ISSN : 0974-7419 Volume 9 Issue 4

Analytical CHEMISTRY
An Indian Journal

Trade Science Inc.

Full Paper

ACAIJ, 9(4) 2010 [434-437]

Electrochemical study of the stability constants and thermodynamic parameters of bivalent metal ions with aspirin

Harminder Kaur*[1], Anita Singla[2]

Deptt. of Applied Sciences, PEC University of Technology, Chandigarh, 160 012, (INDIA)

E-mail : hkaur@pec.ac.in

Received: 28th May, 2010 ; Accepted: 7th June, 2010

ABSTRACT

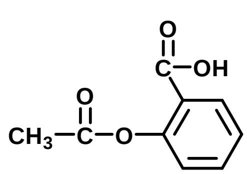
The interaction of bivalent metal ions: Co^{2+}, Ni^{2+}, Cu^{2+}, Zn^{2+}, Cd^{2+}, Sn^{2+}, $Hg^{2+}$and Pb^{2+} with aspirin in mixed aqueous solution (50% v/v water-ethanol medium) was studied by pH-potentiometry keeping fixed ionic strength at 0.1 mol/l (KNO_3) and T = 298K, 308K and 318K. The metal–ligand stability constants of complexes were evaluated using modified Irving-Rossotti technique. The corresponding thermodynamic parameters (free energy, enthalpy and entropy) have also been evaluated to confirm the feasibility of complex formation.

KEYWORDS

Aspirin;
Bivalent metal ions;
Potentiometry;
Stability constant;
Thermodynamic parameters.

INTRODUCTION

Aspirin ($C_9H_8O_4$; chemically 2- acetyloxybenzoic acid) M.W. 180.160 g/mol was the first - discovered member of the class of drugs known as non - steroidal anti-inflammatory drugs (NSAIDs).The medicinal effects of the plants were discovered by people through life experiences. Aspirin has many therapeutic effects. At over-the-counter dosage (one or two grams), it relieves fever and minor aches and pains. At dosages three or four times higher, available by prescription only, it reduces swelling and is used to treat gout, rheumatoid arthritis, and inflammatory ailments[1].



Aspirin (2- acetyloxybenzoic acid)

Many people take low dosages (below 100 milligrams) daily for preventing recurrent stroke or heart attack[2]. Recent studies found it effective in reducing risks for colon and breast cancers[3]. Evidence is accumulating for similar effects in Alzheimer and other diseases[4,5]. A great complaint against aspirin is that it irritates the stomach and in some cases cause ulcers and internal bleeding. Data revealed that those who had taken aspirin bleed longer and the platelets in their blood aggregated less. Hawkey suggested that aspirin may also prevent artery blood clots[6].

The process of complex formation of a drug with bio-metals can be successfully studied by potentiometry. A potentiometric sensor was reported for the Hg (II) detection, which uses substituted thiourea – functionalized nanoporous silica as the sensitive material[7]. The apparent dissociation constants of four 2-hydroxy napthoquinones, differently substituted at C-3 were determined by pH-metric / UV titration methods. The possible effects of the measured parameters on the biological activity were also studied[8].

Potentiometric studies had also been carried on the complexes of gabapentin, a drug widely used to relieve neuropathic pain, with bivalent metal ions in dioxane-water medium[9]. The protonation constants of some diamines were determined on the basis of Calvin and Bjerrum method and the effects of solvents were determined[10]. The proton-ligand dissociation constants and stepwise stability constants of various antibiotics metal complexes were determined potentiometrically[11-16]. During the last few years various aspects concerning the absorption, transport, activity, biological transformations, toxicity and excretion of different biometals is extensively studied[17-20].

Since for the epilepsy and neuropathic pain in our body, some of the metal ions such as Lead and Copper are responsible. These metal ions can be metabolized in our biological system by forming complexes with some drugs. The purpose of present paper is the determination of the stability constants of the complexes and thermodynamic parameters of Co^{2+}, Ni^{2+}, Cu^{2+}, Zn^{2+}, Cd^{2+}, Sn^{2+}, $Hg^{2+}$and Pb^{2+} complexes with aspirin using potentiometric pH technique in 50% v/v ethanol water medium at three different temperatures and at an ionic strength of 0.1 mol/l (KNO_3). The method of Calvin and Bjerrum[21,22] as adopted by Irving and Rossotti[23] has been employed to determine log K values.

