HETEROGENEOUS CATALYSIS OF GLUCOSE MUTAROTATION
BY ALUMINA

by

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ABSTRACT

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The kinetics and mechanism of the heterogeneous catalysis of the mutarotation of glucose by alumina have been investigated. Various types of aluminas held in suspension, in dimethyl sulfoxide, were used.

At 25.0°C, the first order kinetic plots for mutarotation by alumina neutral (for thin layer chromatography; γ-form) were curved due, first, to relatively slow adsorption of glucose on alumina and, second, to progressive deactivation of the catalyst. Partially deactivated catalysts produced linear first order plots over three half lives and hence, glucose mutarotation by alumina is a first order reaction. Further, there were <1% side products formed during the surface reaction.

Dehydration of the catalyst at low temperatures (i.e. upto 600°C) decreased the catalytic activity, unlike the other reactions studied on alumina surfaces. On further dehydration at higher temperatures the catalytic activity increased, and the activity per unit area of α-alumina (= 3.6 x 10^-5 sec^-1 m^-2) formed at 1250°C was about 26 times that of the standard alumina neutral. High catalytic activity for the α-form of alumina compared with the γ-form was previously virtually unknown. Further, this α-alumina did not deactivate during catalysis and produced linear first order plots over three half lives.

Adsorption studies showed the presence of (0.70 ± 0.02) x 10^-4 moles of irreversible adsorption sites on the surface of a gram of alumina neutral. The isotherm for adsorption of glucose on alumina neutral showed only monolayer adsorption. The Langmuir plot for reversible adsorption of glucose on the surface showed the presence of two types of reversible adsorption
sites; $(1.0 \pm 0.1) \times 10^{-4}$ moles of strong adsorption sites with an equilibrium constant for adsorption $K_1 = (8.2 \pm 1.2) \times 10^2$ litre mole$^{-1}$, and $(1.4 \pm 0.2) \times 10^{-4}$ moles of weak adsorption sites with an equilibrium constant for adsorption $K_2 = 44 \pm 3$ litre mole$^{-1}$. The study of the variation of initial rate with concentration of $\alpha$-D-glucose showed that only the weak adsorption sites are catalytically active. Hence the active site density on alumina neutral was obtained as $(5.4 \pm 0.6) \times 10^{13}$ sites/cm$^2$. The turnover number of a catalytic site was determined to be $2 \times 10^{-3}$ molecules/site/sec. This is one of the highest turnover numbers for a reaction catalyzed by an alumina surface.

The observed first order rate constant for the surface reaction,

$$G_\alpha + C \xrightleftharpoons[k_2]{k_1} G_\alpha \cdot C \xrightarrow[k_4]{k_3} G_\beta \cdot C \xrightleftharpoons[k_1]{k_2} G_\beta + C$$

was shown to be $k_{\text{obs}} = \frac{(k_3 + k_4) \text{[Catalytic Sites]}}{k_2/k_1 + \text{[Glucose]}}$, where $k_1 = K_2$. The catalytic constant $(k_3 + k_4)$ for the interconversion of $\alpha$-D-glucose ($G_\alpha$) and $\beta$-D-glucose ($G_\beta$) on the surface was determined to be $5 \times 10^{-3}$ sec$^{-1}$. Comparison with the catalytic constant for mutarotation in pure water ($= 4 \times 10^{-4}$ sec$^{-1}$) showed that the alumina surface offers a better medium for mutarotation than water. Further, the activity of the catalytic sites on alumina neutral is about 9 times that of strong acids in water.

The inhibitory effects of 'neutral' molecules (water, methanol, methyl glucoside, inositol etc.) indicate that the glucose adsorption sites on alumina neutral are relatively specific for adsorption of polyhydroxy compounds. In addition, aldehyde (e.g. hexanal) groups seem to interact preferentially with the catalytic sites on the alumina surface.

Studies with acidic (carbon dioxide) and basic (e.g. pyridine,
n-butylamine) inhibitor molecules suggest that the catalytic activity of alumina towards glucose mutarotation is due to the presence of basic oxide ions and weak Brönsted acid sites on the surface. About 85% of the activity of α-alumina formed at 1250°C is due to these basic sites, while the weak Brönsted acid sites give rise to about 90% of the activity of alumina neutral. The observed high catalytic activity of these weak Brönsted acid sites is probably due to the stabilization of the transition state leading to acyclic intermediate by the polar alumina surface.

Normal deuterium isotope effects were observed with alumina neutral ($k_H/k_D = 1.3$) and the other aluminas prepared by dehydration of alumina neutral (e.g. $k_H/k_D = 1.9$ with α-alumina formed at 1250°C). There was no isotope effect on the adsorption-desorption process and hence, the observed isotope effect is due to the catalytic reaction on the surface. Therefore, the observed normal isotope effects indicate that glucose mutarotation on alumina surface is a general acid-base catalyzed reaction, and occurs by a consecutive mechanism, via the acyclic intermediate. There is probably no bifunctional catalysis of the glucose mutarotation on the alumina surface. Further, the acid sites seem to show an isotope effect of 1.2 and the basic sites an isotope effect of 2.1.

Hence, these studies have shown that glucose mutarotation differs (e.g. higher catalytic activity of the hydrated surface, high catalytic activity of α-alumina) from many other reactions studied on alumina surfaces. This difference in behaviour under mild conditions is probably due to the high sensitivity of the mutarotation reaction to the weak acidic and basic sites on the alumina surface.
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Dedicated

to

My Parents
I INTRODUCTION
I. INTRODUCTION

Catalysis is the phenomenon in which relatively small amounts of foreign material called a catalyst, augments the rate of a chemical reaction without the catalyst itself being consumed. In heterogeneous catalysis the catalyst forms one phase, usually a solid, and the reactants and products are present in one or more fluid phases (gas or liquid)\(^1\). The catalytic reaction occurs on the surface of the solid and the catalyst provides a mechanistic pathway not present in its absence.

Heterogeneous catalytic systems have many advantages and also some disadvantages over homogeneous catalysts. Work up is easy with a solid catalyst since filtration would remove the catalyst and for reactions at gas/solid interface there is no solvent involved. Because of strong surface forces, reactions normally not possible in solution, can take place on a solid catalyst. They are also easily adapted to continuous processes which is a great advantage for industrial applications. In spite of these advantages heterogeneous catalysts are easily poisoned and they may not be used as efficiently as a homogeneous catalyst since only the atoms on the surface are available for the catalytic reaction.

I.1 Catalytic Aluminas

Aluminas (i.e. various forms of Al\(_2\)O\(_3\)) have been used extensively as adsorbents, active catalysts, and catalyst supports\(^2\). Already in 1797 the alumina-catalyzed dehydration of ethanol was discovered by Dutch chemists and as early as 1914, Sabatier\(^3,4\) reviewed the use of aluminas as active catalysts for various reactions. Since that time the application of aluminas in catalytic processes have increased tremendously. In industrial catalytic
processes, aluminas are mostly used as catalyst supports\textsuperscript{5,6}. Oxides and mixed oxides as well as transition metals and noble metals are supported on alumina. Thus chromia-alumina catalysts are being used for the conversion of paraffins and olefinic hydrocarbons, in hydrodealkylation of aromatics and to a lesser extent in catalytic reforming. The latter process is also catalyzed by molybdena-alumina, a catalyst system which is also active for making toluene and other aromatics from saturated hydrocarbons. It also catalyzes the isomerization of paraffins. Another catalytic system of enormous importance is cobalt oxide-molybdenum oxide-alumina which is widely used for hydrodesulfurization, hydrodenitrogenation and hydrocracking reactions. In all these cases there is ample evidence that alumina is far from a passive inert support\textsuperscript{7,8}.

The application of pure aluminas as catalysts in industrial processes is of less importance, although they are used, for example, in the catalytic conversion of the side products in the oxo process, such as alkyl esters and high boiling condensation products\textsuperscript{2}. In academic research, pure aluminas are used widely for several groups of reactions, some of which are summarized in Table 1.

The reactions compiled in the table show that aluminas are able to activate hydrogen-hydrogen, carbon-hydrogen, carbon-carbon, carbon-oxygen and oxygen-hydrogen bonds, although with varying efficiency. Thus o-p-H\textsubscript{2} and H\textsubscript{2}-D\textsubscript{2} equilibration reactions occur at very low temperatures, and C-H bond activation in exchange and isomerization reactions is effective near room temperature. For C-C bond activation, for example in skeletal isomerization, aluminas are less active. These reactions require much higher temperatures roughly above 325°C.
TABLE 1

REACTIONS CATALYZED BY ALUMINAS

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Temperature, °C</th>
</tr>
</thead>
<tbody>
<tr>
<td>o-H₂ + p-H₂</td>
<td>-195</td>
</tr>
<tr>
<td>H₂ + D₂ → 2 HD</td>
<td>-125</td>
</tr>
<tr>
<td>CH₄/CD₄ isotopic scrambling</td>
<td>25</td>
</tr>
<tr>
<td>Alkene + D₂ → alkene-d + HD</td>
<td>25</td>
</tr>
<tr>
<td>Benzene + D₂ → benzene-d + HD</td>
<td>25</td>
</tr>
<tr>
<td>Double-bond isomerization of alkenes</td>
<td>25</td>
</tr>
<tr>
<td>Cis/trans isomerization of alkenes</td>
<td>25</td>
</tr>
<tr>
<td>Cyclopropane → propene</td>
<td>100</td>
</tr>
<tr>
<td>Alcohols → alkenes + H₂O</td>
<td>75</td>
</tr>
<tr>
<td>2 alcohols → ether + H₂O</td>
<td>125</td>
</tr>
<tr>
<td>Skeletal isomerization of alkenes</td>
<td>325</td>
</tr>
<tr>
<td>o-Xylene isomerization</td>
<td>500</td>
</tr>
</tbody>
</table>

In addition to those mentioned above many other reactions, often involving complex organic compounds, have been observed on alumina during chromatography. This led to the use of chromatographic alumina to cause various adsorbed organic molecules to undergo many different types of often unanticipated chemical reactions. Interest has been rekindled in this area by recent developments involving deliberate placement of different
reagents on solid alumina and use of this doped alumina to cause diverse organic reactions heterogeneously at the alumina surface under unusually mild conditions. A few of the more interesting examples are discussed below.

(i) The reaction of (+)-2α,3α-epoxypinane I on exposure to active alumina in hexane gave three main products II, III and IV in the ratio 1:4:2 besides traces (½%) of trans glycol. The alcohol IV resulted from Cannizzaro type reaction of aldehyde V on alumina, the aldehyde itself being first formed from the oxide via a carbonium ion rearrangement. The occurrence of a Cannizzaro reaction was confirmed by isolation of the Cannizzaro acid VI as the methyl ester from the spent alumina.

(ii) Posner et al. have shown that a stirred slurry of commercially available Woelm 200 neutral chromatographic alumina catalyzes opening of a very wide variety of epoxides by only a few equivalents of RZ-H nucleophiles reproducibly and under exceedingly mild conditions (10 min to 1 h 25°C, diethyl ether solvent). Nucleophiles successfully incorporated under these conditions include alcohols, thiols, benzene selenol, amines and acetic acid.
Alumina impregnated with a few equivalents of these nucleophiles opened epoxides regioselectively at the less substituted epoxide carbon atom and stereospecifically (trans) to give the corresponding β-functionalized alcohol cleanly and in good yield.

For example cyclohexene oxide reacted with alumina carrying 4% of RZ-H to give the corresponding trans-2-functionalized cyclohexanols reproducibly in good yield as the only product (Eq. 1). No allyl alcohol, 1,2-diol or cis-isomer was detected. In several cases cyclohexene oxide reactions with RZ-H doped alumina were superior to homogeneous methods for synthesis of the corresponding 2-RZ-cyclohexanols.

<table>
<thead>
<tr>
<th>ZR</th>
<th>Yield (%)</th>
</tr>
</thead>
<tbody>
<tr>
<td>OMe</td>
<td>66</td>
</tr>
<tr>
<td>OCH$_2$Ph</td>
<td>47</td>
</tr>
<tr>
<td>SEt</td>
<td>78</td>
</tr>
<tr>
<td>SPh</td>
<td>70</td>
</tr>
<tr>
<td>SePh</td>
<td>95</td>
</tr>
<tr>
<td>NH-n-Bu</td>
<td>73</td>
</tr>
<tr>
<td>N</td>
<td>40</td>
</tr>
</tbody>
</table>

In addition to the reactions mentioned above, alumina has been used for intramolecular addition of OH groups, intra- and intermolecular addition of CH groups, in oxidation-reduction reactions, substitution reactions, elimination reactions, decarboxylation reactions and skeletal rearrangements.

In spite of the wide variety of organic reactions catalyzed by alumina in solution, there have been little or no studies of the kinetics and mechanism of alumina catalyzed organic reactions in solution. Most of
the mechanistic studies of organic reactions on alumina surface are confined to gas phase reactions of alcohols and alkenes. The only kinetic study of a substrate in solution reacting on an alumina surface is by Leffler and Miller who investigated the reaction of diacyl peroxides with chromatographic alumina. They observed that several diacyl peroxides as well as perbenzoic acid and hydrogen peroxide reacted rapidly with alumina surfaces to give a nonextractable oxidant of equivalent oxidizing power.

Surface hydroperoxides \( \text{Al}_s\text{OOH} \) were suggested as the titrable oxidant groups formed on alumina because the decay reaction released molecular oxygen. First order plots for decomposition of the surface oxidant were curved with high initial rates which decreased rapidly both within a run and as a function of concentration of the initial surface oxidant. They explained these results in terms of surface sites with different properties; the sites were divided into categories I and II. According to the model, Type I sites reacted with diacyl peroxide more rapidly than the Type II sites (Eqs. 2 and 3). However, the oxidant at Type I sites was considered to decompose more slowly than the oxidant at the Type II sites (Eqs. 4 and 5) giving rise to the observed kinetics.

\[
\begin{align*}
\text{R-C-O-C-R + Al}_2\text{O}_3 (I) & \xrightarrow{\text{fast}} 2\text{RCOOH} + \text{Al}_s\text{OOH}(I) \\
\text{R-C-O-C-R + Al}_2\text{O}_3 (II) & \xrightarrow{\text{slow}} 2\text{RCOOH} + \text{Al}_s\text{OOH} (II)
\end{align*}
\]

\[
\begin{align*}
\text{Al}_s\text{OOH}(I) & \xrightarrow{\text{slow}} \text{Al}_s\text{OH} + \frac{1}{2}\text{O}_2 \\
\text{Al}_s\text{OOH} (II) & \xrightarrow{\text{fast}} \text{Al}_s\text{OH} + \frac{1}{2}\text{O}_2
\end{align*}
\]

In order to obtain a more complete study of the kinetics and
mechanism of an alumina catalyzed organic reaction the study of glucose mutarotation by aluminum oxide was systematically undertaken for this thesis. Further, the study of glucose mutarotation by alumina should provide information on the nature of active sites on the surface and hence, might find uses as a method for characterizing alumina catalysts.

