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Thermodynamic Parameters of Electrochemical Oxidation of L-DOPA: Experimental and Theoretical Studies

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Thermodynamic Parameters of Electrochemical Oxidation of L-DOPA:

Experimental and Theoretical Studies

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Abstract:

Electrode potential and thermodynamic parameters of the electrochemical reaction of L-DOPA in aqueous solution are obtained experimentally by cyclic voltammetry method and also calculated theoretically using accurate ab initio calculations (G3MP2//B3LYP) along with the available solvation model of CPCM. The pH dependence of the redox activity of L-DOPA in aqueous solution at temperatures in the range of 10-30°C was used for the experimentally determination of the standard reduction potential, changes of entropy, enthalpy and Gibbs free energy for the studied reaction. The experimental formal redox potential of two-proton-two-electron reduction process was obtained to be 0.745 V versus standard hydrogen electrode (SHE). The theoretical and experimental values (0.728 V and 0.745 V) for the standard electrode potential of L-DOPA are in agreement with each other. The difference between the peak potential of the L-DOPA and the products, which are produced by chemical reactions, has been measured experimentally and also calculated theoretically. There is also an agreement between experimental and theoretical potential difference. Also in this work the changes of thermodynamic functions of solvation are calculated from the differences between the solution-phase experimental values and the gas-phase theoretical values.

Keyword: L-DOPA; Thermodynamic parameters; Standard redox potentials; G3MP2//B3LYP; Ab initio calculations

1- Introduction

The catecholamines act as neuromodulators in the central nervous system and as hormones in the blood circulation. Catecholamine drugs are also used to treat hypertension, bronchial asthma, organic heart disease and are used in cardiac surgery and myocardial infarction.[1-7] L-3,4-dihydroxyphenylalanine (L-DOPA) has known as an important catecholamine and is a chemical that is made and used as part of the normal biology of some animals and plants. L-DOPA is considered to be the most effective medicine drug available for Parkinson's disease. [8-11] Electrochemical behavior of L-DOPA plays important roles in its physiological functions, and is a key factor in diagnosis of some diseases in clinical medicine.[12-14] Hence, it is desirable to develop an electrochemical method to study electron-transfer processes for this compound. Furthermore, the measurement of the electrode potential of L-DOPA, which could be carried out by cyclic voltammetry method, will be beneficial toward understanding and interpreting the mechanism of electron-transfer of this compound.

There are several spectrophotometric[15-18] and electrochemical methods for the determination of L-DOPA in literature.[8, 14, 19-21] Despite of important biological properties of L-DOPA, there are a few studies on the electrochemical behavior of L-DOPA in aqueous solution.[22-24] In the present work, which builds on our previous studies of other catechol derivatives,[25-28] we have studied the effect of temperature and pH on electrochemical redox reaction of L-DOPA in aqueous solution experimentally. The redox potential of L-DOPA and changes of thermodynamic functions have been also studied theoretically and the results have been compared with the experimental data. A key finding of this work is to further reveal the electrooxidation mechanism of L-DOPA by employing accurate theoretical methods. To identify

the electrooxidation mechanism of a solute at the aqueous solution, the produced chemical compounds, are usually identified by classical methods and instruments. But, in the present work, we have used standard ab initio calculations in order to verify the proposed electrochemical mechanism. According to the literature and to the best of our knowledge, this is the first such study on L-DOPA and should also be relevant to other catecholamines derivatives. We have also calculated the changes of solvation energy for the electrochemical reaction of L-DOPA using a method we proposed recently,[28] and have compared the results with those have been calculated using available solution model. The agreement between theory and experiment is excellent and confirms the validity of theoretical models used here and the accuracy of experimental methods.

2. Experimental section

2.1. Equipment and reagents

Electrochemical measurements were carried out using a computer-controlled potentiostat, Autolab/PGSTAT 101 (Eco-Chimie) with the powerful NOVA software. A three-electrode electrochemical cell was employed for all the electrochemical measurements. The working, counter, and reference electrodes were a glassy carbon electrode with a diameter of 2 mm, a graphite electrode, and a saturated calomel reference electrode (SCE) respectively. All the potentials in the text were reported with respect to this reference electrode. All potentials have been measured versus this reference electrode. Room temperature and elevated temperature experiments were performed in a double walled glass cell, of 20 cm^3 volume, which was connected to a heated water bath at the appropriate temperature. pH measurements were

performed using a Metrohm 691 pH/mV meter. In order to obtain a clean renewed electrode surface, the glassy carbon electrode was cleaned by successive polishing with 0.05 μm alumina-water slurry prior to its use in electrochemical experiments. L-DOPA and all other reagents were purchased from Merck and used without further purification. The solutions were prepared with doubly distilled water. Phosphate buffer solutions (0.1 M) were prepared from 0.1 M H_3PO_4-NaH_2PO_4 and pH adjusted with 0.1 M H_3PO_4 or 2.0 M NaOH.