EXPERIMENTAL

Chemicals and ligand used were of analytical grade. Ligand solution was prepared in twice distilled deionized carbon dioxide free water. Metal salt solutions were prepared by dissolving the corresponding metal salt in twice distilled deionized water and standardized by standard volumetric methods. The potentiometric titrations were carried out in a jacketed cell. The free hydrogen ion concentrations were measured with a combined glass electrode attached to a EI pH meter model 112. 50% v/v ethanol-water medium is used at three temperatures T = 298K, 308K and 318K and at an ionic strength of 0.1 mol/l (KNO_3). The pH meter was calibrated with suitable buffers before use. The three solutions (total volume 50 mL in each case) were prepared as follows: (a) 5.0 ml of 0.005 mol/l HNO_3 (b) 5.0 ml of 0.005 mol/l HNO_3 + 5.0 ml of 0.0025 mol/l ligand (c) 5.0 ml of 0.005 mol/l HNO_3 + 5.0 ml of 0.0025 mol/l Ligand + 5.0 ml of 0.005 mol/l metal ion solution.

An appropriate quantity of potassium nitrate solution (1.0 mol/l) was added to maintain the desired ionic strength (0.1 mol/l). Above three solutions were titrated against potassium hydroxide (0.05 mol/l). The solution to be titrated was taken in a cell and immersed in the thermostat for half an hour before the titration so that it attained the required temperature. After the addition of each portion of alkali the highest pH reading which remained steady was recorded in all cases.

Calculation of $\overline{n_A}$, $\bar{n}$ and pL

The values of $\overline{n_A}$ (the degree of formation of the proton complex) was calculated by employing the following equation (1):

$$\overline{n_A} = Y + \frac{(V' - V'')(N + E^o)}{(V^o + V')T_L{}^o} \qquad (1)$$

Where Y = number of replaceable hydrogen ion, V^o = total volume 50 ml, V' = volume of alkali used by acid V'' = Volume of alkali used by acid and ligand, N = concentration of alkali, E^o = total strength of acid, $T_L{}^o$ = total concentration of ligand.

The proton ligand formation curve was obtained by plotting the degree of formation ($\overline{n_A}$) of the proton complex against pH values, The values of log $K_1{}^H$ were obtained from the curves corresponding to $\overline{n_A}$ values of 0.5. The stability constants at three different temperatures were calculated by various computations methods[24,25] and are summarized in TABLE 1. The values of $\bar{n}$ (average number of ligand molecules attached per metal ion) were calculated using equation (2)

$$\bar{n} = \frac{(V''' - V'')(N + E^o)}{(V^o + V')\overline{n_A}T_M{}^o} \qquad (2)$$

where V''' = volume of alkali used for acid + ligand + metal ion, $T_M{}^o$ = total concentration of the metal ion, rest of term symbols are as given in equation (1). The free ligand exponent, pL was calculated using equation 3 as given below:

$$pL = \log_{10}\left[\frac{\sum_{n=o}^{n=j}\beta^H n(1/\text{anti}\log\beta)^n}{T_L{}^O - \bar{n}T_M{}^O}\ \frac{V^O + V^m}{V^O}\right] \qquad (3)$$

The formation curves obtained by plotting $\overline{n_A}$ against

Full Paper

pH values for proton – ligand are shown in Figure 1 at 298K, 308K and 318K. From the proton – ligand formation curve the approximate values of the pk_1^H were calculated by Bjerrum half integral method at $\overline{n_A} = 0.5$ or 1.5. The accurate values were calculated by interpolation of various $\overline{n_A}$ values[26].