To gain an insight into the action of alumina as a catalyst it is important to understand both the physical and chemical properties of alumina, specially of its surface. The following sections will review the present knowledge of the preparation, structure and nature of active sites on alumina.

I.1.1 Preparation and Crystalline Structure of Alumina

Aluminas are usually obtained by dehydration of aluminum trihydroxides (gibbsite, bayerite and nordstrandite) obtained by precipitation from aqueous solutions containing aluminum ions. Depending on the method of aging the gelatinous hydroxide and the pyrolysis temperature it can have several crystalline forms (Fig. I).

The method of formation of the hydroxide generally determines the impurities present in the resultant alumina and hence will affect its catalytic properties. Aluminum hydroxide obtained by hydrolysis of aluminum isopropoxide which has been distilled in vacuum will produce alumina virtually free of ionic impurities. But the hydroxide produced by adding ammonium hydroxide to a solution of an ammonium salt contains entrapped anions. On the other hand, alumina prepared from gibbsite or from potassium or sodium aluminate contains alkali in the amount of 0.08 to 0.65%.

During dehydration of the trihydroxide adjacent hydroxyl groups
Fig. 1
Schematic Representation of Formation of Various $\text{Al}_2\text{O}_3$-hydrates$^{8,17}$. (Temperature in °C).
combine forming water vapour which gives rise to small pores within the crystal lattice. As dehydration proceeds the pores from different parts of the lattice join forming a network and eventually the water vapour finds its way out of the lattice connecting the internal surface with the external surface. Thus a porous sample of alumina with high surface area per unit weight is produced. Since a heterogeneous catalytic reaction proceeds on the surface of the catalyst, the presence of a high surface area increases its efficiency and the presence of active sites in pores can also lead to a 'solvating effect'\textsuperscript{14} of alumina and reactions normally not possible on a surface may proceed within it.

I.1.2 Classification of Aluminas

Aluminas can be classified according to the temperature at which they were obtained from the hydroxide or based upon the crystallographic structure of the alumina\textsuperscript{17}.

Aluminas are divided into two groups according to the temperature at which they were formed.

(a) Low temperature aluminas or the \( \gamma \)-group. \( \text{Al}_2\text{O}_3\cdot n\text{H}_2\text{O} \) in which \( 0<n<0.6 \); obtained by dehydration at temperatures not exceeding 600°C.

(b) High temperature aluminas, called the \( \delta \)-group. These are nearly anhydrous aluminas obtained at temperatures between 900°C and 1000°C.

To group (a) belong \( \omega \)-, \( \chi \)-, \( \eta \)- and \( \gamma \)-aluminas and to group (b) belong \( \kappa \)-, \( \theta \)- and \( \delta \)-aluminas.

A classification based on the crystallographic structure of alumina was proposed by Krischner in 1966\textsuperscript{17}. As these structures are all based on a more or less close-packed oxygen lattice with aluminum ions in the octahedral and tetrahedral interstices, three series could be distinguished viz:
α-series with hexagonal close-packed lattice, schematically ABAB...

β-series with alternating close-packed lattice, schematically ABAC-ABAC or ABAC-CABA

γ-series with cubic close-packed lattice, schematically ABCABC

The only representative of the α-series is α-alumina. The β-series consists of alkali or alkaline earth oxide containing β-alumina and, χ- and κ-alumina. The γ-series can be sub-divided to a γ- or low-temperature group (consists of η- and γ-alumina) and a δ- or high-temperature group (δ- and θ-alumina).

Out of these different crystalline forms of alumina only γ- and η-phases are important catalytically. The catalytic activity of η-alumina usually turns out to be higher than that of γ-alumina. The other main form of alumina, viz. α-alumina is considered to be most inert of all aluminum oxides, and is used mainly as a catalyst support (inert carrier) because of its high temperature stability.

As mentioned above γ- and η-aluminas consist of cubic close-packed structure of oxide ions. Lippens who did an extensive X-ray crystallographic study of aluminas proposed models for γ- and η-aluminas. The most important structural characteristic of alumina in catalysis is the surface and since alumina occurs in the form of lamellae it is most probable that only one type of surface plane is predominant. According to Lippens this is the (111)-plane for η-alumina and the (110)- or (100)-plane for γ-alumina. In practice the (111)- and (110)-planes are considered to form the surface layers of η- and γ-aluminas, respectively.

It has been shown that the (111)-plane of η-alumina and (110)-plane of γ-alumina have aluminum ions arranged in both tetrahedral and octahedral positions. But there is a higher density of aluminum ions in tetrahedral
positions in η-alumina and this is considered to give rise to the higher acidity and catalytic activity observed with η-alumina \(^8,23\).

Thus both γ- and η-aluminas consist of cubic close-packed oxygen lattice with the aluminum ions distributed in octahedral and tetrahedral sites and some cation sites are left vacant for stoichiometric reasons. Therefore they are said to have 'defect' spinel structure after the mineral spinel (MgAl\(_2\)O\(_4\)) which has a similar structure \(^8\).

As mentioned above α-alumina consists of a hexagonal close-packed oxygen lattice and, unlike η- and γ-forms, all the Al\(^{3+}\) ions are located in octahedral sites with one in every three cation sites vacant \(^8,17\). The presence of Al\(^{3+}\) ions only in octahedral sites probably leads to the low catalytic activity observed in α-alumina, since the Al\(^{3+}\) ions in octahedral sites are not as acidic as those in tetrahedral sites because of the higher coordination number.

I.1.3 The Surface Structure of Alumina

The chemical nature of the alumina surface is of primary importance in its catalytic and adsorptive properties. So-called 'active-alumina' \(^17\), which is alumina used as adsorbents and catalysts, is not pure alumina but contains, depending upon temperature and water vapour pressure, from few tenths to about 5% water. Depending on preparative conditions, other components may be present too, e.g. alkali oxide, iron oxide and sulphate \(^17\). The presence of even minute amounts of sodium oxide was found to decrease the catalytic effect of alumina on the dehydration of propanol and butanol \(^17\). Dehydration of cyclohexanol over pure alumina prepared from aluminum isopropoxide produced cyclohexene and upto 60% methylcyclopentenes. But over alumina containing about 0.4% of sodium or potassium ions, cyclohexene was
the only product owing to the absence of strong acid sites which were neutralized by the alkali metal ions. The presence of sulphate or other anions in general is considered to increase the 'acidic nature' of alumina. It is also known to affect the rate of dehydration and the rate of sintering of the catalyst.

Active alumina adsorbs water either as hydroxyl ions or as water molecules on the surface, depending upon the temperature. When exposed to water vapour at about room temperature γ-alumina adsorbs water as undissociated molecules bonded with strong hydrogen bonds to the underlying surface. On γ-alumina which is catalytically more active, Borello et al. have shown that in addition to molecular adsorption of water, hydroxyl ion formation on the alumina surface occurs by dissociative adsorption of water. At higher water vapour pressures more water is adsorbed in a multilayer physical adsorption but this water can be removed easily at about 120°C. The strongly adsorbed water which cannot be removed at 120°C was defined as 'chemisorbed' water by de Boer et al. It was found to be the same (25 mg/100 m²) independent of the temperature of dehydration of alumina. Peri and Hannan presented infra-red spectroscopic evidence for the occurrence of both hydroxyl groups and undissociated water molecules on the surface of γ-alumina at low temperatures. During the process of drying by heating water molecules, not desorbed and removed from the surface, react to form surface hydroxyl groups. Hence, initially, there is a decrease in the number of water molecules and an increase in the concentration of hydroxyl groups on the surface. All water molecules are removed after evacuation at 400°C. At higher temperatures the hydroxyl groups are gradually expelled as water but even at 800 to 1000°C and vacuum some tenth of a percent of water is still retained in the alumina. γ-alumina heated to 650-700°C
showed three major bands of hydroxyl stretching frequencies at 3698, 3737, and 3795 cm\(^{-1}\) due to 'isolated' (not hydrogen bonded) hydroxyl groups. Two more bands were seen under high resolution at 3780 cm\(^{-1}\) and 3733 cm\(^{-1}\) in samples that were well dried\(^{28}\). For all alumina modifications so far investigated a total of five bands have been observed\(^{23}\). This indicates the presence of isolated hydroxyl groups in five different environments on the alumina surface. The lowest frequency band was found to be the most acidic because it exchanged most easily with deuterium between 250° and 500°C, and with butene at 200°C. In γ-alumina the hydroxyl band at 3737 cm\(^{-1}\) was removed more rapidly on drying while with η-alumina the bands at 3780 and 3700 were removed more readily\(^{23,26}\), but all three major bands clearly remained after drying at 850°C. These results show that the hydroxyl groups are present in different chemical environments (these may play different roles in catalytic reactions). The adsorption of large molecules (e.g. CCl\(_4\)) generally reduced the peak intensity of the hydroxyl bands and broadened and shifted them to lower frequencies. The hydroxyl groups must therefore be on the surface, rather than within the alumina lattice. The number of hydroxyl groups on the surface was also determined\(^{25}\) by measuring the number of hydrogen atoms that could be exchanged with deuterium gas. The results showed that the surface is 40% covered with hydroxyl groups after drying at 400°C, 15% at 600°C, 2% at 800°C and 1% or less above 900°C. Hydrogen bonding appeared to exist on alumina dried below 600°C.

Adsorption of water vapour at room temperature on γ-alumina previously dried at 800°C produced absorption bands similar to those of the stretching and bending vibrations of liquid water. The hydroxyl band at 3795 cm\(^{-1}\) was replaced by a band near 3500 cm\(^{-1}\) but most of the isolated hydroxyl groups apparently were not perturbed by the adsorbed water.
1.1.3.1 **A Model for the Surface of Alumina**

Using his infrared and gravimetric data as a guide, Peri in 1965 proposed a model for the surface of γ-alumina where he considered only the (100)-plane on the surface. This model was later improved by Knözinger and Ratnasamy who considered all three low index planes (100), (110) and (111) of alumina. For energetic reasons only anion layers will terminate a crystallite and it was shown with the aid of Pauling's electrostatic valence rule (which states that the net charge in a stable ionic structure should be equal or nearly equal to zero) that these surface layers will most favourably consist of OH groups (Fig. 2).

The variety of different surface hydroxyl groups will be briefly considered in the next section.

I.1.3.1.1 **Surface Hydroxyl Groups**

As indicated above Knözinger and Ratnasamy have shown that on the surface of γ- or η-alumina there are five types of hydroxyl groups; Type Ia, a terminal OH group coordinated to a single tetrahedral $\text{Al}^{3+}$ ion:

Type Ia

$$\text{OH} \quad \text{Al}$$

Type IIa, a bridging OH group which links a tetrahedral and an octahedral cation:

Type IIa

$$\text{H} \quad \text{O} \quad \text{Al} \quad \text{Al}$$
Type IIb where the OH group links to two cations in octahedral positions:

```
    H
   /\ |
  /  \ Al
 /    |
|     |
O
```

Type III where it is coordinated to three cations in octahedral interstices:

```
    H
   /\ |
  /  \ Al
 /    |
|     |
O
```

and Type Ib where the OH group is coordinated to a single cation in an octahedral interstice

```
    OH
   /   |
  /    |
|     |
Al
```

Table II summarizes the five possible OH configurations, the coordination number of surface anions and also the net charge on the oxygen when it exists as an oxide ion and when it is protonated to form a hydroxyl group. These values have been obtained as the sum of the negative charge of the anion and sum of the strengths of the electrostatic bonds (= cation charge divided by coordination number) to the anion from adjacent cations. Knözinger and Ratnasamy making use of the net charge on OH groups assigned the stretching frequencies given in the last column of Table II to the five hydroxyl groups. They correspond closely to the five stretching bands observed with alumina as mentioned earlier. The band of highest wavenumber (3800 cm$^{-1}$) was assigned to the configuration Ib which bears the most negative charge (-0.5) while the band of lowest wavenumber (3700 cm$^{-1}$)
### TABLE II

POSSIBLE OH CONFIGURATIONS ON ALUMINA SURFACE

<table>
<thead>
<tr>
<th>Configuration</th>
<th>Coordination Numbers of Surface Anion to Al(VI); to Al(IV)</th>
<th>Net Charge at 0</th>
<th>Net Charge at OH</th>
<th>v(OH) cm(^{-1})</th>
</tr>
</thead>
<tbody>
<tr>
<td>III</td>
<td>3 - -0.5 +0.5</td>
<td>3700 - 3710</td>
<td></td>
<td></td>
</tr>
<tr>
<td>IIb</td>
<td>2 - -1.0 0</td>
<td>3740 - 3745</td>
<td></td>
<td></td>
</tr>
<tr>
<td>IIa</td>
<td>1 1 -0.75 +0.25</td>
<td>3730 - 3735</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ia</td>
<td>- 1 -1.25 -0.25</td>
<td>3760 - 3780</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ib</td>
<td>1 - -1.5 -0.5</td>
<td>3785 - 3800</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

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\(^a\) Large circle: oxide ion or hydroxyl group; • : Al(VI); ○ : Al(IV)
was attributed to configuration III which exhibits the most positive charge (+0.5). The remaining three bands were assigned with decreasing wavenumber to the corresponding configurations with increasing positive net charge.

In addition the net charge on the hydroxyl group should determine the relative acidities and basicities of the hydroxyl groups. One would expect the OH configuration of Type III with a net positive of +0.5 to be the most acidic. The protonic acidity of the OH groups should decrease as the net charge on them becomes more negative and their basicity should increase at the same time. The ease of removal of the OH groups should parallel their basicity since the remaining net positive charge at the anion vacancy is lower the higher the net negative charge of the leaving OH group. The exceptional lability of the Ia and Ib OH groups is reflected in their ease of exchange with C$^{18}$O$_2$, even at room temperature$^{23}$. The oxygens of other configurations were exchanged much more slowly.

I.1.3.1.2 Surface Dehydration

According to Knözinger and Ratnasamy$^{23}$ proton acidity and the ease of removal of OH groups should govern the dehydration process, at least in the initial stages at low temperatures. Thus a proton from an acidic OH group (e.g. Type IIa) would combine with a neighbouring basic hydroxyl group (e.g. Type Ia) to form a water molecule. Therefore, the charge defects created are as small as possible and initial dehydration leads to the formation of weak Lewis acid sites and neighboring weak basic sites. In fact it has been observed that on dehydration intensity of the

\[
\begin{align*}
\text{Al} & \quad \text{O} \quad \text{H} \\
\text{Type IIa} & \quad \text{Al} & \quad \text{O} \quad \text{H} \\
\text{Type Ia} & \quad \text{Al} & \quad \text{Al} & \quad \text{Al} & + \quad \text{H}_2\text{O}
\end{align*}
\]
infrared bands due to acidic and basic hydroxyl groups decrease faster than those of the central bands. According to this model dehydration would proceed without the formation of high energy multiple vacancies (adjacent \( \text{Al}^{3+} \) ions) and clusters of oxide ions. They showed that it can proceed until about 50 to 65% dehydroxylation has occurred, depending on the crystal planes involved. By studies of dehydration of different alumina samples it has been shown that this degree of dehydroxylation occurs at temperatures between 300° and 400°C (Fig 2).

