

Experimental Investigation of the Kinetics and Thermodynamics of the Production of Methanol from the Hydrolysis of Methyl Acetate Catalyzed with Hydrochloric Acid

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Abstract

The production of methanol from different sources as alternate biofuel as well as gasoline-blend fuel component will help in mitigating the adverse effect of using gasoline from hydrocarbon as fuel for internal combustion engines. Kinetics and thermodynamics studies were carried out in this study using the process of hydrolysis of methyl acetate catalyzed with hydrochloric acid to produce methanol. The method employed was backward-titration which was used to determine the concentration of methyl acetate catalyzed with hydrochloric acid in a batch reactor. The titration was performed using a known volume (5 ml) of the reaction mixture at different time intervals (10 mins – 50 mins at intervals of 10 mins) against sodium hydroxide. The results obtained were fitted into first, second and n^{th} order kinetic models. The analysis revealed that the hydrolysis followed first order kinetic model initially, but with increases in temperature, the second order kinetic model became evident and gave the best fit for the hydrolysis at the optimum temperature of 40 °C. The n^{th} order kinetic model showed that at the investigated temperatures, the order of the hydrolysis of methyl acetate was < 1 . The values of the changes in enthalpy (2.04 KJ/mole), entropy (-1.52 J/Mol.K) and Gibbs free energy (2.50 KJ/mole) obtained showed that the hydrolysis was endothermic; the degree of disorderliness was negative (-1.52J/Mol.K); and the reaction process was non-spontaneous. These studies revealed that the kinetics and thermodynamics behaviors of the hydrolysis of methyl acetate catalyzed with hydrochloric acid would be a good source for the production of methanol.

Key words: Hydrochloric acid; kinetic; methanol; methyl acetate; thermodynamic

1.0 Introduction

Methanol is a strategic bulk chemical in the chemical industry. It has a very high global demand which is estimated to be over 32 million metric tons and arising [1, 2]. It is used mainly for the production of formaldehyde, acetic acid and application products such as polymers and paints. It is similarly used in most chemical industries as solvent and reactant and finds application in the production of many household products such as paints, varnishes, and cleaning products. It can also be used as a clean and renewable energy carrier produced mainly from syngas; a mixture of H₂, CO and minor quantities of CO₂ and CH₄. Green methanol has environmental benefits and leads to considerable cost reductions as long as the biomass resource has a low or even negative value. It is the first and simplest form of alcohol with CH₃OH as its molecular formula [1, 3]. Similarly, methanol can also be used as alternative biofuel for internal combustion engines or as motor fuel or gasoline blending component [1, 4]. Methanol-fueled vehicles make use of a blend of 85% methanol with 15% unleaded gasoline (M85). Laboratory and road- tests indicate that addition of 10% methanol can raise octane number from 2 to 3. Conventionally, methanol is produced from petroleum product (synthesis gas) through

hydrogenation of CO and CO₂, and reversed water-gas shift reaction. This synthesis gas process uses CuO/ZnO/Al₂O₃ as catalyst. The commercial CuO/ZnO/Al₂O₃ together with a zeolite membrane reactor can provide higher CO₂ conversion and methanol yield when selectively compared with a traditional reactor [1, 5].