The thermodynamic stability of a species is the measure of extent of its formation under a particular set of conditions. In the language of thermodynamics, the equilibrium constant of a reaction is the measure of the heat expelled from the reaction system and entropy change during the reaction. The entropy of a system is the measure of the amount of disorder. The greater the amount of disorder in the products of a reaction relative to the reactants, the greater will be the increase in entropy during the reaction and higher is the stability of products.

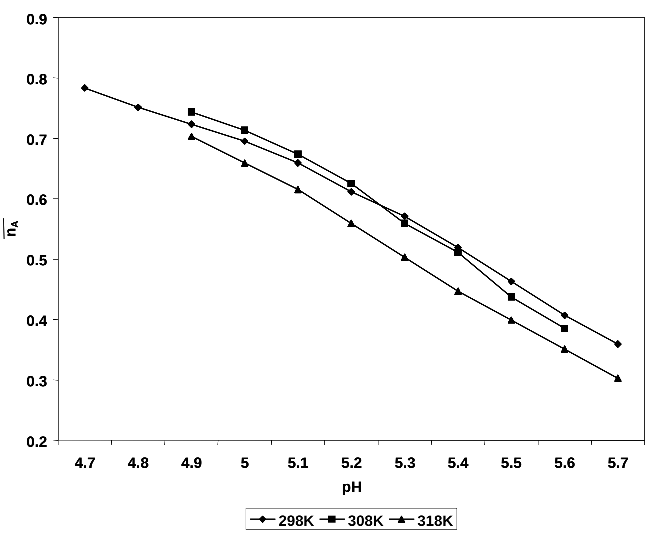


Figure 1 : Proton-ligand formation curves at three different temperatures for aspirin.

RESULTS AND DISCUSSION

Effect of pH

The titration data indicate that the ligand curve is slightly shifted to the left of the acid titration curve at lower pH value. The shift is due to the interaction of proton with the ligand and then with the metal ion.

Effect of temperature

The perusal of data in TABLE 1 shows that the values of stability constants decreases with an increase in temperature. So the degree of dissociation of ligand increases with rise in temperature but there is decrease in the stability constants of the complexes of these metal ions with ligand with temperature. The degree of ionization also increases for a ligand with temperature.

Order of stability

The order of metal – ligand stability constants (log K_1) has been found to be $Cu^{2+} > Sn^{2+} > Hg^{2+} > Pb^{2+} > Co^{2+} \approx Ni^{2+} > Zn^{2+} > Cd^{2+}$. This order is in accordance with irving - Williams order of stability[27]. The stability of Cu^{2+} complexes is much higher than what can be expected from its ionic radius, due to Jahn – Teller stabilization.

Thermodynamic functions

The values of free energies of formation of the complexes becomes more negative with increase in temperature. This shows that the complex formation is a

TABLE 1 : Metal-ligand stability constants of complexes at three temperatures

Cation		Stability Constant at three different Temperatures			$-\Delta G$ (KJ mol^{-1}) $=2.303\ R\ T \log k$			ΔH (K$Jmol^{-1}$)	ΔS ($Jmol^{-1}\ deg^{-1}$)
		298K	308K	318K	298K	308K	318K	308K	308K
H^+	pK_1^H	5.45	5.42	5.35	-	-	-	-	-
Co^{2+}	$\log K_1$	2.55	2.51	2.47	14.543	14.795	15.032	- 7.026	25.2
Ni^{2+}	$\log K_1$	2.55	2.50	2.45	14.543	14.736	14.910	- 8.782	19.3
Cu^{2+}	$\log K_1$	2.86	2.81	2.80	16.310	16.563	17.040	- 8.782	25.2
Zn^{2+}	$\log K_1$	2.52	2.48	2.44	14.372	14.618	14.850	- 7.026	24.6
Cd^{2+}	$\log K_1$	2.49	2.47	2.45	14.201	14.441	14.910	- 3.513	35.5
Sn^{2+}	$\log K_1$	2.69	2.68	2.61	15.341	15.797	15.884	-1.76	45.5
Hg^{2+}	$\log K_1$	2.66	2.61	2.51	15.170	15.384	15.275	-8.78	21.4
Pb^{2+}	$\log K_1$	2.58	2.52	2.46	14.714	14.854	14.971	- 10.539	14.0

spontaneous process and spontaneity increases with temperature. The entropy changes are positive for the system under study and thus the complexes of these metal ions with Aspirin are stabilized by relatively large positive entropy changes.

