

# Greener Alternatives to Selected Organic Oxidation Reactions

By

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The undersigned certify that they have read, and recommended to the Environmental Science Unit (Division of Science) for acceptance, a thesis entitled "Greener Alternatives to Selected Organic Oxidation Reactions" submitted by Ryan P. Sheppard in partial fulfilment of the requirements for the degree of the Bachelor of Science, Honours.

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## 1.0: Abstract

The study of green chemistry is dedicated to eliminating or reducing toxic waste. One route to accomplish this goal is to explore alternative reaction conditions and parameters resulting in the development of more benign synthetic routes and reagents. The primary focus of this research is to find optimal reaction conditions for the oxidation of a primary alcohol to an aldehyde. As a case study, the oxidation of benzyl alcohol to benzaldehyde, a common industrial process, was examined. Traditionally carried out using the Jones Reagent, commonly referred to as chromium (IV) oxide or chromium trioxide ( $\text{CrO}_3$ ) in sulphuric acid, a great deal of research went into utilizing less toxic reagents, such as  $\text{MnO}_2$  or  $\text{KMnO}_4$  supported on a clay base. This research has led to an improvement on these alternatives, using a lithium chloride ( $\text{LiCl}$ ) catalyst in a montmorillonite K10 clay solid phase, together with the oxidizing agent hydrogen peroxide, as even greener alternatives to these traditional oxidizing agents. Experiments were carried out to determine the lifetime of this  $\text{LiCl}/\text{clay}$  system as compared to  $\text{MnO}_2$  and  $\text{KMnO}_4$ , to investigate its ability to catalyze the oxidation of other aromatic alcohols (such as 4-methoxybenzyl alcohol and diphenylmethanol), and to further improve the system's adherence to green chemistry principles. Green solvent alternatives were examined by replacing the toluene solvent with dimethylcarbonate (DMC), and reaction conditions were optimized to improve product yield. It was determined that the  $\text{LiCl}/\text{H}_2\text{O}_2$  system was, in most cases, equally as effective at catalyzing the oxidation of benzyl alcohol to benzaldehyde. Although the catalyst and oxidizing agent eliminated the toxic waste generated from chromium reagents, it offered significant challenges in product isolation, because of an aqueous-organic phase separation.

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## 2.0: Introduction

One of the major goals of modern industrial chemistry has been to maximize efficiency through optimizing reaction conditions. A primary focus of this optimization has been environmental protection and pollution prevention. Research in general towards green chemistry involves innovating new reaction types and conditions, to increase energy efficiency and decrease waste. This will result in both a reduced environmental and health impact and economic savings by reducing waste generation. In this context, this thesis applies principles and practices of green chemistry to a number of selected oxidation-reduction reactions.

### 2.1: Green Chemistry

Green Chemistry is based on twelve core principles, designed and developed to help reduce environmental waste<sup>1-3</sup>:

1. Prevention: Avoiding the generation of waste altogether.
2. Atom Efficiency: Maximizing yield while minimizing waste generation.
3. Non-Toxic Synthesis: Use and production of substances with little or no toxicity to people or the environment.
4. Safer Chemicals: Designing products to maximize their effects while reducing toxicity.
5. Safer Solvents and Auxiliaries: When not eliminated, the solvents and auxiliary substances used should be as non-toxic as possible, and in minimal volumes.
6. Energy Efficiency: The reduction of the energy required for a process, as well as the ideal setting of atmospheric temperature and pressure.
7. Renewable Feedstocks: Use of a feedstock or starting material derived from a renewable source is preferable.

8. Reduction of Derivatization: Unnecessary derivatizations, such as blocking and protecting groups, should be avoided.
9. Catalysts: Catalytic reagents are preferential to stoichiometric reagents.
10. Degradability: Ideally, chemical products should be decomposable or degradable into benign products.
11. Real-Time Analysis: Use of analytical techniques to monitor methods in real-time, to ensure the prevention of waste formation.
12. Safer Chemistry: Minimizing such risks as explosions, fires, and other accidents.

These principles of green chemistry form a basis to carry out and evaluate research. To provide some background, a number of these factors will be examined in greater detail. The need for prevention of toxic waste is the primary goal of this research, and several of laboratory techniques were used to accomplish this. These include:

- Examination of the use of a nontoxic oxidizing agent
- Comparison of a variety of catalyst preparation methods
- Comparison of reaction conditions in terms of yield, purity, and energy efficiency

To compare the effectiveness of these methods, some form of quantification is needed. One way to accomplish this involves analyzing a variety of factors of the reaction and its products, known as metrics. By using these metrics, a quick and convenient comparison of methods may be conducted.

## 2.2: Quantifying Green Chemistry

To properly compare and contrast any alternative pathway in the context of environmental protection, it is necessary to compare them quantitatively. By using these calculations, the advantages and disadvantages of a method be adequately assessed. The most traditional of all metrics for assessing a reaction, the percent yield, measures the yield obtained from a reaction experimentally, and expresses it as a ratio against the calculated theoretical yield, based on the reaction's stoichiometry. While this relationship gives a good indication of how effective a reaction is, it does not give any information about a number of crucial parameters. These include the amount of waste generated, the energy required, reaction and extraction solvent, and other important environmental factors. As these can quickly translate into economic costs, the development of more specialized metrics was required. While there are a wide variety of such parameters, the most prominent in this thesis are atom efficiency and environmental factor, chosen for their versatility and effectiveness.

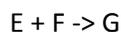
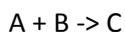
### 2.2.1: Atom Efficiency<sup>1,4</sup>

Atom efficiency (or atom economy) takes into account the stoichiometrics of a reaction when examining conversion from reactant to product. It uses the molecular weight of the final product of a synthesis, divided by the sum of the molecular weights of the reactants from which the product is created, to measure the waste produced as a by-product of a chemical reaction. Atom efficiency (AE) was introduced in 1991, by Trost<sup>1</sup>, and is widely used in green chemistry reaction calculations. It is a useful measure to compare the effectiveness of synthetic pathways, since it alludes to exactly how much of your starting material is incorporated into the final product. For a simple one-step reaction,  $A + B \rightarrow C$ , the AE is calculated as such:

$$\text{Atom economy} = \left( \frac{\text{m.w. of product C}}{\text{m.w. of A} + \text{m.w. of B}} \right) \times 100$$

4

This gets more complex with multi-step processes, whereby the sum of all non-intermediate products is taken as the denominator. Thus, for the multistep process<sup>2</sup>:



The following AE may be calculated:

$$= \left( \frac{\text{m.w. of product G}}{\text{m.w. of A} + \text{m.w. of B} + \text{m.w. of D} + \text{m.w. of F}} \right) \times 100$$

4

Here, C and E are excluded, as they are intermediate reactants which are produced in one step, and then immediately consumed in the next.

As illustrated, the AE is often expressed as a percent. The value allows for a quick and effective comparison of a number of synthetic pathways and processes. An AE of close to 100% indicates that all atoms of the starting material have been converted into products, while a lower value indicates a less efficient conversion. In addition, AE is noteworthy in that it considers catalytic components and acid/base addition as additional reactants, and furthermore, considers stoichiometric ratios and optical purity into the final result as well, applying the appropriate ratio to the molecular weight accordingly.

### 2.2.2: E-Factor<sup>4-5</sup>

A direct measure of the waste generated by a process can be calculated by its environmental factor, or E-Factor. Since this metric includes mass derived from solvents and other auxiliaries, it acts to quantify the mass of desired product against the mass of this waste, as such<sup>4</sup>:

$$\text{E Factor} = \frac{\text{Total waste (kg)}}{\text{kg product}}$$

Being a ratio of waste, here a higher E-Factor for a reaction is indicative of a higher level of waste comparative to the product formed, and as such is unfavorable. In general, this metric is fairly practical, as it offers a quick assessment of the waste a process will produce, despite the use of mass (kilograms) instead of the more chemically descriptive mole.