I.1.4 Nature of Active Sites

It is important to determine whether these weak individual Lewis acid and basic sites, formed during the initial dehydroxylation, are the catalytically active sites. It has been very clearly demonstrated by many authors that alumina should be pretreated at elevated temperatures for the development of catalytic activity in most reactions. For example, alumina had to be pretreated at temperatures of roughly 300° to 400°C in vacuum for double bond isomerization and skeletal isomerization of 1-pentene. van Cauwelaert and Hall have shown that catalysis of ortho-para hydrogen conversion occurs on alumina pyrolyzed above 300°C and that the rate constant increased by a factor of about 10 corresponding to a steep decrease in surface hydroxyl concentration. These results clearly demonstrate that the removal of water and/or OH groups from alumina surface is essential for the development of the catalytic activity.

Although catalytic activity develops on activation at temperatures above 300° to 400°C, it has been shown that \( 3.7 \times 10^{14} \text{ cm}^{-2} \) anion vacancies and coordinatively unsaturated (cus) oxygen atoms have already
been formed at 300°C. Therefore, Knözinger and Ratnasamy concluded that Lewis acid and basic sites produced during regular dehydroxylation can hardly be involved in most catalytic reactions as catalytic sites. The same conclusion was reached by comparing the number of Lewis acid and basic sites formed during dehydration (≈10^{14} \text{ cm}^{-2}) with the number of catalytically active sites determined by specific poisoning experiments (10^{12} to 10^{13} \text{ cm}^{-2}). For comparison, the lattice site densities in ideal planes are of the order of 10^{15} \text{ cm}^{-2}.

Hence, Knözinger and Ratnasamy assumed that at temperatures between 300° to 400°C special site configurations of low probability are beginning to develop which possess structural and energetic properties required for an active site. These special configurations have been identified with defects in the partially dehydroxylated surface; that is with multiple vacancies and clusters of oxygen atoms in certain environments on the surface. They showed that such defects should be formed on further dehydration above 300° to 400°C, due to condensation of equivalent OH groups with the formation of triplet vacancies and neighboring oxide ions (Fig. 2). It has also been postulated\textsuperscript{23,31} that the mobility of ions as the temperature is increased increases the chances of forming defects but this mobility can also redistribute the high surface energy of defect sites. These two opposing effects can lead to an optimum activation temperature of alumina, for certain reactions (e.g. isomerization of alkenes\textsuperscript{32}).

I.1.4.1 Different Types of Active Sites

From the above discussion it is clear that the alumina surface should consist of both acidic and basic sites of different strengths.
Completely hydroxylated (III)-face

Hydroxyl group on top layer

Oxide ion on lower layer

Aluminum ion

300°-400°C

(III)-Face at 50% dehydroxylation

Oxide ion on top layer

> 400°C

Triplet vacancy in (III)-face

Neighboring oxide ions

Triplet vacancy

Fig. 2 (III)-Face of Alumina at Different Stages of Dehydroxylation
The acidic sites are the (a) Brönsted Acid sites, for example Type III OH groups, and (b) Simple Lewis Acid sites (individual Al$^{3+}$ ions), both present at relatively low temperatures, and (c) Triple vacancies (three neighboring Al$^{3+}$ ions) present at temperatures greater than 300° to 400°C.

The basic sites are the (d) Basic hydroxyl groups, for example Type Ib hydroxyl groups, (e) Individual oxide ions, both of which are again present at relatively low temperatures, and (f) Clusters of oxide ions formed at temperatures greater than 300° to 400°C.

As discussed later (Section 1.2.3) in this Introduction, mutarotation of glucose in homogeneous solution is catalyzed by Brönsted acids, bases such as hydroxide ions and carboxylate ions and even Lewis acids such as Al$^{3+}$, Zn$^{2+}$, Cu$^{2+}$, and Ni$^{2+}$. Hence all these acidic and basic sites are potential catalytic sites for the mutarotation reaction.

Evidence for the presence of those active sites has come from adsorption studies of acid or base sensitive substances$^{24}$. The use of such substances as specific poisons of catalytic reactions has shown that these sites actually participate in the surface reactions$^{33}$. Since the same methods can be used to determine the active sites for glucose mutarotation they will be described in the next section.

I.1.4.2 Evidence for the Presence of Active Sites

I.1.4.2.1 Acid Sites$^{24}$

The presence of acid sites on the alumina surface has been shown by titration of the solid acid suspended in dry benzene with n-butylamine. The use of various indicators with different $pK_a$ values to detect the end
point has enabled the determination of the amount of acid at various acid strengths. The acid strengths ($H_0$) as measured by amine titration ranged from +3.3 to -5.6 for a sample of pure alumina prepared from aluminum isopropoxide and pyrolyzed between 300° and 1000°C. The amount of acid at various strengths was a function of pyrolysis temperature. It increased on heating above ≈200°C but decreased at high temperatures (≈1000°C).

The amount of acid sites has also been determined by titration with potassium hydroxide, dioxane and also by chemisorption of ammonia, trimethylamine and pyridine. Each of these methods gave acidity values for alumina which apparently approximate those of silica-alumina.

Pines and Haag determined the number of acid sites on pure alumina prepared from aluminum isopropoxide and activated at different temperatures, by chemisorption of trimethylamine at 300°C. There were $1.6 \times 10^{13}$ sites/cm$^2$ when pyrolyzed at 400°C, $2 \times 10^{13}$ sites/cm$^2$ at 600°C, $1.4 \times 10^{13}$ sites/cm$^2$ at 700°C and $0.3 \times 10^{13}$ sites/cm$^2$ at 900°C, assuming that each adsorbed molecule corresponds to one acid site. These values are in good agreement with those obtained by n-butylamine titration ($1.1$ to $2.5 \times 10^{13}$ sites/cm$^2$).

(i) **The Nature of Acid Sites**

The methods described above do not distinguish between Lewis acid sites and Brønsted acid sites. The nature of the acid sites has been intensively investigated by different techniques during the last two decades. Hammett indicators ($pK_a$ +6.8 to -8.2) which produced acid colour with synthetic cracking catalysts and natural clays gave no acid colour with alumina. To test whether protonic acidity is developed only at higher temperatures, the colour test was performed in refluxing xylene (b. pt. 144°C) using neutral red ($pK_a$ +6.8). However, no acid colour was
produced. Hence, Pines and Haag concluded that Brønsted acid sites on alumina, if present at all, are of very low acid strength. The same conclusion was reached from the failure of alumina to undergo cation exchange with ammonium acetate.

However, when they used a set of indicators (triphenylmethane derivatives) which give colour with Lewis acids (by hydride ion abstraction) but not with Brønsted acids, alumina pyrolyzed at 700°C showed the presence of Lewis acid sites. Exposure of the catalyst to the humidity of the atmosphere before testing inhibited the development of colour in the alumina. The active centres apparently were poisoned by the strong adsorption of water which is removed by heating at high temperatures.

More useful information on the nature of the acid sites on alumina has come from the use of more sensitive instrumental methods. For example, infrared and nuclear magnetic resonance spectroscopy have been used to determine the nature of the chemisorbed bases on the alumina surface, as described below.

(a) Ammonia

Ammonia is a strong Lewis base \(K_b = 1.8 \times 10^{-5}\) in water at 25°C and it is small in size. By infrared spectra of adsorbed species on alumina four surface species have been identified, beside weakly held ammonia which was simply hydrogen bonded. The dominant species was coordinately held ammonia that adsorbs on Lewis acid sites. A limiting form of this species viz. \(+\text{NH}_3\), has been observed on strongly dehydrated (>500°C) alumina. It might represent ammonia associated with strongly positive charged triple vacant sites. At high pretreatment temperatures (>500°C) and low hydroxyl densities, \(\text{-NH}_2\) groups are formed probably due to dissociative chemisorption.
on acid-base pair sites. On alumina containing high OH concentrations (pretreatment <400°C) protonated ammonia (NH$_4^+$) has been observed due to weak Brønsted acid sites.

(b) **Pyridine**

Pyridine is less basic (K$_b$ = 2.3 x 10$^{-9}$ in water at 25°C) than ammonia but it is still a fairly hard base. The infrared spectra of adsorbed pyridine have shown the presence of Lewis acid sites on alumina. But, according to most authors, no intrinsic Brønsted sites could be found by pyridine adsorption since pyridinium ion (PyH$^+$) was not formed at a detectable level due to lower basicity of pyridine. Even when the spectra were recorded at temperatures up to 300°C, the PyH$^+$ species could not be detected indicating that the protonic acidity of alumina was not apparently increased in this temperature range. Only Bremer and co-workers claim to detect PyH$^+$ species on η-alumina.

However, Dewing et al. detected protonated species due to weak Brønsted acid sites, in addition to coordinated species, by adsorption of 2,6-ditertiary-butyl pyridine on γ-alumina. More recently, Pearson used wide line nuclear magnetic resonance to study the nature of deuterated pyridine adsorbed on alumina at 0°C. He detected the protonated species on alumina (2.6 x 10$^{13}$ sites cm$^{-2}$ on alumina activated at 600°C) probably because of the greater sensitivity of the n.m.r. method.

At temperatures greater than 350°C pyridine reacted with surface OH groups forming surface pyridone species with the production of hydrogen. This indicates the presence of strongly basic -OH ions held to certain sites on the alumina surface, their number being of the order of magnitude 10$^{13}$ cm$^{-2}$. Additional evidence for the existence of these reactive and
strongly basic −OH ions has come from adsorption of CO₂, nitriles and ketones on alumina.

(c) Butylamine

The infrared spectra of n-butylamine (Kₜ = 4.8 x 10⁻⁴ in water at 25°C) adsorbed at room temperature on alumina pretreated at 500°C showed the presence of Lewis acid sites. But no Brönsted sites were observed even on heating to 500°C. Adsorption of pyridine on the same sample showed the presence of Lewis acid sites but no Brönsted sites. When n-butylamine was adsorbed at room temperature on alumina pretreated at 100°C both Brönsted and Lewis acid sites were observed. When pyridine was adsorbed on the same sample only Lewis acid sites were observed due to the lower basicity of pyridine.

The above discussion shows that Lewis acid sites are the predominant acid sites on alumina, and also that there are weak Brönsted acid sites on the strongly hydroxylated (Note: according to the model discussed in Section I.1.3.1. acidic protons are removed during the initial dehydration) surface. Formation of +NH₃ and −NH₂ probably gives evidence for the presence of defect sites on strongly dehydrated alumina and the presence of strongly basic hydroxide ions are indicated by the formation of pyridone on the surface.
1.1.4.2.2 Basic Sites

The basic sites on alumina have been investigated using carbon
dioxide, acetic acid and tetracyanoethylene.\(^{33}\)

(a) Carbon Dioxide

Carbon dioxide is a fairly small molecule with acidic properties and
has frequently been used as a probe molecule for basic surface sites and
as a poison in catalytic reactions. After heat treatment at roughly
below 500°C, the alumina surface is still strongly hydroxylated and carbon
dioxide adsorption leads to the formation of a surface bicarbonate ion
predominantly.\(^{33}\) On formation of this species, \(\text{CO}_2\) selectively reacts with
the highest-frequency (3800 cm\(^{-1}\)) OH groups of the alumina surface. It is
assumed that the bicarbonate ion forms on an Al-OH pair site which was called
"X-site" by Fink.\(^{40}\) The number of these sites varied between 1.2 and 1.8 \(\times\)
\(10^{13}/\text{cm}^2\) and may be identical to those that convert pyridine to pyridone.

\[
\text{Al} \quad \text{OH} \quad \text{CO}_2 \quad \text{CO}_2
\]

A free carbonate ion has also been observed on the surface.

On more extensively dehydroxylated alumina surfaces, bidentate (on
acid-base pair sites, possibly \(\alpha\)-sites \(\text{Al}_2\)) and unidentate carbonate
groups have been observed. Gregg and Ramsay\(^{41}\) showed that on \(\kappa\)-alumina
surface heat treated at 1000°C, only 1 in 10 oxide ions is reactive. Hence
it appears that only a small percentage of the surface oxide ions are
located in suitable environments for carbonate formation. Further, it has
also been suggested that carbon dioxide gets adsorbed on Lewis acid sites
on dehydroxylated alumina surfaces\textsuperscript{33}.

(b) **Tetracyanoethylene (TCNE)**

Tetracyanoethylene has been used to detect donor sites on oxide surfaces. Electronic and ESR spectra of the adsorbed acceptor molecules are characteristic of the surface anion radicals which are assumed to be formed according to Eq. (6)\textsuperscript{33}. The basic $^-$OH ions on a hydroxyl rich surface and the oxide ions on a strongly dehydroxylated surface can act as donor sites. Hence as the surface is dehydroxylated the spin concentration of the anion radical passes through two maxima: the first is located between $400^\circ$ and $500^\circ$C ($^-$OH donor sites) and the second (brought about by the oxide ions) is between $600^\circ$ and $700^\circ$C.

These results indicate that there are two types of basic sites on the alumina surface. The predominant species is determined by the temperature of dehydration of the catalyst. Thus the basic hydroxide ions predominate at temperatures less than $500^\circ$C (Note: according to the model discussed in Section I.1.3.1 basic OH groups are removed during initial dehydration), while the reactive oxide ions are more abundant at higher dehydration temperatures.

I.1.4.2.3 **Evidence for the Presence of Defect Sites**

The existence of defect sites was investigated by Della Getta
et al. using carbon monoxide as a surface probe. Carbon monoxide is a rather soft base and therefore should throw light on the presence of strong Lewis acid sites on the surface.

When η-alumina was dehydroxylated below 400°C, no CO adsorption was detected at 36°C. Dehydration at 400°C led to weak reversible adsorption of CO. A second CO species, which is strongly held at 36°C, first appeared on heat treatment around 500°C. Both species were described as CO molecules linked coordinatively to cus surface cations through σ-dative bonds. The site density for the strongly adsorbed species was determined to be $2 \times 10^{13}$ cm$^{-2}$ and it was shown that this species was attached to X-sites on the surface.

The CO adsorption on γ-alumina leads also to the formation of two species, a more strongly held form and a less energetic form. The characteristic difference between the γ- and η-forms is the finding that the number of strong sites was only about $6 \times 10^{12}$ cm$^{-2}$ on γ-alumina. Even the number of weakly adsorbing sites was much lower than on η-alumina. This lower density of defect sites is probably the main reason for lower catalytic activity of γ-alumina.

I.1.5 Evidence for the Participation of these Acidic and Basic Sites in Catalytic Reactions

In the above discussion the presence of different types of acidic and basic sites on the alumina surface activated at different temperatures has been shown using specific surface probes. The mutarotation of glucose is catalyzed by both acids and bases in solution. Hence all these sites are potential catalytic sites for the mutarotation reaction.

Some of the evidence concerning active sites available from various
reactions previously studied on alumina will be briefly discussed below. This discussion will point out the advantages of using the mutarotation reaction as a probe for surface active sites.

