Kinetics of the Reaction of Pure CO\textsubscript{2} with N-Methyldiethanolamine in Aqueous Solutions

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ABSTRACT: The process of pure carbon dioxide absorption is analyzed in aqueous solutions of N-methyldiethanolamine (MDEA). The experiments were made in a stirred tank reactor with a plane and known interface area. The variables considered were the MDEA concentration within the range 0.1–3.0 M and the temperature in the interval 288–313 K. From the results, we deduce that the process takes place under isothermal conditions and moderately fast regime, with second-order kinetics. We determined a reaction order of one with respect to the amine, and an expression for the kinetic constant valid throughout the entire range of temperatures and concentrations assayed \( \ln k = 22.4 - 6243.5/T \). © 2008 Wiley Periodicals, Inc. Int J Chem Kinet 41: 204–214, 2009

INTRODUCTION

The use of tertiary alkanolamines in the carbon dioxide absorption processes is more and more frequent because of the advantages that they present for certain applications (i.e., CO\textsubscript{2} removal from natural gas at high pressure) as opposed to the most widely used primary and secondary ones (monoethanolamine (MEA) or diethanolamine (DEA)). That is, the latter, apart from having a high cost also presents other drawbacks that should be taken into account, such as the high energy consumption for the regeneration of the absorbent (for MEA 900 kcal/kg CO\textsubscript{2} [1]), the foam production, corrosion of materials, quick evaporation, and rapid degradation in the presence of oxygen. All these factors translate to high cost, while tertiary amines reduce these problems despite the slower reaction with carbon dioxide [2,3].

The most widely studied tertiary alkanolamines are triethanolamine (TEA) and N-methyldiethanolamine (MDEA), all research groups coinciding with regard to the order of reaction, one with respect to the carbon dioxide and two for the overall reaction, both in TEA as well as in MDEA.

Many laboratory studies have been performed on the CO\textsubscript{2} absorption process using MDEA and water as solvent, but there are still some discrepancies in the literature as to the interpretation of the kinetic data [4]. For instance, the second-order rate constant for the MDEA–CO\textsubscript{2} at 20°C ranges from 1.35 [5] to
4.7 m⁶/kmol² s [6]. Some of the discrepancies may come from the different absorber apparatus or from the inconsistency of the physical data such as solubility and diffusivity of CO₂ in MDEA aqueous solutions [7]. Besides, there is no consensus on the effect that the reaction of hydroxide with carbon dioxide exerts on the absorption rate of carbon dioxide in aqueous MDEA solutions and, more importantly, on the estimation of the kinetic rate coefficient of the reaction between CO₂ and MDEA. Littel et al. [8] claim that the effect of this reaction is negligible, whereas others [9,10] consider this reaction important, treating it as a pseudo-first-order irreversible reaction.

On an industrial scale, MDEA is an alkanolamine of great interest in acidic-gas absorption and thus presents less corrosive effects than the primary or secondary alkanolamines [11,12]. Furthermore, because of the instantaneous reaction rate of the H₂S with MDEA and the very low rate for the reaction of CO₂ with MDEA, the aqueous solutions of this alkanolamine are frequently used as selective absorbents of the H₂S on the CO₂ [13], when these are found present in a gas current and their separation is sought for later use of the H₂S, or when the CO₂ is not necessary or economically feasible. In this sense, Pacheco and Rochelle [14] report different degrees of selectivity of the process, depending on the type of contactor used.

On the other hand, MDEA has some advantages compared to other alkanolamines used as absorbents such as monoethanolamine (MEA), because MDEA presents higher equilibrium capacity conversion (1 mol of CO₂/1 mol of MDEA) than does MEA, which is limited by the stoichiometry (0.5 mol of CO₂/1 mol of MEA) [15].

Similarly, due to the reactive characteristics, the CO₂–MDEA process takes place with a lower reaction enthalpy than for aqueous solutions of primary or secondary alkanolamines, and this is causing its use to increase [16]. Furthermore, they require less regeneration energy and are resistant to the thermal and chemical degradation [15,17]. Such aspects as those cited above have caused much research to be devoted to tertiary alkanolamines.

In this work, we analyze the absorption process of pure carbon dioxide in aqueous solutions of MDEA and the kinetic parameters of the process.

