Hydrothermal Carbon−Carbon Bond Formation and Disproportionations of C1 Aldehydes: Formaldehyde and Formic Acid

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Hydrothermal reaction pathways and kinetics of C1 (carbon-one) aldehydes, formaldehyde (HCHO) and formic acid (HCOOH = HOCHO), are studied at 225 °C without and with hydrochloric acid (HCl) up to 0.6 M (mol dm^{-3}). Reactions unveiled are the following: (i) the self-disproportionation forming methanol and formic acid, a redox reaction between two formaldehydes, (ii) the cross-disproportionation forming methanol and carbonic acid, a redox reaction between formaldehyde and formic acid, and (iii) the acid-catalyzed C−C bond formation producing glycolic acid ($HOCH_2COOH$) as a precursor of the simplest amino acid, glycine. Reaction iii is a hydrothermally induced chemical evolution step from C1 aldehydes, formaldehyde and formic acid. Disproportionations i and ii are found to proceed even without base catalysts unlike the classical Cannizzaro reaction. Acid catalyzes the self-disproportionation (i) and the C−C bond formation (iii), but retards the cross-disproportionation (ii). The rate constants of noncatalyzed and acid/base-catalyzed paths for reactions i, ii, and iii are given additively as $2 \times 10^{-4} + (2 \times 10^{-3})[H^+]$, $10^{-4} + 10^3[OH^-]$, and $(2 \times 10^{-3})[H^+]$ M^{-1} s^{-1}, respectively; the concentrations of proton [H$^+$] and hydroxide ion [OH$^-$] are expressed in M. The rate constant of the noncatalytic (neutral) cross-disproportionation is 1 order of magnitude larger than that of the self-disproportionation. The reaction pathways are controlled on the basis of the kinetic analysis to make the glycolic acid and methanol productions dominant by tuning the concentrations of formaldehyde, formic acid, and HCl. The conversion to glycolic acid reaches ∼90% when formaldehyde, HCl, and formic acid are mixed in the ratio of 1:2:17. The conversion of formaldehyde to methanol reaches ∼80% when formic acid is added in excess to formaldehyde.

1. Introduction

Chemical evolution prior to the origin of life involves the transformation of simple inorganic molecules into small and simple organics called C1 compounds followed by the buildup of chemical bonds from C1 to C2, C3, ...[1,2] It is not well demonstrated, however, how the elementary steps took place on the primitive earth. Photochemical reactions are considered in one scenario and hydrothermal ones in another.[3,4] To demonstrate a chemical evolution process in hot water without metal catalysts, here we focus on the hydrothermal reactions of such C1 aldehydes as formaldehyde and formic acid (hydroxyl formaldehyde); these species are astronomically available as interstellar molecules. The chemical evolution process from C1 to C2 compound we find here is important in that the C−C bond formation proceeds uniquely in hot water without any organic solvents or metal catalysts. Learning from the chemical evolution processes, we attempt to develop the hydrothermal C1 chemistry as a clean and earth-friendly process for approaching the energy and environmental issues.

Current energy issue is due to the mismatch of the time scales between the cycles of the fossil fuel production of nature and the energy consumption of mankind. One of the important approaches to this issue is to introduce a clean material cycle process for creating a renewable energy resource like methanol. The new recycle path should be rapid enough to match with the current energy consumption and have the possibility of liberating us from relying upon the slow fossil fuel production in nature. Coupled with the energy problem, utilization of C1 reactions as a future source of fuel has been the major topic. Such typical reactions are associated with synthesis gas (a mixture of CO and H_2),[5,6] which is obtained from coal, natural gas, crude oil, biomass, and organic wastes.[7−9] The main purpose has been the pursuit of catalysts with a high yield and selectivity to synthesize alcohols, aldehydes, olefins, transportation fuels, and so on, as represented by the Fischer−Tropsch reaction.[9] Apart from the conventional C1 chemistry, here we expand C1 chemistry into a hydrothermal regime and develop the clean reaction of C1 aldehydes to produce alcohol and hydroxy carboxylic acid without any organic solvents or metal catalysts.

One of the most important targets of C1 chemistry is to transform C1 into C2 compound. Conventionally, C−C bond formation has been performed by the popular Grignard, Friedel−Crafts, and Diels−Alder reactions.[10] These reactions, however, require severe conditions or are restricted to particular pairs of donor and acceptor. Here we report that a new C−C bond can be formed between formaldehyde and formic acid in acidic hot water; such conditions can be found in the deep sea.[11] In our hydrothermal method, neither organic solvents nor metal catalysts are used. The hydrothermal method is applicable to an aldehyde to obtain α-hydroxy carboxylic acid that can be aminated to the corresponding amino acid.[1,12] The C−C bond formation in the reaction temperature range of 200−250 °C has been reported in previous short communications; glycolic acid can be formed from such formaldehyde producers as *s*-trioxane, formalin, paraformaldehyde, and dichloromethane.[1,12] In this work, we perform the kinetic analysis of the C−C bond

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formation to optimize the glycolic acid yield by controlling the reaction pathways. The reaction temperature is fixed at 225 °C so that the reaction can be quantitatively analyzed in the time scale of minutes.

To control the pathways, it is indispensable to kinetically analyze the disproportionations which compete with the C−C bond formation. As previously found,[1,12−14] C1 aldehydes undergo the self- and cross-disproportionations in hot water without base catalysts. In the bimolecular self-disproportionation, one formaldehyde molecule is reduced to methanol and the other oxidized to formic acid. In the cross-disproportionation, formic acid (hydroxyl formaldehyde) reduces formaldehyde to methanol, being oxidized itself to carbonic acid. The noncatalytic hydrothermal disproportionations are novel as a method for producing alcohol in a manner friendly to earth. In this study, we have investigated the pathways and kinetics for hydrothermal reactions between formaldehyde and formic acid in detail by applying ^{1}H and ^{13}C NMR spectroscopy to all of the products in the gas and liquid phases. We substantiate the pathway control between the glycolic acid and methanol formations as a step toward green C1 chemistry.

The experimental procedure is in Section 2. In Section 3.1, we discuss the reaction scheme on the basis of the product distribution at a fixed reaction time. In Section 3.2, the hydrothermal disproportionations and the C−C bond formation are kinetically analyzed to determine their rate constants. On the basis of these rate constants, we perform the pathway control in Section 3.3. Conclusions are given in Section 4.

2. Experimental Section

Materials. 1,3,5-Trioxane (a formaldehyde trimer, called *s*-trioxane hereafter), HCl (2 M), and formic acid (99%) were obtained from Nacalai and used without further purification. CH_2Cl_2 was also purchased from Nacalai and was washed with distilled water to remove methanol contained as a polymerization inhibitor. ^{13}C-enriched formalin (99 atom % and 20% in H_2O) and deutrated formic acid (DCOOH, 99 atom % and 95% in H_2O) were obtained from ISOTEC and were used as received.

When the starting material was *s*-trioxane or ^{13}C-enriched formalin, it was sealed in a quartz NMR tube of 2.2 mm i.d. and 3.0 mm o.d. under argon atmosphere with aqueous HCl solution under a variety of experimental conditions shown in Table 1. Both *s*-trioxane and formalin produce formaldehyde, and their concentrations were adjusted so that they may provide monomeric formaldehyde of 0.3 M (ambient). The sample filling factor, defined as the volume ratio of the solution to the total sample tube, was set to 71%. When CH_2Cl_2 was the starting material, it was sealed together with water in a quartz NMR tube under argon atmosphere. The concentration was set to 1.0 M when the reaction mixture becomes homogeneous. The filling factor of this sample was set to 85% at room temperature to make the gas-phase volume negligible at the reaction temperature.

