Kinetics of the Permanganate Oxidation of Furfural to Furoic Acid in Alcoholic Medium

Yogeshwar N. Dethe[1*], Hritika Dhotre[2], Rushikesh R. Warkad[3], Bhagwansing S. Dobhal[4], A. A. Pathan[1], Ravindra P. Shimpi[4*]

[1]Moulana Azad College, Chh. Sambhajinagar (MS) India
[2]Kokan Gyanpeeth Rahul Dharkar College of Pharmacy & Research Institute Karjat (MS), India
[3]Srinath College of Pharmacy, Chhatrapati Sambhaji Nagar, India
[4]Badrinarayan Barwale College, Jalna (MS) India.
***Corresponding Author**
Email: yogeshdethe222@gmail.com, raviindra336@gmail.com
Received: 25th May, 2026; **Revised:** 6th June, 2026; **Accepted:** 15th July, 2026; **Available Online:** 10th Aug, 2026

ABSTRACT
The kinetics and mechanism of the oxidation of furfural by acidic potassium permanganate were investigated in an alcoholic medium using UV-Visible spectrophotometry by monitoring the decrease in permanganate absorbance at 526 nm. The reaction was studied under pseudo-first-order conditions with furfural and hydrogen ions in excess over permanganate. Furoic acid (2-furoic acid) was obtained as the final oxidation product, with an overall stoichiometry of two moles of permanganate per five moles of furfural. Kinetic studies revealed first-order dependence on both furfural and permanganate concentrations, whereas a fractional-order dependence on hydrogen ion concentration indicated a protonation equilibrium preceding the rate-determining step. The reaction rate increased with increasing dielectric constant but was independent of ionic strength, suggesting the involvement of a polar activated complex and non-ionic species. A mechanism involving protonation of furfural, formation of a cyclic manganate ester intermediate, and its slow decomposition to furoic acid is proposed. Activation parameters supported the formation of a highly ordered transition state and provided further evidence for the proposed reaction pathway.
Keywords: Furfural; Potassium permanganate; Oxidation kinetics; Furoic acid; Heterocyclic aldehydes; UV-Visible spectrophotometry.
How to cite this article: Dethe YN, Dhotre H, Warkad RR, Dobhal BS, Pathan AA, Shimpi RP. Kinetics of the permanganate oxidation of furfural to furoic acid in alcoholic medium. Int J Drug Deliv Technol. 2026;16(74s): 560-568. DOI: 10.25258/ijddt.16.74s.66

Source of support: Nil.
Conflict of interest: Nil.

INTRODUCTION

Furfural (2-furaldehyde) is one of the most prominent platform chemicals derived from renewable lignocellulosic biomass. It is produced industrially through the acid-catalysed dehydration of pentose sugars, primarily xylose, which is obtained from hemicellulose present in agricultural residues such as corncobs, oat hulls, rice husks, bagasse, and cottonseed hulls [1]. The global production of furfural exceeds 300,000 tons annually, and its market value continues to grow due to increasing demand for bio-based chemicals and materials. The U.S. Department of Energy has recognised furfural as one of the top bio-based platform chemicals, owing to its versatility as a feedstock for the synthesis of various value-added products, including furfuryl alcohol, tetrahydrofuran, furoic acid, and numerous pharmaceutical intermediates [2]. The unique molecular structure of furfural, comprising a furan ring with an aldehyde substituent at the 2-position, endows it with diverse chemical reactivity. The aldehyde functionality renders furfural susceptible to various transformations, including hydrogenation, oxidation, decarbonylation, and condensation reactions [3]. Among these, the oxidation of furfural to furoic acid (2-furoic acid) is of particular industrial and academic interest. Furoic acid is a valuable chemical intermediate used in the production of pharmaceuticals, agrochemicals, preservatives, and flavouring agents [4]. Additionally, it serves as a precursor to the synthesis of biodegradable polymers and other speciality chemicals, making its efficient production highly desirable in the context of green chemistry and sustainable development.