Hydrolysis of methyl acetate in the production of methanol is a major reaction in the synthesis of polyvinyl alcohol (PVA) and other specialty chemicals [2, 3]. Methyl acetate can readily be hydrolyzed to produce very useful products which can serve as value-added products such as methanol and acetic acid. The industrial and laboratory applications of these products are appreciable in reactive distillation researches [6 - 11]. As a result of the nature of the hydrolysis reaction, the rate could be increased by employing H⁺ ions as the requisite catalyst [6, 8 -10]. Experimental researches have been conducted on reactive distillation [4, 8 - 10]. The studies were on experimental synthesis and hydrolysis of methyl acetate using a structured catalytic packing (Katapak-S) with an acidic ion-exchange resin (Amberlyst 15) as the heterogeneous catalyst. Three different set-ups were used: one-feed and two-feed setups for the methyl acetate synthesis and another two-feed setup for the hydrolysis reaction. The influences of several variables, such as reaction kinetics, separation efficiency, residence time distribution, and heat loss, on the simulation results were analyzed. An investigation was conducted [12] on the hydrolysis of methyl acetate catalyzed by Amberlyst 15 in terms of the reaction kinetics and chemical equilibrium of the reversible catalytic hydrolysis reaction. The effects of the process parameters in terms of the temperature, catalyst weight and feed molar ratios on the reaction kinetics were also investigated. The results revealed a reaction rate that was strongly-temperature dependent. Similarly, the authors catalyzed the reaction homogeneously by acetic acid itself and heterogeneously by an acidic ion-exchange resin (Amberlyst 15). They measured the chemical equilibrium composition for various temperatures and starting compositions of the reactants and products. They obtained kinetic information at temperatures between 303.15 K and 343.15 K at various starting compositions covering concentration ranges from the stoichiometric regime to the dilute regions. Similarly, other authors [6] carried-out a study on methanol production pathways via biomass gasification and CO₂ recovery from the flue gasses of a fossil fuel-fired power station. They found out that the cost of methanol production from biomass lie in the range of 300–400 €/ton of methanol, and the production cost of CO₂ - based methanol was found to be between 500 and 600 €/ton. Also, another author [13] carried-out an investigation on evaluation of the reaction kinetics and thermodynamics and highlighted the interest of CO₂ conversion into methanol through a complete techno-economic and environmental assessment of the entire process chain. They treated the CO₂ coming from a conventional cement plant with a MEA-based CO₂ capture process, and the captured CO₂ was directly sent to the conversion unit for its catalytic conversion. They combined the two units which led to relevant integrations, especially regarding the reuse of the heat provided by the exothermal methanol reactions for the regeneration of the CO₂ capture solvent. Likewise, another researcher [14] proposed a model for the production and emission of methanol to the atmosphere by flowering plants based on plant structure and metabolic properties. They reported that their model provided a framework to extend seven sets of measurements of methanol emission rates to the global terrestrial biosphere. They calculated a global rate of release of methanol from plants to the atmosphere as 100 Tg/y and also reported that their model successfully reproduces the methanol concentrations in the continental boundary-layer and the free atmosphere, including the inter-hemispheric gradient in the free atmosphere. Yu *et al* [9] presented methyl acetate hydrolysis and esterification reaction rates. They implemented a packed bed column as a fixed bed reactor instead of a batch system with a continuous experimental situation. The reaction rates obtained were only valid for the experimental conditions applied.

Chemical kinetic is globally acknowledged as the rate at which chemical reactions occur [15, 16]. The hydrolysis of methyl acetate catalyzed with hydrochloric acid has been established at low temperatures of 15 °C 20 °C and 25 °C to follow the first order kinetic model [17]. But there were no record to show the actual rate of the reaction at higher temperatures, and the hydrolysis at each temperature involving the emission of energy, which includes but not limited to the activation energy. However, other thermodynamic parameters such as entropy, enthalpy and the Gibbs free energy were evaluated. Several researchers [18 – 20] have also worked on kinetics and equilibrium studies of the hydrolysis of methyl acetate with acetic acid but none has considered the thermodynamic study of the hydrolysis with HCl. This research identified this salient knowledge gap and aims to fill it.

In this work, we explored the hydrolysis of methyl acetate catalyzed with HCl and studied its behaviors from the kinetic and thermodynamic point of view.

2.0 Theory

The rate of a chemical reaction can be measured analytically and experimentally and thereafter treated as a mathematical analysis or model. Likewise, the specific reaction rate constants can be calculated at two or more different temperatures and thereafter, the activation energy is determined. In the prediction of physical – chemical systems behaviour, knowledge of the equilibrium conditions as well as the speed with which the equilibrium is attained is paramount [9, 10, 11, 12]. In some reactions, equilibrium is attained in a short period of time that the reaction appears to be instantaneous while in other reactions, the lag-time parameter is quite important and lengthy.