**MATERIALS AND METHODS**

**Experimental Device and Procedure**

The carbon dioxide used in this work was CO₂ N-48 (99.998%; Air Liquid, Madrid, Spain), with oxygen, water steam, and hydrocarbons as impurities; MDEA was an Acros Organics (Geel, Belgium) product with nominal purity of 99%. Aqueous solutions of alkanolamine were prepared with distilled-deionized water (resistivity 18.2 MΩ-cm).

The experiments were made in a stirred tank reactor at a stirring rate of 80 ± 1 rpm, batch with respect to the gas–liquid phases, and with a flat and known interface area. The characteristics of the reactor are internal height 20.00 cm, external height 21.50 cm, external diameter 9.50 cm, internal diameter 6.70 cm, interface area 35.26 cm², and volume 500.00 mL.

The experimental device used is described in Fig. 1. The total pressure, P, was measured in a previously
calibrated barometer placed close to the experimental installation [18]. The soap flowmeter was composed of a burette of 50-cm³ capacity with a lateral exit in its lower part, connected by a hose to the mouth of the reactor. To form the bubbles, the lower part of the column contains a rubber deposit with water and soap solution. Volumes of 100 mL of the aqueous solutions of MDEA were used in the range 0.1–3.0 M, and the temperature was set within the interval 288–313 K.

**Physical and Transport Properties**

Under these experimental conditions, we measured the viscosity of the amine solutions and their density [19].

The calculation of the initial partial pressure of the CO₂ is given as follows:

\[ p_A = P - p_v \]  

where \( P \) is the total pressure and \( p_v \) is the vapor pressure of the water.

The solubility of CO₂ in the liquid phase was measured by Al-Ghawas et al. [20] using the \( \text{N}_2\text{O} \) analogy method. Data were collected over the temperature range 288–328 K and for MDEA concentrations up to 8 M. The data can be fitted to the equation of Al-Ghawas et al. [20],

\[ \ln H_e = B + \frac{C}{T} + \frac{D}{T^2} \]  

where

\[ B = 2.01874 - 2.37638 \times 10^4 w + 2.900992 \times 10^2 w^2 - 4.80196 \times 10^7 w^3 \]

\[ C = 3.13549 \times 10^2 - 1.54931 \times 10^4 w - 1.83987 \times 10^5 w^2 + 3.00562 \times 10^5 w^3 \]

\[ D = -8.13702 \times 10^5 - 2.48081 \times 10^4 w - 2.92013 \times 10^7 w^2 - 4.70852 \times 10^7 w^3 \]

and \( w \) is the amine mass fraction in the solution.

The diffusion coefficients of CO₂ in the aqueous solution were measured using the relationships of Sada et al. [21] and Versteeg and van Swaaij [22].

\[ \frac{D_{\text{CO}_2}}{D_{\text{CO}_2, w}} = \frac{D_{\text{N}_2\text{O}}}{D_{\text{N}_2\text{O}, w}} \]  

where

\[ D_{\text{CO}_2, w} = 2.35 \times 10^{-6} e^{-\frac{2119}{T}} \]  

\[ D_{\text{N}_2\text{O}, w} = 5.07 \times 10^{-6} e^{-\frac{2371}{T}} \]

\[ D_{\text{N}_2\text{O}} \mu_{\text{w}} = D_{\text{N}_2\text{O}, w} \mu_{\text{w}} \]

where \( \gamma \) = 0.8 [9].

The diffusion coefficients of MDEA in aqueous solution were calculated by means of the relationship of Wilke and Chang [23],

\[ D_B = 2.31 \times 10^{-15} \frac{T}{\mu_B} \]

**Analytical Methods**

Given that it is necessary to determine the concentration of the amine available at the onset of the carbon dioxide absorption, the analyses programmed for all the samples of the alkanolamine solutions were performed to know the exact concentration of amine available. The analyses were made in duplicate at time zero once the system was established.

The initial concentration of amine was determined by titration with HCl solutions using methyl orange as the indicator. The CO₂ concentration in the liquid was determined by standard titration methods. Excess NaOH and the excess BaCl₂ solutions were added to the liquid sample, and the excess NaOH was titrated with the HCl solution using phenolphthalein as the indicator.