2.2. Thermodynamic analysis

L-DOPA, see Scheme 1A for structure, can be oxidized to its oxidized form by a one-step, two-electron redox reaction including deprotonation in aqueous solution.[28-30] The overall redox reaction of L-DOPA relative to the standard hydrogen electrode (SHE) is:

$$\text{L-DOPAquinone }_{(soln)} + H_{2(g)} \rightleftharpoons \text{L-DOPA }_{(soln)} \tag{1}$$
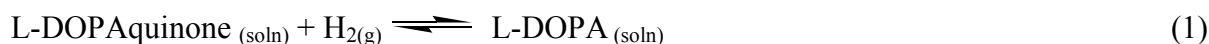
Subtracting the half-reaction of SHE, this two-electron process is

$$\text{L-DOPAquinone }_{(soln)} + 2e^- + 2H^+_{(soln)} \rightleftharpoons \text{L-DOPA }_{(soln)} \tag{2}$$
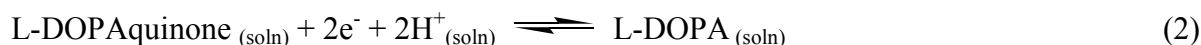
For this reaction, the conditional formal potential, $E°'$, is given by the following equation [31]

$$E°' = E° - (2.303mRT/2F) \text{ pH} \tag{3}$$

where $E°$ is the standard redox potential (or formal potential at pH 0); R, T, and F are the gas constant, temperature, and Faraday constant, respectively. As can be seen from Eq. 3, the conditional formal potential, which is approximated by the midpoint potential between the anodic and cathodic peaks, $E°' = (E_{p,a} + E_{p,c})/2$, of L-DOPA is pH-dependent and it is shifted to negative potential as pH increases. The standard redox potential of L-DOPA , $E°$, can be obtained from the intercept of Eq. 3, at pH 0. Standard redox potentials has been measured at

different temperatures in order to calculate changes of enthalpy, entropy, Gibbs free energy and equilibrium constant of studied reaction based on the following equations.[31, 32]

$$\Delta S° = nF(\partial E°/\partial T)_P \qquad (4)$$

$$\Delta H° = -nF[E° - T(\partial E°/\partial T)_P] \qquad (5)$$

$$\Delta G° = -nFE° \qquad (6)$$

$$LnK = -\Delta G°/RT \qquad (7)$$

where T, F and n are the temperature, Faraday constant (96485 C mol^{-1} or 23.061 kcal mol^{-1} V^{-1}), and number of electrons transferred, respectively. In this study, electrode potentials have been measured versus SCE, then converted to electrode potentials vs SHE using the exact potential values of SCE at each temperature. [33] Using the standard redox potentials and its variation relative to temperature, $\Delta H°$ and $\Delta S°$ of the corresponding to reaction (1) have been obtained.

3. Computational details

G3MP2//B3LYP, one of recent methods in the Gaussian-*n* series, provides improved accuracy for relatively large molecules.[34] At the G3MP2//B3LYP method, first the molecular geometries are optimized and the vibrational frequencies are calculated at the B3LYP level of theory using the 6-31G(d) basis set. The energy of different conformers of each species has been investigated and the most stable conformer has been chosen for further calculations and extra care has been pained in order to find the global minimum for each species. Next, single point calculations are performed at increasing levels of theory and basis sets to obtain the electronic energy so that QCISD(T) with the 6-31G(d) basis set and MP2 with the GTMP2Large basis set are applied to derive the electronic energies, enthalpy and Gibbs free energy of the molecules in

the gas phase [34]. In order to calculate solvation energies, a continuum model of solvation, the conductor-like polarizable continuum model (CPCM) has been used at the B3LYP/6-311+G(d,p) level of theory.[35-37] This level of theory has been recommended for the calculation of solvation energies.[37] The radii and the rest of the parameters have been kept as default values. All geometries of the studied species have been optimized fully in the presence of solvent. Gaussin-03[38] has been used for all ab initio molecular orbital theory and density functional theory calculations.[39,40]

4. Results and discussion

4.1. Influence of pH and temperature on the L-DOPA voltammetric response

The effect of pH on the L-DOPA response was investigated by cyclic voltammetry using a 0.1 M phosphate buffer solution containing 10.0 µM L-DOPA at different temperatures (10-30 °C) and at various pH values ranging from 2.0 to 9.0. The voltammmograms corresponding to temperature of 10-30 °C are presented in Supporting Information (Figs. S1-S5). As it is shown in Fig. S1 of the Supporting Information, the first redox couple (A_1/C_1) is related to oxidation and reduction of L-DOPA and the second redox couple (A_2/C_2 or A_3/C_3) belongs to the oxidation and reduction of the new compounds produced by subsequent chemical reactions.[22,41] The experimental results also indicate that the first redox couple of L-DOPA is well-developed quasireversible peak in acidic pHs (Fig. S1 of suppoting information, voltammograms a and b) while in neutral and basic pHs, it shows an irreversible behavior with no well-defined cathodic peak (Fig. S1 of suppoting information, voltammograms c and d). Therefore, in this study we used conditional formal potential in the pH range of 2.0 − 5.0 for calculating the formal redox

