JOURNAL OF
CHEMICAL &
ENGINEERING
DATA

ARTICLE

pubs.acs.org/jced

Density, Viscosity, Heat Capacity, Surface Tension, and Solubility of CO_2 in Aqueous Solutions of Potassium Serinate

Ho-Jun Song, Min-Gu Lee, Hyuntae Kim, Ankur Gaur, and Jin-Won Park*

Department of Chemical and Biomolecular Engineering, Yonsei University, 262 Seongsanno, Seodaemun-gu, Seoul 120-749, South Korea

ABSTRACT: The density, viscosity, and heat capacity of aqueous potassium serinate solutions were measured over the temperature range from (298.15 to 353.15) K in the mass fraction range of (7.2 to 57.3) %. Each physical property was correlated with temperature and mass fraction, by regression analysis, and the corresponding coefficients for each property are presented here. The surface tension of the aqueous potassium serinate solution was also measured over the mass fraction range of (7.2 to 57.3) % at 298.15 K. In addition, the solubility of carbon dioxide in a 14.3 mass % aqueous potassium serinate solution was measured at (313.15, 343.15, and 373.15) K over partial pressures of CO_2 ranging from (0.1 to 400) kPa. It was shown that the aqueous potassium serinate solution could be a good CO_2 absorbent because of its high cyclic CO_2 absorption capacity. Furthermore, it might be suitable for membrane gas absorption due to its higher surface tension.

INTRODUCTION

The world is facing an uncontrolled environmental problem: global warming. The increase in atmospheric concentration of carbon dioxide, due to the continuous combustion of fossil fuels for power generation and various industrial activities, is accelerating global warming. According to energy technology perspectives of the International Energy Agency (IEA), although the consumption of fossil fuels will inevitably continue, the employment of CO_2 capture and storage (CCS) in power generation, industry, and transformation sectors could reduce the amount of CO_2 emission by 19 % (11.8 Gt on a global scale) by 2050.[1] Of the available techniques for the removal of acid from the gas stream, the wet chemical absorption method using aqueous alkanolamine is currently thought to be the most feasible option for the capture of postcombustion CO_2 from large stationary sources.[2] The wet chemical absorption process consists basically of an absorber, a desorber, and heat exchangers. The CO_2 removal occurs in the absorber, by contacting CO_2-containing flue gas and absorbent in the countercurrent at a low temperature (e.g., 40 °C). The CO_2-free flue gas, then, is released into the atmosphere. Simultaneously, the CO_2-rich absorbent solution is transferred to the desorber and regenerated by the reboiler heat; the bottom temperature of the desorber is around 120 °C, releasing almost pure CO_2 overhead. Industrially important alkanolamines for the acid gas (e.g., H_2S and CO_2) treatment are monoethanolamine (MEA), diethanolamine (DEA), *N*-methyldiethanolamine (MDEA), 2-amino-2-methyl-1-propanol (AMP), and so forth. In the last half century, there have been extensive and intensive investigations into alkanolamines for acid gas treatment purposes.[3] Nevertheless, alkanolamines have a lot of problems which must be solved for them to be utilized as an economically practical CCS technique. Alkanolamine suffers from thermal and oxidative degradation during the cyclic CO_2 absorption−desorption operation.[4,5] Corrosion of process equipment composed of carbon and stainless steel is another problem caused by dissolved CO_2 and O_2 and alkanolamine itself.[6] Furthermore, some alkanolamines such as MEA need a large amount of thermal energy to recover the aqueous alkanolamine saturated with CO_2.[7]

Recently, environmentally friendly amino acid salts have been widely investigated as acid gas scrubber candidates. Amino acid salts have demonstrated better resistance to the oxidative degradation when they come in contact with oxygen-rich flue gas.[8] Due to the ionic structure of amino acid salt, the vapor pressure of the aqueous solution is relatively low, so the vapor loss might be reduced at a high temperature.[9] An unique advantage of aqueous amino acid salt over aqueous alkanolamine is its high surface tension. Surface tensions of some aqueous amino acid salts are even higher than that of pure water, thus making them suitable as CO_2 absorbing liquids for the membrane-gas absorption lowering pore wetting problem.[10,11] Membrane pore wetting behavior can be simply explained by the following Laplace−Young equation:

$$\Delta P = -\frac{4\sigma_L \cos\theta}{d_{max}} \qquad (1)$$

where ΔP is the minimum pressure (breakthrough pressure) to be applied to the absorbing liquid to enter the membrane pore, σ_L is the surface tension of the absorbing liquid, θ is the contact angle between the absorbing liquid and the membrane, and d_{max} is the maximum diameter of the pore. Thus, the higher the surface tension value of the absorbing liquid is, the lower the pore wettability.

Lee et al.[12−15] and Song et al.[16,17] have intensively studied the absorption of carbon dioxide into the aqueous sodium glycinate (SG) solution. They conducted investigations into the reaction

Special Issue: John M. Prausnitz Festschrift

Received: October 30, 2010
Accepted: February 15, 2011
Published: March 07, 2011

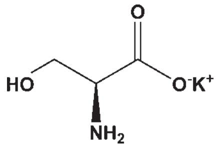

Figure 1. Chemical structure of potassium serinate.