**Figure 2** CO₂ absorbed per unit of surface and time in the experiments with MDEA at 313 K.
RESULTS AND DISCUSSION

The flow densities, \( N_A \), were calculated assuming that the gas follows ideal behavior, using the following expression:

\[
N_A = \frac{n'}{A} = \frac{PQ'}{RTA}
\]  

(8)

The value of the volumetric flow coincided with the value of the slope of the straight lines on representing the volume of the \( \text{CO}_2 \) absorbed against time (\( t \)) for each experiment. The determinations were made using the linear regression method of the experimental results. As an example, Fig. 2 offers a graphic representation of the \( \text{CO}_2 \) absorbed per unit of surface area against time for the experiments conducted at 313 K. A constant operating temperature and partial gas pressure were maintained, varying the initial MDEA concentration between 0.11 and 2.78 M.

In the absorption experiments with pure \( \text{CO}_2 \) in aqueous MDEA solutions, we analyzed the influence of the initial alkanolamine concentration and the operating temperature.

Absorption in MDEA Solutions

MDEA, \( \text{H}_3\text{C}-\text{N}-(\text{CH}_2-\text{CH}_2-\text{OH})_2 \), is a tertiary alkanolamine in which the nitrogen is bonded to two ethanol groups and a methyl group.

Tertiary alkanolamines are differentiated from the primary and secondary ones in their reaction with \( \text{CO}_2 \), resulting in a slower reaction compared with that of the other amine types, the carbamate formation not being possible.

It has been amply accepted that the reaction order of the chemical absorption process of carbon dioxide in aqueous solutions of MDEA is one with respect to the carbon dioxide and the alkanolamine, giving an overall reaction order of two. Donaldson and Nguyen [24] proposed that the reaction might be described with a kind of base catalysis of the \( \text{CO}_2 \) hydration. This catalytic effect is based on the formation of a hydrogen bond between the amine and water. This weakens the bond between hydroxyl group (\( -\text{OH} \)) and hydrogen, and then increases the water nucleophilic reactivity toward carbon dioxide. The mechanism followed is given as

\[
\text{R}_3\text{N} + \text{H}_2\text{O} \leftrightarrow \text{R}_3\text{NH} + \text{HOH}^-
\]

Savage et al. [25] consider the process to take place into two stages:

\[
\text{H}_2\text{O} + \text{R}_3\text{N} \rightleftharpoons \text{OH}^- + \text{R}_3\text{NH}^+ \text{(instantaneous)}
\]

(10)

\[
\text{OH}^- + \text{CO}_2 \rightleftharpoons \text{HCO}_3^- \text{(slow)}
\]

(11)

Figure 3 (a) Variation in pH of the MDEA solution with time, over the absorption process at 298 K at corresponding concentrations. (b) Variation in pH of the MDEA solution measured at the same time with the amine concentration at different temperatures.
The overall mechanism also included the reactions of the carbonic acid formation (12) and the bicarbonate ion formation (13),

\[
\begin{align*}
\text{CO}_2 + \text{H}_2\text{O} & \rightleftharpoons \text{H}_2\text{CO}_3 \quad (12) \\
\text{CO}_2 + \text{OH}^- & \rightleftharpoons \text{HCO}_3^- \quad (13)
\end{align*}
\]

The relative importance of these reactions depends on their velocity constants and the amine concentrations. In this mechanism, the CO\(_2\) hydration reaction (12) is slow, with a \(k_w = 0.0265 \text{ s}^{-1}\) at 298 K \[26\], and, according to Blauwhoff et al. \[27\], usually negligible. However, this research group considers the reaction forming the bicarbonate ion (13) to be the one that most contributes to the process \[28\], according to a direct reaction of the CO\(_2\) with water, followed by a rapid dissociation of carbonic acid (14),

\[
\text{CO}_2 + \text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{CO}_3 \rightleftharpoons \text{H}^+ + \text{HCO}_3^- \quad (14)
\]

exerting the CO\(_2\) reaction with the hydroxyl group, the dominant role when pH values exceed 8.0 \[29\].

Given the possibility that in the absorption process a monoalkylcarbonate could form (15), as suggested by Jørgensen and Faurholt \[30\],

\[
\text{CO}_2 + \text{OH}^- + \text{R}_2\text{NCH}_2\text{OH} \quad \rightleftharpoons \text{R}_2\text{NCH}_2\text{OCOO}^- + \text{H}_2\text{O} \quad (15)
\]

Some research groups \[28,31\] hold that this reaction contributes negligibly to the overall velocity for pH values lower than 12 and occurs only at extremely high pH values, as reflected by the study of Jørgensen and Faurholt \[30\].

