



# MONASH University

## ***Single-step Conversion of Synthesis Gas into Formaldehyde***

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A thesis submitted for the degree of *Doctor of Philosophy* at  
Monash University in 2016  
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## Abstract

Formaldehyde is known as the building block in many industries including resins, polymers, paints and adhesives. It is widely used in furniture and wood processing. The annual production rate of formaldehyde is in the range of 30 million tons globally and the demand of formaldehyde has grown by 2-3 % per year over the past two decades. Industrially, formaldehyde is produced via methanol partial oxidation. Methanol in turn is produced from synthesis gas which is produced via steam reforming of natural gas. Both methanol synthesis and steam reforming process suffer from high exergy loss due to high temperature processes and large purification units. Considering the large quantity of formaldehyde produced in the world, when combined with the high exergy losses, leads to high energy losses and CO<sub>2</sub> emissions globally. Therefore, alternative method for the production process of formaldehyde is needed. This project develops a novel method of formaldehyde production by using equimolar quantities of carbon monoxide and hydrogen in a catalytic slurry phase reaction. Promoted nickel-based catalysts – Ru-Ni/Al<sub>2</sub>O<sub>3</sub> and Pd-Ni/Al<sub>2</sub>O<sub>3</sub> were used in this project due to their high activity in hydrogenation reactions.

It was observed in this project that formaldehyde yield was significantly higher in the slurry phase (4.6 mmol.L<sup>-1</sup>.g<sub>cat</sub><sup>-1</sup>) compared to the gas phase (8.2×10<sup>-3</sup> mmol.L<sup>-1</sup>.g<sub>cat</sub><sup>-1</sup>). Thermodynamics analysis showed that the reaction is equilibrium limited in the gas phase whereas it is kinetically limited in the liquid phase. It was found that in the slurry reactor increasing the temperature increases the reaction rate but the formaldehyde yield peaks at 353 K, above which the yield reduces. This is in agreement with the fact that the equilibrium constant of the reaction decreases as the temperature increases. Increasing pressure and stirring speed increased the formaldehyde yield.

The reaction mechanism was investigated based on deuterium labelling technique. It was shown that the reactant gases dissolve in the solvent and are adsorbed on the catalyst surface. The adsorbed gases react on the catalyst surface to form formaldehyde which desorbs and immediately hydrates in aqueous conditions to form methylene glycol.

The effect of solvents with high CO and H<sub>2</sub> solubility was investigated in this study since it was found that increasing the solubility of the gases was important to achieve higher production. However, it was observed that in addition to the gases solubility, formaldehyde yield also depends on the

reactivity of the solvents with desorbed formaldehyde. The last step is important to shift the equilibrium of the CO hydrogenation reaction.

In a nutshell, higher pressures were in favour of the reaction but the operating pressure in this study was limited to 100 bar by the maximum available gas cylinder pressure. The most suitable solvent was found to be methanol since it has higher solubility of CO and H<sub>2</sub> compared to water, and similar to water, it also reacts with formaldehyde to shift the equilibrium. Highest yield was achieved by using pure methanol as a solvent at 363 K and 100 bar, which resulted in formaldehyde yield of 15.58 mmol.L<sup>-1</sup>.g<sub>cat</sub><sup>-1</sup>.

## **Declaration**

This thesis contains no material which has been accepted for the award of any other degree or diploma at any university or equivalent institution and that, to the best of my knowledge and belief, this thesis contains no material previously published or written by another person, except where due reference is made in the text of the thesis.

## Monash University

### Declaration for thesis based or partially based on conjointly published or unpublished work

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This thesis includes two original papers published in peer reviewed journals and one unpublished publications. The core theme of the thesis is single-step conversion of synthesis gas into formaldehyde. The ideas, development and writing up of all the papers in the thesis were the principal responsibility of myself, the candidate, working within the Chemical Engineering Department under the supervision of Dr Akshat Tanksale and A/Prof Andrew Hoadley.

The inclusion of co-authors reflects the fact that the work came from active collaboration between researchers and acknowledges input into team-based research.

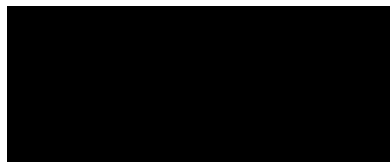
In the case of Chapter 2 and Chapter 3, my contribution to the work involved the following:

Thesis chapter	Publication title	Publication status*	Nature and extent of candidate's contribution
2	Critical review and exergy analysis of formaldehyde production processes	Published	Initiation 50% Modelling 100% Writing 70%
3	Formaldehyde production via hydrogenation of carbon monoxide in the aqueous phase	Publihsed	Initiation 50% Experimental works 100% Writing 70%

I have renumbered sections of submitted or published papers in order to generate a consistent presentation within the thesis.

Signed:

Date:



## Acknowledgement

I express my sincere appreciation to everyone who helped me to complete my thesis. I sincerely thank my supervisors Dr. Akshat Tanksale and A/Prof Andrew Hoadley for their incredible help and support and their great contribution to this work. It was impossible for me to conduct this research without their supervision and their valuable feedback. It was a great pleasure for me to work with them and learn from them during my candidature. I also would like to thank Monash Institute of Graduate Research (MIGR) for the financial support and giving me this opportunity to conduct my research in one of the best universities in Australia.

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## List of Publications

### *Journal Publications:*

**A.M. Bahmanpour**, A. Hoadley, and A. Tanksale, *Critical review and exergy analysis of formaldehyde production processes*. **Reviews in Chemical Engineering**, 2014. **30**(6): p. 583-604.

**A.M. Bahmanpour**, A. Hoadley, and A. Tanksale, *Formaldehyde production via hydrogenation of carbon monoxide in the aqueous phase*. **Green Chemistry**, 2015. **17**(6): p. 3500-3507.

### *Conference Publications:*

**A.M. Bahmanpour**, A. Hoadley, and A. Tanksale, *Direct formaldehyde production from Synthesis gas*, **AIChE annual meeting**, Atlanta, GA, USA, 2014.

## Abbreviations

AW	Atomic Weight
BASF	Badische Anilin- und Soda-Fabrik
BET	Brunauer- Emmett- Teller
BFM	Bubble Flow Meter
BPR	Back Pressure Regulator
DFT	Density Functional Theory
FORMOX	Formaldehyde by oxidation
GC	Gas Chromatography
HPLC	High Performance Liquid Chromatography
ICI	Imperial Chemical Industries
MFC	Mass Flow Controller
MS	Mass Spectrometer
NIH	National Institutes of Health
NMR	Nuclear Magnetic Resonance
PF	Particle Filter
PG	Pressure Gauge
TEM	Transmission Electron Microscopy
TOF	Turnover Frequency
TPRS	Temperature Programmed Reaction Spectroscopy
WF	Weight Fraction
XRD	X-ray Diffraction
XRF	X-ray Fluorescence



## Symbols

%vol	-	Volume Percent
%wt	-	Weight Percent
$g_{cat}$	g	Weight of Catalyst in Grams
$k_i$	$s^{-1}$	Rate Constant
K	$kPa^{-1}$	Equilibrium Constant
$M_i$	$\mu mol.g^{-1}$	Amount of Adsorbed Gas
$N_s$	$\mu mol.m^{-2}$	Surface Density of Redox Site
P	bar	Pressure
$p^*$	kPa	Vapor Pressure
T	K	Temperature
X <sub>e</sub>	%	Equilibrium Conversion
X	%	Conversion
$\Delta G$	$kJ.mol^{-1}$	Gibbs Free Energy
$\Delta H$	$kJ.mol^{-1}$	Enthalpy Change
$\Delta S$	$J.mol^{-1}.K^{-1}$	Entropy Change

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# **Chapter 1:**

## **Introduction**

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## 1.1 Background

Due to the increasing global energy demand, many researchers have focused on reduction of energy consumption in chemical industry. It is always essential to search for more efficient, energy saving processes for the production of valuable chemicals. In this way, various production methods have been optimized or changed. Many aspects can be considered and evaluated during the production process. Although the yield of the product and the conversion of the reactants are two of the most important factors to consider in a chemical process, other factors such as optimizing the energy consumption rate, lowering the production capital costs and simplifying the downstream process are yet to be considered.

Formaldehyde is considered as one of the main building blocks in petrochemical industry. As one of the main intermediate products, formaldehyde is used to produce phenol-formaldehyde, urea-formaldehyde, and melamine-formaldehyde. The annual production rate of formaldehyde is in the range of 30 million tons globally and the demand of formaldehyde has grown by 2-3 % per year over the past two decades (Ulukardesler et al., 2010). Over 80% of the produced formaldehyde is used as an intermediate product for resin production (National Toxicology, 2010). Recently, it is shown that the industrial process of formaldehyde production suffers from ~57% losses of exergy (i.e. quality energy) (Bahmanpour et al., 2014). Considering the large amount of formaldehyde annually produced, the total energy loss of the formaldehyde production process is considerably high.

It is known that formaldehyde is produced through a relatively long sequence of processes from natural gas. Formaldehyde is produced in three stages: (a) steam reforming of natural gas to produce synthesis gas (b) methanol synthesis and (c) partial oxidation and/or dehydrogenation of methanol to produce formaldehyde. Although many have studied the possibility of production of formaldehyde through a single step process from natural gas to reduce the capital costs and energy losses, relying on natural gas as a depleting fossil fuel is still considered as a disadvantage of this method. Moreover, the low methane conversion and/or formaldehyde selectivity is still considered as the main issue in this method.

Very limited studies on the production of formaldehyde directly from synthesis gas are available (Newton & Dodge, 1933). However, this process was reported not to be successful since only trace

## Chapter 1

amount of formaldehyde was reported to be produced. The highest carbon monoxide conversion reported was quite insignificant (0.2%) (Morgan et al., 1932). Since direct formaldehyde production from synthesis gas reduces the number of stages for formaldehyde production (by bypassing methanol production step), successful formaldehyde production in this method can reduce the capital costs of the process. In addition, synthesis gas can be produced through biomass gasification process. Therefore, formaldehyde production through biomass gasification followed by direct synthesis from syngas can lead to a sustainable method of formaldehyde production. The presented thesis is focused on the single step production of formaldehyde from synthesis gas.

### 1.2 Research Aims

The general aim of this research is to develop and establish a new method of formaldehyde production based on direct synthesis from syngas. The first step is to investigate the possibility of the production of formaldehyde in this way theoretically and experimentally. It is critical to understand the conditions of the process in order to improve the yield. The previous attempts to convert synthesis gas directly into formaldehyde were done by using copper-zinc catalyst. It is very important to find a suitable catalyst for the process based on the conditions at which the experiment is conducted. Optimization of the operating conditions of the process to achieve higher yield of formaldehyde is the next step. A critical step is to investigate the reaction pathway since the bottleneck of the process may be identified through the reaction mechanism. The main aims of this study are summarised here:

- ❖ Feasibility study of the direct conversion of synthesis gas to formaldehyde
- ❖ Direct conversion of synthesis gas to formaldehyde in the gas phase and the slurry phase
- ❖ Investigation of the synthesis gas to formaldehyde reaction pathway
- ❖ The effect of solvents on the yield of formaldehyde in the slurry phase reaction of synthesis gas to formaldehyde

The results then can be used for further improvement of this process which may potentially lead to development of a novel renewable method of formaldehyde production. In the bigger picture, the results of this research can be the elementary steps towards the production of green formaldehyde.

## **1.3 Thesis Structure**

The outline of this thesis is presented here:

### **Chapter 1: Introduction**

This chapter includes the research background and aims as well as the thesis structure as presented here.

### **Chapter 2: Literature Review**

This chapter contains a thorough literature review and discussion about different methods of formaldehyde production and it extensively explains the kinetics of the methanol partial oxidation reaction to formaldehyde. Some of the recent advancements in formaldehyde production are presented and other methods of formaldehyde synthesis are looked into. The contents of this chapter along with the simulation and exergy analysis study done on the methanol ballast process have been published as a journal paper in *Reviews in Chemical Engineering*.

### **Chapter 3: Formaldehyde Production via CO Hydrogenation in the Aqueous Phase**

In chapter 3, the feasibility of direct formaldehyde production from synthesis gas is investigated theoretically using HSC Chemistry software <sup>TM</sup>. The experimental results of the formaldehyde production processes in the gas phase and in the slurry phase are presented and discussed. The catalyst characterization results are presented and the effects of temperature, pressure, Stirring speed, and the mass of the catalyst are studied. Some of the contents of this chapter are published as a journal paper in *Green Chemistry*.

### **Chapter 4: Reaction Mechanism of CO Hydrogenation to Formaldehyde in the Aqueous Phase**

In this chapter, the reaction mechanism of the carbon monoxide hydrogenation reaction to formaldehyde in the aqueous phase is investigated. The reaction mechanism is studied based on the potential intermediate products in various samples collected from the reaction system during the process as well as determination of the source of atoms participating in the reaction. Based on this study, a paper is currently under preparation for publication.

## Chapter 1

### Chapter 5: Effect of Solvents on the Formaldehyde Production in the Slurry Reactor

The main aim of chapter 5 is to study the effect of solvents on the formaldehyde yield in the slurry reactor. The idea of catalytic carbon dioxide hydrogenation to formaldehyde is also introduced in this chapter for further studies. Based on this study, a paper is currently under preparation for publication.

### Chapter 6: Conclusions

This chapter includes the outcomes of this research. It can be considered as the fulfilment of the aims of this project.

### Chapter 7: Recommendations and Future Studies

In this chapter, the main recommendations for the future studies are presented.

### Appendix: Published Journal Article Front Pages

## 1.4 References

- Bahmanpour, A. M., Hoadley, A., & Tanksale, A. (2014). Critical review and exergy analysis of formaldehyde production processes. *Reviews in Chemical Engineering*, 30(6), 583-604.
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## Declaration for Thesis Chapter 2

### Declaration by candidate

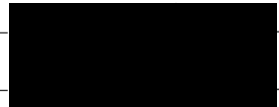
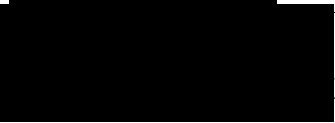
In the case of Chapter 2, the nature and extent of my contribution to the work was the following:

Nature of contribution	Extent of contribution (%)
Initiation	50
Modelling	100
Writing	70

The following co-authors contributed to the work. If co-authors are students at Monash University, the extent of their contribution in percentage terms must be stated:

Name	Nature of contribution	Extent of contribution (%) for student co-authors only
Akshat	Initiation	
Tanksale	Reviewing and Editing the Paper	
Andrew	Initiation	
Hoadley	Reviewing and Editing the Paper	

The undersigned hereby certify that the above declaration correctly reflects the nature and extent of the candidate's and co-authors' contributions to this work.

Candidate's Signature		Date 24/09/18
Main Supervisor's Signature		Date 24/09/15

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# Chapter 2:

## Literature Review

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### Abstract

*Formaldehyde is one of the most important intermediate chemicals and has been produced industrially since 1889. Formaldehyde is a key feedstock in several industries like resins, polymers, adhesives and paints, making it one of the most valuable chemicals in the world. In this chapter, we study the leading commercial processes for formaldehyde production and compare them with recent advancements in catalysis and novel processes. This chapter compares, in extensive detail, the reaction mechanisms and kinetics of water ballast process (or BASF process), methanol ballast process, and Formox process. Thermodynamics of the reactions involved in the formaldehyde production process was investigated using HSC Chemistry™ software.*

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## 2.1 Introduction

Formaldehyde (HCHO) is one of the simplest molecules in organic chemistry which has a wide range of applications in chemical processes. It is produced in nature by photochemical processes in atmosphere or by incomplete combustion of organic materials (Reuss et al., 2003). HCHO gas is colorless at ambient temperature with an irritating odour which unfavorably affects eyes and skin. The effects of HCHO on human health have been studied. HCHO is toxic and probably a carcinogen since there is enough evidence that it can cause nasopharyngeal cancer (Cogliano et al., 2005). Some physical properties of monomeric HCHO can be found in Table 2.1 (Reuss et al., 2003):

Table 2.1: Physical properties of monomeric formaldehyde (Reuss et al., 2003)

Properties	Values
Heat of formation at 298.15 K	$-115.9 \pm 6.3 \text{ kJ.mol}^{-1}$
Gibbs energy at 298.15 K	$-109.9 \text{ kJ.mol}^{-1}$
Entropy at 298.15 K	$218.8 \pm 0.4 \text{ kJ.mol}^{-1}.\text{K}^{-1}$
Heat of combustion at 298.15 K	$-561.5 \text{ kJ.mol}^{-1}$
Heat of vaporization at 253.95 K	$23.32 \text{ kJ.mol}^{-1}$
Specific heat capacity at 298.15 K, $C_p$	$35.425 \text{ J.mol}^{-1}.\text{K}^{-1}$
Heat of solution at 296.15 K	
in water	$-62 \text{ kJ.mol}^{-1}$
in methanol	$-62.8 \text{ kJ.mol}^{-1}$
in 1-propanol	$-59.5 \text{ kJ.mol}^{-1}$
in 1-butanol	$-62.4 \text{ kJ.mol}^{-1}$
Cubic expansion coefficient	$2.83 \times 10^{-3} \text{ K}^{-1}$
Specific magnetic susceptibility	$-0.62 \times 10^6$
Vapour density relative to air	1.04

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HCHO heat of formation, heat capacity, and entropy are given in Table 2.2 for a wide range of temperatures and atmospheric pressure (J. F. Walker, 1967):

Table 2.2: Heat of formation, heat capacity, and entropy of formaldehyde (J. F. Walker, 1967)

Temperature (K)	Heat of formation (kJ.mol <sup>-1</sup> )	Heat capacity (J.mol <sup>-1</sup> .K <sup>-1</sup> )	Entropy (J.mol <sup>-1</sup> .K <sup>-1</sup> )
100	113.5	33.3	182.0
200	114.5	33.5	205.1
298.15	116.0	35.4	218.8
300	116.0	35.5	219.0
400	117.7	39.3	229.7
500	119.3	43.8	239.0
600	120.8	48.2	247.4
800	123.3	56.0	262.3
1000	125.0	62.0	275.5
2000	128.6	754.0	323.8
3000	130.9	79.6	355.4
4000	134.6	81.1	378.5
5000	140.5	81.9	396.7
6000	148.6	82.3	411.7

The gas mixture with air is flammable at 293.15 K or higher. HCHO is considered as a reactive compound. At high temperatures (>423.15 K) it tends to undergo heterogeneous decomposition to produce methanol (CH<sub>3</sub>OH) and carbon dioxide (CO<sub>2</sub>). But if the temperature goes higher (>623.15 K), it may tend to form synthesis gas (hydrogen (H<sub>2</sub>) and carbon monoxide (CO) mixture) (Bone & Smith, 1905; Calvert & Steacie, 1951). It can be easily involved in both oxidation and reduction reactions. For example, in the presence of H<sub>2</sub> it can be reduced to form CH<sub>3</sub>OH over a Ni catalyst. Nitric acid (HNO<sub>3</sub>), potassium permanganate (KMnO<sub>4</sub>), potassium dichromate (K<sub>2</sub>Cr<sub>2</sub>O<sub>7</sub>), or oxygen (O<sub>2</sub>) can readily oxidize HCHO to form formic acid (HCOOH), CO<sub>2</sub> and water (J. F. Walker, 1967). At atmospheric pressure, HCHO condenses at 253.95 K and freezes at 155.15 K as a white paste. In the liquid phase, its density is 0.8153 g/cm<sup>3</sup> at 253.15K and 0.9172 g/cm<sup>3</sup> at 193.15 K. The vapour pressure of the liquefied HCHO has been measured for a wide range of temperatures by Spence and Wild (1935). Measured partial pressures for the temperature range of 163.75 K to 250.85 K are tabulated in Table 2.3:

Table 2.3: Partial pressure of formaldehyde for the temperature range of 163.75 K to 250.85 K (Spence &amp; Wild, 1935)

Temperature (K)	Pressure (kPa)
163.75	0.13
168.75	0.25
174.85	0.48
177.95	0.65
184.05	1.16
187.55	1.63
194.25	2.80
194.85	2.95
201.25	4.72
204.65	6.19
207.85	7.86
208.55	8.22
209.45	8.69
217.35	14.80
219.15	16.62
223.85	21.74
232.55	35.54
234.05	38.74
238.85	49.17
250.85	88.55

According to the measured values, the vapour pressure of liquid HCHO can be determined using the following equation (Spence & Wild, 1935):

$$\log_{10} p^* = -1429/T + 1.75 \log T - 0.0063T + 5.023 \quad (2.1)$$

where  $p^*$  is the HCHO vapour pressure in kPa and  $T$  is in K.

Many different common uses and industrial applications can be named for HCHO as it is one of the most important compounds in chemical manufacturing industry. HCHO is a building block in the synthesis of industrial resins such as melamine-formaldehyde, urea-formaldehyde and phenol-formaldehyde. Due to the applications of HCHO derivatives in wood processing and furniture, carpeting and textiles, and construction, this chemical is considered as an important material in the global economy (Tang et al., 2009). HCHO can also be used directly to produce other chemicals; glycolonitrile ( $C_2H_3NO$ ) can be produced by HCHO reaction with hydrocyanic acid (HCN) (Kirk & Othmer, 1991). Acid catalysed reaction between HCHO and olefins may lead to the production of  $\alpha$ -

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hydroxymethylated adducts (J. F. Walker, 1967). Methylamine can also be produced by the reaction of HCHO with amines. Due to its high ability to polymerize during long term storage, HCHO is commercially available as an aqueous solution of 37 wt% to 50 wt% concentration known as formalin. CH<sub>3</sub>OH is usually added (10-12 wt%) in the solution as the polymerization inhibitor.

The HCHO production and consumption rate have significantly increased during recent years. From 2003 to 2006, the HCHO consumption rate grew at an average of 5.4% worldwide (Bizzari, 2007). In 2007, China as the largest HCHO producer and consumer produced 12 million metric tonnes of HCHO (Tang et al., 2009). Total global HCHO production is approximately 25 to 27 million metric tonnes per year and the growth of the consumption rate of this chemical has been fairly steady over the past twenty years (Uluokardesler et al., 2010). Considering its steady consumption rate and its application as a base chemical in more than 50 industries (Reuss et al., 2003), it is important to review the current production processes and identify novel technologies which may lead to more efficient and sustainable production in the future.

## 2.2 Formaldehyde Production

### 2.2.1 History

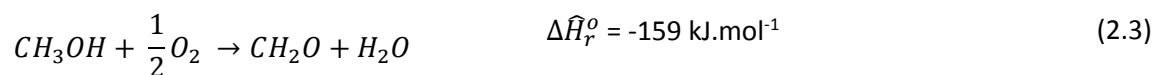
HCHO was first accidentally produced by Alexander Mikhailovich Butlerov in 1859 while he was attempting to produce methylene glycol (CH<sub>2</sub>(OH)<sub>2</sub>) by the hydrolysis of methylene acetate (Butlerov, 1859). In aqueous solutions, CH<sub>2</sub>(OH)<sub>2</sub> is stable while isolation of this chemical leads to the decomposition of the unstable glycol to HCHO and water (F. Walker, 1933). Butlerov also produced solid HCHO polymers through the reaction between methylene iodide (CH<sub>2</sub>I<sub>2</sub>) and silver oxalate (Ag<sub>2</sub>C<sub>2</sub>O<sub>4</sub>), but he did not realize that the monomer can be achieved by vaporization. He also prepared a new compound or polymer through the reaction between CH<sub>2</sub>I<sub>2</sub> and silver oxide (Ag<sub>2</sub>O). The new compound was reacted with ammonia (NH<sub>3</sub>) to form a crystalline compound. Although he did not realise that HCHO was produced through his experiments, he stated that these reactions are expected to produce the unknown “formyl-aldehyde”. In 1868, August Wilhelm von Hofmann produced and identified HCHO by passing CH<sub>3</sub>OH vapour and air over a heated Pt spiral (Hofmann, 1868). Therefore he was the first person to produce HCHO by CH<sub>3</sub>OH oxidation and identify the product. This process is still mainly used for HCHO production, with the changes in catalysts. By reacting hydrogen sulphide



(H<sub>2</sub>S) and its acid solution, Hofmann was able to produce trithioformaldehyde ((CH<sub>2</sub>S)<sub>3</sub>) and later he was able to form HCHO polymer by vaporization of this chemical (Hofmann, 1869). According to Tollens proposal for regulation of the CH<sub>3</sub>OH to air ratio and increasing the yield of the process, industrial production of HCHO was achieved in 1882 (Reuss et al., 2003). HCHO was first produced commercially by catalytic partial oxidation of CH<sub>3</sub>OH with air over an unsupported Cu catalyst at atmospheric pressure. Later Cu was replaced by Ag to increase the yield from 43 - 47% yield using Cu to 67% using Ag) (Roeda, 2003; Trillat, 1903). There was an attempt to produce HCHO by non-catalytic oxidation of propane (C<sub>3</sub>H<sub>8</sub>) and butane (C<sub>4</sub>H<sub>10</sub>) in the United States. However, a wide range of co-products were produced and it required a complicated and costly separation system. Therefore, CH<sub>3</sub>OH partial oxidation was the preferred choice of process (J. F. Walker, 1967) which is discussed further.

### 2.2.2 Industrial Methods

The sequence of HCHO production is made of three main stages- steam reforming of natural gas leads to the production of synthesis gas which is converted to CH<sub>3</sub>OH via CH<sub>3</sub>OH synthesis reaction (hydrogenation of CO) and eventually partial oxidation of CH<sub>3</sub>OH leads to HCHO production. HCHO is industrially produced mainly through two major reactions; one is CH<sub>3</sub>OH dehydrogenation (equation 2.2) and the other one is partial oxidation of CH<sub>3</sub>OH (equation 2.3). These reactions are as follows:



Currently HCHO is industrially produced from CH<sub>3</sub>OH using two major processes; the air deficient process or silver contact process through which CH<sub>3</sub>OH goes through both dehydrogenation and partial oxidation reaction to produce HCHO (used by BASF, Borden, Bayer, Degussa, ICI, Celanese, DuPont, Mitsubishi and Mitsui) (Brenk, 2010; Qian et al., 2003). The other process involves an excess of air, known as the FORMOX process, in which only partial oxidation of CH<sub>3</sub>OH takes place (used by Lummus, Montecatini, Hiag/Lurgi, and Perstorp/Reichold) (Brenk, 2010; Qian et al., 2003; Reuss et al., 2003). These two process routes are shown in a tree diagram in Figure 2.1. The air deficient process

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can be further classified into two types – the methanol ballast process and the water ballast process (or BASF process). In the methanol ballast process only  $\text{CH}_3\text{OH}$  and air are used in the reactant feed stream and low single pass  $\text{CH}_3\text{OH}$  conversion is achieved (used by ICI and Degussa). In the water ballast process, steam is also added to the feed stream and comparatively higher single pass  $\text{CH}_3\text{OH}$  conversion is achieved.

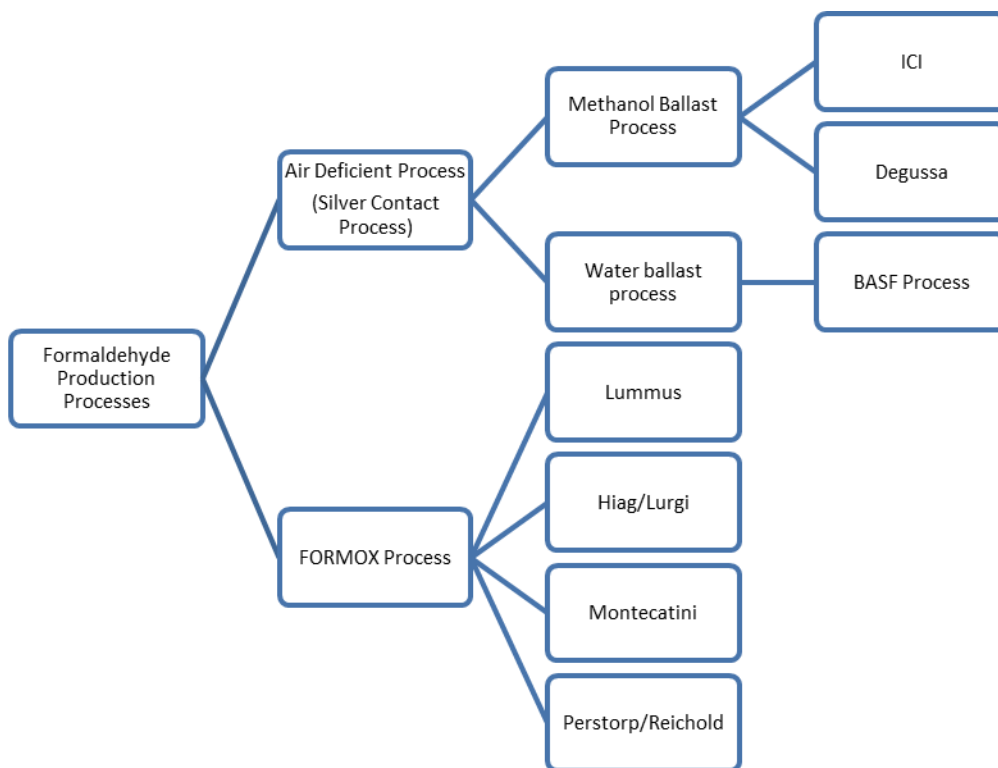


Figure 2.1: Tree diagram of industrial processes for formaldehyde production

### 2.2.2.1 Methanol Ballast Process

Methanol ballast process has been mentioned as one of the common methods of  $\text{HCHO}$  production in industry, used by some of the known producers of  $\text{HCHO}$  (e.g. ICI and Degussa)(Qian et al., 2003). In the ICI process (first patented in 1940 (Arthur & Maurice, 1940)), the silver crystals are formed in a 1 cm thick layer inside the reactor which is directly mounted on the heat recovery system. Silver crystal held by silver metal gauze is used in the Degussa process. The reactor operates at 788.15-823.15 K. In this method, heat is immediately recovered after the product stream leaves the reactor (Chauvel &

Lefebvre, 1989). Single pass conversion is relatively low in methanol ballast process and unconverted  $\text{CH}_3\text{OH}$  is recycled for higher overall conversion achievement.

#### **2.2.2.2 BASF Process**

The BASF process is the most widely applied industrial  $\text{HCHO}$  synthesis process and involves passing  $\text{CH}_3\text{OH}$  over a silver catalyst. A schematic diagram of the BASF process (Fig. 2.2) shows that a fresh mixture of  $\text{CH}_3\text{OH}$  and water enters an evaporator along with process air and is mixed with off-gas recycled from the absorption column. Part of the required heat in the evaporator is supplied by using the first stage of the absorption column. The reactor used in this process contains a thin layer of silver grains placed between silver gauze. A cooler is located immediately after the reactor to prevent further possible reactions.

In order to control the temperature of the reactor, the water content of the fresh feed and  $\text{CH}_3\text{OH}$  to  $\text{O}_2$  ratio can be adjusted. The ratio controls the temperature of the process and determines the dominant reaction (either equation (2.2) or (2.3)) through which  $\text{HCHO}$  is produced. "Scheme A" shown in Fig. 2.2, contains overhead light components of the absorption column and heavier components can be recycled in "Scheme B". The desirable composition of the final formalin solution defines whether or not any of the mentioned streams should be used.

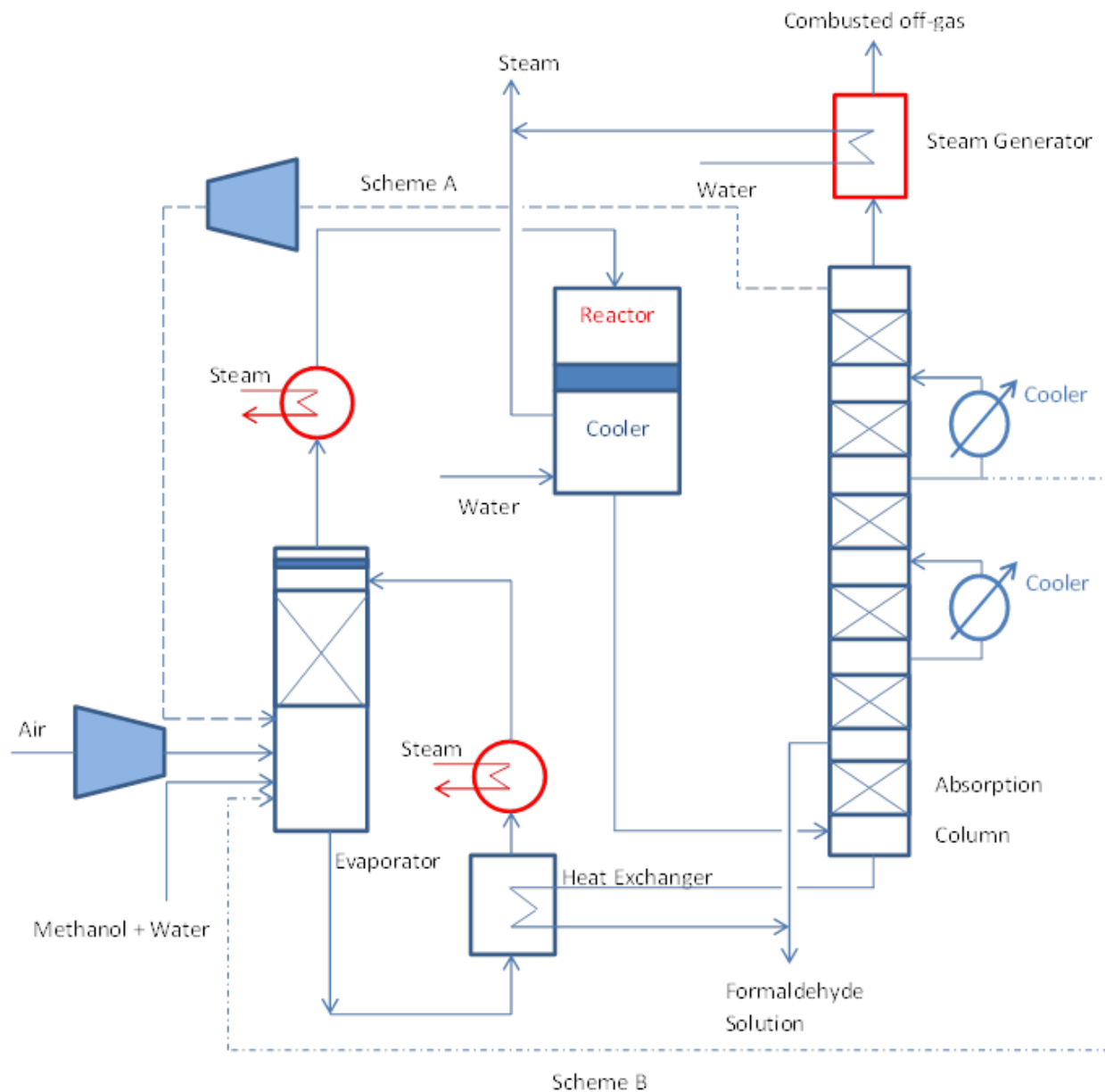


Figure 2.2: BASF process schematic diagram

The role of each reaction in this process can be determined by controlling the temperature of the reactor, amount of water entering with  $\text{CH}_3\text{OH}$  as the fresh feed, and the  $\text{CH}_3\text{OH}$  to  $\text{O}_2$  ratio. Depending on the  $\text{CH}_3\text{OH}$  to  $\text{O}_2$  ratio, at high temperatures  $\text{CH}_3\text{OH}$  may be dehydrogenated directly (equation (2.2)). The effect of water content was studied in a research done by Qian et al. (2003).