potential of two-proton-two-electron reduction process of L-DOPA. The value of anodic and cathodic peak potential of the first redox couple (A_1/C_1 redox couple) at the pH range of 2.0-5.0 are listed in Table. S1 of Supporting Information. As it is shown in Fig. 1, at all temperatures, $E°'$ is linearly pH dependent and the slope of this plots is in a good agreement with the theoretical value of -0.059 V/pH (the slope of Eq. 3). In order to obtain standard redox potentials, $E°$, of L-DOPA, the formal redox potentials have been extrapolated to pH 0, which corresponds to unit activity of proton in solution, $a^{H+} = 1.0$. The standard redox potential at 25 °C is found to be 0.745 V versus SHE for L-DOPA. Fig. 2 shows the resulting variation of redox potentials with temperature. The slope of E^0 versus T, presented in this figure, has been used to calculate the entropy and enthalpy changes of reaction 1 using Eqs. 4 and 5. The results are listed in Table 1. The changes of entropy and enthalpy for this reaction were found to be -92.24 cal mol^{-1} K^{-1} and -61.87 kcal mol^{-1}. These values are temperature independent in the studied range of temperature.

4.2. Theoretical calculations of electrode potential

The change of standard Gibbs energy of reaction 1 can be computed using the thermodynamic cycle that is shown in Scheme 2. From this cycle, ΔG is computed by the following expression:

$$\Delta G_{(soln)} = \Delta G° (gas) + \Delta G(solv) \tag{8}$$

where $\Delta G°_{gas}$ is standard Gibbs energy of reaction 1 in the gas phase and ΔG_{solv} is the net solvation energy of reaction 1 and are defined as follows:

$$\Delta G°(gas) = G°(\text{L-DOPA, gas}) - G°(\text{L-DOPAquinone, gas}) - G°(H_2, gas) \tag{9}$$

$$\Delta G(solv) = G(\text{L-DOPA, solv}) - G(\text{L-DOPAquinone, solv}) \tag{10}$$

As detailed in section 3, the gas phase Gibbs energies are calculated accurately using ab initio calculations at the G3MP2//B3LYP level of theory and solvation energies are calculated using the CPCM model of solvation. For the standard state correction of solvation energies, the value of 1.9 kcal mol^{-1} (RT Ln 24.2) has been considered for the change in standard state from 1 bar to 1 mol L^{-1}. [26] G3MP2//B3LYP results for the studied species together with CPCM/6-311+G(d,p) solvation energies are presented in Table 2. These calculated values of the thermodynamic parameters associated with the reaction 1 are presented in Table 3. Based on Eq. 6, and using the G3MP2//B3LYP calculated gas-phase energy together with CPCM solvation energies, the redox potential for the two-proton-two-electron reduction, reaction 1, has been calculated as 0.728 V, which shows a small deviation of just 0.017 V in comparison with the experimental value of 0.745 V for the studied reaction. This agreement between theoretical redox potential and the experimental value shows that the approximations have been made in calculation of electrode potential are justified.

The components of Gibbs free energy of solvation, $\Delta H_{(solv)}$ and $\Delta S_{(solv)}$, can be calculated from the differences between the solution-phase experimental values and the gas-phase theoretical values, which both have been obtained accurately.

$$\Delta H_{(solv)} = \Delta H_{(tot, expt)} - \Delta H_{(gas, theory)} \tag{11}$$

$$\Delta S_{(solv)} = \Delta S_{(tot, expt)} - \Delta S_{(gas, theory)} \tag{12}$$

$$\Delta G_{(solv)} = \Delta H_{(solv)} - T\Delta S_{(solv)} \tag{13}$$

$\Delta H_{(solv)}$, $\Delta S_{(solv)}$ and $\Delta G_{(solv)}$ are the solvation enthalpy, entropy and Gibbs free energy of reaction 1. Using the G3MP2//B3LYP results presented in Table 3 for the gaseous thermodynamic quantities, together with experimental value obtained from Fig. 2, $\Delta H_{(solv)}$ and $\Delta S_{(solv)}$ are calculated as -12.2 kcal mol^{-1} and -58.8 cal mol^{-1} K^{-1} for reaction 1, respectively.

Using these values and Eq. 13, we have calculated the changes of Gibbs energy of solvation for reaction 1, $\Delta G_{(solv)}$, and it is found to be 5.4 kcal mol^{-1}. We have also calculated the $\Delta G_{(solv)}$ for reaction 1 using CPCM model of solvation at the B3LYP/6-311+G(d,p) level of theory. The calculated $\Delta G_{(solv)}$ for reaction 1 obtained by CPCM model of solvation was also found to be 6.1 kcal mol^{-1}, in agreement with the value obtained by Eqs. 11-13. The perfect agreement between these two results, indicate that CPCM-B3LYP+6-311G(d,p) has successfully predicted accurate aqueous solvation energies for the catecholamine compounds.