Haimour et al. \[32\], although contending that the mechanism between the carbon dioxide and the tertiary alkanolamines in aqueous solution is not completely understood, for the case of CO\(_2\) absorption in aqueous solutions of MDEA, indicate that at the stoichiometric level the process that takes place is described in reaction (16).

\[
\text{CO}_2 + \text{H}_2\text{O} + \text{R}_3\text{N} \rightleftharpoons \text{R}_3\text{NH}^+ + \text{HCO}_3^- \quad (16)
\]

Consequently, the extent of the mechanism proposed by Donaldson and Nguyen \[24\] for aqueous methyl diethanolamine solutions is assumed by different research groups that have studied this system, as agrees with the results by several of them \[5,28,33\].

From all the above, it can be deduced that the tertiary amines do not react directly with CO\(_2\), and consequently MDEA does not form carbamate ions. Recent studies agree with the above commentary (referring to the reactions (12,13)), considering the reaction of bicarbonate formation to be fast and capable of augmenting the mass transfer, even when the concentration of hydroxyl is low and can contribute significantly to the observed absorption rate \[34\].

As a result, for the process between CO\(_2\) and MDEA, the reaction rate would be given by

\[
r = k_{2\text{MDEA}[\text{CO}_2]}[\text{MDEA}] + k_{\text{OH}^-[\text{CO}_2]}[\text{OH}^-] = k[\text{CO}_2] \quad (17)
\]

\[
k = k_{2\text{MDEA}[\text{MDEA}]} + k_{\text{OH}^-[\text{OH}^-]}[\text{OH}^-] \quad (18)
\]

\[
k_{\text{app}} = k - k_{\text{OH}^-[\text{OH}^-]} = k_{2\text{MDEA}[\text{MDEA}]} \quad (19)
\]

Finally, although there is unanimity regarding the mechanism of the reaction, there are nevertheless many discrepancies in the literature on the interpretation of the kinetic results, which, in the opinion of Ko and Li \[7\], may come from both the different absorbers used as well as from the inconsistency of the physical data (solubility and CO\(_2\)-diffusion coefficient in the aqueous MDEA solutions).

To analyze the behavior of the pH over the CO\(_2\)-absorption process in aqueous MDEA solutions, we made experiments for which the results at 298 K are shown as an example in Fig. 3, reflecting a
decline in pH over the absorption process at a certain concentration.

**Reaction Regime**

For the determination of the reaction regime, representations were made of values of flux density, \( N_A \), against \( C_{Bo} \), as shown in Fig. 4. No linear relationship was appreciated between the flux density and the amine concentration. This could indicate a noninstantaneous reaction regime for all the concentration and temperature intervals.

In addition, the results from different research groups that have worked with methyldiethanolamine and carbon dioxide under different conditions and contactors [5–10,31,35] express no doubts concerning whether the absorption process takes place in the fast-reaction regime.

**Table I** Flux Densities in the Absorption of CO\(_2\) in N-Methyldiethanolamine Solutions

<table>
<thead>
<tr>
<th>( T ) (K)</th>
<th>( C_{Bo} ) (kmol/m(^3))</th>
<th>( p_A ) (kPa)</th>
<th>( N_A ) ( (\times 10^6 ) kmol/m(^2) s)</th>
<th>( H_e ) (kPa m(^3)/kmol)</th>
<th>( D_A ) ( 10^9 ) (m(^2)/s)</th>
<th>( \log \left( \frac{N_A^2 H_e}{p_A D_A} \right) )</th>
<th>( \log C_{Bo} )</th>
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<td>1.54</td>
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</table>
Figure 5  Study of the reaction regime for MDEA at different temperatures.

To test in this study whether the process takes place in a fast-reaction regime, we assume as our initial hypothesis the fast-reaction regime, in which order $m$ is one with respect to the CO$_2$ and $n$ with respect to the MDEA. The molar flux is given by Eq. (20) [36],

$$N_A = \sqrt{\frac{2}{m+1}} D_A k_{m,n} C_A^{m+1} C_B^n$$  \hspace{1cm} (20)

Assuming that the order of the reaction is one with respect to CO$_2$ in the absorption with aqueous solutions of MDEA, then Eq. (20) is reduced to

$$N_A = C_A e \sqrt{D_A k_{1,n} C_B^n}$$  \hspace{1cm} (21)

if we consider that $C_A^*$ is the CO$_2$ concentration in equilibrium with the gaseous phase. This can be evaluated by Henry’s law ($p_A = H e C_A^*$), with the result

$$\log \left( \frac{N_A^2 H e^2}{p_A^2 D_A} \right) = \log k_{1,n} + n \log C_B$$  \hspace{1cm} (22)

The results of $\log \left( \frac{N_A^2 H e^2}{(D_A p_A^2)} \right)$ and $C_B$ are listed in Table I.