It was claimed that  $\text{CH}_3\text{OH}$  conversion and  $\text{HCHO}$  selectivity increase by addition of water compared with the same case with no added water. However, some aspects such as the concentration of the

final product are affected by the water content of the fresh feed. Therefore, an optimum ratio should be used. The water to CH<sub>3</sub>OH ratio of 2:3 or 0.67 has been suggested in literature (Reuss et al., 2003).

As mentioned before, both partial oxidation and direct dehydrogenation of CH<sub>3</sub>OH take place when silver catalysts are used. However, the selectivity of HCHO can be limited considering the following reactions:



Methyl formate (C<sub>2</sub>H<sub>4</sub>O<sub>2</sub>) and HCOOH are the other main by-products of this process and are generally removed as impurities by anion exchange.

### 2.2.2.3 FORMOX Process

Many researchers have discussed the effects of using other catalysts. The most important catalysts which have been discussed are iron/molybdenum or vanadium oxide in the Formox process. The Mo/Fe atomic ratio in this process is usually between 1.5 and 3. Small amounts of other materials such as V<sub>2</sub>O<sub>5</sub>, Cr<sub>2</sub>O<sub>3</sub>, CuO, P<sub>2</sub>O<sub>5</sub> and CoO are also present in the catalyst. At atmospheric pressure, CH<sub>3</sub>OH is almost completely converted in the temperature range of 543.15-673.15 K. However, if the temperature exceeds 743.15 K, oxidation of HCHO to CO becomes significant:



A schematic diagram of the FORMOX process is shown in Fig. 2.3. As it can be seen, the CH<sub>3</sub>OH feed stream passes through a steam evaporator and enters the reactor. Similar to the BASF process, a cooler is located directly after the reactor. Products are cooled down to 383.15 K before entering the absorption column. One of the key controllers of the HCHO concentration in this process is the

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amount of water added at the top of absorption column. The product stream goes through an anion-exchange in order to minimize the formation of  $\text{HCOOH}$ . The final product stream contains up to 55 wt%  $\text{HCHO}$  and 0.5 – 1.5 wt%  $\text{CH}_3\text{OH}$ .  $\text{CH}_3\text{OH}$  conversion is around 98 to 99% in case of high selectivity and activity. The overall plant yield can be between 88 and 91%.

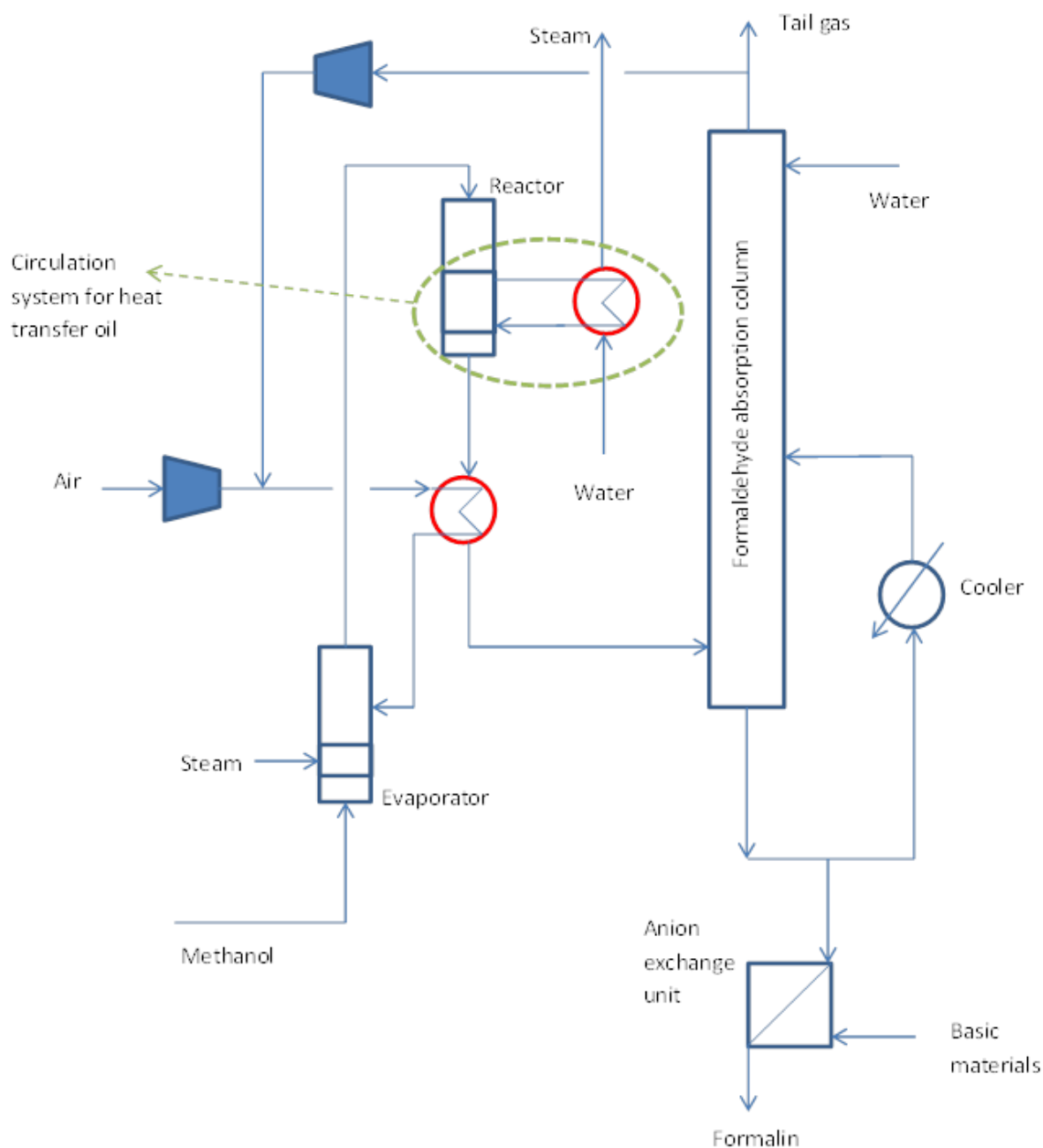


Figure 2.3: FORMOX process schematic diagram

Due to the lower temperature and the higher stability of the catalysts in the direct oxidation of  $\text{CH}_3\text{OH}$  to  $\text{HCHO}$ , this method is attracting more attention compared with the silver-based catalytic process. Adkins and Peterson (1931) reported Iron molybdenum as a suitable catalyst for  $\text{CH}_3\text{OH}$  oxidation to  $\text{HCHO}$ . It was commercially used as one of the main catalysts in the  $\text{HCHO}$  production process in the 1950s. In industrial processes, a combination of  $\text{Fe}_2(\text{MoO}_4)_3$  and  $\text{MoO}_3$  has been used as the catalyst this combination may have a higher life time and higher selectivity compared with pure  $\text{Fe}_2(\text{MoO}_4)_3$  catalyst. The effects of different Mo to Fe atomic ratios in  $\text{Fe}_2(\text{MoO}_4)_3$  catalysts on catalyst activity and selectivity towards  $\text{HCHO}$  have been investigated in literature and will be discussed in kinetics and mechanism section. It is found out that ratios above 1.5:1 provide desirable results (Soares et al., 2001).

### **2.2.3 Other Methods**

There are many other methods through which  $\text{HCHO}$  can be produced. In many of these methods either  $\text{CH}_3\text{OH}$  or hydrocarbon gases are used as the source of  $\text{HCHO}$  production. Therefore, these methods are categorized in two groups as indicated below:

#### **2.2.3.1 Other Formaldehyde Production Methods from Methanol**

Many researchers have investigated other methods of  $\text{HCHO}$  production from  $\text{CH}_3\text{OH}$  using different operating conditions. Brackman (1959) reported that by passing  $\text{O}_2$  through a stirred mixture of liquid  $\text{CH}_3\text{OH}$  and cupric nitrate tri-hydrate complex at 298.15 K, 80% yield of  $\text{HCHO}$  based on the amount of  $\text{O}_2$  consumed can be achieved. Dehydrogenation of  $\text{CH}_3\text{OH}$  is another subject which has been discussed by some researchers (Newton & Dodge, 1933; Sabatier & Mailhe, 1910; B. A. Sexton, 1981; Su et al., 1992; Zaza et al., 1994); Sabatier and Mailhe (1910) studied this process and declared that this reaction is a reversible one. Later, Newton and Dodge (1933) derived the following equilibrium relation for this reversible reaction:

$$\log_{10} K_p = \left( \frac{4600}{T} - 8.48 \right) \quad (2.8)$$

Where  $T$  (K) is the equilibrium temperature and  $K_p$  is the equilibrium constant for this reaction. However, there are three undesirable side reactions which might occur during this process. They can

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be controlled by controlling the operating conditions. These reactions are pyrolytic decomposition of HCHO and oxidation to HCOOH, CO, CO<sub>2</sub> and water;



Aries (1960) described that elevated pressures in CH<sub>3</sub>OH dehydrogenation process prevents the formation of HCOOH in the products stream. Passing CH<sub>3</sub>OH and air over a finely divided silver catalyst at 202.65-1013.25 kPa and 973.15 to 1073.15 K, the product gas mixture contained HCHO, traces of water and unreacted CH<sub>3</sub>OH, but no HCOOH.

Punderson (1960) suggested a combination of Ag, Cu and Si as a new CH<sub>3</sub>OH dehydrogenation catalyst. According to his patent, by combination of 98.8% Ag, 1% Cu, and 0.2% Si, an active catalyst was made through which 70% CH<sub>3</sub>OH conversion and 82% HCHO yield was possible. By using the combination of 97.8% Ag, 2% Cu, and 0.2% Si even better results were achieved, 71.5% CH<sub>3</sub>OH conversion at the end of the seventh cycle in the process was achieved using the latter catalyst while it was exposed to pure CH<sub>3</sub>OH vapour in the temperature range of 873.15-973.15 K.

Considering the difficulty of the separation of HCHO from the water produced by partial oxidation of CH<sub>3</sub>OH to HCHO, CH<sub>3</sub>OH dehydrogenation has been studied as an alternative method using types of reactors such as inorganic membrane reactors (Deng & Wu, 1994) and slurry reactors (Shreiber & Roberts, 2000) using different catalysts.

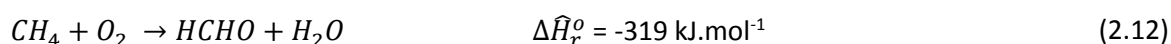
### 2.2.3.2 Formaldehyde Production from Methane

Direct oxidation of methane (CH<sub>4</sub>) to form HCHO is another method which has been extensively discussed (Blair & Wheeler, 1923; Polnišar et al., 2011; H. Zhang et al., 2006; X. Zhang et al., 2003). This process has the obvious advantage of producing HCHO in a single stage, however low selectivity and yield of HCHO are the main issues faced in this process (Bobrova et al., 2007). Many attempts have been made to increase the selectivity of HCHO and in some studies approximately 80% selectivity was achieved (J. Zhang et al., 2011), but the issue of low CH<sub>4</sub> conversion (<2%) remained unresolved. In a recent study (J. Zhang et al., 2012a), NO<sub>2</sub> has been introduced as a promoter and the process took



place in a very short residence time. Highest CH<sub>4</sub> conversion (24.3%) and yield (9%) was achieved in this study for this process so far. CO is the main by-product of this process.

Among the catalysts which have been used for partial oxidation of CH<sub>4</sub> to HCHO, iron oxide, molybdenum oxide (MoO<sub>3</sub>), and vanadium oxide (V<sub>2</sub>O<sub>5</sub>) are the most common catalysts (He et al., 2009; Launay et al., 2007; X. Zhang et al., 2003). It is difficult to activate the relatively inert CH<sub>4</sub> on the surface of a catalyst in order to make HCHO considering the weak chemisorption of CH<sub>4</sub> on the catalyst surface (Michalkiewicz et al., 2008). CH<sub>4</sub> oxidation (eq. (2.12)) and HCHO oxidation (eq. (2.6)) occur in this process dominantly:



As it can be seen, further oxidation of HCHO may lead to CO<sub>2</sub> production for which oxygen should be the limiting reactant which significantly affects CH<sub>4</sub> conversion. Further considerations are that metal-oxygen bonds on the catalyst surface may lead to higher catalyst activity, but lower HCHO selectivity (Fajardo et al., 2008). High temperature increases the CH<sub>4</sub> activation and the kinetics of the process, but also leads to complete oxidation of CH<sub>4</sub> to CO<sub>2</sub>, instead of HCHO synthesis (A. W. Sexton et al., 1998). In order to prevent unfavourable oxidation of HCHO to CO<sub>2</sub>, a short residence time is required. One positive effect is that the high surface area of the catalyst support has been shown to improve CH<sub>4</sub> conversion (Bobrova et al., 2007).

## 2.3 Kinetics and Mechanism

In the following section, the reactions through which HCHO is being produced in industrial processes- namely partial oxidation and dehydrogenation of CH<sub>3</sub>OH- are investigated and the kinetics and mechanism of the mentioned reactions are discussed.

### 2.3.1 Thermodynamic Investigation of the Reactions

The two main reactions which participate in HCHO production in industrial processes are partial oxidation of CH<sub>3</sub>OH and CH<sub>3</sub>OH dehydrogenation. Therefore, these reactions will be investigated thermodynamically in this section. The specifications of each reaction, including enthalpy, entropy,

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Gibbs free energy change, and equilibrium coefficients at different temperatures have been investigated using HSC Chemistry™ software. The results for the range of temperature of 273.15 to 1273.15 K for these reactions are tabulated in Tables 2.4 and 2.5:

Table 2.4: Thermodynamic conditions for CH<sub>3</sub>OH partial oxidation

$CH_3OH(g) + \frac{1}{2} O_2 \leftrightarrow CH_2O(g) + H_2O(g)$				
T	$\Delta H$	$\Delta S$	$\Delta G$	Log(K)
K	kJ.mol <sup>-1</sup>	J.mol <sup>-1</sup> .K <sup>-1</sup>	kJ.mol <sup>-1</sup>	
273.15	-149.694	64.454	-167.299	31.995
323.15	-149.182	66.180	-170.568	27.573
373.15	-148.761	67.392	-173.909	24.346
423.15	-148.424	68.243	-177.301	21.888
473.15	-148.162	68.829	-180.729	19.954
523.15	-147.971	69.214	-184.180	18.391
573.15	-147.845	69.446	-187.647	17.103
623.15	-147.777	69.560	-191.123	16.022
673.15	-147.760	69.586	-194.602	15.102
723.15	-147.788	69.546	-198.080	14.309
773.15	-147.854	69.458	-201.556	13.618
823.15	-147.952	69.336	-205.026	13.011
873.15	-148.075	69.191	-208.489	12.474
923.15	-148.219	69.030	-211.944	11.993
973.15	-148.382	68.858	-215.392	11.562
1023.15	-148.561	68.679	-218.830	11.173
1073.15	-148.754	68.495	-222.260	10.819
1123.15	-148.961	68.307	-225.680	10.497
1173.15	-149.172	68.123	-229.090	10.201
1223.15	-149.388	67.943	-232.492	9.929
1273.15	-149.610	67.765	-235.885	9.679

As it can be seen, the reaction is promising from the thermodynamics point of view. Enthalpy change does not change significantly with temperature. High equilibrium coefficients and highly negative Gibbs free energy change show high possibility of reaction. Negative Gibbs free energy change which decreases with temperature shows that increasing temperature makes the reaction “more spontaneous”. However, considering the reaction as an exothermic one, temperature increase lowers the equilibrium coefficient based on Le Chatelier’s principle.

Table 2.5: Thermodynamic conditions for CH<sub>3</sub>OH dehydrogenation

$CH_3OH(g) \leftrightarrow CH_2O(g) + H_2(g)$				
T	$\Delta H$	$\Delta S$	$\Delta G$	Log(K)
K	kJ/mol	J/mol.K	kJ/mol	
273.15	91.887	108.018	62.382	-11.930
323.15	92.895	111.409	56.893	-9.197
373.15	93.805	114.030	51.254	-7.175
423.15	94.619	116.079	45.500	-5.617
473.15	95.346	117.706	39.654	-4.378
523.15	95.993	119.008	33.734	-3.369
573.15	96.566	120.054	27.757	-2.530
623.15	97.071	120.899	21.732	-1.822
673.15	97.513	121.583	15.670	-1.216
723.15	97.901	122.139	9.576	-0.692
773.15	98.238	122.589	3.458	-0.234
823.15	98.530	122.956	-2.681	0.170
873.15	98.782	123.254	-8.837	0.529
923.15	98.998	123.494	-15.006	0.849
973.15	99.180	123.686	-21.185	1.137
1023.15	99.329	123.836	-27.374	1.398
1073.15	99.446	123.948	-33.568	1.634
1123.15	99.532	124.026	-39.768	1.850
1173.15	99.595	124.081	-45.971	2.047
1223.15	99.633	124.113	-52.176	2.228
1273.15	99.649	124.125	-58.382	2.395

In this reaction, CH<sub>3</sub>OH is dehydrogenated on the surface of the catalyst. The temperature increase is clearly in favour of this reaction. At low temperatures, the low equilibrium coefficient and the high positive Gibbs free energy show that the reaction may not occur at ambient conditions. In the temperature range of 1123.15-1173.15 K, the reaction becomes fairly spontaneous.

### 2.3.2 Kinetics of Reactions and Mechanism

Regarding the kinetics and mechanism of CH<sub>3</sub>OH conversion to HCHO, the reactions will be discussed in different groups classified based on the catalysts used; Cu, Ag, and metal oxides. Since in the

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classical process of HCHO production Ag or Cu catalysts have been used, first the kinetics and mechanism related to the reactions of this group will be discussed and then metal oxide catalytic reactions for HCHO production are extensively studied.

### 2.3.2.1 Copper and Silver as Catalysts

HCHO was first produced commercially by catalytic partial oxidation of CH<sub>3</sub>OH with air over an unsupported Cu catalyst at atmospheric pressure. Later Cu was replaced by Ag due the higher yield when using the latter as the catalyst (43 to 47% yield obtained using Cu and 67% using Ag) (Roeda, 2003; Trillat, 1903). Ag is also more selective towards forming HCHO compared with Cu (Thomas, 1920). A detailed study on kinetics and mechanism of CH<sub>3</sub>OH oxidation on the surface of Cu (110) by flash decomposition spectroscopy was done by Wachs and Madix (1978b). In order to distinguish between the H atom on the hydroxyl group and the methyl group for mechanistic study, deuterium-substituted methanol (CH<sub>3</sub>OD) was used. According to this study, chemisorption of CH<sub>3</sub>OH on the surface of oxygen-free Cu was lower compared to the surface of pre-oxide Cu (110). CH<sub>3</sub>OH was oxidized while adsorbed to make methoxide (CH<sub>3</sub>O) and deuterium oxide (D<sub>2</sub>O). Adsorption occurs at 180 K on the surface and produced CH<sub>3</sub>O which was then decomposed to form HCHO and H<sub>2</sub> at 365 K. In a parallel way, but to a lesser extent, some of the chemisorbed CH<sub>3</sub>OH was oxidized to form HCOO which in turn decomposed to CO<sub>2</sub> and H<sub>2</sub>. Based on this study, both decomposition reactions occurred at a first order rate with the following rate constants:

$$k_{(H_2CO/CH_3OH)} = 5.2 \pm 1.6 \times 10^{12} e^{\left(\frac{-92.5 \pm 0.4 (kJ.mol^{-1})}{RT}\right)} (s^{-1}) \quad (2.13)$$

$$k_{(CO_2/CH_3OH)} = 8.0 \pm 2.0 \times 10^{13} e^{\left(\frac{-129.3 \pm 0.8 (kJ.mol^{-1})}{RT}\right)} (s^{-1}) \quad (2.14)$$

As the researchers mentioned, the HCHO selectivity of the oxidation of CH<sub>3</sub>OD on Cu (110) was 80% and CO<sub>2</sub> selectivity was 20%. Based on this study, the following mechanism was proposed for formation of HCHO from CH<sub>3</sub>OH on a Cu surface:





A similar study was done by Wachs and Madix (1978a) on a silver surface, using temperature programmed reaction spectroscopy (TPRS) to distinguish between hydroxyl and methyl H atoms. They showed that the pre-oxide Ag surface showed a higher ability to dissociatively chemisorb CH<sub>3</sub>OH on the catalyst surface. The process of CH<sub>3</sub>OH oxidation on Ag (110) and Cu (110) was investigated and similarities and differences are verified. Based on this research, the following reactions are observed for Ag (110):



The amount of adsorbed CH<sub>3</sub>OH increases as the O<sub>2</sub> exposure increases as a result of the interaction of the hydroxyl group of CH<sub>3</sub>OH and surface O atom. CH<sub>3</sub>O and D<sub>2</sub>O form on the catalyst surface in both cases. However, unlike Cu, the adsorbed D<sub>2</sub>O and O<sub>2</sub> on the Ag surface desorb eventually due to weaker binding compared with Cu. Based on the work by Wachs and Madix, the mechanism of CH<sub>3</sub>OH oxidation over Cu and Ag do not differ significantly. However, considering different activities of these catalysts, different selectivity for each material is achieved. The Ag surface is more active for decomposition of intermediate products compared with Cu. Reaction 2.24 between CH<sub>3</sub>O and HCHO

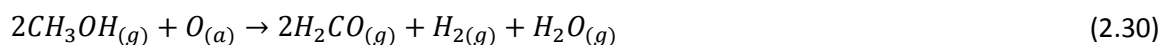
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to form  $C_2H_4O_2$  was observed on the Ag surface, but not on Cu. This is due to the higher activity of Ag for the decomposition of adsorbed  $CH_3O$  to HCHO and hence HCHO is produced at a lower substrate temperature which increases the residence time of HCHO on the Ag surface.

A study by Francis et al. (1994) on  $CH_3OH$  oxidation over oxygenated Cu surface confirmed the results achieved earlier (Leibsl et al., 1994) that the kinetics of this reaction depends on initial oxygen exposure of the Cu surface and the surface temperature. At low temperatures (150-350 K)  $CH_3OH$  reacts with adsorbed oxygen in a 2:1 ratio to generate the adsorbed  $CH_3O$  species based on the following reaction:



The formed  $CH_3O$  is stable at this temperature range, however, as the temperature increases the stability of the formed  $CH_3O$  decreases and at medium temperatures (350-400 K),  $CH_3O$  decomposes to form HCHO and  $H_2$ . Therefore, the following reaction is achieved:



At high temperatures (above 500 K), the stoichiometry of the reaction changes and  $CH_3OH$  reacts with adsorbed oxygen atom in a 1:1 ratio to form HCHO according to the following reaction:



This mechanism is confirmed by Barnes et al. (1990) because of the lack of  $H_2$  production at temperatures below 320 K and above 470 K. The presence of surface oxygen is believed to be critical in this process. It is believed that pre-oxidizing is necessary in order to activate the surface of the catalyst. But over exposure of the catalyst to  $O_2$  can poison the active sites and reduce the catalyst activity. Researchers proposed that as a dual site combination is necessary for high reactivity, overloading the catalyst with surface oxygen may disable the  $Cu^0$  sites and consequently reduce the activity of the catalyst.

The effect of  $O_2$  on the activation of Ag catalyst in the  $CH_3OH$  oxidation reaction has been investigated by many researchers. Based on the study done by Schlögl et. al. (Bao, Barth, et al., 1993; Bao, Muhler, et al., 1993; A. Nagy et al., 1998; A. J. Nagy et al., 1999), three types of O-Ag interactions can be found. (1) Dissociative chemisorption of molecular  $O_2$  leads to atomic oxygen species adsorbed on the surface of the Ag (111) and Ag (110). This species is referenced as  $O_a$  and it participates in the  $CH_3OH$  oxidation

mechanism and oxygen desorption proposed by Wachs and Madix (1978a). (2) Atomic oxygen species, referenced as  $O_\beta$ , dissolves in Ag bulk and does not participate in the oxidation reaction directly. (3) Investigation of Ag (111) at temperatures around 900 K shows that atomic oxygen segregates from the bulk catalyst to the surface by interstitial diffusion to form the atomic oxygen species referenced as  $O_\gamma$ . There is a general agreement among the researchers that  $O_\beta$  does not participate in the oxidation reaction directly. However, there is a discussion about the effects of  $O_\alpha$  and  $O_\gamma$  on the mechanism of this reaction. Some researchers have claimed that  $O_\gamma$  is likely to be the active agent in  $CH_3OH$  oxidative dehydrogenation since  $O_\alpha$  desorbs at 600 K and it cannot be the dominant species on the catalyst surface for  $CH_3OH$  oxidation at 900 K (Bao, Muhler, et al., 1993). Based on Wachs and Madix (1978a), Andreasen et al. (2003) asserted that presence of  $O_\alpha$  can justify the industrial HCHO synthesis process and  $O_\gamma$  is not necessarily involved, and that Schlögl group's conclusion about the leading role of  $O_\gamma$  is only valid at ultrahigh vacuum in which they have conducted their research. According to the study done by Waterhouse et al. (2004) the population of each atomic oxygen species depends on temperature and structure of the Ag catalyst. They declared that  $O_\alpha$  species is dominant at low temperatures, while at industrial operating conditions,  $O_\gamma$  plays the leading role and drives the process towards oxidative dehydrogenation reaction. Van Veen et al. (2002) used a temporal-analysis-of-products method and pulse experiment to investigate the mechanism of oxidative hydrogenation of  $CH_3OH$  on Ag catalyst. Based on their findings, at temperatures higher than 550 K  $O_\gamma$  provides a highly selective oxidative hydrogenation reaction to generate HCHO where at lower temperatures  $O_\alpha$  plays an important role in non-selective oxidation of  $CH_3OH$  at high surface coverage. However, according to their research, low coverage of  $O_\alpha$  may improve the selectivity significantly. Some researchers have investigated the effects of different operating conditions on oxidation of  $CH_3OH$  and HCHO production using Ag catalyst. Based on a study done by Cao and Gavrilidis (2005), as reactor temperature increases from 500 K to 620 K,  $CO_2$  production rate increases and HCHO selectivity decreases. This is in agreement with the results achieved by Van Veen et al. (2002) as they indicated a strong role of  $O_\alpha$  in  $CO_2$  production and low selectivity towards HCHO which is unlike the role of  $O_\gamma$ . As the reactor temperature exceeds 620 K, HCHO selectivity increases and  $CO_2$  concentration decreases which proves the role of  $O_\gamma$  as indicated in the results by Van Veen et al. (2002). The  $O_2$  concentration in the gas phase was another aspect which was investigated by Cao and Gavrilidis (2005). As the  $O_2$  concentration increased from 1.6 to 3.68 Vol%,  $CO_2$  concentration and  $CH_3OH$  conversion increased and HCHO selectivity decreased.

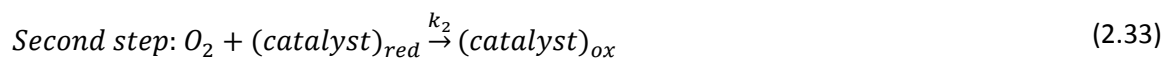
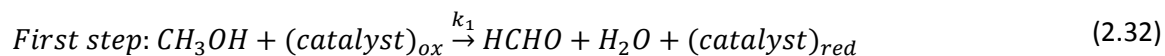
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Qian et al. (2003) investigated the effects of temperature on CH<sub>3</sub>OH oxidation process and HCHO yield. HCHO yield gradually increased until it reached its maximum at 923 K. At higher temperatures, HCHO yield decreased which was accompanied by an increase in the yield of CO and H<sub>2</sub> which showed HCHO decomposition. They also investigated the effects of residence time on CH<sub>3</sub>OH conversion and HCHO selectivity which showed that in the range of 0.06-0.45 s residence time, CH<sub>3</sub>OH conversion increased while HCHO selectivity decreased at higher residence times. The effect of water vapour was also quantified by increasing the H<sub>2</sub>O/CH<sub>3</sub>OH molar ratio from 0 to 2. While CH<sub>3</sub>OH conversion constantly increased with H<sub>2</sub>O/CH<sub>3</sub>OH ratio, HCHO selectivity reached its maximum at around the molar ratio of 0.75.

While many studies have discussed the kinetics, mechanism, and optimum operating conditions of the oxidation reaction on Cu and Ag catalysts, partial oxidation of CH<sub>3</sub>OH on metal oxides has attracted higher attention. In the following section we discuss the application of metal oxides as catalysts.

### 2.3.2.2 Metal Oxides as Catalysts

Bailey and Craver (1921) patented a method of CH<sub>3</sub>OH oxidation into HCHO using metal oxides (V<sub>2</sub>O<sub>5</sub> in particular) as catalysts. Since then many researchers have investigated the kinetics and mechanism of the reactions which occur on the surface of these metal oxides in the presence of CH<sub>3</sub>OH and other materials. In a kinetic study done by Bhattacharyya et al. (1967), the CH<sub>3</sub>OH oxidation reaction was investigated at 519.15 K, 537.15 K and 554.15 K using V<sub>2</sub>O<sub>5</sub> as the catalyst. Based on their study, the Mars-van Krevelen redox mechanism (Mars & van Krevelen, 1954) was suggested in which CH<sub>3</sub>OH reduced the catalyst and formed HCHO and water (equation 2.32) and in the next step, O<sub>2</sub> exposure re-oxidized the reduced catalyst (equation 2.33):



According to their study, the following rate of reaction was obtained assuming a steady state pattern:

$$r = \frac{2k_1k_2P_mP_{O_2}}{k_1P_m + 2k_2P_{O_2}} \quad (2.34)$$

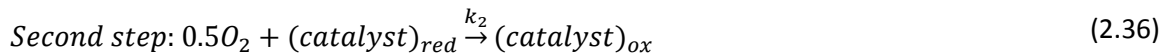
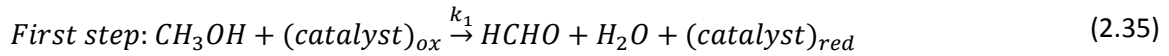


Where  $k_1$  and  $k_2$  refer to rate constants for the first and second step, respectively, which were evaluated at each temperature at which the reaction was investigated. The results are given in Table 2.6:

Table 2.6: Values for the first and the second rate constants for  $\text{CH}_3\text{OH}$  partial oxidation (Bhattacharyya et al., 1967)

Rate constants moles of $\text{HCHO}/(\text{moles of}$ $\text{CH}_3\text{OH.s. m}^3.\text{kPa})$	Temperature (K)		
	519.15	537.15	554.15
$k_1$	2.96	5.93	15.3
$k_2$	0.26	0.49	1.09

In a more detailed study, Mann and Dosi (1973) studied the kinetics of the same reaction on vanadium oxide-molybdenum oxide ( $\text{V}_2\text{O}_5\text{-MoO}_3$ ) in the range of temperatures (523.15 K to 803.15 K), catalyst metal oxides ratio (0 to 1), and  $\text{CH}_3\text{OH}$  concentration in air (5 to 10%). Based on their study, the best results were obtained at 733.15 K and by  $\text{V}_2\text{O}_5/\text{MoO}_3$  weight ratio of 1/4. Optimum  $\text{CH}_3\text{OH}$  concentration in air is detected to be 8%. The same two-stage redox mechanism which was applied by Bhattacharyya et al. (1967), has been considered in this study, too:



Based on this research, the following rate equation is achieved for this reaction:

$$r = \frac{k_1 P_m}{1 + \left( k_1 P_m / 2k_2 P_{\text{O}_2}^{0.5} \right)} \quad (2.37)$$

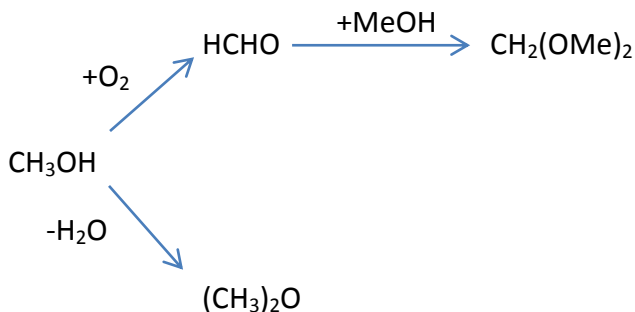
Where  $k_1$  and  $k_2$  are the rate constants for first and second step, respectively. The value of  $k_1$  and  $k_2$  can be found based on the Arrhenius-law-based equations given for these two rate constants:

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$$\log k_1 = 0.773 - \frac{2.358 \times 10^3}{T(K)} \quad (2.38)$$

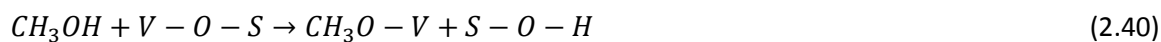
$$\log k_2 = 0.157 - \frac{2.04 \times 10^3}{T(K)} \quad (2.39)$$

Tatibouët and Germain (1981) investigated MoO<sub>3</sub> as catalyst for CH<sub>3</sub>OH oxidation reaction. They used dilute CH<sub>3</sub>OH–O<sub>2</sub> mixture (CH<sub>3</sub>OH /O<sub>2</sub>/He molar ratio = 8.2/19.7/72) in the range of 473.15-573.15 K and at atmospheric pressure and concluded that reaction may take two different paths based on the polymorphic crystalline form of MoO<sub>3</sub>. On hexagonal surface of MoO<sub>3</sub> surface, CH<sub>3</sub>OH goes through dehydration path and dimethyl ether (DME) is produced. Whereas on orthorhombic MoO<sub>3</sub>, CH<sub>3</sub>OH goes through partial oxidation to form HCHO which may react further with CH<sub>3</sub>OH to form its acetal:



In an FTIR study, Busca (1989) observed formation of CH<sub>3</sub>O groups upon CH<sub>3</sub>OH adsorption over V<sub>2</sub>O<sub>5</sub> catalyst. CH<sub>3</sub>O groups in turn oxidized to form dioxymethylene (H<sub>2</sub>CO<sub>2</sub>). HCHO adsorption on the surface of V<sub>2</sub>O<sub>5</sub> also formed H<sub>2</sub>CO<sub>2</sub> which implies that desorption of H<sub>2</sub>CO<sub>2</sub> formed during CH<sub>3</sub>OH oxidation may lead to HCHO formation. However, Forzatti et al. (1997) found that at low temperatures (< 435 K) and high CH<sub>3</sub>OH availability, H<sub>2</sub>CO<sub>2</sub> may react with CH<sub>3</sub>OH to form Dimethoxymethane (DMM) while it can preferably desorb as HCHO at high temperatures (> 435 K) and low CH<sub>3</sub>OH availability. Therefore, high selectivity towards HCHO is achieved over weak HCHO binding sites of the catalyst. The type of support- such as acidic, basic, or redox- also affects the reaction pathways substantially. Deo and Wachs (1994), showed that different products, such as HCHO, DMM, DME, C<sub>2</sub>H<sub>4</sub>O<sub>2</sub>, and oxides of carbon can be produced over different supports. These products depend upon the pathways taken by surface CH<sub>3</sub>O over different supports. Acid sites cause CH<sub>3</sub>OH dehydration and DME production. Basic sites oxidize CH<sub>3</sub>OH to CH<sub>3</sub> and further convert it into CO and CO<sub>2</sub>. On redox

sites, partial oxidation occurs and HCHO, DMM and C<sub>2</sub>H<sub>4</sub>O<sub>2</sub> can be produced. The desired pathway for HCHO formation is shown below in equations 40 to 42 (Burcham & Wachs, 1999).