Procedure. Each sample tube was heated in a programmable electric furnace kept at the reaction temperature of 225 ± 1 °C. After a desired reaction time, the sample was removed from the furnace and cooled to room temperature by air within a minute. The reaction time was fixed at 2 h in the analysis of the reaction scheme. After the reaction, the sample was put in a Pyrex NMR tube, and the ^{1}H and proton-decoupled ^{13}C NMR spectra were measured with 400, 500, and 600 MHz NMR (JEOL). In the reaction tube, the liquid and gas phases coexist and the measurements of both phases are necessary to identify all the products. The liquid and gas phases were separately measured as described elsewhere.[15]

TABLE 1: Reaction Conditions and the Concentrations[a] of the Reactant at 225 °C

formaldehyde generators	aldehyde[b] (M)	HCl (M)	HCOOH (M)	kinetics[c] A/B/F
s-trioxane	0.3	0	0	A
s-trioxane	0.3	0.15	0	B
s-trioxane	0.3	0.225	0	B
s-trioxane	0.3	0.3	0	B
s-trioxane	0.3	0.6	0	F
s-trioxane	0.3	0.15	0.3	A
s-trioxane	0.3	0.225	0.3	A
s-trioxane	0.3	0.3	0.3	A
s-trioxane	0.3	0.5	0.3	A
s-trioxane	0.3	0	2.0	F
s-trioxane	0.3	0.6	1.0	F
s-trioxane[d]	0.3	0.6	3.0	F
s-trioxane[d]	0.3	0.3	5.0	F
s-trioxane[d]	0.3	0.6	5.0	F
^{13}C formalin[d]	0.3	0.6	0	F
^{13}C formalin[d]	0.3	0	2.0	F
^{13}C formalin[d]	0.3	0.6	5.0	F
dichloromethane	1.0	0	0	A

[a] The concentrations shown here are those at room temperature. [b] The aldehyde concentrations are those after the conversion of reactant into monomer. [c] In method A, the reaction progress was studied by the repetition of heating, cooling, and NMR measurement of a single sample, and in method B, by assigning one reaction time to one sample. The letter F indicates that samples are analyzed only at a fixed reaction time of 2 h. [d] Quartz NMR tubes of 1.5 mm i.d. and 3.0 mm o.d. are used because of pressure.

The self- and cross-disproportionations and the glycolic acid formation were kinetically analyzed under the conditions summarized in Table 1. For the cross-disproportionation and the C−C bond formation, the reaction time evolution was observed at time intervals of 10−15 min by the repetition of heating, cooling, and measuring of the same sample. The results were averaged for two reaction samples. For the self-disproportionation, the reaction was determined at the time step of 1−2 min. This is because the self-disproportionation can be selectively picked up and analyzed only in the early stage of reaction with *s*-trioxane and HCl. When the reaction time is as short as 1 min, the progress of the reaction cannot be accurately traced by the repetition of heating, cooling, and measuring of a single sample. In this case, a series of samples were used as follows: since it takes about 1 min until the sample attains the desired reaction temperature due to the low thermal conductivity of quartz, this "dead time" affects the determined rate constants. In the study of the self-disproportionation, therefore, a number of samples were prepared for each reaction mixture and every sample was assigned to a different reaction time. The reaction time was set to 2 min or longer; 2 min is somewhat longer than the estimated dead time (∼1 min). Then, the time evolution of product concentrations was obtained by subtracting the product concentration of the 2-min sample from that of every other reacted sample. The time-scale normalization treatment cancels out the dead time from the time evolution of product concentrations.

3. Results and Discussion

It is important to control the reaction pathways of the C−C bond formation and the disproportionations in hot water to a desirable direction. We elucidate the reaction scheme and kinetics on the basis of the reaction products, the distribution, and the time evolution. In doing so, we have comprehensively applied ^{1}H and ^{13}C NMR spectroscopy to all of the products not only in the liquid but also in the gas phase. The time

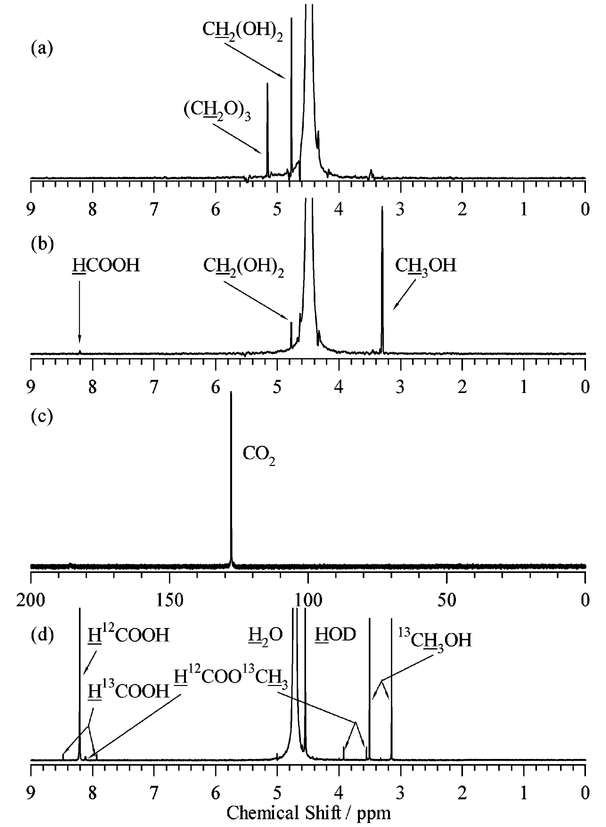

Figure 1. (a) ^{1}H NMR spectrum of the liquid phase at 30 °C for the experiment with 0.1 M *s*-trioxane after 15 min of reaction at 225 °C. (b) ^{1}H NMR spectrum of the liquid phase for the experiment with 0.1 M *s*-trioxane after 18 h of reaction at 225 °C. (c) Proton-decoupled ^{13}C NMR spectrum of the gas phase for the experiment with 0.3 M ^{13}C formalin after 18 h of reaction at 225 °C. (d) ^{1}H NMR spectrum of the liquid phase for the experiment with 0.3 M ^{13}C formalin and 2.0 M HCOOH after 2 h of reaction at 225 °C. The peaks of $H^{13}COOH$ are the satellite of added $H^{12}COOH$. $H^{12}COO^{13}CH_3$ is the byproduct due to the esterification between $^{13}CH_3OH$ and $H^{12}COOH$.

evolution of all the reaction products is monitored by the structural, elemental, and quantitative analysis due to the ^{1}H and ^{13}C NMR.

3.1. Products Analysis and Reaction Scheme. We have elucidated the reaction scheme of C1 aldehydes in the hydrothermal condition on the basis of the product species and distributions. Here we show that the following reactions take place in hot water: self-disproportionation, cross-disproportionation, and acid-catalyzed C−C bond formation producing glycolic acid ($HOCH_2COOH$).

3.1.1. Hydrothermal Disproportionations of Formaldehyde. First let us see what products are generated by the hydrothermal reactions of formaldehyde at 0.3 M (0.1 M *s*-trioxane) in neutral conditions at 225 °C. Figure 1a shows the liquid-phase ^{1}H NMR spectrum at an early reaction time of 15 min. Before the disproportionations proceed, methanediol is detected as a reactive intermediate, only this is formed by the hydration of formaldehyde produced from *s*-trioxane.[14,16−18] After 18 h of reaction, methanol is found to form dominantly in neutral hydrothermal condition as indicated by the liquid-phase ^{1}H NMR spectrum in Figure 1b and the gas-phase ^{13}C NMR spectrum in Figure 1c. Product yields are shown in the decreasing order

$$CH_3OH\ (60) > CO_2\ (28) > CH_2(OH)_2\ (11) > HCOOH\ (1)$$

where the numbers in parentheses are the product yield (%). Methanol is considered to be generated by the reduction of formaldehyde, and formic acid by the oxidation.[19] These products can be generated as follows:

$$2HCHO + H_2O \rightarrow CH_2(OH)_2 + HCHO \rightarrow CH_3OH + HCOOH \quad (1)$$

This is the aldehyde self-disproportionation, a redox reaction between aldehyde molecules of the same kind. It is significant that hot water alone can reduce formaldehyde in such a high yield without base catalysts. Despite the common product species, the absence of base catalysts differentiates the hydrothermal self-disproportionation from the classical Cannizzaro reaction, the base-catalyzed disproportionation in ambient conditions. Even in acidic condition, furthermore, the hydrothermal self-disproportionation proceeds.