The oxidation of furfural can be achieved using various oxidising agents, including molecular oxygen, hydrogen peroxide, and metal oxidants such as permanganate, dichromate, and ceric ammonium

nitrate [5]. Among these, potassium permanganate ($KMnO_4$) is a powerful and widely used oxidising agent in organic chemistry, known for its ability to oxidise aldehydes, alcohols, and other functional groups under mild conditions [8]. The permanganate ion (MnO_4^-) undergoes stepwise reduction through various manganese oxidation states, including Mn(VI), Mn(V), Mn(IV), and ultimately Mn(II) in acidic media [6]. This versatility makes permanganate an attractive reagent for mechanistic studies, as the reaction kinetics and intermediate species can be conveniently monitored using UV-Visible spectrophotometry due to the intense purple colour of MnO_4^-, which exhibits a characteristic absorption maximum at 526 nm [10]. The oxidation of aldehydes by permanganate generally proceeds through the formation of a cyclic manganate diester intermediate via nucleophilic addition of MnO_4^- to the carbonyl group, followed by a rate-determining two-electron transfer decomposition to yield the corresponding carboxylic acid [11,12]. Earlier kinetic studies have reported that the permanganate oxidation of furfural follows a first-order dependence on the concentrations of both furfural and permanganate [6]. However, the oxidation of furfural presents unique features due to the heterocyclic nature of the substrate and the electron-donating properties of the furan ring. The presence of the ring oxygen atom influences the electronic distribution within the molecule, potentially affecting the reactivity and mechanism of the oxidation process [7]. Despite the industrial relevance of furfural, relatively few detailed kinetic and mechanistic studies have been reported on its oxidation by permanganate, creating a gap in understanding of its oxidation chemistry.

The present study was therefore undertaken to investigate the kinetics and mechanism of the oxidation of furfural by acidic potassium permanganate using UV-Visible spectrophotometry [8]. The objectives include determining the rate law and evaluating kinetic parameters such as reaction order, rate constant, and activation energy, elucidating the reaction mechanism through spectroscopic monitoring and identification of intermediate species, and evaluating the effect of various parameters, including acid concentration, temperature, dielectric constant, and ionic strength, on the reaction rate. The proposed mechanism involves protonation of the aldehyde oxygen to form an activated furfural, followed by nucleophilic attack by MnO_4^- to generate a cyclic manganate diester intermediate, which undergoes rate-determining two-electron-transfer decomposition to yield furoic acid [9] and reduced manganese species, which are subsequently converted to Mn^{2+} in the acidic medium. The findings of this study are expected to contribute to the development of efficient and selective oxidation protocols for biomass-derived aldehydes, thereby supporting the transition toward sustainable chemical manufacturing [10,11]. The paper is organised into experimental, results, discussion, and conclusion sections, presenting a comprehensive account of the kinetic and mechanistic investigation of furfural oxidation by acidic potassium permanganate [12].

2. Materials

2.1. Chemicals & Reagents:

All chemicals used were of Analytical Reagent (A.R.) grade. Furfural (C_4H_3OCHO, MW = 96.08 g/mol), potassium permanganate ($KMnO_4$, MW = 158.034 g/mol), and concentrated sulfuric acid (H_2SO_4, 98%) were obtained from Sigma-Aldrich. Salts including potassium chloride (KCl, MW = 74.55 g/mol), potassium bromide (KBr, MW = 119.00 g/mol), potassium sulfate (K_2SO_4, MW = 174.26 g/mol), magnesium chloride hexahydrate ($MgCl_2 \cdot 6H_2O$, MW = 203.30 g/mol), calcium nitrate tetrahydrate ($Ca(NO_3)_2 \cdot 4H_2O$, MW = 236.15 g/mol), and aluminum nitrate nonahydrate ($Al(NO_3)_3 \cdot 9H_2O$, MW = 375.13 g/mol) were procured from. Double-distilled water was used throughout the experiments.

Chemical Name	Molecular Formula	Molecular Weight (g/mol)	Purity	Purpose
Furfural	C_4H_3OCHO	96.08	Analytical Reagent (≥99%)	Substrate (oxidizable compound)
Potassium Permanganate	$KMnO_4$	158.034	Analytical Reagent (≥99.5%)	Oxidizing agent
Sulfuric Acid	H_2SO_4	98.08	Analytical Reagent (98%)	Acidic medium provider
Potassium Chloride	KCl	74.55	Analytical Reagent (≥99%)	Salt effect study (neutral salt)

Potassium Bromide	KBr	119.00	Analytical Reagent (≥99%)	Salt effect study (halide inhibition)
Potassium Sulfate	K_2SO_4	174.26	Analytical Reagent (≥99%)	Salt effect study (dianion effect)
Magnesium Chloride Hexahydrate	$MgCl_2·6H_2O$	203.30	Analytical Reagent (≥98%)	Salt effect study (divalent cation)
Calcium Nitrate Tetrahydrate	$Ca(NO_3)_2·4H_2O$	236.15	Analytical Reagent (≥99%)	Salt effect study (nitrate anion)
Aluminum Nitrate Nonahydrate	$Al(NO_3)_3·9H_2O$	375.13	Analytical Reagent (≥98.5%)	Salt effect study (trivalent cation)
Distilled Water	H_2O	18.02	Double-distilled	Solvent for all solutions

2.2. Instrument:

A Labman LMSP UV-1900S double beam UV-Visible spectrophotometer with UV LabSolutions software (version 2.0) was used for the kinetic measurements. An analytical balance (Shimadzu ATX224, ±0.0001 g accuracy) was utilised for weighing.