SiO<sub>2</sub>, ZrO<sub>2</sub>, TiO<sub>2</sub>, Nb<sub>2</sub>O<sub>5</sub>, and Al<sub>2</sub>O<sub>3</sub> supported V<sub>2</sub>O<sub>5</sub> were investigated by Deo and Wachs (1994) and the best results for each catalyst based on HCHO selectivity are tabulated in Table 2.7:

**Table 2.7: Effects of different supports on the V<sub>2</sub>O<sub>5</sub> performance in CH<sub>3</sub>OH partial oxidation**

Sample	TOF (s <sup>-1</sup> )	Selectivity%				
		FA	MF	DMM	DME	CO <sub>2</sub> and CO
5% V <sub>2</sub> O <sub>5</sub> /SiO <sub>2</sub>	15.1×10 <sup>-3</sup>	88	0	2	8	2
5% V <sub>2</sub> O <sub>5</sub> /ZrO <sub>2</sub>	1.6	94	5	Trace	Trace	Trace
1% V <sub>2</sub> O <sub>5</sub> /TiO <sub>2</sub>	2	99+	0	0	Trace	0
2% V <sub>2</sub> O <sub>5</sub> /Nb <sub>2</sub> O <sub>5</sub>	1.2	97	0	1	2	0
20% V <sub>2</sub> O <sub>5</sub> /Al <sub>2</sub> O <sub>3</sub>	6.8×10 <sup>-2</sup>	46	0	3	50	1

According to these results, the highest HCHO selectivity (more than 99%) and the highest turnover frequency (2 s<sup>-1</sup>) was achieved by using TiO<sub>2</sub> as the support. Nb<sub>2</sub>O<sub>5</sub> also showed high selectivity towards HCHO (97%). Based on the results gained, the reactivity of the supported V<sub>2</sub>O<sub>5</sub> as CH<sub>3</sub>OH oxidation catalyst depended on the strength of the V-O-Support bond, meaning that the weaker the V-O-Support bond was the higher catalytic reactivity was achieved.

Adkins and Peterson (1931) reported 38% O<sub>2</sub> conversion and 100% O<sub>2</sub> selectivity for HCHO production over MoO<sub>3</sub> catalyst using CH<sub>3</sub>OH partial oxidation. Higher O<sub>2</sub> conversion was achieved over FeO<sub>2</sub> but mostly CO<sub>2</sub> was formed which meant complete oxidation of CH<sub>3</sub>OH. Mixed oxides of Fe and Mo showed high conversion (>90%) and high selectivity for HCHO and CO. Soares et al. (2001) studied the effects of Mo to Fe atomic ratio by comparing industrial Fe-Mo oxides catalyst (Mo/Fe = 3) with stoichiometric composition (Mo/Fe = 1.5). Higher Mo/Fe atomic ratio resulted in higher surface area and higher selectivity of the catalyst towards HCHO production. However, the activity of the catalyst did not change significantly. Therefore it can be concluded that the Fe<sub>2</sub>(MoO<sub>4</sub>)<sub>3</sub> stoichiometric phase is the active phase of the catalyst. Wachs' group (Routray et al., 2010) showed that if excess MoO<sub>x</sub> is

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added to crystalline  $\text{Fe}_2(\text{MoO}_4)_3$  it is deposited as surface monolayer which improves the catalytic performance towards HCHO production. Contrary to Soares et al. (2001), Wach's group (Routray et al., 2010) found that stoichiometric  $\text{Fe}_2(\text{MoO}_4)_3$  (Mo/Fe=1.5) has higher surface area compared with catalysts with excess molybdenum oxide. Wach's group (Routray et al., 2010) achieved highest selectivity towards HCHO by using the catalyst Mo/Fe = 2. However, the turnover frequency was higher when using the catalyst Mo/Fe = 1.7. Based on the mechanism suggested for  $\text{CH}_3\text{OH}$  oxidation reaction over  $\text{FeO}_2$  and the one for Fe-Mo oxide by Wach's group (Routray et al., 2010), they have concluded that due to the dependence of the HCHO production pathway on the presence of bulk lattice oxygen and considering the incapability of  $\text{FeO}_2$  to supply bulk lattice oxygen, the chance of formation of HCHO on  $\text{FeO}_2$  is low. On the other hand, considering the high availability of bulk lattice oxygen on the surface of the Fe-Mo catalyst, HCHO and probably DMM may form on the surface of the Fe-Mo catalyst. The following rate of reaction is suggested by the researchers:

$$r = k_{rds} K_{ads} P_{\text{CH}_3\text{OH}} N_s \quad (2.43)$$

Where  $k_{rds}$  is the first order rate constant ( $\text{s}^{-1}$ ),  $K_{ads}$  is the  $\text{CH}_3\text{OH}$  adsorption equilibrium constant ( $\text{kPa}^{-1}$ ),  $P$  represents  $\text{CH}_3\text{OH}$  partial pressure (kPa), and  $N_s$  is the surface density of redox sites ( $\mu\text{mol.m}^{-2}$ ).

Based on the studies done by many researchers, many improvements and developments have occurred in this field. However, there are still some issues that can be addressed.

### 2.4 Latest Advancements in Formaldehyde Production

The recent studies done on HCHO production can mainly be divided into two sections; some of these studies investigated the effects of different modifications done on industrial catalysts for partial oxidation of  $\text{CH}_3\text{OH}$  to produce HCHO. Bell's group (Vining et al., 2011, 2012) investigated the effects of additional zirconium dioxide ( $\text{ZrO}_2$ ) and cerium dioxide ( $\text{CeO}_2$ ) on  $\text{SiO}_2$  supported  $\text{V}_2\text{O}_5$  catalyst for  $\text{CH}_3\text{OH}$  partial oxidation. They found out that by using these bi-layered catalysts, the turnover frequency increased by nearly two orders of magnitude for  $\text{CH}_3\text{OH}$  oxidation compared with mono-layered catalyst. They proposed that the addition of  $\text{CeO}_2$  and  $\text{ZrO}_2$  can affect the rate-limiting step in the mechanism of  $\text{CH}_3\text{OH}$  partial oxidation which is hydrogen abstraction from V-OCH<sub>3</sub> group. Other researchers have studied other methods for HCHO production. Some recent studies (Beznis et al., 2011; Carlsson et al., 2012; Polnišer et al., 2011; Wang et al., 2012; J. Zhang et al., 2012a, 2012b) have

targeted the  $\text{CH}_4$  reactivity and the selectivity towards HCHO production in the process of direct oxidation of  $\text{CH}_4$  to HCHO. Exposing the reactants to  $\text{NO}_2$  using low residence time (J. Zhang et al., 2012a, 2012b), using periodic operating conditions (Carlsson et al., 2012), and using Cu-Fe pyrophosphate as the catalyst (Polnišar et al., 2011) were suggested. However, simultaneous increase of  $\text{CH}_4$  conversion and HCHO selectivity remained unsolved. Lou et al (Lou, Wang, et al., 2007; Lou, Zhang, et al., 2007) suggested using  $\text{C}_2\text{H}_6$  instead of  $\text{CH}_4$ . They used a mesoporous  $\text{SiO}_2$  supported  $\text{MoO}_3$  catalyst to achieve 30% HCHO yield by oxidation of  $\text{C}_2\text{H}_4$  to HCHO at 836 K. Using a double bed catalyst for simultaneous dehydrogenation of  $\text{C}_2\text{H}_6$  to  $\text{C}_2\text{H}_4$  and oxidation of  $\text{C}_2\text{H}_4$  to HCHO, maximum of 14% HCHO yield was achieved. Since anhydrous dehydrogenation of  $\text{CH}_3\text{OH}$  to HCHO has been an interesting process for researchers because it avoids the formation of water, a recent study was dedicated to the theoretical investigation of this process by using a web-based application called “CatApp”(Lausche et al., 2012). Based on this study, the mechanism of  $\text{CH}_3\text{OH}$  dehydrogenation reaction and different binding energies were investigated over different metal surfaces. They claimed that on the surface of Pt, Pd, and Rh, there is a higher tendency to dehydrogenate HCHO compared with desorption of this chemical. This may lead to CO formation on the surface of these metals as indicated in a study by Tolmachev et al. (2009).

Although hydrogenation of  $\text{CO}_2$  to HCHO was investigated before (Lunov et al., 2001), a high yield of HCHO was never achieved. Recently, Nakata et al. (2014) succeeded to produce HCHO by reduction of  $\text{CO}_2$  in an electrochemical procedure. A boron-doped diamond (BDD) electrode was used for electrochemical process under ambient conditions and  $\text{HCOOH}$  was produced as an intermediate. Relatively high faradaic efficiency (up to 74%) was achieved by using  $\text{CH}_3\text{OH}$ , aqueous salt, or seawater as the electrolyte. Other electrodes such as glassy carbon, tin (Sn), Cu, Ag, and tungsten (W) were used for comparison. The highest faradaic efficiency (74%) for HCHO was achieved by using BDD electrode while  $\text{HCOOH}$  faradaic efficiency was 15%. Faradaic efficiency for HCHO was about 20% using glassy carbon electrode. No HCHO production was noticed using metal electrodes.

Hydrogenation of  $\text{CO}_2$  is not the only method of producing HCHO from  $\text{CO}_2$ . Recently, in a study by Lambert's group (Gómez-Ramírez et al., 2014), low temperature HCHO production was achieved by partial oxidation of  $\text{C}_2\text{H}_6$  by using  $\text{CO}_2$  as the oxygen source. The catalyst for this process was  $\text{V}_2\text{O}_5/\text{Al}_2\text{O}_3$  which was dispersed on barium titanate ( $\text{BaTiO}_3$ ). The catalyst was loaded on a dielectric barrier discharge reactor. Catalyst was activated at near ambient temperature by plasma-excited  $\text{CO}_2$  which led to generation of an activated oxide surface. Through the activated catalyst, oxy-

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dehydrogenation of  $C_2H_6$  took place to produce  $C_2H_4$ . Discharge-activated  $CO_2$  molecules reacted with the produced  $C_2H_4$  molecules at this level to generate acetaldehyde ( $CH_3CHO$ ) which in turn converted to  $HCHO$  through oxidative decomposition. 11.4%  $HCHO$  yield was achieved through this process upon complete conversion of  $C_2H_6$ .

In a recent study, the industrial process of  $HCHO$  production (starting from natural gas steam reforming followed by  $CH_3OH$  synthesis and methanol ballast process to form  $HCHO$ ) was simulated. An exergy analysis was done to measure the exergy efficiency of the process. It is shown that the industrial process suffers from ~57% losses of exergy (Bahmanpour et al., 2014). Therefore, there is a need to improve the  $HCHO$  production process.

## 2.5 Conclusion

Formaldehyde ( $HCHO$ ) is a highly valuable chemical in the global economy. Methanol synthesis followed by partial oxidation into  $HCHO$  is still the leading technology after more than a century; however, new catalysts and process modifications have been made over the years to increase conversion and selectivity of  $HCHO$  synthesis. The mechanism of  $CH_3OH$  partial oxidation showed that the reaction proceeds via oxidation of gaseous  $CH_3OH$  by adsorbed oxygen or hydroxyl species on the catalyst surface to form adsorbed  $CH_3O$  species, which decomposes to form  $HCHO$  depending upon the reaction temperature. Pre-oxidation of Cu or Ag catalysts to form metal oxides were found to be more active than metals as catalysts for partial oxidation of  $CH_3OH$  to  $HCHO$ . However, since  $CH_3O$  is stable at low temperatures, the reaction is carried out at above 623.15 K. Different types of atomic oxygen species were found to be active in the  $HCHO$  production reactions when Ag catalyst was used. Mars-van Krevelen redox mechanism was suggested for the formation of  $HCHO$  on the surface of metal oxide catalysts by some researchers and the effects of the addition of the excess  $MoO_3$  in Fe-Mo oxide catalysts were argued. Many recent attempts at improving the  $HCHO$  synthesis process have focused on developing new redox catalysts to improve the  $CH_3OH$  partial oxidation process. Other proposed methods for  $HCHO$  production such as hydrogenation of CO and  $CO_2$ , and  $C_2H_4$  oxidation hold much promise but extensive research is required in catalyst development for these processes.

## 2.6 References

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Monash University

### Declaration for Thesis Chapter 3

#### Declaration by candidate

In the case of Chapter 3, the nature and extent of my contribution to the work was the following:

Nature of contribution	Extent of contribution (%)
Initiation	50
Experiments	100
Analysis	80
Writing	70*

The following co-authors contributed to the work. If co-authors are students at Monash University, the extent of their contribution in percentage terms must be stated:

Name	Nature of contribution	Extent of contribution (%) for student co-authors only
Akshat Tanksale	Initiation	
	Analysis	
	Reviewing and Editing the Paper	
Andrew Hoadley	Initiation	
	Analysis	
	Reviewing and Editing the Paper	

The undersigned hereby certify that the above declaration correctly reflects the nature and extent of the candidate's and co-authors' contributions to this work.

Candidate's  
Signature



Date

24/09/15

Main  
Supervisor's  
Signature



Date

24/09/15

## Chapter 3:

# Formaldehyde Production via CO Hydrogenation in the Aqueous Phase

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### Abstract

*In order to study a new chemical process, one may need to study the possibility of the process theoretically and experimentally first before any further studies. Once the possibility has been approved, there might be a chance to improve the production yield by optimizing the operating conditions. In this chapter, the thermodynamic conditions of the carbon monoxide hydrogenation reaction to formaldehyde are discussed using the HSC Chemistry <sup>TM</sup> software and experimental data for both gas phase and aqueous phase reactions. The feasibility of producing formaldehyde directly from synthesis gas is proved. However, gas phase hydrogenation of carbon monoxide into formaldehyde is thermodynamically limited and therefore, results in low carbon monoxide conversion of only  $1.02 \times 10^{-4}$  %. On the other hand, the aqueous phase hydrogenation of carbon monoxide into formaldehyde is found to be thermodynamically favourable and kinetically limited. Highest carbon monoxide conversion of 19.14% and selectivity of 100% is achieved by using Ru-Ni/Al<sub>2</sub>O<sub>3</sub> catalyst at 353 K and 100 bar in the aqueous phase. The rapid hydration of formaldehyde in the aqueous phase to form methylene glycol shifts the carbon monoxide hydrogenation reaction equilibrium towards formaldehyde formation. Increasing the pressure and stirring speed increase the production yield of formaldehyde.*

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### 3.1 Introduction

As discussed in chapter 2, formaldehyde (HCHO) is produced in three stages – (a) Steam reforming of natural gas to produce syngas (Table 3.1, Eq.3.1), (b) Methanol (CH<sub>3</sub>OH) Synthesis (Eq. 3.2) and (c) partial oxidation of CH<sub>3</sub>OH to produce HCHO (Eq. 3.3). Alternatively, HCHO is produced via dehydrogenation of CH<sub>3</sub>OH (Eq. 3.4) (Reuss et al., 2003; Walker, 1967). However, these are all high temperature and high pressure reactions which require combustion, compression and large process units for purification, which are the root cause of energy losses (Li et al., 2014; Rosen & Scott, 1988). We have recently shown that this series of processes from natural gas to HCHO production suffers from ~57% losses in exergy (i.e. energy quality) (Bahmanpour et al., 2014). Given the large amount of HCHO produced in the world, when combined with the high losses in exergy, leads to high energy losses and also high CO<sub>2</sub> emissions, globally. As mentioned in chapter 2, the effort of the researchers to overcome this issue by finding ways to produce HCHO directly from natural gas by partial oxidation of CH<sub>4</sub> (Eq. 3.5) was not successful due to low CH<sub>4</sub> conversion and/or poor selectivity towards formation of HCHO (Blair & Wheeler, 1923; Dake & Chaudhari, 1985; Purwanto et al., 1996) since the rate of HCHO decomposition into CO and H<sub>2</sub> is much greater than the rate of partial oxidation of CH<sub>4</sub>, especially at temperatures in excess of 373 K, which means that in order to produce HCHO selectively, one must limit the conversion of CH<sub>4</sub> in Eq. 3.5 to a very small value. Moreover, if we consider that natural gas is a valuable energy resource, it becomes evident that an alternative feedstock for HCHO production is much needed. An alternative is direct conversion of synthesis gas into HCHO. Syngas can be produced from a range of sources including biomass and allows the mole fractions of CO to H<sub>2</sub> to be controlled more easily through the use of H<sub>2</sub>O and CO<sub>2</sub> which is helpful in climate change abatement (Chan & Tanksale, 2014).

## Chapter 3

**Table 3.1: Chemical reactions used for the production of HCHO in the gas phase**

Reaction Name	Reaction Stoichiometry	Equation No.
Steam Reforming	$CH_4 + H_2O \rightleftharpoons CO + 3H_2$ $\Delta\hat{H}_r^o = +206 \text{ kJ.mol}^{-1}$	(3.1)
Methanol Synthesis	$CO + 2H_2 \rightleftharpoons CH_3OH$ $\Delta\hat{H}_r^o = -91 \text{ kJ.mol}^{-1}$	(3.2)
Methanol Partial Oxidation	$CH_3OH + \frac{1}{2}O_2 \rightarrow HCHO + H_2O$ $\Delta\hat{H}_r^o = -159 \text{ kJ.mol}^{-1}$	(3.3)
Methanol Dehydrogenation	$CH_3OH \rightleftharpoons HCHO + H_2$ $\Delta\hat{H}_r^o = +84 \text{ kJ.mol}^{-1}$	(3.4)
Methane Partial Oxidation	$CH_4 + O_2 \rightarrow HCHO + H_2O$ $\Delta\hat{H}_r^o = -319 \text{ kJ.mol}^{-1}$	(3.5)
Syngas to Formaldehyde in Gas Phase	$CO + H_2 \rightleftharpoons HCHO$ $\Delta\hat{H}_r^o = -5.4 \text{ kJ.mol}^{-1}$ , $\Delta G^o = +34.6 \text{ kJ.mol}^{-1}$	(3.6)

Gas phase hydrogenation of CO to produce HCHO (Equation 3.6) suffers from positive Gibbs free energy change of the reaction (Newton & Dodge, 1933). Only trace amount of HCHO in the product has been reported with the highest CO conversion of 0.2% (Morgan et al., 1932). Therefore, direct conversion of synthesis gas into HCHO has not been studied extensively. In this chapter, hydrogenation of CO into HCHO in a slurry reactor is presented as a viable alternative. By comparing with gas phase conversion in a fixed bed reactor, it is demonstrated that the thermodynamic limitation can be overcome in the slurry reactor. A low temperature active catalyst is desirable for the slurry phase reaction as the reaction was found to be favourable below 373 K. Generally, Ni Pd and Ru are considered as active hydrogenation catalysts in the literature for many reactions (Abdur-Rashid et al., 2002; Araki & Ponc, 1976; Fatsikostas & Verykios, 2004; Goodman et al., 1980; Hashiguchi et al., 1995; Jessop et al., 1994; Kidambi et al., 2004; Kroll et al., 1996; Niu et al., 2001; Noyori & Hashiguchi, 1997). Based on the density functional theory (DFT) studies done in the literature, Pd and Ni have been shown to produce HCHO as an intermediate of CO hydrogenation to produce CH<sub>3</sub>OH (Neurock, 1999; Remediakis et al., 2004). Previous studies have also shown that bi-metallic Pd-Ni catalyst has better reducibility and higher metal surface area (Tanksale et al., 2008). Therefore, Pd-Ni and Ru-Ni supported on  $\gamma$ -Al<sub>2</sub>O<sub>3</sub> were used as the catalysts in this study.



## 3.2 Experiments and Methods

### 3.2.1 Catalyst Preparation and Characterization

The catalysts used in this process were produced by the wet impregnation method followed by calcination. Nickel nitrate ( $(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O}$ , Sigma-Aldrich), ruthenium chloride ( $\text{RuCl}_3 \cdot x\text{H}_2\text{O}$ , Sigma-Aldrich) and palladium nitrate (10 wt%  $\text{Pd}(\text{NO}_3)_2$  in 10wt% nitric acid, Sigma-Aldrich) were used as Ni, Ru and Pd precursors, respectively. The desirable amounts of Ni and noble metal precursors were added simultaneously to commercial  $\gamma\text{-Al}_2\text{O}_3$  suspended in 20 ml water. The mixture was stirred at 333 K for 6 h. The suspension was then dried at 373 K overnight. The dried catalyst was calcined at 873 K for 6 h.

Fixed bed reactor tests were started by reducing the catalyst *in-situ* prior to the experiment by flowing  $50 \text{ ml} \cdot \text{min}^{-1}$  of  $\text{H}_2$  through the catalyst bed at 673 K for 5 h followed by purging with Ar at 673 K for 1 h. The catalyst bed was cooled to room temperature under Ar flow overnight. For the slurry reactor tests, the same procedure of catalyst reduction was carried out *ex-situ*.

BET surface area of the support was measured by the  $\text{N}_2$  physisorption method in Micrometric ASAP2020 at 77 K. CO chemisorption was used to measure the amount of active sites of each catalyst and the metal dispersion percentage using ASAP2020 (Micrometrics). X-ray Florescence (XRF) spectroscopy was used to determine the actual mass percentage of the metals using an Ametek Spectro iQ II XRF. X-ray Diffraction (XRD) patterns of fresh calcined and reduced catalysts were recorded using a REGAKU MiniFlex 600 X-ray diffraction instrument equipped with a Ni-filtered  $\text{Cu K}\alpha$  radiation in order to study the situation of the catalyst before and after reduction as well as the catalyst support stability. XRD patterns were gained for  $2\theta$  using step size of 0.01. Transmission Electron Microscopy (TEM) was used to evaluate the particle size of each promoter (Pd and Ru) in the fresh calcined catalysts and ImageJ 1.48 (National Institutes of Health) was used to generate the particle size distribution. The catalyst characterization results are discussed in the results and discussion section.

### 3.2.2 Fixed Bed Reactor

A schematic diagram of the custom made fixed bed reactor setup is illustrated in Figure 3.1. CO and H<sub>2</sub> gases were mixed in 1:1 mole ratio (30 ml.min<sup>-1</sup> each) using mass flow controllers in a gas manifold prior to feeding it in the fixed bed reactor. The reactor was Swagelok ¼" OD seamless tube in which 1 g of catalyst was fixed using quartz wool. The reactor was heated by a tube furnace which was used for *in-situ* reduction of the catalyst prior to the test. A back pressure regulator controlled the desired pressure upstream in the system. The gas stream leaving the back pressure regulator was passed through a scrubber containing 40 ml of 5 vol% CH<sub>3</sub>OH in water to recover HCHO. CH<sub>3</sub>OH was used in order to prevent hydrated HCHO molecules from polymerization. At the end of the run, the HCHO concentration was measured using the photometric cell test kit (Merck Millipore) in a DR 5000™ UV/VIS spectrophotometer (HACH Company, USA) at 575nm wavelength using a chromotropic acid method (Kennedy, 1994). The tests were conducted at different operating conditions. A range of pressures (20bar, 40bar, 85bar, and 117bar) and temperatures (293 K and 313 K) were tested.

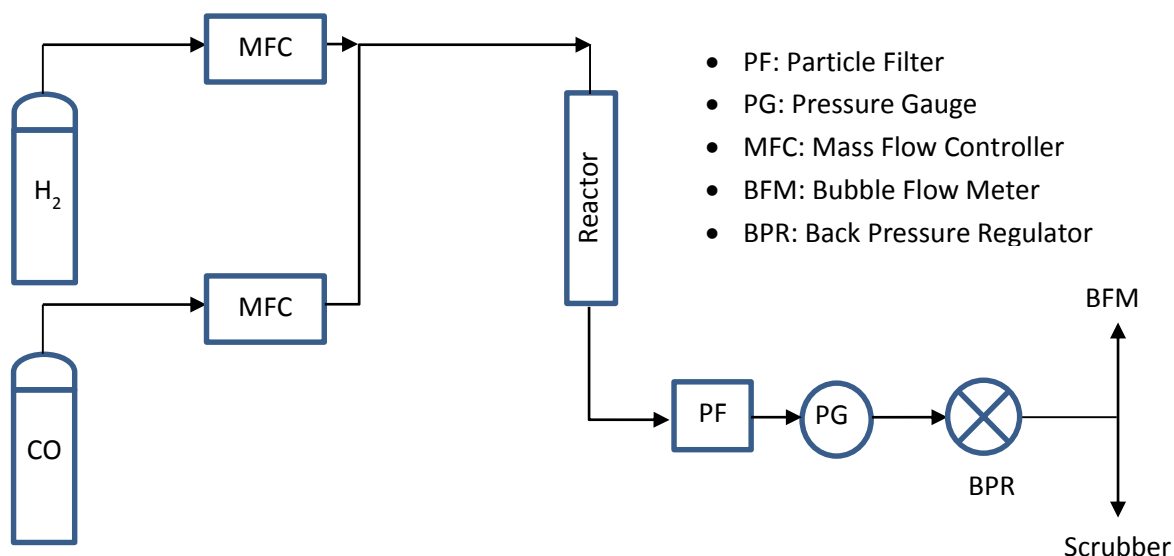


Figure 3.1: Schematic diagram of the fixed bed reactor setup

### 3.2.3 Slurry Reactor

The slurry reactor (batch reactor), illustrated in Figure 3.2, was charged with 40 ml of 5 vol% CH<sub>3</sub>OH in water. 1 g of the desired catalyst was reduced *ex-situ* before adding it to the batch reactor. The reactor was subsequently pressurized to 100 bar with equimolar mixture of CO and H<sub>2</sub>. Subsequently the reactor was heated to the desired temperature (293 K, 333 K, 353 K, 373 K, or 403 K). During the run, liquid samples were collected at regular intervals through the dip tube. the HCHO concentration was measured using the photometric cell test kit (Merck Millipore) with a DR 5000™ UV/VIS spectrophotometer (HACH Company, USA) at 575 nm wavelength using a chromotropic acid method. In this method, 4.5 ml of 75% sulphuric acid (H<sub>2</sub>SO<sub>4</sub>) was measured in a plastic tube. 0.1 g of chromotropic acid disodium salt dehydrate (C<sub>10</sub>H<sub>10</sub>Na<sub>2</sub>O<sub>10</sub>S<sub>2</sub>) was added to the tube and was shaken vigorously. 3 ml of the sample (diluted if required) was added to the mixture. The intensity of the violet colour resulting from the reaction was checked using the spectrophotometer and quantified based on the pre-prepared calibration curve.

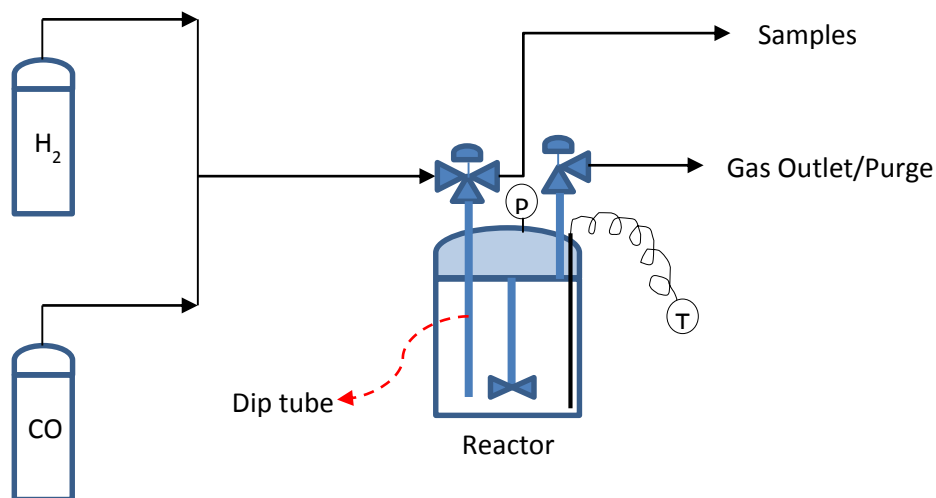


Figure 3.2: Schematic diagram of the slurry reactor setup

## 3.3 Investigation of the Thermodynamic Conditions

Hydrogenation of CO into HCHO in the gas phase and aqueous phase was thermodynamically investigated using the HSC Chemistry® version 7.11 (Outotec, Finland) software and published data

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(Oelkers et al., 1995) for a wide range of temperatures (298-623 K) and pressures (50-500 bar). The results (Table 3.2) show that the Gibbs free energy ( $\Delta G$ ) of the reaction is positive in the gas phase and it increases with increasing temperature. Therefore the reaction is non-spontaneous at all temperatures above 298 K. The equilibrium constant of the gas phase reaction is significantly low ( $8.76 \times 10^{-7} \text{ mol}^{-1}$  at 298 K) which, along with positive  $\Delta G$ , suggests that the forward reaction is not favourable. In the aqueous phase however,  $\Delta G$  of the reaction is negative at low temperatures and the equilibrium constant is high ( $17.33 \text{ mol}^{-1}$  at 298 K), which suggests that HCHO formation is favourable in the aqueous phase. This is because the heat of solution of HCHO is  $-62 \text{ kJ.mol}^{-1}$  (Reuss et al., 2003) which is relatively high while the heat of solution of  $\text{H}_2$  and CO is insignificant in comparison. This results in the improvement of the equilibrium conversion. The effect of pressure on the Gibbs free energy of the reaction in the aqueous phase is shown in Figure 3.3. The data presented in this Figure are calculated based on the interpolation of experimental data presented by Oelkers et al (Oelkers et al., 1995). It can be seen that as the operating pressure increases, the Gibbs free energy of the reaction decreases. Although it is concluded that higher pressures are favourable for this reaction, the operating pressure in this study was limited to 100 bar by the maximum available gas cylinder pressure.

**Table 3.2: Thermodynamic conditions of CO hydrogenation reaction to HCHO**

$\text{CO} + \text{H}_2 \leftrightarrow \text{CH}_2\text{O}$						
Gas Phase				Aqueous Phase		
$T$	$X_e$	$\Delta G$	$K$	$X_e$	$\Delta G$	$K$
K	%	$\text{kJ.mol}^{-1}$	$\text{mol}^{-1}$	%	$\text{kJ.mol}^{-1}$	$\text{mol}^{-1}$
298	$4.38 \times 10^{-3}$	34.565	$8.760 \times 10^{-7}$	31.40	-7.071	17.332
323	$4.67 \times 10^{-3}$	37.297	$9.348 \times 10^{-7}$	18.42	-5.648	8.186
373	$4.88 \times 10^{-3}$	42.933	$9.764 \times 10^{-7}$	3.86	-1.297	1.519
423	$4.86 \times 10^{-3}$	48.705	$9.711 \times 10^{-7}$	$5.84 \times 10^{-1}$	4.853	$2.52 \times 10^{-1}$
473	$4.70 \times 10^{-3}$	54.586	$9.405 \times 10^{-7}$	$8.88 \times 10^{-2}$	12.343	$4.30 \times 10^{-2}$
523	$4.49 \times 10^{-3}$	60.554	$8.983 \times 10^{-7}$	$1.50 \times 10^{-2}$	21.129	$8.00 \times 10^{-3}$
573	$4.26 \times 10^{-3}$	66.592	$8.523 \times 10^{-7}$	$1.72 \times 10^{-3}$	31.171	$1.00 \times 10^{-3}$
623	$4.03 \times 10^{-3}$	72.685	$8.068 \times 10^{-7}$	$3.95 \times 10^{-4}$	43.053	$2.46 \times 10^{-4}$

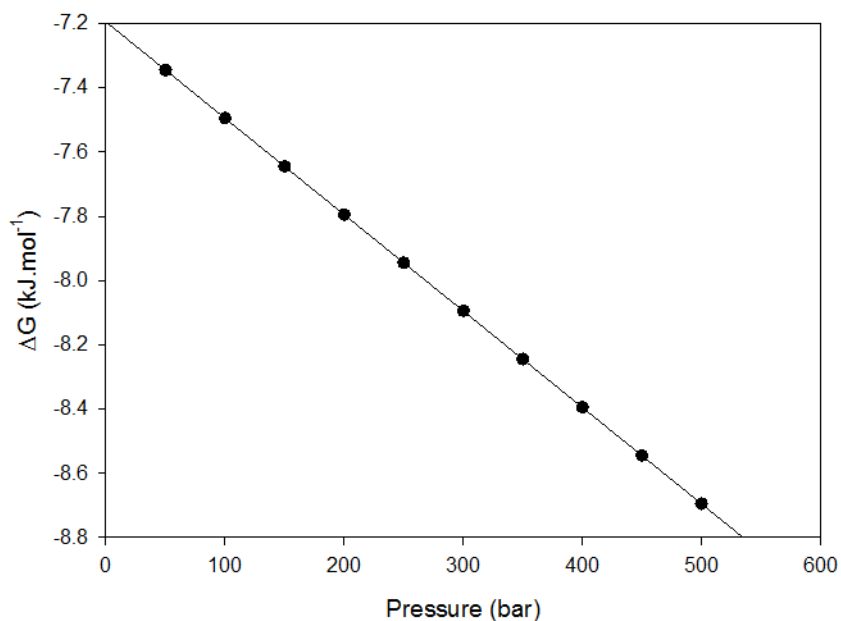


Figure 3.3: Effect of pressure on the Gibbs free energy of the reaction in the aqueous phase, T= 298 K (Oelkers et al., 1995)

Direct hydrogenation of CO to produce HCHO is not practical in the gas phase because of the low equilibrium conversion of the reaction. However, as it can be seen in Table 3.2, the equilibrium conversion of the reaction is significantly higher in the aqueous phase, which is due to low  $\Delta G$  of the components in the aqueous media. Therefore, the reaction is thermodynamically feasible in the liquid phase and can achieve sufficiently high equilibrium conversion to make this process viable. The reaction is therefore expected to be kinetically limited, which can be improved by the use of an appropriate catalyst.