### 2.3: A Summary of Green Chemistry Metrics

There are a variety of ways to quantify the environmental impact of a chemical process. Each of the metrics relevant to this thesis are tabulated below:

Table 1: A Summary of Green Metrics

Metric	Measurement	Formula <sup>4</sup>
E-Factor (E-F)	Ratio of waste to product (kg/kg)	$\text{E Factor} = \frac{\text{Total waste (kg)}}{\text{kg product}}$
Atom Efficiency (AE)	Percent of starting material in product	$\text{Atom economy} = \left( \frac{\text{m.w. of product C}}{\text{m.w. of A} + \text{m.w. of B}} \right) \times 100$

There are other metrics which exist, often in the form of a highly specialized formula to yield specific quantities or figures. An example of this is Reaction Mass Efficiency (RME)<sup>4</sup>. Perhaps the most important of those not mentioned is the Life Cycle Assessment (LCA); this is an important consideration for all reagents, as it tracks the overall impact a species has on the environment, from creation to transport to disposal<sup>5</sup>. However, this metric is more relevant to industry, particularly the pharmaceutical

industry, and is less relevant for small-scale laboratory use. LCAs can also be difficult to quantify, and as such they will not be formally discussed as a means of quantitative analysis in this report.

These metrics help to draw a number of conclusions about the efficiency and environmental friendliness of an experiment, namely that these parameters rely heavily on experimental design. As a result, the main focus of these chemical reactions should be:

1. Maximization of main product
2. Minimization of by-products
3. Minimization of solvents and auxiliaries

For the purpose of this thesis project, all three of these factors were major considerations in experimental design, with the end goal being to reduce or eliminate toxic waste. Once this goal was established, a number of alternatives which may help support it were examined, and these will be discussed below in section 3.0.

### **3.0: Greener Approaches and Alternatives to Reactions**

With the goal of reducing toxic waste established, it is logical to proceed to the design and implementation of these factors into an experimental setting. There are many techniques which can be used to accomplish this, with three being examined in greater detail: The reduction or elimination of solvents, the reduction of energy usage, and the use of catalysts. Each of these are important in green chemistry, and may be used independently or in tandem to improve reaction efficiency.

#### **3.1: Use and Toxicity of Solvents and Reagents**

Solvents have always been a crucial part of any reaction. However, where possible, the use of solvents should be reduced or eliminated from green chemistry reactions. However, given the nature of organic synthesis, elimination is not always possible, therefore the choice and volume of solvent must be carefully considered. In choosing the best solvent for a given, its renewability, toxicity, and effectiveness must all be taken into account<sup>3</sup>. Often the most effective solvent is not always the safest or most renewable, and as an example, Table 2 on the following page contrasts the physical properties of a number of important industrial solvents:

Table 2: Physical Properties of Select Solvents<sup>1,5</sup>

Solvent	Boiling Point (°C)	Flash Point (°C)	Dielectric Constant (e')	Threshold Limit Value (ppm)	Hazards	Pros
Methanol	64	12	32.6	200	Toxic, flammable	Can be renewable
Ethanol	78	16	24.3	1000	Irritant, flammable	Can be renewable
Isopropanol	96	15	18.3	400	Irritant, flammable	
1-Butanol	117	12	17.1	100	Harmful, flammable	
Ethyl Acetate	76	-2	6.0	400	Harmful, flammable	
Ethyl Lactate	154	46		Unknown	Irritant, flammable	Renewable
THF	65	-17	7.4	200	Irritant, flammable	
2-Methyl THF	80	-11		Unknown	Irritant, flammable	Renewable
2-Butanone	80	-3		200	Irritant, flammable	
Dichloromethane	40	n/a	9.1	100	Toxic, harmful, suspected carcinogen	
Chloroform	61	n/a	4.8	10	Possible carcinogen	
Toluene	110	4	2.4	50	Irritant, teratogen, flammable	
Hexane	68	-26	1.9	50	Irritant, reproductive hazard, flammable	
Heptane	98	-4		400	Irritant, flammable	
Water	100	n/a	80.4	n/a	n/a	Renewable, non-flammable
Carbon Dioxide	n/a	n/a		5000	Compressed gas	Renewable, non-flammable
PEG-1000	n/a	n/a		n/a		Non-toxic, non-volatile

Many of the common organic solvents are at the very least flammable, with a large number of them being irritants or hazardous. Ultimately, solvents are vital to many reactions, and on a broad scale, to many other chemical processes<sup>5</sup>. This makes their outright elimination difficult, despite the dangers many pose. Thus, part of the research in green chemistry has been towards reducing the use of solvents, and to replacing them with more environmentally benign compounds whenever applicable<sup>2</sup>. While

perhaps not feasible for all reactions, there are cases where solvent elimination can occur, falling under the category of solid-state reactions.

Solventless and solvent-free reaction conditions both follow the guidelines of green chemistry, in that they incorporate a minimal use of solvents. However, the terms are often used interchangeably, and the difference lies in experimental process. Many of the common solventless reactions utilize a liquid reagent, or involve the generation of a liquid in situ, and still others utilize aqueous reagents, such as the 30% hydrogen peroxide used later in this study<sup>5</sup>. Another factor influencing the designation of solventless or solvent-free lies in purification and extraction<sup>1</sup>, as often a great deal of solvent is required to achieve these results following what may be a very green, solvent-free reaction. The issue of terminology between the two concepts can be reduced to a general lack of solvent, despite some liquid or highly concentrated aqueous reactants, as per a solvent-free reaction, compared to a strict lack of solvent and liquids, following a solventless protocol<sup>5</sup>. Regardless of which is employed, and to what degree the reaction is truly “solvent-free”, the important factor in this concept is in the reduction of solvent. Any reduction is a step in the right direction, and part of the problem lies in the scale of these reduced solvent conditions. Industry is able to conduct solvent-free reactions on the kg-scale, but it is still difficult to carry these procedures out on a true industrial level with our current technology<sup>5</sup>.

### **3.2: Energy Reduction**

Just as reagents and solvents can be assessed based on environmental impact, so too can the energy used to fuel these reactions. Regardless of type, the generation and transportation involved with generating electricity creates some degree of environmental impact, and as such, any reduction to energy usage can be seen as an improvement in process greenness, as stated in section 3.1. As such, two alternatives to traditional heating, microwave and sonication, will be examined briefly. While

neither were used in the research of this thesis, both are still effective at reducing energy costs, and form an integral part of the application of green chemistry principles.

### 3.2.1: Microwave Synthesis

The main principle behind microwave synthesis lies in a more efficient means of heating the reactants, thus a more efficient reaction overall. The Arrhenius equation describes this relationship, and is given as<sup>1</sup>:

$$k = Ae^{(-E_a/RT)}$$

From this, a definitive link between the temperature, activation energy, and reaction rate can be drawn<sup>1</sup>, and as such, the importance of proper heating is underlined. The main factor is the inefficiency of a classical reflux setup, as being a mechanical form of heating, be it a heating mantle, water bath, or hot plate, involves the transfer of heat from the medium through the glassware to the reactants. This causes a number of issues, both in the incomplete transfer of heat through the reaction vessel, as well as an uneven distribution of heat through the reactants<sup>1</sup>. Mixing can help abate this somewhat, but the issue remains that the heat transfer will not be quantitative, and physical heat transfer tends to be relatively slow when compared with a microwave technique<sup>1</sup>. By contrast, microwave radiation is able to penetrate most Teflon and glass vessels, and as such, much of the efficiency of this heating type depends on the boiling point and dielectric constant of the solvent<sup>6</sup>.