The hydrolysis of methyl acetate is catalyzed by hydrogen ions from the hydrochloric acid. It does not proceed with any measurable velocity in pure water [16]. Although, two molecules are involved in the hydrolysis, water is present in such excess that only the methyl acetate appears to change in concentration. Also, the large excess of water prevents any appreciable backward reaction which will result to formation of methyl acetate. As each molecule of methyl acetate is hydrolyzed, one molecule of methanol and one molecule of acetic acid are produced. The increase in acidity is a direct measure of the amount of methyl acetate that has reacted [16]. Methanol boils at 64.7 °C, water boils at 100 °C while acetic acid boils at 117.9 °C, and this makes the separation of methanol from the mixture much easier.

The kinetic study of the hydrolysis provides information on the mechanism and characteristics of the hydrolysis reaction [17]. The pseudo first order and second order kinetic models need to be tested to determine which model is in good agreement with the experiment ‘a – x’ (amount of methyl acetate hydrolyzed) value, thus suggesting which model the hydrolysis system follows.

2.1 Pseudo first order equation

A first-order reaction is one in which the rate of reaction is found; experimentally, to be directly proportional to the concentration of the reacting substance. Many reactions follow the first-order equation given as:

The first order reaction can be represented by

$A \rightarrow \text{Products}$

Taking ‘C’ as the concentration reactant ‘A’ at time ‘t’, the rate at which ‘A’ is used up can be represented as

$$\frac{-dC}{dt} = k_1 C \quad (1)$$

In terms of moles

$$\frac{-d(a-x)}{dt} = k_1 (a - x) \quad (2)$$

$$\frac{dx}{dt} = k_1 (a - x) \quad (3)$$

$$\frac{dx}{a-x} = k_1 dt$$

Upon integration

$$\ln \frac{a}{(a-x)} = k_1 t \quad (4)$$

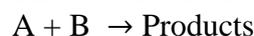
The amount of hydrochloric acid remains unchanged throughout the process. If T_0 , T_t and T_∞ are the titres/ml of both hydrochloric acid and acetic acid at times 0, t and 48 hours respectively, then, the concentration of methyl acetate at time 't' is proportional to $(T_\infty - T_t)$ while the initial concentration of methyl acetate 'a' is proportional to $(T_\infty - T_0)$. Therefore,

$$\ln \left(\frac{T_\infty - T_0}{T_\infty - T_t} \right) = k_1 t \quad (5)$$

$$\text{Log} \left(\frac{T_\infty - T_0}{T_\infty - T_t} \right) = \frac{k}{2.303} t \quad (6)$$

2.2 Pseudo second order equation

The second order reaction can be represented by



General definition = fundamental postulate

$$\frac{dx}{dt} = K' (a - x) (b - x) \quad (7)$$

$$\frac{dx}{(a-x)(b-x)} = K' dt \quad (8)$$

Applying partial fraction method

$$\frac{1}{(a-x)(b-x)} = \frac{A}{(a-x)} + \frac{B}{(b-x)}$$

$$= \frac{1}{(a-b)} \left(\frac{-1}{(a-x)} + \frac{1}{(b-x)} \right)$$

$$\frac{1}{(a-b)} \int \frac{-1}{(a-x)} + \frac{1}{(b-x)} = K' \int dt$$

$$\frac{1}{(a-b)} \int \left(\frac{-dx}{(a-x)} + \frac{dx}{(b-x)} \right) = K' t$$

$$\frac{1}{(a-b)} (\ln(a-x)_0^x - \ln(b-x)_0^x) = K' t$$

$$\frac{1}{(a-b)} (\ln(a-x) - \ln a - (\ln(b-x) - \ln b)) = K' t$$

$$\frac{1}{(a-b)} \left(\ln \frac{(a-x)}{a} + \ln \frac{b}{(b-x)} \right) = K' t$$

$$\frac{1}{(a-b)} \left(\ln \frac{(a-x)}{a} \frac{b}{(b-x)} \right) = K' t$$

$$\frac{1}{(a-b)} \left(\ln \frac{b(a-x)}{a(b-x)} \right) = K' t \quad (9)$$

Where $(a - x) = T_t$

$$b = T_0 - T_\infty$$

$$(b - x) = T_t - T_\infty$$

$$a = T_0$$

$$a - b = T_0 - T_0 - T_\infty = T_0 + T_\infty$$

$$t = \frac{2.303}{K(a-b)} \text{Log} \frac{(a-x)b}{(b-x)a}$$

$$= \frac{2.303}{KV_\infty} \text{Log} \frac{(T_0 - T_\infty)T_t}{(T_t - T_\infty)T_0}$$