kinetics[14] and solubility[16] of CO_2 in an aqueous SG solution as well as the measurements of some physical properties of the aqueous SG solution.[12,13] Recently, the reaction enthalpies, CO_2 solubilities, and densities of aqueous SG solutions before and after CO_2 absorption have been reported by researchers.[18,19] The CO_2 absorption characteristics of the aqueous potassium salts of glycine and taurine have been well-researched by Kumar et al.[10,20−23] By the breakthrough pressure measurement of aqueous potassium taurate for a hydrophobic polyolefin membrane, it was revealed that the aqueous solution of potassium taurate could be an excellent CO_2 absorbing liquid which can carry out the membrane-gas absorption process in a stable condition.[10] Ahn et al.[24] investigated the corrosion characteristics of carbon steel in the aqueous potassium salts of glycine and taurine and the aqueous blends of potassium taurate and piperazine. Van Holst et al.[25] screened salts of various common amino acids by plotting the overall rate constant (k_{ov}) versus the acidity constant (pK_a) and found that the aqueous salts of sacosine and proline are the most promising solvents. Muñoz et al.[26] conducted cyclic CO_2 absorption and desorption experiments on the aqueous potassium salts of synthesized amino acids as well as some common amino acids. In the study by Muñoz et al., some synthetic amino acids based on a piperazine ring showed outstanding performance. In their study, the tested aqueous salts of common amino acids, namely serine, histidine, arginine, and ornithine, showed superior cyclic CO_2 absorption capacity to MEA. Zhang et al.[27] synthesized a new type of ionic liquid by the addition of serine, glycine, alanine, and lysine as an anion to phosphonium ($[PR_1R_2R_3R_4]^+$). The synthesized amino acid ionic liquids were doped on the porous silicagel support, and their CO_2 sorption performances were tested.

As shown in Figure 1, the potassium salt of serine might be a good CO_2 absorbent due to its structural features. Potassium serinate is a primary amine which enables easier access of CO_2 to its nitrogen atom. The hydroxymethyl group ($-CH_2OH$) attached to the α-carbon could cause a steric hindrance effect, thus resulting in lower carbamate stability and enhanced CO_2 absorption capacity.[8] These structural advantages of potassium serinate prompted us to investigate the aqueous potassium serinate solution as a potential CO_2 absorbent. In the present study, CO_2 solubility as well as some physical properties of an aqueous potassium serinate solution were measured. The measured physicochemical properties are essential for the design of CO_2 absorption facilities such as a gas−liquid contactor and heat exchanger and could be used for further research on issues such as the estimation of gas diffusivity, reaction rate constant, and so forth.

■ EXPERIMENTAL SECTION

Aqueous potassium serinate solutions were prepared by neutralizing serine (Sigma-Aldrich, ≥ 98 %), which had been dissolved in demineralized water, with an equimolar quantity of potassium hydroxide (J. T. Baker, 45 wt % aqueous solution). The mass fraction of potassium serinate in the aqueous solution was determined by titrating it against 0.1 M H_2SO_4. The measured mass fraction was accurate to ± 0.5 %. MEA (≥ 98 %) and H_2SO_4 were purchased from Sigma-Aldrich and Duksan Chemical Co., respectively. The CO_2 and N_2 gases used for the measurement of CO_2 solubility into the aqueous potassium serinate solution were purchased from Gas Valley. The purity of both gases was 99.9 %. The selected solution concentrations for the measurements, 7.2, 14.3, 28.6, 43.0, and 57.3 mass %, were identical with the corresponding molarities, 0.5, 1.0, 2.0, 3.0, and 4.0 M, and the corresponding mole fractions, 0.0096, 0.0206, 0.0481, 0.0867, and 0.1444, respectively.

CO_2 Solubility. The solubilities of carbon dioxide in 14.3 mass % aqueous potassium serinate solutions were measured at (313.15, 343.15, and 373.15) K over partial pressures of CO_2 ranging from (0.1 to 400) kPa. The experimental setup for the measurement of CO_2 solubility was mainly composed of a 3960 cm^3 gas loading cylinder and a 503 cm^3 equilibrium cell. Both were equipped with a K-type thermocouple and a pressure transducer. Before each run, the whole system was flushed with pure nitrogen to eliminate trace gases. After that, the aqueous absorbent solution and CO_2 gas were fed into the equilibrium cell and loading cylinder, respectively, and maintained at the same designated temperature in a forced convection air bath. By opening the valve between the equilibrium cell and the gas loading cylinder, CO_2 was fed from the gas cylinder to the equilibrium cell which contained the aqueous absorbent solution. Then, the stirrer of the cell was turned on to reach the vapor−liquid equilibrium quickly. During the run, pressures and temperatures of the cell and the gas cylinder were continuously monitored. When the vapor−liquid equilibrium was attained, 0.5 cm^3 of gas phase was extracted from the cell and analyzed by gas chromatography analysis to determine the N_2/CO_2 ratio. Then, the CO_2 loading (mol of CO_2/mol of amine) was determined by the application of experimental $P-T$ data to the thermodynamic virial equation.[28] The materials, experimental procedures, and analysis methods are basically identical to those used in our previous study, which measured the solubility of CO_2 in an aqueous sodium glycinate solution.[17] The uncertainty of the CO_2 solubility measurements was 2.5 %.

Density and Viscosity. The densities of the aqueous potassium serinate solutions were measured at 10 K intervals, from (298.15 to 353.15) K, using a 36.940 cm^3 (at 298.15 K) Gay-Lussac pycnometer. The temperature of the thermostatic water bath (RBC-11, Woori Science Instrument Co.) was controlled within ± 0.1 K of the designated temperature. The internal volume of the pycnometer was calibrated using pure, degassed water, and the densities of the pure water samples were compared with data reported in the literature.[29] The uncertainty of the density measurements was 0.04 %. The dynamic viscosities were measured directly using a commercial viscometer (RheoStress1, HAAKE Instruments Inc.) equipped with a DG 41 spindle, in which the temperature was kept within ± 0.01 K of the target temperature. The uncertainty of the viscosity measurements was 0.9 %. The reported values are the average of at least three measurements.