2.3. Preparation of solutions:

All stock solutions were prepared using analytical-grade chemicals and double-distilled water. The concentrated acid was carefully diluted to create a 1.0 M sulfuric acid solution. Potassium permanganate stock solution (1.0×10^{-2} M) was prepared weekly and stored in amber bottles. Furfural solution (0.1 M) was prepared daily to ensure freshness due to its sensitivity to oxidation. KCl, KBr, K_2SO_4, $MgCl_2$, $Ca(NO_3)_2$, and $Al(NO_3)_3$ salt solutions (0.1 M each) were made using commercially available hydrated forms. All volumetric measurements were conducted using Class A glassware, and solutions were held under proper conditions until use.

Solution	Conc.	Mass	Storage condition
Furfural	0.1M	0.385 g	Fresh daily (4°C)
Potassium Permanganate	1.0×10^{-2} M	0.1580 g	Ambered colour Volumetric Flask, RT
Sulfuric Acid	1.0 M	5.56 mL conc.	Volumetric Flask, RT
KCl Solution	0.1 M	0.0746 g	Volumetric Flask, RT
KBr Solution	0.1 M	0.1190 g	Ambered colour Volumetric Flask, RT
K_2SO_4 Solution	0.1M	0.1743 g	Volumetric Flask, RT
$MgCl_2$ Solution	0.1M	0.2033 g	Volumetric Flask, RT
$Ca(NO_3)_2$ Solution	0.1M	0.2362 g	Volumetric Flask, RT
$Al(NO_3)_3$ Solution	0.1 M	0.3751 g	Volumetric Flask, RT

3. Kinetic Measurements

3.1. Determination of λ_{max}

The absorption maximum of $KMnO_4$ was found by scanning a 1.0×10^{-3} M solution in 0.1 M H_2SO_4 from 450 to 600 nm using the spectrophotometer's spectrum mode. For all kinetic data, a strong peak was seen at 530 nm.

3.2. General Kinetics Methods:

Furfural was used in significant excess over $KMnO_4$ in all kinetic simulations conducted under pseudo-first-order conditions [13]. A predetermined amount of $KMnO_4$ solution was quickly added to a thermostatic mixture of distilled water, sulfuric acid, and furfural (with or without additional salts) to initiate the reaction. The total reaction volume was maintained at 50.0mL [14].

The UV LabSolutions software's Time Course (kinetics) mode was used to track the reduction in absorbance at 530 nm using the following parameters:

Parameters	Observation

Experimental Variable Studies:

- **Impact of Furfural Concentration:** Varied between $2.0 \times 10^{-3}$and 8.0×10^{-3}M, while keeping the oxidant concentration and sulfuric acid concentration constant.
- **Impact of Sulphuric Acid Concentration:** Varied from 0.05 to 0.3 M, while keeping [furfural] = 4.0×10^{-3}M and [$KMnO_4$] constant.
- **Impact of Added Salts:** Examined using KCl, KBr, K_2SO_4, $MgCl_2$, $Ca(NO_3)_2$, and $Al(NO_3)_3$ concentrations of 0.01 and 0.02 M.
- **Impact of Oxidant Concentration:** If you are using $KMnO_4$ for furfural oxidation, include a separate study where [$KMnO_4$] is varied while furfural and acid are held constant.

3.3. Data analysis:

The slope of the linear plot of ln (Absorbance) vs time was used to calculate the pseudo-first-order rate constant (k') using the following equation [15]:

$$\ln(A_t) = -k't + \ln(A_0)$$

A_t = Absorbance at time t.

A_0 = Initial absorbance.

k' = Pseudo-first-order rate constant.