I.1.5.1 Hydrogen-Deuterium Exchange

As already mentioned alumina catalyzes the $\text{H}_2$-$\text{D}_2$ equilibration and also $\alpha$-$\beta$-$\gamma$ conversion at low temperatures ($\approx -75^\circ\text{C}$ for $\text{H}_2$-$\text{D}_2$ equilibration) provided the oxides are dehydroxylated at temperatures higher than $325^\circ\text{C}$. This suggests that defect sites of low probability must be involved in these reactions.

It was shown that the hydrogen species which led to exchange with $\text{D}_2$ was a species chemisorbed atomically onto the surface with a saturation coverage at $-75^\circ\text{C}$ of $9.7 \times 10^{12}$ hydrogen atom cm$^{-2}$. The sites responsible for this type of hydrogen chemisorption should exhibit a very strong gradient of electric field strength. Thus the most probable catalytic sites are the defect sites. The necessity of oxide ions for these reactions was shown by the inhibitory effect of carbon dioxide on these reactions.

I.1.5.2 Isomerization and Exchange Reactions of Hydrocarbons

Aluminas can also activate C-H bonds in saturated and unsaturated hydrocarbons. They are therefore active catalysts for double bond and cis-trans isomerization reactions and also for exchange reactions such as $\text{D}_2$ exchange with hydrocarbons and deuterium scrambling (e.g. $\text{C}_6\text{H}_6/\text{C}_6\text{D}_6$ or $\text{CH}_4/\text{CD}_4$). The behaviour of aluminas in these reactions has turned out to be extremely complex, and only a few specific examples will be described here.

Olefins and aromatic hydrocarbons such as benzene and toluene undergo
exchange reactions with deuterium gas around room temperature on alumina activated at 500°C and higher temperatures\textsuperscript{23}.

Deuterium exchange of benzene and toluene occurs via heterolytic ring C–H bond cleavage, with loss of H\textsuperscript{+} to the catalyst surface and the formation of a carbanion like species. The exchange with alkyl side chain was much slower than with ring hydrogen\textsuperscript{23}.

With terminal olefins methylenic hydrogen atoms are preferentially exchanged. Therefore alken-1-yl species have been suggested as intermediates for the exchange reactions which must be formed via a dissociative chemisorption step.

Hence it is clear that heterolytic cleavage of unsaturated hydrocarbon C–H bonds occurs on activated alumina. As possible sites one has to assume strong acid-base pair sites which exhibit very high gradients of electric field strength. As expected a new OH group was formed from H\textsuperscript{+} split off during chemisorption of benzene and olefins, giving rise to a new OH stretching band at 3590 cm\textsuperscript{-1}\textsuperscript{23}. Furthermore, the characteristic high-frequency band of the Type Ia OH groups was perturbed by the chemisorbed alkenes or benzene. Precoverage of the alumina surface by alkene or benzene blocked the hydrogen chemisorption. Hence hydrogen (or deuterium) and hydrocarbon molecules are chemisorbed on the same defect X-sites, with the small hydrogen or deuterium molecule being located within the multiple anion vacancies underneath the hydrocarbon molecule.

Carbon dioxide has been used as a poison for identification of active sites of exchange reactions such as CH\textsubscript{4} + D\textsubscript{2}, CH\textsubscript{4} + CD\textsubscript{4}, exchange of olefins with deuterium and benzene with deuterium. These exchange reactions were strongly influenced by the carbon dioxide chemisorption and from the infrared spectra of adsorbed carbon dioxide the active site has
has been identified as a reactive oxide ion. The number of such sites active in the exchange reactions as determined by carbon dioxide poisoning vary between $2.4 \times 10^{12} \text{ cm}^{-2}$ for the exchange of methane with $\text{D}_2$ and about $1.4 \times 10^{13} \text{ cm}^{-2}$ for the exchange of olefins with $\text{D}_2$. Hence, again, low probability site configurations were suggested as the active sites. Since $\text{CO}_2$ still adsorbs to a large extent as a surface bicarbonate species after activation at 530°C, the active site for the exchange reactions was assumed additionally to involve an adjacent hydroxyl group. This type of active site has been schematically pictured as

\[ \begin{array}{c}
\text{H} \\
\text{Al} \\
\text{O} \\
\end{array} \]

Hightower and Hall$^{43,44}$ have shown that with alumina activated at 530°C, deuterium exchange and the intramolecular double-bond isomerization of olefins are independent processes since the sites that catalyze $\text{D}_2$ exchange with olefins are blocked by $\text{CO}_2$ but sites active in double-bond isomerization remain unaffected by $\text{CO}_2$ chemisorption. On the other hand $\text{H}_2\text{S}$ and methyl mercaptan are effective poisons at 25°C for sites that catalyze double-bond migration and cis/trans isomerization of olefins but had no effect on exchange sites. The isomerization sites have a surface density of $5 \times 10^{13} \text{ cm}^{-2}$ and consists of certain exposed Al$^{3+}$ ions$^{45}$. Deuterium exchange sites are much less numerous $(0.3$ to $0.8 \times 10^{13} \text{ cm}^{-2})$ and are associated with a very small fraction of highly energetic surface oxide ions, as mentioned earlier. These studies have led to the postulation of dual nature of sites on alumina surface.

The need of Lewis acid sites for double bond isomerization has been clearly demonstrated by the retarding effect of ammonia, triethylamine and pyridine, on such reactions$^{33}$. 
1.1.5.3 **Dehydration of Alcohols**

The elimination of water from aliphatic alcohols on alumina is known to proceed through two possible routes\(^3\), namely, monomolecular olefin formation,

\[
\begin{align*}
\text{C–C–OH} & \quad \rightarrow \quad \text{C=CH} + \text{H}_2\text{O} \\
\end{align*}
\]

and bimolecular ether formation

\[
\begin{align*}
2 \text{ROH} & \quad \leftrightarrow \quad \text{ROR} + \text{H}_2\text{O} \\
\end{align*}
\]

The active sites in the two reaction routes will be discussed separately.

1.1.5.3.1 **Olefin Formation**

Knözinger\(^4\) and Deo et al.\(^5\) have proposed that olefin formation proceeds through hydrogen bonded adsorbed species. Although Lewis acid sites had earlier been considered as the adsorption sites of alcohol undergoing dehydration to olefins, it was shown that pyridine adsorbed on alumina had no effect on the rate of olefin formation\(^4\). The importance of hydroxyl groups was shown by the fact that dehydration activity of alumina has been shown to exhibit an optimum value for a given pretreatment temperature of alumina. The necessity of basic sites was shown by poisoning effect of TCNE on olefin formation from isopropanol\(^3\). Based on these results Knözinger proposed the \(E_2\) mechanism given below for olefin formation on alumina. That the reaction is by \(E_2\) mechanism was shown by the deuterium isotope effect on dehydration of 2-methyl propan-1-ol, butan-2-ol and t-butanol\(^7,49\). No isotope effects from deuteration of the OH groups of these reactants were found. Very high \(\beta\)-deuterium isotope effects were observed; the greatest effect being exhibited by primary
alcohol (3.44 at 150°C) and the smallest by the tertiary alcohol (2.42 at 150°C). The sequence suggested that the mechanism for dehydration of the primary alcohol possessed the highest E₂ character. A shift towards E₁ character was seen on moving to the secondary and tertiary alcohols, but even so the mechanism of dehydration of the tertiary alcohol was still predominantly of the E₂ type. The observed kinetic isotope effects were lower at higher temperatures, this being associated with an increase in E₁ character of the elimination reactions. It was suggested that this tendency may be due to the increase of Brönsted acidity of the surface hydroxyl groups on alumina at higher temperatures which would favour ionic contributions to the reaction mechanism.

I.1.5.3.2 Ether Formation

It has been shown that only those alcohols that form detectable surface alcoholate species on alumina undergo bimolecular dehydration with ether and water as reaction products. Alcoholate formation is due to the dissociative chemisorption step of the alcohol that occurs on Al–O pair site.
Figuera Roca and co-workers\textsuperscript{33} showed that the rate of ether formation from methanol and ethanol responded very sensitively to poisoning with TCNE. This proves the participation of basic sites in the bimolecular alcohol dehydration. Participation of Lewis acid sites was shown by poisoning with pyridine\textsuperscript{33}. Pyridine influenced the surface concentration of a surface ethanolate as shown by infrared spectroscopy and the rate of diethyl ether formation of the pyridine-poisoned alumina catalyst was found to be directly proportional to the number of alcohoolate groups in the surface. This indicates that only one alcohoolate group participates per reaction step. The second molecule was assumed to be a hydrogen bonded alcohol molecule, since it has been shown\textsuperscript{47}, as in olefin formation, that OH groups are also necessary as active centers for the ether formation.

The above discussion gives clear evidence for the active participation of Lewis acid and base sites in reactions catalyzed by aluminas. The only example of possible Brönsted acid catalysis was in the dehydration of alcohols. The reaction mechanism changed from $E_2$ to $E_1$ when the temperature was increased from 120\degree to 200\degree C, probably because of the increase of Brönsted acidity at higher temperatures. At 120\degree C the proton acidity was too low and t-butanol dehydrated mainly by an $E_2$ mechanism. Direct evidence for the participation of Brönsted acid sites in catalytic reactions was obtained by John et.al.\textsuperscript{50}. They showed that alumina dried at 450\degree or 750\degree C and treated with D\textsubscript{2}O at $>180$\degree C catalyzed double bond migration of propene at 210\degree C by a carbonium ion mechanism involving Brönsted acid sites. No such mechanism occurred at 25\degree C. Hence it appears that although weak Brönsted acid sites had been detected in significant quantities ($2.6 \times 10^{13}$ sites cm$^{-2}$ on alumina activated at 600\degree C) at room
temperature using sensitive surface probe molecules and sensitive techniques, catalytic activity of such sites has not been observed at ordinary temperatures. According to Knözinger the lifetime of a protonated species of a probe molecule may be very low due to the high mobility of surface protons and may contribute to low detectability of the protonated form of the poison. He also stated that protons that can hardly be detected directly by protonated probe molecules may well initiate catalytic reactions due to their polarizing action during their fluctuations. In other words, it may be possible to detect catalytic activity of those weak Brönsted acid sites even at room temperature. According to John et al. "the controversy that has hitherto centred on the presence or otherwise of Brönsted centres on alumina as investigated by spectroscopic techniques, seems to emphasize the need for the use of sensitive catalytic reactions as a direct probe to reaction mechanism, rather than its inference from indirect techniques".

A reaction that has been very extensively studied in homogeneous solution, which is catalyzed by both Brönsted acids and bases, is the mutarotation of glucose. Brönsted and Guggenheim in their studies of acid and base catalysis stated that the reaction to be investigated should be very sensitive to $\text{H}_3\text{O}^+$ and $\text{OH}^-$ in order that the effect of weak acids and bases may be detectable. They selected mutarotation of glucose as the model reaction since it fulfilled those conditions of sensitivity.

I.2 Mutarotation of Glucose

The term mutarotation refers to the change in the optical rotation of a solution to an equilibrium value. The mutarotation of a solution of D-glucose was discovered in 1846 by Dubraunfaut, and since then various
theories were advanced to explain the mutarotation of sugars.

Although glucose was at first considered to be a six carbon straight chain pentahydroxy aldehyde VII, it failed to undergo certain reactions typical of aldehydes. Colley in 1870 and Tollen in 1883 explained the lack of typical aldehyde reactions as arising from an alteration of the aldehyde group by the formation of an inner hemiacetal type linkage. Fischer in 1893 and von Lippmann pointed out that ring formation would produce a new asymmetric carbon atom and thus the existence of α- and β-isomers VIII and IX of D-glucose and its derivatives was clarified. Initially the ring formation was considered to be of 1,4-type (furanose), but in 1926 Haworth showed that it is between 0-5 of glucose and its aldehyde group, i.e. it is a pyranose form.

\[ \begin{align*}
\text{CHO} & \quad \text{H} \\
\text{OH} & \quad \text{H} \\
\text{HO} & \quad \text{OH} \\
\text{H} & \quad \text{HO} \\
\text{CH}_2\text{OH} & \quad \text{OH} \\
\end{align*} \]

\[ \begin{align*}
\alpha-\text{D-Glucopyranose} & \quad \beta-\text{D-Glucopyranose} \\
\end{align*} \]

VII \quad VIII \quad IX

In 1899 Lowry recognized that mutarotations shown by reducing sugars are due to interconversion of ring isomers α- and β-forms to the equilibrium mixture of the two forms.

\[ \alpha-\text{D-glucose} \xrightarrow[k_1]{k_2} \beta-\text{D-glucose} \] (7)

I.2.1 Kinetics of Mutarotation

Urech, in 1883, using first order equation derived by Wilhelmy for
following the inversion of sucrose, showed that the mutarotation of D-glucose obeys first order kinetics reasonably well \(^{51(a)}\).

In 1903, Hudson \(^{51(a), 57}\) showed that the velocity of the first order reversible reaction (Eq. 7) is given by,

\[
\frac{dx}{dt} = k_1 (a - x) - k_2 x
\]  

(8)

where \(a\) is the initial concentration of \(\alpha\)-D-glucose and \(x\) = concentration of \(\beta\)-form at time \(t\). Upon integration Equation (8) becomes,

\[
k_1 + k_2 = \frac{1}{t} \ln \left( \frac{x_e}{x_e - x} \right)
\]

where \(x_e\) = equilibrium concentration of the \(\beta\)-isomer. He showed that this equation can be expressed in terms of optical rotation as,

\[
k_1 + k_2 = \frac{1}{t} \ln \left( \frac{\alpha_o - \alpha_x}{\alpha_t - \alpha_x} \right)
\]  

(9)

where \(\alpha_o\) is the initial rotation, \(\alpha_t\) is the rotation at time \(t\) and \(\alpha_x\) is the equilibrium rotation. Therefore,

\[
\ln (\alpha_t - \alpha_x) = -(k_1 + k_2)t + \ln (\alpha_o - \alpha_x)
\]  

(10)

Hence a plot of \(\ln (\alpha_t - \alpha_x)\) vs time should be a straight line and the kinetics of mutarotation can be easily and conveniently followed by measuring the change in optical rotation with time. Kinetics of mutarotation of glucose has also been measured by nonpolarimetric methods such as changes in refractive index, volume, electrical conductivity, pH, calorimetric methods, solubility methods, nmr and infrared spectra, and gas-liquid
chromatography \(^{51(a), 58, 59, 60}\). However, the measurement of optical rotation is still the simplest and most satisfactory means of studying the kinetics of the reaction \(^{51(a)}\).

It has been shown by many authors \(^{51(a)}\) that the same value of mutarotation constant \((k_1 + k_2)\) is obtained by mutarotation of the \(\alpha\)- and \(\beta\)-anomers. Except for two reports, by Reine et al. \(^{61}\) and Cox and Natarajan \(^{62}\) for mutarotation in dimethylformamide, in all the studies of mutarotation of \(\alpha\)-D-glucose so far, with different solvents, catalysts and at different temperatures, the kinetics has been found to obey the first order equation derived by Hudson. Because of these characteristics the mutarotation of glucose "has long been a proving ground for the testing of kinetic theories".\(^{63}\).