If methanol is produced only by the self-disproportionation, the yields of methanol and formic acid should be equal and at most 50% according to eq 1. The yield of methanol in Figure 1b, however, exceeds 50% and is ∼60 times larger than that of formic acid. This clearly indicates that the methanol formation in hot water involves some other reaction pathways than eq 1. The difference in the amount of methanol and formic acid formed is twice as much as the amount of carbon dioxide. This indicates the presence of methanol formation expressed as:

$$HCHO + HCOOH + H_2O \rightarrow CH_2(OH)_2 + HCOOH \rightarrow CH_3OH + HOCOOH \rightarrow CH_3OH + CO_2 + H_2O \quad (2)$$

This is the aldehyde cross-disproportionation, a redox reaction between aldehyde molecules of different kinds. Formic acid, the oxidized form of formaldehyde, still remains a member of the C1 aldehyde family.

We should confirm that formic acid reduces formaldehyde as in eq 2. To verify this, we have reacted ^{13}C-enriched formalin (formaldehyde) with an excess amount of ordinary formic acid. We can scrutinize the cross-disproportionation between ^{13}C-labeled formalin and ^{12}C formic acid added in excess; the methanol formation from the self-disproportionation can be avoided in this condition. The formation of ^{13}C-enriched methanol thus means the reduction of formaldehyde by formic acid; instead, the formation of ^{13}C-enriched formic acid and ^{12}C methanediol means the other way around. If eq 2 is the case, therefore, ^{13}C-enriched methanol should be dominant. As indicated by Figure 1d, the major product is ^{13}C-enriched methanol. ^{13}C-enriched formic acid is produced to a negligible amount; the yield is less than 3% as seen in Table 2. This is a clear indication that formic acid reduces formaldehyde as expressed by eq 2.

Furthermore, it is important to know whether the proton used for the reduction is directly transferred from formic acid to formaldehyde. To clarify this, formaldehyde is reacted with an excess amount of ^{2}H-enriched formic acid (DCOOH). The major product is deuterated methanol (CDH_2OH). It is clearly indicated that formic acid reduces formaldehyde by transferring proton directly attached to the carbonyl group. In particular, the redox cross-disproportionation between aldehyde and formic acid is useful as a new and green way of alcohol preparation because aldehyde can be transformed into the corresponding alcohol in such a high degree larger than 50%, the upper limit of the self-disproportionation. We have achieved a yield of 80% when formaldehyde and formic acid are reacted in the ratio of 1:7 as seen in Table 2. By an excessive addition of formic acid, we can even expect the perfect conversion to alcohol. Thus various

TABLE 2: The Product Distribution of Experiments with ^{13}C-Labeled Formaldehyde or ^{2}H-Labeled Formic Acid after 2 h of Reaction at 225 °C[a]

products[b]	0.3 M ^{13}C formalin 2.0 M HCOOH	0.1 M *s*-trioxane 2.0 M DCOOH	0.3 M ^{13}C formalin 6.0 M HCl 5.0 M HCOOH
$HO^{13}CH_2{}^{13}COOH$	0	0	4×10^{-3}
$HO^{13}CH_2{}^{12}COOH$	0	0	2.6×10^{-1}
$^{12}CH_3OH$	2×10^{-3}	1.9×10^{-2}	0
$^{13}CH_3OH$	2.3×10^{-1}	0	3.1×10^{-2}
$^{12}CDH_2OH$	0	1.5×10^{-1}	0
$CH_2(OH)_2$	2.3×10^{-2} ($^{13}CH_2(OH)_2$)	7.0×10^{-2} ($^{12}CH_2(OH)_2$)	0
$H^{12}COOH$	1.5	2.6×10^{-2}	1.5
$H^{13}COOH$	9×10^{-3}	0	3.5×10^{-2}
$H^{12}COO^{13}CH_3$	3.0×10^{-2}	0	7×10^{-3}
$D^{12}COO^{12}CDH_2$	0	2.7×10^{-2}	0
$D^{12}COO^{12}CH_3$	0	3×10^{-3}	0
$^{13}CH_3Cl$	0	0	4×10^{-3}

[a] The product distribution is expressed in concentration (M) at room temperature. [b] Only the liquid phase is analyzed. CO and CO_2 in the gas phase are not detected.

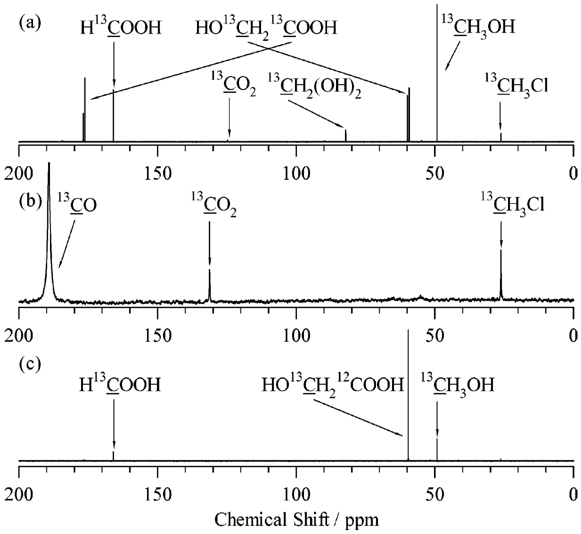

Figure 2. (a) Proton-decoupled ^{13}C NMR spectrum of the liquid phase for the experiment with 0.3 M ^{13}C formalin and 0.6 M HCl after 2 h of reaction at 225 °C. The ratio of integration of doublet at 60 ppm is 1:1, and so is the doublet at 176 ppm. These peaks appear as doublets due to $^{13}C-^{13}C$ coupling in glycolic acid. (b) The proton-decoupled ^{13}C NMR spectrum of the gas phase for the experiment with 0.3 M ^{13}C formalin and 0.6 M HCl after 2 h of reaction at 225 °C. (c) The proton-decoupled ^{13}C NMR spectrum of the liquid phase for the experiment with 0.3 M ^{13}C formalin, 0.6 M HCl, and 5.0 M $H^{12}COOH$ after 2 h of reaction at 225 °C.

alcohols can be obtained in a way friendly to earth through the reduction of corresponding aldehydes using formic acid, an oxidized C1 compound.[13−15,20−22]

3.1.2. Acid-Catalyzed C−C Bond Formation. Now we discuss how the product species and the distribution are modified by the presence of acid, HCl. When the pH is changed to an acidic condition, glycolic acid is found to be formed intermolecularly with a new C−C bond from C1 compounds. This clearly shows that the addition of HCl opens a new reaction path of hydrothermal chemical evolution. As illustrated by panels a and b of Figure 2, the hydrothermal reaction of formaldehyde (^{13}C-enriched formalin, 0.3 M) generates the following products in the presence of 0.6 M HCl for 2 h:

$$HOCH_2COOH\ (43) > CH_3OH\ (32) > CH_2(OH)_2\ (8) > HCOOH\ (7) > CO\ (5) > CH_3Cl\ (4) > CO_2\ (1)$$

where the numbers in parentheses (%) indicate the product distribution based on carbon atom. Glycolic acid is considered to be formed through the reaction

$$HCHO + HCOOH + H_2O \rightarrow CH_2(OH)_2 + HCOOH \xrightarrow[HCl]{} HOCH_2COOH + H_2O \quad (3)$$

Here formic acid is provided through the self-disproportionation of formaldehyde (eq 1) as discussed above. As seen in Figure 2a, the formation of the C−C bond in glycolic acid is identified by the signals of methylene carbon at ∼60 ppm and of carbonyl carbon at ∼176 ppm; the signals are split into the doublet due to the $^{13}C-^{13}C$ coupling. No glycolic acid is formed in the absence of HCl despite the presence of the weak acid, HCOOH.[23] In fact, the yield of glycolic acid decreases dramatically (by ∼60%) when the HCl concentration is lowered from 0.6 to 0.3 M. When formic acid is mixed in excess to formaldehyde, a high yield of glycolic acid is attained. When 0.3 M formaldehyde (0.1 M *s*-trioxane) and 5.0 M formic acid are treated in the presence of 0.6 M HCl, the yield of glycolic acid reaches ∼90%. The optimized yield is ∼4 times larger than that in the absence of added formic acid.