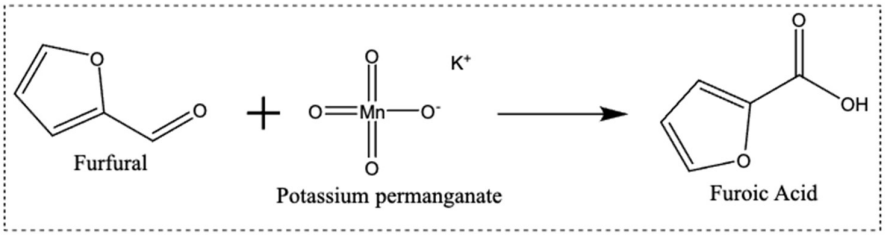
Linear regression analysis was performed in Microsoft Excel 365, and correlation coefficients (R^2) exceeding 0.99 were obtained for all runs. Plots of log k' vs log (concentration) were used to identify the sequence of reaction with regard to each reactant. Using rate constants found at various temperatures (25, 30, 35, and 40°C), activation parameters were computed using the Arrhenius and Eyring equations [15–18].

4. Proposed Mechanism for the Oxidation of Furfural

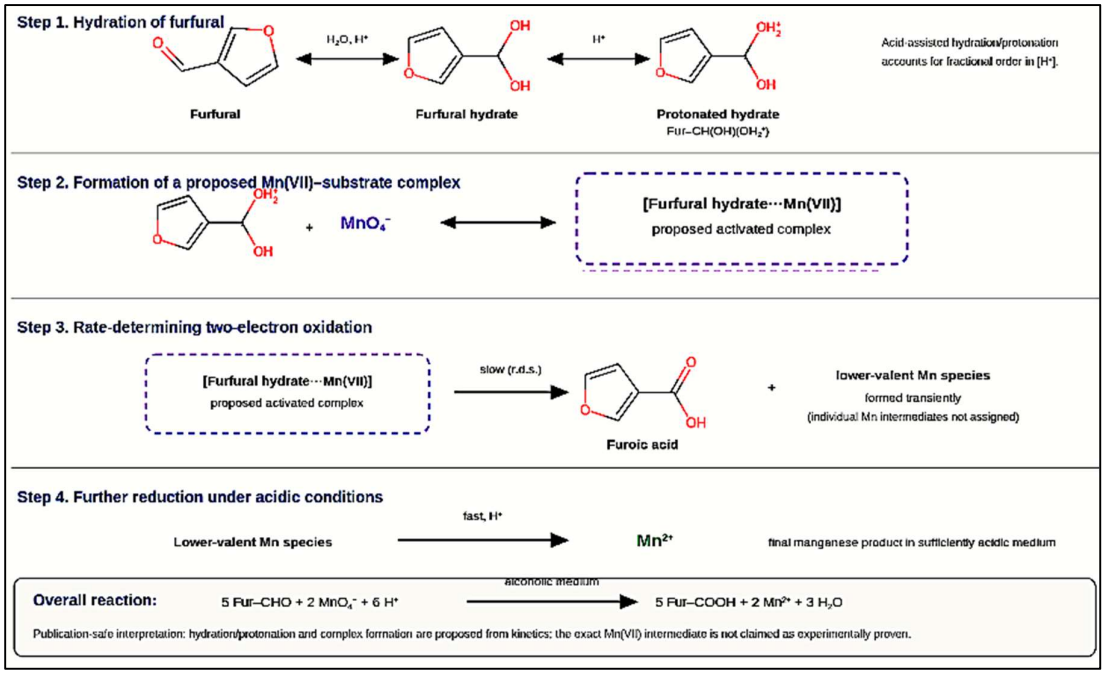
Permanganate oxidations are generally multistep electron-transfer processes in which manganese(VII) is reduced through lower oxidation states, depending on the substrate and reaction medium. Under strongly acidic conditions, permanganate is ultimately reduced to Mn^{2+} through an overall five-electron reduction process [19]. Since oxidation of an aldehyde to a carboxylic acid is a two-electron process, several furfural oxidation events are coupled with the complete reduction of permanganate. The kinetic results obtained in the present study, together with previous reports on aldehyde oxidation by permanganate, support the following reaction mechanism.

Overall Oxidation Reaction:

In the acidic alcoholic medium, furfural is activated through reversible hydration of the aldehyde group and protonation of the carbonyl oxygen. These pre-equilibrium steps increase the susceptibility of the substrate toward oxidation by permanganate and are consistent with the observed acid dependence of the reaction rate [20].



Proposed Mechanism for the Oxidation of Furfural to Furoic Acid by Acidic Permanganate in Alcoholic Medium



Step I. Hydration and Acid Activation of Furfural

In the aqueous alcoholic acidic medium, furfural exists in equilibrium with its hydrated (gem-diol) form. Protonation of the carbonyl oxygen further facilitates hydration and activates the substrate toward oxidation.

The observed fractional-order dependence on hydrogen-ion concentration is consistent with this acid-assisted activation equilibrium preceding the rate-determining step.