It has been shown that sugars such as D-glucose which obey Equation (9) contains \(\alpha\)- and \(\beta\)-pyranose forms in solution with only a small proportion of other species. Such mutarotations are designated simple mutarotations \(^{51(a)}\), and can also be expressed by Equation (11), the exponential form of Equation (9):

\[
[a]_t = Ae^{-k't} + c \tag{11}
\]

where \(A\) is the difference between the initial and final rotations, \(c\) is the equilibrium rotation and \(k' = (k_1 + k_2)\) of Equation (9).

The mutarotation of sugars containing substantial proportions of the furanose forms in addition to the \(\alpha\)- and \(\beta\)-pyranose forms require an equation having two exponential terms. Such mutarotations are called complex mutarotations.
I.2.2  Forms of Glucose in Solution

I.2.2.1  Cyclic Forms of Glucose

As mentioned earlier glucose forms intramolecular cyclic hemiacetal structures by the reaction of the carbonyl group with neighboring C-5 hydroxyl group. The solutions and crystals of glucose consist primarily of only the anomers of D-glucopyranose. This has been established by X-ray analysis of crystals, uv and infrared measurements of freeze-dried samples and by nmr of equilibrated solutions.

It has been established that in aqueous solution glucopyranose occurs in the chair conformation and that the chair conformation with the OH groups attached to C-2, C-3, C-4 and CH$_2$OH on C-5 are in equatorial positions, is more stable. The other chair conformation with those groups in axial positions is less stable because of interactions among 1,3 axial substituents on the same side of pyranose ring.

The anomeric hydroxyl group on C-1 can be either axial (α-form VIII) or equatorial (β-form IX). The equatorial position is favored by steric effects and the axial position is favored by electrostatic repulsion between the C=O dipole due to anomeric hydroxyl group and the dipole of the unshared electron pairs on the ring oxygen (anomeric effect). The position of equilibrium is changed by the dielectric constant of the solvent. In solvents of high dielectric constant, e.g. water ($\varepsilon = 78.5$), the dipole-dipole interaction is lowered and the β-form is favored ($\alpha : \beta$ ratio = 36 : 64). In DMSO, a highly polar solvent ($\varepsilon = 49$), the equilibrium composition is almost the same. According to Kabayama and Patterson, the aldopyranose having an equatorial OH group fits into the tridymite structure of water, whereas the isomer having an axial anomeric OH does not. Thus the
coordination of the equatorial OH group with solvent also tends to counteract the anomeric effect. The anomeric effect is enhanced in less polar solvents and thus in methanol \((\varepsilon = 32.6)\) \(\alpha\)- and \(\beta\)-D-glucopyranoses have equal stabilities.\(^{64}\)

Solvent can also affect the pyranoid to furanoid ring ratio. Since bulky groups can be accommodated better on pyranoid than on furanoid rings, reducing the size of ring substituents favours furanose forms. Thus in aprotic solvents like DMSO and dimethylformamide, hydroxyl groups of sugars are less well solvated than in water and are consequently less bulky. Therefore the proportion of five-membered compounds increases.\(^{64}\) Although the furanose form has not been detected (i.e. \(<1\%\)) in aqueous solutions of glucose,\(^{64,66}\) about 2.3\% has been detected in dimethylformamide at 20\(^\circ\)C and about 3\% in refluxing pyridine,\(^{68}\) by gas chromatographic methods.

I.2.2.2 Acyclic Forms

Because of rapid reversible reactions, direct chemical methods for measuring the concentration of the acyclic form of a sugar in solution are not generally satisfactory. Determination of the concentration of the acyclic form by a physical means, which does not alter any equilibrium in solution, is more satisfactory.

Measurement of uv absorption clearly shows that under normal conditions the concentration of the free carbonyl form must be low\(^{51}(a)\). Infrared spectra of solvent-free equilibrated glucose showed faint absorption bands that can be attributed to the presence of traces of the free aldehyde sugar together with absorptions characteristic of \(\alpha\)- and \(\beta\)-pyranose forms.

The amount of reducible form present in solutions of several aldohexoses was measured by Los, Simpson and Wiesner\(^{69}\) by polarographic
methods for a three component equilibrium (Eq. 12). The proportion of the

\[
\begin{align*}
\alpha & \overset{k_1}{\underset{k_1}{\rightleftharpoons}} \gamma & \overset{k_2}{\underset{k_2}{\rightleftharpoons}} \beta
\end{align*}
\]  

reducible form (γ) averaged 0.0026% of the total D-glucose present, at 25°C in phosphate buffer of pH 6.9. The rate constants in 0.0153 M buffer are 

\[k_1 = 9.8 \times 10^{-3} \text{ min}^{-1}, \quad k_2 = 3.2 \times 10^{-3} \text{ min}^{-1}, \quad k_{-1} = 139 \text{ min}^{-1} \text{ and } k_{-2} = 78 \text{ min}^{-1}.\]  

The small proportion of the "γ" form is a reflection of the extreme rapidity of the ring closure reactions \((k_{-1} \text{ and } k_{-2})\). The absence of a substantial proportion of the aldehydrol in aqueous solution shows that the ring forms are far more stable than the aldehydrol. Formation of aldehydrol requires introduction of a hydroxyl group from the solvent. The observation that in \(\text{H}_2\text{O}^{18}\), one oxygen atom of D-glucose is exchanged only slowly \(^{70}\) and D-glucose-1-\(^{18}\)O exchanges its \(^{18}\)O with water at a rate lower than that of the mutarotation reaction \(^{71}\), shows that solvation of the acyclic form is slower than ring interconversion.

I.2.2.3 Ionic Forms

Reducing sugars are amphoteric. In aqueous solution they yield cations by addition of a proton, and anions by removal of a proton. The basic properties arise in large measure from the nucleophilic character of the ring O-atom; the acidic properties are attributable to the OH groups especially the anomeric OH group since the anomeric carbon is attached to ring oxygen as well. The \(pK_a\) for dissociation of the anomeric OH group on D-glucose at 25°C is 12.1. From this it can be shown that in a molar solution of D-glucose at 25°C there are 8.0 \(\mu\)moles of D-glucose anion/litre at \(\text{pH } 7^{51}(a)\).
I.2.3 Catalysis of Mutarotation by Acids and Bases

Mutarotation of glucose occurs in pure water and is catalyzed by acids and bases. The acceleration by acids was first reported by Erdmann in 1880 and its catalysis both by acids and bases was described by Urech in 1882.

Lowry and co-workers studied the mutarotation of tetra-O-methyl-\(\alpha\)-D-glucopyranose and found that the rate of reaction is low in dry pyridine or in dry cresol but high in a mixture of the two solvents or in either solvent when moist. Lowry and Smith concluded that the mutarotation requires an acid catalyst and a base catalyst and that amphoteric solvents are complete catalysts for the process whereas aprotic solvents are not. Thus in DMSO mutarotation of glucose proceeds extremely slowly, if at all, in the absence of acids and bases.

Although early workers attributed catalytic activity to acids and bases, later work revealed that catalysis of mutarotation is not the exclusive property of \(H^+\) ions and \(-\text{OH}\) ions. For example, Lowry and co-workers showed that molecules of undissociated acids (e.g. acetic acid), cations of weak bases (e.g. ammonium ion) and anions of weak acids (e.g. acetate ion) have catalytic properties. Much the same concept was developed by Brönsted and Guggenheim and came to be known as general acid and base catalysis. It was found that the rate of mutarotation of a sugar in the presence of a mixture of several catalysts may be represented by an equation of the type,

\[
(k_1 + k_2) = k_{H_2O}[H_2O] + k_{HA_j}[HA_j] + k_{B_n}[B_n] 
\]

where the symbols in brackets represent the concentrations (activities) of the catalysts and the coefficients \(k_{H_j}\) and \(k_{B_n}\) represent the catalytic
activities of the acid and base catalysts, respectively. This kind of rate equation arises from a stepwise process where the acid catalyst and the base catalyst act separately. If they act together, both the acid and the base catalysts take part in the transition state with the addition of a proton at one point in the molecule and with the elimination of a proton at another point. Theoretically, the rate constant for a concerted reaction involving several acid-base combinations yield an equation of the form \[ k = \sum \sum k_{H_j} k_{B_n} [HA_j][B_n] \] (14)

where the observed velocity depends on the product of the velocities of acid and base catalysts.

Numerous workers have examined the rate constants for the mutarotation of D-glucose in the presence of acetic acid and sodium acetate in an attempt to ascertain whether Equation (13) or (14) applies.

Although the two equations differ widely, it is not easy to distinguish experimentally between them. It is probable that the concerted mechanism is of no significance in the acid and base catalyzed mutarotation of sugars in aqueous solution. In solvents of low dielectric constants the formation of ionic intermediates becomes less favoured, and the concerted mechanism may apply (as discussed in Section 1.2.5.1.3). Some catalytic constants for the mutarotation of D-glucose in different solvents are given in Table III.

Comparison of results for benzoate catalysis in different solvents shows that there is almost no rate increase on going from water to DMSO or DMF as solvent. Thus for catalysis by benzoate the catalytic constant in DMSO is approximately equal to that in water (allowing for small temperature
TABLE III

CATALYTIC CONSTANTS \( \left(10^4 k_{\text{cat}} \right) \text{ LITRE MOLE}^{-1} \text{ SEC}^{-1}\) FOR MUTARotation OF GLUCOSE OR TETRAMETHYLGLUCOSE

<table>
<thead>
<tr>
<th>Catalyst</th>
<th>Water</th>
<th>DMSO</th>
<th>DMF</th>
<th>Dioxan</th>
<th>Benzene</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydronium Ion</td>
<td>100 (25°C)\textsuperscript{77}</td>
<td>140 (25°C)\textsuperscript{65}</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Acetic Acid</td>
<td>0.91 (18°C)\textsuperscript{52}</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ammonium Ion</td>
<td>0.20 (20°C)\textsuperscript{75(b)}</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Benzoate Ion</td>
<td>5.8 (18°C)\textsuperscript{52}</td>
<td>10.2 (25°C)\textsuperscript{62}</td>
<td>83.7 (25°C)\textsuperscript{62}</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Acetate Ion</td>
<td>10.2 (18°C)\textsuperscript{52}</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Hydroxide Ion</td>
<td>145 x 10\textsuperscript{4} (25°C)\textsuperscript{78}</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>2-Hydroxypyridine</td>
<td>1.47 (30°C)\textsuperscript{79}</td>
<td>0.88 (30°C)\textsuperscript{79}</td>
<td>48 (30°C)\textsuperscript{79}</td>
<td>700 (25°C)\textsuperscript{80}</td>
<td></td>
</tr>
<tr>
<td>Mutarotase</td>
<td>1625 x 10\textsuperscript{4} (21.5°C)\textsuperscript{81}\textsuperscript{b}</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Rate constant for catalysis by pure water\textsuperscript{77} at 25°C is \(4.1 \times 10^{-4}\) sec\(^{-1}\)

\(a\) Concentration of tetramethylglucose = 0.091 M

\(b\) Concentration of glucose = 0.11 M
difference) and that in DMF is approximately eight times large. This is in sharp contrast to studies involving anionic bases and nucleophiles in dipolar aprotic solvents where rate increases of approximately $10^6$ fold have been commonly observed. This unusual result could be due to the fact that the proton transfer in water from the base to the glucose molecule could occur via hydrogen bonded water molecule, there being no need for substantial desolvation of the base.

The unusually high activity of 2-hydroxypyridine in benzene will be discussed in Section 1.2.5.1.3.

I.2.3.1 Lewis Acids as Catalysts

There are a few reports of Lewis acid catalysis of the glucose mutarotation. According to Broser and Ruecker $^{82}$ AlCl$_3$ and ZnCl$_2$ in acetic acid/sodium acetate buffer pH 3.7, showed catalytic activity due to their action as Lewis acids. MgCl$_2$ showed minor catalytic effect and NaCl had no effect at all. According to Mitzner and Behrenwald $^{83}$ mutarotation of sugars is catalyzed by Ni$^{2+}$ ion due to the formation of a sugar-metal ion complex. Bobbio et al. $^{84}$ have reported that the mutarotation of glucose is catalyzed by Cu$^{2+}$ due to formation of a complex in which C-1 of the sugar is involved.

I.2.3.2 Catalytic Effect of Salts

Ordinarily, salts of strong acids and strong bases have little influence on the rate of mutarotation in water. $^{52}$ Thus mutarotation of glucose is retarded slightly by sodium chloride or lithium chloride $^{51(b)}$. Under some conditions however, large effects are observed. Thus Eastham and co-workers $^{85}$ found that the mutarotation of tetramethylglucose in
pyridine or in nitromethane is slow, but that it is enhanced by a factor of ten on addition of 2 mM lithium perchlorate. It has been suggested that the salt stabilizes the transition state by formation of an ion-pair complex. Further, Rony et al.\textsuperscript{86} have shown that ion pairs such as PhO\textsuperscript{-} + NR\textsubscript{4}\textsuperscript{+} are effective general base catalysts for mutarotation of tetramethylglucose in benzene.

I.2.4 \textbf{Isotope Effects in Mutarotation Reactions}\textsuperscript{51(b),87}

Deuterium isotope effects, \( k_{H}/k_{D} \), arise from a combination of kinetic and solvent isotope effects. Kinetic isotope effects are caused by differences in the energy required for alteration of the normal and isotopic bond in the corresponding transition states; solvent isotope effects can exist when the isotopic compound is used both as a reactant and as a solvent. The observed isotope effects may have values smaller than, equal to, or greater than unity. The differences in the rates of mutarotation of sugars in water and in deuterium oxide have provided valuable means for studying the mechanisms of mutarotation reactions.

I.2.4.1 \textbf{Kinetic Isotope Effect}

A kinetic isotope effect is large when the bond joining the isotopic atom to the substrate is formed or is broken in the rate-determining step. In general, the stronger the altered bond, the greater the isotope effect. Ordinarily, the heavier isotope gives the lower reaction rate\textsuperscript{88}, and, hence, values of \( k_{H}/k_{D} \) greater than unity are designated as normal isotope effects and values less than unity as inverse isotope effects.

But if the proton transfer occurs after the rate-controlling step it will have no primary kinetic consequence. If proton addition occurs
before the rate-controlling step and no proton transfer occurs in the transition state (i.e. no kinetic isotope effect), the reaction is subject to hydronium ion catalysis and an inverse isotope effect (due to solvent) is observed.

