Interaction Sites of OS and OS-Water Systems

Name	Interacting sites	IR stretching frequencies (cm <sup>-1</sup> )		
		Normal (OS)	OS-H <sub>2</sub> O system	
			Model-I	Model-II
Acetaldehyde	C=O	1777.72	1766.14	-
Acetone	C=O	1770.59	1761.62	-
Acetyl chloride	C=O	1840.63	1828.39	-
Acetic acid	C=O	1819.52	1831.99	1786.42
Formic acid	C=O	1844.25	1810.57	1829.01
Propanoic acid	C=O	1811.94	1820.48	1779.35
Chloroacetic acid	C=O	1841.98	1971.59	1810.09
Methenamine	N-H	3554.07	3541.57	3541.58
Ethanamine	N-H	3534.46	3520.51	3534.85
Aniline	N-H	3581.81	3559.20	3556.06
	N-H	3618.09	3589.37	3544.11
Ethanamide	C=O	1781.00	1867.30	1755.27

angles (H-N-H) of the interaction sites of the amine (-NH<sub>2</sub>) group. However, a significant change was observed in the interaction sites, *i.e.*, N-H bond lengths and H-N-H bond angles, of the amide-water system due to its mild acidic nature (Table 2).

The above finding was further examined by computing the IR stretching frequency of the functional groups of OS and OS-water systems. The above comparative analysis also confirmed the structural changes that had taken place in the OS-water systems. The IR stretching frequencies of the functional groups are shown in Table 3.

The computed IR stretching frequencies of functional groups may not always be close to the experimental values, but they are still quite reliable because they allow us to investigate the changes occurring in the systems.

### Analysis of Standard Enthalpy Change ( $\Delta H^\circ$ ) in OS-Water Systems

The interaction between the organic solvents and water

molecules cause many physicochemical changes in the OS-water systems. One of the major physicochemical properties is called standard enthalpy change ( $\Delta H^\circ$ ). In this study, the standard enthalpy change of all OS-water systems showed negative values, indicating that such interactions were basically exothermic in nature. Although the exothermic behavior of such systems have already been experimentally established, the same exothermic behavior was observed in the present study with the help of quantum mechanical calculations. The overall analysis of the standard enthalpy change for different carboxylic acid-water models revealed that the proton donor mode (model-II) resulted in more negative enthalpy change values than not needed the proton donor mode (model-I). Also, the acidic nature of carboxylic acid-water systems give such systems a more exothermic nature. The computed standard enthalpy change ( $\Delta H^\circ$ ) values of proton donor mode (model-II) for formic acid ( $-7.71 \text{ kcal mol}^{-1}$ ), acetic acid ( $-8.58 \text{ kcal mol}^{-1}$ ), propanoic

acid ( $-8.61 \text{ kcal mol}^{-1}$ ), and chloroacetic acid ( $-9.25 \text{ kcal mol}^{-1}$ ) reflected negative interaction energy and their purely exothermic nature. Therefore, the enthalpy change ( $\Delta H^\circ$ ) analysis for model-II predictably followed an almost similar trend to that of model-I, as shown in Table 4.

Formic acid < Acetic acid ~ Propanoic acid < Chloroacetic acid

Again, organic solvents such as amine, which showed basic character, underwent a similar kind of standard enthalpy change and exhibited similar exothermic behavior observed in OS-water systems. Model-I resulted in more negative enthalpy change values; that is, the proton acceptor model (model-I) exhibited more exothermic behavior than not needed the proton donor model (model-II), as shown in Table 4. For example, aliphatic amine-water systems showed more negative enthalpy changes than aromatic

**Table 4.** Computed Free Energy, Enthalpy, and Zero-Point Energy Changes for OS-Water Systems in the Gas Phase at 298K and 1 atm