Heat Capacity. The heat capacities were determined at 5 K intervals, from (303.15 to 353.15) K, using a differential scanning calorimeter (Q-10, TA Instruments). The method employed in this study is similar to that of Chiu et al.[30] The calorimeter was calibrated with a 22.5 mg standard sapphire window (TA Instrument - Waters LLC). The heat capacities of sapphire for

Table 1. Solubility of CO_2 α in Aqueous 15.3 mass % MEA Solution

T = 313.15 K		T = 373.15 K	
α^a	$P_{CO_2}{}^b$/kPa	α^a	$P_{CO_2}{}^b$/kPa
0.347	0.1	0.093	0.5
0.394	0.3	0.163	2.1
0.439	0.8	0.223	4.1
0.480	2.4	0.297	9.9
0.551	16.0	0.359	20.7
0.635	65.0	0.426	47.3
0.720	182.0	0.513	129.4
0.804	402.7	0.587	290.0

[a] mol of CO_2/mol of MEA. [b] N_2 was present in the gas phase for the gas chromatography analysis.

Table 2. Solubility of CO_2 α in Aqueous 14.3 mass % Potassium Serinate Solution

T = 313.15 K		T = 343.15 K		T = 373.15 K	
α^a	$P_{CO_2}{}^b$/kPa	α^a	$P_{CO_2}{}^b$/kPa	α^a	$P_{CO_2}{}^b$/kPa
0.273	0.2	0.112	0.2	0.030	0.1
0.392	1.0	0.212	1.1	0.091	0.9
0.502	3.9	0.305	3.9	0.145	3.2
0.605	12.9	0.386	9.6	0.248	13.5
0.692	31.3	0.461	19.6	0.339	32.0
0.761	59.9	0.582	56.4	0.450	79.6
0.866	151.2	0.650	96.4	0.542	156.9
0.996	433.5	0.845	393.0	0.683	387.2

[a] mol of CO_2/mol of potassium serinate [b] N_2 was present in the gas phase for the gas chromatography analysis.

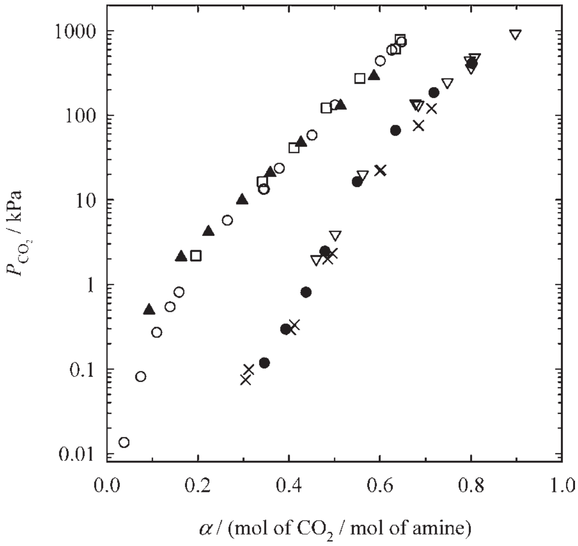

Figure 2. Comparison of the solubility of CO_2 in aqueous 15.3 mass % MEA solution. Experimental data: ●, 313.15 K; ▲, 373.15 K. Data of Jones et al.:[32] ×, 313.15 K; ○, 373.15 K. Data of Lee et al.;[33] ▽, 313.15 K; □, 373.15 K.

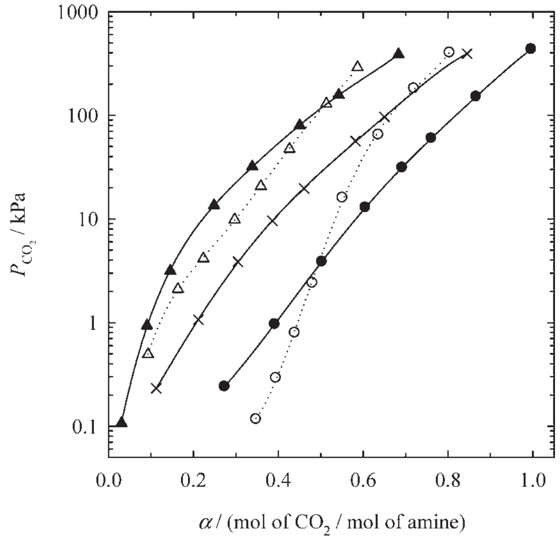

Figure 3. Comparison of the solubility of CO_2 in aqueous amine solutions. The points show the experimental data. The lines show the smoothed values. Experimental data of aqueous 14.3 mass % potassium serinate solution: ●, 313.15 K; ×, 343.15 K; ▲, 373.15 K. Experimental data of aqueous 15.3 mass % MEA solution: ○, 313.15 K; △, 373.15 K.

the calibration were taken from the literature.[31] The aqueous potassium serinate solution was prepared in a hermetic pan to resist the vapor pressure of the aqueous solution, and the mass of the aqueous solution was from (20 to 30) mg. The sample-containing hermetic pan was kept in a nitrogen-flushed atmosphere, with a nitrogen flow rate of 50 $cm^3 \cdot min^{-1}$. The experiment was initiated by maintaining the sample-containing hermetic pan at 298.15 K for 10 min. Then, the pan was heated at a rate of 5 $K \cdot min^{-1}$ until it reached 355.15 K (82 °C) and was then held isothermally for 10 min. The heat flow to the sample and the temperature of the pan were continuously recorded during the run. The heat capacities of the aqueous potassium serinate solutions can be calculated using the following equation:

$$C_P = \left[\frac{60E}{H_r}\right]\frac{\Delta H}{m} \qquad (2)$$

where C_P is the specific heat (in $J \cdot g^{-1} \cdot K^{-1}$), E is the cell calibration coefficient at the temperature, H_r is the heating rate (in $K \cdot min^{-1}$), ΔH is the heat flow difference between the sample-containing pan and an empty hermetic pan (in mW), and m is the sample mass (in mg). The reported values are the averages of at least six measurements. The uncertainty in the heat capacity measurements was 0.9 %.