Another attribute of microwave heating lies in the ability to superheat the solvent, due to the inability of hot solvent to escape when contained in the central solvent body<sup>1,6</sup>. On average, this can result in some normally unsuitable, but relatively green solvents, such as methanol and ethanol, to be heated to nearly twice their normal boiling point (65 °C and 76 °C respectively), allowing for their use over other higher boiling-point compounds such as toluene<sup>1</sup>.

The applications of microwave synthesis are intriguing, however, it is not without fault. A number of safety concerns arise when using a standard microwave for such a use, among them are arcing from vapours, and the risk of vessel rupture and chemical spill<sup>1</sup>. While remediated somewhat by careful laboratory conduct, and the use of one sample at a time, it is also difficult to accurately monitor temperature over the course of reaction<sup>1</sup>. As such, industrial model microwaves were innovated for use in chemical reactions. Much more sophisticated than a regular commercial model, these large-scale microwaves possess built-in temperature and pressure monitoring, rotating racks, locking mechanisms, and corrosion-resistant linings in case of spills or ruptures. This is offset by a cost of approximately \$20,000 per unit<sup>1</sup>. However, this is a long-term investment, and a great deal of research has been done in the area of microwave synthesis in the past 20 years<sup>6</sup>, attesting to its efficiency and environmental friendliness.

### 3.2.2: Sonication

Similar to microwave synthesis, sonication is another potential alternative to traditional heating apparatuses. While less widely used, there are examples of research being conducted by using sonication techniques to enhance yields and product formation, and energy in the range of 500-800 kHz at 20-60 W is sufficient to cleave significant organic bonds, such as a glycosidic bond<sup>7</sup>. Literature examples of sonication, in the form of ultrasonic baths, proves promising, with significantly reduced reaction times and increased selectivity and yields, proving its utility in both synthesis and extraction<sup>8-9</sup>. It is for this reason that sonication is mentioned here as an alternative to traditional heating techniques.

### 3.3: Use of Catalysts

A catalyst is a substance which lowers the activation energy of a reaction by providing an alternate pathway. This is important for green chemistry, as the proper use of a catalyst offers the ability to reduce energy cost, allowing for a reaction to occur under less extreme conditions. Another key feature of a catalyst is its regeneration as part of the reaction; that is, a true catalyst is not consumed as a part of the reaction it is involved in, again contributing to the greenness of a reaction by not producing excess waste. Typically, a catalyst can be divided into a homogeneous or heterogeneous classification<sup>3</sup>. This refers to the catalyst being in the same phase as the reactant, such as a liquid catalyst in an aqueous reaction (homogeneous), or in a different phase, such as a solid metal powder in an aqueous reaction (heterogeneous)<sup>3</sup>. These heterogeneous catalysts are used in the methods of this thesis, as the use of a solid catalyst in a liquid medium has been the focus of research.

Entire textbooks have been written on the subject of catalysis, particularly in the context of green chemistry, and thus the focus of this thesis will be on those using new catalysts to prevent waste generation. The history of the catalysts in this experiment lie in their use in a typically difficult reaction, and the reduction of their toxicity is the goal of this thesis, as catalyst toxicity still contributes to the reaction as a whole.

## 4.0: Experimental Technique

In the experimental focus of this research, a number of green chemistry techniques and metrics were applied to a series of reactions themed around organic oxidation. Preliminarily, a number of reactions were carried out to determine the viability of alternative catalysts to a common reaction. Once this was established, the catalyst was examined by means of lifetime and versatility, and more effort was put in to further adhere to the principles of green chemistry. These reactions were themed around the selective oxidation of a primary alcohol, benzyl alcohol, to the aldehyde product, using lithium chloride (LiCl) and hydrogen peroxide (H<sub>2</sub>O<sub>2</sub>) as an alternative to potassium permanganate (KMnO<sub>4</sub>).

### 4.1: Establishing Experimental Conditions

Using classical organic chemistry methods and reagents, the oxidation of a primary alcohol to an aldehyde form was fairly difficult to accomplish. Unless great care was put into the choice of reagent, the corresponding carboxylic acid was generally the result of this reaction, as there is a strong tendency for oxidation to continue after the aldehyde is formed. This is illustrated in Figure 1:

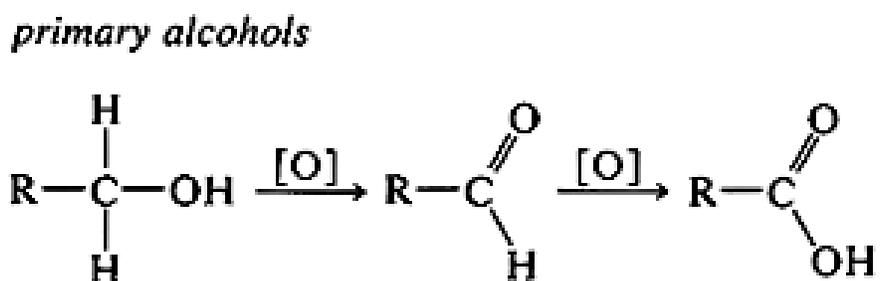


Figure 1: Schematic for the oxidation of a primary alcohol<sup>18</sup>.

One reliable way to obtain the aldehyde product was through use of the Jones Reagent, a CrO<sub>3</sub> solution in H<sub>2</sub>SO<sub>4</sub><sup>10,17</sup>, with other ways involving pyridine chlorochromate/dichromate (PCC/PDC)<sup>17</sup>. These reactions were quick and easy, with good yields, however problems arose due to the high toxicity

of chromium. In addition to being highly oxidative and corrosive, the  $\text{Cr}^{6+}$  reagent which was used possesses a high water solubility, and furthermore, this compound is highly toxic to humans, being a confirmed carcinogen<sup>11</sup>. As such, alternatives such as  $\text{KMnO}_4$  are frequently employed instead, due to their lower environmental and human toxicity. However, this is somewhat less convenient due to the need for a solid-phase support, such as a montmorillonite clay.

Unfortunately, manganese is not entirely environmentally benign either. Despite being an essential nutrient for soil organisms, it can be highly toxic to plants and soil microorganisms, interfering with a number of key metabolic pathways in bacteria<sup>12</sup>, and causing severe localized oxidation when absorbed in excess by plants<sup>13</sup>. As such, a great deal of research has been done into improving both the methodology and nature of this reaction and its catalysts. Recently a paper has cited the use of  $\text{LiCl}$  alongside  $\text{H}_2\text{O}_2$  as being an effective alternative to  $\text{KMnO}_4$ <sup>14</sup>, and these compounds are both much greener reagents as well.  $\text{LiCl}$  is analogous to  $\text{NaCl}$ , both metals being low-mass, monovalent alkali metals, and  $\text{H}_2\text{O}_2$  rapidly decomposes to water and oxygen gas, both excellent examples of green mineralization. As such, the primary focus of this thesis has been on the establishment of the  $\text{LiCl}/\text{H}_2\text{O}_2$  catalyst system in comparison to  $\text{KMnO}_4$ , and the extensions of it for future reactions.

A solid-phase support is needed for this reaction as well, and as such, a brief discussion on the montmorillonite K-10 support is needed to fully establish the catalyst properties.

#### 4.1.1: The Montmorillonite K-10 Catalyst

The clay catalyst used in conjunction with  $\text{KMnO}_4$ , a montmorillonite K10 clay, is of the zeolite branch of catalysts. Being a 2:1 aluminosilicate clay, it consists of an octagonal inner aluminum unit, with tetrahedral silicon units on either side, as illustrated in figure 2 on the following page. As a result of this, the clay possesses a large interior surface area, and a high surface charge, allowing for effective ion exchange capacity<sup>17</sup>. As these smectite-type clays have the ability to act as both an acid and base, this

allows for electron-accepting/oxidizing sites to be found along the edges, or embedded in the inner structure, of the clay<sup>15</sup>.