Complexes	Models	Free energy change ( $\Delta G^\circ$ ) (kcal mol <sup>-1</sup> )	Enthalpy change ( $\Delta H^\circ$ ) (kcal mol <sup>-1</sup> )	Zero-point energy ( $\Delta ZPE$ ) (kcal mol <sup>-1</sup> )
Acetaldehyde-water	-	3.07	-4.19	-4.01
Acetone-water	-	2.59	-5.01	-4.82
Acetyl chloride-water	-	2.78	-2.57	-2.81
Ethanamide-water	Model-I	6.01	-2.82	-2.11
	Model-II	0.69	-8.19	-7.65
Methenamine-water	Model-I	0.35	-6.30	-5.99
	Model-II	2.08	-5.63	-5.96
Ethanamine-water	Model-I	0.67	-6.88	-6.11
	Model-II	1.56	-6.27	-5.15
Aniline-water	Model-I	1.29	-5.52	-5.38
	Model-II	4.21	-5.17	-4.36
Formic acid-water	Model-I	2.58	-6.00	-5.32
	Model-II	0.30	-7.71	-7.52
Acetic acid-water	Model-I	6.77	-3.29	-3.13
	Model-II	2.68	-8.58	-7.98
Cl-Acetic acid-water	Model-I	7.46	-2.34	-1.43
	Model-II	0.04	-9.25	-8.70
Propanoic acid-water	Model-I	4.76	-2.77	-2.41
	Model-II	0.48	-8.61	-8.04

amine-water systems. Furthermore, methenamine and ethanamine-water systems resulted in more negative enthalpy values (-6.30 and -6.88 kcal mol<sup>-1</sup>, respectively) than not needed the aromatic amine(aniline)-water system (-5.52 kcal mol<sup>-1</sup>). Also, the enthalpy changes ( $\Delta H^\circ$ ) for such systems were analyzed at the interaction site (-NH<sub>2</sub>-H<sub>2</sub>O), and the trend observed for such systems are shown below:

Aniline < Methenamine < Ethanamine

Contrary to that, in the ethanamide-water system, the proton donor mode (model-II), compared to the proton acceptor mode (model-I), resulted in a much higher negative enthalpy value (-8.19 kcal mol<sup>-1</sup> vs. -2.82 kcal mol<sup>-1</sup>, respectively) and more exothermic behavior. This could be due to the mild acidic nature of the -CONH<sub>2</sub> functional group (Table 1).

Moreover, organic solvents such as acetyl chloride, acetaldehyde, and acetone showed only one type of hydrogen-bonding interaction with water molecules. The standard enthalpy changes ( $\Delta H^\circ$ ) of the above systems were computed, and it was found that these systems also had exothermic behavior in the OS-water interaction, as evidenced in their negative enthalpy values. The acetone-water system had a high interaction energy value, resulting in more exothermic behavior with an enthalpy value of -5.01 kcal mol<sup>-1</sup>. The acid chloride-water system had a low negative enthalpy value (-2.57 kcal mol<sup>-1</sup>) with less exothermic behavior compared to other systems. The computed standard enthalpy change ( $\Delta H^\circ$ ) trend for these systems in the gas phase are shown below:

Acid chloride < Acetaldehyde < Acetone

### **Analysis of Standard Gibbs Free Energy Change ( $\Delta G^\circ$ ) in OS-Water Systems**

The analysis of the standard Gibbs free energy change ( $\Delta G^\circ$ ) of such OS-water systems revealed the feasibility of molecular interaction and stability of such systems. Generally, while higher negative  $\Delta G^\circ$  values reveal that the OS-water interaction is more feasible at room temperature, lower negative or positive  $\Delta G^\circ$  values of OS-water systems result in less stable or unfavored molecular interaction. In this study, the computed free energy for most OS-water

systems were found to be positive (Table 4). This could be due to the fact that all the experiments were carried out in only the gas phase. Interestingly, however, the stability and feasibility trends for all molecular interactions in OS-water systems were quite similar to those of such systems.