These representations (see Fig. 5) reflect an acceptable linear relationship for all temperatures and concentrations assayed, indicating that we would be operating in a fast-reaction regime, which is consistent with the literature.

**Thermal Effects and Kinetic Parameters**

The control of the temperature in the bulk liquid phase, in the experiments of the greatest MDEA concentration, indicates that there are no differences with respect to the operating temperature, so we can consider the operating conditions to be isothermal.

Representations were made of the first member of Eq. (22) against $\log C_B$ (see Fig. 5), and, applying the least-squares method, we calculated the slope ($n$) and $y$-intercept (0.0 $\equiv \log k_{1,n}$), which determined a kinetic constant (see Table II).

The values of the slope ranged from 0.82 to 0.97 at the temperatures of 303 and 288 K, respectively; these values being close to those reported by Ko and Li [7] that give values of 0.91, 0.94, and 0.88 at 303, 308, and 313 K, respectively. Thus, the reaction can be considered a pseudo-first-order reaction, according to different authors [7,9,10,27,35].

The values of the kinetic constants determined were of the same order of magnitude as those reported by other authors, although there are some discrepancies,
such as the second-order constant at 298 K ranging from 1.35 [5] to 4.70 [6].

Finally, from the values of Table II, we fit the value of \( k \) and \( T \) to an Arrhenius-type equation. Thus, in Fig. 6, \( \ln k \) versus \( 1/T \) are represented, the application of the least-squares method leads to an acceptable fit, enabling the deduction of the following expression, similar to those found in the literature (see Table III):

\[
\ln k_2 = 22.4 - \frac{6243.5}{T}
\]  

(23)

Considering the equation for the kinetic constant proposed (Eq. (23)) and the same found in the literature, we made the representation (Fig. 6) in which the agreement of our results with previous works can be appreciated, including certain values for the kinetic constant determined at particular temperatures.

From Eq. (23), we can deduce the apparent activation energy \((E_A)_a\) of 51.9 kJ/mol, value which is close to those cited by most of research groups (see Table III). Also, from the \( y \)-intercept the frequency factor can be determined, giving a value of \( 5.34 \times 10^9 \text{ m}^3/\text{kmol s} \).

## Enhancement Factors and the Hatta Number

Assuming valid the expression (23) for the kinetic constant, the Hatta number can be calculated by Eq. (24):

\[
Ha = \sqrt{\frac{k_2 \beta Bo}{2k_L}}
\]  

(24)

where the liquid-phase mass-transfer coefficients \( (k_L) \) were obtained in the previous work [39] at the same operating conditions and were related, by the modified Sherwood number, to the numbers of Schmidt and Reynolds [18].

Thus, the instantaneous enhancement factor, \( E_i \), in agreement with the film theory, can be evaluated at each temperature,

\[
E_i = 1 + \frac{D_B}{z D_A} \frac{C_{Bo} He}{p_A}
\]  

(25)

where \( z = 1 \) according to the reaction stoichiometry.

Once \( E_i \) and \( Ha \) are known, the enhancement factor, \( E \), can be determined from the DeCoursey expression [40],

\[
E = \frac{Ha^2}{2(E_i - 1)} + \frac{Ha^4}{4(E_i - 1)^2} + \frac{E_i Ha^2}{E_i - 1} + 1
\]  

(26)

Table IV lists the values of \( E \), \( E_i \), and \( Ha \), corresponding to the entire interval of temperatures and concentrations assayed.

The analysis of these results confirms that the reaction regime is not instantaneous because the following condition is satisfied:

\[
E_i \gg E
\]
Table IV  Enhancement Factors and the Hatta Number in the MDEA Experiments

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<tr>
<th>$T$(K)</th>
<th>$C_{Bo}$(kmol/m$^3$)</th>
<th>$E_i$</th>
<th>$Ha$</th>
<th>$E$</th>
<th>$Ha + 1$</th>
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</table>

Nevertheless not fulfilled, according to which $E = Ha$. However, except in the experiments made at concentrations higher than 1 M at 303 K and 0.5 M from 308 to 313 K, it is true that

$$0.3 < Ha < 3.0$$

indicating that the process occurs in the moderately fast-reaction regime [36], although in the experiments mentioned above, $Ha > 3$, it could be accepted that the absorption occurs in the fast-reaction pseudo-first-order regime as the following condition is satisfied [41]:

$$3 < Ha < E_i/2$$
Similarly, we confirmed that, in most cases, the relationship between the enhancement factor and the Hatta number is \( E \approx Ha + 1 \), a fact that, for \( Ha \) values between the above-cited limits, it would agree with a moderately fast-reaction regime.