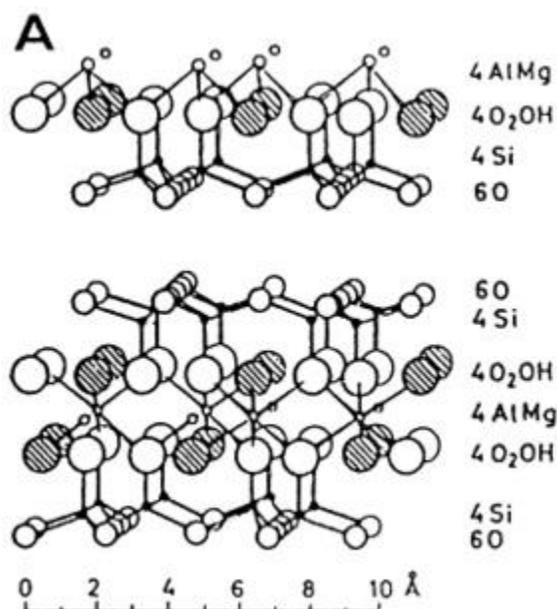


Figure 2: The interior structure of a typical montmorillonite clay<sup>12</sup>.

#### 4.1.2: Experimental Design

For the reactions conducted in this phase of research, benzyl alcohol was reacted with both the  $\text{KMnO}_4$  and  $\text{LiCl}/\text{H}_2\text{O}_2$  systems under various conditions to selectively form benzaldehyde. Summarized below are the reactions, as well as the conditions and catalyst system used:

Table 3: A Summary of Establishing Green Parameters

Catalyst System	Reaction Type
Montmorillonite – $\text{KMnO}_4$	Room temperature mixing
Montmorillonite – $\text{LiCl}/\text{H}_2\text{O}_2$ w/Mixed Catalyst	Room temperature mixing
Montmorillonite – $\text{LiCl}/\text{H}_2\text{O}_2$ w/Mixed Catalyst	Reflux
Montmorillonite – $\text{LiCl}/\text{H}_2\text{O}_2$ w/Refluxed Catalyst	Room temperature mixing
Montmorillonite – $\text{LiCl}/\text{H}_2\text{O}_2$ w/Refluxed Catalyst	Reflux

Experimental procedure for the room temperature mixing was conducted as follows<sup>17</sup>: For the initial solid-phase reaction using  $\text{KMnO}_4$ , 9.0 g of a 2:1 montmorillonite – K10 clay: $\text{KMnO}_4$  reagent was collected in a 100.0 mL beaker, to which was added 0.6000 mL of benzyl alcohol, collected via a 1.0000 mL Eppendorf pipette. The mixture was stirred for 5 minutes, then sealed and left to react for a period of 5-7 days. Following reaction, the benzaldehyde was extracted via filtration, using dichloromethane as the solvent, and the excess solvent was extracted via evaporation under pressure to yield pure benzaldehyde. A similar procedure was repeated using K10 clay and LiCl, using approximately 7 g of a 5:1 K10 clay:LiCl, along with 5.0 mL of 30%  $\text{H}_2\text{O}_2$  in place of  $\text{KMnO}_4$ .

Reflux reactions were conducted differently, on a smaller scale, and were carried out as follows<sup>14</sup>: 0.1000 mL of benzyl alcohol, collected via a 1.0000 mL Eppendorf pipette, was added to a 250 mL round bottom flask, to which was added 15.0 mL of toluene and 0.25 g of the montmorillonite – K10 catalyst. Following this, 5.0 mL of 30%  $\text{H}_2\text{O}_2$  was added while the solution was stirred in an ice bath, and after a five minute period, the solution was refluxed with vigorous stirring for approximately one hour. The solution was filtered to remove the K10 catalyst, and washed with dichloromethane. To this, excess anhydrous magnesium sulphate ( $\text{MgSO}_4$ ) was added to remove the aqueous component, followed by a second filtration. Finally, the excess toluene and dichloromethane was removed via evaporation under pressure to yield pure benzaldehyde as the product.

Two different procedures were also used to prepare the solid catalysts. Catalyst A was prepared as per the CHEM 2400 laboratory manual, with the mechanical mixing of a 20:80 mass/mass mixture of LiCl:Montmorillonite, and a 1:2 ratio of  $\text{KMnO}_4$ :Montmorillonite, being prepared in a mortar and pestle<sup>15</sup>. Catalyst B was prepared by a more sophisticated method, as per Khalili, Abdollahifar, and Hashemi in 2007, with a 20:80 proportion of LiCl:Montmorillonite being prepared by dissolving 1.95 g of LiCl in 100.0 mL of deionized  $\text{H}_2\text{O}$ . Following this, 8.02 g of Montmorillonite was added, and the solution was refluxed with stirring for approximately three hours<sup>15</sup>. The resulting solution was decanted, then

evaporated under pressure to remove as much water as possible, followed by oven-drying for a 24h period at 120 °C. The exact configuration of this catalyst is discussed below, as well as the application of similar zeolite catalysts.

#### 4.1.3: Green Chemistry Metrics of the Establishment Reaction Conditions

Two procedures were conducted as part of the establishment of experimental conditions: The solid-phase oxidation of benzyl alcohol to benzaldehyde using both a K10-Permanganate and a K10-LiCl/H<sub>2</sub>O<sub>2</sub> catalyst, as well as the oxidation of benzyl alcohol to benzaldehyde using the same two catalysts under reflux conditions. Given that the catalyst for each reaction was not collected for reuse following the reactions, it will be considered as part of the waste. Further experimentation was conducted to analyse the collection and reuse of this solid phase, and will be discussed in section 4.2.

Metrics were calculated and tabulated below, as per the formulas in table 1, using the following reactions:

- A) Solid-Phase: C<sub>6</sub>H<sub>5</sub>CH<sub>2</sub>OH (108.14 g/mol) + CAT → C<sub>6</sub>H<sub>5</sub>CHO (106.12 g/mol) + Waste Cat (9.0 g)
- 2x 15mL dichloromethane (1.325 g/mL; 84.93 g/mol) washes
- B) Reflux: C<sub>6</sub>H<sub>5</sub>CH<sub>2</sub>OH (108.14 g/mol) in 15.0 mL Toluene (0.865 g/mL, 92.14 g/mol) + CAT → C<sub>6</sub>H<sub>5</sub>CHO (106.12 g/mol) in Toluene + Waste Cat (0.25 g)
- 4x 15mL dichloromethane washes
  - ~20g of MgSO<sub>4</sub> (120.36 g/mol)

Table 4: Theoretical Green Chemistry Metrics from Established Experimental Conditions

Metric	Solid-Phase	Reflux
E-Factor	79.2 kg/kg	1099.8 kg/kg
Atom Efficiency	98.13%	98.13%

#### 4.1.4: Results and Discussion of the Establishment of Reaction Conditions

Samples were analyzed using both thin-layer chromatography (TLC) and infrared (IR) analysis, and percent yields were calculated based on the mass obtained after rotary evaporation of excess solvent. TLC analysis was conducted using 50:50 dichloromethane:diethyl ether solution, and IR analysis was conducted using neat samples of each liquid product, between two NaCl disks, using a Nicolette 6700 IR spectrophotometer. The results are tabulated below:

Table 5: Results from Results from Establishing Experimental Conditions:

Reaction	Mass recovered (g)	Expected Recovery (g)	Percent yield (%)	TLC Purity	IR Purity
Solid-Phase w/Mixed Catalyst	0.7195	0.6153	116.9%	Fair	Poor
Reflux w/Mixed Catalyst	0.1050	0.1025	102.4%	Fair	Good
Reflux w/Reflux Catalyst	0.2716	0.1025	265.0%	Fair	Good
Solid-Phase w/Reflux Catalyst	1.1330	0.6153	181.2%	Fair	Poor

As evidenced above, the percent yields for all experiments were suspect, and it is likely that there are impurities remaining in the solution; in the case of the reflux experiments, it is likely that there is residual toluene present in the solution. The solid-phase reactions may contain some residual water, dichloromethane, or other unknown contaminants. Furthermore, the yields from the use of the reflux catalyst are dramatically higher than their predicted yields, by a factor of close to two in both cases, and as a result it is possible that some catalyst contamination has occurred. As a result of this, the procedure involving the simple mixing of LiCl and K10 clay was determined to be the most reliable catalyst preparation technique.