It is worthwhile to consider the mechanism of the new C−C bond formation. Let us examine whether the carbonyl carbon of glycolic acid originates from formic acid or not. As discussed above, glycolic acid with 100% $^{13}C-^{13}C$ bond is formed from ^{13}C-enriched formalin (formaldehyde); note that formic acid is supplied from the self-disproportionation of ^{13}C-enriched formalin. To elucidate the origin of each carbon in the C−C bond formation, we distinguish carbon atoms by labeling formalin (formaldehyde) with ^{13}C, and mix it with an excess amount of ^{12}C formic acid. With the excess amount of ^{12}C formic acid, ^{13}C-enriched formaldehyde (formalin) should give rise to the $^{13}C-^{12}C$ bond in glycolic acid. Comparison of panels a and c in Figure 2 shows that the composition of glycolic acid obtained is $HO^{13}CH_2{}^{12}COOH$. The presence of the $^{13}C-^{12}C$ bond is evidenced by the loss of the doublet of the carbonyl carbon of glycolic acid at 176 ppm and the coupling of the methylene carbon at 60 ppm. In the C−C bond formation, thus, glycolic acid carries the carboxyl group from formic acid and the methylene group from methanediol. The glycolic acid formation from formaldehyde and formic acid we show is an innovative reaction, which proceeds in mild condition without any high-pressure gas or metal catalysts.[24−26] The reaction mechanism may be similar to that of aldol condensation or the Friedel−Crafts reaction.[10]

Here we show how the C−C bond formation process from formaldehyde can be applied to a hydrothermal recycling process

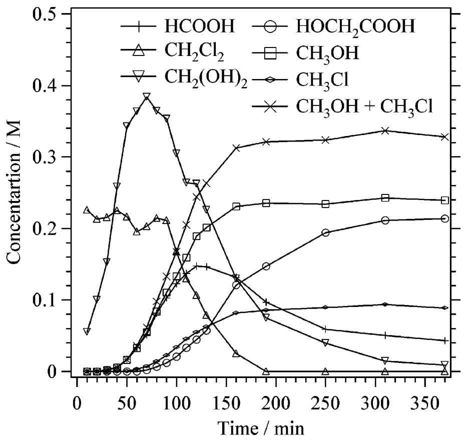

Figure 3. The time evolution of reactant and product concentrations for the reaction of 1 M CH_2Cl_2 at 200 °C. There appears to be a large amount of missing mass in the early stage of reaction because the sample was measurement by NMR at room temperature. Until the concentration of the remaining CH_2Cl_2 becomes lower than ~0.2 M, the reaction mixture is separated into two phases due to the low solubility of CH_2Cl_2 in water, and the decrease of CH_2Cl_2 cannot be detected by the NMR measurement at room temperature.

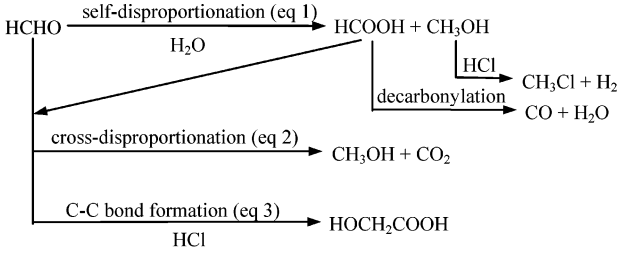

Figure 4. The pathways of the hydrothermal reaction of C1 aldehydes at 225 °C.

TABLE 3: The Product Distribution of 0.3 M Formalin and 0.6 M HCl after 2 h of Reaction at 225 °C

products	distribution (% C)[a]
$HOCH_2COOH$	43
CH_3OH	32
$CH_2(OH)_2$	8
HCOOH	7
CO	5
CH_3Cl	4
CO_2	1
total	100

[a] The mass balance is maintained within 5% before and after the reaction. The distribution is normalized by setting the total product concentration to 100%.

of CH_2Cl_2. The hydrothermal treatment of CH_2Cl_2 leads to the spontaneous formation of a C−C bond in glycolic acid without any added catalysts in a temperature range of 200−250 °C.[12] The glycolic acid formation from CH_2Cl_2 is a one-pot reaction. This reaction can be explained by the same reaction mechanism as the acidic hydrothermal reaction of formaldehyde. Figure 3 shows the time evolution of reactant and product concentrations in the experiment of 1 M CH_2Cl_2 at 200 °C. The concentration of CH_2Cl_2 is set to be higher than that of formaldehyde in the previous subsections for the purpose of detoxication or dechlorination of CH_2Cl_2; the larger the amount of decomposed CH_2Cl_2 the higher the applicability of this reaction. As seen in Figure 3, methanediol is formed by the hydrolysis of CH_2Cl_2:

$$CH_2Cl_2 + 2H_2O \rightarrow CH_2(OH)_2 + 2HCl \qquad (4)$$

Though the formation of HCl cannot be detected directly by NMR, it is indicated by the peak shift of water observed during the reaction due to the variation of pH. Since the hydration of CH_2Cl_2 simultaneously prepares methanediol and HCl, glycolic acid forms spontaneously from CH_2Cl_2. After the methanediol formation, methanol and formic acid start to form by the self- and cross-disproportionations (eqs 1 and 2). The concentration of formic acid increases until 2 h, and after reaching the maximum, it begins to decrease. Corresponding to the behavior of formic acid concentration, the glycolic acid concentration gradually increases. This correlation among the methanol, formic acid, and glycolic acid concentrations indicates the glycolic acid formation from formaldehyde and formic acid. CH_2Cl_2 reaction is significant because we can convert CH_2Cl_2, a hazardous chlorinated organic compound, to glycolic acid, a useful and recyclable organic compound, by a simple hydrothermal treatment.

3.1.3. Reaction Scheme Based on Mass Balance. The reaction scheme can be constructed as shown in Figure 4 according to the abovementioned NMR analysis of the reaction products and mechanisms.[27] Before going to the kinetic study, here we check the reliability of the analysis scheme on the basis of the mass balance. In Table 3, we show the distribution of all the products including both liquid and gaseous ones after the reaction of 0.3 M formaldehyde (^{13}C-enriched formalin) and 0.6 M HCl. The decarbonylation of formic acid to CO and H_2O is minor under the present reaction conditions of concentration, pH, time, and temperature; the yield of CO is only 4.5% in terms of carbon atom percentage.[5,6] CH_3Cl is another minor (3.5%) product of the chlorination of methanol. When the reaction scheme is valid, the following mass balance equation is to be satisfied:

$$([CH_3OH] + [CH_3Cl] - [CO_2]) - ([HCOOH] + [CO_2] + [CO]) = [HOCH_2COOH] \qquad (5)$$

In fact, the left- and right-hand sides of this equation are equal within an error of 8%. It is concluded therefore that there are no other important pathways than those shown in Figure 4.

3.2. Kinetics and Rate Constants. In this section, we perform the kinetic analysis of the C−C bond formation and the self- and cross-disproportionations to elucidate the weights of these reactions for the path control. The rate constants of these paths have been determined by analyzing the time evolution of the relevant products on the basis of the rate laws. The kinetic analysis of the complicated reaction scheme is possible because hydrochloric acid promotes the C−C bond formation and the self-disproportionation, and because it retards the cross-disproportionation.