(10)

Plotting $\text{Log} \frac{(T - T_\infty)T}{(T_t - T_\infty)T_0}$ against t

2.3 The nth order of the adsorption

$$\frac{dx}{dt} = k_n (a - x)^n \quad (11)$$

Where n is the order of the hydrolysis,
k_n is the rate constant of the nth order hydrolysis,

$$\frac{dx}{(a - x)^n} = k_n dt \quad (12)$$

$$\int_0^x \frac{dx}{(a - x)^n} = k_n \int_0^t dt \quad (13)$$

$$\left[\frac{(a - x)^{-(n-1)}}{(n-1)} \right]_0^x = k_n t \quad (14)$$

$$\frac{(a - x)^{-(n-1)}}{n-1} - \frac{a^{-(n-1)}}{n-1} = k_n t \quad (15)$$

$$(a - x)^{-(n-1)} - a^{-(n-1)} = (n - 1) k_n t$$

$$\text{Log } t = (n - 1) \text{Log} \left(\frac{a}{a - x} \right) - \text{Log} (n - 1) k_n \quad (16)$$

$$\text{Log } t = (n - 1) \text{Log} \left(\frac{T_\infty - T_0}{T_\infty - T_t} \right) - \text{Log } k_n (n - 1) \quad (17)$$

2.4 Arrhenius equation and activation energy

From the Van't Hoff equation $\frac{d(\ln k)}{dT} = \frac{\Delta H}{RT^2}$

$$(18)$$

Under isobaric and isochoric conditions,

$$k = \frac{k_2}{k_1}$$

Where 'k' is chemical equilibrium constant,

k₁ and k₂ are rate constants of forward and reverse reactions,

H = Heat of reaction,

R = Universal gas constant

T = Absolute temperature

Substituting $\frac{k_2}{k_1}$ for k,

$$\frac{d \ln \frac{k_2}{k_1}}{dT} = \frac{d \ln k_1 - d \ln k_2}{dT} = \frac{\Delta H}{RT^2} \quad (19)$$

The heat of a chemical reaction can be taken as the difference of the two quantities,

$$\Delta H = H_1 - H_2 \quad (20)$$

This gives:

$$\frac{d \ln k_1}{dT} - \frac{d \ln k_2}{dT} = \frac{H_1}{RT^2} - \frac{H_2}{RT^2} = \frac{\Delta H}{RT^2} \quad (21)$$

This can be considered with accuracy that

$$\frac{d \ln k_1}{dT} = \frac{H_1}{RT^2} + C \quad (22)$$

$$\frac{d \ln k_2}{dT} = \frac{H_2}{RT^2} + C \quad (23)$$

Where $C = 0$, hence

$$\frac{d \ln K}{dT} = \frac{H}{RT^2} \quad (24)$$

Equation (24) is Arrhenius equation which characterizes the dependence of the rate constants 'k' of a chemical reaction on the temperature.

The quantity 'H' has the dimensions of energy and is known as the "activation energy". It is defined as the excess energy in comparison with the mean energy of the molecules at a given temperature which a molecule must have in order to enter into a chemical reaction.

Integration of equation (24) and assuming 'H' to be constant yields:

$$\ln k = \frac{-H}{RT} + \ln A \quad (25)$$

Where $\ln A$ = integration constant

A plot of $\ln k$ against $\frac{1}{T}$ gives a straight line with slope equal to $-\frac{H}{R}$ and intercept equal to $\ln A$.