The pH of the isoelectric point (PI) of L-DOPA, as an amino acid, is about 6.0.[42] Therefore, at a low pH, the molecules will be in acidic form. But in moderate acidic and neutral pHs (which is near to PI), most of the L-DOPA molecules are in neutral form and they have both positive and negative charge. In order to understanding of the mechanisms involved in the electrochemical oxidation of L-DOPA on a glassy carbon electrode, the effects of pH, L-DOPA concentration, time window of the chosen electrochemical method and temperature on the second cycle of the voltammetric response of L-DOPA have been studied. [41] Investigation of the electrochemical behavior of L-DOPA in different pHs indicated that the process of electrooxidation of L-DOPA is followed by the coupled chemical reactions. The results indicated that the coupled chemical reactions were dependent on the solution pH. As it is noted, at a low pH, the amine and carboxyl groups will be protonated, and the molecules will be in acidic form. Therefore, the nucleophilic property of the amine group is removed through protonation, and the intramolecular Micheal reaction does not take place in strong acidic pHs. Therefore, the following chemical reaction of L-DOPA oxidation is the 1,4 addition of H_2O on L-DOPAquinone. Since the hydroxyl group is a strong electron-donating group, the product of the hydrolysis reaction is more easily oxidized than the starting material (L-DOPA). Thus, the A_2 and C_2 peaks in strong acidic media (Fig. S1

of Supporting Information, voltammogram a, A_2/C_2 redox couples), may be related to the oxidation and the reduction of the hydrolysis chemical reaction product (see Scheme 1B for the structure). The voltammogram presents that the difference between the redox couple formal potentials of A_1/C_1 and A_2/C_2 (ΔE_{pB}) is 0.26 V. The difference between the peak potentials of redox couples of A_2/C_2 and A_1/C_1, ΔE_{pB}, has been calculated theoretically and compared with the experimental value. The gas phase Gibbs energies of these compounds are calculated accurately using ab initio calculations at the MP2//G3MP2large levels of theory, considering the size of compounds, and the solvation energies are calculated using the CPCM model of solvation. The ΔE_{pB} has been calculated as 0.28 V which is in agreement with the experimental value of 0.26 V. As the pH increases, L-DOPA molecules with deprotonated amine group ($R-NH_2$) increase. Therefore, with increasing the pH to moderate acidic and neutral pHs, the concentration of $R-NH_2$ increases and the intramolecular following chemical reaction can take place. The product of the intramolecular chemical reaction formed via the intramolecular Michael addition reaction of amine group with *o*-benzoquinone ring (see Scheme 1C for the structure) and is more easily oxidized because of presence of an electron donating group in its structure. Considering that the nucleophilic property of the amine group is more than that of the hydroxyl group, it could be concluded that at moderate acidic pHs, the A_3/C_3 redox couple (Fig. S1 of Supporting Information, voltammogram b-d, A_3/C_3 redox couple) corresponds to oxidation/reduction of the intramolecular chemical reaction product. Disappearance of the A_2/C_2 redox couple also indicates that the hydroxyl chemical reaction does not take place at moderate acidic pHs. These results are supported by the other spectrophotometric, voltammetric and simulation methods.[22, 41, 43] The theoretical results show that ΔE_{pC} ($\Delta E_{pC} = E°'(A_1/C_1) - E°'(A_3/C_3)$) is equal to 0.41 V. In a similar way, the difference between the peak potentials of redox couple of A_3/C_3 and A_1/C_1,

ΔE_{pC}, (see Fig. S1 of Supporting Information, voltammogram b) has been calculated as 0.53 V which is in agreement with the experimental value of 0.41 V. The agreement between the theoretical and experimental values of ΔE_{pB} and ΔE_{pC}, approves the kinds of subsequent chemical reactions which occur after the electrooxidation of L-DOPA in different pHs.

The originality of the present work is that we have used accurate ab intio calculations in order to support the kinds of subsequent chemical reactions, which are predicted by the experimental evidences. To identify the products of the following chemical reactions, they are usually separated and characterized by the classical and instrumental methods, e.g. melting point, FT-IR, UV-Vis, and NMR spectroscopy. Although they are capable of giving good results, these methods are expensive, time-consuming and may involve difficult separations. In this study, accurate theoretical methods, which are relatively fast and cheap, are used as an alternative for the experimental methods in order to support the kinds of following chemical reactions.

5- Conclusion

In this work electrochemical behavior of L-DOPA, in an aqueous solution has been studied both experimentally and theoretically. L-DOPA is oxidized to the corresponding quinone by a one-step, two-electron redox reaction including deprotonation. The formal redox potential of L-DOPA, E°′, is shifted to more negative values as the pH is increased. The standard redox potential (i.e., at pH 0) of L-DOPA for the two-electron-two-proton process is 0.745 V. As part of this work, high-level ab initio molecular orbital calculations were used to calculate the standard redox potential of L-DOPA. All theoretical calculations were in agreement with experiment (within 0.017 V), mutually verifying the experimental and theoretical studies.

Changes of solvation enthalpies, entropies and Gibbs free energies for the studied reaction have been determined and the results have been compared with the result of solvation model of CPCM. There is a good agreement between solvation Gibbs energy obtained by the method we proposed here and CPCM solvation model. Investigation of electrochemical behavior of L-DOPA indicates that there are subsequent chemical reactions following the oxidation of L-DOPA. These products are oxidized at less positive potential with respect to L-DOPA.