### 3.4 Catalyst Characterization

BET surface area of the support and CO-chemisorption results for Pd-Ni and Ru-Ni are shown in Table 3.3. BET surface area is given for the commercial  $\gamma$ -Al<sub>2</sub>O<sub>3</sub> before metal impregnation. It is expected that the surface area reduced to some extent after impregnation (Tanksale et al., 2008).

The nominal and actual metal content of the catalysts are presented in Table 3.3, which shows good agreement. The promoter content was calculated based on the mass balance as the actual content of

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the NiO and Al<sub>2</sub>O<sub>3</sub> was measured by XRF. Based on the amount of CO adsorption on the catalyst surface, metal dispersion was calculated according to the following formula:

$$Dispersion(\%) = \frac{M_{CO} \times AW}{WF} \times 100 \quad (3.7)$$

where  $M_{CO}$  is the amount of adsorbed CO ( $\mu\text{mol.g}^{-1}$ ),  $AW$  is the atomic weight of the metals ( $\text{g} \cdot \mu\text{mol}^{-1}$ ), and  $WF$  is the weight fraction of the metals in the catalysts. Table 3.3 shows that CO uptake and metal dispersion of Ru-Ni/Al<sub>2</sub>O<sub>3</sub> was significantly higher than Pd-Ni/Al<sub>2</sub>O<sub>3</sub>, even though similar metal loading was used. This suggests that the supported Ru-Ni nanoparticles are much smaller than the Pd-Ni nanoparticles. This is confirmed from the TEM results shown in Figure 3.4. The nanoparticles sizes were estimated using ImageJ software (NIH) and shown in the Figure. The mean particle size of Pd-Ni nanoparticles was 3.69 nm compared to the mean particle size of Ru-Ni nanoparticles of 2.14 nm. Lower particle size and hence higher dispersion of the catalyst nanoparticles is favourable for HCHO formation because it would provide higher surface area for CO and H<sub>2</sub> adsorption which is expected to be a necessary step for hydrogenation reaction. The particles size of the powder catalyst is calculated to be in the range of 10-60  $\mu\text{m}$  based on the SEM images of the catalysts. The SEM images are presented in appendix A.4.

**Table 3.3: BET surface area, metal loading, and CO chemisorption data**

Catalyst	Nominal Metal Loading (%w/w)	Actual Metal Loading* (%w/w)	Total CO uptake ( $\mu\text{mol.g}^{-1}$ )	Metal Dispersion (%)	BET Surface area of the Support ( $\text{m}^2.\text{g}^{-1}$ )
Ru-Ni/Al <sub>2</sub> O <sub>3</sub>	Ni-10	Ni-9.3	41.96	2.33	108.51
	Ru-1	Ru-1.4			
Pd-Ni/Al <sub>2</sub> O <sub>3</sub>	Ni-10	Ni-9.5	26.50	1.47	
	Pd-1	Pd-1.3			

\* Al<sub>2</sub>O<sub>3</sub> and NiO content are measured by XRF. Promoter content is calculated based on mass balance

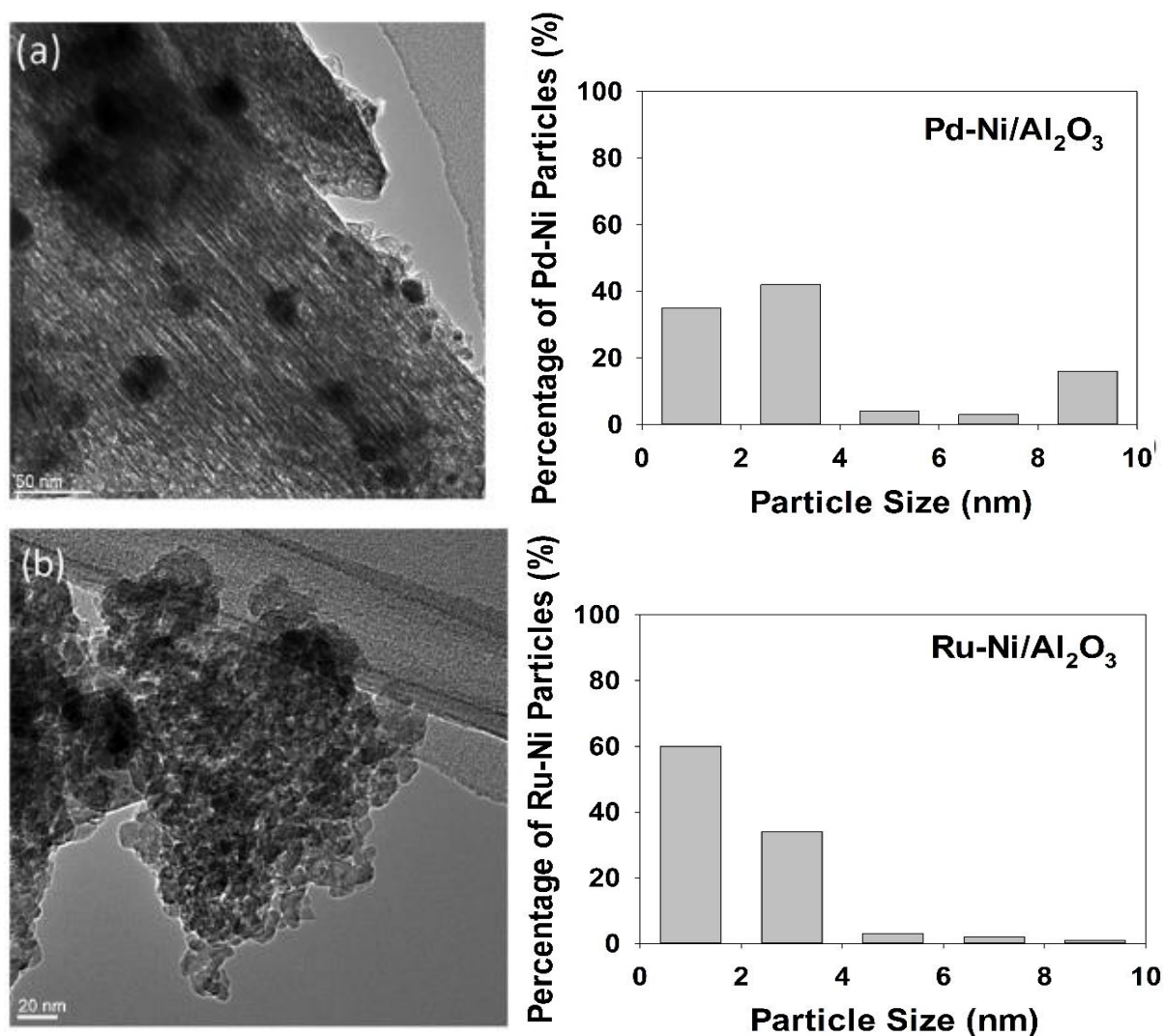


Figure 3.4: TEM image and size distribution of a) Pd-Ni/Al<sub>2</sub>O<sub>3</sub>, Scale bar= 50nm and b) Ru-Ni/Al<sub>2</sub>O<sub>3</sub>, Scale bar= 20nm

The XRD patterns of the calcined and reduced Ru-Ni/Al<sub>2</sub>O<sub>3</sub> and Pd-Ni/Al<sub>2</sub>O<sub>3</sub> catalysts are presented in Figure 3.5. The peaks representing PdO (34.06°) and RuO<sub>2</sub> (28.19° and 34.06°) disappeared after reduction and peaks representing Pd<sup>0</sup> (40.15°) and Ru<sup>0</sup> (43.98°) were observed, which confirmed complete reduction of the catalysts (Abu Bakar et al., 2012; Wojcieszak et al., 2010). Broad NiO and Ni<sup>0</sup> peaks were also observed at 62.88° and 51.83° 2θ angles, which also confirm that the Ni nanoparticles are finely dispersed (Hanson & Norby, 2013).

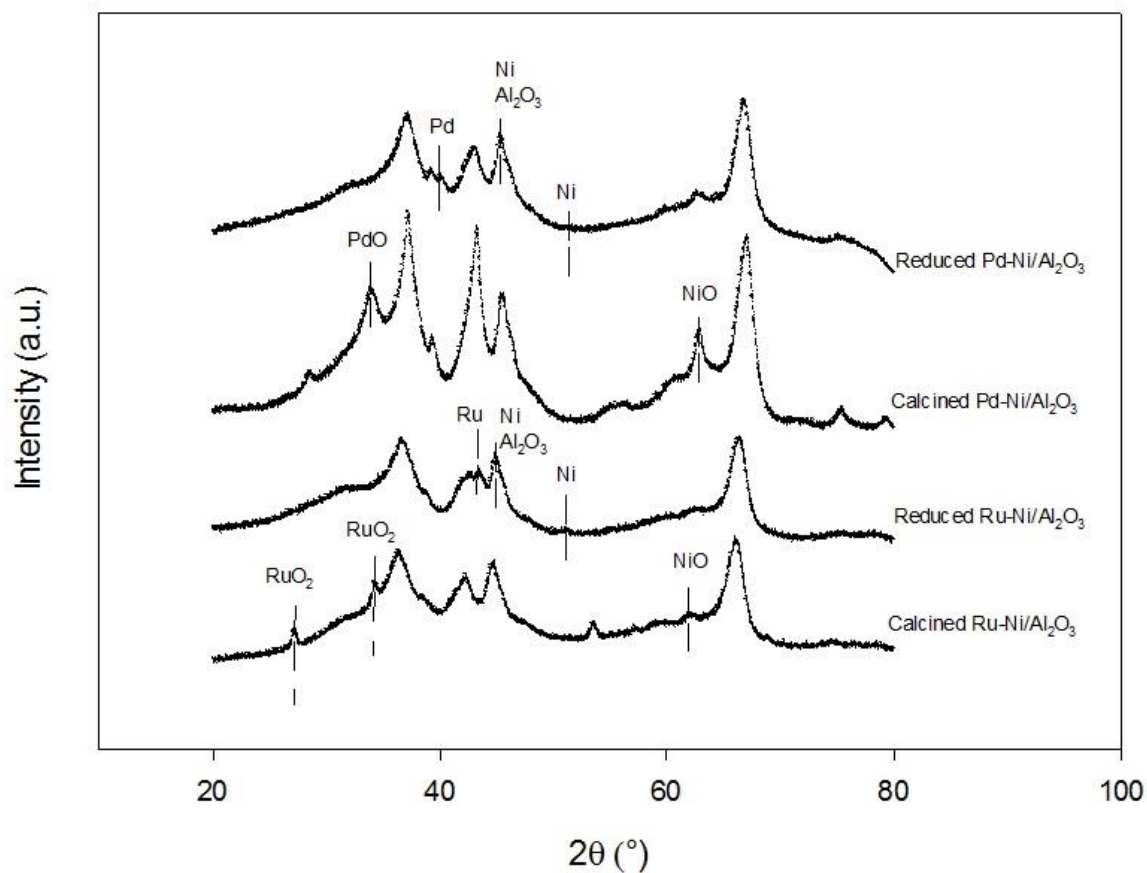


Figure 3.5: XRD patterns of calcined and reduced Pd-Ni/Al<sub>2</sub>O<sub>3</sub> and Ru-Ni/Al<sub>2</sub>O<sub>3</sub>

## 3.5 Comparison of Formaldehyde Production in Fixed Bed and Slurry Reactors

### 3.5.1 Fixed Bed Reactor

Figure 3.6 shows the effect of pressure (Figure 3.6a) and temperature (Figure 3.6b) on the molar yield of HCHO using Pd-Ni/Al<sub>2</sub>O<sub>3</sub> as the catalyst in the gas phase reaction. It is evident that higher pressures are favourable for CO hydrogenation to HCHO, which is theoretically expected based on the Le



Chatelier's Principle and the previous studies (Newton & Dodge, 1933; Walker, 1967). Since two moles of reactants are combining to form one mole of the product, higher pressure would favour the forward reaction. The highest yield of HCHO ( $8.2 \times 10^{-3} \text{ mmol.L}^{-1} \cdot \text{g}_{\text{cat}}^{-1}$ ) in fixed bed reactor was obtained at 117 bar pressure at 293 K, which equates to conversion of only  $1.02 \times 10^{-4} \%$ , which is well below the equilibrium conversion. It was also confirmed from the experiments, as shown in Figure 3.6b, higher temperatures are unfavourable for the reaction in the gas phase because the overall yield of HCHO reduced from  $4.84 \times 10^{-3} \text{ mmol.L}^{-1} \cdot \text{g}_{\text{cat}}^{-1}$  to less than  $2.12 \times 10^{-3} \text{ mmol.L}^{-1} \cdot \text{g}_{\text{cat}}^{-1}$  after 180 min reaction time.  $t = 0$  is the control sample which is the sample taken before the start of the experiment.

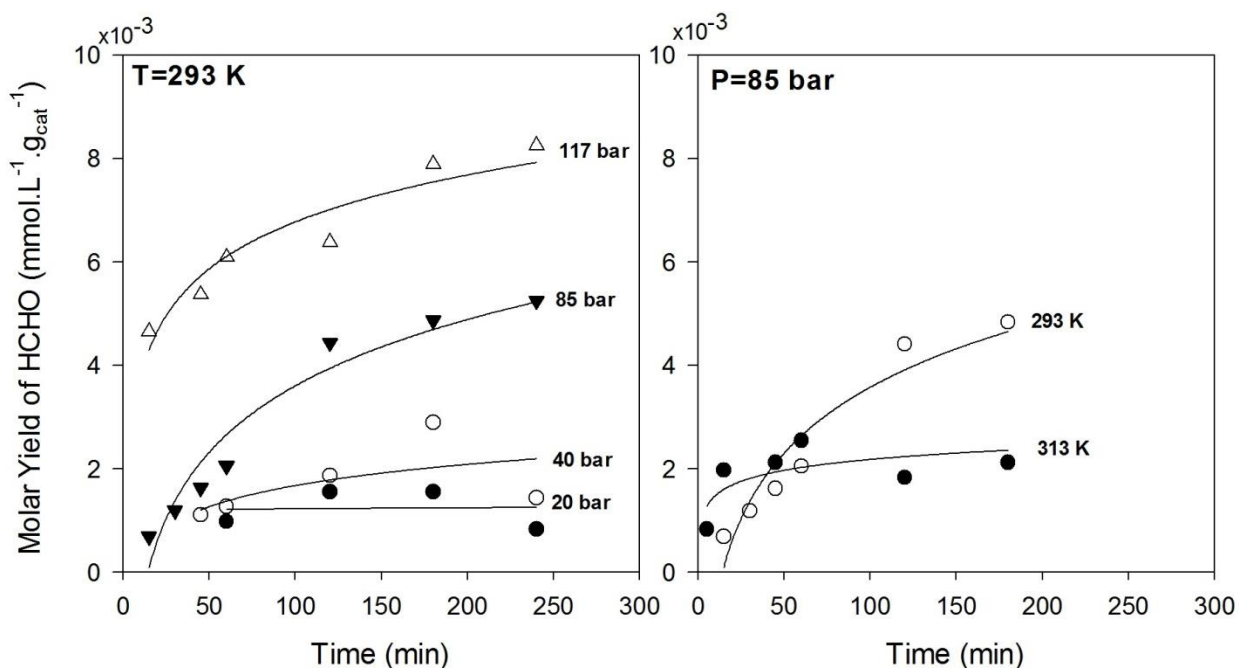


Figure 3.6: Effect of (a) pressure and (b) temperature on the molar yield of HCHO in the fixed bed reactor using Pd-Ni/Al<sub>2</sub>O<sub>3</sub>

### 3.5.2 Slurry Reactor

Figure 3.7 shows the Arrhenius plot based on the rate of reaction calculated after  $t = 2 \text{ h}$ , assuming first order batch reactor kinetics. The rate of formation is calculated based on the yield of HCHO per unit time. The initial concentration of the reactant gases is calculated based on Henry's law. The highest rate of reaction in the first two hours was  $0.0378 \text{ } \mu\text{mol.L}^{-1} \cdot \text{s}^{-1}$  at 403 K. The activation energy was calculated to be  $27.58 \text{ kJ.mol}^{-1}$  for the Ru-Ni/Al<sub>2</sub>O<sub>3</sub> as the catalyst. Figure 3.8 and Figure 3.9 show

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the effect of temperature on the yield of HCHO as a function of time. Although the equilibrium constant decreased with increasing temperature in the aqueous phase, HCHO production rate increased which proves that the process was kinetically limited, unlike the gas phase which was thermodynamically limited. After 48 hours, it was observed that HCHO yield was higher at 353 K and 373 K compared to 403 K.

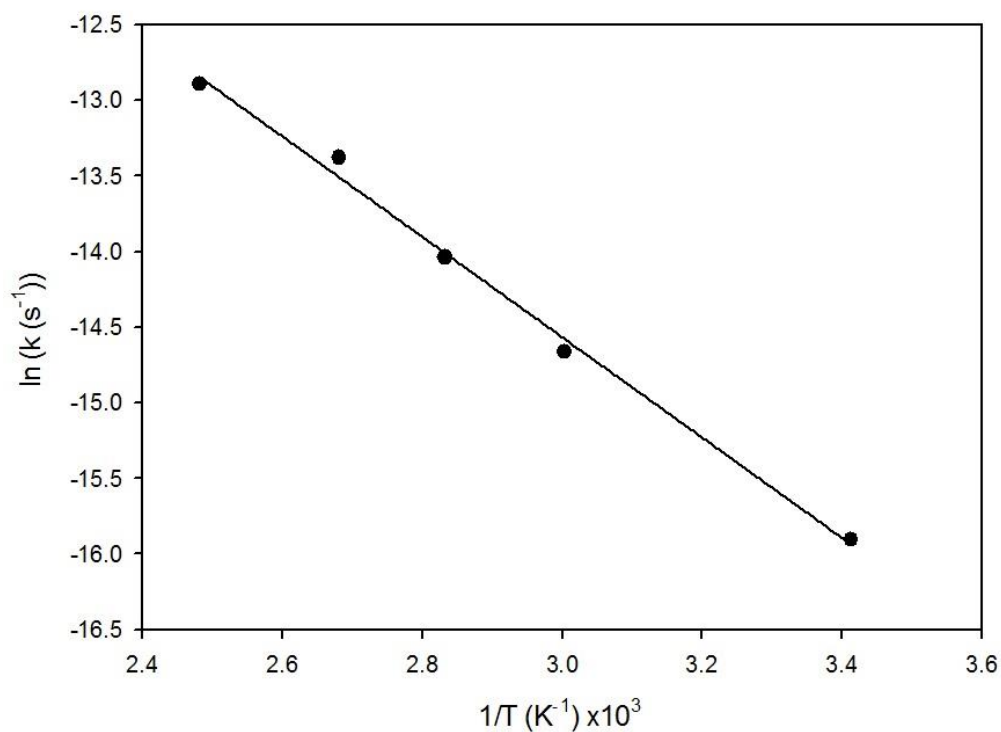


Figure 3.7: Reaction constant based on the first 2 hour of the reaction, P=100 bar, Catalyst: Ru-Ni/Al<sub>2</sub>O<sub>3</sub>

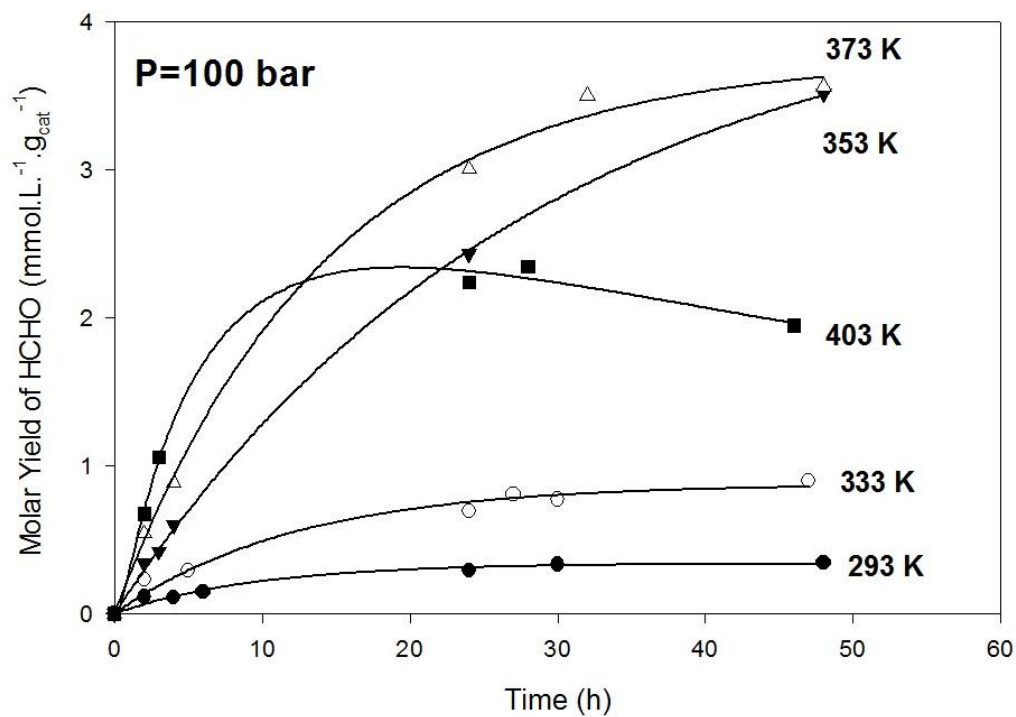


Figure 3.8: Effect of temperature on the molar yield of HCHO in the slurry reactor after 48 hours using Ru-Ni/Al<sub>2</sub>O<sub>3</sub>

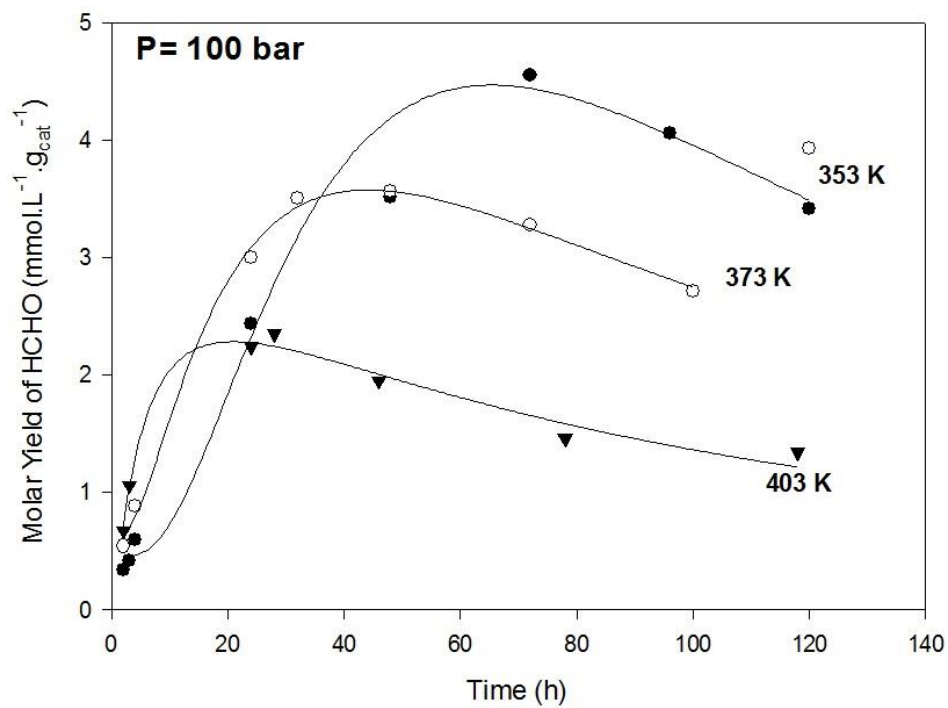


Figure 3.9: Effect of temperature on molar yield of HCHO in the slurry reactor after 120 hours using Ru-Ni/Al<sub>2</sub>O<sub>3</sub>

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As shown in Figure 3.9, the highest level of the HCHO yield was  $4.55 \text{ mmol.L}^{-1}.\text{g}_{\text{cat}}^{-1}$  at 72 hours of operation at 353 K, which is equal to turn over frequency (TOF) of  $0.0602 \text{ h}^{-1}$ . The highest conversion of soluble CO was 19.14% which is significantly higher compared to published works on gas phase CO hydrogenation (Morgan et al., 1932; Newton & Dodge, 1933). This conversion is also higher than the theoretical equilibrium conversion because the equilibrium is shifted when HCHO is dissolved in water. HCHO reacts with water within 70 ms to form methylene glycol ( $\text{CH}_2(\text{OH})_2$ ) (Boyer et al., 2013). In the solution, HCHO and  $\text{CH}_2(\text{OH})_2$  co-exists in dynamic equilibrium with a HCHO/ $\text{CH}_2(\text{OH})_2$  ratio of 1:2499 at STP and pH = 7 (Moret et al., 2014), which means 99.96% of HCHO is converted to  $\text{CH}_2(\text{OH})_2$ . Therefore, HCHO produced in this method is instantly absorbed in water, shifting the reaction equilibrium in the forward direction, as shown in Figure 3.10. Conversion of HCHO into  $\text{CH}_2(\text{OH})_2$  has been well studied in the literature (Boyer et al., 2013; Hasse et al., 1990; Hasse & Maurer, 1991; Maurer, 1986; Winkelman et al., 2000; Winkelman et al., 2002).  $\text{CH}_2(\text{OH})_2$  also polymerizes to form polyoxymethylene ( $(\text{CH}_2\text{O})_n$ ) (Maurer, 1986). However, the polymerization reaction rate is much lower compared with the rate of hydration and dehydration reactions and it can be inhibited by  $\text{CH}_3\text{OH}$  (Boyer et al., 2013). Based on the studies done by Winkelman et al, the HCHO hydration rate ( $k_h$ ) and the equilibrium constant for hydration ( $K_h$ ) are calculated to be as follows (Winkelman et al., 2000; Winkelman et al., 2002):

$$k_h = 2.04 \times 10^5 \times e^{\frac{-2936}{T}} \quad (3.8)$$

$$K_h = e^{\frac{3769}{T} - 5.494} \quad (3.9)$$

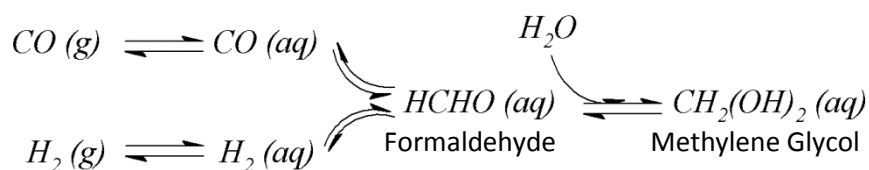


Figure 3.10: Shifting the reaction equilibrium in favour of producing HCHO

It is demonstrated in Figure 3.9 that the HCHO yield peaked in all cases, and that peak shifted towards lower time as the temperature was increased. The final product was checked for the presence of other possible compounds such as ethylene glycol. No other compound was detected in the liquid phase

which indicates loss of HCHO into the gas phase. At high temperatures and low pH values, the  $\text{HCHO}/\text{CH}_2(\text{OH})_2$  equilibrium shifts towards HCHO which may vaporise into the gas phase (Boyer et al., 2013).  $\text{CH}_2(\text{OH})_2$  is known to be very unstable in the gas phase because it tends to dehydrate rapidly to HCHO and water (Kent Iv et al., 2003). Therefore, heating a solution of HCHO and  $\text{CH}_2(\text{OH})_2$  may lead to HCHO emission (Boyer et al., 2013).

Comparison of the HCHO yield in the fixed bed reactor and the slurry reactor at identical operating conditions are presented in Figure 3.11. The yield of HCHO in the slurry reactor at room temperature was more than an order of magnitude higher compared to the fixed bed reactor. The TOF of HCHO production was higher for the Ru-Ni/ $\text{Al}_2\text{O}_3$  than Pd-Ni/ $\text{Al}_2\text{O}_3$  catalyst in the slurry reactor, which suggests that the former catalyst is more active in the aqueous conditions (Boudart, 1995). At room temperature in the aqueous phase, TOF for Ru-Ni/ $\text{Al}_2\text{O}_3$  was  $0.0475 \text{ h}^{-1}$  compared to  $0.0319 \text{ h}^{-1}$  for Pd-Ni/ $\text{Al}_2\text{O}_3$ .

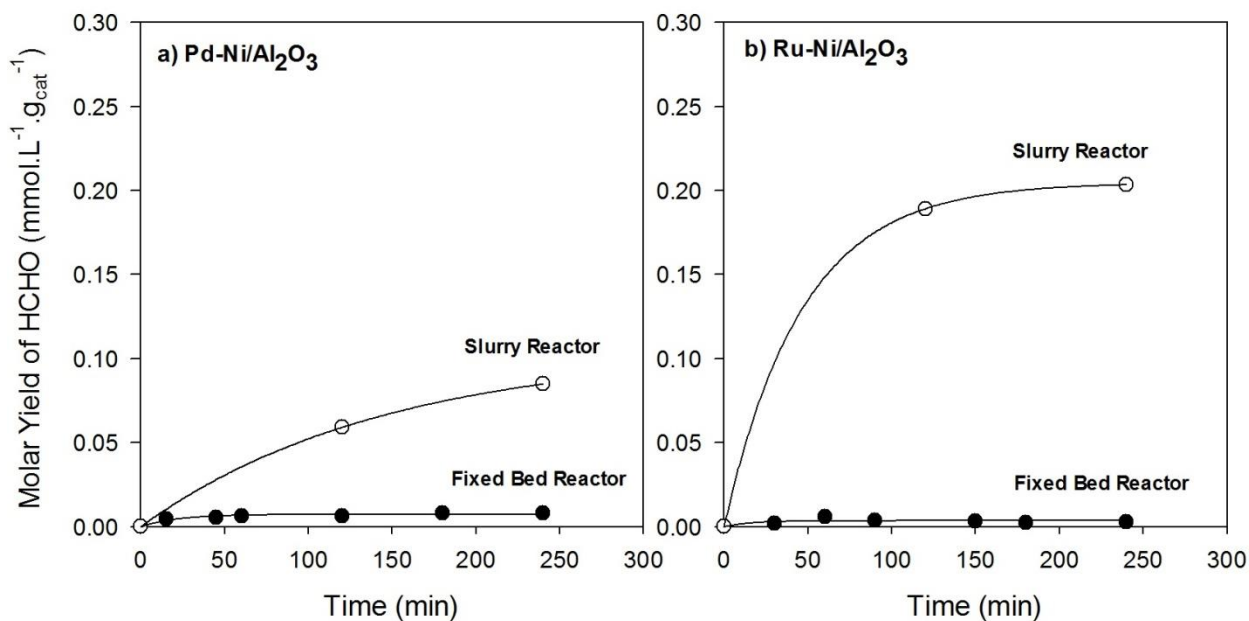


Figure 3.11: Comparison between the results of the fixed bed reactor and the slurry reactor at 293 K and 100 bar for a) Pd-Ni/ $\text{Al}_2\text{O}_3$  and b) Ru-Ni/ $\text{Al}_2\text{O}_3$

### 3.5.3 Effect of the Stirring Speed in the Slurry Reactor

Effect of stirring speed was tested by varying RPM= 0, 400, 800, 1200 at 298 K and 100 bar using Ru-Ni/Al<sub>2</sub>O<sub>3</sub>. Figure 3.12 shows that the yield of HCHO increased from 0.26 mmol.L<sup>-1</sup>.g<sub>cat</sub><sup>-1</sup> to 0.4 mmol.L<sup>-1</sup>.g<sub>cat</sub><sup>-1</sup> as the stirring rate increased from 0 RPM to 800 RPM. But further increasing the RPM did not increase the yield of HCHO. There are many factors which affects catalytic conversion in a slurry reaction. In a non-stirred reactor the catalyst particles may settle at the bottom of the reactor which makes the process diffusion limited. Stirring can decrease the mass transfer limitation by increasing the convective mass transfer coefficient and exposing the catalyst surface to the dissolved gases. Increasing the mass transfer coefficient increases the apparent global rate of reaction and this was observed in Figure 3.12. However, once the rate of mass transfer is sufficiently high, further increasing the stirring rate has no impact on the global rate of reaction because the reaction is kinetically controlled (Lou et al., 2010; S Kislik, 2010). Therefore, it can be concluded that at 800 RPM the test was conducted in a kinetically controlled regime.