For the adsorption process, the natural logarithm of the rate constant (k) could be represented as a straight line function of $\frac{1}{T}$

$$\ln k = \frac{-H}{RT} + \ln A \quad (26)$$

Where 'k' is the equilibrium constant, 'A' is a frequency factor, 'R' is the universal gas constant ($8.314 \text{ J.K}^{-1}.\text{mol}^{-1}$) and 'T' is the absolute temperature. The value of 'H' is calculated from the slope of plotting $\ln k$ versus $\frac{1}{T}$, and 'A' (min^{-1}) is determined from the intercept.

Equation (24) can also be integrated within the limits T_1 to T_2 to give

$$\ln \frac{k_{T_2}}{k_{T_1}} = \frac{H(T_2 - T_1)}{R T_1 T_2} \quad (27)$$

Where k_{T_2} = rate constant of chemical reaction at T_2 ,

k_{T_1} = rate constant of chemical reaction at T_1 [12].

2.5 Thermodynamic studies

The determination of the basic thermodynamic parameters of the hydrolysis, such as changes in enthalpy (ΔH), Gibbs free energy (ΔG) and entropy (ΔS), is important as it allows to estimate if the hydrolysis is favorable or not from thermodynamic point of view, to assess the spontaneity of the system, and to ascertain whether the hydrolysis is exothermic or endothermic in nature [20]. These thermodynamic parameters of the hydrolysis will be determined from the experimental data obtained at equilibrium after the hydrolysis has gone into completion at various temperatures using the following Gibbs free energy equations:

$$\Delta G^0 = \Delta H^0 - T \Delta S^0 \quad (28)$$

Where 'T' is the temperature of the system

Also, Gibbs free energy isotherm equation is given by:

$$\Delta G = -RT \ln K_d \quad (29)$$

$$K_d = \frac{q_e}{c_e} \quad (30)$$

$$\ln K_d = \frac{\Delta S}{R} - \frac{\Delta H}{RT} \quad (31)$$

K_d = distribution coefficient for the hydrolysis. $q_e = c_o - c_e =$ amount of methyl acetate hydrolyzed

c_e = Concentration of methyl acetate after the hydrolysis has attained equilibrium for a particular temperature,

' T ' is the absolute temperature,

' R ' is gas constant ($8.314 \text{ J/K}^{-1} \cdot \text{mol}^{-1}$).

The values of changes in enthalpy (ΔH°) and entropy (ΔS°) are obtained from the slope and intercept of $\ln K_d$ versus $\frac{1}{T}$ plots [20].

3.0 Methodology

3.1 Design of experiment

The design of experiment was carried out using temperature, time and volume as the variables, while the Titre value was obtained as the response.

Table 1: Central composite design for 5-level factor code for the hydrolysis of methyl acetate

Factor	-2	-1	0	+1	+2
Temperature ($^\circ\text{C}$)	30	35	40	45	50
Time (min)	10	20	30	40	50
Vol. of methyl acetate (ml)	1	2	3	4	5

Table 2: Design of experiment for the hydrolysis of methyl acetate

Run	Variables			Response
	Temperature ($^\circ\text{C}$)	Time (min)	Volume (ml)	Titre (ml)
1	40	50	3	37.90
2	40	30	3	23.30
3	35	40	2	23.70
4	35	20	4	23.20
5	40	30	1	22.50
6	35	40	4	24.10
7	40	30	5	23.80
8	40	30	3	23.30
9	45	20	4	23.40
10	45	40	2	25.00
11	35	20	2	23.50
12	45	40	4	28.80
13	50	30	3	28.80
14	40	30	3	23.30
15	45	20	2	24.60
16	40	30	3	23.30
17	30	30	3	17.80
18	40	10	3	8.30
19	40	30	3	23.30
20	40	30	3	23.30

3.2 Materials

Thermostatic water-bath, 250 ml Erlenmeyer conical flasks, 5 ml pipettes, graduated burettes, retort stand, stop watch, methyl acetate, 2 litres 0.2 N sodium hydroxide, 500 ml 2N hydrochloric acid, 1 litre distilled water, phenolphthalein indicator.