Step II. Formation of a Proposed Mn(VII)–Substrate Complex

The activated furfural hydrate subsequently interacts with the permanganate ion to form a transient Mn(VII)–substrate complex (or permanganate ester-like intermediate) [21]. Although this intermediate was not directly observed, its formation is consistent with the kinetic behaviour of the reaction. The rapid decrease in absorbance at **526 nm** confirms the consumption of permanganate during the reaction but does not independently establish the exact structure of the intermediate.

Step III. Rate-Determining Oxidation

The proposed Mn(VII)-substrate complex undergoes slow decomposition through a two-electron electron-transfer process, involving cleavage of the aldehydic C-H bond to produce furoic acid together with transient lower-valent manganese species [9].

This step is considered the rate-determining step of the overall oxidation process.

Step IV. Reduction of Manganese Species

The transient lower-valent manganese species generated during oxidation undergo rapid subsequent electron-transfer reactions and are ultimately converted to Mn^{2+} under the acidic reaction conditions. Complete disappearance of the characteristic purple colour of permanganate is consistent with the reduction of Mn(VII), although the individual manganese intermediates cannot be identified solely from the kinetic measurements.

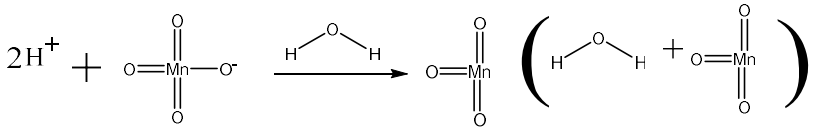
The proposed mechanism satisfactorily accounts for the observed first-order dependence on furfural and permanganate concentrations, the fractional-order dependence on hydrogen-ion concentration, the negligible effect of ionic strength, and the increase in reaction rate with increasing dielectric constant of the medium. These observations support a mechanism involving acid-assisted activation of furfural, formation of a transient Mn(VII)–substrate complex, and a slow electron-transfer step leading to the formation of furoic acid.

4.1 Nature of the Oxidising Species

The identity of the active oxidising species in acidic permanganate solutions has been the subject of extensive investigation because the oxidation behaviour of permanganate is strongly influenced by the acidity of the medium. Under the experimental conditions employed in the present study, the permanganate ion (MnO_4^-) is considered the principal oxidising species. In acidic solution, protonation of permanganate may increase its electrophilic character, thereby facilitating electron transfer from the substrate [22]. The observed fractional-order dependence on hydrogen-ion concentration suggests that acid catalysis plays an important role in activating the reaction through protonation of the substrate and/or enhancement of the oxidising ability of permanganate.

Consequently, oxidation is proposed to proceed through the interaction of activated furfural with Mn(VII), followed by formation of a transient Mn(VII)–substrate complex that undergoes slow decomposition to produce furoic acid and reduced manganese species.

The following equilibrium might be present in acidic permanganate solutions:

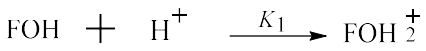


Permanganate ion, which is produced when O_3^+ MnO_3^+ forms, is a far more potent oxidising agent than $KMnO_4$. The several oxidation states of Mn involved in the reaction pathway may explain why all permanganate oxidations are typically complex.

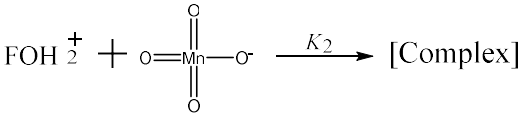
Rate law & rate equations:

Assuming that furfural and permanganate ion form an intermediate complex, the rate law and rate equations may be written as follows:

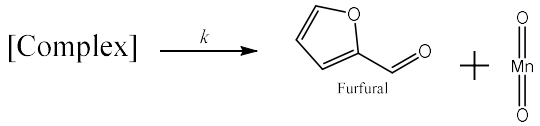
Step 1: Protonation of substrate.



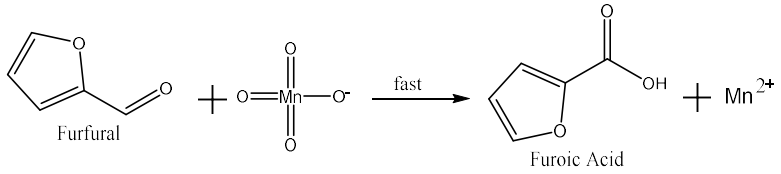
Step 2: Formation of Complex.



Step-3: Slow Decomposition of Complex.



Step-4: Fast Oxidation to Furoic Acid.