Surface Tension. The surface tensions were measured directly at 298.15 K, using a bubble pressure meter (BP2, Krüss) which uses the maximum bubble pressure method. The surface tension was measured for over 50 000 ms, to determine the equilibrium value of the surface tension. The temperature of the aqueous solutions was controlled within ± 0.1 K. The uncertainty of the surface tension measurements was 0.3 %.

■ RESULTS AND DISCUSSION

CO_2 Solubility. The solubilities of CO_2 in aqueous 15.3 mass % MEA solutions at (313.15 and 373.15) K were measured and are shown in Table 1. The experimental solubility data for the aqueous 15.3 mass % MEA solution are also depicted with those of Jones et al.[32] and Lee et al.[33] in Figure 2. As shown in Figure 2,

Table 3. Density ρ, Viscosity η, and Heat Capacity C_P of Pure MEA

T/K	ρ/g·cm^{-3}		η/mPa·s		C_P/J·g^{-1}·K^{-1}	
	this work	ref 34	this work	ref 35	this work	ref 30
303.15	1.0087	1.0089	15.36	15.11	2.68	2.74
313.15	1.0004	1.0002	10.26	10.02	2.75	2.78
323.15	0.9925	0.9918	7.190	6.972	2.82	2.83

Table 4. Density ρ of Potassium Serinate (1) + Water (2)

T/K	100 w_1				
	7.2	14.3	28.6	43.0	57.3
			ρ/g·cm^{-3}		
298.15	1.0347	1.0687	1.1370	1.2012	1.2635
303.15	1.0331	1.0667	1.1344	1.1988	1.2605
313.15	1.0288	1.0619	1.1291	1.1931	1.2544
323.15	1.0246	1.0569	1.1238	1.1875	1.2485
333.15	1.0191	1.0515	1.1180	1.1816	1.2423
343.15	1.0139	1.0463	1.1125	1.1759	1.2363
353.15	1.0078	1.0396	1.1060	1.1696	1.2296

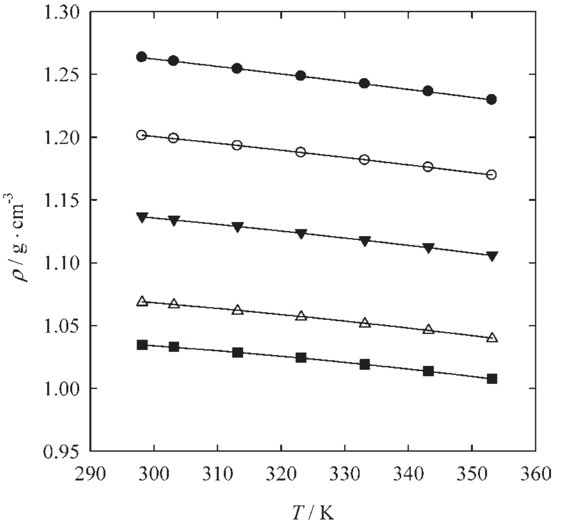

Figure 4. Density of potassium serinate (1) + water (2) at different temperatures and mass fractions. ■, w_1 = 7.2 %; △, w_1 = 14.3 %; ▼, w_1 = 28.6 %; ○, w_1 = 43.0 %; ●, w_1 = 57.3 %; the curved lines are correlations from eq 3.

Table 5. Viscosity η of Potassium Serinate (1) + Water (2)

T/K	100 w_1				
	7.2	14.3	28.6	43.0	57.3
			η/mPa·s		
298.15	1.121	1.372	1.913	2.757	4.202
303.15	1.032	1.261	1.736	2.485	3.685
313.15	0.870	1.043	1.402	1.996	2.927
323.15	0.730	0.896	1.140	1.767	2.556
333.15	0.645	0.806	0.989	1.560	2.278
343.15	0.564	0.671	0.855	1.423	2.063
353.15	0.494	0.564	0.732	1.255	1.919

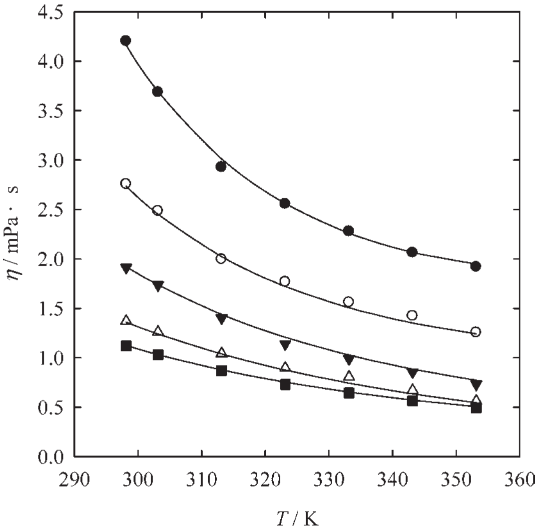

Figure 5. Viscosity of potassium serinate (1) + water (2) at different temperatures and mass fractions. ■, w_1 = 7.2 %; △, w_1 = 14.3 %; ▼, w_1 = 28.6 %; ○, w_1 = 43.0 %; ●, w_1 = 57.3 %; the curved lines are correlations from eq 4.

the CO_2 solubilities at (313.15 and 373.15) K of this study are in good agreement with the values found in the literature.