Most results from TLC analysis indicated a fair conversion rate to benzaldehyde, but with some detectable impurities. IR analysis; however, yielded the relative impurities present in all samples, with only the two reflux reactions showing any quantitative conversion to benzaldehyde.

In terms of green chemistry metrics, the reflux reactions are fairly inefficient in terms of auxiliary reactants, solvents, and catalysts, requiring a large amount of wash solvent to properly extract the products. As a result, the E-Factor for this reaction is incredibly high, despite the high purity of product obtained. By contrast, the solid-phase reactions possessed a very low E-Factor, as no reaction solvent was used in the procedure, only a small volume of dichloromethane as an extraction solvent. Unfortunately, the downside of this is the relatively low purity of products obtained. The need for a drying agent further hindered the E-Factor of both of these reactions, and is an issue when using hydrogen peroxide in these processes, introducing water into otherwise anhydrous conditions. Finally, the extraction solvent used in both reactions would ideally be reduced or eliminated, as an extensive amount of washing was required for both procedures to obtain pure product, and the dichloromethane used in this is neither green nor safe.

The results from this experiment clearly indicate the effectiveness of the K10-LiCl/H<sub>2</sub>O<sub>2</sub> catalyst system, and are comparable to KMnO<sub>4</sub> for selectively oxidized aldehyde product. Following these preliminary successes, this technique was extended to examine the reusability and versatility of this K10-LiCl/H<sub>2</sub>O<sub>2</sub> catalyst, as well as to attempt to improve the green metrics of the reaction.

## 4.2: Determination and Comparison of Catalyst Lifetimes

The next experiment was conducted to determine the lifetime of the LiCl/K10 solid catalyst. Given the relative success of the reflux reaction when used with a freshly ground catalyst, this experimental model was used to conduct repeated experiments, collecting and reusing the solid catalyst each time.  $\text{KMnO}_4$  and  $\text{MnO}_2$ , two common oxidizing agents for a similar primary oxidation reaction, were also ground with montmorillonite K10 clay. Again, these were reacted, collected, and reused in a similar fashion, with the goal being the selective oxidation of benzyl alcohol to benzaldehyde.

### 4.2.1: Experimental Design

For each reaction, approximately 1.0 g of the solid-phase catalyst was added to 15 mL of toluene and 0.600 mL of benzyl alcohol. In addition, to the LiCl system, 5.0 mL of 30%  $\text{H}_2\text{O}_2$  was added, and each solution was then refluxed for 3 – 3.5 hours. Following this, the solid catalyst was filtered from the liquid organic-phase, which was then subjected to rotary evaporation.

The LiCl/ $\text{H}_2\text{O}_2$  reactions involved a more complicated refinement, with a distinct aqueous-organic interface developing from the use of  $\text{H}_2\text{O}_2$  in conjunction with the toluene/benzyl alcohol reactants. This made collection of the solid LiCl/K10 catalyst difficult, and a number of methods were attempted to streamline this. Collection of the entire aqueous phase, which contained the solid catalyst, proved to be modestly effective, while any attempt to filter both phases at once resulted in poor yields and loss of product and catalyst. Addition of anhydrous magnesium sulphate ( $\text{MgSO}_4$ ) quickly and effectively separated the two phases, but resulted in difficulty capturing the catalyst for reuse.

The resulting benzaldehyde solutions from each trial were analyzed using TLC, with a 1:1 solution of hexane isomers:dichloromethane as the TLC extraction solvent, testing against both benzaldehyde and benzyl alcohol, using dichloromethane to dilute the samples. As well, IR analysis was conducted using a Nicolette 6700 infrared spectrophotometer, with solid NaCl disks as the matrix.

#### 4.2.2: Results and Discussion of Catalyst Lifetime Comparison

As each reaction was repeated three times for each catalyst system, this resulted in nine reactions in total for this procedure, which are summarized below:

Table 6: Results from the Comparison of Catalyst Lifetimes

Reaction	Mass recovered (g)	Expected Recovery (g)	Percent yield (%)	TLC Purity	IR Purity
KMnO <sub>4</sub> – Trial 1	0.4324	1.025	42.19%	Fair	Good
KMnO <sub>4</sub> – Trial 2	2.7834	1.025	271.6%	Fair	Fair
KMnO <sub>4</sub> – Trial 3	5.0594	1.025	493.6%	Fair	Poor
MnO <sub>2</sub> – Trial 1	0.4438	1.025	43.30%	Fair	Good
MnO <sub>2</sub> – Trial 2	0.5977	1.025	58.31%	Poor	Good
MnO <sub>2</sub> – Trial 3	6.0262	1.025	587.9%	Poor	Good
LiCl – Trial 1	0.6360	1.025	62.05%	Fair	Fair
LiCl – Trial 2	0.0978	1.025	9.541%	Fair	Poor
LiCl – Trial 3	1.6997	1.025	165.8%	Fair	Fair

TLC plates and IR spectra were judged qualitatively, from poor to fair to good to excellent, against pure benzaldehyde and benzyl alcohol. For a number of reasons, the yields and purity, in general, declined over each subsequent reaction, and these issues were different for each catalyst used.

Percent Yields are somewhat erratic, which may be attributed to solvent contamination of some samples. In particular, all three of the products for the third catalyst use exhibit masses far beyond what would normally be expected. Again, this is likely due to a combination of solvent contamination, perhaps in conjunction with a low product volume, as extended rotary evaporation did little to remove this excess volume. Also worthy of note is the result from LiCl trial 2, in which an aqueous LiCl/K10 clay catalyst was recovered from trial 1 and used in place of the usual oven-tried solid. As this yield is particularly small, it is evident that this technique is unsuitable to the reaction, and it highlights the need for a dry catalyst.

Reactions with  $\text{KMnO}_4$  were moderately successful when used repeatedly. Easily filtered, dried, and reused, it can be seen as the standard to which  $\text{MnO}_2$  and  $\text{LiCl}/\text{H}_2\text{O}_2$  can be compared in this regard. However, after the second trial, the catalytic capability of this reagent declined significantly, following the prediction of declining capability, which can be observed by the sharp decline in the aldehyde spectral peaks.

Reactions using  $\text{MnO}_2$  exhibited strong IR peaks in all three trials in this region. However, TLC testing for the second and third trial reveal that the product spots traveled further than even the benzaldehyde spot, suggesting that an alternative oxidation product had been formed. It is unclear exactly what occurred in this reaction, however it can be determined that  $\text{MnO}_2$  is not suitable alongside the K10 clay for this reaction over prolonged usage.

The  $\text{LiCl}/\text{H}_2\text{O}_2$  system, finally, has exhibited fairly good conversion rates, even over prolonged usage. However, the main concern with the reuse of this catalyst lies in the nature of the system itself, as mentioned in experimental design. The resulting solution following reflux with this catalyst, unlike the other strictly solid-phase systems, is an organic-aqueous mixture resulting from the addition of  $\text{H}_2\text{O}_2$ , with the solid catalyst having a strong affinity towards this aqueous phase. As such, it is difficult to filter and recover this catalyst, as it has a strong tendency to clog filter paper with immiscible layers. The aqueous layer was captured and reused as is for the second trial, however this resulted in a poor IR spectral quality, due to contamination with water. Yield for the third trial was likely low due to attempts to improve filtration speed. Among solvents used as a wash to attempt to break the immiscible layers, such as water and dichloromethane, there was none which could adequately solve this issue. Addition of approximately 10 g of anhydrous  $\text{MgSO}_4$  was the most efficient way to separate the organic and aqueous layers, as this technique successfully bound both the aqueous phase and the leftover solid catalyst, leaving the product in an organic phase. This technique did, however, render the solid catalyst unusable for future trials, and thus cannot be used when recovery is desired.