Rate = k [complex]

$[FOH_2^+] = K\,[FOH]\,[H^+]$

$[Complex] = K_1\,[FOH_2^+]\,[Mn^{7+}]$

$= K_1K_2\,[FOH]\,[H^+]\,[Mn^{7+}]$

Rate= $k\,KK_1\,[FOH]\,[H^+]\,[Mn^{7+}]$

In pseudo- first-order conditions, where [FOH] & $[H^+]$ are in excess:

$k' = k\,K_1K_2\,[FOH]\,[H^+]$

where k' is observed pseudo-first-order rate constant.

The first sharp decline in absorbance can be explained by the formation of the condensed product (manganate ester), which has a lower molar absorptivity than the permanganate ion at 526 nm [23].

A more thorough rate expression can be expressed as follows, taking into account the potential for various equilibria and distinct manganese species:

$$\text{Rate} = \frac{kK_1K_2[FOH][H^+][MnO_4^-]_T}{\{1+K_1[H^+]\}\{1+K_2[MnO_4^-]\}}$$

$[MnO_4^-]_T$ = the total concentration of permanganate.

4.2. Stoichiometry Determination

A mixture comprising excess furfural (0.01 M) and $KMnO_4$ (0.02 M) in 1 M H_2SO_4 was allowed to react for 24 hours at room temperature to determine the reaction stoichiometry. The unreacted $KMnO_4$ was quantified spectrophotometrically, and the product was identified as furoic acid by thin-layer chromatography and melting-point comparison.

4.3. Statistical Analysis

The findings are presented as mean ± standard deviation, and all kinetic runs were performed in triplicate. Statistical significance was assessed using Student's t-test; $p < 0.05$ was considered significant.

5. Result and Discussion

At 25°C in an acidic solution, the oxidation of Furfural (furan-2-carbaldehyde) by potassium permanganate was examined. A UV-Visible double-beam spectrophotometer was used to monitor the reaction by measuring the decrease in absorbance of $[KMnO_4]$ at its λ_{max} of 526 nm. When the concentration of furfural was higher than that of the oxidant, the kinetic runs were conducted under

pseudo-first-order conditions.

5.1. Effect of Oxidant $KMnO_4$ Concentration:

The concentrations of furfural and H_2SO_4 were held constant, while the concentration of $KMnO_4$ was varied to examine the impact of the oxidation. An increase in $KMnO_4$ was shown to raise the pseudo-first-order rate constant (k'). When log(rate) was plotted against log $KMnO_4$, a straight line with a slope of around one was produced. This demonstrates that the reaction is first-order with respect to $KMnO_4$. This finding is in line with kinetic analyses of related phenolic compounds.

The absorbance data for each run of furfural oxidation by $KMnO_4$ at 526 nm were analysed.

To calculate the pseudo-first-order rate constant (k'), we use the integrated rate law:

$$\ln(A_t) = \ln(A_0) - k't$$

Runs	Furfural (ml)	H_2SO_4 (ml)	DI Water (ml)	$KMnO_4$ (ml)	k' × 10^{-2} (sec^{-1})
I	2	5	27.5	2.5	0.638
II	4	5	25.5	2.5	0.956
III	6	5	23.5	2.5	1.089
IV	4	2.5	28	1	1.059
V	4	10	20	1	0.494
VI	4	15	15	1	-0.069

Observation:

I. The absorbance consistently decreases from Run I to Run IV, demonstrating a definite oxidation process. The rate constants vary between 1.089 × 10^{-2} sec^{-1} and 0.638 × 10^{-2} sec^{-1}, indicating that furfural undergoes oxidation by potassium permanganate under the experimental conditions.

II. Run V shows a significantly lower rate constant (0.494 × 10^{-2} sec^{-1}) compared to Run IV (1.059 × 10^{-2} sec^{-1}), which is attributed to the inhibitory effect of excess acid concentration. The decrease in rate suggests that while moderate acid facilitates protonation of the aldehyde group, higher acid concentrations suppress the reaction, possibly due to the formation of less reactive permanganate species.

III. Run VI exhibits a negative rate constant (-0.069 × 10^{-2} sec^{-1}), which is not feasible for a genuine decay curve. This negative value indicates that no measurable oxidation occurred at very high acid concentration (15.0 mL H_2SO_4). The absorbance slightly increased over time, suggesting complete inhibition of the reaction, likely due to decomposition of the manganate intermediate or the formation of unreactive Mn species under highly acidic conditions.