Supporting Information Available: B3-LYP/6-31G(d) optimized geometries in the form of GAUSSIAN archive entries, and corresponding total energies as well as experimental cyclic voltammetry results. This information is available free of charge via the Internet at http://pubs.acs.org.

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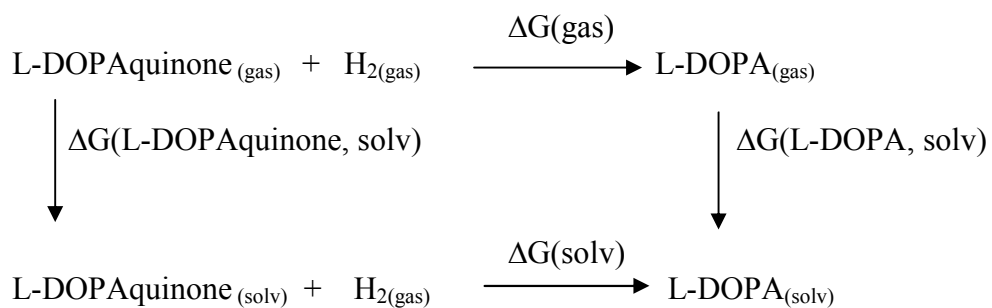46
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References:

(1) Croll, R. P. *J. Comp. Neurol.* **2001**, *441*, 91-105.

(2) Lechin, F.; Van Der Dijs, B.; Lechin, A. *J. Appl. Res. Clin. Experiment.* **2005**, *5*, 605-621.

(3) Kobayashi, K. *J. Invest. Derm. Symp.* **2001** , *6*,115-121.

(4) Martin, P. R.; Weingartner, H.; Gordon, E.K.; Burns, R.S. Linnoila, M.; Kopin, I. J.; Ebert,

M. H. *Ann. Neurol.* **1984** , *15*,184-187.

(5) Vogt, M. *Brit. Med. Bull.* **1973** , *29*,168-171.

(6) Agarwal, R.; Chandna, V.; Engelking, L.; Lightbown, K.; Kumar, M.; *Brain Res. Bull.* **1993**,

32, 285-291.

(7) Viskin, S.; Belhassen, B. *Prog. Cardiovasc. Dis.* **1998** , *41*,17-34.

(8) Shahrokhian, S.; Asadian, E. *J. Electroanal. Chem.* 636 **2009**, *636*, 40-46.

(9) Aviles-Olmos, I.; Martinez-Fernandez, R.; Foltynie, T. *Eur. Neurol. J.* **2010**, *2*,91-100.

(10) Wolf, J.; Bouhaddi, M.; Louisy, F.; Mikehiev, A.; Mourot, L.; Cappelle, S.; Vuillier, F.;

Andre, P.; Rumbach, L. *J. Regnard, Clin. Sci.* **2006**, *110*, 369-377.

(11) Fabbrini,G.; Brotchie, J. M.; Grandas, F.; Nomoto, M.; Goetz, C.G. *Movement disord.* **2007**,

22, 1379-1389.

(12) Wring, S. A.; Hart, J. P. *Analyst* **1992**, *117,* 1215-1229.

(13) Demircigil, B. T.; Dogan-Topal, B.; Ozkan, S. A. *Collect. Czech. Chem. Commun.* **2011**, *76*,

1717-1736.

(14) Yan, X. X.; Pang, D. W.; Lu, Z. X.; Lu, J. Q.; Tong, H. *J. Electroanal. Chem.* **2004**, *569*,

47-52.

(15) Paik, M. J.; Nguyen, D. T.; Yoon, J.; Cho, I. S.; Shim, W.; Kim, K. R.; Cho, K. H.; Choi, S.;

Lee, G. *Bull. Korean Chem. Soc.* **2011**, *32*, 977-980.

(16) Teixeira, M. F. S.; Marcolino-Junior, L.; Fatibello-Filho, O.; Dockal, E.; Bergamini, M. F. *Sensor Actuat. B-Chemical* **2007**, *122*, 549-555.

(17) Zhao, S.; Bai, W.; Wang, B.; He, M. *Talanta* **2007**, *73*, 142-146.

(18) Dinē, E.; Kaya, S.; Doganay, T.; Baleanu, D. *J. Pharmaceut. Biomed.* **2007**, *44*, 991-995.

(19) Teixeira, M. F. S.; Bergamini, M. F.; Marques, C. M. P.; Bocchi, N. *Talanta* **2004**, *63*, 1083-1088.

(20) Agüí, L.; Gonzalez-Cortes, A.; Yanez-Sedeno, P.; Pingarron, *J. Anal. Chim. Acta* **1999**, *401*, 145-154.

(21) Jung, M. E.; Lazarova, T. I. *J. Org. Chem.* **1997**, *62*, 1553-1555.

(22) Afkhami, A.; Nematollahi, D.; Khalafi, L.; Rafiee, M. *Int. J. Chem. Kinet.* **2005**, *37*, 17-24.