Thus, in theory, the LiCl/H<sub>2</sub>O<sub>2</sub> catalyst functions quite well after repeated reuse, outperforming the KMnO<sub>4</sub> catalyst in terms of product quality after the third use. However this catalyst can be troublesome to recover and reuse due to the more complex nature of the solution. Efficient use required a large quantity of anhydrous MgSO<sub>4</sub>, heightening its waste production significantly.

### **4.3: Application of Green Principles**

Given the focus on green chemistry, the next series of experiments was aimed improving the environmental friendliness of the procedures established in the first set of reactions. To that end, two main experiments were designed: The improvement of yield, and thus the E-Factor, by increasing the quantity of reactant, and the replacement of toluene with dimethylcarbonate (DMC), a less toxic and renewable alternative. In addition to utilizing reflux reactions to improve yields and E-Factors, a number of solid-phase reactions were also conducted to attempt to improve both the reaction efficiency and E-Factor of this process.

#### **4.3.1: Experimental Design**

For the first series of experiments, three samples of 1.00 mL, 2.00 mL, and 3.00 mL of benzyl alcohol were collected in an Eppendorf pipette and was dissolved in enough toluene to bring the total volume to 15 mL. To this was added 1.0 g of a prepared 2:1 montmorillonite K10:LiCl solid matrix, followed by 5.0 mL of H<sub>2</sub>O<sub>2</sub> dropwise under cooling and stirring. Once addition was complete, each reaction was refluxed for three hours, approximately 10 g of anhydrous MgSO<sub>4</sub> was added to remove the aqueous layer, after which the organic layer was filtered using dichloromethane as a wash solvent. The resulting solution was evaporated under pressure using a rotovap, then analyzed via TLC and IR.

Similar to this, a second set of experiments was conducted to compare toluene and DMC as reaction solvents, with an identical set-up. Two vessels were prepared with 1.00 mL of benzyl alcohol,

with 15 mL of toluene added to a control, and 15 mL of DMC added to the test vessel. Otherwise, reaction and processing procedures were identical to those above.

Finally, a series of solid-phase reactions were conducted to assess the viability of this procedure. Two sets of reactions, one using 1.0 g of 2:1 montmorillonite K10/LiCl solid matrix, and the second using 5.0 g of the same matrix, were carried out using increments of 1.00 mL, 2.00 mL, and 3.00 mL of benzyl alcohol. Each aliquot was added to a separate 50 mL beaker under cooling and stirring, to which was added 5.0 mL of H<sub>2</sub>O<sub>2</sub>. The resulting mixture was stirred for approximately 10 minutes, then left to react for one week before extraction and analysis. Approximately 10 g of anhydrous MgSO<sub>4</sub> was added to each of the 1 g trial samples, with no drying agent being used in the 5 g trials. Samples from both experiment were extracted with dichloromethane, then filtered and evaporated under pressure using a rotovap. Finally, each sample was analyzed using TLC and IR analysis.

TLC for all products was conducted using a 1:1 solution of methylpentane mixtures and dichloromethane, using the latter to dilute each sample before spotting against a sample of benzaldehyde and benzyl alcohol. IR analysis was accomplished using a Nicolette 6700 Infrared Spectrophotometer, using NaCl plates as a matrix for the liquid-phase organic samples.

#### 4.3.2: Results and Discussion of the Application of Green Principles

Following the completion of all three aspects of this experiment, the results were tabulated below:

Table 7: Results from the Application of Green Principles

Reaction	Mass recovered (g)	Expected Recovery (g)	Percent yield (%)	TLC Purity	IR Purity
Reflux – 1 mL	2.6974	1.025	263.1%	Good	Good
Reflux – 2 mL	7.0074	2.050	341.8%	Good	Fair
Reflux – 3 mL	1.9411	3.075	63.13%	Poor	Fair
Reflux - Toluene	1.0417	1.025	101.6%	Good	Good
Reflux - DMC	1.2411	1.025	121.1%	Poor	Poor
1 g Solid-Phase – 1 mL	No product	1.025	N/A	No reaction	N/A
1 g Solid-Phase – 2 mL	No product	2.050	N/A	No reaction	N/A
1 g Solid-Phase – 3 mL	No product	3.075	N/A	No reaction	N/A
5 g Solid-Phase – 1 mL	0.2549	1.025	24.87%	Good	Fair
5 g Solid-Phase – 2 mL	0.6652	2.050	32.45%	Good	Fair
5 g Solid-Phase – 3 mL	0.8901	3.075	28.95%	Good	Fair

Both TLC plates and IR spectra were judged qualitatively, from poor to fair to good to excellent, and both were compared to pure samples of both benzaldehyde and benzyl alcohol. The set of solid-phase reactions using 1 g of catalyst did not yield any product. As such, it is likely that there was insufficient catalyst to allow the reaction to occur, thus a 5 g series of reactions was conducted to further test the abilities of a solid-phase setting to carry out this reaction.

Percent yields for the reflux-phase reactions, as seen previously in sections 4.1 and 4.2, are erratic, which is likely due to the contamination of the product with the toluene solvent. Repeated rotary evaporation did little to remove this excess volume, nor did extended idle evaporation. Yields from the solvent trials are likely also due to solvent presence in the product, though to a lesser extent. The 5 g solid-phase reactions, however, display accurate percent yields, albeit somewhat disappointingly low. This low yield may be due to the entrapment of some product in the solid matrix, as the result after a week of reaction was a fairly solid mass of catalyst and product.

The solvent comparison study shows that the use of DMC as a solvent in this reaction is unsuitable. TLC analysis showed that the majority of product was more closely in the benzyl alcohol phase, versus the benzaldehyde, and IR analysis corroborates this, showing much weaker signals in the aldehyde range when compared to the toluene trial. This may be due to a side-reaction occurring between the DMC solvent and the oxidizing agent, however the exact nature of this interference remains unknown at this time.

For the reflux trials, TLC analysis shows a marginal decline in reaction completeness as the reactant volume increases. IR analysis again supports this, as there is a significant decline in signal strength between the 2 mL and 3 mL trials in the aldehyde range, confirming a decline in conversion. As such, 2 mL of benzyl alcohol per 1 g of solid catalyst is likely the upper limit of reactive capacity. Any further addition of product will lead to a decline in product purity, as seen from these results.

TLC and IR analysis of the 5 g solid-phase trials indicated the presence of product, albeit with a relatively high contamination with what is likely water. It appears that this technique was moderately successful, with TLC plates indicating a good, but incomplete conversion of the alcohol to aldehyde, and IR indicating a strong water band above the aldehyde region. This technique can thus be appraised as a moderate success, given the low yields observed and the intermediate purity, with the main advantage being the lack of solvent necessary to facilitate the reaction.

#### 4.4: Versatility of the LiCl/H<sub>2</sub>O<sub>2</sub> Catalyst

Thus far, the reactions of this LiCl/H<sub>2</sub>O<sub>2</sub> catalyst have been restricted to that of the oxidation of benzyl alcohol to benzaldehyde. This last series of reactions attempted to test the ability of this system to oxidize other alcohols, using a number of similar reactants in identical reaction procedures. 4-methoxybenzyl alcohol, possessing an electron-donating CH<sub>3</sub>O- group at the para position to the CH<sub>2</sub>OH on the benzene ring, as well as 4-nitrobenzyl alcohol, possessing an electron-withdrawing O<sub>2</sub>N- group. In addition to these two, diphenylmethanol, a secondary alcohol, was also reacted to test the ability of the LiCl/H<sub>2</sub>O<sub>2</sub> catalyst to oxidize alcohols to ketones.