3.2.1. Hydrothermal Disproportionations in Neutral Condition. Panels a and b in Figure 5 show the time evolution of the reactant and product concentrations when 0.3 M formaldehyde (0.1 M *s*-trioxane) is reacted; the short-time region of Figure 5a is expanded in Figure 5b. As shown in Figure 1a, first the *s*-trioxane ring opens to be decomposed into the hydrated monomers (methanediol). It takes ~40 min for all the *s*-trioxane to be converted into methanediol; as seen in Figure 5c, the hydrolysis is completed within 5 min in the presence of HCl at 0.3 M. In the neutral condition, the main product is methanol as shown in Figure 5a. After 30 h of reaction, the yield of methanol reaches ~60% whereas formic acid yield is essentially low throughout the reaction: it is 7% or lower. Thus the cross-disproportionation overwhelms the self-disproportionation in neutral hot water.

The observed rate constants of the self- (SD) and cross-disproportionations (CD) in neutral (N) hydrothermal condition, k_{SDN} and k_{CDN}, are determined to be (1.6 ± 0.6) × 10^{-4} and

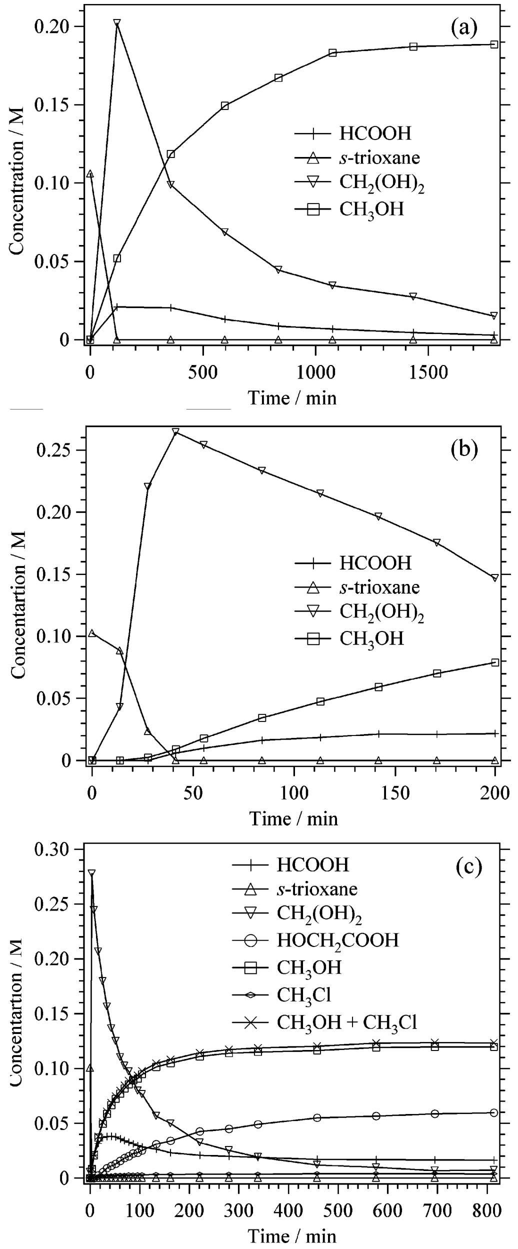

Figure 5. (a) The time evolution of reactant and product concentrations for the reaction of 0.1 M *s*-trioxane at 225 °C. (b) The expanded figure of panel a in a short-time region. (c) The time evolution of reactant and product concentrations for the reaction of 0.1 M *s*-trioxane and 0.3 M HCl at 225 °C.

$(1.3 \pm 0.2) \times 10^{-3}$ M^{-1} s^{-1}, respectively, as summarized in Table 4; see Appendix A for the rate constant determination. It is to be noted that the observed value of k_{CDN} is 1 order of magnitude larger than that of k_{SDN}. This explains why formic acid decreases after 360 min of reaction in Figure 5a: the cross-disproportionation proceeds faster than the self-disproportionation in neutral condition despite the low concentration of formic acid (cf. eqs 1 and 2).

It is of great interest that formic acid, the oxidized form of formaldehyde, still has a reducing ability much higher than that of formaldehyde itself; also in the case of other aldehydes, the cross-disproportionation proceeds faster than the self-disproportionation in neutral hydrothermal condition.[15,22] Formic acid thus acts as an effective reducer for aldehyde. Various alcohols can be produced by the hydrothermal cross-disproportionation, which proceeds in the absence of any added catalysts.

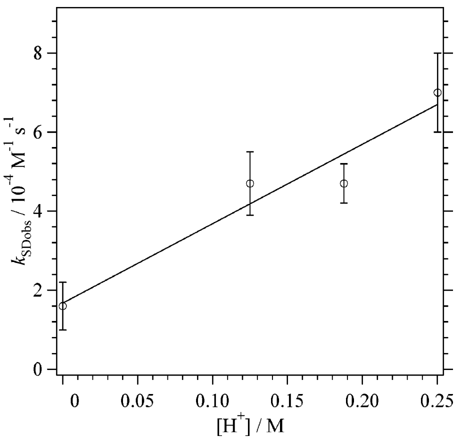

Figure 6. The plot of k_{SDobs} against the H^+ concentration. The H^+ concentrations are those in the hydrothermal condition.

3.2.2. Hydrothermal Disproportionations in Acidic Conditions. To elucidate the effect of the strong acid, we determine the rate constants of the self- and cross-disproportionations in the presence of HCl at concentrations of 0.15−0.3 M; see Appendix A for the details of the rate constant determination. Table 4 lists the rate constants of the self- and cross-disproportionations with and without HCl at 225 °C. The acid catalyzes the self-disproportionation, whereas it retards the cross-disproportionation.[29] This contrasts sharply with the common notion of the classical Cannizzaro reaction, which is catalyzed only by strong base in ambient condition.

To reveal the effect of acid more precisely, we divide k_{SDobs}, the observed rate constant of the self-disproportionation in acidic condition, into $k_{SDwater}$, the rate constant in the absence of H^+, and k_{SDacid}, the rate constant of the acid-catalyzed path.[30] When the reaction order of the proton concentration $[H^+]$ is unity,[31] k_{SDobs} is expressed as:

$$k_{SDobs} = k_{SDwater} + k_{SDacid}[H^+] \tag{6}$$

The first term of the right-hand side indicates that the self-disproportionation in neutral condition is induced by the undissociated form of water; we call this reaction the water-induced path of self-disproportionation.[32] The value of $k_{SDwater}$ is obtained by hypothetically setting $[H^+] = 0$. $k_{SDwater}$ is different from k_{SDN} since even in neutral condition, $[H^+]$ is on the order of 10^{-6} M at 225 °C due to the autoprotolysis of water.[33] When HCl is added to the reaction mixture, $[H^+]$ is equal to [HCl] since HCl completely dissociates at 225 °C.[33] Figure 6 shows the plot of k_{SDobs} against $[H^+]$. The values of $k_{SDwater}$ and k_{SDacid} are obtained as $(2 \pm 1) \times 10^{-4}$ M^{-1} s^{-1} and $(2 \pm 1) \times 10^{-3}$ M^{-2} s^{-1}, respectively; see Table 5. The contribution of the acid-catalyzed path is negligible in the neutral hydrothermal condition: $k_{SDacid}[H^+] = 2 \times 10^{-9}$ M^{-1} s^{-1} at

TABLE 4: Observed Rate Constants of the Self- and Cross-Disproportionations and the C−C Bond Formation in Neutral and Acidic Conditions at 225 °C

	neutral	0.15 M HCl	0.225 M HCl	0.3 M HCl	0.5 M HCl
k_{SDobs} (M^{-1} s^{-1})	$(1.6 \pm 0.6) \times 10^{-4}$	$(5 \pm 1) \times 10^{-4}$	$(5 \pm 1) \times 10^{-4}$	$(7 \pm 1) \times 10^{-4}$	
k_{CDobs} (M^{-1} s^{-1})	$(1.3 \pm 0.2) \times 10^{-3}$	$(1 \pm 1) \times 10^{-4}$	$(2 \pm 1) \times 10^{-4}$	$(1 \pm 2) \times 10^{-4}$	
k_{Gobs} (M^{-1} s^{-1})		$(3 \pm 2) \times 10^{-4}$	$(4 \pm 2) \times 10^{-4}$	$(4 \pm 3) \times 10^{-4}$	$(1 \pm 1) \times 10^{-3}$