The measured solubilities of CO_2 in the aqueous 14.3 mass % potassium serinate at (313.15, 343.15, and 373.15) K are shown in Table 2. The solution concentration of 14.3 mass % was selected, based on experience with CO_2 solubility measurements, because the concentration was safely low enough for the prevention of precipitation during CO_2 absorption. In addition, the solubility of CO_2 into an aqueous 14.3 mass % potassium serinate could be simply compared to that of 15.3 mass % (2.5 M) MEA, and there is plenty of available literature data to support this. No precipitate was observed for both absorbents during the measurements at the experimental conditions. To compare the CO_2 absorption capacity, the solubilities of CO_2 in the aqueous 14.3 mass % potassium serinate and the 15.3 mass % MEA solutions are depicted in Figure 3. For both absorbent solutions, the CO_2 solubilities increase with an increase in CO_2 partial pressure and decrease with an increase in temperature. As can be seen in Figure 3, at the absorber condition in the real acid gas treatment process (i.e., at 313.15 K), the CO_2 solubility in the aqueous potassium serinate solution is much higher than that of the MEA above 4 kPa of partial pressure of CO_2. However, the CO_2 solubility in the aqueous potassium serinate solution at 373.15 K is slightly lower than that of the MEA. Thus, the cyclic CO_2 absorption capacity of potassium serinate seems to be higher than that of MEA. The similar tendency of potassium serinate and MEA can also be seen in the results of Muñoz et al.[26]

Density and Viscosity. To validate the experimental methods for the measurements of density, viscosity, and heat capacity, each property of MEA was measured at (303.15, 313.15, and 323.15) K and compared with values from the literature,[30,34,35] as reported in Table 3. The measured data for MEA were in good agreement with the data in the literature. The measured densities and viscosities of the aqueous potassium serinate solutions, over

Table 6. Heat Capacity C_P of Potassium Serinate (1) + Water (2)

T/K	100 w_1				
	7.2	14.3	28.6	43.0	57.3
	C_P/J·g^{-1}·K^{-1}				
303.15	3.88	3.72	3.35	3.06	2.74
308.15	3.91	3.77	3.38	3.08	2.76
313.15	3.93	3.79	3.40	3.11	2.78
318.15	3.96	3.82	3.42	3.13	2.80
323.15	3.98	3.84	3.44	3.16	2.81
328.15	4.00	3.86	3.47	3.18	2.83
333.15	4.02	3.89	3.49	3.20	2.84
338.15	4.04	3.91	3.51	3.22	2.85
343.15	4.06	3.94	3.53	3.23	2.86
348.18	4.07	3.96	3.55	3.25	2.86
353.15	4.08	3.97	3.56	3.26	2.90

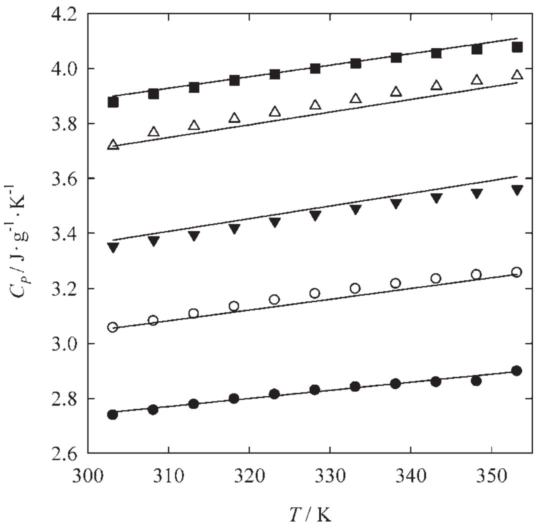

Figure 6. Heat capacity of potassium serinate (1) + water (2) at different temperatures and mass fractions. ■, w_1 = 7.2 %; △, w_1 = 14.3 %; ▼, w_1 = 28.6 %; ○, w_1 = 43.0 %; ●, w_1 = 57.3 %; the straight lines are correlations from eq 5.

the temperature range from (298.15 to 353.15) K in the mass fraction range of (7.2 to 57.3) %, are shown in Table 4 and Figure 4 and in Table 5 and Figure 5, respectively. The symbol w_1, in Tables 4 and 5 and Figures 4 and 5, is the mass fraction of the potassium serinate in aqueous solutions. The density decreases with an increase in temperature but increases with potassium serinate mass fraction. The viscosity decreases exponentially with an increase in temperature and increases with an increase in mass fraction. The density and viscosity are correlated to temperature and mass fraction by eqs 3 and 4,[36] respectively.

$$\rho/\text{g}\cdot\text{cm}^{-3} = K_1 + K_2(T/\text{K}) + K_3(T/\text{K})^2 \qquad (3)$$

$$\ln \eta/\text{mPa}\cdot\text{s} = K_4 + K_5/(T/\text{K}) + K_6(T/\text{K}) \qquad (4)$$

The curved lines in Figures 4 and 5 for the aqueous potassium serinate solutions correspond to the values correlated by eq 3 and 4, respectively.