All of the above reinforces the initial hypothesis, according to which the absorption of pure \( \text{CO}_2 \) by aqueous solutions of MDEA takes place in the fast-reaction regime, in which a moderate reaction in the film occurs and fast in the bulk liquid phase interior of the solution [36]. Finally, the representation of logarithmic coordinates of \( E \) versus \( Ha \) (Fig. 7) confirms the fast-reaction regime assumed.

**CONCLUSION**

During the absorption of \( \text{CO}_2 \) in aqueous solutions of MDEA, the control of the temperature in the bulk liquid phase indicates that there are no differences with respect to the operating temperature, so we can consider the operating conditions to be isothermal. We deduce that the process takes place under moderately fast regime, with second-order kinetics. Finally, we fit the value of the kinetic constant and temperature to an Arrhenius-type equation. Considering this expression, we can deduce the apparent activation energy of 51.9 kJ/mol.

**NOMENCLATURE**

- \( A \) Interfacial area (m\(^2\))
- \( B \) Correlation parameter in Eq. (2)
- \( C \) Correlation parameter in Eq. (2)
- \( C_{Ao} \) Initial concentration of component A (\( \text{CO}_2 \)) (kmol/m\(^3\))
- \( C_A^* \) Concentration component A (\( \text{CO}_2 \)) in equilibrium with the gaseous phase (kmol/m\(^3\))
- \( C_{Bo} \) Initial concentration of amine in the aqueous phase (kmol/m\(^3\))
- \( D \) Correlation parameter in Eq. (2)
- \( D_{A(D_{CO}_2)} \) Diffusion coefficient of component A (\( \text{CO}_2 \)) in the aqueous alkanolamine solution (m\(^2\)/s)
- \( D_B \) Diffusion coefficient of alkanolamine in the liquid phase (m\(^2\)/s)
- \( D_{CO}_2,w \) Diffusion coefficient of \( \text{CO}_2 \) in water (m\(^2\)/s)
- \( D_{N_2,O} \) Diffusion coefficient of \( \text{N}_2\text{O} \) in the alkanolamine solution (m\(^2\)/s)
- \( D_{N_2,O,w} \) Diffusion coefficient of \( \text{N}_2\text{O} \) in water (m\(^2\)/s)
- \( E \) Enhancement factor
- \( (E_A)_{ap} \) Apparent activation energy (kJ/mol)
- \( E_i \) Instantaneous-enhancement factor
- \( Ha \) Hatta number
- \( He \) Henry’s law constant (kPa m\(^3\)/kmol)
- \( k \) Reaction rate constant
- \( k_L \) Liquid-phase mass-transfer coefficient (m/s)
- \( m \) Order of reaction with respect to \( \text{CO}_2 \)
- \( n \) Order of reaction with respect to amine
- \( n' \) Absorption rate of \( \text{CO}_2 \) (kmol/s)
- \( N_A \) Absorption rate per unit interfacial area of component A (\( \text{CO}_2 \)) (kmol/m\(^2\)/s)
- \( P \) Total pressure (kPa)
- \( P_{A} \) Partial pressure of component A (\( \text{CO}_2 \)) (kPa)
- \( p_v \) Vapor pressure of the water (kPa)
- \( Q' \) Volumetric flow rate of absorbed \( \text{CO}_2 \)(m\(^3\)/s)
- \( r^2 \) Linear regression coefficient.
- \( R \) Gas constant (kPa m\(^3\)/K kmol)
- \( t \) Time (s)
- \( T \) Temperature (K)
- \( w \) Amine mass fraction in solution
- \( z \) Stoichiometric coefficient
- \( \gamma \) Constant defined in Eq. (6)
- \( \mu_B \) Viscosity of the solution amine (Pa s)
- \( \mu_w \) Viscosity of the pure water (Pa s)

**BIBLIOGRAPHY**