TABLE 5: The Water-Induced and Acid-Catalyzed Components of the Rate Constants of the Self-Disproportionation and the C−C Bond Formation at 225 °C

	k_{water} (M^{-1} s^{-1})	k_{acid} (M^{-2} s^{-1})
self-disproportionation	$(2 \pm 1) \times 10^{-4}$	$(2 \pm 1) \times 10^{-3}$
C−C bond formation		$(2 \pm 2) \times 10^{-3}$

225 °C. The effect of the neutral water has been observed also in the case of the hydrothermal dehydration of diol to cyclic ether. In the previous study of 1,4-butanediol and tetrahydrofuran, the water-induced path is found to be dominant over the acid-catalyzed path in the neutral condition.[30,34]

3.2.3. Acid-Catalyzed C−C Bond Formation. Before the determination of the rate constant of the glycolic acid formation, let us study how the time evolution of the reactant and product concentrations is modified by the addition of HCl; see Figure 5a−c. Glycolic acid starts to form appreciably only after 20 min of reaction in the presence of HCl at 0.3 M. Methanol and formic acid equally increase in the first 10 min, reflecting the self-disproportionation path (eq 1). Formic acid subsequently decreases around 50 min, corresponding to the consumption of formic acid due to both the cross-disproportionation and the C−C bond formation.

Now let us determine the rate constant of the glycolic acid formation. The C−C bond is formed between hydrated formaldehyde and formic acid as expressed by eq 3. The rate constant for this process is denoted by k_G. The value of k_G has been determined from the observed [HCl]-dependent quantity $k_{\mathrm{Gobs}} = k_G\,[\mathrm{HCl}]^m$; see Appendix A. The reaction order m of HCl concentration is determined as 1, and the value of k_G is approximately obtained to be $(2 \pm 2) \times 10^{-3}$ M^{-2} s^{-1} as summarized in Table 5.

3.3. Reaction Pathway Control. When we apply the hydrothermal reactions of C1 aldehydes for selective production of glycolic acid and methanol, the pathway control becomes of essential importance. In this section, we perform the pathway control by tuning the kinetic variables, such as the reactant concentrations, the pH conditions, and the reaction time under given temperature and rate constants.

3.3.1. Glycolic Acid Formation. Glycolic acid formation can be selected by maximizing the path weight of the C−C bond formation and suppressing the methanol formation in acidic condition through the kinetic variables. Equations A4 and A5 in Appendix A enable us to understand that the C−C bond formation is enhanced by the addition of either HCl or formic acid. The addition of HCl, however, enhances the self-disproportionation as well as the C−C bond formation; cf. eqs 6 and A5. The formation of glycolic acid can be selectively accelerated by the addition of formic acid but not by that of HCl because of the similarity of the value of k_{SDacid} and k_G in Table 5. The selective formation of glycolic acid can be achieved only by addition of formic acid in large excess.

We examine in more detail the reaction conditions for the selective formation of glycolic acid. The validity of the relation

$$\frac{\mathrm{d}[\mathrm{HOCH_2COOH}]}{\mathrm{d}t} > \frac{\mathrm{d}[\mathrm{CH_3OH}]}{\mathrm{d}t} \qquad (7)$$

throughout the reaction is a sufficient condition for the dominant production of glycolic acid. When the inequalities

$$k_G[\mathrm{HCl}] - k_{\mathrm{CD}} > 0 \qquad (8)$$

and

$$\frac{[\mathrm{HCOOH}]_0}{[\mathrm{HCHO}]_0} > \max\left(1, \frac{k_{\mathrm{SDwater}} + k_{\mathrm{SDacid}}[\mathrm{HCl}]}{k_G[\mathrm{HCl}] - k_{\mathrm{CD}}}\right) \qquad (9)$$

are both valid,[35] eq 7 is satisfied over the entire range of reaction time. Here $[\mathrm{HCHO}]_0$ and $[\mathrm{HCOOH}]_0$ denote the initial concentrations of formaldehyde and formic acid, respectively; the verification is shown in Appendix B. According to Tables 4 and 5, eq 8 provides [HCl] > 0.2 M.[36] The initial concentrations of formaldehyde and formic acid can be chosen for a given [HCl] with eq 9. For example, when $[\mathrm{HCHO}]_0 = 0.3$ M, [HCl] = 0.6 M, and $[\mathrm{HCOOH}]_0 > 0.5$ M, both eqs 8 and 9 are satisfied. Consequently, the dominance of glycolic acid can be attained as seen in Figure 7. When the concentration of formic acid is ∼17 times higher than that of formaldehyde in the presence of 0.6 M HCl, the yield of glycolic acid is ∼90%.[37] The hydrothermal synthesis of glycolic acid is highly advantageous in that no environmentally harmful solvent or catalyst is needed and in that the operational procedure is simple. Acidic hydrothermal conditions exist in the ocean, and the C1 to C2 process found here is expected to constitute a molecular step toward the origin of life in the primitive ocean.

3.3.2. Methanol Formation. As can be seen in Figure 4, the methanol yield can be maximized by enhancing two kinds of disproportionations (eqs 1 and 2) without opening the C−C bond formation path. This is achieved by treating the system in the absence of HCl. Since the rate constant of the cross-disproportionation k_{CDN} is 1 order of magnitude larger than that of the self-disproportionation k_{SDN}, the maximum methanol yield can exceed 50%. Theoretically, the maximum methanol yield is achieved when $^2/_3$ of the initial formaldehyde disproportionates to produce methanol and formic acid equally and when all the formic acid produced cross-disproportionates with the remaining formaldehyde ($^1/_3$ of the initial amount). The maximum methanol yield is thus ∼70% ($^2/_3$ of the initial amount of formaldehyde).

The experimental result indicates that the neutral hydrothermal reaction of formaldehyde achieves a methanol yield almost equal to the maximum described above. As shown in Figure 5a, the concentration of formic acid is as low as 20 mM. The cross-disproportionation does proceed faster than the self-disproportionation in neutral condition despite the low formic acid concentration compared to formaldehyde. The methanol yield is 60% after 30 h of reaction, corresponding to the yield of 63% at infinite reaction time. It is striking that a simple hydrothermal treatment reduces more than 60% of formaldehyde into methanol

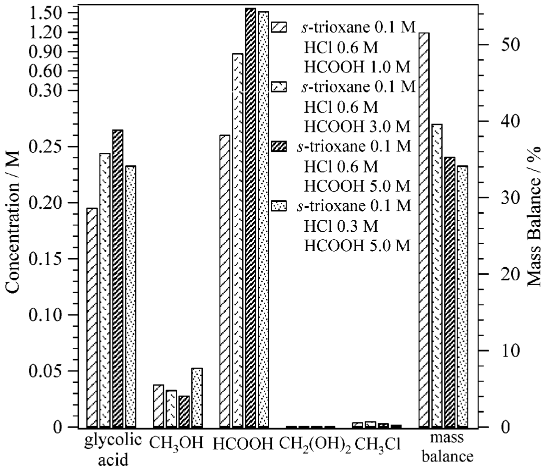

Figure 7. The reactant and product concentrations after 2 h of reaction at 225 °C. The ordinate scale is changed at 0.3 M. The mass balance is based on the liquid-phase ^{1}H NMR.

without any added catalysts. To achieve higher methanol yield or to shorten the reaction time, we only need to add formic acid to aqueous formaldehyde solution. This is experimentally demonstrated; the addition of 2.0 M formic acid, which is 7-fold of formaldehyde in amount, attains ∼80% of methanol yield within 2 h as seen in Table 2.

4. Conclusions

In this work, the hydrothermal C−C bond formation and the disproportionations of formaldehyde and formic acid are studied from the viewpoint of kinetics and path control.