Table 7. Coefficients and the AADs for the Correlations According to eqs 3 to 6

i/j	$k_{i,j}$			
	1	2	3	4
	Density (ρ), eq 3			
1	$7.730\cdot10^{-1}$	$2.322\cdot10^{0}$	$-4.107\cdot10^{0}$	$3.410\cdot10^{0}$
2	$1.774\cdot10^{-3}$	$-1.060\cdot10^{-2}$	$2.330\cdot10^{-2}$	$-1.950\cdot10^{-2}$
3	$-3.397\cdot10^{-6}$	$1.489\cdot10^{-5}$	$-3.236\cdot10^{-5}$	$2.668\cdot10^{-1}$
100 AAD	0.02			
	Viscosity (η), eq 4			
4	$-1.760\cdot10^{1}$	$2.230\cdot10^{2}$	$-8.978\cdot10^{2}$	$8.005\cdot10^{2}$
5	$3.333\cdot10^{3}$	$-3.198\cdot10^{4}$	$1.328\cdot10^{5}$	$-1.181\cdot10^{5}$
6	$2.120\cdot10^{-2}$	$-3.776\cdot10^{-1}$	$1.508\cdot10^{0}$	$-1.335\cdot10^{0}$
100 AAD	2.16			
	Heat Capacity (C_P), eq 5			
7	$3.044\cdot10^{0}$	$-6.730\cdot10^{0}$	$1.276\cdot10^{1}$	$-8.079\cdot10^{0}$
8	$3.451\cdot10^{-3}$	$1.320\cdot10^{-2}$	$-3.900\cdot10^{-2}$	$2.510\cdot10^{-2}$
100 AAD	0.22			

Table 8. Surface Tension σ of Potassium Serinate (1) + Water (2) at 298.15 K

T/K	100 w_1				
	7.2	14.3	28.6	43.0	57.3
	σ/mN·m^{-1}				
298.15	73.2	74.1	77.9	80.4	82.2

Heat Capacity. The measured heat capacities, over the temperature range from (303.15 to 353.15) K in the mass fraction range of (7.2 to 57.3) %, of the aqueous potassium serinate solutions are presented in Table 6 and depicted in Figure 6. The heat capacity increases with an increase in temperature and decreases with an increase in potassium serinate mass fraction. The heat capacity is correlated by eq 5[30,36] as

$$C_P/\text{J}\cdot\text{g}^{-1}\cdot\text{K}^{-1} = K_7 + K_8(T/\text{K}) \qquad (5)$$

The straight lines in Figure 6 correspond to the values correlated by eq 5.

The K_i values used in eqs 3 to 5 are given as[36]

$$K_i = k_{i,1} + k_{i,2}w_1 + k_{i,3}{w_1}^2 + k_{i,4}{w_1}^3 \qquad (6)$$

The coefficient, $k_{i,j}$, of eq 6 and the absolute average deviations (AADs) for density, viscosity, and heat capacity are presented in Table 7. The absolute average deviation (AAD) was calculated as

$$\text{AAD} = \frac{1}{n}\sum\frac{|y_i^{\text{exptl}} - y_i^{\text{calcd}}|}{y_i^{\text{exptl}}} \qquad (7)$$

where n is the number of data points taken in the measurement and y_i^{exptl} and y_i^{calcd} represent the experimental and calculated values. The AAD is a convenient measure to show how well the data are correlated with the corresponding equation.[12,34,35] In the

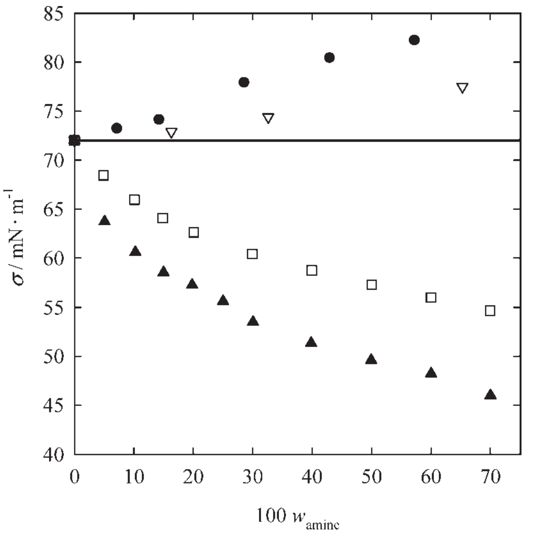

Figure 7. Surface tension of aqueous solutions of amino acid salts and alkanolamines at 298.15 K. — (horizontal line, 72.01 mN·m^{-1}), pure water;[37] ▽, potassium taurate;[10] □, MEA;[37] ▲, MDEA;[38] ●, potassium serinate (this study).

present study, the AADs for the density, viscosity, and heat capacity values are low, as shown in Table 7.

Surface Tension. The measured values for the surface tension of aqueous potassium serinate solutions at 298.15 K are listed in Table 8. The surface tension values of aqueous potassium serinate solutions increase with an increase in solution concentration. In Figure 7, the surface tension of the aqueous potassium serinate solution as a function of mass fraction was compared with other aqueous alkanolamine solutions at 298.15 K.[10,37,38] Unlike the aqueous solutions of alkanolamine, the surface tensions of the aqueous salts of serine and taurine increase with an increase in solution mass fraction. Kumar et al.[10] and Yan et al.[11] pointed out that the salts of taurine and glycine might be suitable CO_2 absorbents for the membrane-gas absorption (MGA) system, because of their high surface tensions resulting in a reduced pore wettability. Aqueous potassium serinate solution has an even higher surface tension than the aqueous potassium taurate solution and pure water, so it could also be a suitable absorbing liquid for MGA.

CONCLUSION

In this study, potassium serinate was proposed as a novel candidate material for CO_2 absorption. The densities, viscosities, and heat capacities of aqueous potassium serinate solutions were measured over a temperature range from (298.15 to 353.15) K and at the mass fractions of (7.2, 14.3, 28.6, 43.0, and 57.3) %. The surface tensions of the solutions were also measured at 298.15 K. The CO_2 solubility in the aqueous 14.3 mass % potassium serinate solution was measured at (313.15, 343.15, and 373.15) K and compared with those of the aqueous 15.3 mass % MEA solution. The aqueous potassium serinate solution is thought to be a good CO_2 absorbent due to its high cyclic CO_2 absorption capacity. Furthermore, its high surface tension values would be desirable and suitable for the membrane gas absorption process.