##### 4.4.1: Experimental Setup

As before, reflux reactions were conducted using each of the alcohols above. Furthermore, samples were ran for both KMnO<sub>4</sub> as well as LiCl, with an extra benzyl alcohol reaction being conducted as a control for the KMnO<sub>4</sub> system. Molar equivalents standardized for 0.100 mL of benzyl alcohol were used for these reactions, resulting in 0.971 mmol of each alcohol. This solid was dissolved in 15 mL of toluene, followed by the addition of 1.0 g of the solid-phase 2:1 K10 clay:KMnO<sub>4</sub>. For the LiCl samples, a further addition of 5.0 mL of H<sub>2</sub>O<sub>2</sub> was conducted under stirring in an ice bath. Each reaction was then refluxed for 3-3.5 hours, after which the KMnO<sub>4</sub> samples were corked and stored for three days before filtration, given inclement weather. The LiCl samples were processed immediately after reflux was complete. Following cooling, both samples were filtered, with approximately 10 g of anhydrous MgSO<sub>4</sub> being added to the LiCl samples beforehand to separate the two phases. Finally, excess solvent was removed under rotary evaporation. Similar to before, the resulting products were analyzed with TLC, using a 1:1 methylpentane:dichloromethane extraction solution, dichloromethane as the dilution solvent, and benzyl alcohol and benzaldehyde as reference spots. IR analysis was also performed using the same Nicolette 6700 series infrared spectrophotometer on NaCl disks.

#### 4.4.2 Results and Discussion of the Versatility of the LiCl/H<sub>2</sub>O<sub>2</sub> Catalyst

Results from this experiment are tabulated below, as well as the mass and molar ratio used for each reaction:

Table 8: Results from the Versatility of the LiCl/H<sub>2</sub>O<sub>2</sub> Catalyst

Reaction	Mass used (g)	Moles used (mmol)	Expected Recovery (g)	Mass recovered (g)	Percent yield (%)	TLC Purity	IR Purity
KMnO <sub>4</sub> /Benzyl Alcohol*	0.105	0.971	0.1025	0.2204	215.0%	Fair	Fair
KMnO <sub>4</sub> /4-methoxybenzyl alcohol	0.1351	0.978	0.1332	1.6102	1208%	Good	Poor
KMnO <sub>4</sub> /4-nitrobenzyl alcohol	0.1507	0.984	0.1487	0.3947	265.4%	Good	Poor
KMnO <sub>4</sub> /diphenylmethanol	0.1790	0.972	0.1771	0.2381	134.4%	Fair	Good
LiCl/4-methoxybenzyl alcohol	0.1345	0.974	0.1326	0.0943	71.11%	Good	Good
LiCl/4-nitrobenzyl alcohol	0.1504	0.982	0.1484	0.2945	198.5%	Good	Good
LiCl/diphenylmethanol	0.1793	0.973	0.1773	0.3062	172.7%	Fair	Poor

\*Benzyl Alcohol was collected as a liquid, all other alcohols were collected as solids.

TLC plates were judged qualitatively, from poor to fair to good to excellent. IR spectra were compared against those on the National Institute of Standards and Technology (NIST) website<sup>19</sup>.

Percent yields are, again, indicative of a large solvent volume present in each sample, save for the LiCl-catalyzed 4-methoxybenzyl alcohol reaction. For all KMnO<sub>4</sub> samples, this was quite evident, as each took the form of a clear, colourless solution, in contrast to the clear yellow solutions formed from the LiCl samples (Save for 4-nitrobenzaldehyde, which formed as dark yellow crystals in a clear yellow solution). Given this, it is possible that the KMnO<sub>4</sub> catalyst which was prepared was flawed in some way, or there was some other variable at work to yield this abnormal product. This may also be due to the waiting period the KMnO<sub>4</sub> samples experienced before filtration and evaporation, however this is somewhat unlikely. Regardless, the inability to completely remove the toluene solvent remains to be a problem whenever reflux reactions are conducted.

When the IR spectra of these products is compared to their respective pure compound, it appears that the LiCl/H<sub>2</sub>O<sub>2</sub> catalyst is more efficient at converting other primary alcohols, while KMnO<sub>4</sub> is better at carrying out the secondary alcohol reaction. Furthermore, while not significant, it also appears that the 4-nitrobenzaldehyde product is spectrally less pure than that of the 4-methoxybenzaldehyde. This indicates a possible effect of electron donation/withdrawal on the oxidation process. It is possible that the electron donor effect of the methoxy group assists in the oxidation process, though the precise nature of this interaction falls outside the scope of this research.

The reaction producing diphenylmethanone, commonly referred to as benzophenone, is another interesting example, as the LiCl/H<sub>2</sub>O<sub>2</sub> catalyst failed to oxidize this compound as readily as the KMnO<sub>4</sub> reagent. It is unclear as to the reason behind this behaviour, as the opposite was displayed by the KMnO<sub>4</sub> system, perhaps indicating a preference towards primary alcohols.

## 5.0: Discussion

A recurring issue in the analysis of the data obtained from these experiments lies with the contamination of liquid-phase samples with the reaction solvent, toluene in most cases. While toluene's high boiling point ( $>100^{\circ}\text{C}$ ) allows it to function as an excellent reflux solvent, this also causes issues in terms of removal, as rotary evaporation can be quite slow to remove this compound. Furthermore, in some samples, the excess solvent failed to evaporate given extensive time spent evaporating under pressure, potentially indicating issues with the rotovap operation or parameters. Naturally, solid-phase reactions avoid this issue altogether, however these types of reactions, as explored in experiments 4.1 and 4.3, typically have issues with product yield and purity instead, as well as the extended period of time which must be allotted for these reactions to be completed.

Based on the data obtained from this research, the ability of the  $\text{LiCl}/\text{H}_2\text{O}_2$  catalyst to oxidize primary alcohols appears to be similar to that of  $\text{KMnO}_4$ . In terms of dealing with catalyst lifetime, while  $\text{LiCl}/\text{H}_2\text{O}_2$  appears to maintain its catalytic activity longer than  $\text{KMnO}_4$ , the difficulty isolating the catalyst after use causes a great many problems in this area. The aqueous-organic interface caused by the use of  $\text{H}_2\text{O}_2$  also resulted in difficulty filtering the organic product, and the isolation of the aqueous phase with the solid catalyst leads to subpar product formation.  $\text{KMnO}_4$  can simply be filtered from the organic phase, dried, and reused, despite its catalytic degradation. This streamlines the process greatly when compared with the use of  $\text{LiCl} \cdot \text{MnO}_2$ , the last oxidizing agent tested in this trial, had good lifetime and a similar ease of use to  $\text{KMnO}_4$ , but yielded anomalous product after the second and third reuse, a possible indication of its inability to selectively oxidize to the aldehyde product. Neither of these systems are the perfect choice for a reusable solid-phase catalyst for this reaction, as each has clear drawbacks. However, given the possible development of a more effective separation and recovery technique, it is possible that  $\text{LiCl}/\text{H}_2\text{O}_2$  may be the best choice here, given its apparent lack of catalytic degradation. This

trait may be due to the addition of fresh  $\text{H}_2\text{O}_2$  over each trial, effectively acting as a new oxidizing agent; this is somewhat troublesome in terms of waste, as while the LiCl/K10 matrix can be recovered with some effort, the  $\text{H}_2\text{O}_2$  loses all effective oxidizing power as it quickly degrades to water. This represents the main advantage of using this system, as it makes the process much less toxic than the other two catalyst systems.