3.3 Experimental procedure

The concentration of methyl acetate at any time was determined by titration of samples with 0.2N sodium hydroxide. The sodium hydroxide solution used was prepared from saturated sodium hydroxide in order to reduce the amount of carbonate present. 100 ml of 0.5 mol./dm³ hydrochloric acid was placed in a 150 ml flask clamped in a thermostatic water bath at 30 °C. Another solution was placed in a second for check determination. Some quantity of methyl acetate in a test tube was also set into the thermostatic water bath. After allowing 10 min to 15 min for the solutions to achieve the temperature of the thermostatic water bath, exactly 5 ml of the methyl acetate was pipetted into one of the flasks. It was shaken to mix thoroughly and a sample of exactly 5 ml was withdrawn immediately and run into a 25 ml of ice cold water to arrest the reaction. The dilution with ice cold water was to arrest the reaction though the titration was carried out immediately. The time at which the pipette was half emptied into the water of the titration flask was recorded accurately along with the titration conducted. The titration was further repeated with 5.0 ml samples after 10, 20, 30, 40, and 50 mins. About half an hour after starting the first reaction, the check determination was started, and the two reactions were carried along. The remaining solutions were stoppered and allowed to stand in the desk for a couple of days in order to permit a final titration when reaction mixture had reached equilibrium. Additional determinations were carried out at other investigated temperatures. Again, samples of the remaining solutions at the investigated temperatures were retained for final titration after standing for at least two days.

4.0 Results and discussion

The concentrations of methyl acetate left at various time interval during the hydrolysis can be determined from the titration values obtained. These titre values obtained from the hydrolysis of methyl acetate catalyzed by HCl are presented in Table 3.

Table 3: Titre values at the investigated temperatures

	30 °C	35 °C	40 °C
Time (min)	Titre value (ml)	Titre value (ml)	Titre value (ml)
0	26.10	26.60	27.80
10	27.50	27.70	29.10
20	27.70	28.10	28.90
30	28.10	28.20	29.60
40	27.90	28.90	30.90
50	28.40	29.20	31.20
∞	33.60	34.60	34.30

The optimal data obtained from the design of experiment shows that maximum production of methanol occurred in laboratory scale when 5 ml of methyl acetate was used at 40 °C and 30 mins residence time. The first, second and nth order kinetic models were used for fitting the experimental data obtained. The first order kinetic model provided good fit for the hydrolysis, which agrees with a study conducted [4]. The study showed that the hydrolysis of methyl acetate catalyzed with hydrochloric acid is a first order kinetic model.