AUTHOR INFORMATION

Corresponding Author

*Tel.: +82-2-364-1807. Fax: +82-2-312-6401. E-mail: jwpark@yonsei.ac.kr.

Funding Sources

This work was supported by the Power Generation and Electricity Delivery of the Korea Institute of Energy Technology Evaluation and Planning (KETEP) grant funded by the Ministry of Knowledge Economy, Republic of Korea (No. 20091010-10006B).

REFERENCES

(1) International Energy Agency. *Energy Technology Perspectives 2008 Scenarios & Strategies to 2050*; International Energy Agency: France, 2008.

(2) Rostam-Abadi, M.; Chen, S.; Lu, Y. Assessment of Carbon Capture Options for Power Plants, *Proceedings of the 4th Annual Conference on Carbon Capture & Sequestration*, Alexandria, VA, USA, 2005.

(3) Kohl, A. L.; Nielsen, R. B. *Gas Purification*, 5th ed.; Gulf Publishing Company: Houston, TX, 1997.

(4) Dawodu, O. F.; Meisen, A. Degradation of Alkanolamine Blends by Carbon Dioxide. *Can. J. Chem. Eng.* **1996**, *74*, 960–966.

(5) Lepaumier, H.; Picq, D.; Carrette, P. L. New Amines for CO_2 Capture. II. Oxidative Degradation Mechanisms. *Ind. Eng. Chem. Res.* **2009**, *48*, 9068–9075.

(6) Veawab, A.; Tontiwachwuthikul, P.; Bhole, S. D. Studies of Corrosion and Corrosion Control in a CO_2-2-Amino-2-methyl-1-propanol (AMP) Environment. *Ind. Eng. Chem. Res.* **1997**, *36*, 264–269.

(7) Sakwattanapong, R.; Arronwilas, A.; Veawab, A. Behavior of Reboiler Heat Duty for CO_2 Capture Plants Using Regenerable Single and Blended Alkanolamines. *Ind. Eng. Chem. Res.* **2005**, *44*, 4465–4473.

(8) Hook, R. J. An Investigation of Some Sterically Hindered Amines as Potential Carbon Dioxide Scrubbing Compunds. *Ind. Eng. Chem. Res.* **1997**, *36*, 1779–1790.

(9) Goan, J. C.; Miller, R. R.; Piatt, V. R. *Alkazid M as a Regenerative Carbon Dioxide Absorbent*; NRL Report 5465; Naval Research Laboratory: Washington, DC, 1960.

(10) Kumar, P. S.; Hogendoorn, J. A.; Feron, P. H. M.; Versteeg, G. F. New Absorption Liquids for the Removal of CO_2 from Dilute Gas Streams using Membrane Contactors. *Chem. Eng. Sci.* **2002**, *57*, 1639–1651.

(11) Yan, S. P.; Fang, M. X.; Zhang, W. F.; Wang, S. Y.; Xu, Z. K.; Luo, Z. Y.; Cen, K. F. Experimental Study on the Separation of CO_2 from Flue Gas using Hollow Fiber Membrane Contactors without Wetting. *Fuel Process. Technol.* **2007**, *88*, 501–511.

(12) Lee, S.; Choi, S. I.; Maken, S.; Song, H. J.; Shin, H. C.; Park, J. W.; Jang, K. R.; Kim, J. H. Physical Properties of Aqueous Sodium Glycinate Solution as an Absorbent for Carbon Dioxide Removal *J. Chem. Eng. Data* **2005**, *50*, 1773–1776.

(13) Lee, S.; Song, H. J.; Maken, S.; Shin, H. C.; Song, H. C.; Park, J. W. Physical Solubility and Diffusivity of N_2O and CO_2 in Aqueous Sodium Glycinate Solutions. *J. Chem. Eng. Data* **2006**, *51*, 504–509.

(14) Lee, S.; Song, H. J.; Maken, S.; Park, J. W. Kinetics of CO_2 Absorption in Aqueous Sodium Glycinate Solutions. *Ind. Eng. Chem. Res.* **2007**, *46*, 1578–1583.

(15) Lee, S.; Song, H. J.; Maken, S.; Yoo, S. K.; Park, J. W.; Kim, S.; Shim, J. G.; Jang, K. R. Simulation of CO_2 Removal with Aqueous Sodium Glycinate Solutions in a Pilot Plant. *Korean J. Chem. Eng.* **2008**, *25*, 1–6.

(16) Song, H. J.; Lee, S.; Maken, S.; Park, J. J.; Park, J. W. Solubilities of Carbon Dioxide in Aqueous Solutions of Sodium Glycinate. *Fluid Phase Equilib.* **2006**, *246*, 1–5.

(17) Song, H. J.; Lee, S.; Park, K.; Lee, J.; Spah, D. C.; Park, J. W.; Filburn, T. P. Simplified Estimation of Regeneration Energy of 30 wt % Sodium Glycinate Solution for Carbon Dioxide Absorption. *Ind. Eng. Chem. Res.* **2008**, *47*, 9925–9930.

(18) Salazar, V.; Sánchez-Vicente, Y.; Pando, C.; Renuncio, J. A. R.; Cabañas, A. Enthalpies of Absorption of Carbon Dioxide in Aqueous Sodium Glycinate Solutions at Temperatures of (313.15 and 323.15) K. *J. Chem. Eng. Data* **2010**, *55*, 1215–1218.

(19) Harris, F.; Kurnia, K. A.; Mutalib, M. I. A.; Thanapalan, M. Solubilities of Carbon Dioxide and Densities of Aqueous Sodium Glycinate Solutions before and after CO_2 Absorption. *J. Chem. Eng. Data* **2009**, *54*, 144–147.