The acid-catalyzed hydrothermal C−C bond formation is a clean chemical evolution, which generates glycolic acid, a C2 compound, from formaldehyde and formic acid, C1 compounds, without the use of organic solvent or metal catalyst. The amination of glycolic acid generates glycine, the simplest amino acid. The hydrothermal C−C bond formation gives a new synthetic method and at the same time it may be a step toward the origin of life on the primitive earth. From the kinetic analysis, the rate constants of the self- and cross-disproportionations and the C−C bond formation are determined. The cross-disproportionation is found to proceed faster than the self-disproportionation in neutral condition, which indicates a stronger reducing ability of formic acid than that of formaldehyde. Acid is found to promote the self-disproportionation. In hot water, the undissociated form of water is found to induce both self- and cross-disproportionations. These are in contrast to the classical Cannizzaro reaction in ambient condition. The cross-disproportionation in neutral condition should be noted as a new method for alcohol production. Simple hydrothermal treatment with formic acid reduces aldehyde into alcohol.

On the basis of the rate constants, the pathway control between glycolic acid and methanol formations is performed. Glycolic acid yield is enhanced up to 90% by tuning the ratio of the input concentrations of formaldehyde, HCl, and formic acid as 1:2:17. A methanol yield of 80% is achieved by reacting formaldehyde and formic acid in the ratio of 1:7.

As illustrated by the hydrothermal recycling of CH_2Cl_2 into glycolic acid and methanol, the hydrothermal reaction of C1 aldehydes is important from synthetic, energy, and environmental concerns. In the present work, we have presented a possible cure for the recent concerns.

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Appendix A: Determination of Rate Constants

First, the rate constants k_{SDN} and k_{CDN} of the self- and cross-disproportionations in neutral condition are determined. In this Appendix, the rate constants with suffix "SD" or "CD" refer to the observed ones. The following equations are derived from eqs 1 and 2.

$$\frac{d[HCHO]}{dt} = -2k_{SDN}[HCHO]^2 - k_{CDN}[HCHO][HCOOH] \quad (A1)$$

$$\frac{d[HCOOH]}{dt} = k_{SDN}[HCHO]^2 - k_{CDN}[HCHO][HCOOH] \quad (A2)$$

As seen in Figure 5b, [HCOOH] varies within 10% and [HCHO] linearly decreases by ∼20% as the reaction time passes from 110 to 200 min. The linear fitting of [HCHO] and [HCOOH] in this time region then gives the rate constants by accounting for the small changes in [HCHO] and [HCOOH].

Second, the rate constants of the self- and cross-disproportionations in acidic condition are determined. When both the self- and the cross-disproportionations take place, the rate equation for the methanol formation is written as

$$\frac{d[CH_3OH]}{dt} = k_{SD0.3}[HCHO]^2 + k_{CD0.3}[HCHO][HCOOH] \quad (A3)$$

where $k_{SD0.3}$ and $k_{CD0.3}$ represent the rate constants of the self- and cross-disproportionations, respectively, in the presence of 0.3 M HCl. As seen in Figure 5c, the progress of methanol and formic acid concentrations are in accord in the first 10 min of the reaction, indicating the progress only of the self-disproportionation. The second term of eq A3 can thus be neglected in this time region. In the very beginning of the reaction, [HCHO] can be considered constant and therefore the rate constant $k_{SD0.3}$ is obtained from the linear fitting of the methanol concentration against the reaction time. The value of $k_{CD0.3}$ is then obtained by using $k_{SD0.3}$ and eq A3. The initial rate of the methanol formation is estimated in the experiment where 0.3 M formaldehyde (0.1 M *s*-trioxane), 0.3 M HCl, and 0.3 M formic acid are reacted. Since [HCHO] and [HCOOH] are considered constant at the early stage of the reaction, the rate constant $k_{CD0.3}$ is obtained by linearly fitting the experimental plot of the methanol concentration against the reaction time. When the HCl concentration is 0.15 or 0.225 M, the rate constants are determined in the same way.

Third, the rate constant of the glycolic acid formation is determined by using the observed rate constant k_{Gobs} as an intermediate parameter. Since the hydration of HCHO to $CH_2(OH)_2$ is fast in eq 3, the rate equation for the glycolic acid formation is written as

$$\frac{d[HOCH_2COOH]}{dt} = k_{Gobs}[HCHO][HCOOH] \quad (A4)$$

The rate constant k_G can be expressed by using k_{Gobs} as

$$k_{Gobs} = k_G[HCl]^m \quad (A5)$$

where *m* is the reaction order of HCl concentration. To obtain k_{Gobs}, 0.3 M formaldehyde (0.1 M *s*-trioxane) and 0.3 M formic acid are reacted with various concentrations of HCl (0.15, 0.225, 0.3, and 0.5 M). The k_{Gobs} is given by linearly fitting the experimental plot of the glycolic acid concentration against the reaction time at the early stage of the reaction since [HCHO] and [HCOOH] are considered constant. Table 4 shows the observed rate constant k_{Gobs} at given HCl concentrations. The value of k_G is obtained by plotting the logarithmic value of k_{Gobs} in Table 4 against that of HCl concentration as shown in Figure 8.

Appendix B: Pathway Control for the Glycolic Acid Formation

Here, we elucidate that eqs 8 and 9 are sufficient for the selective formation of glycolic acid. Glycolic acid is provided

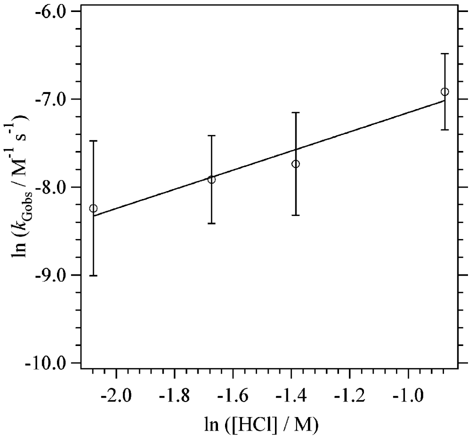

Figure 8. The plot of the logarithm of k_{Gobs} against that of HCl concentration. HCl concentrations are those in the hydrothermal condition.

by the C−C bond formation, and methanol, by the self- and cross-disproportionations. Therefore, eq 7 is rewritten by virtue of eqs A3−A5 as

$$k_G[HCHO][HCOOH][HCl] > k_{SD}[HCHO]^2 + k_{CD}[HCHO][HCOOH] \quad (B1)$$

where k_{SD} and k_{CD} denote the observed rate constants of the self- and cross-disproportionations including the effect of added acid. By substituting ($k_{SDwater}$+ k_{SDacid}[HCl]) for k_{SD}, eq B1 leads to

$$[HCOOH](k_G[HCl] - k_{CD}) > (k_{SDwater} + k_{SDacid}[HCl])[HCHO] \quad (B2)$$

Equation B2 indicates that eq 8 needs to be satisfied because the right-hand side of eq B2 never becomes negative. (k_G[HCl] − k_{CD}) represents the magnitude relation between the rates of the C−C bond formation and the cross-disproportionation. If it is positive, the rate of the C−C bond formation is larger than that of the cross-disproportionation. Thus we can control the reaction path weight between the cross-disproportionation and the C−C bond formation by modifying HCl concentration.

The disproportionations and the C−C bond formation are found to be the main reactions in acidic hydrothermal conditions. Therefore, the shifts of [HCHO] and [HCOOH] are expressed as follows:

$$\frac{d[HCHO]}{dt} = -2k_{SD}[HCHO]^2 - k_{CD}[HCHO][HCOOH] - k_G[HCHO][HCOOH][HCl] \quad (B3)$$

$$\frac{d[HCOOH]}{dt} = k_{SD}[HCHO]^2 - k_{CD}[HCHO][HCOOH] - k_G[HCHO][HCOOH][HCl] \quad (B4)$$

By subtracting eq B3 from eq B4, we obtain

$$\frac{d([HCOOH] - [HCHO])}{dt} = 3k_{SD}[HCHO]^2 \quad (B5)$$

This equation indicates that the decreasing rate of formaldehyde is always greater than that of formic acid.[38] Also, $[HCHO]_0$ and $[HCOOH]_0$ in eq 9 satisfy

$$[HCOOH]_0 > [HCHO]_0 \quad (B6)$$

Since eq B3 is negative, we obtain

$$\frac{d[HCOOH]}{dt} > \frac{[HCOOH]}{[HCHO]}\frac{d[HCHO]}{dt} \quad (B7)$$

from eqs B5 and B6. Therefore,

$$\frac{d}{dt}\frac{[HCOOH]}{[HCHO]} > 0 \quad (B8)$$

Equation B8 indicates that the ratio of [HCOOH] against [HCHO] increases during the reaction. From eqs 8, 9, and B8,

$$\frac{[HCOOH]}{[HCHO]} > \frac{[HCOOH]_0}{[HCHO]_0} > \frac{k_{SDwater} + k_{SDacid}[HCl]}{k_G[HCl] - k_{CD}} \quad (B9)$$

is obtained. Equations B8 and B9 show not only that eq 7 holds throughout the reaction but also that the relative rate of the C−C bond formation against the methanol formation is enhanced as the reaction proceeds.