REFERENCES

[1] G.Weissman; Scientific American, **264(1)**, 84 **(1991)**.
[2] V.Fuster, M.Cohen, J.Halperin; New England Journal of Medicine, **321**, 183 **(1989)**.
[3] L.J.Marnett, R.N.DuBois; Annual Review of Pharmacology and Toxicology, **42**, 55 **(2002)**.
[4] S.Renfrey, S.Downton, J.Featherstone; Nature Review: Drug Discovery, **2**, 175 **(2003)**.
[5] P.L.McGeer, G.E.McGeer; Neurobiology of Aging, **28(5)**, 639 **(2007)**.
[6] C.J.Hawkey; The Lancet, **374**, 93 **(2009)**.
[7] M.Javanbhakht, F.Divsar, A.Badiei, M.R.Ganjali, P.Norouzi, G.M.Ziarani, M.Chaloosi, A.A.Jahangir; Analytical Sciences, **25**, 789 **(2009)**.
[8] T.Ossowskil et al.; Journal of the Brazilian Chemical Society, **19**, 1 **(2008)**.
[9] A.Kumar, A.K.Malik; Eurasion Journal of Analytical Chem., **2**, 2 **(2007)**.
[10] S.Sharma, M.C.Shah, N.Patel, D.Dalwadi, J.J.Vora; E-Journal of Chemistry, **4(3)**, 313 **(2007)**.
[11] B.Jeragh, D.Al-Wahaib, A.A.El-Sherif, J.El-Dissouky; Chem.Eng.Data, **52(5)**, 1609 **(2007)**.
[12] A.Dogan, E.Kilic; Turk J.Chem., **29**, 41 **(2005)**.
[13] F.Borges, C.Guimareas, J.L.F.C.Lima, I.Pinto, S.Reis; Talanta, **66(3)**, 670 **(2005)**.
[14] M.M.Khalil, A.E.Attia; J.Chem.Eng.Data, **45(6)**, 1108 **(2000)**.
[15] A.A.A.Gaber, O.A.Farghaly, M.A.Ghadour, H.S.El-Said; Monatshefte fur Chemie, **131**, 1031 **(2000)**.
[16] A.P.Arnold, A.L.Canty, R.S.Reid, D.L.Rebenstein; Can.J.Chem., **63**, 2430 **(1985)**.
[17] G.Zampella, P.Fantucci, V.L.Pecoraro, L.D.Gioia; J.Am.Chem.Soc., **127**, 953 **(2005)**.
[18] T.Sismanoglu; Chinese Chemical Letters, **14(11)**, 1207 **(2003)**.
[19] A.Dean et al.; Dalton Trans., 1689 **(2008)**.
[20] E.Mohajeri, F.Moattar; Asian J.of Chemistry, **21(7)**, 4987 **(2009)**.
[21] M.Calvin, K.W.Wilson; J.Amer.Chem.Soc., **67**, 2003 **(1945)**.
[22] J.Bjerrum; 'Metal Amine Formation in Aqueous Solution'. P Hasse and Son Coppenhagen, 298 **(1941)**.
[23] H.Irving, H.S.Rossotti; J.Amer.Chem.Soc., 2904 **(1954)**.
[24] A.I.Vogel; 'Text Book of Practical Organic Chemistry', Longman, London **(1968)**.
[25] H.Irving, H.S.Rossotti; J.Amer.Chem.Soc., 3397 **(1953)**.
[26] L.G.Van Uitert, C.G.Haas; J.Amer.Chem.Soc., **75**, 451 **(1953)**.
[27] R.C.Agarwal, H.N.Singh, R.C.Saxena; Thermochim Acta, **56**, 371 **(1982)**.