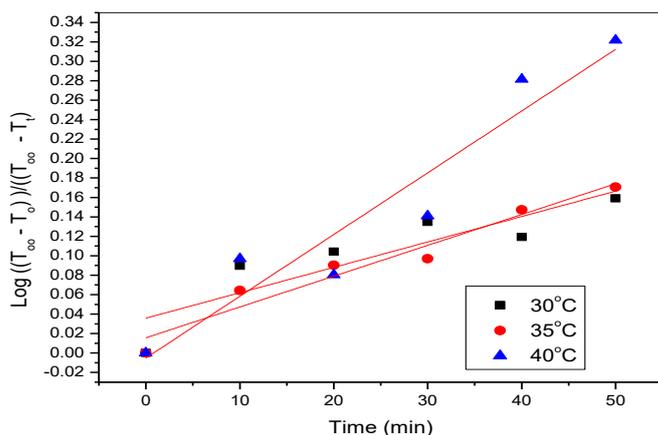


Figure 1: Pseudo 1st order reaction kinetic at the investigated temperatures

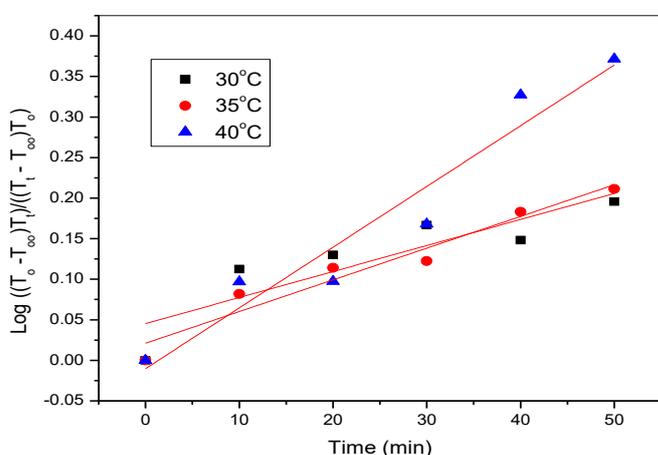


Figure 2: Pseudo 2nd order reaction kinetic at the investigated temperatures

Figures 1 and 2 shows that as the temperature increases, the second order kinetic model becomes better and at 40 °C, it provided the best fit with a correlation coefficient of 0.9373 than the first order kinetic model with correlation coefficient of 0.9176. This is clearly seen from the data obtained with the nth order kinetic, which shows that hydrolysis of methyl acetate catalyzed with hydrochloric acid is not absolutely first order or second order, but a little higher than the first order. This was observed to be in agreement with studies conducted [1, 2, 3].

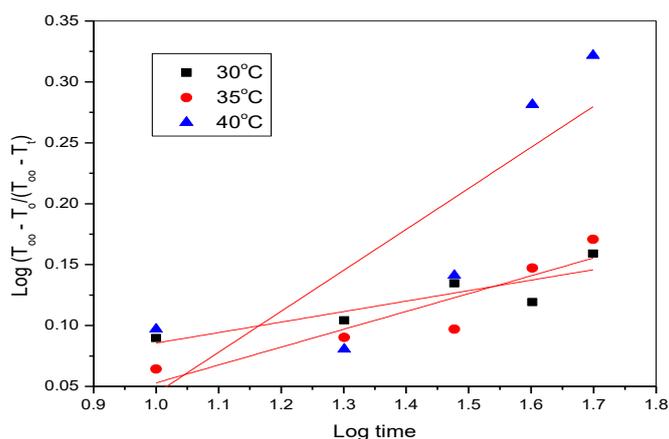


Figure 3: n^{th} order kinetic model at the investigated temperatures

The second order kinetic model also provided good fit for the hydrolysis [1], but the first order kinetic model provided better fit at 30 °C and 35 °C with correlation coefficient of 0.7870 and 0.9523 respectively. This necessitated the use of the n^{th} order kinetic model for the experimental data as shown in Figure 3.

Table 4: Rate constants at the investigated temperatures

Temperature (K)	303	308	313
K_1 (min^{-1})	0.00599	0.00737	0.0145
R^2	0.7870	0.9523	0.9176
K_2 ($\text{min}^{-1} \cdot \text{cm}^{-3}$)	0.00022	0.00026	0.00050
R^2	0.7798	0.9485	0.9373