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(16) Before studying the hydrothermal reaction, *s*-trioxane is treated in the absence of solvent for 6 h at 225 °C. The mass balance is maintained before and after the reaction, and only formaldehyde is produced from *s*-trioxane by decomposition. In the absence of water, thus, the decarbonylation of formaldehyde expressed as $HCHO \rightarrow CO + H_2$ does not take place; no reaction other than the conversion of *s*-trioxane to the monomeric formaldehyde proceeds at 225 °C in our time scale. This result agrees with a recent ab initio calculation of the formaldehyde reaction; the activation energy of the decarbonylation of formaldehyde is shown to be too high to overcome at 225 °C.[17] According to these results and the retarding effect of hot water on the decarbonylation,[15,18] we can consider that formaldehyde does not decarbonylate in hydrothermal conditions.

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(19) The possible causes of formic acid formation are the disproportionation of formaldehyde (eq 1) and the decarbonylation of formaldehyde (the reaction scheme of ref 16) followed by the formation of formic acid from CO and H_2O.[5,6] H_2 is not observed in both neutral and acidic hydrothermal conditions as seen in Table 3, which means that formaldehyde does not decarbonylate. Formic acid is thus supplied only by the self-disproportionation. The reactivity of neat formaldehyde at 225 °C also indicates that the decarbonylation of formaldehyde does not proceed in hydrothermal conditions.[14−18]

(20) We have achieved the ethanol yield of ∼70% by treating 0.3 M acetaldehyde with 2.0 M HCOOH for 2 h at 250 °C.

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(23) When *s*-trioxane and formic acid are treated in the absence of HCl, the glycolic acid formation is also expected besides the cross-disproportionation because formic acid may work as an acid catalyst. As seen in Figure 1d, however, no glycolic acid is detected in this experiment, which suggests that the acidity (proton productivity) of formic acid is insufficient for the C−C bond formation.

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(27) The decarbonylation of formic acid ($HCOOH \rightarrow CO + H_2O$) is the source of CO,[5,6] and CO_2 is the product of the cross-disproportionation. Though the decarbonylation of formaldehyde to CO and H_2 can produce CO, this reaction does not occur in the present reaction condition; H_2 is not detected as seen in Table 3. Besides the cross-disproportionation, the decarboxylation of formic acid can generate CO_2 and H_2.[5,6] The absence of H_2, however, indicates that this reaction does not proceed in the present reaction condition. Indeed, the decarboxylation of formic acid near our reaction temperature is in the order of 50 h and is too slow for our time scale of a few hours.[5] When the reaction temperature becomes higher and/or the water density is lowered, decarboxylation of formic acid must be taken into account as a competing reaction.[5,6,18,28]

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(29) As seen in Table 4, k_{CD} values in acidic condition are only roughly determined due to a large error. Even taking the errors into account, however, the table no doubt indicates that acid retards the cross-disproportionation.

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(31) Though $[H^+]$ is not zero even in the neutral condition due to the autoprotolysis of water, k_{SDN} is found to be almost equal to $k_{SDwater}$. Here we estimate the reaction order of $[H^+]$ by treating k_{SDN} as the rate constant in the absence of H^+ ($k_{SDwaeter}$). Since $k_{SDwater}$ is substituted with k_{SDN} in eq 6, the equation is transformed as $\ln(k_{SDobs} - k_{SDN}) = \ln k_{SDacid} + n \ln[H^+]$, where n represents the order of $[H^+]$. The value of n is obtained as 0.8 by the plot of $\ln(k_{SDobs} - k_{SDN})$ against $\ln[H^+]$. Considering the experimental error, this result is close enough to unity.

(32) In the case of the cross-disproportionation, the observed rate constant is reduced by the addition of acid, as shown in Table 4. This is an indication of the presence of the base-catalyzed path. When the observed rate constant k_{CDobs} for the cross-disproportionation is written as $k_{CDobs} = k_{CDwater} + k_{CDbase}[OH^-]$; $k_{CDwater}$ and k_{CDbase} are obtained as 10^{-4} M^{-1} s^{-1} and 10^3 M^{-2} s^{-1}, respectively. With the addition of 0.15−0.3 M HCl, $[OH^-]$ is on the order of 10^{-11} M, where it is around 10^{-6} M in neutral condition. The base-catalyzed path is thus dominant in neutral condition. Still, a nonzero $k_{CDwater}$ indicates that the cross-disproportionation proceeds in the absence of OH^-.

(33) Pitzer, K. S. *Activity Coefficients in Electrolyte Solutions*, 2nd ed.; CRC Press: Boca Raton, FL, 1991.

(34) In the present experiment, the concentration of added HCl is higher than that in the study of 1,4-butanediol by 3−4 orders.[30] This is because k_{SDacid} in the present study is smaller than k_{acid} in the reaction of 1,4-butanediol by ∼4 orders. On the other hand, $k_{SDwater}$ of the present study is on the order of k_{water} in the reaction of 1,4-butanediol. Though the reaction temperature is different between the two studies, the difference in the order of k_{SDacid} and k_{acid} indicates that the hydrothermal self-disproportionation is less sensitive to acid than the dehydration of 1,4-butanediol.

(35) In our experimental condition, HCl completely dissociates in water and the concentration of [HCl] we treat is so large that it can be treated as the proton concentration $[H^+]$.

(36) We adopted 3×10^{-4} M^{-1} s^{-1} for the value of k_{CD}. This is because it is the largest value of k_{CD} in acidic condition. By doing so, we avoid underestimation of the required HCl concentration in eq 8.

(37) The low proton-based mass balance in Figure 7 is due to the decarbonylation of formic acid to CO and H_2O. Since H_2O is abundant from the outset as a solvent, newly generated H_2O cannot be detected and this leads to the loss of mass balance. The decarbonylation of formic acid is no longer a minor path in the presence of the excess amount of formic acid.[5,6] Though both C−C bond formation and decarbonylation of formic acid are catalyzed by acid, the former is found to be more sensitive to HCl concentration than the latter. Figure 7 shows that in the presence of 0.1 M *s*-trioxane and 5.0 M HCOOH, the yield of glycolic acid decreases by 12% upon the change of HCl concentration from 0.6 to 0.3 M whereas the mass balance does not vary significantly. On the selective formation of glycolic acid, therefore, the decarbonylation of formic acid is inevitable. This reaction produces CO and increases the pressure in the reactor. We thus need to balance the final yield of the glycolic acid and the pressure capacity of the reactor in the application.

(38) When formic acid is added in excess to the reaction mixture, eq B5 should be modified as $d([HCOOH] - [HCHO])/dt = 3k_{SD}[HCHO] - \Delta$, due to the decomposition of formic acid (Δ). Even in this case, glycolic acid can be selectively produced as long as $[HCOOH]_0$ is large enough to hold $[HCOOH]/[HCHO] > (k_{SDwater} + k_{SDacid}[HCl])/(k_G[HCl] - k_{CD})$ until formaldehyde runs out. In fact, we successfully obtain glycolic acid as the main product when added formic acid partly decomposes as seen in Figure 7.