I.2.4.2 Solvent Isotope Effect

Change of solvent from $\text{H}_2\text{O}$ to $\text{D}_2\text{O}$ causes changes in the degree of ionization and solvation both of the reactants and the catalyst. The reactions that involve a rapid pre-equilibrium, in which proton transfer from the catalyst to the substrate (HS) occurs before rate-determining step, will have a higher concentration of the intermediate conjugate acid in $\text{D}_2\text{O}$ than in $\text{H}_2\text{O}$ because $\text{D}_3\text{O}^+$ is a stronger acid than $\text{H}_3\text{O}^+$. Therefore, whether the observed isotope effect is normal or inverse will depend on the nature of rate-determining step. If the rate-determining step does not involve a proton transfer, an inverse isotope effect will be observed due to the higher concentration of protonated substrate ($\text{HSD}^+$) in $\text{D}_2\text{O}$ and the reaction is subject to specific acid catalysis.

\[
\begin{align*}
\text{HS} + \text{H}_3\text{O}^+ & \rightleftharpoons \text{HSH} + \text{H}_2\text{O} \\
\text{HS} + \text{D}_3\text{O}^+ & \rightleftharpoons \text{HSD} + \text{D}_2\text{O} \\
\text{HSH} & \rightarrow \text{products} \\
\text{HSD} & \rightarrow \text{products}
\end{align*}
\]

However, a proton (or deuteron) transfer in the rate-determining step will give rise to a normal kinetic isotope effect for that step. Hence the observed isotope effect can be normal if the kinetic isotope effect of the rate-determining step is large enough. Such mechanisms are subject to
general acid catalysis.

The results of deuterium isotope effect studies with glucose mutarotation will be discussed in the next section where the possible mechanisms of mutarotation are considered.

I.2.5 The Mechanism of Mutarotation of Glucose

There are at least three mechanisms possible for the mutarotation of glucose.\(^{58,65}\)

1. Mutarotation can proceed through the free aldehyde form (\(\gamma\)-glucose):

\[
\begin{align*}
\alpha-\text{Glucose} & \quad \rightarrow \quad \gamma-\text{Glucose} \\
\text{Acids can catalyze the reaction by protonation of the ring oxygen; thereby the ease with which C-O bond can be broken is increased. Bases can catalyze by removing the proton from the anomeric hydroxyl group.} \\
\beta-\text{Glucose} & 
\end{align*}
\]

2. According to Christansen,\(^{90}\) the reaction proceeds as a result of the removal of the C-1 hydrogen, although it is difficult to understand why C-1 hydrogen should ionize over hydrogen on the anomeric OH group.
Acid catalysis could presumably result from the prior protonation of the neighboring oxygen.

(3) Another pathway is the direct displacement of anomeric OH by a water molecule or hydroxide ion. It can occur via carbonium ion formation in the case of acid catalysis.

\[ \text{H}_2\text{O}^+ + \text{H} \]

All the evidence available from mechanistic studies in solution supports Mechanism (1). The best evidence for this is the observation that glucose-1-\(^{18}\)O undergoes oxygen exchange with water more than 30 times slower than the mutarotation reaction\(^ {58,71}\). If the reaction involved elimination and addition of a OH group at the anomeric carbon atom (Mechanism (3)), the rate of exchange of \(^{18}\)O in D-glucose-1-\(^{18}\)O should be equal to the rate of mutarotation, and according to Mechanism (2), the anomeric hydroxyl group should not exchange at all. Further, if Mechanism (3) occurs by S\(^\text{N}2\) type displacement of anomeric OH group by OH ions, then there should not be any deuterium isotope effect. While, acid catalyzed carbonium ion type mechanism should give rise to an inverse isotope effect due to solvent as discussed above. However, normal deuterium isotope effects have been observed for glucose mutarotation under all conditions studied so far\(^ {51(b)}\).

The small proportion of \(^{18}\)O exchanged may occur by reversible formation of a gem-diol of the intermediate acyclic sugar. The existence
of the free aldehyde form has been confirmed by physical and chemical methods as mentioned in Section 1.2.2.2. The observation that exchange of $^{18}O$ is relatively slow indicates that the nascent carbonyl group of the acyclic intermediate reacts faster, intramolecularly, with neighboring hydroxyl groups than with the hydroxyl groups of the solvent. To explain the rapidity of the ring closure reaction and the slow exchange of anomeric oxygen with solvent, Isbell et al.\textsuperscript{91} postulated that the cyclic sugar passes through a so-called "pseudo-acyclic" intermediate during mutarotation.

1.2.5.1 Specific Mechanisms

The acyclic intermediate postulated in Mechanism (1) can be formed by acid or base catalysis or by pure water acting as a catalyst. Thus it can be formed by several specific mechanisms which are discussed below.

1.2.5.1.1 Catalysis by Acids

\[
\begin{align*}
\text{O} & \quad \text{H} + \text{HA} \quad \text{rapid} \quad \text{H} + \text{A}^- \quad \text{slow} \quad \text{OH} \quad \text{H} + \text{A}^- \\
\text{OH} & \\
\text{OH} & \quad \text{OH} \\
\text{O} & \quad \text{H} + \text{HA}
\end{align*}
\]

If all rate constants are of the same order this mechanism leads to complicated kinetics. But first order kinetics has been observed for glucose
mutarotation as mentioned earlier.

If the first step is rapid and ring opening reaction is rate-determining, then the rate is proportional to \([\text{glucose}][H_3O^+]\). Therefore the reaction is subject to specific acid catalysis and should show an inverse isotope effect due to higher concentration of deuteronated glucose in D$_2$O.

On the other hand, if the conjugate base $A^-$ takes part in the rate-determining step by removal of the proton on anomeric OH group (Equation 16), the rate is proportional to \([\text{glucose}][HA]\) and the reaction is subject to general acid catalysis. Further, the second step would show a normal kinetic isotope effect and thus the overall reaction can have a normal isotope effect. Hence the mechanism shown in Equation (16) is a possible route for the mutarotation catalysis by acids. For catalysis by strong acids $HA = H_3O^+$ and $A^- = H_2O$. For strong acid catalysis in water, deuterium isotope effect observed is about 1.37$^{92}$ at 25°C while for weak acids (e.g. acetic acid) it is around 2.6$^{51}$ (b) at 25°C. The higher isotope effect with the weak acid is due to the formation of a stronger H-$A$ bond (compared with weaker H-OH$_2$ bond) in the transition state.

Another possible mechanism for acid catalysis is shown in Equation (17), where the ring opening occurs during H-$A$ bond rupture. This mechanism also leads to general acid catalysis, first order kinetics and a normal
isotope effect. But Capon and Walker have shown that for such a mechanism, the rate constant \( k_R \) for the reverse of rate-determining step would be higher than the diffusion controlled limit for a strong acid, weak acid or even for the water catalyzed mutarotation of glucose. Hence the mechanism shown in Equation (16) is the preferred mechanism for acid catalyzed mutarotation. However, they also showed that for the water catalyzed reaction by Mechanism (16) the rate constant for the rate-determining step would be above the rate constant for diffusion. The mechanism in Equation (16) (as well as Equation 17) is therefore not possible for water itself.

Electron withdrawing substituents at 6-position (i.e. 6-substituted 6-deoxy glucoses) decreased the rate of mutarotation by oxonium ions and by water. The \( \rho \) value of -2.87 for acid catalysis is consistent with a mechanism in which proton transfer to the ring oxygen atom is further advanced than breaking of the ring C(1)-O bond. The \( \rho \) value for water catalysis was -2.92.

In addition, 2-substituted 2-deoxy-D-glucoses gave results consistent with Mechanism in Equation (16). Thus 2-deoxy-D-glucose showed a higher rate, and conjugate acid of 2-amino-2-deoxy-D-glucose showed a lower rate for mutarotation than D-glucose itself.

\[ \text{Base Catalysis} \]

Again, two mechanisms, similar to those considered for acid
Both the mechanisms would show first order kinetics, general base catalysis and normal isotope effects. The deuterium isotope effects observed for base catalysis are generally higher than that observed for strong acid catalysis because of the stronger H-B bond formed. The observed $k_H/k_D$ values for acetate, pyridine and hydroxide (without correcting for catalysis by glucosate anion) are $2.38^{94}$, $2.86^{95}$ and $1.83^{87}$ respectively.

Capon and Walker showed that for mechanism in Equation (19) the rate constant $k_R$ for the reverse of the rate-controlling step would be close to or greater than the diffusion controlled limit, while for mechanism in Equation (18) the rate constant for the slow step is well within the diffusion controlled limit. Hence mechanism in Equation (18) is favoured for base catalysis.
In contrast to acid catalysis, electron withdrawing substituents at the 6-position increased the rate of mutarotation catalyzed by bases. The p values for catalysis by Tris and hydroxide ions are +3.97 and +6.22 respectively.

1.2.5.1.3 Concerted Mechanism

The mechanisms discussed above for acid and base catalysis of mutarotation in aqueous solution involve ionic intermediates which can be easily stabilized by polar water molecules. But in solvents of low dielectric constant the formation of free ions becomes less likely, so that a concerted mechanism may be favoured even for reactions which do not follow it in water or similar solvents. In fact, the concerted mechanism was first put forward by Lowry on the basis of observations on the mutarotation of tetramethylglucose in media of low dielectric constant. As mentioned earlier (Section 1.2.3), this reaction was very slow in dry pyridine (possessing no acid properties) or in dry cresol (possessing hardly any basic properties), but was rapid in a mixture of the two solvents or in either solvent when moist, suggesting that both an acid and a base must take part in the reaction.

Swain and Brown studied the same reaction in dilute benzene (ε = 2.3) solutions of amine and phenol. They found that the reaction was kinetically of the 3rd order, the velocity being proportional to the product of the concentrations of phenol, amine and tetramethylglucose as expected from Equation (14). But later Pocker pointed out that the rate proportional to the product of concentrations of amine and phenol could be attributed to base catalysis by phenoxide ion in an ion pair such as PhO⁻·NH₃⁺. This is supported by the demonstration that ion-pairs such as PhO⁻·NR₄⁺ are effective basic catalysts for mutarotation of tetramethylglucose in
benzene, although they contain no acidic group.

Swain and Brown \(^{80}\) also showed that 2-hydroxypyridine is a powerful specific catalyst for the mutarotation reaction; at a concentration of 0.001 M it is 7000 times as effective as a mixture of 0.001 M pyridine and 0.001 M phenol, though it is only one ten-thousandth as strong a base as pyridine and one-hundredth as strong an acid as phenol. Since, further, the velocity is proportional to the first power of the concentration of 2-hydroxypyridine, it is clear that the operation of the concerted mechanism where both the proton transfers occur simultaneously (Equation 20) is facilitated by the presence of an acidic and a basic group in the same catalyst molecule. On the other hand, 3- or 4- hydroxypyridines where the nitrogen and hydroxyl groups are too far apart, do not show bifunctional catalysis, and consequently they are poor catalysts \(^{80}\).

\[
\text{Catalyst Molecule} \quad \begin{array}{c}
\text{OH} \\
\text{O} \\
\text{H} \\
\text{HN} \\
\text{O} \\
\text{O} \\
\text{O} \\
\end{array} \quad \text{+} \quad \begin{array}{c}
\text{OH} \\
\text{O} \\
\end{array} \quad \text{Mutarotation} \\
\]

Since two O-H bonds are broken in the transition state one would expect a deuterium isotope effect higher than that observed for bases. In fact, the observed isotope effect for 2-hydroxypyridine catalysis of the mutarotation of tetramethylglucose in benzene is 3.5 \(^{99}\).

It has been shown that bifunctional catalysts are effective in non-polar solvents only when they can interact with the substrate without formation of high energy dipolar intermediates \(^{100,101,102}\); this implies that they can exist in two tautomeric forms of comparable energy, and hence
are also described as tautomeric catalysts. Other examples of bifunctional catalysts in non-polar media are carboxylic acids, pentane-2,4-dione, and pyrazole.\(^\text{102}\)

The above catalysts act as effective bifunctional catalysts in benzene because of the absence of effective acidic and basic groups on the solvent. No such bifunctional catalysis has been observed for these catalysts in water\(^\text{79,80,103}\) since it can stabilize any polar intermediates formed and also can act as an acid or base. Thus 2-hydroxypyridine does not show bifunctional catalysis in water and the observed deuterium isotope effect is 2.3\(^\text{99}\) (compared with \(k_H/k_D = 3.5\) in benzene, 2.91 in DMSO-\(\text{D}_2\text{O}\) mixture with \(\text{D}_2\text{O}\) concentration = 22.2 mol dm\(^{-3}\) and 2.53 in DMSO-\(\text{D}_2\text{O}\) mixture with \(\text{D}_2\text{O}\) concentration = 44.4 mol dm\(^{-3}\))\(^\text{79}\).

There are other factors that determine whether catalysis will occur by a stepwise or a concerted mechanism. For example, Fiandanese and Naso\(^\text{104}\) have shown that benzamidine, a potential bifunctional catalyst, is a powerful catalyst for tetramethylglucose mutarotation in benzene. They also showed that catalytic activity is not due to a bifunctional action, but due to its activity as a base. They concluded that the presence of both acidic and basic centres suitably located in a tautomeric system does not represent a sufficient condition for a bifunctional intervention in the mutarotation process. With benzamidine, the high strength and reactivity of the basic centres make irrelevant the advantage occurring from its favourable position relative to the acidic function.

I.2.5.1.3.1 Catalysis by Water

It was mentioned in Section I.2.5.1.1 that although the substituent effects for acid catalysis and water catalysis are the same, the
rate constant for the rate-determining step, for water catalysis by the mechanism in Equation (16), would be above the rate constant for diffusion. Also, the difference in the substituent effects between water and base catalyzed mutarotation of the 6-substituted 6-deoxy glucoses makes it unlikely that these reactions follow the same mechanism.

Additional evidence against the conventional stepwise mechanism has come from the observation that some of the necessary fractionation factors for mutarotation of α-D-tetramethylglucose in mixed H₂O-D₂O solvents were difficult to interpret when the stepwise mechanism was considered. However, a cyclic concerted mechanism involving two or three water molecules led to plausible fractionation factors.

Direct evidence for participation of several water molecules in the transition state has come from the effect of solvent composition, in mixtures of water and an organic solvent, on the mutarotation reaction. For example, Ballash and Robertson found that in DMSO-water mixtures, the order with respect to water is about three. But for acid catalysis under similar conditions the order with respect to water was zero. Chin and Huang studied the mutarotation of α-D-tetramethylglucose in aqueous dioxan and DMSO solutions. They found that for both solvents the order with respect to water was about two. For pyridine catalyzed reaction the order with respect to water was one, suggesting that one molecule of water has been replaced by a molecule of base. Hence the following concerted mechanism involving two water molecules, one acting as an acid and the other as a base, has been postulated (see page 59).

It is important to note that a cyclic transition state involving a bifunctional catalyst can catalyze a reaction in a concerted (i.e. correspondence between the proton transfers within times of <10⁻¹⁰ sec).
or stepwise manner. For example, Bell and his co-workers\textsuperscript{107} have calculated the activation energies for proton transfers by concerted and stepwise mechanisms for the reversible hydration of 1,3-dichloroacetone. Their results indicate that the activation energy for the concerted mechanism is about twice as great for the stepwise process. Hence they concluded that the latter would be preferred. This conclusion was supported by the fact that the observed deuterium isotope effect for the uncatalyzed reaction is 2.7. This is low compared with the abnormally high isotope effect expected for the concerted process. The same argument may apply for the mutarotation reaction. For example, although the observed deuterium isotope effect for water catalysis is about 3.8\textsuperscript{51(b)} at 25°C, the isotope effect expected for a concerted process, according to calculations by Schowen\textsuperscript{99}, is 12. Hence Chin and Huang\textsuperscript{105} have suggested that water catalyzed mutarotation occurs by an intimate stepwise mechanism rather than a concerted process.