(23) Rafiee, M.; Khalafi, L. *Electrochimica. Acta* **2010**, 55, 1809-1813.

(24) Quintino, M. S. M.; Yamashita, M.; Angnes, L. *Electroanal. Chem.* **2006**, *18*, 655-661.

(25) Zare, H. R.; Namazian, M.; Nasirizadeh, N. *J. Electroanal. Chem.* **2005**, *584*, 77-83.

(26) Zare, H. R.; Eslami, M.; Namazian, M. *Electrochim. Acta* **2010**, *56*, 2160-2164.

(27) Zare, H. R.; Namazian, M.; Coote, M. L. *Electrochim. Acta* **2009**, *54*, 5353-5357.

(28) Zare, H. R.; Eslami, M.; Namazian, M.; Coote, M. L. *J. Phys. Chem. B* **2009**, *113*, 8080-8085.

(29) Zare, H. R.; Sobhani, Z.; Mazloum-Ardakani, M. *Sensor Actuat B-Chemical* **2007**, 126, 641-647.

(30) Golabi, S. M.; Zare, H. R. *J. Electroanal. Chem.* **1999**, *465*, 168-176.

(31) Bard, A. J.; Faulkner, L. R. *Electrochemical methods Fundamentals and Applications*, 2nd ed., Wiley, New York, 2001.

(32) Atkins, P. *Physical Chemistry*, Oxford University Press, Oxford, 1999.

(33) Gokel, G.W. *Dean's handbook of organic chemistry*, McGraw-Hill, New York, 2004 (see

Table 8.8 in this reference).

(34) Anantharaman, B.; Melius, C. F. *J. Phys. Chem. A* **2005**, *109*, 1734-1747.

(35) Cossi, M.; Crescenzi, O. *Theor. Chim. Acta* **2004**, 111, 162-167.

(36) Klamt, A.; Schueuermann, G. *J. Chem. Soc., Perkin Trans. 2,* **1993**, 799-805.

(37) Takano, Y.; Houk, K. *J. Chem. Theory Comput.* **2005**, *1*, 70-77.

(38) Frisch, M. J.; Trucks, G. W.; Schlegel, H. B.; Scuseria, G. E.; Robb, M. A.; Cheeseman, J.

R.; Montgomery, J. A.; T. Vreven Jr.; Kudin, K. N.; Burant, J. C.; et. al. *GAUSSIAN 03,*

Revision C.02; Gaussian, Inc.: Wallingford. CT, 2004.

(40) Parr, R. G.; Yang, W. *Density-functional theory of atoms and molecules*, Oxford University

Press, Oxford, 1989.

(41) Eslami, M.; Namazian, M.; Zare, H. R. Unpublished work, in preparation.

(42) Liu, H.; Zhang, R.; Yao, X.; Liu, M.; Hu, Z.; Fan, B. *J. Chem. Inf. Comp. Sci.* **2004,** *44*,

161-167.

(43) Brun, A.; Rosset, R. *J. Electroanal. Chem.* **1974,** *49*, 287-300.

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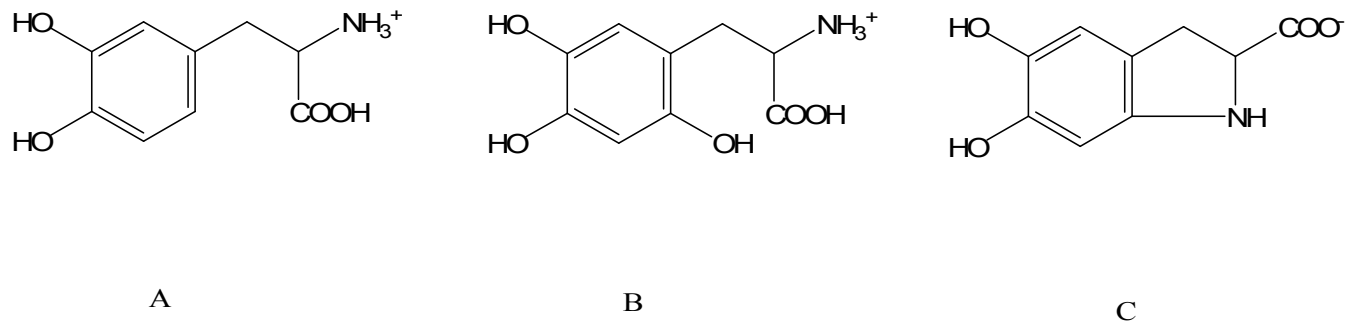
Scheme 1. Structure of (A) L-DOPA, (B) the product of hydrolysis following chemical reaction and (C) the product of intramolecular following chemical reaction.

$$\text{L-DOPAquinone}_{(gas)} \; + \; H_{2(gas)} \xrightarrow{\;\Delta G(gas)\;} \text{L-DOPA}_{(gas)}$$

$$\big\downarrow \Delta G(\text{L-DOPAquinone, solv}) \qquad\qquad\qquad \big\downarrow \Delta G(\text{L-DOPA, solv})$$

$$\text{L-DOPAquinone}_{(solv)} \; + \; H_{2(gas)} \xrightarrow{\;\Delta G(solv)\;} \text{L-DOPA}_{(solv)}$$