5.2. Effects of Salts on Reaction Rate:

In certain instances, the effect of salt on the rate constant (Table) shows a noticeable drop, which may be due to a negative catalytic effect. The rate constant is found to be independent of ionic strength in the presence of K_2SO_4, $MgCl_2$, and $Ca(NO_3)_2$. This suggests that the rate-determining step includes both an ion and a neutral molecule. For KBr, the rate constant initially increases at lower concentrations and then decreases at higher concentrations. On the contrary, there is no discernible sequence or pattern of rate constants for KI and KCl. The participation of I^- and Cl^- ions in competing oxidation processes with permanganate, which generate reactive halogen species that block the primary oxidation route, might be the cause of this erratic behaviour.

[Furfural] = 1 × 10^{-4} M, [H_2SO_4] = 1 M, [$KMnO_4$] = 1 × 10^{-3} M, Temperature = 25°C

Salt Conc.	Rate Constant in presence of salt (per sec)							
(mol/dm^3)	K_2SO_4	KBr	$AlCl_3$	KI	$MgCl_2$	$Ca(NO_3)_2$	KCl	$Al(NO_3)_3$
0.01	0.0324	0.0795	0.0411	0.0384	0.0539	0.0520	0.0492	0.0483
0.02	0.0346	0.1281	0.0482	0.0452	0.0530	0.0432	0.0555	0.0461
0.03	0.0392	0.0940	0.0436	0.0331	0.0483	0.0488	0.0594	0.0490
0.04	0.0490	0.0831	0.0450	0.0713	0.0484	0.0540	0.0470	0.0423
0.05	0.0569	0.0610	0.0472	0.0368	0.0422	0.0488	0.0438	0.0448
0.06	0.0596	0.0177	0.0428	0.0239	0.0555	0.0531	0.0509	0.0493
0.07	0.0408	0.0506	0.0487	0.0209	0.0493	0.0498	0.0463	0.0443
0.08	0.0593	0.0523	0.0436	0.0210	0.0468	0.0488	0.0493	0.0499
0.09	0.0427	0.0357	0.0453	0.0562	0.0468	0.0491	0.0446	0.0546

5.3. Effects of Temperature on Furfural Oxidation:

Furfural was oxidised by potassium permanganate at four different temperatures (293 K, 303 K, 313 K,

and 323 K) with constant reactant concentrations ([Furfural] = 1×10^{-1} M, [H_2SO_4] = 1 M, [$KMnO_4$] = 1×10^{-3}. The reaction's adherence to the Arrhenius law was confirmed by the observation that the rate constant rises with temperature.The activation energy (Ea#), activation enthalpy (ΔH#), activation entropy (ΔS#), and activation free energy (ΔG#) were derived from rate constants measured at various temperatures.

Rate k (per sec)	T (K)	1/T × 10^{-3} (K^{-1})	log k	$\Delta H^{\#}$ ($Jmol^{-1}$)	$\Delta S^{\#}$ ($J\ mol^{-1}$)	$\Delta G^{\#}$ ($J\ mol^{-1}$)
0.0202	287	3.41	-1.6946	-33052.28	-169.38	16575.40
0.0513	305	3.30	-1.2898	-33135.42	-164.64	16749.31
0.0646	317	3.19	-1.1897	-33218.56	-162.52	17649.90
0.0672	328	3.10	-1.1726	-33301.70	-161.18	18757.53
Average				**-33676.99**	**-164.43**	**17433.11**

6. References

[1] K.A. Gardner, L.L. Kuehnert, J.M. Mayer, Hydrogen Atom Abstraction by Permanganate: Oxidations of Arylalkanes in Organic Solvents, Inorg. Chem. 36 (1997) 2069–2078. https://doi.org/10.1021/IC961297Y.

[2] A.J. Fatiadi, The classical permanganate ion: Still a novel oxidant in organic chemistry, Synthesis (Germany) 1987 (1987) 85–127. https://doi.org/10.1055/S-1987-27859.

[3] S. Dash, S. Patel, B.K. Mishra, Oxidation by permanganate: synthetic and mechanistic aspects, Tetrahedron 65 (2009) 707–739. https://doi.org/10.1016/J.TET.2008.10.038.

[4] K.B. Wiberg, R. Stewart, The Mechanisms of Permanganate Oxidation. I. The Oxidation of Some Aromatic Aldehydes, J. Am. Chem. Soc. 77 (1955) 1786–1795. https://doi.org/10.1021/JA01612A020.

[5] U.A. Spitzer, Mechanism of permanganate oxidation of alkanes, arenes and related compounds, (1972). https://doi.org/10.14288/1.0060242.