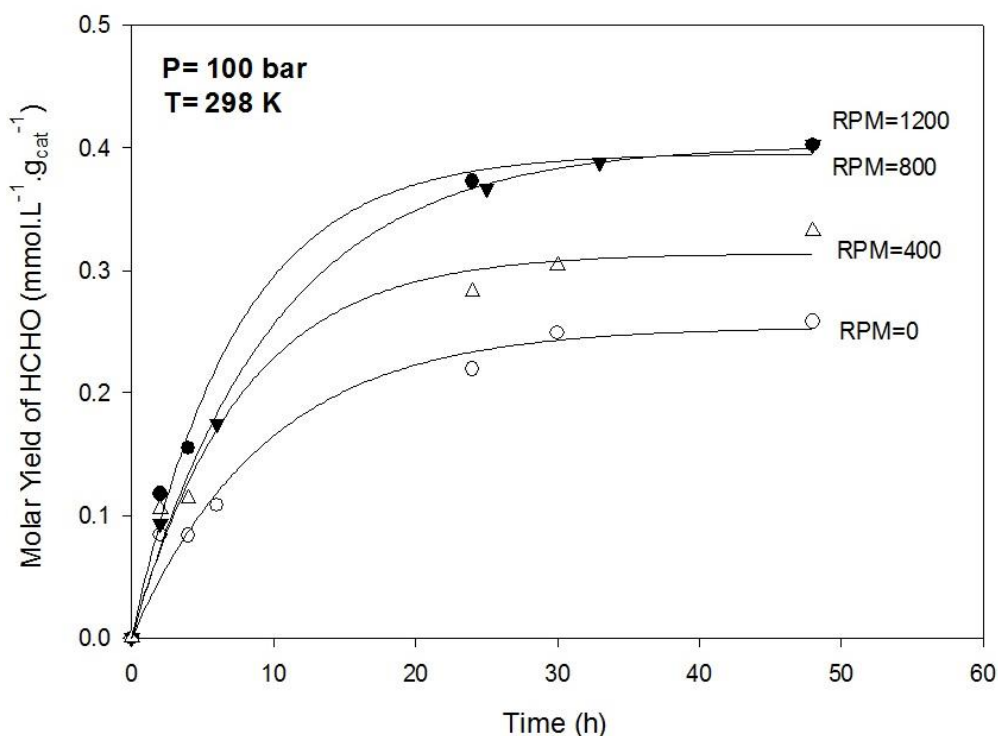


Figure 3.12: Effect of the stirrer rotation speed on the molar yield of HCHO using Ru-Ni/Al<sub>2</sub>O<sub>3</sub>

### 3.5.4 Effect of the Mass of the Catalyst in the Slurry Reactor

In order to check the effect of total accessible active sites, the effect of the mass of the catalyst was studied. Three separate tests were conducted at 353 K and 100 bar. 0.5 g, 1 g, and 2 g of Ru-Ni/Al<sub>2</sub>O<sub>3</sub> were used in each test. The results are presented in Figure 3.13. Based on the following results, it is clear that the HCHO yield increased as the mass of the catalyst increased from 0.5 g to 1 g. However, the yield remained constant as the mass of the catalyst increased further to 2 g. This shows that by using 1 g of the catalyst, sufficient active sites are accessed by the reactants.

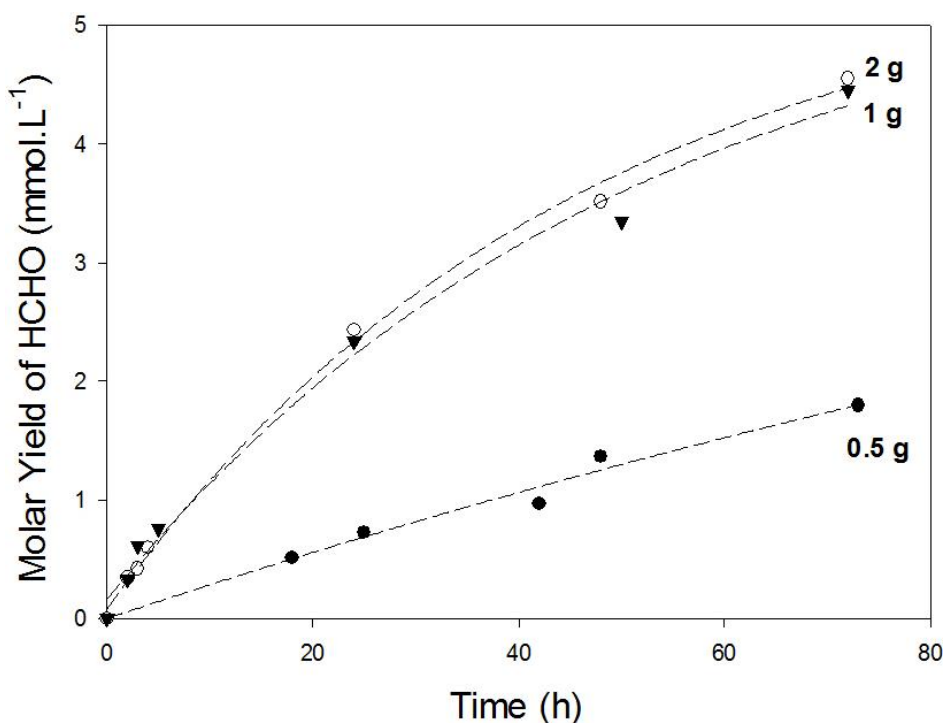


Figure 3.13: Effect of the mass of the catalyst on the HCHO yield, Catalyst: Ru-Ni/Al<sub>2</sub>O<sub>3</sub>, P = 100 bar, T = 353 K

### 3.6 Catalyst Support Stability in the Slurry Reactor

The stability of the catalyst support was investigated by studying the XRD patterns of the fresh and spent Ru-Ni/Al<sub>2</sub>O<sub>3</sub> catalyst. The XRD patterns are shown in Figure 3.14. It is noticed that the XRD pattern of the spent catalyst at 403 K shows sharp peaks at  $2\theta = 14.5^\circ$ ,  $28.2^\circ$ ,  $38.3^\circ$ ,  $49^\circ$ , and  $49.3^\circ$  while these peaks are not observed neither in the fresh catalyst nor in the spent catalyst at 353 K. This

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phenomenon was studied and based on previous studies in the literature, it was discovered that the  $\text{Al}_2\text{O}_3$  support changes phase to form boehmite ( $\text{AlOOH}$ ). Based on previous studies (Ravenelle et al., 2011), the phase change of  $\text{Al}_2\text{O}_3$  to  $\text{AlOOH}$  in hot water results in a decrease in surface area. Also it is shown that the coordination of the Al atoms changes through this structural change. Both tetrahedral and octahedral coordination can be observed in  $\text{Al}_2\text{O}_3$  structure. However, only octahedral coordination of the Al atoms can be seen in boehmite.

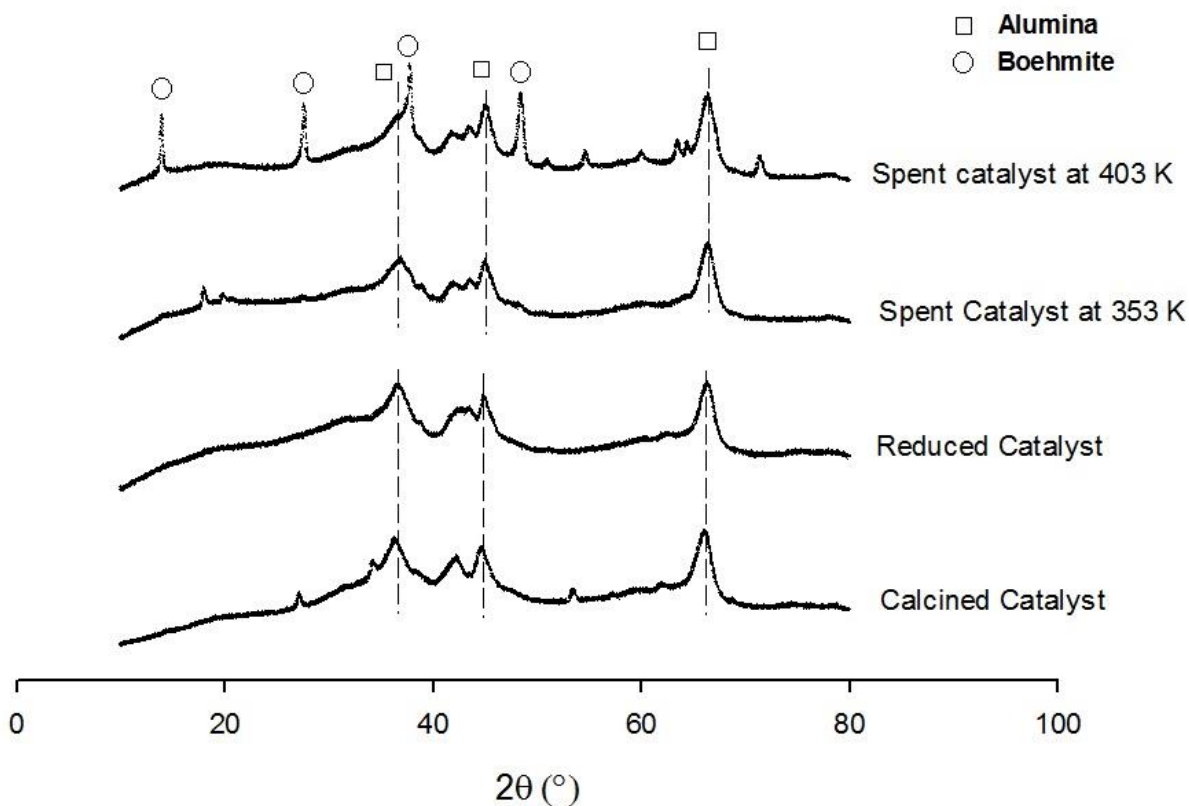


Figure 3.14: XRD patterns of fresh and spent Ru-Ni/ $\text{Al}_2\text{O}_3$  in the slurry phase process

The BET surface area of the fresh calcined Ru-Ni and the spent catalyst at 403 K was measured. It was observed that the surface area of the fresh calcined catalyst is  $95 \text{ m}^2 \cdot \text{g}^{-1}$  where this value is  $83 \text{ m}^2 \cdot \text{g}^{-1}$  for the spent catalyst at 403 K. This measurement shows that the surface area of the catalyst decreases by almost 13% at 403 K. The  $^{27}\text{Al}$  NMR spectra of the fresh and the spent catalyst are shown in Figure 3.15. As it can be seen, both tetrahedral and octahedral coordination can be observed for Al



atoms of the fresh and calcined catalyst. However, the Al atoms of the spent catalyst at 403 K only coordinate octahedrally which shows the formation of boehmite.

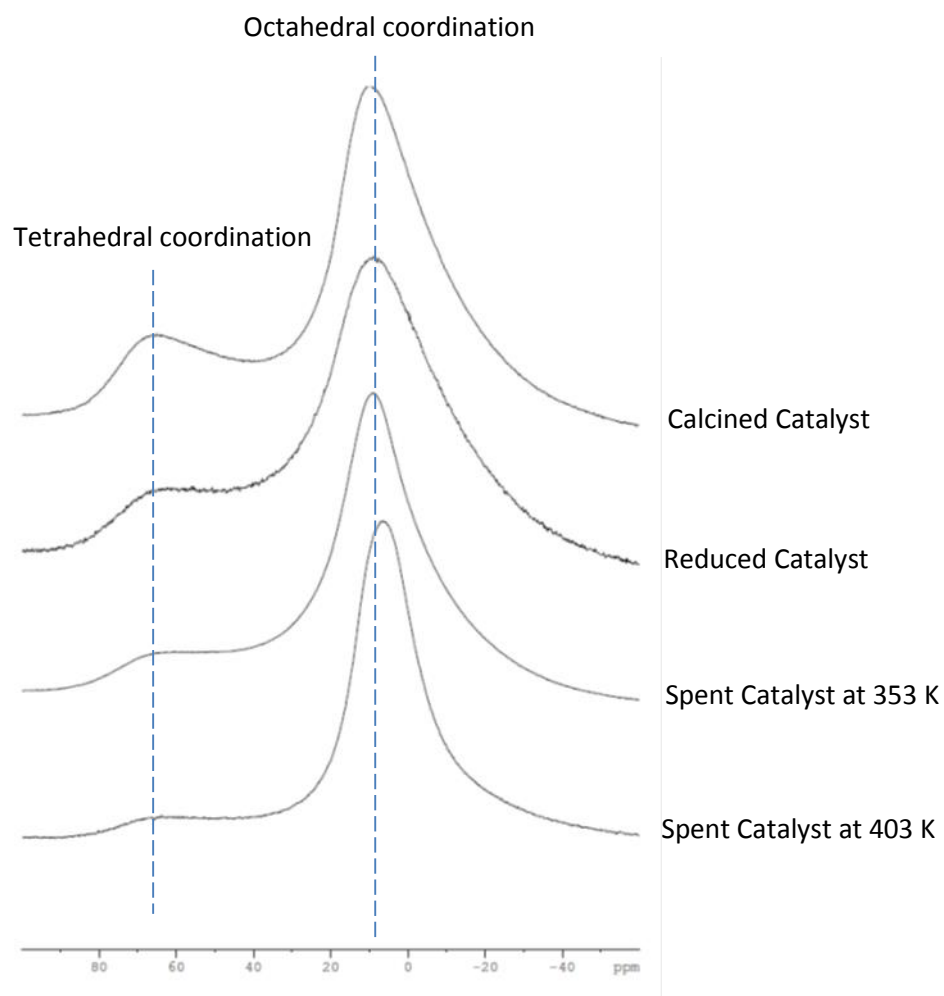


Figure 3.15:  $^{27}\text{Al}$  NMR spectra of the fresh and the spent Ru-Ni/ $\text{Al}_2\text{O}_3$  in the slurry phase process

Hydration of  $\text{Al}_2\text{O}_3$  to  $\text{AlOOH}$  was shown before in the literature at the reforming conditions (Jongorius et al., 2013). However, in this study it is shown that the hydration of  $\text{Al}_2\text{O}_3$  can occur at lower temperatures in longer test runs.

### 3.7 Conclusion

In this chapter, the direct formaldehyde (HCHO) production from synthesis gas in a slurry reactor was investigated. Thermodynamic investigation showed that CO hydrogenation in the gas phase is limited by positive Gibbs free energy ( $\Delta G$ ) at all temperatures above 298 K, whereas in the aqueous phase the reaction is thermodynamically favourable because the  $\Delta G$  is negative below 383 K. This resulted in low yield of HCHO in the fixed bed reactor, and significantly higher yield in the slurry reactor ( $8.25 \times 10^{-3}$  and  $8.48 \times 10^{-2}$  mmol.L<sup>-1</sup>.g<sub>cat</sub><sup>-1</sup>, respectively, at 298 K in 4 h using Pd-Ni/Al<sub>2</sub>O<sub>3</sub>). The HCHO yield reduced with temperature in the fixed bed reactor, where it significantly increased with temperature in the slurry reactor. The highest yield of the HCHO was 4.55 mmol.L<sup>-1</sup>.g<sub>cat</sub><sup>-1</sup> at 353 K after 72 h, which equates to conversion of 19.14% of soluble CO. This conversion is higher than the equilibrium conversion at this temperature because HCHO produced in the aqueous phase is rapidly absorbed by water and hydrated to produce methylene glycol which shifts the equilibrium of CO hydrogenation reaction towards formaldehyde production. It is concluded that Al<sub>2</sub>O<sub>3</sub> is stable at the recommended reaction conditions (T = 353 K, P= 100 bar) in the slurry phase reaction. However, higher temperatures may unfavorably affect the catalyst stability and should be the subject of further investigation. The slurry phase method presented here may be a viable alternative for HCHO production which bypasses the methanol synthesis route. In order to have a better vision of the process and finding the obstacles preventing us from achieving higher production yield, further investigation in the reaction pathway of this process is necessary. Therefore, the reaction mechanism is studied in the next chapter.

### 3.8 References

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## Chapter 4:

# Reaction Mechanism of CO Hydrogenation to Formaldehyde in the Aqueous Phase

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### Abstract

*In this chapter, a reaction mechanism is proposed for hydrogenation of carbon monoxide into formaldehyde. Ideally, direct detection of intermediate species of the reaction in-situ is required for deducing reaction mechanism in heterogeneous catalysis. However, considering the harsh operating conditions of the process (i.e. high pressure and aqueous environment), an in-situ surface study was not feasible. Therefore, the reaction mechanism is investigated in this study based on ex-situ investigation of the potential intermediate products in the samples collected during the reaction in addition to isotope labelling. In separate studies deuterated water and methanol were used with hydrogen gas and normal water and methanol were used with deuterium gas to identify the hydrogen transfer steps in formaldehyde formation using a mass spectrometer. The results suggest that both carbon monoxide and hydrogen gases are dissolved into the reaction mixture and adsorbed on the surface of the catalyst. It is proposed that while carbon monoxide is adsorbed as a molecule, hydrogen adsorbs dissociatively. The adsorbed atoms of hydrogen then attach to the adsorbed carbon monoxide to form formaldehyde molecule. It is shown that in the absence of methanol in the reaction mixture, the adsorbed formaldehyde molecules are further hydrogenated to form methanol. However, the selectivity towards formaldehyde is 100% when the solvent contains at least 5 vol% methanol.*

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## 4.1 Introduction

The conventional process for the production of formaldehyde was presented in detail in chapter 2. The long sequence of production and the high rate of energy loss due to combustion, compression and large purification units in the conventional process leads to relatively low exergy efficiency (~43%) (Bahmanpour et al., 2014; Rosen & Scott, 1988). A novel method of HCHO production was proposed in chapter 3 which could potentially increase the exergy efficiency. Since this is a new method of HCHO production, understanding its reaction mechanism is important so that suitable catalysts and solvents can be identified. Therefore, in this chapter the reaction mechanism of HCHO production through CO hydrogenation is investigated. *In-situ* study of the catalyst surface would be normally necessary to identify the intermediate adsorbed species of reactions. *In-situ* ATR-FTIR has been used in similar processes to investigate the reactions on the catalyst surface (Copeland et al., 2013; Ebbesen et al., 2007). However, considering the harsh conditions in the current work (high pressure and the situation of the catalyst in the aqueous phase), an *in-situ* study of the catalyst surface was not feasible. Therefore, *ex-situ* methods are applied to investigate the reaction mechanism.

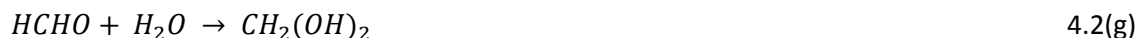
The first step is to consider all possible reaction mechanisms in the process. Based on the source of hydrogen atoms forming HCHO molecule, different reaction mechanisms are applicable. If the source of H atoms is the gaseous H<sub>2</sub> in the reactor, the reaction pathway 4.1 may be proposed:



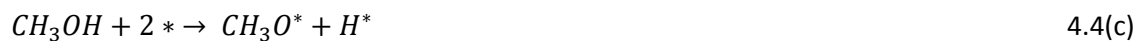
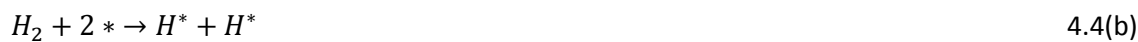
Another source of H atoms can be dissociative adsorption of water molecules on the catalyst surface as H and OH groups. In which case, the HCHO molecule forms on the surface of the catalyst based on the pathway 4.2 with the formation of formic acid (HCOOH) as the intermediate or hydrated form of

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HCHO known as methylene glycol ( $\text{CH}_2(\text{OH})_2$ ) may be formed directly on the surface of the catalyst based on the reaction pathway 4.3



The last possible mechanism is through carboxylation of  $\text{CH}_3\text{OH}$  through which methyl formate ( $\text{CH}_3\text{OCOH}$ ) is formed as an intermediate:







All mentioned reaction mechanisms are summarized in Figure 4.1. Formation of HCHO through dehydrogenation or partial oxidation of CH<sub>3</sub>OH in this process is highly unlikely considering the reducing environment in which the reaction occurs. The four mentioned possible mechanisms can be distinguished based on the source of H atoms forming the HCHO molecules and the potential intermediates which may form during the reaction. These mechanisms can be distinguished by using deuterium labelling technique. Table 4.1 presents the possible products of each experiment based on each possible mechanism.

**Table 4.1: Possible intermediate and final products based on feedstock and possible mechanisms**

Feedstock	Solvent	Products			
		Mechanism	Mechanism	Mechanism	Mechanism
		4.1	4.2	4.3	4.4
<b>H<sub>2</sub> and CO</b>	<b>5% CH<sub>3</sub>OH in H<sub>2</sub>O</b>	HCHO	HCHO	CH <sub>2</sub> (OH) <sub>2</sub>	HCHO
		Detectable	Detectable	Detectable	Detectable
		Intermediate: -	Intermediate: (HCOOH)	Intermediate: -	Intermediate: (CH <sub>3</sub> OCOH)
<b>D<sub>2</sub> and CO</b>	<b>5% CH<sub>3</sub>OH in H<sub>2</sub>O</b>	DCDO	HCHO	CHDO <sub>2</sub> HD	HCDO
		Detectable	Detectable	Detectable	Detectable
		Intermediate: -	Intermediate: (HCOOH)	Intermediate: -	Intermediate: (CH <sub>3</sub> OCOH/ CH <sub>3</sub> OCOD)
<b>H<sub>2</sub> and CO</b>	<b>5% CD<sub>3</sub>OD in D<sub>2</sub>O</b>	HCHO	DCDO	CHDO <sub>2</sub> HD	HCDO
		Detectable	Detectable	Detectable	Detectable
		Intermediate: -	Intermediate: (DCOOD)	Intermediate: -	Intermediate: (CD <sub>3</sub> OCOH/ CD <sub>3</sub> OCOD)

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Mechanism 4.1 is shown to be feasible based on DFT studies on CO hydrogenation to CH<sub>3</sub>OH in the literature (Neurock, 1999; Remediakis et al., 2004). In these studies, it is shown that HCHO can possibly form as an intermediate before conversion to CH<sub>3</sub>OH. Hydrogenation of HCOOH to HCHO (mechanism 4.2) is also shown to be feasible in previous studies (Behrens et al., 2012). It is shown that carbon can receive a proton before separation of the hydroxyl group from the carbonyl side. Mechanisms 4.3 and 4.4 are also considered to cover all possible ways of production of HCHO (or CH<sub>2</sub>(OH)<sub>2</sub>).

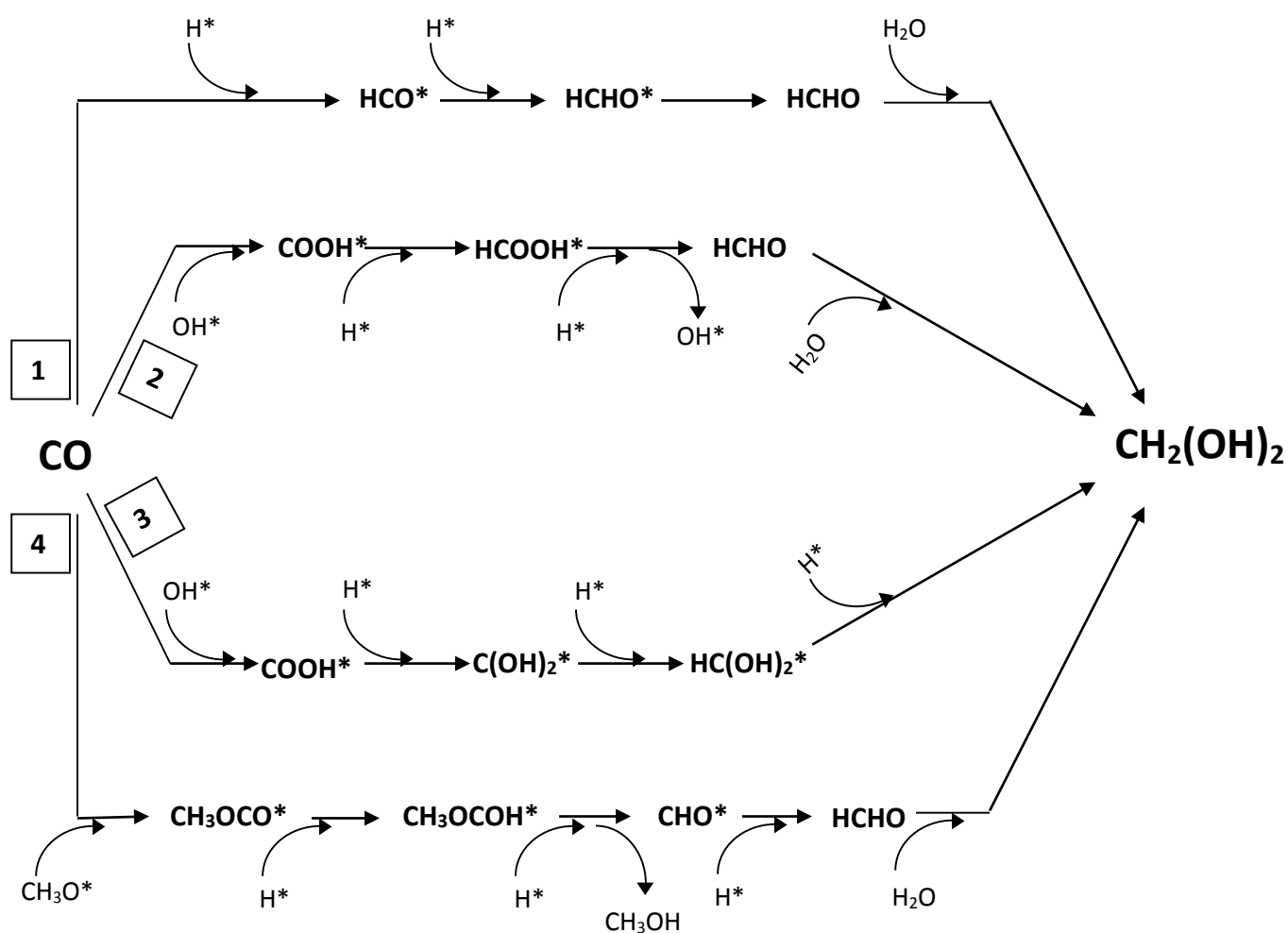


Figure 4.1: Potential reaction mechanisms of CO hydrogenation reaction to HCHO

## 4.2 Experiments and Methods

The process of catalyst generation, reduction, and examination has been described in chapter 3. Detection of HCHO derivatives and other potential products was done by using RHM-Monosaccharide column (Phenomenex) on HPLC (Agilent) using a refractive Index Detector (RID). HCHO solution (Sigma-Aldrich) was used and diluted to  $0.8 \text{ mmol.L}^{-1}$ ,  $1.6 \text{ mmol.L}^{-1}$ ,  $2.5 \text{ mmol.L}^{-1}$ , and  $3.3 \text{ mmol.L}^{-1}$  to generate a calibration curve for detection of methylene glycol ( $\text{CH}_2(\text{OH})_2$ ) which is the hydrated form of HCHO.  $10 \mu\text{l}$  of each solution was injected into the column using milli-Q water as the carrier phase.  $0.5 \text{ ml/min}$  of carrier phase was used and the column was kept at  $338 \text{ K}$ . RID temperature was kept constant at  $318 \text{ K}$ .  $\text{D}_2$  (Sigma-Aldrich),  $\text{D}_2\text{O}$  and  $\text{CD}_3\text{OD}$  (Merck Millipore) were used in the batch reactor for detection of the source of hydrogen transfer in the HCHO production.

The blank solutions (containing  $\text{H}_2\text{O}$  and  $\text{CH}_3\text{OH}$  in the  $\text{H}_2$  replacement experiment and  $\text{D}_2\text{O}$  and  $\text{CD}_3\text{OD}$  in the solvent replacement experiment) as well as the samples taken from the process were heated to  $338 \text{ K}$  for  $10 \text{ min}$ . the headspace of each vial was injected into an injection port situated in the carrier gas line. Carrier gas was connected to vacuum chamber through a mass flow controller. The gas stream was sampled using a capillary tube connected to the vacuum chamber which was supplied with an RGA-300 quadrupole mass spectrometer. The carrier gas flow was set to  $70 \text{ ml/min}$ . The gas samples were checked constantly as the partial pressure of the gas passing through the vacuum chamber was plotted against time. Figure 4.2 shows the schematic diagram of the setup equipped with MS.

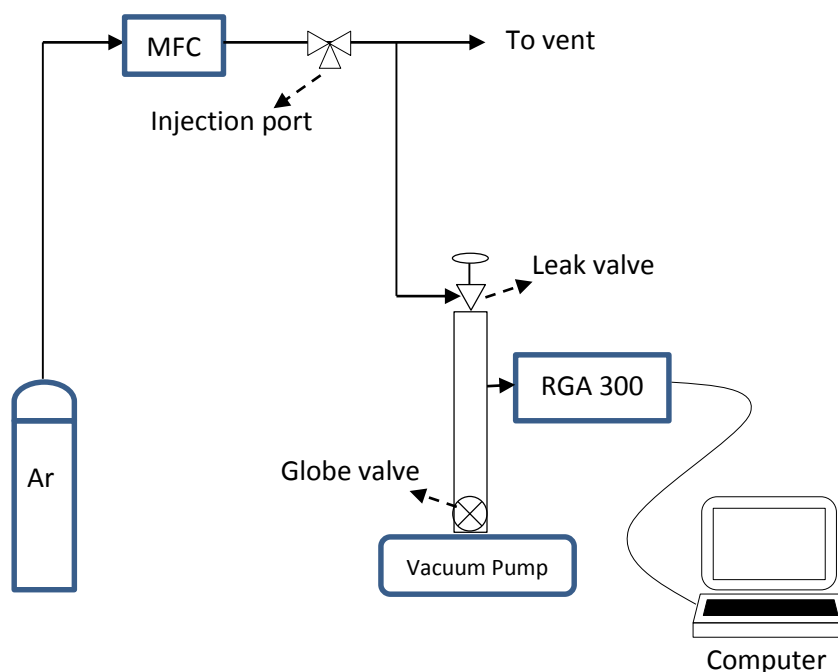


Figure 4.2: schematic diagram of the setup equipped with MS

To detect HCHO or its derivatives on HPLC, different standard formalin solutions were prepared at 0.8 mmol.L<sup>-1</sup>, 1.6 mmol.L<sup>-1</sup>, 2.5 mmol.L<sup>-1</sup>, and 3.3 mmol.L<sup>-1</sup>. The chromatograms are presented in Figure 4.3. Samples were checked on HPLC and the area underneath the peak representing CH<sub>2</sub>(OH)<sub>2</sub> was calibrated based on the concentration. The retention time of the peak was 17.6 min and the calibration curve (Figure 4.4) was linear. This result by itself is interesting since many researchers have tried to detect HCHO or its hydrated form (CH<sub>2</sub>(OH)<sub>2</sub>) by complicated derivatization methods (Arias-Pérez et al., 1992; Dojahn et al., 2001; Soman et al., 2008). To the author's knowledge, detection of underivatized HCHO on HPLC is reported for the first time in this study.

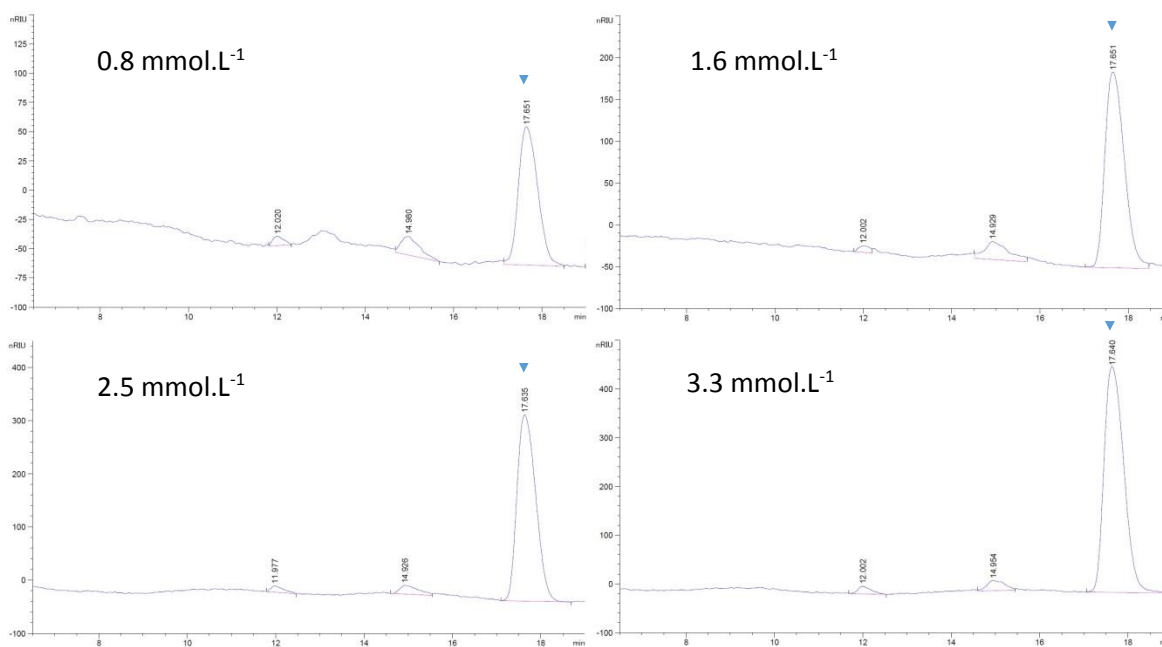


Figure 4.3: Chromatograms of the standard HCHO solutions (▼ = peak at 17.6 min which was calibrated for HCHO)

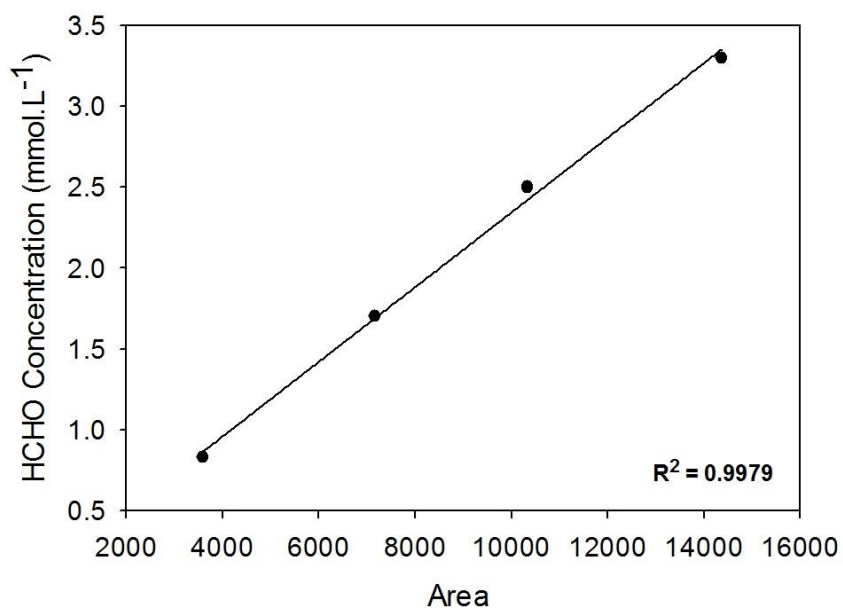


Figure 4.4: Calibration Curve of HCHO on HPLC

### 4.3 Sample Investigation

The samples taken from the process were checked on HPLC for detection of  $\text{CH}_2(\text{OH})_2$  and any potential intermediate products. The chromatograms are shown in Figure 4.5. Based on the achieved chromatograms, it can be seen that no peak is detected before  $\text{CH}_2(\text{OH})_2$  appears at 1.5 h of operation. This indicates that no intermediate products are formed before the formation of HCHO (and consequently  $\text{CH}_2(\text{OH})_2$ ).