### **Analysis of Zero-Point Energy Change ( $\Delta ZPE$ ) in OS-Water Systems**

Zero-point energy change ( $\Delta ZPE$ ) analysis of OS-water systems is also an important tool to understand their molecular vibration.  $\Delta ZPE$  values of OS-water mixtures were calculated using an empirical approach dependent on molecular interaction. The computed  $\Delta ZPE$  value showed almost similar results to those  $\Delta H^\circ$  values obtained previously. Accordingly, the more stable and favored OS-water systems exhibited higher negative or less positive ZPE values (Table 4).

In the overall analysis of ZPE change for different carboxylic acid-water systems, the proton donor mode (model-II) resulted in higher negative enthalpy change values than not required the proton donor mode (model-I). Such an analysis was also important in assessing the stability of carboxylic acid-water systems. The more acidic behavior of carboxylic acid-water systems resulted in higher negative ZPE values. The computed ZPE values of the proton donor mode (model-II) for formic acid (-7.52 kcal mol<sup>-1</sup>), acetic acid (-7.98 kcal mol<sup>-1</sup>), propanoic acid (-8.04 kcal mol<sup>-1</sup>), and chloroacetic acid (-8.70 kcal mol<sup>-1</sup>) were found to be negative, which reveals the formation of stable zero-point vibrational energy for such systems. Thus, the ZPE analysis of the proton donor mode followed an almost similar trend to that observed in the enthalpy analysis, as represented below:

Formic acid < Acetic acid < Propanoic acid < Chloroacetic acid

Amines also were found to follow similar ZPE changes in OS-water systems. The proton acceptor mode (model-I) resulted in more negative ZPE change values than not needed the proton donor mode (model-II). For example, aliphatic amine-water systems had higher negative ZPE values than the aromatic amine-water systems. Methenamine- and ethanamine-water systems exhibited

higher negative ZPE values (-5.99 kcal mol<sup>-1</sup> and -6.11 kcal mol<sup>-1</sup>, respectively) than the aromatic amine(aniline)-water system (-5.38 kcal mol<sup>-1</sup>). The ZPE analysis was performed for the ground-state molecular vibrational modes of the interaction site of -NH<sub>2</sub>-H<sub>2</sub>O, and the trend of such zero-point vibrations are shown below:



However, for the mild acidic ethanamide-water system, the proton donor mode (model-II) resulted in a much higher negative ZPE value (-7.65 kcal mol<sup>-1</sup>) compared to that obtained *via* the proton acceptor mode (model-I; 2.11 kcal mol<sup>-1</sup>) (Table 4). The ethanamide-water system exhibited the most stable zero-point vibration than other aliphatic or aromatic amines.

As acetyl chloride, acetaldehyde, and acetone showed only one type of hydrogen-bonding interaction with water molecules, the zero-point molecular vibration for these systems were quite stable. The ZPE changes in these systems were also computed, and it was found that they also had stable OS-water interaction due to their negative ZPE values. The acetone-water system was found to have a higher ZPE value (-4.82 kcal mol<sup>-1</sup>) than the aldehyde-water system (-4.01 kcal mol<sup>-1</sup>). The acid chloride-water system exhibited a lower negative ZPE value (-2.81 kcal mol<sup>-1</sup>) and had a less stable zero-point vibration than those of other systems. The computed ZPE trend for the above complexes in the gas phase is shown below:



### The Effect of Protonation on OS

Protonation of organic solvents in the aqueous phase plays a very important role in changing their physicochemical properties. In this research, quantum mechanical calculations were used to investigate the acidic and basic nature and the hydrogen-bonding interaction of some organic solvents. The theoretical analysis of the protonated OS-H<sub>3</sub>O<sup>+</sup> systems in the aqueous phase were compared with their respective interaction energies to analyze the strength of hydrogen-bonding interaction and the stability and acid-base trends (Table 5). In the aqueous phase, since the hydrogen bond formation was very strong

**Table 5.** Computed Interaction Energies (kcal mol<sup>-1</sup>) of OS-H<sub>3</sub>O<sup>+</sup> Systems in the Aqueous Phase at 298 K and 1 atm