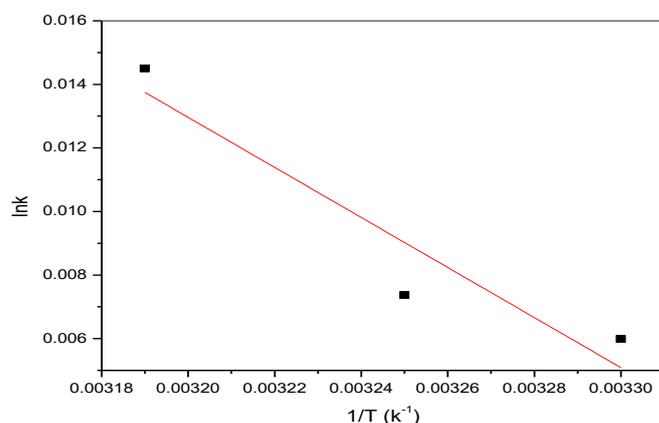


Figure 4: Arrhenius equation for 1st order reaction

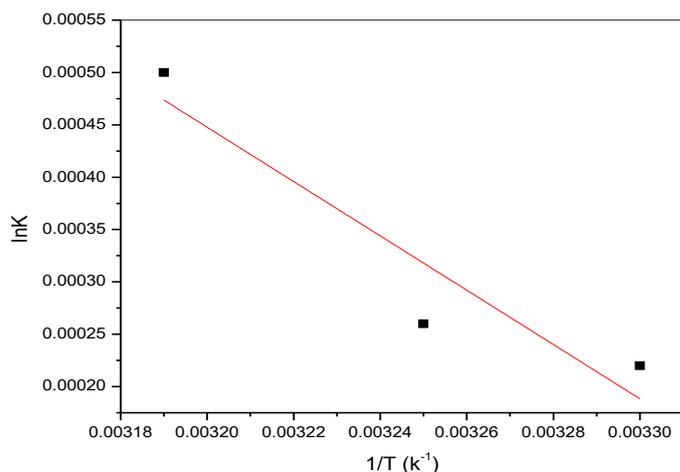


Figure 5: Arrhenius equation for 2nd order reaction

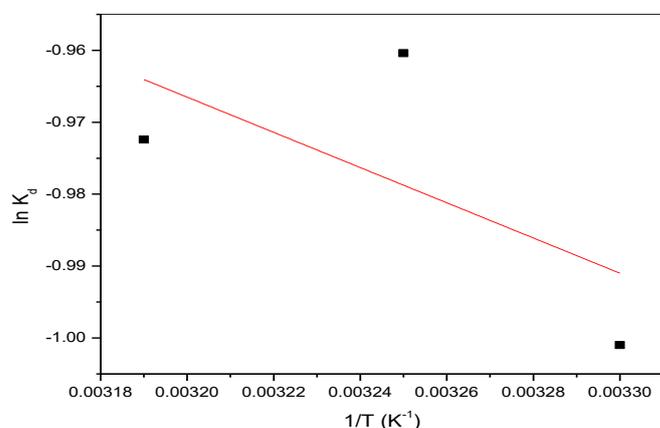


Figure 6: Thermodynamic effect on the hydrolysis of methyl acetate catalyzed with HCl

4.1 Discussion of results

Using equation (17), at 30 °C, the order of the hydrolysis was evaluated as 1.0861 with rate constant of 11.628. At 35 °C, the order was evaluated as 1.1465 with rate constant of 8.4657, and at 40 °C the order of the hydrolysis was evaluated as 1.3366 with rate constant of 5.8235. The activation energy for the first order model was obtained as 654.57J/mole while that of the second order was obtained as 21.56J/mol. These results showed that if the first order model was applied, more energy would be required before the hydrolysis can take place compare to when the second order kinetic model is used.

The thermodynamic parameters such as changes in enthalpy, entropy and Gibbs free energy were determined from Figure 6 using equations (29) and (31). The change in enthalpy for the hydrolysis was obtained as 2.04KJ/mole, showing that the hydrolysis is an endothermic reaction. This is higher than the value of obtained by Yu *et al* [8]. The change in entropy was obtained as -1.52J/Mol.K, which shows that the degree of disorderliness of the hydrolysis is minimal and in the negative. The Gibbs free energy was obtained as 2.50KJ/mole confirming that the hydrolysis of methyl acetate catalyzed by hydrochloric acid is non-spontaneous.

5.0 Conclusion

The kinetic and thermodynamic behavior of the hydrolysis of methyl acetate catalyzed with HCl was undertaken in this research. At 30 °C and 35 °C, the first order kinetic model gave better fit for the hydrolysis than the second order kinetic model. However, at 40 °C, the second order model gave better fit for the hydrolysis, indicating that at higher temperatures, hydrolysis of methyl acetate catalyzed with hydrochloric acid does not follow the first order kinetic model. Furthermore, the results obtained with the n^{th} order kinetic model showed that the hydrolysis of methyl acetate catalyzed with hydrochloric acid in the production of methanol does not exactly follow the first order kinetic model. The enthalpy of the hydrolysis showed that it is endothermic while the Gibbs free energy revealed that the hydrolysis is slow in nature.

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