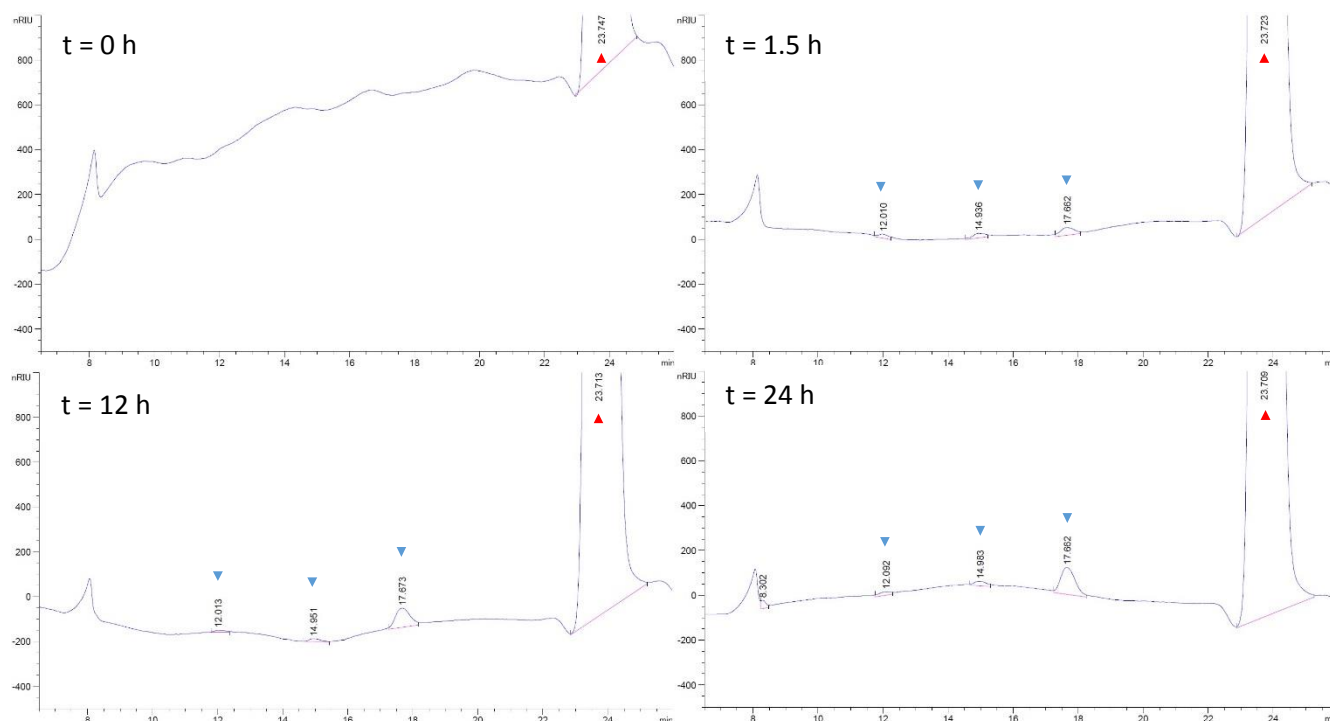


Figure 4.5: HPLC chromatograms of samples taken at various reaction times. Reaction conditions:  $T = 353 \text{ K}$ ,  $P = 100 \text{ bar}$ .

▼ = peak associated with standard formalin solution (only peak at 17.6 min was calibrated for HCHO); ▲ = peak associated with methanol

However, in order to confirm that the potential intermediates based on the discussed reaction mechanisms were not present in the samples, a standard solution containing 400 ppm of formalin solution as well as 100 ppm of  $\text{HCOOH}$  and  $\text{CH}_3\text{COOH}$  was checked on HPLC. The peaks representing  $\text{CH}_3\text{COOH}$  and  $\text{HCOOH}$  were observed at 7.36 min and 11.46 min, respectively, which were not detected in the samples' chromatograms. The chromatogram of the standard solution is shown in appendix A.5.

## 4.4 Determination of the Source of Hydrogen Atoms

### 4.4.1 Labelling H Atoms by Using D<sub>2</sub>

First hypothesis is that hydrogenation of CO happens via reaction mechanism 4.1 in which molecular hydrogen gas is dissolved in water and then adsorbed dissociatively on the catalysts surface, followed by surface reaction of adsorbed H and CO. H<sub>2</sub> gas was replaced with D<sub>2</sub> in the reaction whereas H<sub>2</sub>O and 5 vol% CH<sub>3</sub>OH were used as solvents in the reaction mixture. Temperature was set to 353 K during the process and the reactor was pressurized to 100 bar (CO:H<sub>2</sub> 50:50). Three samples were collected during the process at 24 h, 30 h, and 48 h. The headspace of the blank solution and the reaction product was injected in the injection port at 95 s, 313 s, 588 s, and 887 s. Ten different mass to charge ratio were monitored simultaneously on the RGA software. The mass to charge ratios ( $m/z$ ) of 20 and 34 were selected to monitor D<sub>2</sub>O and CD<sub>3</sub>OD profiles.  $m/z = 29, 30$  and  $31$  were selected to monitor HCHO, DCDO, and CH<sub>3</sub>OH profiles, respectively.

The mass to charge ratios of 29, 30, and 31 are shown in Figure 4.6. As expected, no peak was observed for HCHO and DCDO (or DCHO) when the blank solution headspace was injected. This confirms that the blank solution did not contain any form of formaldehyde (HCHO, DCHO, or DCDO). Sample injections showed peaks at  $m/z = 30$  whereas no peak was observed at  $m/z = 29$ . This confirms the first hypothesis that the reaction takes place according to mechanism 4.1. Absence of peak at  $m/z = 29$  rules out possibility of formation of HCHO and DCHO. Therefore, the hydrogenation of CO proceeded via transfer of adsorbed D atoms only.

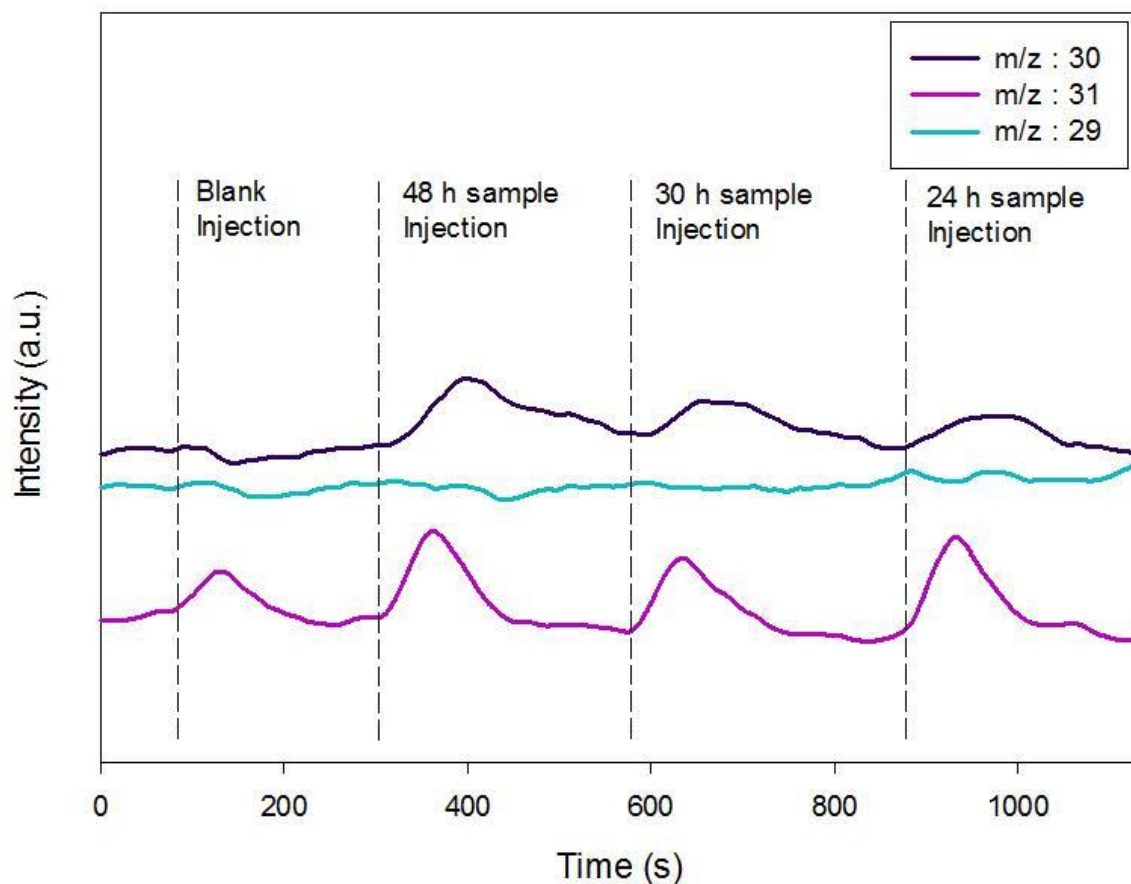


Figure 4.6: Mass to charge ratio profiles for HCHO (29), DCDO (30), and CH<sub>3</sub>OH (31) using 5 vol% CH<sub>3</sub>OH in H<sub>2</sub>O as the reaction mixture

Profiles for  $m/z = 20$  and  $34$ , corresponding to D<sub>2</sub>O and CD<sub>3</sub>OD, were also observed to test for any side products of the reaction (Figure 4.7). No peaks were observed at either  $m/z$  ratios which shows that neither D<sub>2</sub>O nor CD<sub>3</sub>OD were produced in this reaction. This confirms that in the presence of methanol in the reaction mixture HCHO would not hydrogenate further to form methanol since the equilibrium is not favourable.



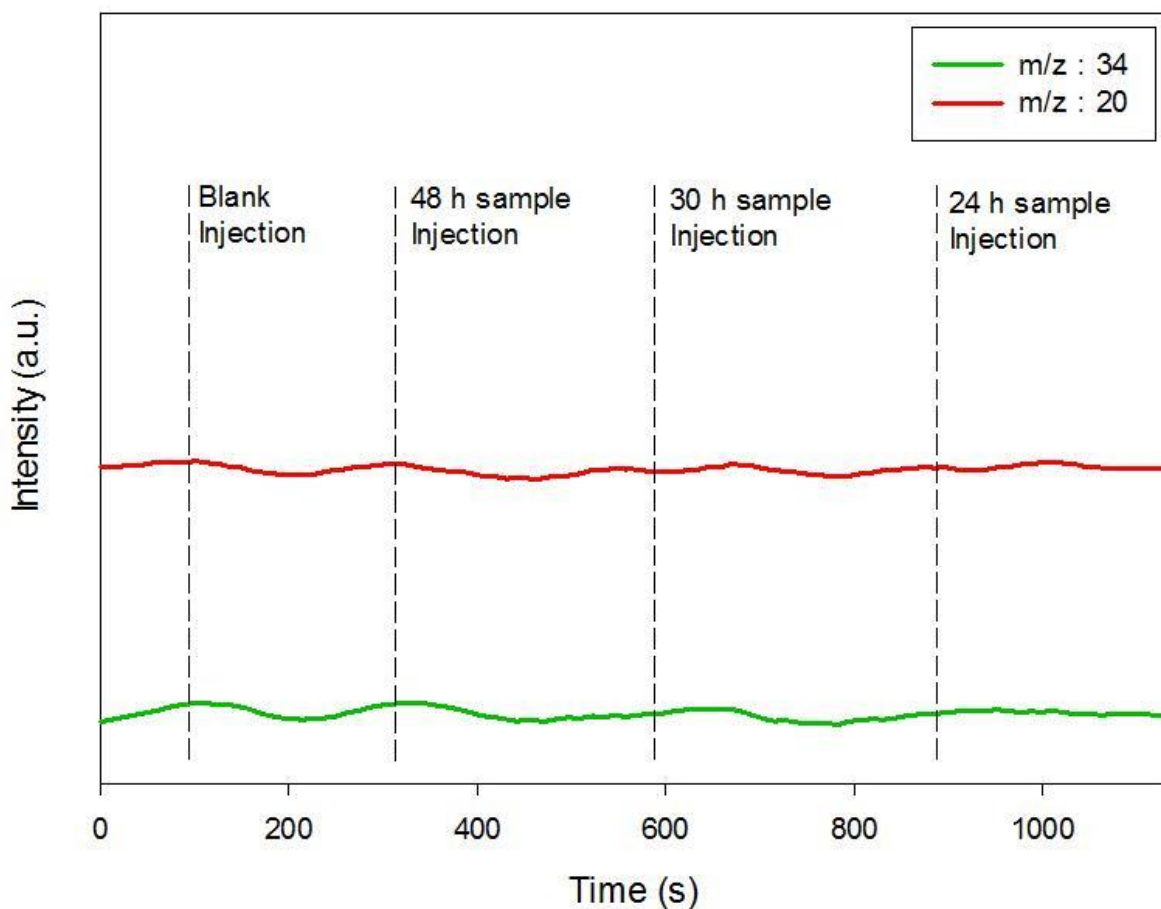


Figure 4.7: Mass to charge ratio profiles for  $D_2O$  (20) and  $CD_3OD$  (34) using 5 vol%  $CH_3OH$  in  $H_2O$  as the reaction mixture

#### 4.4.2 Labelling H Atoms by Using $D_2O$ and $CD_3OD$

To confirm the results of the  $H_2$  replacement experiment, 5 vol%  $CD_3OD$  in  $D_2O$  was used as the solvent in the solvent replacement experiment.  $H_2$  gas was used in the reactor in this test. In this experiment, the headspace of the blank solution and the samples was injected in the injection port at 93 s, 332 s, 650 s, and 936 s, respectively. The m/z ratios monitored during the injection were 20 for  $D_2O$  34 for  $CD_3OD$ , 29 for  $HCHO$  30 for  $DCDO$  (or  $DCHO$ ) and 31 for  $CH_3OH$ . Figure 4.8 shows peaks at 20 and 34 which shows the presence of  $D_2O$  and  $CD_3OD$  in each sample.

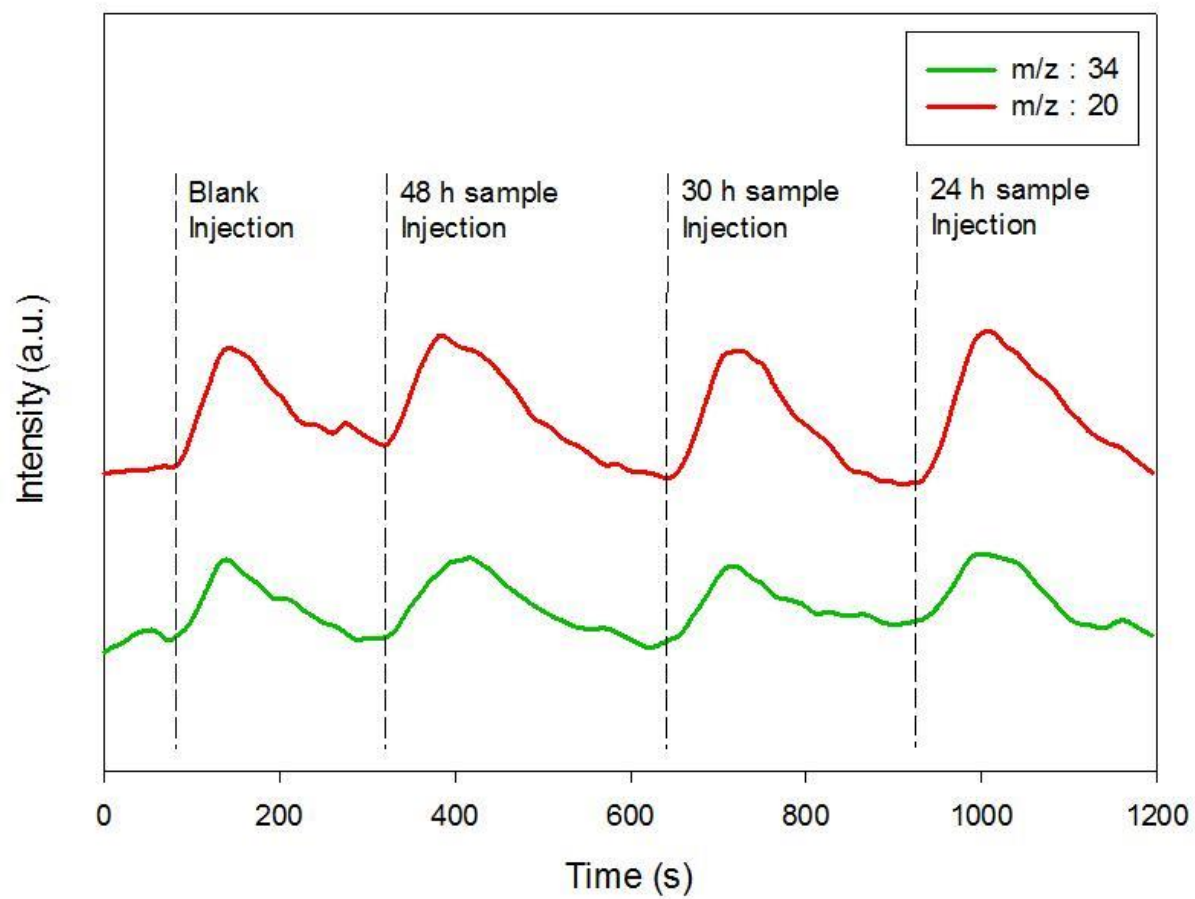


Figure 4.8: Mass to charge ratio profiles for  $D_2O$  (20) and  $CD_3OD$  (34) using 5 vol%  $CD_3OD$  in  $D_2O$  as the reaction mixture

Figure 4.9 shows the  $m/z$  profiles of 29, 30, and 31.

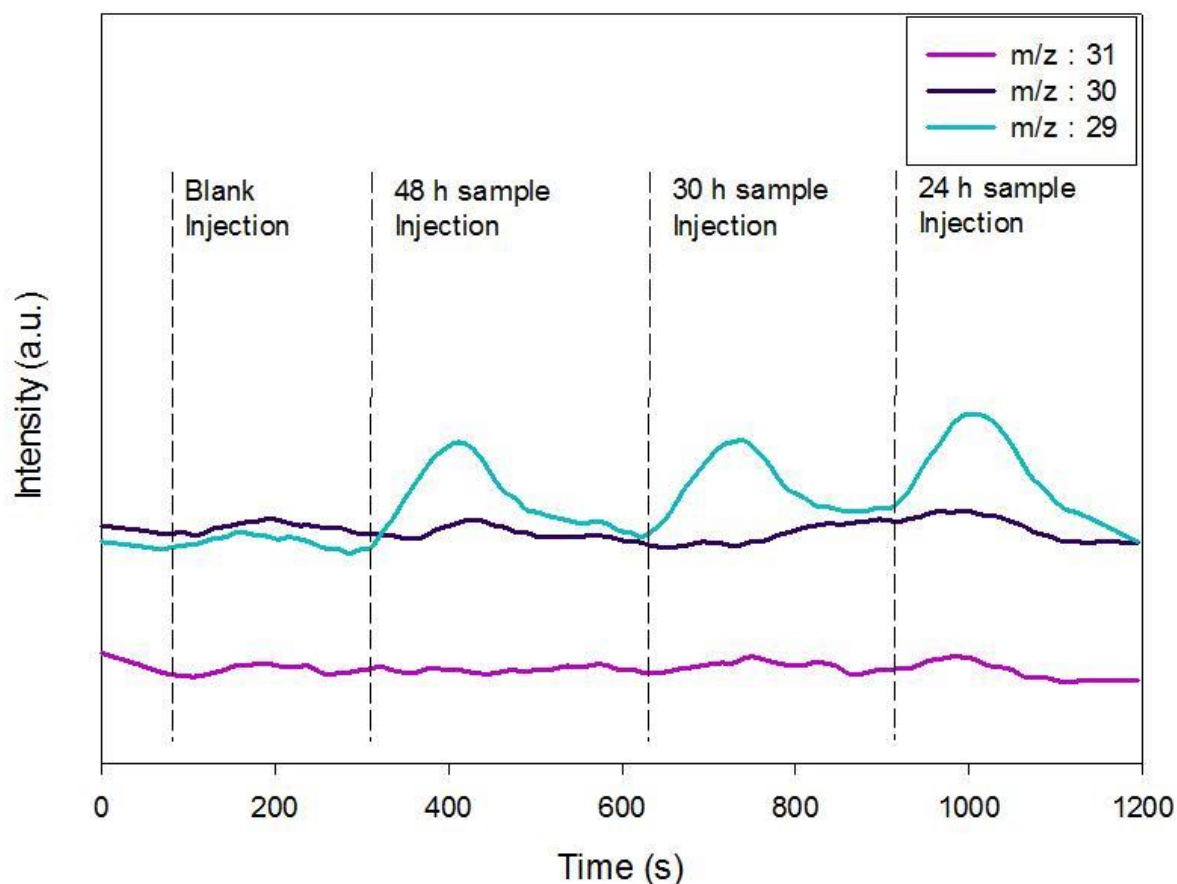


Figure 4.9: Mass to charge ratio profiles for HCHO (29), DCDO (30), and CH<sub>3</sub>OH (31) using 5 vol% CD<sub>3</sub>OD in D<sub>2</sub>O as the reaction mixture

No peak is observed when the blank sample was injected which confirms the absence of HCHO and CH<sub>3</sub>OH. HCHO peak at  $m/z = 29$  is observed for the samples taken after 48 h, 30 h and 24 h reaction time. This is in agreement with the results achieved in the H<sub>2</sub> replacement experiment. DCDO (or DCHO) peak at  $m/z = 30$  was not observed in these samples which indicate that hydrogenation of CO did not take place via deuterium transfer from D<sub>2</sub>O or CD<sub>3</sub>OD. Figure 4.9 also shows that CH<sub>3</sub>OH peak at  $m/z = 31$  is not observed which indicate that HCHO did not go through further hydrogenation to form CH<sub>3</sub>OH. It is worth noting that CH<sub>2</sub>(OH)<sub>2</sub> (or its deuterated form) cannot be detected by this method since CH<sub>2</sub>(OH)<sub>2</sub> does not exist in the gas phase. When the liquid sample is heated, CH<sub>2</sub>(OH)<sub>2</sub> dehydrates to form HCHO in the gas phase.

## Chapter 4

Based on the above mentioned results, it is concluded that no intermediate product forms during the process before the formation of  $\text{CH}_2(\text{OH})_2$  and the source of H atoms participating in the formation of HCHO is the gaseous  $\text{H}_2$  in the reactor. Considering all possible reaction mechanisms for this reaction, the only possible reaction mechanism which satisfies these criteria is mechanism 4.1.

Therefore, it is proposed that CO and  $\text{H}_2$  gas molecules are partially dissolved in the reaction mixture. The dissolved gases are adsorbed on the surface of the catalyst. While CO is adsorbed in its molecular form,  $\text{H}_2$  adsorbs dissociatively forming two H atoms. In the next step, H atoms attach to the C atom forming HCHO which desorbs from the surface of the catalyst and hydrates to form  $\text{CH}_2(\text{OH})_2$ .

### 4.5 Reaction Mechanism in the Absence of $\text{CH}_3\text{OH}$

Although the results show that HCHO does not go through further hydrogenation to form  $\text{CH}_3\text{OH}$ , this fact can be affected in the absence of  $\text{CH}_3\text{OH}$  in the reaction mixture. Therefore, another test run was conducted using pure water as the solvent. The procedure of this test was the same as before. Figure 4.10 shows the comparison of the HCHO yield in the presence and absence of  $\text{CH}_3\text{OH}$  as a solvent.

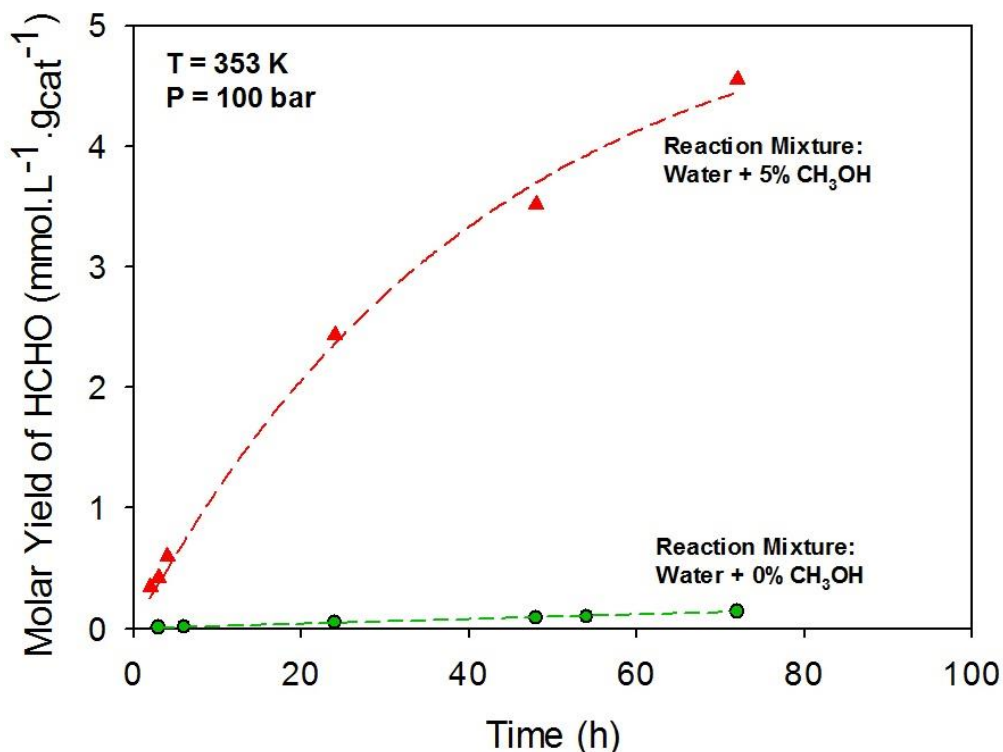


Figure 4.10: The effect of the presence of  $\text{CH}_3\text{OH}$  in the reaction mixture on the molar yield of HCHO

The HCHO yield in the absence of CH<sub>3</sub>OH is significantly lower than the yield in the presence of 5 vol% CH<sub>3</sub>OH. It was hypothesised that HCHO undergoes further hydrogenation into CH<sub>3</sub>OH in the absence of CH<sub>3</sub>OH as a solvent. Liquid samples collected at 6 h, 48 h, and 72 h of the run were injected into HPLC to see if CH<sub>3</sub>OH is produced through this experiment. Figure 4.11 confirms the production of CH<sub>3</sub>OH through this process in the absence of CH<sub>3</sub>OH as a solvent. When 5 vol% CH<sub>3</sub>OH is used as a solvent, the initial concentration of CH<sub>3</sub>OH shifts the equilibrium in the favour of HCHO production. However, in the absence of CH<sub>3</sub>OH, HCHO is hydrogenated further into CH<sub>3</sub>OH.

The source of HCHO hydrogenation was confirmed by D<sub>2</sub> labelling in the gas phase. The samples collected at 5 h, 48 h and 72 h reaction time were tested in the mass spectrometer (Figure 4.12). Injection of the blank sample did not show any peak, whereas, peaks at  $m/z = 34$  and  $30$ , corresponding to CD<sub>3</sub>OD and DCDO is observed for the all the samples. This confirms the hypothesis that in the absence of CH<sub>3</sub>OH as a solvent formaldehyde is hydrogenated further in to methanol.

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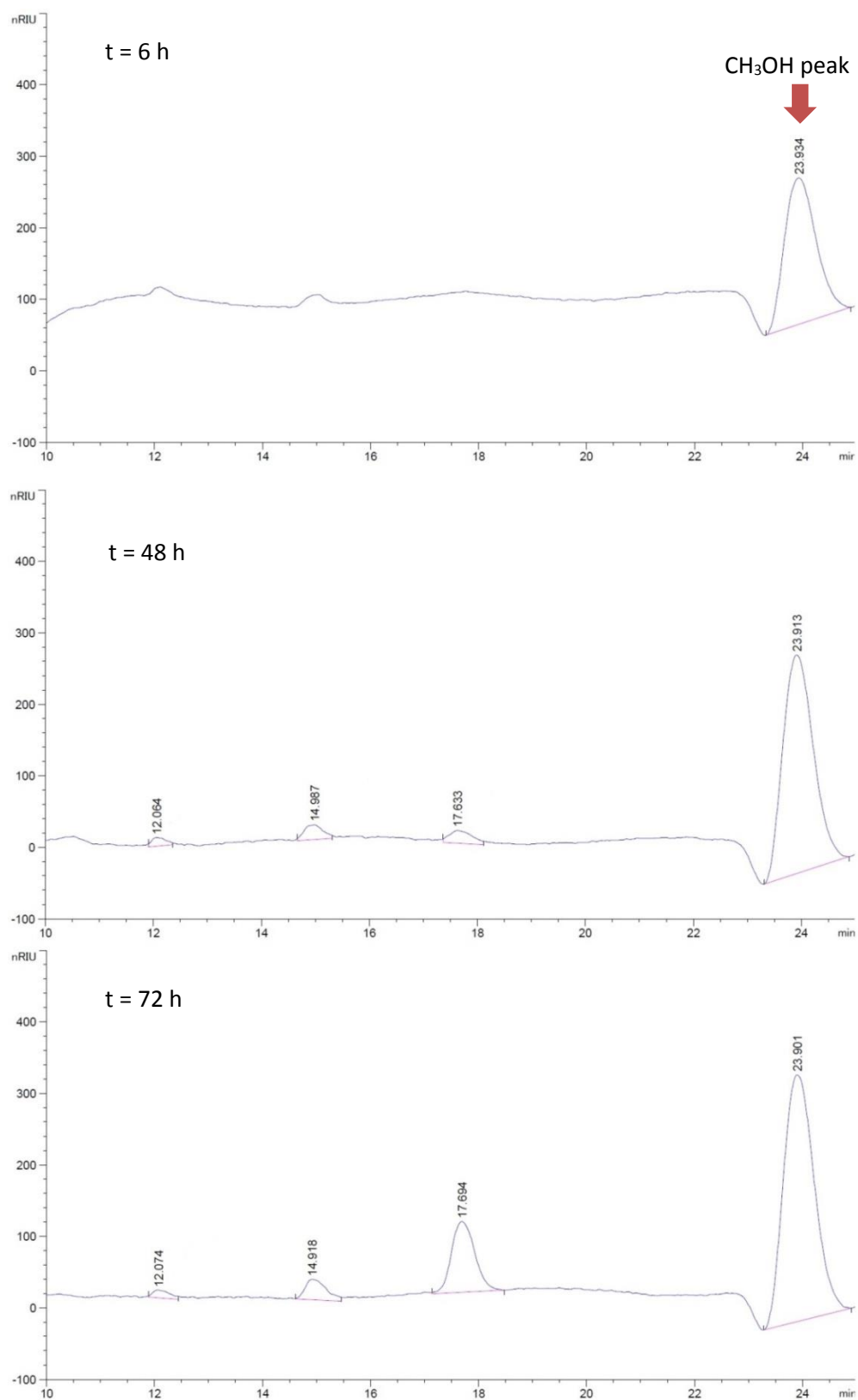


Figure 4.11: The chromatograms of the samples of the run with pure water (no methanol),  $T=353\text{ K}$ ,  $P=100\text{ bar}$