Scheme 2. The thermodynamic cycle proposed to convert standard Gibbs energy of redox reaction in the gas phase to solution phase.
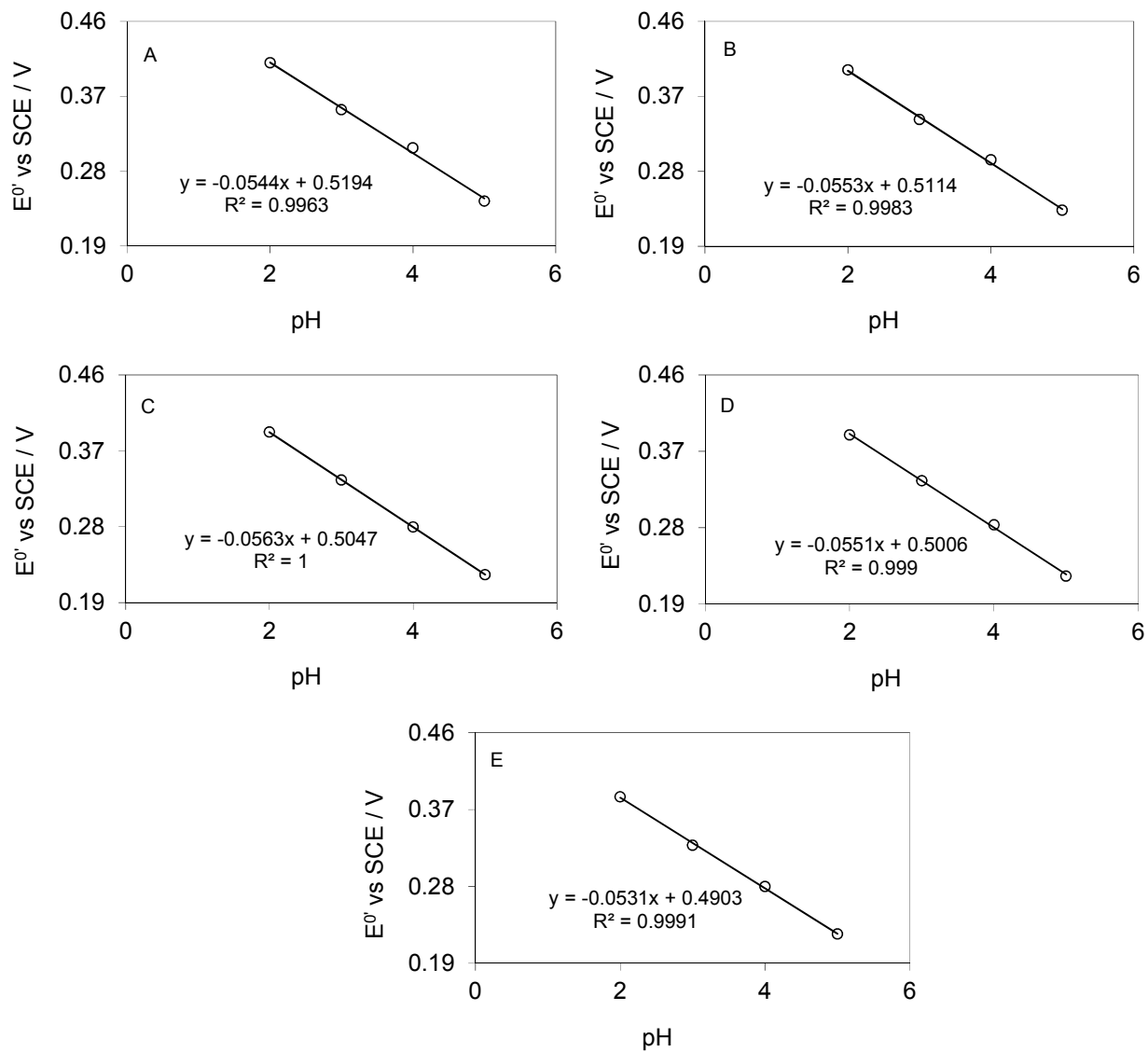
Fig. 1. Plot of the formal potential versus pH at different temperature of A) 10 °C, B) 15 °C, C) 20 °C, D) 25 °C and E) 30 °C.
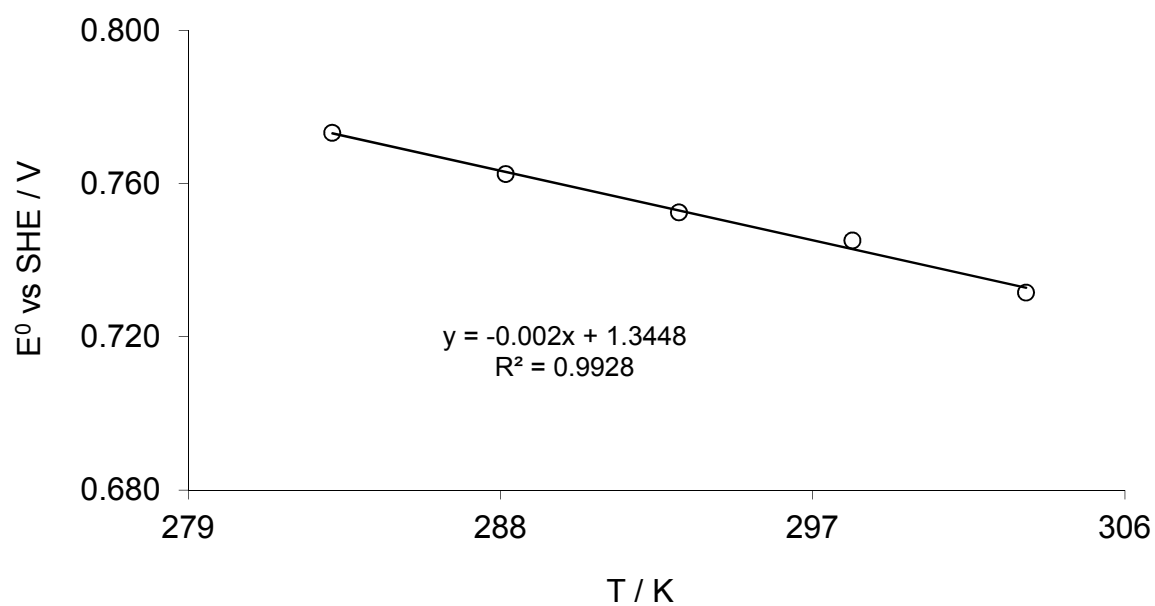
Fig. 2. The effect of temperature on the redox potential of L-DOPA.