(20) Kumar, P. S.; Hogendoorn, J. A.; Feron, P. H. M.; Versteeg, G. F. Density, Viscosity, Solubility, and Diffusivity of N_2O in Aqueous Amino Acid Salt Solutions. *J. Chem. Eng. Data* **2001**, *46*, 1357–1361.

(21) Kumar, P. S.; Hogendoorn, J. A.; Feron, P. H. M.; Versteeg, G. F. Equilibrium Solubility of CO_2 in Aqueous Potassium Taurate Solutions: Part 1. Crystallization in Carbon Dioxide Loaded Aqueous Salt Solutions of Amino Acids. *Ind. Eng. Chem. Res.* **2003**, *42*, 2832–2840.

(22) Kumar, P. S.; Hogendoorn, J. A.; Timmer, S. J.; Feron, P. H. M.; Versteeg, G. F. Equilibrium Solubility of CO_2 in Aqueous Potassium Taurate Solutions: Part 2. Experimental VLE Data and Model. *Ind. Eng. Chem. Res.* **2003**, *42*, 2841–2852.

(23) Kumar, P. S.; Hogendoorn, J. A.; Versteeg, G. F.; Feron, P. H. M. Kinetics of the Reaction of CO_2 with Aqueous Potassium Salt of Taurine and Glycine. *AIChE J.* **2003**, *49*, 203–213.

(24) Ahn, S.; Song, H. J.; Park, J. W.; Lee, J. H.; Lee, I. Y.; Jang, K. R. Characterization of Metal Corrosion by Aqueous Amino Acid Salts for the Capture of CO_2. *Korean J. Chem. Eng.* **2010**, *27*, 1576–1580.

(25) Van Holst, J.; Versteeg, G. F.; Brilman, D. W. F.; Hogendoorn, J. A. Kinetic Study of CO_2 with Various Amino Acid Salts in Aqueous Solution. *Chem. Eng. Sci.* **2009**, *64*, 59–68.

(26) Muñoz, D. M.; Portugal, A. F.; Lozano, A. E.; De La Campa, J. G.; De Abajo, J. New Liquid Absorbents for the Removal of CO_2 from Gas Mixtures. *Energy Environ. Sci.* **2009**, *2*, 883–891.

(27) Zhang, J.; Zhang, S.; Dong, K.; Zhang, Y.; Shen, Y.; Lv, X. Supported Absorption of CO_2 by Tetrabutylphosphonium Amino Acid Ionic Liquids. *Chem.—Eur. J.* **2006**, *12*, 4021–4026.

(28) Reid, R. C.; Prausnitz, J. M.; Poling, B. E. *The Properties of Gases & Liquids*, 4th ed.; McGraw-Hill: New York, 1987.

(29) Lide, D. R. *Handbook of Chemistry and Physics*, 88th ed.; CRC Press: Boca Raton, FL, 2008.

(30) Chiu, L. F.; Liu, H. F.; Li, M. H. Heat Capacity of Alkanolamines by Differential Scanning Calorimetry. *J. Chem. Eng. Data* **1999**, *44*, 631–636.

(31) Ditmars, D. A.; Bernstein, G.; Chang, S. S.; Ishihara, S.; West, E. D. Enthalpy and Heat-Capacity Standard Reference Material-Synthetic Sapphire (α-Al_2O_3) from 10 to 2250 K. *J. Res. Nat. Bur. Stand* **1982**, *87*, 159–163.

(32) Jones, J. H.; Froning, H. R.; Claytor, E. E., Jr. Solubility of Acidic Gases in Aqueous Monoethanolamine. *J. Chem. Eng. Data* **1959**, *4*, 85–92.

(33) Lee, J. I.; Otto, F. D.; Mather, A. E. The Solubility of H_2S and CO_2 in Aqueous Monoethanolamine Solutions. *Can. J. Chem. Eng.* **1974**, *52*, 803–805.

(34) Murrieta-Guevara, F.; Rodríguez, A. T. Liquid Density as a Function of Temperature of Five Organic Solvents. *J. Chem. Eng. Data* **1984**, *29*, 204–206.

(35) Mandal, B. P.; Kundu, M.; Bandyopadhyay, S. S. Density and Viscosity of Aqueous Solutions of (N-Methyldiethanolamine + Monoethanolamine), (N-Methyldiethanolamine + Diethanolamine), (2-Amino-2-methyl-1-propanol + Monoethanolamine), and (2-Amino-2-methyl-1-propanol + Diethanolamine). *J. Chem. Eng. Data* **2003**, *48*, 703–707.

(36) Al-Ghawas, H. A.; Hagewiesche, D. P.; Ruiz-Ibanez, G.; Sandall, O. C. Physicochemical Properties for Carbon Dioxide Absorption in Aqueous Methyldiethanolamine. *J. Chem. Eng. Data* **1989**, *34*, 385–391.

(37) Vázquez, G.; Alvarez, E.; Navaza, J. M.; Rendo, R.; Romero, E. Surface Tension of Binary Mixtures of Water + Monoethanolamine and Water + 2-Amino-2-methyl-1-propanol and Tertiary Mixtures of These Amines with Water from 25 to 50 °C. *J. Chem. Eng. Data* **1997**, *42*, 57–59.

(38) Alvarez, E.; Rendo, R.; Sanjurjo, B.; Sánchez-Vilas, M.; Navaza, J. M. Surface Tension of Binary Mixtures of Water + *N*-Methyldiethanolamine and Tertiary Mixtures of This Amine and Water with Monoethanolamine, Diethanolamine, and 2-Amino-2-methyl-1-propanol from 25 to 50 °C. *J. Chem. Eng. Data* **1998**, *43*, 1027–1029.