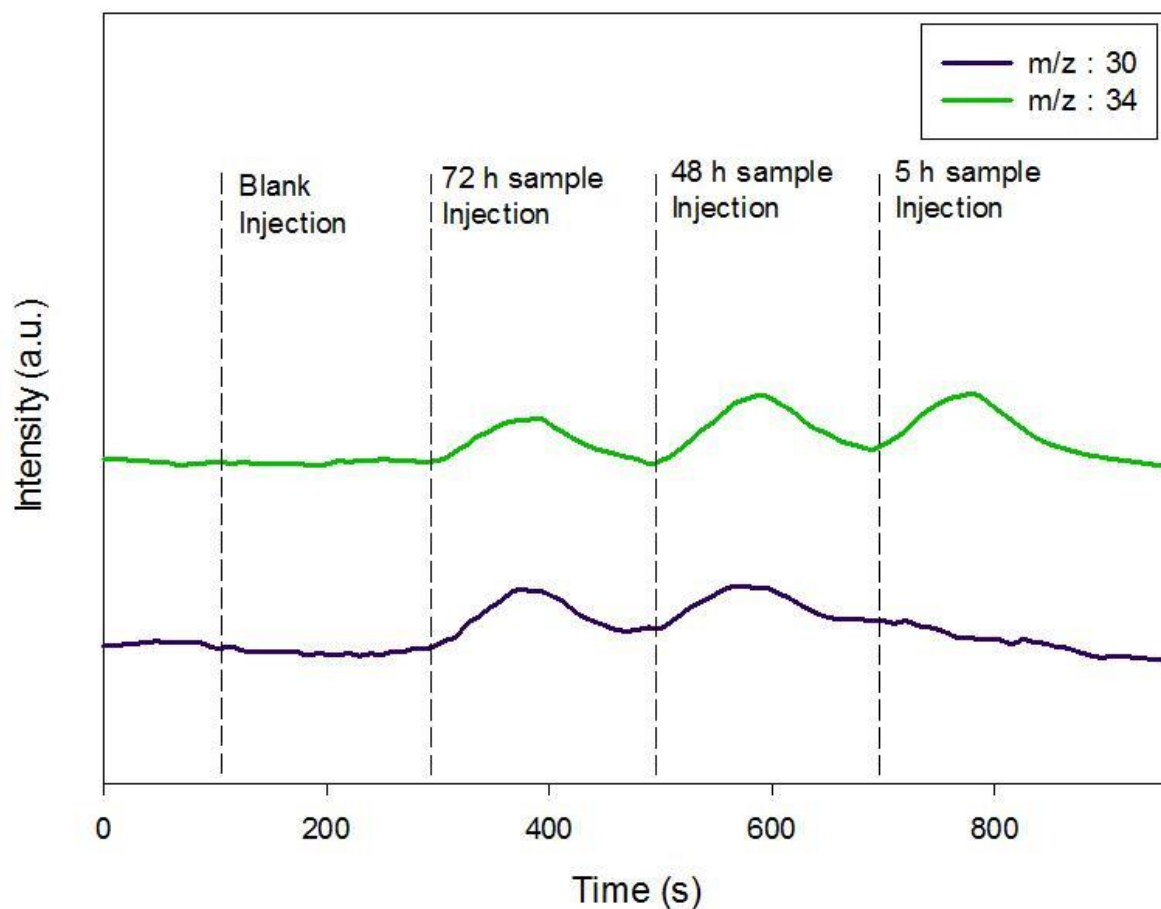
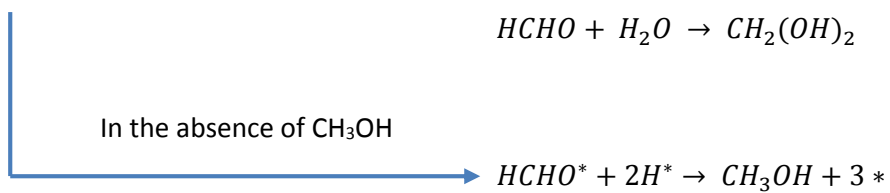
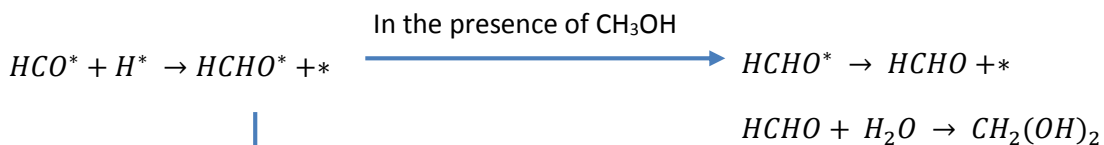
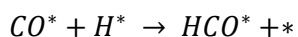
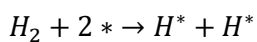
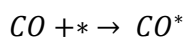


Figure 4.12: Mass to charge ratio profiles for DCDO (30) and CD<sub>3</sub>OD (34) using pure H<sub>2</sub>O as the solvent

Therefore, the proposed reaction mechanism in this process depends on the presence of CH<sub>3</sub>OH as a solvent in the initial reaction mixture. The general reaction mechanism for this process is presented here:



### 4.6 Conclusion

In this chapter, the reaction mechanism of the CO hydrogenation reaction to HCHO was studied. Considering the harsh conditions of the process, an *in-situ* study of the catalyst surface was not feasible. Therefore, by investigating the samples taken during the experiment at regular intervals as well as isotope labelling of H<sub>2</sub> gas, a reaction mechanism was proposed for this reaction. Analysis of the samples showed that no intermediate product was detected before the formation of HCHO in the reaction environment. To be able to identify the source of H atoms participating in the reaction, two separate test runs were conducted (H<sub>2</sub> replacement and solvent replacement). H<sub>2</sub> was replaced by D<sub>2</sub> in the H<sub>2</sub> replacement run while H<sub>2</sub>O and CH<sub>3</sub>OH were used as the reaction mixture. To confirm the results of this experiment, H<sub>2</sub>O and CH<sub>3</sub>OH were replaced by D<sub>2</sub>O and CD<sub>3</sub>OD as the reaction mixture while H<sub>2</sub> was used in the solvent replacement run. By investigation of the samples on mass spectrometer, it was shown that the source of the H atoms which took part in the formation of HCHO was the gaseous H<sub>2</sub> molecules. Based on the mechanistic investigation of the CO hydrogenation reaction, it was clear that this reaction follows mechanism 4.1. It was proposed that CO and H<sub>2</sub> gases are dissolved in the solvent and adsorb on the surface of the catalyst. As opposed to CO, H<sub>2</sub> adsorbs dissociatively on the catalyst surface to form H atoms. H atoms are attached to CO to form HCHO molecule. In the absence of CH<sub>3</sub>OH in the reaction mixture, it was shown that the formed HCHO molecules are further hydrogenated to form CH<sub>3</sub>OH which then desorbs from the catalyst surface. However, in the presence of CH<sub>3</sub>OH in the reaction mixture, it can be seen that the selectivity towards the formation of HCHO is 100% since no CH<sub>3</sub>OH is formed. Desorbed HCHO rapidly reacts with water to form CH<sub>2</sub>(OH)<sub>2</sub>. Considering the proposed reaction mechanism, solution of gases into the solvent is an important step in this process. Low solubility of CO and H<sub>2</sub> in water is one of limitations of the current method. However, solubility of these reactants may be improved by using other solvents. This possibility is studied in the next chapter.



## 4.7 References

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# Chapter 5

## Effect of Solvents on the Formaldehyde Production in the Slurry Reactor

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### Abstract

*Considering the relatively low solubility of the reactant gases in water, the effect of organic solvents on the yield of formaldehyde is studied in this chapter. Higher solubility of the reactants in organic solvents may increase the yield, however, it is equally important to study the effect of solvent interaction with formaldehyde produced over catalyst surface. Formaldehyde rapidly reacts with water and methanol to produce methylene glycol and hemiformal, respectively and in the process shifts the equilibrium in favour of formaldehyde production. Surprisingly, using water-methanol mixtures reduced the formaldehyde yield. It is theorized that since water molecules fill the space in methanol rings, the stability of methylene glycol and hemiformal, which are stabilized by the hydrogen bonding of the surrounding molecules, is reduced. The highest formaldehyde yield achieved through carbon monoxide hydrogenation was  $15.58 \text{ mmol.L}^{-1}.\text{g}_{\text{cat}}^{-1}$  by using methanol as a solvent at 363 K. This yield is four times higher when compared to water as the solvent. Carbon dioxide hydrogenation to formaldehyde is also examined in this study. Although this process is at its preliminary stage at the moment, it shows promising results for future studies. The highest achieved yield through carbon dioxide hydrogenation is  $0.62 \text{ mmol.L}^{-1}.\text{g}_{\text{cat}}^{-1}$ .*

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## 5.1 Introduction

A novel method of formaldehyde (HCHO) production through CO hydrogenation was introduced in Chapter 3. In order to have an insight into this process, the reaction mechanism of CO hydrogenation to HCHO in the aqueous phase was studied in Chapter 4. Based on the results obtained in Chapter 4, it was realized that CO and H<sub>2</sub> gases are dissolved in the solvent and then adsorb on the surface of the catalyst. In the absence of CH<sub>3</sub>OH in the reaction mixture, it was observed that the formed HCHO molecules are hydrogenated further to form CH<sub>3</sub>OH which then desorbs from the catalyst surface. However, in the presence of CH<sub>3</sub>OH in the reaction mixture, was observed that the selectivity towards the formation of HCHO is 100%. Considering the relatively low solubility of CO and H<sub>2</sub> in water at 353 K (0.023 mol.L<sup>-1</sup> and 0.029 mol.L<sup>-1</sup>, respectively, calculated based on Henry's law), it is proposed that the yield of HCHO might improve by using other solvents in which the reactant gases are more soluble. Therefore, the effect of solvents on the reaction is studied in this chapter.

The effect of solvents on hydrogenation reactions have been studied in detail (Akpa et al., 2012; Augustine & Techasauvapak, 1994; Bertero et al., 2011; Jianming et al., 2003; Moret et al., 2014; Mukherjee & Vannice, 2006; Rautanen et al., 2000). The solvent effects have been shown to be significant in a similar study of CO<sub>2</sub> hydrogenation into formic acid (HCOOH) in aqueous phase (Moret et al., 2014). It was shown that the HCOOH yield increased by approximately an order of magnitude by using dimethyl sulfoxide (DMSO)-water solution compared to pure water due to the higher activity of the catalyst in DMSO. Liquid phase hydrogenation reaction of toluene (C<sub>7</sub>H<sub>8</sub>) on Ni/Al<sub>2</sub>O<sub>3</sub> showed higher rate using isooctane ((CH<sub>3</sub>)<sub>3</sub>CCH<sub>2</sub>CH(CH<sub>3</sub>)<sub>2</sub>) and n-heptane (C<sub>7</sub>H<sub>16</sub>) compared with cyclohexane (C<sub>6</sub>H<sub>12</sub>) due to 40% lower hydrogen solubility in C<sub>6</sub>H<sub>12</sub> compared with other solvents (Rautanen et al., 2000). However, solvent effects on hydrogenation reactions are not always dependant on solubility only. Other parameters such as solvent polarity, catalyst-solvent interactions, and solvent molecular structure may affect the reaction. Although non-polar reactants (such as H<sub>2</sub> in hydrogenation reactions) are more soluble in non-polar solvents, it is known that polar solvents may enhance the adsorption of non-polar reactants on the catalyst surface (Augustine & Techasauvapak, 1994; Singh & Vannice, 2001). The interaction between catalyst support and the solvent can affect the reaction. An interesting example of this case is the effect of hydrophobic and hydrophilic catalyst supports on hydrogenation reaction in two phase water mixtures. Hydrogenation of cyclohexene (C<sub>6</sub>H<sub>10</sub>) and cyclohexanone (C<sub>6</sub>H<sub>10</sub>O) in water/cyclohexane mixture showed that using hydrophobic catalyst

support (such as activated carbon) favours the hydrogenation of  $C_6H_{10}$  whereas by using hydrophilic catalyst support (such as  $SiO_2$ )  $C_6H_{10}O$  is favourably hydrogenated (Koopman et al., 1981). Effect of co-solvents on the molecular structure of the solvation medium is found to have an impact on the selectivity towards the products. Molecular dynamic simulation has shown that glucose selective conversion into 5-hydroxymethyl furfural (HMF) is enhanced by using co-solvents (dimethyl sulfoxide (DMSO), tetrahydrofuran (THF) and N,N-dimethyl formamide (DMF)) due to the arrangement of co-solvent molecules around the H atom of glucose hydroxyl groups (Vasudevan & Mushrif, 2015). It is therefore concluded from literature review that solvents may have complex effects on the conversion and selectivity of the hydrogenation reactions.

Many types of solvents including polar protic (water, methanol and ethanol), polar aprotic (DMSO), and non-polar ( $C_7H_8$ ) solvents were used in the present study.

Synthesis gas may contain impurities like  $CO_2$ , therefore, the effect of solvents on the hydrogenation of  $CO_2$  in the slurry reactor was also investigated.  $CO_2$  hydrogenation into HCHO has been reported before in the literature (Lunev et al., 2001; Nakata et al., 2014). However, catalytic formation of HCHO directly from  $CO_2$  hydrogenation in the slurry reactor is reported in this study for the first time.  $CO_2$  can react with  $H_2$  to form HCHO in the following way (Eq. 5.1):



Simultaneous hydrogenation of CO and  $CO_2$  resulted from the biomass gasification process could lead to a new process for the production of HCHO.

## 5.2 Experiments and Methods

$Ru-Ni/Al_2O_3$  was used as the catalyst in all the experiments since it provided better yield in the experiments reported in Chapter 3. The catalyst production procedure as well as the characterization results is reported in Chapter 3. In each experiment, 40 ml of the solution was added to a Parr 100 mL non-stirred batch reactor (Parr-4838). 5% v/v  $CH_3OH$  was added to each solution to stabilise the produced HCHO molecules and to prevent further hydrogenation of HCHO to  $CH_3OH$ . The only exception was when pure  $CH_3OH$  was used as the solvent. 1 g of the catalyst was reduced *ex-situ* and added to the reaction mixture. The reactor was pressurized with equimolar  $CO-H_2$  mixture at a total pressure of 100 bar. The pressurized reactor was then heated to the desired temperature at  $10^\circ C/min$ .

During the reaction, liquid samples were collected using a dip tube and the HCHO concentration was measured using a FluoroQuik fluorometer version 4.3.A using 360nm/490nm as the excitation/emission wavelengths. In this method, a working reagent was prepared by addition of 22  $\mu\text{l}$  of acetoacetanilide ( $\text{C}_{10}\text{H}_{11}\text{NO}_2$ ) and DMSO solution and 33  $\mu\text{l}$  of ammonium acetate ( $\text{C}_2\text{H}_3\text{O}_2\text{NH}_4$ ) solution (Amisience Corporation). 50  $\mu\text{l}$  of the working reagent was mixed with 50  $\mu\text{l}$  of each sample and the mixtures were incubated in the dark for 30 min before measuring the fluorescence intensity. This method of HCHO detection and quantification is used based on previous studies in the literature (Li et al., 2007). Equation 5.1 shows the stoichiometric reaction of  $\text{CO}_2$  hydrogenation into HCHO. Since the stoichiometric ratio of  $\text{CO}_2$  to  $\text{H}_2$  is 1:2, the partial pressure of  $\text{CO}_2$  was set to 33 bar and the partial pressure of  $\text{H}_2$  was set to 66 bar. The reaction was studied at 353 K. 5% v/v  $\text{CH}_3\text{OH}$  in water and  $\text{C}_2\text{H}_5\text{OH}$ , and pure  $\text{CH}_3\text{OH}$  were used as the solvents in this case.

### 5.3 Effect of Solvents on Carbon Monoxide Hydrogenation Process

The effect of solvents on the yield of HCHO production is shown in Figure 5.1. The highest HCHO yield is achieved by using pure  $\text{CH}_3\text{OH}$  followed by 5%v/v  $\text{CH}_3\text{OH}$ /water solution. The rapid reaction of HCHO with water and  $\text{CH}_3\text{OH}$  to form methylene glycol ( $\text{CH}_2(\text{OH})_2$ ) and hemiformal (also known as methoxymethanol) ( $\text{C}_2\text{H}_6\text{O}_2$ ), respectively, shifts the CO hydrogenation reaction equilibrium towards HCHO side which results in higher HCHO yield compared with using other solvents (Bahmanpour et al., 2015).

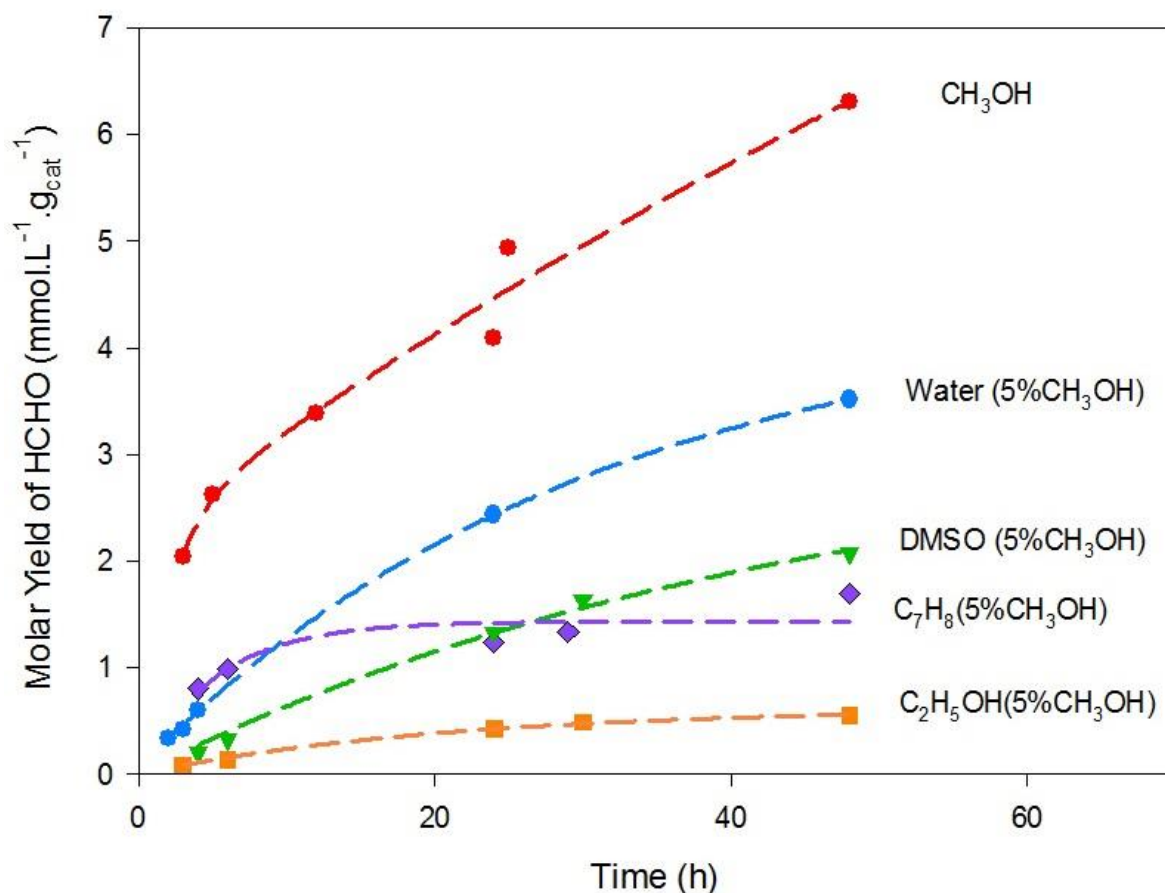


Figure 5.1: The effect of using different solvents on HCHO production yield, T=353 K, P=100 bar, Catalyst: Ru-Ni/Al<sub>2</sub>O<sub>3</sub>

Higher yield of HCHO in pure CH<sub>3</sub>OH solvent may be due to the higher solubility of CO and H<sub>2</sub> compared to water. However, the CO and H<sub>2</sub> solubility is also very high in C<sub>2</sub>H<sub>5</sub>OH, but the yield of HCHO in this solvent was the lowest (Dake & Chaudharl, 1985; Radhakrishnan et al., 1983). Therefore, it can be concluded that the HCHO yield is a function of reactant solubility as well as the HCHO reactivity and stability in the solvent.

Figure 5.2 shows the effect of composition of water- CH<sub>3</sub>OH mixture on the yield of HCHO. It is evident that pure CH<sub>3</sub>OH as a solvent results in the highest HCHO yield and the addition of water to CH<sub>3</sub>OH in any composition decreases the HCHO yield. Moreover, from Figure 5.3 an interesting trend is observed which shows that the yield of HCHO at t = 48 h is higher at mole fractions of CH<sub>3</sub>OH approaching 0 or 1, but the yield is lower for 0 < x < 1.0. The lowest value of x in Figure 5.3 was set to 4.5×10<sup>-4</sup> (equal to 0.1 v/v% CH<sub>3</sub>OH) which was added to favour the equilibrium towards HCHO.



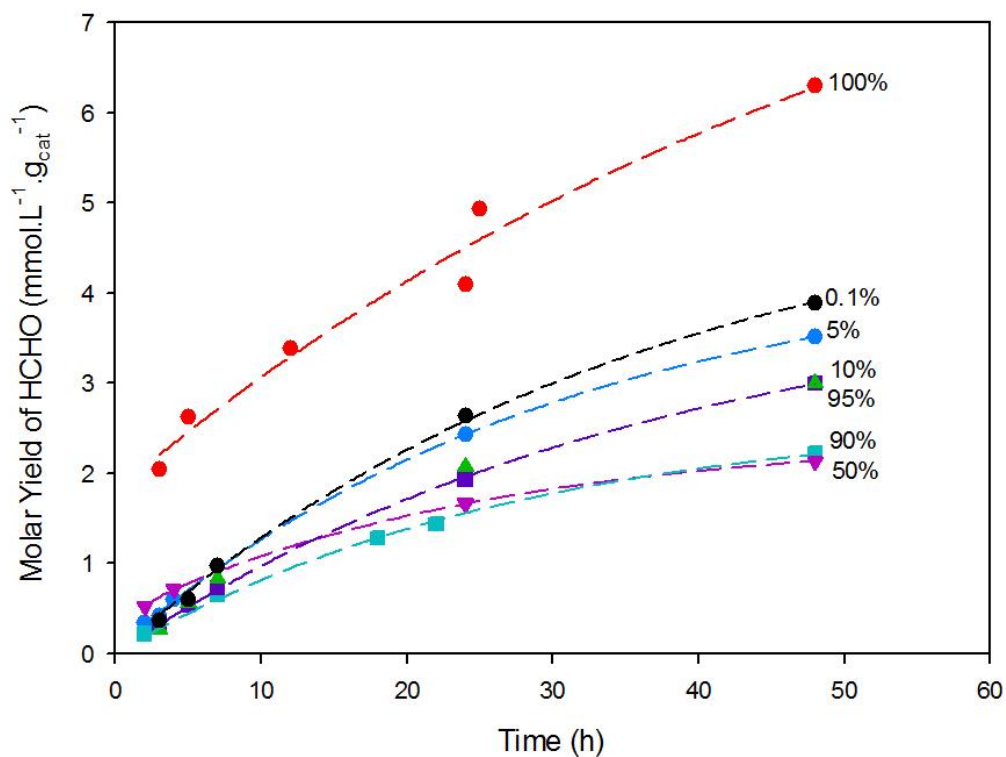


Figure 5.2: Effect of  $\text{CH}_3\text{OH}$  concentration (% v/v) on HCHO yield at  $T = 353 \text{ K}$ ,  $P = 100 \text{ bar}$ , Catalyst:  $\text{Ru-Ni/Al}_2\text{O}_3$

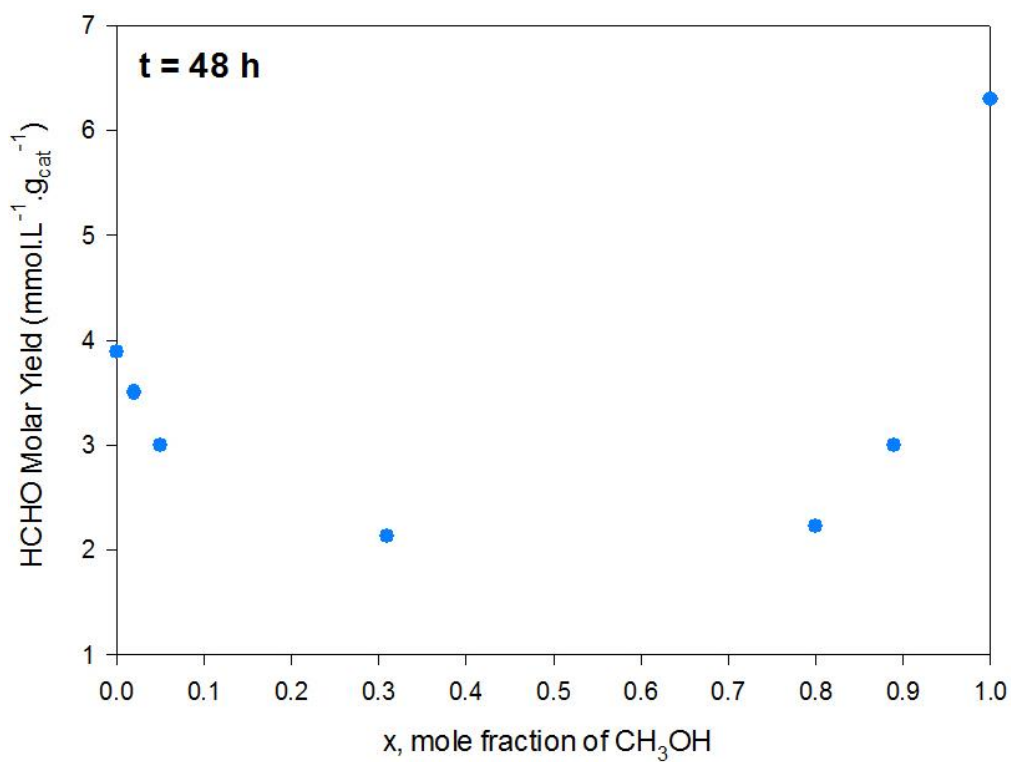


Figure 5.3: Effect of the  $\text{CH}_3\text{OH}$  mole fraction in the solvent mixture on the HCHO yield at  $t = 48 \text{ h}$ , the lowest value of  $x$  tested is  $4.5 \times 10^{-4}$

It is known that  $\text{CH}_2(\text{OH})_2$  and  $\text{C}_2\text{H}_6\text{O}_2$  which are formed through  $\text{HCHO}$  reacting with water and  $\text{CH}_3\text{OH}$ , respectively are unstable molecules. They are stabilized by the hydrogen bonding with the surrounding water or methanol molecules (Boyer et al., 2013). The trend that is observed in Figure 5.3 may be explained by understanding this stabilisation effect of water and methanol via hydrogen bonding. In this thesis it is theorised that the  $\text{CH}_2(\text{OH})_2$  molecules are stabilised by hydrogen bonding with water molecules which form cage structure (Izgorodina et al., 2010); and the  $\text{C}_2\text{H}_6\text{O}_2$  molecules are stabilised by hydrogen bonding with methanol molecules which form cyclic oligomer structure (Guo et al., 2003). Figure 5.4 shows the theorised process of formation, desorption, and stabilisation of the  $\text{HCHO}$  molecules in  $\text{CH}_3\text{OH}$  cyclic oligomers when pure  $\text{CH}_3\text{OH}$  was used as a solvent. It is known from literature that water molecules form three dimensional network that dynamically change to form cages and chains (Izgorodina et al., 2010) and  $\text{CH}_3\text{OH}$  molecules form cyclic oligomers of 6 and/or 8 molecules which provide strong hydrogen bonding ( $191 \text{ kJ.mol}^{-1}$ ) (Guo et al., 2003; Ludwig, 2005). It is also known that the interspace of the  $\text{CH}_3\text{OH}$  cyclic oligomers can be filled with water molecules in water- $\text{CH}_3\text{OH}$  mixtures (Guo et al., 2003). According to Gao et al., 2003, although  $\text{CH}_3\text{OH}$  is fully soluble in water in all proportions, the mixture is not homogeneously mixed at the molecular level. A chain of 8  $\text{CH}_3\text{OH}$  molecules bonds with 1-2 water molecules to form a ring and a chain 6  $\text{CH}_3\text{OH}$  molecules bonds with 2-4 water molecules to form a ring. Their findings also suggest that “free swimming” water molecules (without hydrogen bonding) do not exist in any appreciable amount. Therefore, from the trend observed in Figure 5.3 it can be theorised that the tendency of the  $\text{CH}_3\text{OH}$  cyclic oligomers to accommodate the  $\text{C}_2\text{H}_6\text{O}_2$  molecules may decrease in the presence of water molecules, since water molecules may preferentially fill the interspace of the  $\text{CH}_3\text{OH}$  cyclic oligomers, resulting in reduced stability of  $\text{C}_2\text{H}_6\text{O}_2$  in water- $\text{CH}_3\text{OH}$  solutions. The highest mole fraction of  $\text{CH}_3\text{OH}$  at which water would fill the interspace of the  $\text{CH}_3\text{OH}$  cyclic oligomers is 0.89 (configuration of 8  $\text{CH}_3\text{OH}$  molecules with 1 water molecule). Above this mole fraction ‘water free’ cyclic oligomers of  $\text{CH}_3\text{OH}$  may be found. Figure 5.3 shows that the yield of  $\text{HCHO}$  increased slightly when the mole fraction of  $\text{CH}_3\text{OH}$  increased from 0.8 to 0.9, however, the yield increased significantly when the mole fraction of  $\text{CH}_3\text{OH}$  was increased to 1.0. This theory may therefore be able to explain the effect of water on the  $\text{HCHO}$  yield in concentrated  $\text{CH}_3\text{OH}$  solutions. Figure 5.3 closely resembles the excess molar volume of water- $\text{CH}_3\text{OH}$  mixture which shows non-ideal mixing of water and  $\text{CH}_3\text{OH}$  at the molecular level (Xiao et al., 1997). However, to provide strong evidence to support this theory, further experimental and molecular modelling work similar to the X-ray absorption and X-ray emission

spectroscopy, and density functional theory work of Gao et al. 2003 is necessary. This is beyond the scope of the current thesis.

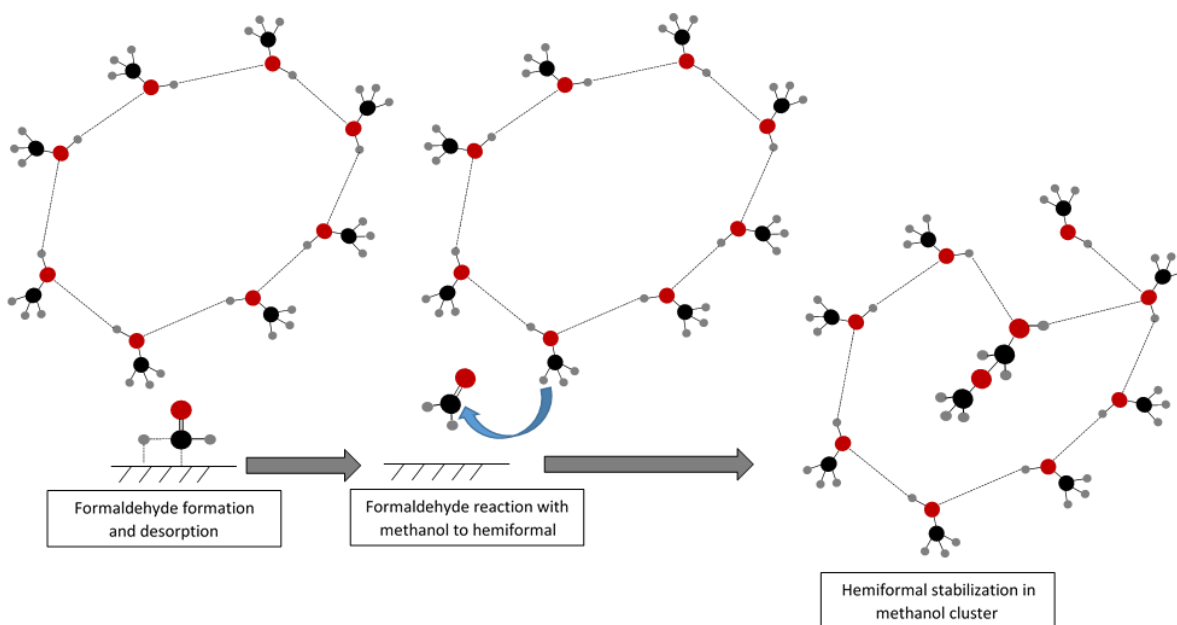


Figure 5.4: A simplified schematic of HCHO formation, desorption and stabilization in  $\text{CH}_3\text{OH}$ . Adapted from (Guo et al., 2003)

The molecular structure of  $\text{C}_2\text{H}_5\text{OH}$  molecules may not provide the required molecular space and hydrogen bonding for the produced HCHO molecules since they form winding chain like structure instead of cyclic oligomers or cages (Benmore & Loh, 2000) which may explain the poor yield of HCHO in this solvent.

## 5.4 Effect of temperature on the HCHO yield in methanol solvent

Figure 5.5 shows the result of HCHO yield at 353 K, 363 K, 373 K, and 403 K using pure  $\text{CH}_3\text{OH}$ . From the initial slope of the curves it can be seen that the rate of reaction increased with temperature, which is in agreement with the previous results reported in Chapter 3. Figure 5.6 shows the Arrhenius plot based on the rate of reaction calculated after  $t = 5$  h, assuming first order batch reactor kinetics. The rate of formation is calculated based on the yield of HCHO per unit time. The apparent activation energy,  $E_a$ , is calculated to be 8.17 kJ/mol based on this graph. The highest yield is measured to be  $15.58 \text{ mmol} \cdot \text{L}^{-1} \cdot \text{g}_{\text{cat}}^{-1}$  which is achieved at 363 K and  $t = 100$  h. However, as it can be seen in Figure 5.5, the yield approaches a maximum value at each temperature at long residence times. Moreover, the

maximum yield reduces as the temperature increases, even though the initial rate is higher at higher temperatures.