Table . 1. Experimental redox potentials and thermodynamic functions for reactions (1) for 10.0 μM L-DOPA in an aqueous solution in the range of 10 to 30 °C (283-303 K).

property	Temperature / K				
	283.15	288.15	293.15	298.15	303.15
$E^0_1{}^a$ / V	0.773	0.763	0.752	0.745	0.731
$E_{SCE}{}^b$ / V	0.2538	0.2511	0.2477	0.2445	0.2411
$\Delta H^0_1{}^c$ / kcal mol^{-1}	-61.7	-61.7	-61.7	-61.8	-61.7
$\Delta S^0_1{}^c$ / cal mol^{-1} K^{-1}	-92.24	-92.24	-92.24	-92.24	-92.24
ΔG^0_1 / kcal mol^{-1}	-35.66	-35.17	-34.70	-34.37	-33.74

[a] Relative redox potentials vs SHE; the conversion has been carried out using the exact values of SCE at each temperature [33] [b] The exact potential values of SCE at each temperature [33] [c]The changes of enthalpies and entropies for reactions 1, are temperature independent in the studied range of temperature (linear relation of E vs T, see figure 2).

Table. 2. Calculated thermodynamic parameters of studied species using G3MP2//B3LYP level of theory together with CPCM/B3LYP/6-311+G(d,p) solvation energies.

Compound name	$G_{solv.}$ (kcal mol^{-1})	$G^{\circ \rightarrow *}$ (kcal mol^{-1}) [a]	G° (Hartree)	H° (Hartree)	S° (cal mol^{-1} K^{-1})
		Thermodynamic quantities [b]			
L-DOPA	-73.38	1.9	-704.58908	-704.53355	116.9
L-DOPAquinone	-79.51	1.9	-703.34364	-703.28702	119.1
H_2	0.0	0.0	-1.18213	-1.16734	31.1

[a] The standard state correction of 1.9 kcal mol^{-1} (RT Ln P)[26] has been considered to Gibbs energy of those species which contribute to solvation process.
[b] These quantities are calculated at pressure of 1 atm, but they can be also regarded conveniently at 1 bar.

Table. 3. G3MP2//B3LYP calculations of enthalpy and Gibbs free energy of studied reactions together with calculated/experimental redox potentials at 298 K.

Thermodynamic property	G3MP2B3
$\Delta H^0_{1\ (gas)}$ / kcal mol^{-1}	-49.7
$\Delta S^0_{1\ (gas)}$ / cal mol^{-1} K^{-1}	-33.4
$\Delta G^0_{1\ (gas)}$ / kcal mol^{-1}	-39.7
$\Delta G^0_{1\ (soln)}{}^a$ / kcal mol^{-1}	-33.6
E^0_1 (calcd)[b]/ V	0.728
E^0_1 (exp)[b, c]/ V	0.745

[a] Considering changes of Gibbs free energies of solvation of -6.1 kcal mol^{-1} for reactions 1,calculated using CPCM at B3LYP/6-311+G(d,p) level of theory. [b] Standard redox potentials are relative to SHE. [c] Experimental redox potentials have been measured in this work using cyclic voltammetry.

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$$\Delta S° = nF(\partial E°/\partial T)_P$$

$$\Delta G° = -nFE°$$

$$\Delta H° = -nF[E° - T(\partial E°/\partial T)_P]$$

E⁰ / V vs SHE

0.800

0.760

0.720

0.680

279   288   297   306

T / K