\textbf{1.2.6 Heterogeneous Catalysis of the Glucose Mutarotation}

All the results discussed above refer to studies of glucose mutarotation carried out using homogeneous catalysts. On the other hand, there are only a few reports of heterogeneous catalysis of the glucose mutarotation by solid catalysts. Thus Tanabe et al.\textsuperscript{108} have shown that silica-alumina, nickel sulphate, and solid phosphoric acid catalyze the
mutarotation of tetramethyl-α-D-glucopyranose in benzene solution, with the observed activity depending on the acid properties of the catalyst. Silica-alumina which has the highest acid strength showed the greatest catalytic activity. They suggested that silica-alumina and nickel sulphate may act as bifunctional catalysts since they both have acidic and basic sites. Nickel sulphate heat treated at 250°C showed higher catalytic activity than that heated at 350°C, indicating that the catalytically active sites are the Brønsted type but not the Lewis type.

According to Rustamov and Tyrina the mutarotation of glucose in water was accelerated by nickel prepared by reduction of $\text{Ni}_2(\text{OH})_2\text{CO}_3$ in hydrogen at 300°C. The activation energies for the forward and reverse reactions were decreased from 13.65 to 7.32 and 14.63 to 9.38 Kcal./mole respectively. Heating nickel in vacuo to remove the adsorbed hydrogen did not affect its activity.

Rustamov and Usmanov have shown that the mutarotation of glucose in water is not catalyzed by anion exchange resins AN-2F, EDE-10P and PE-9.

According to Schmid and Bauer mutarotation of α-D-glucose is catalyzed by finely divided copper but the activation energy for heterogeneous copper catalysis (19 Kcal./mole) agreed, within experimental error, with that for homogeneous catalysis by water. Hence they concluded that heterogeneous copper catalysis is by water molecules adsorbed on the copper surface.

Hence it is clear that there is only a little work done on heterogeneous catalysis of glucose mutarotation and they are also incomplete. Because of the lack of any systematic study of the kinetics and mechanisms of the catalysis of glucose mutarotation by solid catalysts, it was decided to study the effect of aluminum oxide on the above reaction. It should,
for the first time, provide information on the kinetics and mechanism of an alumina catalyzed simple organic reaction in solution. The results obtained can be compared with the results already available on the mechanism of homogeneous catalysis, to determine similarities and differences in the mechanisms. The same approach, to understand the action of solid acid/base catalysts, has been used in the study of alcohol dehydration on solid surfaces based on advances made in understanding its mechanism in homogeneous systems.

Since the mutarotation reaction is more sensitive to Brønsted acids than any of the reactions that have been studied on alumina, it could be a sensitive probe for the presence of catalytically active Brønsted acid sites on the surface. Deuterium isotope effect studies can provide evidence for the presence of general or specific acid-base catalysis, cyclic or acyclic mechanism and bifunctional catalysis. Presence of bifunctional catalysis would indicate the presence of acidic and basic sites suitably located on the surface.

The solvent chosen was dimethyl sulfoxide. It is a highly polar ($\varepsilon = 49$) aprotic solvent which dissolves glucose easily and does not catalyze the mutarotation. Hence catalysis can be easily arrested by filtering the slurry. The non-aqueous solvent will also enable the study of the effect of water on the surface reaction.

Commercially available Woelm alumina neutral for thin layer chromatography (hence forth referred to as alumina neutral) was chosen as the catalyst, since it is widely available and is similar to the alumina used by Posner (Woelm 200 neutral chromatographic alumina) in his studies of alumina catalyzed organic reactions.

The physical characteristics of alumina neutral and other aluminas
prepared by dehydration of alumina neutral are given in the next section. This thesis is then developed by a general study of the kinetics of glucose mutarotation by alumina. It is followed by a study of glucose adsorption onto the surface, and then, by investigations into the nature of the catalytically active sites. Finally, the mechanism of the surface catalyzed reaction is discussed.
II CHARACTERIZATION OF ALUMINAS
II CHARACTERIZATION OF ALUMINAS

Alumina neutral and other aluminas prepared by pyrolysis of alumina neutral at different temperatures (Section XII) were characterized by determining surface areas, pore size and particle size distributions, crystalline structure and trace impurities present. The theory and experimental methods are described in Appendix A. Hence only the results of these determinations are given in this section.

II.1 Surface Areas

The surface areas were determined by the multipoint BET method using nitrogen gas as adsorbate. The results given in Table IV show that specific surface areas have decreased with increase of pyrolysis temperature. The extent of sintering (i.e. adhesion of the particles of a solid to form aggregates) has increased markedly above the Tammann temperature ($\approx 870°C$) and aluminas pyrolyzed at 1100°C and 1250°C possess low specific surface areas.

II.2 Pore Size Distributions

The adsorption and desorption isotherms of alumina neutral and alumina pyrolyzed at 800°C showed the presence of both capillary condensation and hysteresis. These aluminas are therefore porous and the pore size distributions, calculated using the desorption isotherms, are given in Fig. 3. This figure shows that most of the pore volume in alumina neutral is in pores of radius $\approx 23 \, \text{Å}$ while most of the pores have radii between 18 and 27 Å. With alumina pyrolyzed at 800°C most of the pore volume is in pores of radius $\approx 28 \, \text{Å}$ while most of the pores
### TABLE IV

**BET SURFACE AREAS**

<table>
<thead>
<tr>
<th>Sample Preparation</th>
<th>BET Surface Area $m^2/g$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Alumina Neutral</td>
<td>140 ± 5</td>
</tr>
<tr>
<td>Dehydrated at 150 ± 5°C, 0.01 mm pressure, 2 days</td>
<td>130 ± 5</td>
</tr>
<tr>
<td>Dehydrated at 600 ± 50°C, 4 hours under dry N$_2$</td>
<td>116 ± 4</td>
</tr>
<tr>
<td>Dehydrated at 800 ± 50°C, 4 hours under dry N$_2$</td>
<td>100 ± 4</td>
</tr>
<tr>
<td>Dehydrated at 1100 ± 50°C, 3 hours under dry N$_2$</td>
<td>14 ± 1</td>
</tr>
<tr>
<td>Dehydrated at 1250 ± 50°C, 6 hours under dry N$_2$</td>
<td>6.2 ± 0.1</td>
</tr>
</tbody>
</table>

have radii between 22 and 31 Å. Further, the shapes of the hysteresis loops can be taken to indicate that the pores are cylindrical.\textsuperscript{114,115}

On the other hand, the adsorption and desorption isotherms of alumina pyrolyzed at 1250°C showed no capillary condensation or hysteresis. This indicates the absence of a porous structure and is due to sintering of alumina particles at high temperatures.\textsuperscript{113,114,115,116}
Fig. 3 Pore Size Distributions of (a) Alumina Neutral and (b) Alumina Pyrolyzed at 800°C for 4 hours
II.3 Particle Size Distributions

II.3.1 Particle Size by use of an Electrozone Celloscope

The particle size distributions of alumina neutral and sintered (1250°C) alumina were determined using an Electrozone Celloscope and are given in Fig. 4. The characteristics of the plots are shown in Table V. They show that the mean particle diameters are 7.92 microns for alumina neutral and 7.39 microns for 1250°C alumina. The approximately 7% decrease in the particle size on pyrolysis at 1250°C is due to sintering of particles. Further, >99% of the particles have diameters between 2.8 and 30 microns.

<table>
<thead>
<tr>
<th>Particle Diameter (microns)</th>
<th>Log mean</th>
<th>Mode</th>
<th>Median</th>
</tr>
</thead>
<tbody>
<tr>
<td>Alumina Neutral</td>
<td>7.92</td>
<td>8.49</td>
<td>7.74</td>
</tr>
<tr>
<td>1250°C Alumina</td>
<td>7.39</td>
<td>6.59</td>
<td>7.22</td>
</tr>
</tbody>
</table>

II.3.2 Particle Size by use of an Electron Microscope

The alumina particles were also observed with a Scanning Electron Microscope. As described in other reports, particles of sintered alumina (Fig. 6) show smoother and rounder edges than those of alumina neutral (Fig. 5). Further, the particles shown in Figs. 5 and 6 have average diameters ranging from 9 to 12 microns which fall within the range
Fig. 4 Particle Size Distributions of (a) Alumina Neutral and (b) Alumina Sintered at 1250°C for 6 hours.
Fig. 5  Electron Micrograph of Alumina Neutral for TLC (x 4000)

Fig. 6  Electron Micrograph of Sintered (1250°C) Alumina (x 8000)
of diameters determined using Electrozone Celloscope.

II.4 Crystalline Structure of Aluminas

X-ray powder photographs showed that alumina neutral is mainly γ-alumina with some κ- (and Χ-) alumina. On the other hand, alumina sintered at 1250°C contains only α-alumina.

II.5 Trace Impurities Present

Trace metals present in the alumina samples were determined using an Inductively Coupled Argon Plasma Spectrograph (by Can Test Ltd.). The results showed that there is 0.2% trace metals (by weight) in alumina neutral, the main impurities being Ca (0.11%), Fe (0.03%), Na (0.02%), and Mg (0.02%).
III    PRELIMINARY KINETIC STUDIES
III PRELIMINARY KINETIC STUDIES

The kinetics of the heterogeneous catalysis was carried out by stirring a 0.05 M solution of α-D-glucose in dimethyl sulfoxide with alumina (26.7 mg/ml) at 25.0°C as described in the Experimental Section. A typical first order plot of $\ln 10^3 (\alpha_t - \alpha_\infty)$ versus time (mins), where $\alpha_t =$ optical rotation at time $t$ and $\alpha_\infty =$ optical rotation when equilibrium is finally reached, for the catalysis of mutarotation of α-D-glucose by alumina neutral is shown in Fig. 7a, and that for β-D-glucose in Fig. 7b. The reproducibility of the results was excellent (see Section XI).

There are two characteristic features in these plots. First, with both α- and β-D-glucose, there is an initial rapid decrease in optical rotation. In the case of β-D-glucose the optical rotation decreases initially, although the actual mutarotation of β-D-glucose which leads to the formation of α-D-glucose causes an increase in optical rotation. Secondly, unlike the well known linear first order plots obtained for homogeneous catalysis of glucose mutarotation in solution, the first order plots are curved throughout the runs. This continued curvature in the first order plots means either the surface reaction is not first order in glucose or the catalyst is undergoing progressive deactivation during the reaction.

For both these plots in Fig. 7 the experimental infinity value of the optical rotation ($\alpha_\infty$) was used. It was the same in both cases (+0.124° at $\lambda = 365$ nm) and was less than the infinity value observed (+0.144°) for homogeneous catalysis (see Experimental Section). This difference in the infinity values could be due to adsorption of glucose onto alumina, which might also give rise to the initial rapid decrease in optical rotation in both plots, or because there are side reactions occurring on the surface of alumina which might also be responsible for the non-linear first order plots.
Fig. 7  First Order Kinetic Plots for Mutarotation of (a) α-D-Glucose and (b) β-D-Glucose by Alumina Neutral
Therefore an analysis of the reaction products in solution was carried out to determine whether any side products are formed during the mutarotation catalyzed by alumina.
IV  PRODUCT ANALYSIS
IV PRODUCT ANALYSIS

The products of the mutarotation reaction were determined by glc analysis of trimethylsilyl ethers of glucose, prepared according to the procedure of Sweeley et al.\textsuperscript{68}, as described in the Experimental Section.

Two main peaks and a small peak were observed. The three peaks A, B and C had relative retention times 0.82, 1.00 and 1.35 which accounted for 2.2%, 40.3% and 57.5%, respectively, of the total area. The two main peaks B and C were identified as penta-O-trimethylsilyl-\(\alpha\)- and -\(\beta\)-D-glucopyranose, respectively, by comparisons of their retention times with those of the trimethylsilyl ethers of pure \(\alpha\)- and \(\beta\)-D-glucopyranose. These results agree well with those of Hveding et al.\textsuperscript{67} who carried out the mutarotation (homogeneous catalysis by solvent) in dimethylformamide. They identified a small peak with relative retention time 0.82 (area about 2.3% when the reaction was carried out at 20\degree C), as the trimethylsilyl ethers of \(\alpha\)- and \(\beta\)-D-glucofuranose.

These results show that the catalysis of the mutarotation of \(\alpha\)-D-glucose by alumina produces \(\beta\)-D-glucose and that side reactions are negligible. Therefore the cause of the almost 14\% discrepancy between the infinity values must be the adsorption of glucose onto alumina (see adsorption studies in the next section).
V ADSORPTION OF GLUCOSE ON ALUMINA SURFACE

PART I PRELIMINARY STUDIES
V ADSORPTION OF GLUCOSE ON ALUMINA SURFACE

PART I PRELIMINARY STUDIES

A heterogeneous catalytic reaction in solution consists of five categories of elementary steps (Equation (21)); (i) diffusion of reactant(s), \( S \), from solution to the surface of the catalyst, (ii) adsorption of the reactant(s) on the surface, (iii) reaction on the surface forming the product(s), \( P \), (iv) desorption of product(s) from the surface, and (v) diffusion of product(s) from the surface to the bulk solution.

\[
\begin{array}{c}
\text{diffusion} \quad S \rightarrow S + C \\
\text{adsorption} \quad SC \rightarrow P + C \\
\text{reaction} \quad SC \rightarrow PC \\
\text{desorption} \quad PC \rightarrow P \\
\text{diffusion} \quad P \rightarrow P + C
\end{array}
\]

(21)

Therefore, before a surface catalyzed reaction can occur the reactant must get adsorbed on the catalyst surface. Once the surface reaction has occurred the product formed must be able to desorb into the solution. Hence reversible adsorption of glucose on alumina is an essential pre-requisite for mutarotation catalysis by alumina.

To study the adsorption of glucose on alumina surface, a solution of \( \alpha \)- and \( \beta \)-D-glucose, of equilibrium composition (henceforth referred to as the \( \alpha,\beta \) mixture), in DMSO was prepared as described in the Experimental Section. Sixty ml 0.05 M solutions of \( \alpha,\beta \) mixture were stirred with different weights of alumina and the change in optical rotation with time was followed until there was no further change. Since the solutions are completely equilibrated any decrease in optical rotation must be due to adsorption of glucose on the alumina surface. In all cases a decrease in optical rotation was observed. The greater the weight of alumina used
the greater was the amount adsorbed. A graph of change in optical rotation \((\alpha_0 - \alpha_{\text{eq}})\), which is proportional to the amount of glucose adsorbed, where \(\alpha_0\) is the initial optical rotation of the solution of \(\alpha,\beta\) mixture and \(\alpha_{\text{eq}}\) the equilibrium optical rotation of the slurry, versus the weight of alumina used is given in Fig. 8. A plot of \(1/\alpha_{\text{eq}}\) versus the weight of alumina used is shown in Fig. 9. It is clear that the plot in Fig. 9 is linear when the weight of alumina is low but curves upwards when the weight of alumina is increased, while the plot in Fig. 8 is curved throughout the range of concentrations of alumina used. Further experiments showed that the upward curvature in Fig. 9 can be reproduced whether the slurry is stirred slowly or rapidly, indicating that the increase in adsorption observed at higher concentrations of alumina is not due to break-up of particles due to a rapid rate of stirring.