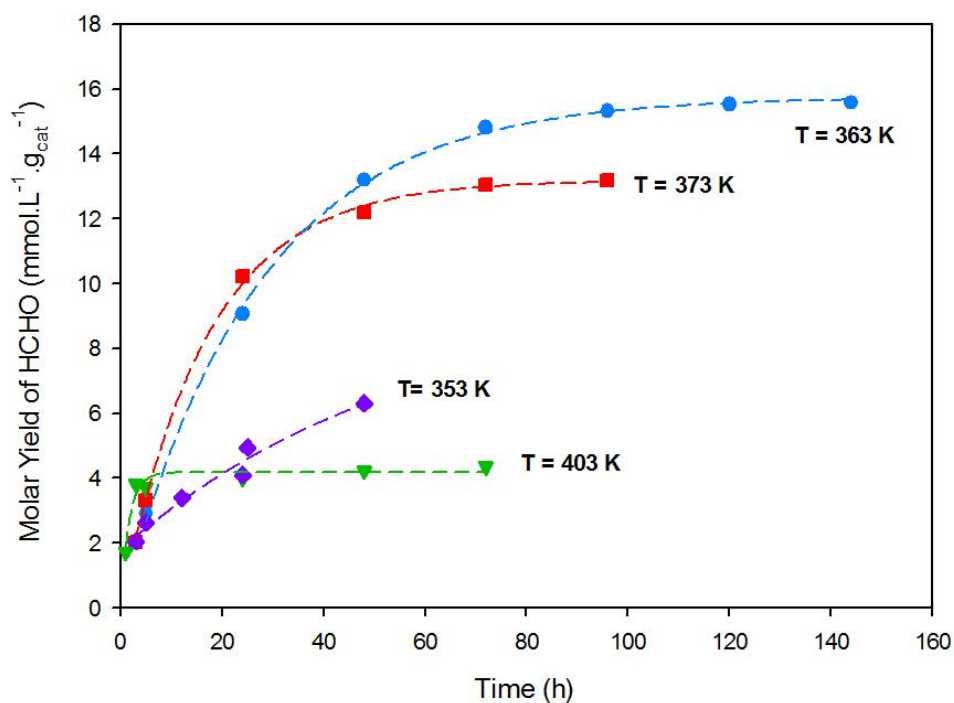


Figure 5.5: Effect of temperature on the molar yield of HCHO, solvent: CH<sub>3</sub>OH, P = 100 bar

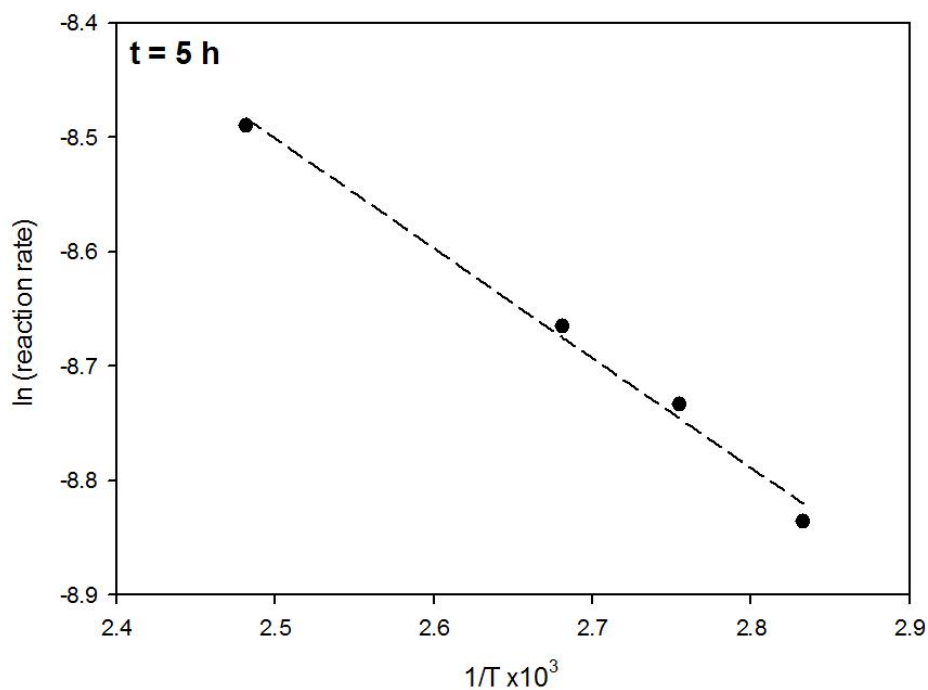


Figure 5.6: Arrhenius plot based on the rate of reaction calculated at t = 5 h, solvent: CH<sub>3</sub>OH, P = 100 bar, Catalyst: Ru-Ni/Al<sub>2</sub>O<sub>3</sub>

## 5.5 Effect of Solvents on Carbon Dioxide Hydrogenation Process

Hydrogenation of  $\text{CO}_2$  to produce HCHO was studied by using the identical reaction conditions used in CO hydrogenation in Chapter 3. The results are shown in Figure 5.7. Although the HCHO yield is significantly lower compared to CO hydrogenation, it was important to note that the catalytic hydrogenation of  $\text{CO}_2$  to produce HCHO is feasible in the slurry reactor. Similar to CO hydrogenation, the highest HCHO yield is achieved by using  $\text{CH}_3\text{OH}$  as the solvent.

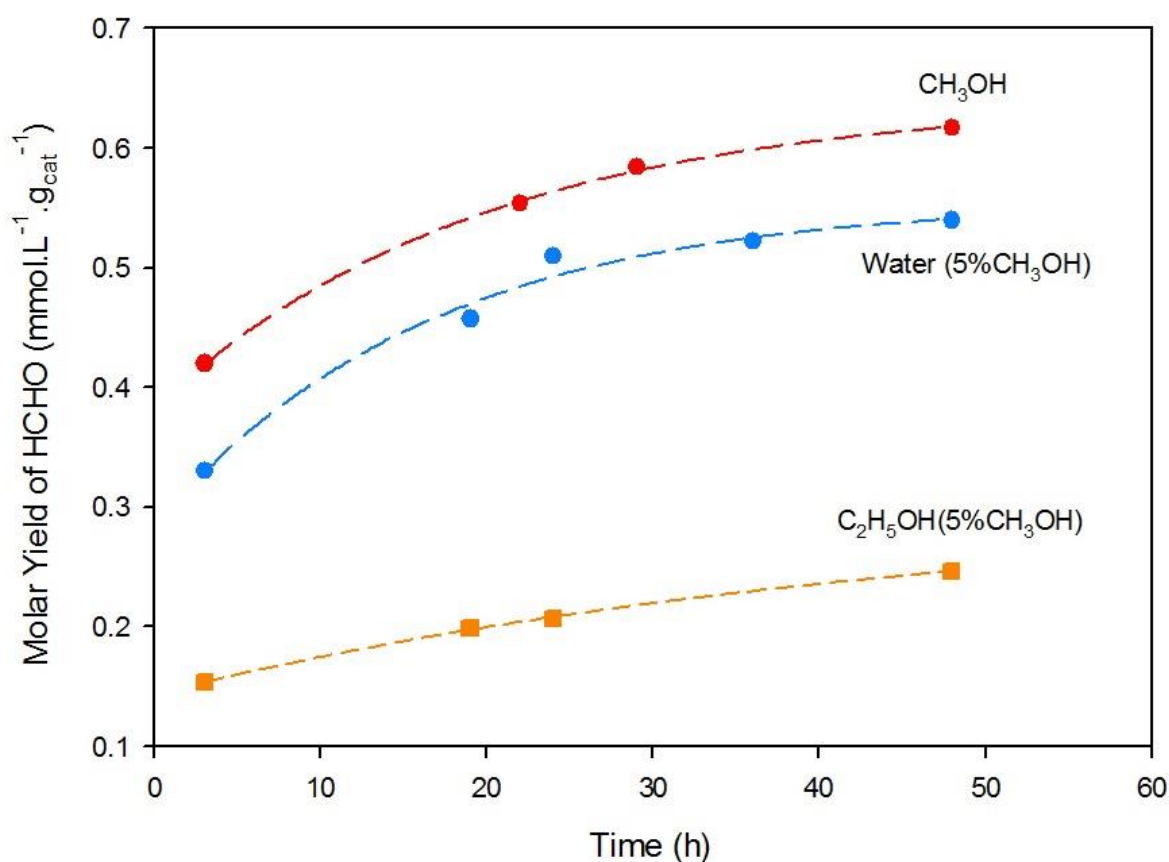


Figure 5.7:  $\text{CO}_2$  hydrogenation in the slurry reactor,  $T=353\text{ K}$ ,  $P=100\text{ bar}$ , Catalyst:  $\text{Ru-Ni/Al}_2\text{O}_3$

The  $\text{CO}_2$  hydrogenation results presented here are the preliminary results of this study. More detailed studies can be done on optimization, catalyst development, and the mechanism of this process in the future.

## 5.6 Conclusion

The effects of solvents on the hydrogenation of CO and CO<sub>2</sub> to produce HCHO in the slurry reactor were studied in this chapter. Many types of solvents including polar protic (water, methanol and ethanol), polar aprotic (DMSO), and non-polar (C<sub>7</sub>H<sub>8</sub>) solvents were used in the present study. It was shown that the highest HCHO yield was achieved by using pure CH<sub>3</sub>OH as the solvent followed by 0.1%v/v CH<sub>3</sub>OH-water mixture. HCHO reacts with CH<sub>3</sub>OH and water to produce C<sub>2</sub>H<sub>6</sub>O<sub>2</sub> and CH<sub>2</sub>(OH)<sub>2</sub>, respectively, thereby shifting the equilibrium of hydrogenation towards HCHO production. The CH<sub>3</sub>OH and water solvents stabilise the C<sub>2</sub>H<sub>6</sub>O<sub>2</sub> and CH<sub>2</sub>(OH)<sub>2</sub> molecules, respectively, via hydrogen bonding. Lower yield of HCHO was achieved by using water-CH<sub>3</sub>OH mixtures compared to pure CH<sub>3</sub>OH since water molecules may occupy the CH<sub>3</sub>OH cyclic oligomer rings, reducing the stability of HCHO molecules in the solvent. The effect of temperature on HCHO yield using CH<sub>3</sub>OH as the solvent was studied and the highest HCHO yield was found to be 15.58 mmol.L<sup>-1</sup>.g<sub>cat</sub><sup>-1</sup> at 363 K and 100 bar. It was observed that the reaction rate increased as the temperature increased. However, lower HCHO yield was achieved at higher temperatures and long residence times, since the solubility of the gases as well as the thermodynamic equilibrium of the reaction was less favourable at higher temperatures. CO<sub>2</sub> hydrogenation was investigated to show the feasibility of HCHO production in the catalytic reduction of CO<sub>2</sub> in the slurry reactor. The same trend for HCHO production can be seen in the CO<sub>2</sub> hydrogenation experiments. The highest HCHO yield through CO<sub>2</sub> hydrogenation at 353 K and 100 bar was achieved to be 0.62 mmol.L<sup>-1</sup>.g<sub>cat</sub><sup>-1</sup>.

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# Chapter 6

## Conclusions

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## 6.1 Introduction

In this thesis, a novel method of formaldehyde production was developed. Through this study, the effects of important parameters such as temperature, pressure, and stirring speed were examined. The reaction mechanism was investigated and the effect of solvents on the HCHO yield was studied. In this chapter, the main findings of this research are summarized.

## 6.2 Gas Phase Reaction and Slurry Phase Reaction

From the thermodynamic investigation of the CO hydrogenation reaction into HCHO, it is found that the reaction is non-spontaneous in the gas phase since the Gibbs free energy of the reaction is positive at room temperature and it increases with temperature. This was proven experimentally since the HCHO yield was insignificant with the maximum value of  $8.2 \times 10^{-3} \text{ mmol.L}^{-1}.\text{g}_{\text{cat}}^{-1}$ . Moreover, thermodynamic investigation of the slurry reaction showed significantly higher equilibrium constant of the reaction in the aqueous phase. Experimental results showed that the yield of HCHO in the aqueous phase reached  $4.55 \text{ mmol.L}^{-1}.\text{g}_{\text{cat}}^{-1}$  after  $t = 72 \text{ h}$ .

## 6.3 Effect of Temperature

The Gibbs free energy of CO hydrogenation reaction to form HCHO increases with temperature in both gas and aqueous phase reaction. It was found that in the slurry reactor increasing the temperature increased the reaction rate but the formaldehyde yield reached its maximum value at 363 K, above which the yield reduced. This is in agreement with the thermodynamic calculation. This was observed in all experiments conducted using the slurry reactor. Lower solubility of the reactant gases and formaldehyde in the liquid phase was also responsible to reducing the yield of formaldehyde at high temperatures. The effect of temperature on the stability of the catalyst support was studied in the aqueous phase reaction. Based on the XRD patterns and solid state NMR results, it was shown that the catalyst support went through phase change to form AlOOH after 120 h at 403 K. This phase change was not observed at 353 K which shows the dependence of this phenomenon on temperature.

### 6.4 Effect of Pressure

Based on the experimental data in the literature, it was noted that the Gibbs free energy of the reaction decreases as the pressure of the process increases. The results of this thesis confirmed the effect of pressure in both gas and liquid phase reactions. In the liquid phase higher pressures lead to higher gases solubility which is also an advantage. However, the operating pressure of in this thesis was limited to the maximum gas cylinders pressure available.

### 6.5 Effect of Stirring Speed

Another important parameter to consider was the stirring speed in the slurry phase reactor to study when the reaction changes from mass transfer controlled to kinetically controlled regime. It was observed that increasing the stirring speed from 0 to 800 rpm increased the HCHO yield since stirring decreased the mass transfer limitation by increasing the convective mass transfer coefficient and exposing the catalyst surface to the dissolved gases. However, increasing the stirring speed from 800 rpm to 1200 rpm did not affect the yield because the rate of mass transfer was sufficiently high at 800 rpm and further increasing the stirring rate did not affect the global rate of reaction because the reaction was kinetically controlled.

### 6.6 Reaction Mechanism

In order to investigate the process and improve the HCHO yield, the reaction mechanism was studied. By investigating the liquid samples for the presence of intermediate and by deuterium labelling technique, it was shown that the gases dissolve in the solvent and are adsorbed on the catalyst surface. While CO adsorbs as a molecule,  $H_2$  dissociates on the catalyst surface. Adsorbed CO molecule is then hydrogenated to form adsorbed HCHO. In water and methanol solvent mixture desorption of HCHO results in rapid hydration leading to methylene glycol formation. In the absence of  $CH_3OH$ , HCHO molecules are hydrogenated further to produce  $CH_3OH$ . Based on this reaction mechanism, it was concluded that using solvents with higher  $H_2$  and CO solubility may increase the HCHO yield.

## 6.7 Effect of Solvents

The effect of solvents on the yield of HCHO was studied in this project. Many types of solvents including polar protic (water, methanol, and ethanol), polar aprotic (DMSO), and non-polar ( $C_7H_8$ ) solvents were used in the present study. It was shown that the highest HCHO yield is achieved by using pure  $CH_3OH$  followed by 0.1% v/v  $CH_3OH$ /water solution. The rapid reaction of HCHO with water and  $CH_3OH$  to form  $CH_2(OH)_2$  and  $C_2H_6O_2$ , respectively, shifts the CO hydrogenation reaction equilibrium towards HCHO side which results in higher HCHO yield compared with using other solvents.

In this study, it was found that the solvent effects on CO hydrogenation to HCHO are not solely dependent on solubility. It was shown that using the solvents with higher ability to stabilise the produced HCHO molecules and its derivatives led to higher HCHO yield. It was observed that the HCHO yield in pure  $CH_3OH$  and 0.1% v/v  $CH_3OH$ /water solution was higher compared to concentrated  $CH_3OH$ -water solution. It was theorised that the tendency of the  $CH_3OH$  cyclic oligomers to stabilise the  $C_2H_6O_2$  molecules may decrease in the presence of water molecules, since water molecules may preferentially fill the interspace of the  $CH_3OH$  cyclic oligomers, resulting in reduced stability of  $C_2H_6O_2$  in water- $CH_3OH$  solutions. The highest HCHO yield was measured to be  $15.58 \text{ mmol.L}^{-1}.\text{g}_{\text{cat}}^{-1}$  which was achieved at 363 K using pure  $CH_3OH$ .

## 6.8 $CO_2$ Hydrogenation

The catalytic reduction of  $CO_2$  to HCHO was studied in the same reactor used for CO hydrogenation. In this study, it was shown that the catalytic  $CO_2$  reduction to HCHO was feasible. 5%v/v  $CH_3OH$  solutions of water and  $C_2H_5OH$  and pure  $CH_3OH$  were used for this process. Similar to CO hydrogenation, the highest HCHO yield was achieved by using  $CH_3OH$  as the solvent. The highest HCHO yield through  $CO_2$  hydrogenation at 353 K and 100 bar was achieved to be  $0.62 \text{ mmol.L}^{-1}.\text{g}_{\text{cat}}^{-1}$ . The  $CO_2$  hydrogenation reaction to HCHO should be studied further in future studies.

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# **Chapter 7**

## **Recommendations and Future Studies**

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## 7.1 Introduction

In this research, a novel method of HCHO production was introduced. Direct CO hydrogenation to HCHO was proven to be feasible in this study. However, there is a need for further investigation of this process prior to commercialisation. Considering the novelty of this process, there are several improvements and optimisation necessary. The key aspects to consider for future studies on direct conversion of synthesis gas to HCHO are summarised here.

## 7.2 Carbon Dioxide Hydrogenation

In this study, it was shown that catalytic CO<sub>2</sub> hydrogenation to HCHO is feasible. It is well known that activating CO<sub>2</sub> at low temperatures is more difficult compared to CO. This process needs to be optimized to achieve higher yield of HCHO. Considering the global issues of greenhouse gas emissions and global warming, utilizing CO<sub>2</sub> to produce value-added chemicals can be an interesting approach for future studies. It is recommended to use other catalysts to enhance CO<sub>2</sub> activation and hydrogenation. Using different solvents to absorb CO<sub>2</sub> on catalyst surface is important in this process. All aspects investigated for CO hydrogenation in this research, can separately be applied for CO<sub>2</sub> hydrogenation to optimize the HCHO yield in this process. It is also interesting to examine simultaneous hydrogenation of CO and CO<sub>2</sub> as they both are present in synthesis gas.

## 7.3 Catalyst Study

Only two catalysts, i.e. alumina supported Pd-Ni and Ru-Ni, were studied in this thesis. However, catalyst development is an essential part of a process in catalytic reactions. Other noble metals known for their high capability in hydrogenation reactions such as Pt, Rh, and Re can be examined in this process. Using higher content of noble metals also may improve the yield of HCHO. Moreover, it is important to examine other catalyst supports such as CeO<sub>2</sub>, ZnO<sub>2</sub>, and CeO<sub>2</sub>-ZnO<sub>2</sub>. The catalysts used in this research were prepared using wet impregnation method. Using other methods of catalyst preparation which may enhance the surface area and/or metal dispersion may improve the process.

## Chapter 7

It is important to investigate the mass of the catalyst while using pure  $\text{CH}_3\text{OH}$  to confirm that the experiments are conducted in a kinetically controlled regime.

Addition of noble metals as the promoters may enhance the reducibility and stability of the catalyst. However, it is very important to study the mechanism of promoter effect on the catalyst functionality in detail. It is therefore suggested that the molecular structure of the catalyst as well as the atom bindings should be investigated using experimental methods such as X-ray Absorption Spectroscopy (XAS) analysis.

### 7.4 Process Optimization

The feasibility of CO hydrogenation to HCHO was shown in this research. However, the HCHO molar yield was significantly lower compared to the industrial method of HCHO production. There is a potential to optimize the conditions of this process to overcome the relatively low productivity of this method. Considering the novelty of this method, not enough information was available in the literature to improve the conditions of the process. It is therefore suggested to investigate the conditions of the process using computational modelling and simulation to find the optimum conditions for the process. This approach can provide an insight into the mechanism of the process.

### 7.5 *In-situ* Study of the Reaction Mechanism

In this study, the reaction mechanism was investigated using *ex-situ* investigation of the potential intermediate products in the samples collected during the reaction in addition to isotope labelling. However, it is essential to investigate the catalyst surface to study the formed species and confirm the mechanism. *In-situ* ATR-FTIR methods using high pressure cell is recommended for further investigation of the reaction mechanism. The investigation of the catalyst surface combined with the computational modelling provides a better understanding of the reaction conditions. This may reveal the main obstacles of the process which is the first step to further improvement of HCHO production.

## 7.6 Investigation of the Solvent-Formaldehyde Interaction Theory

The stabilisation of formaldehyde molecules in the water-methanol mixtures was justified based on the previous findings on the water-methanol interactions in the literature. The presented theory in this thesis needs further investigation through molecular dynamic modelling and experiments. It is critical to confirm the developed theory based on theoretical and experimental studies for further improvement of the formaldehyde yield and gaining an insight into the molecular interactions between the solvent molecules and the desorbed formaldehyde molecules

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# Appendix:

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## A.1 Critical Review and Exergy Analysis of Formaldehyde Production Processes:

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Ali Mohammad Bahmanpour, Andrew Hoadley and Akshat Tanksale\*

### Critical review and exergy analysis of formaldehyde production processes

**Abstract:** Formaldehyde is one of the most important intermediate chemicals and has been produced industrially since 1889. Formaldehyde is a key feedstock in several industries like resins, polymers, adhesives, and paints, making it one of the most valuable chemicals in the world. However, not many studies have been dedicated to reviewing the production of this economically important product. In this review paper, we study the leading commercial processes for formaldehyde production and compare them with recent advancements in catalysis and novel processes. This paper compares, in extensive detail, the reaction mechanisms and kinetics of water ballast process (or BASF process), methanol ballast process, and Formox process. The thermodynamics of the reactions involved in the formaldehyde production process was investigated using HSC Chemistry™ software (Outotec Oyj, Espoo, Finland). Exergy analysis was carried out for the natural gas to methanol process and the methanol ballast process for formaldehyde production. The former process was simulated using Aspen HYSYS™ and the latter using Aspen Plus™ software (Aspen technology, Burlington, MA, USA). The yield and product specifications from the simulation results closely matched with published experimental data. The exergy efficiencies of the natural gas to synthesis gas via steam reforming, methanol synthesis, and formaldehyde synthesis processes were calculated as 60.8%, 61.6%, and 66%, respectively. The overall exergy efficiency of natural gas conversion into formaldehyde was found to be only 43.2%. The main sources of exergy losses were the steam reformer and methanol loss in formaldehyde synthesis process. Despite high conversions and selectivities of these processes, the low exergy efficiency suggests that innovations in formaldehyde production processes could give a more sustainable product. Novel methods of direct conversion of natural gas or synthesis gas into formaldehyde will improve the exergy efficiency, but the conversion rate must also be increased with advancements in catalysis.

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**Keywords:** exergy analysis; formaldehyde; kinetics and mechanism; methanol ballast process.

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### 1 Introduction

Formaldehyde is one of the simplest molecules in organic chemistry, which has a wide range of applications in chemical processes. It is produced in nature by photochemical processes in atmosphere or by incomplete combustion of organic materials (Reuss et al. 2003). Formaldehyde gas is colorless at ambient temperature with an irritating odor that unfavorably affects eyes and skin. The effects of formaldehyde on human health have been studied. Formaldehyde is toxic and probably a carcinogen because there is enough evidence that it can cause nasopharyngeal cancer (Cogliano et al. 2005). Some physical properties of monomeric formaldehyde can be found in Table 1 (Reuss et al. 2003).

Formaldehyde's heat of formation, heat capacity, and entropy are given in Table 2 for a wide range of temperatures and atmospheric pressure (Walker 1967).

The gas mixture with air is flammable at 293.15 K or higher (National Fire Protection Association 1994). Formaldehyde is considered as a reactive compound. At high temperatures (>423.15 K), it tends to undergo heterogeneous decomposition to produce methanol and carbon dioxide. However, if the temperature goes higher (>623.15 K), it may tend to form synthesis gas (hydrogen and carbon monoxide mixture) (Bone and Smith 1905, Calvert and Steacie 1951). It can be easily involved in both oxidation and reduction reactions. For example, in the presence of hydrogen, it can be reduced to form methanol over a nickel catalyst. Nitric acid, potassium permanganate, potassium dichromate, or oxygen can readily oxidize formaldehyde to form formic acid, carbon dioxide, and water (Walker 1967).

At atmospheric pressure, formaldehyde condenses at 253.95 K and freezes at 155.15 K as a white paste. In the liquid phase, its density is 0.8153 g/cm<sup>3</sup> at 253.15 K and

## A.2 Formaldehyde Production via Hydrogenation of Carbon Monoxide in the Aqueous Phase



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### Formaldehyde production *via* hydrogenation of carbon monoxide in the aqueous phase

Ali Mohammad Bahmanpour, Andrew Hoadley and Akshat Tanksale\*

Formaldehyde (HCHO) is an essential building block in many industries for producing value-added chemicals like resins, polymers and adhesives. Industrially, formaldehyde is produced *via* partial oxidation and/or dehydrogenation of methanol. Methanol is produced from natural gas in a series of processes, with synthesis gas as an intermediate. This study presents for the first time, formaldehyde production *via* hydrogenation of carbon monoxide in the aqueous phase, which eliminates the need for methanol synthesis, which may potentially save capital costs and reduce energy consumption. Gas phase hydrogenation of CO into formaldehyde is thermodynamically limited and therefore, resulted in a low CO conversion of only  $1.02 \times 10^{-4}\%$ . However, the aqueous phase hydrogenation of CO into formaldehyde was found to be thermodynamically favourable and kinetically limited. The highest CO conversion of 19.14% and selectivity of 100% were achieved by using a Ru–Ni/Al<sub>2</sub>O<sub>3</sub> catalyst at 353 K and 100 bar. The rapid hydration of formaldehyde in the aqueous phase to form methylene glycol shifts the CO hydrogenation reaction equilibrium towards formaldehyde formation. Increasing the pressure and stirring speed increased the yield of formaldehyde, whereas increasing the temperature above 353 K resulted in a lower yield.

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### Introduction

Industrially, formaldehyde (HCHO) is produced in three stages – (a) steam reforming of natural gas to produce syngas (Table 1, eqn (1)), (b) methanol (CH<sub>3</sub>OH) synthesis (eqn (2)) and (c) partial oxidation of CH<sub>3</sub>OH to produce HCHO (eqn (3)). Alternatively, HCHO is industrially produced *via* dehydrogenation of CH<sub>3</sub>OH (eqn (4)).<sup>1,2</sup> However, these are all high temperature reactions which require combustion, compression and large process units for purification, which are the root cause of energy losses.<sup>3,4</sup> We have recently shown that this series of processes from natural gas to HCHO production suffers from ~57% losses in exergy (*i.e.* energy quality).<sup>5</sup> Given the large quantity of HCHO produced in the world, when combined with the high losses in exergy, leads to high energy losses and also high CO<sub>2</sub> emissions, globally. Many researchers have tried to overcome this issue by finding ways to produce HCHO directly from natural gas by partial oxidation of CH<sub>4</sub> (eqn (5)).<sup>6–8</sup> However, there has been no significant progress to date due to low CH<sub>4</sub> conversion and poor selectivity.<sup>6</sup> The rate of HCHO decomposition into CO and H<sub>2</sub> is much greater than the rate of partial oxidation of CH<sub>4</sub>, particularly at

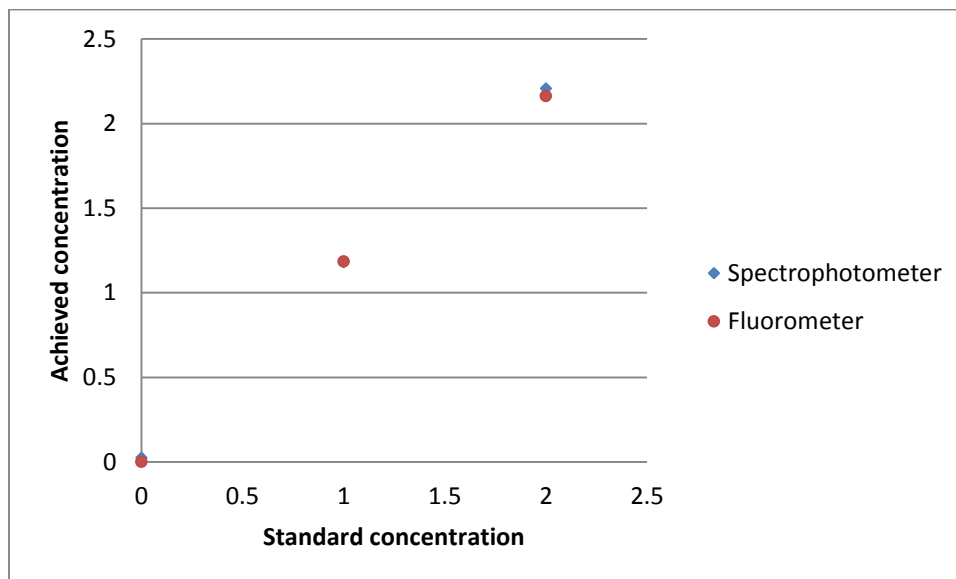
temperatures in excess of 373 K, which means that in order to produce HCHO selectively, one must limit the conversion of CH<sub>4</sub> in eqn (5) to a very small value. Moreover, if we consider that natural gas is a valuable energy resource, it becomes evident that an alternative feedstock for HCHO production is much needed. An alternative is direct conversion of synthesis gas into HCHO. Syngas can be produced from a range of sources including biomass and it allows the mole fractions of CO to H<sub>2</sub> to be controlled more easily by the use of H<sub>2</sub>O and CO<sub>2</sub> which are helpful in climate change abatement.<sup>9</sup> Gas phase hydrogenation of CO to produce HCHO (eqn (6)) is not feasible because of the positive Gibbs free energy change of the reaction.<sup>10</sup> Only a trace amount of HCHO in the product has been reported with the highest CO conversion of 0.2%.<sup>11</sup> Therefore, direct conversion of synthesis gas into HCHO has not been studied extensively. In this report, hydrogenation of CO into HCHO in a slurry reactor is presented as a viable alternative. By comparing with gas phase conversion in a fixed bed reactor, it is demonstrated that the thermodynamic limitation can be overcome in the slurry reactor. A low temperature active catalyst is desirable for the slurry phase reaction as the reaction was found to be favourable below 373 K. In general, Ni Pd and Ru are considered as active hydrogenation catalysts in the literature for many reactions.<sup>12–21</sup> Based on the density functional theory (DFT) studies performed in the literature, Pd and Ni have been shown to produce HCHO as an intermediate of CO hydrogenation to produce CH<sub>3</sub>OH.<sup>22,23</sup> Previous studies

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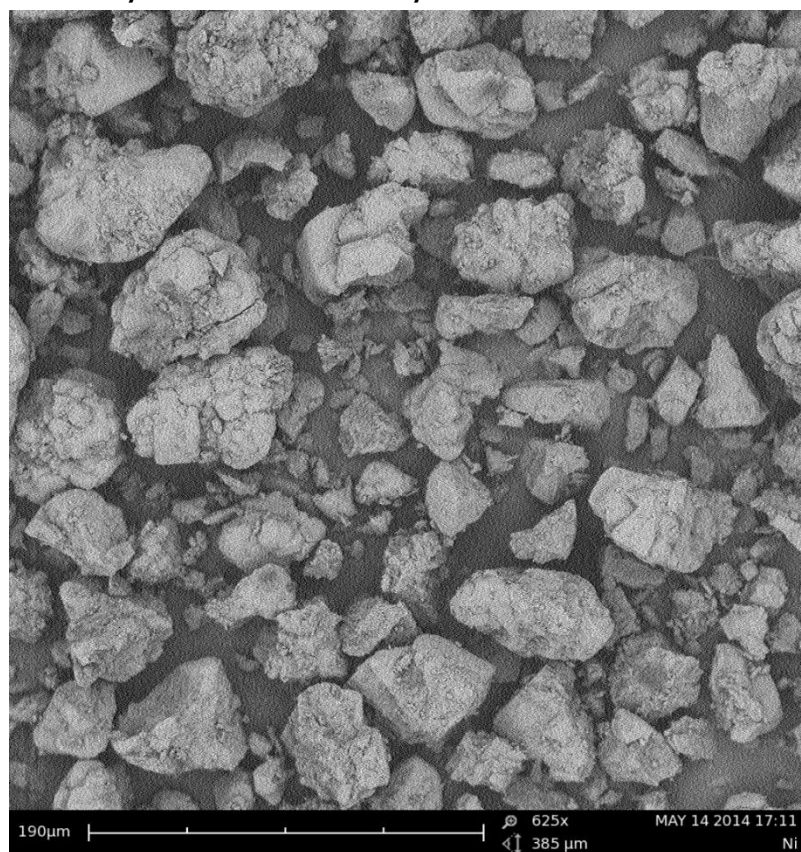
### A.3 Consistency of the Formaldehyde Detection Methods

The formaldehyde detection methods used in this thesis (i.e. Spectrophotometric method and fluorometric method) are compared here to evaluate the consistency and accuracy of these detection methods.

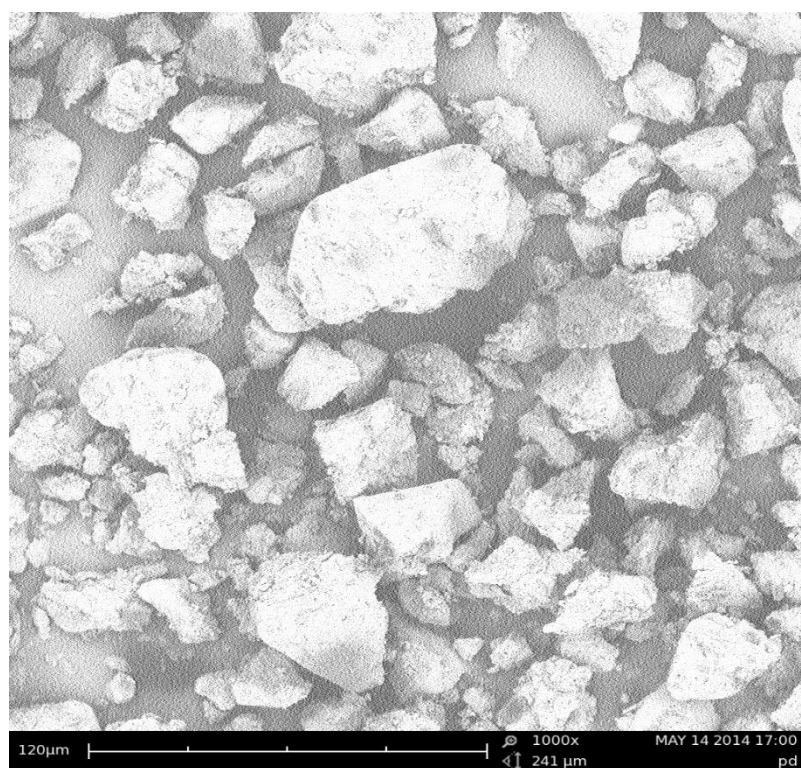


#### A.4 SEM Images of Fresh Ru-Ni/Al<sub>2</sub>O<sub>3</sub> and Pd-Ni/Al<sub>2</sub>O<sub>3</sub>

Ru-Ni/Al<sub>2</sub>O<sub>3</sub>:



Pd-Ni/Al<sub>2</sub>O<sub>3</sub>:



### A.5 Standard solution containing 400 ppm formalin, 100 ppm HCOOH and CH<sub>3</sub>OCOH on HPLC