Examining the improvement of the LiCl/ $\text{H}_2\text{O}_2$  catalyst's green metrics provided some modest results. As discussed on page 29, it was possible to improve the yield of benzaldehyde by increasing the quantity of benzyl alcohol, but this method was effective up to 2 mL per 1 g of catalyst; further addition of alcohol resulted in a decline in product formation. Perhaps most disappointing was the failure of dimethylcarbonate (DMC) to perform as a replacement solvent to toluene, given its status as a renewable solvent, and its lower toxicity. The exact nature of this interference is unknown, and more research is needed to identify the problem. Finally, the solid-phase reactions are an interesting examination of improving this reaction's green metrics, as the use of no solvent and less LiCl/K10 clay would greatly boost this process' E-Factor. However, none of the reactions which used 1 g of LiCl/K10 produced an oxidation product, indicating that without a solvent and heat, it comes down to the quantity of solid present. The 5 g trials confirmed this, as product was formed with the increase in catalyst mass. However, the benzaldehyde isolated from this reaction was in relatively low yield, and upon IR analysis, was found to contain significant quantities of water.

The determination of catalyst versatility was another interesting series of reactions, as this set of procedures probed the ability of LiCl/ $\text{H}_2\text{O}_2$  to oxidize other alcohols, both primary and secondary, compared to the traditional  $\text{KMnO}_4$  reagent. 4-methoxybenzyl alcohol and 4-nitrobenzyl alcohol were chosen for their similarities to benzyl alcohol, as well as to test the effects of electron withdrawal and donation on this oxidation process. In addition, diphenylmethanol was chosen to test the ability of LiCl/ $\text{H}_2\text{O}_2$  to oxidize secondary alcohols, again given its structural similarity to benzyl alcohol. It was

found that, while  $\text{LiCl}/\text{H}_2\text{O}_2$  was able to catalyze these other two primary alcohols quite effectively, it was unable effectively produce the ketone product from the secondary alcohol. In contrast,  $\text{KMnO}_4$  was better suited for the oxidation of this secondary alcohol than that of the primary alcohols tested, indicating a difference in preference for these two procedures. The exact nature of this behaviour is unclear, as one would expect  $\text{H}_2\text{O}_2$  to be a much more powerful oxidizing agent than  $\text{KMnO}_4$ , however it is possible that there was some error in reaction during this phase as well. Each of the solutions resulting from the  $\text{KMnO}_4$  trials were clear and colourless, in contrast to the dark yellow solution which formed from the diphenylmethanol reaction, and the small yellow crystals which formed as a product from both the 4-methoxy and 4-nitrobenzyl alcohol reactions. Traces of product were found in all  $\text{KMnO}_4$  reactions, however, indicating that some reaction did occur, thus the nature of this phenomena is unclear.

## 6.0: Conclusion

The results of this study show that LiCl/H<sub>2</sub>O<sub>2</sub> is an effective oxidizing agent in the selective oxidation of primary alcohols to aldehydes. Further results show that this system is as effective as KMnO<sub>4</sub> supported on a solid catalyst. Neither reagent is perfect and the search will continue for an ideal catalyst for such a reaction: a reagent with persistent catalytic activity, efficient reaction conditions with little clean-up, and minimal environmental toxicity. As it is, while the use of LiCl/H<sub>2</sub>O<sub>2</sub> requires a drying agent, it is much less toxic when compared to chromium and manganese. For this reason, LiCl/H<sub>2</sub>O<sub>2</sub> makes an effective alternative to these metals as an oxidation reagent. Much more research remains to be conducted with this system, thus it will be interesting to examine further studies as it appears to have a great deal of potential in catalyzing these reactions. In terms of waste, the resulting water and LiCl solution is nontoxic compared to the CrO<sub>3</sub> and KMnO<sub>4</sub> alternatives. It is for this reason that the LiCl/H<sub>2</sub>O<sub>2</sub> catalyst offers an alternative to traditional oxidation reagents. As we move forward with our chemical innovations, we now must keep green chemistry at the forefront of our choices with careful consideration being made towards the toxicity and environmental impact of our reagents. This thesis has shown that innovation and research can introduce new reaction conditions and catalysts into the mainstream of chemistry and acts as a practical application of green chemistry.

## 7.0: References

1. *Green Organic Chemistry in Lecture and Laboratory*; Dicks, A.P., Ed.; Boca Raton, FL: CRC Press, 2012, P. 1-21
2. Anastas, P.T., Williamson, T.C. *Green Chemistry: Frontiers in Benign Chemical Syntheses and Processes*; Oxford [England]; New York: Oxford UP, 1998, P. 1-21
3. Anastas, P. T., Williamson, T.C. *Green Chemistry : Designing Chemistry for the Environment.*; Washington, DC: American Chemical Society, 1996, P. 1-16
4. Constable, D. J. C., Curzons, A.D., Cunningham, V.L. Metrics to 'green' Chemistry—which Are the Best? *Green Chem.* **2002** 4(6), 521-27.
5. Kerton, F. M., Clark, J.H., Kraus, G.A. *Alternative Solvents for Green Chemistry*; Cambridge, UK: RSC, 2009, P. 1-21
6. Desai, K. R., Kanetkar, V.R. *Green Chemistry Microwave Synthesis*. Rev. ed; Mumbai [India]: Himalaya Pub. House, 2010, P. 1-9
7. Tanifum, C. T, Chang, C-W. T. Sonication-assisted Oligomannoside Synthesis. *J. of Org. Chem.* **2009** 74(2), 634-44.
8. Fang, R. H, Aryal, S., Hu, C-M. J., Zhang, L. Quick Synthesis of Lipid-polymer Hybrid Nanoparticles with Low Polydispersity Using a Single-step Sonication Method. *Langmuir: ACS J. of Surf. and Coll.* **2010** 26(22), 16958-62.
9. Carrera, C., Ruiz-Rodríguez, A., Palma, M., Barroso, C.G. Ultrasound Assisted Extraction of Phenolic Compounds from Grapes. *Anal. Chim. Acta* **2012** 732, 100-104.
10. Pfitzner, K.E., Moffat, J.G., A New and Selective Oxidation of Alcohols *J. Am. Chem. Soc.* **1963** 85(19), 3027–3028
11. Gad, S C. Acute and Chronic Systemic Chromium Toxicity. *Sci. of the Total Env.* **1989** 86(1-2), 149-157.
12. Hohle, T. H., O'Brian, M.R. Magnesium-dependent Processes Are Targets of Bacterial Manganese Toxicity. *Mole. Micro.* **2014** 93(4), 736-47.
13. Maksimović, D.J., Mojović, M., Maksimović, V, Römheld, V., Nikolic, M. Silicon Ameliorates Manganese Toxicity in Cucumber by Decreasing Hydroxyl Radical Accumulation in the Leaf Apoplast. *J. of Exp.Bot.* **2012** 63(7), 2411-2420.
14. *Chemistry 2400 – Introductory Organic Chemistry I Laboratory Manual*. Corner Brook, NL: Memorial University of Newfoundland, Grenfell Campus, 2012, P. B-56 – B-58
15. Khalili, B., Abdollahifar, A., Hashemi, M. M. Transition Metal Free Oxidation of Alcohols to Carbonyl Compounds Using Hydrogen Peroxide Catalyzed with LiCl on Montmorillonite K10. *ACTA CHIMI.SLOV.* **2007** 54(3), 635-637.
16. March, J. *Advanced Organic Chemistry: Reactions, Mechanisms, and Structure*. 2<sup>nd</sup> ed; New York, NY: McGraw-Hill Inc., 1977, P. 1082-1085
17. Vaccari, A. Preparation and Catalytic Properties of Cationic and Anionic Clays. *Cat.Today* **1998** 41(1-3), 53-71.
18. <http://www.organicchemistry.com/oxidation-methods-of-alcohols/>
19. <http://webbook.nist.gov/chemistry/>