OS-H <sub>3</sub> O <sup>+</sup> systems	Interaction energies (kcal mol <sup>-1</sup> )	
	Model-I	Model-II
Acetaldehyde-H <sub>3</sub> O <sup>+</sup>	-23.20	
Acetone-H <sub>3</sub> O <sup>+</sup>	-26.10	
Acetyl chloride-H <sub>3</sub> O <sup>+</sup>	-13.18	
Formic acid-H <sub>3</sub> O <sup>+</sup>	-2.38	-12.31
Acetic acid-H <sub>3</sub> O <sup>+</sup>	-12.81	-22.83
Cl-Acetic acid-H <sub>3</sub> O <sup>+</sup>	-10.58	-23.25
Propanoic acid-H <sub>3</sub> O <sup>+</sup>	-12.60	-16.07
Ethanamide-H <sub>3</sub> O <sup>+</sup>	-21.08	-24.35
Methenamine-H <sub>3</sub> O <sup>+</sup>	-20.81	-2.48
Ethanamine-H <sub>3</sub> O <sup>+</sup>	-20.97	-2.48
Aniline-H <sub>3</sub> O <sup>+</sup>	-18.41	-2.40

in OS-H<sub>3</sub>O<sup>+</sup> systems, the computed interaction energy, which was dependent on the strength of intermolecular hydrogen bonding between the two entities, for these systems resulted in very high negative values. The high negative interaction energy values for hydrogen bond formation in OS-H<sub>3</sub>O<sup>+</sup> is reflected in Table 5.

The protonation of acetone resulted in the formation of a hydrogen bond as >C=O---H<sub>3</sub>O<sup>+</sup>. As the two -CH<sub>3</sub> groups of acetone molecule also enhanced the electron charge density on the O atom of the >C=O group, the hydrogen bond formation between acetone and H<sub>3</sub>O<sup>+</sup> was quite strong, which is clearly reflected in its interaction energy value (-26.10 kcal mol<sup>-1</sup>). However, in the acetaldehyde-H<sub>3</sub>O<sup>+</sup> system, only one single methyl (-CH<sub>3</sub>) group of aldehyde molecule donated its electron charge density to the O atom of the >C=O group, which resulted in a lower interaction value in the aldehyde-H<sub>3</sub>O<sup>+</sup> system (-23.20 kcal mol<sup>-1</sup>) compared to the acetone-H<sub>3</sub>O<sup>+</sup> system. Interestingly, in the acid chloride-H<sub>3</sub>O<sup>+</sup> system, the electron-donating -CH<sub>3</sub> group on the >C=O functional group increased the electron charge density of the acid chloride molecule whereas the electronegative -Cl group reduced the electron charge density on the >C=O functional group, as a result of which the acid chloride-H<sub>3</sub>O<sup>+</sup> system, compared to the protonated



aldehyde and ketone systems, exhibited the lowest negative interaction energy value ( $-13.18 \text{ kcal mol}^{-1}$ ). The computed interaction energies for these protonated molecules in the aqueous phase showed the following trend:

Acid chloride < Acetaldehyde < Acetone

The protonation of carboxylic acid led to the formation of carboxylic acid- $\text{H}_3\text{O}^+$  systems. Moreover, it was found that the interaction energy for model-II (proton donor), compared to model-I, for all carboxylic acid- $\text{H}_3\text{O}^+$  systems resulted in higher negative interaction energy values ( $-12.31$  to  $-23.25 \text{ kcal mol}^{-1}$  in Model-II compared to  $-2.38$  to  $-12.81 \text{ kcal mol}^{-1}$  in model-I). This could be due to the fact the proton donor mode is theoretically more favorable than the proton acceptor mode (Table 5). This property of carboxylic acid- $\text{H}_3\text{O}^+$  systems in the proton donor mode (model-II) clearly represents the acidic nature of carboxylic acids. The computed interaction energy trend of proton donor mode (model-II) for different carboxylic acid- $\text{H}_3\text{O}^+$  systems in the aqueous phase followed the following trend:

Formic acid < Propanoic acid < Acetic acid < Chloroacetic acid

Among all the carboxylic acid- $\text{H}_3\text{O}^+$  systems, the formic acid- $\text{H}_3\text{O}^+$  system exhibited a lower negative interaction energy value ( $-12.31 \text{ kcal mol}^{-1}$ ). This could be due to its less stable conformation compared to other carboxylic acids. The formic acid- $\text{H}_3\text{O}^+$  system became unstable as formic acid was highly acidic in nature. On the other hand, the highest negative interaction energy value ( $-23.25 \text{ kcal mol}^{-1}$ ) was found in the chloroacetic acid- $\text{H}_3\text{O}^+$  system. This could be due to the presence of the highly electronegative -Cl group, which enhanced the interaction energy of this pair and the formation of a strong hydrogen bond. Interestingly, the interaction energy value for the protonated acetic acid and propanoic acid fell within the range of  $\sim -16$ - $22 \text{ kcal mol}^{-1}$ . It could be due to the electron-donating effect caused by the alkyl groups (*i.e.*,  $-\text{CH}_3$  &  $-\text{C}_2\text{H}_5$ ).

However, amine with the functional group  $-\text{NH}_2$  had different computed interaction energy compared to that of amine- $\text{H}_3\text{O}^+$  systems. This difference can be attributed to

the basic nature of amine. The negative interaction energy values of the protonated methenamine, ethanamine, and aniline were much higher in the proton acceptor mode (model-I;  $-18$ - $20 \text{ kcal mol}^{-1}$ ) than in the proton donor mode (model-II;  $-2$ - $4 \text{ kcal mol}^{-1}$ ). The computed interaction energy trend of different amine- $\text{H}_3\text{O}^+$  systems in the proton acceptor mode (model-I) in the aqueous phase is shown below:

Aniline < Methenamine  $\sim$  Ethanamine

Interestingly, in the protonated ethanamide, the proton donor mode (model-II) resulted in a higher negative interaction energy value ( $-24.35 \text{ kcal mol}^{-1}$ ) than the proton acceptor mode (model-I;  $-21.08 \text{ kcal mol}^{-1}$ ), which shows the mild acidic nature of the  $-\text{CONH}_2$  functional group (Table 1).

## CONCLUSIONS

The above results clearly show that the optimized models of OS-water systems had a strong hydrogen bond formation. In addition, the results confirmed that the molecular interaction between the solvents and  $\text{H}_2\text{O}$  or  $\text{H}_3\text{O}^+$  ions caused significant changes in bond distances and bond angles. The analyses also provided some information about the stability and structural changes in the OS-water systems. An examination of standard enthalpy changes ( $\Delta H^\circ$ ) for all OS-water systems always resulted in negative values, showing that the interactions were basically exothermic in nature. Interestingly, the stability and feasibility trends of all OS-water systems were quite similar, as reflected in the analysis of Gibbs free energy change ( $\Delta G^\circ$ ). The calculation of zero-point energy change ( $\Delta ZPE$ ) partially supported the analysis of  $\Delta H^\circ$  and  $\Delta G^\circ$  values. The very strong hydrogen bond formation in protonated organic solvents to form stable OS- $\text{H}_3\text{O}^+$  systems in the aqueous phase resulted in very high negative interaction energy values. The stability and physicochemical properties of the OS- $\text{H}_3\text{O}^+$  system showed an almost similar trend to that observed in OS- $\text{H}_2\text{O}$  systems. For all carboxylic acid- $\text{H}_3\text{O}^+$  systems, the proton donor mode (model II) led to a more stable structure than not needed the proton acceptor mode (model I). However, in amine- $\text{H}_3\text{O}^+$  systems, the proton

acceptor mode (model-I) was much more favored than the proton donor mode (model-II). Exploring the formation of OS-water systems can be very useful in studying the protonation and deprotonation mechanisms, transition state, electron transfer, and kinetics analysis.

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