[6] U.A. Spitzer, Mechanism of permanganate oxidation of alkanes, arenes and related compounds, (1972). https://doi.org/10.14288/1.0060242.

[7] A.J. Shakir, G. Ionita, P. Ionita, SELECTIVE OXIDATION OF ALCOHOLS WITH STABLE ORGANIC POLYRADICALS, 2017. http://web.icf.ro/rrch/.

[8] N. V. Lawale, B.S. Dobhal, S.S. Gaur, R. Pardeshi, Permanganatic oxidation of neomycine sulphate salt in basic media: a kinetic study, Journal of Drug Delivery and Therapeutics 9 (2019) 159–164. https://doi.org/10.22270/jddt.v9i1.2201.

[9] J.P. Ferraz, L. do Amaral, Acidity Constants of Some 5-Substituted 3-Furoic Acids, Journal of Organic Chemistry 41 (1976) 2350–2352. https://doi.org/10.1021/JO00875A030.

[10] 3-Methyl-2-furoic acid (CAS 4412-96-8) | Pharma Intermediate | BenchChem, (n.d.). https://www.benchchem.com/es/product/b189696 (accessed July 11, 2026).

[11] United States Patent (19) Majoie 54 PROCESS FOR THE PREPARATION OF 2,3-DIBROMFURAN 30 Foreign Application Priority Data, (1969).

[12] S. Bakili, T. Kivevele, B. Kichonge, A.A. Salifu, C.K. King'ondu, Furfural from lignocellulose biomass a comprehensive review of hydrolysis methods production technologies and integration into the circular economy, Discover Sustainability 6 (2025). https://doi.org/10.1007/S43621-025-01644-5.

[13] J.A.C. Broekaert, Daniel C. Harris: Quantitative chemical analysis, 9th ed., Anal. Bioanal. Chem. 407 (2015) 8943–8944. https://doi.org/10.1007/S00216-015-9059-6.

[14] D.A. Skoog, S. University, D.M. West, S. Jose, S. College, W. Holt, Principles of instrumental analysis (Skoog, Douglas A.), J. Chem. Educ. 49 (1972) A362. https://doi.org/10.1021/ED049PA362.1.

[15] Arrhenius equation, The IUPAC Compendium of Chemical Terminology (2008). https://doi.org/10.1351/GOLDBOOK.A00446.

[16] S. Arrhenius, S. Arrhenius, Über die Reaktionsgeschwindigkeit bei der Inversion von Rohrzucker durch Säuren, Zeitschrift Für Physikalische Chemie 4U (1889) 226–248. https://doi.org/10.1515/ZPCH-1889-0416.

[17] A.L. Bacharach, F.A. Robinson, M. Sc, F.I. Tech, The Theory of Rate Processes (Glasstone, Samuel; Laidler, Keith J.; Eyring, Henry), J. Chem. Educ. 19 (1942) 249. https://doi.org/10.1021/ED019P249.1.

[18] HarperCollins Publishers: World-Leading Book Publisher, (n.d.). https://www.harpercollins.com/ (accessed July 11, 2026).

[19] F. Freeman, J.B. Brant, N.B. Hester, A.A. Kamego, M.L. Kasner, T.G. McLaughlin, E.W. Paull, Permanganate oxidations. III. Kinetics and mechanisms of the oxidation of furfurals in alkaline media, Journal of Organic Chemistry 35 (2002) 982–985. https://doi.org/10.1021/JO00829A025.

[20] G.S. Chaubey, B. Kharsyntiew, M.K. Mahanti, Oxidation of substituted 2-furaldehydes by quinolinium dichromate: A kinetic study, J. Phys. Org. Chem. 17 (2004) 83–87. https://doi.org/10.1002/POC.698.

[21] S. Dash, S. Patel, B.K. Mishra, Oxidation by permanganate: synthetic and mechanistic aspects, Tetrahedron 65 (2009) 707–739. https://doi.org/10.1016/J.TET.2008.10.038.

[22] L. Hu, H.M. Martin, T.J. Strathmann, Oxidation

Kinetics of Antibiotics during Water Treatment with Potassium Permanganate, Environ. Sci. Technol. 44 (2010) 6416–6422. https://doi.org/10.1021/ES101331J.

[23] G.S. Chaubey, B. Kharsyntiew, M.K. Mahanti, Oxidation of substituted 2-furaldehydes by quinolinium dichromate: A kinetic study, J. Phys. Org. Chem. 17 (2004) 83–87. https://doi.org/10.1002/POC.698;PAGE:STRING:ARTICLE/CHAPTER.