V.1 Theoretical Study of Adsorption on Reversible and Irreversible Sites

The following model systems were analyzed to investigate the type of adsorption which gives rise to the plots observed in Figs. 8 and 9.

V.1.1 Only Reversible Adsorption

If there are only reversible adsorption sites, at equilibrium we get,

\[
G_{\text{eq}} + C \rightleftharpoons GC
\]  

(22)

where \(G_{\text{eq}}\) is glucose in solution, \(C\) designates the free adsorption sites on catalyst and \(GC\) is the glucose-catalyst complex.

Therefore equilibrium constant \(K = \frac{[GC]}{[G_{\text{eq}}][C]}\) litre mole\(^{-1}\)  

(23)
Fig. 8. Plot of Change in Optical Rotation versus Weight of Alumina for Adsorption of Glucose on Alumina Neutral
Fig. 9  Plot of $\frac{1}{\text{Equilibrium Optical Rotation (degrees)}}$ versus Weight of Alumina for Adsorption of Glucose on Alumina Neutral
when all concentrations are expressed in mole litre\(^{-1}\) and,

\[ [G\text{eq}] = [G_o] - [GC] \]  

(24)

where \([G_o]\) = initial concentration of glucose in solution.

Substituting \([GC] = K [G\text{eq}][C]\) in Equation (24) we get,

\[ [G\text{eq}] = [G_o] - K [G\text{eq}][C] \]

Therefore \([G\text{eq}] = \frac{[G_o]}{1 + K [C]}\)

Rearranging the above equation we get,

\[ \frac{1}{[G\text{eq}]} = \frac{K}{[G_o]} [C] + \frac{1}{[G_o]} \]

Therefore a graph of \(\frac{1}{[G\text{eq}]\text{eq}}\), or of \(\frac{1}{\alpha\text{eq}}\) which is proportional to \(\frac{1}{[G\text{eq}]\text{eq}}\), versus \([C]\) would be a straight line at all catalyst concentrations. At high catalyst concentrations the fraction of the adsorption sites used to form \(GC\) complex would be small and the concentrations of all the sites on the catalyst \([C_o]\) \(\approx [C]\). Hence a graph of \(\frac{1}{\alpha\text{eq}}\) versus \([C_o]\) should be linear at high values of \([C_o]\). For example, the theoretical plot of \(\frac{1}{[G\text{eq}]\text{eq}}\) versus \([C_o]\) for a system where \([G_o] = 1 \text{ and } K = 0.01\) is given in Fig. 10. The values of \([GC]\) (and hence \([G\text{eq}] = [G_o] - [GC]\)) for different values of \(C_o\) were determined by solving Equation \(0.01 = \frac{[GC]}{(1-[GC])([C_o] - [GC])}\), obtained from Equations (23) and (24). It is clear, unlike the experimental plot in Fig. 9, that the theoretical plot for reversible adsorption is completely straight, even at high concentrations of the catalyst, where about 80% adsorption is observed. Note that the weight of catalyst used is proportional to the total number of sites \(C_o\).
$[G_o] = 1 \text{ mole.litre}^{-1}$

$K = 0.01 \text{ mole}^{-1}\text{.litre}$

Fig. 10  Theoretical Plot of $1/[G_{eq}]$ versus $[C_o]$, the Concentration of Adsorption Sites of a Solid Catalyst Containing Only Reversible Adsorption Sites
V.1.2 Only Irreversible Adsorption

If a fraction $k$ of the total sites $C_o$ gives rise to irreversible adsorption, then as long as there is glucose remaining in solution, the amount of glucose adsorbed is proportional to the amount of catalyst, i.e. $[GC] = k[C_o]$ and,

$$[G_{eq}] = [C_o] - k[C_o]$$  \hspace{1cm} (25)

Therefore a plot of $[G_{eq}]$ versus $[C_o]$ would be linear and a graph of $\frac{1}{[G_{eq}]}$ versus $C_o$ will be a rectangular hyperbola. A plot of $\frac{1}{[G_{eq}]}$ versus $C_o$ for a system with only irreversible adsorption where $k = 0.01$ and $[G_o] = 1$ is given in Fig. 11(a). It is clear that irreversible adsorption gives rise to an upward curvature similar to that experimentally observed in Fig. 9.

Since alumina catalyzes the mutarotation of glucose, adsorption on the catalytic sites must be reversible. However, from Fig. 9, there appears to be irreversible sites as well. Therefore, as reported in the following section, the shape of the plot when a catalyst contains both reversible and irreversible sites was studied next.

V.1.3 Both Reversible and Irreversible Adsorption Sites Present

At all concentrations of the catalyst all the irreversible adsorption sites would form $GC$ complex as long as there is glucose remaining in solution. Hence the calculation for $G_{eq}$ can be simplified by considering first the adsorption on irreversible sites (using Equation (25)) and then adsorption on reversible sites. When all the irreversible sites are occupied the glucose concentration in solution is $[G] = [G_o] - k[C_o]$. The amount of glucose adsorbed on reversible adsorption sites can be calculated using the procedure described in Section V.1.1 with $[G]$ replacing $[G_o]$.

The theoretical plot of $\frac{1}{[G_{eq}]}$ versus $[C_o]$ for a system where
Fig. 11  Theoretical Plots of $1/[C_{eq}]$ versus $[C_o]$, the Concentration of Adsorption Sites of a Solid Catalyst Containing (a) Only Irreversible Adsorption Sites, (b) Both Irreversible and Reversible Adsorption Sites, and (c) Only Reversible Adsorption Sites.
$k = 0.01, \ K = 0.01$ and $[G_o] = 1$ is given in Fig. 11(b), and it is clear that it is similar to the experimental plot in Fig. 9.

A further refinement on possibilities is when there are two possible reversible adsorption sites as presented in the next section.

V.1.4 Both Weak and Strong Reversible Adsorption Sites Present

When two types of reversibly adsorbing sites are present we get the following equations for the two equilibria. To simplify calculations the concentrations of the two sites are considered to be equal.

\[
K_1 = \frac{[x]}{([C_o] - [x])([G_o] - [x] - [y])}
\]

\[
K_2 = \frac{[y]}{([C_o] - [y])([G_o] - [x] - [y])}
\]

where $[x]$ = concentration of glucose-catalyst complex on sites of Type 1, $[y]$ = concentration of glucose-catalyst complex on sites of Type 2, and $[C_o] = initial$ concentration of sites of Type 1 = initial concentration of sites of Type 2.

The two equations can be solved for $[x]$ and $[y]$, using a computer, and the equilibrium concentration of glucose $[G_{eq}] = [G_o] - [x] - [y]$ can be calculated. The theoretical plot for a system where $K_1 = 1, K_2 = 0.01$ and $[G_o] = 1$ is shown in Fig. 12. It is linear with a slope greater than that observed for reversible adsorption in Section V.1.1. For comparison, plots for systems with $K = 0.01, [G_o] = 1$, and $K = 1, [G_o] = 1$ are also included in Fig. 12.
Fig. 12  Theoretical Plots of $1/[G_{eq}]$ versus $[C_o]$, the Concentration of Adsorption Sites of a Solid Catalyst Containing (a) Both Weak and Strong Reversible Adsorption Sites, (b) Only Strong Adsorption Sites, and (c) Only Weak Reversible Adsorption Sites.

$K_1 = 1$

$K_2 = 0.01$ (a)

(b) $K_1 = 1$

(c) $K_2 = 0.01$

98.8% adsorption
Since, as given above, straight lines don't fit the experimental observations it should be concluded that alumina neutral contains both reversible and irreversible adsorption sites.

V.2 Determination of the Number of Irreversible Adsorption Sites per Gram of the Catalyst

From Equations (25) and (23) it is clear that as the catalyst concentration is increased at constant initial glucose concentration $[G_0]$, the irreversible sites will increasingly adsorb glucose and hence the equilibrium glucose concentration will decrease. Therefore at high alumina concentrations the irreversible sites will gain glucose from the reversible sites. Finally a stage will be reached when all the glucose molecules are occupying irreversible sites and there is none in solution. At the stage when no glucose can be detected in solution (i.e. when the DMSO solution shows zero optical rotation) there cannot be any glucose in reversible sites. Hence the number of irreversible sites should be equal to the number of glucose molecules in the original solution.

To determine the number of irreversible adsorption sites the experiment described at the beginning of Section V was continued using higher proportions of the catalyst. In order to make filtration easier a 60 ml 0.01 M $\alpha,\beta$ mixture was used instead of 0.05 M solution so that lower concentrations of the catalyst could be used. The optical rotations were measured using a 1 dm path length cell in order to obtain higher sensitivity. The results are given in Fig. 13, which shows that with 60 ml 0.01 M glucose solution zero optical rotation is reached when between 8.0 and 9.0 g of alumina is present. The most probable range is between 8.25 and 8.75 g alumina.
Fig. 13  
Plot of Equilibrium Optical Rotation versus Weight of Alumina Neutral for the Determination of the Number of Irreversible Adsorption Sites on Alumina Neutral
Hence the number of irreversible sites on 8.5 ± 0.25 g of alumina neutral is equal to the number of moles of glucose in 60 ml 0.01 M glucose solution which is 6.0 x 10^{-4} mole. Therefore the number of irreversible adsorption sites is (0.70 ± 0.02) x 10^{-4} mole/g.

V.3 Rate of Adsorption

The rate of adsorption of D-glucose onto alumina surface was studied by stirring a 60 ml 0.05 M solution of the α,β mixture with a pre-weighed sample of alumina, as described in the Experimental Section. The rate of change of optical rotation with time was followed until there was no further change. The procedure was repeated with different weights of alumina. The results given in Fig. 14, although not accurate enough to obtain individual rate constants, show that adsorption is complete in 15 to 20 minutes.

Comparison of these results with a typical first order plot for α-D-glucose mutarotation (Fig. 15) shows that the initial rapid decrease in optical rotation stops at about the same time the adsorption of glucose onto alumina surface is complete. In addition, the decrease in optical rotation when the α,β mixture was stirred with alumina was equal to the difference between the theoretical and experimental infinity values observed in a kinetic run, when the same weight of alumina was used. Hence it is clear that the initial rapid decrease in optical rotation observed in the kinetic runs is due to the adsorption of glucose on alumina surface and the rest of the plot represents the actual catalytic reaction. Therefore in all non-linear first order plots described in this thesis, the rate constants were determined by measuring the slope at $t = 25$ min when the adsorption is complete.
Fig. 14  Rate of Adsorption of Glucose on Alumina Neutral
Fig. 15  First Order Kinetic Plot (left and •) for α-D-Glucose Mutarotation by Alumina Neutral and Rate of Adsorption of Glucose (right and ×) on Alumina Neutral
VI DIFFUSION IN HETEROGENEOUS CATALYSIS
VI DIFFUSION IN HETEROGENEOUS CATALYSIS

As mentioned in Section V, before adsorption and surface reaction can take place, the substrate molecules must diffuse from the solution to the adsorption sites on the surface of the catalyst and the desorbed product must diffuse from the surface to the bulk solution.

\[
\begin{align*}
S & \xrightarrow{\text{diffusion}} S + C & \xrightarrow{\text{adsorption}} SC & \xrightarrow{\text{surface reaction}} PC & \xrightarrow{\text{desorption}} C + P & \xrightarrow{\text{diffusion}} P
\end{align*}
\]

For a heterogeneous catalytic system involving a porous catalyst like alumina (Section II.2), diffusion may be divided to external diffusion of the substrate from solution to the surface of the catalyst particle and internal diffusion from the surface of the catalyst particle through the pores to the adsorption site inside the pore\(^1\). If the catalyst is not rapidly agitated in solution a concentration gradient will be set up between the bulk solution and the catalyst surface, if the rates of adsorption, surface reaction, and desorption are faster than that of diffusion, and in such a case, the rate of diffusion can be rate limiting.

The concentration gradient set up outside the catalyst particles can be eliminated by rapidly stirring the slurry, which makes the solution homogeneous since the substrate is quickly brought to the surface and the desorbed product is quickly removed from the surface. If the rate of a reaction is diffusion controlled, then the rate would increase when the rate of stirring is increased (and hence the concentration gradient is decreased) and finally reaches a plateau value when the rate of the reaction is independent of the diffusion of substrate to the surface\(^1\).

In order to agitate the slurry efficiently, different stirrers, viz.
magnetic stirring bar, overhead Corning Vibrastir, overhead Fischer Dyna-Mix connected to screw type and dispersion stirrers, were employed. The dispersion stirrer was found to be the most efficient and the observed rate constant (determined at \( t = 25 \) min after the adsorption is complete), increased with the increase in rate of stirring and reached a maximum plateau value (Fig. 16). This optimum rate of stirring was used in all the kinetic and adsorption studies described in this thesis.

Although external diffusion can be eliminated by rapid agitation of the slurry, the rate of a reaction in a porous catalyst like alumina neutral might still be controlled by internal diffusion through the pores, since the volume of liquid inside the pores is not mixed by the mechanical stirrer present outside the pores.

To test whether internal diffusion is rate limiting, the method suggested by Boudart and Burwell\(^1\) was used. It consists of breaking down the catalyst particles by grinding to decrease the particle size and hence decrease the pore length. If internal diffusion is rate limiting in the unground catalyst particles, then the observed rate constant should be higher with the ground sample.

About 10 g of alumina neutral was ground well with a pestle and mortar for 15 mins. The particle size distributions of the ground and unground samples were determined with the Electrozone Celloscope. The alumina used was from a second bottle of "alumina neutral for tlc" and showed two maxima at approximately 11.5 microns and 24 microns (Fig. 17(a)). Comparison of the particle size distribution curves in Fig. 17 shows that, by grinding, the size of the peak corresponding to particles of larger diameter has decreased relative to the other peak indicating that the larger particles have been broken down to smaller ones. In addition, the
Fig. 16 The Effect of Rate of Stirring on the Observed Rate Constant

\[
[\alpha\text{-D-Glucose}] = 0.05 \text{ M}
\]

\[
[\text{Alumina}] = 13.3 \text{ mg.ml}^{-1}
\]
Fig. 17  Particle Size Distributions of (a) Unground and (b) Ground Samples of Alumina Neutral
log mean diameter has decreased from 11.2 microns to 10.4 microns, i.e. a 7% decrease in log mean diameter. The cumulative percentages greater than stated diameters given in Table VI also clearly show that the particle size has decreased on grinding. However, kinetic runs were superimposable and therefore no change in the rate constants was observed when equal weights of unground and ground samples were used with 60 ml 0.05 M α-D-glucose solution in DMSO. The rate of catalysis by alumina neutral is therefore not controlled by internal diffusion. External diffusion can be eliminated by rapid stirring as described above.

**TABLE VI**

**CUMULATIVE PERCENTAGES GREATER THAN STATED DIAMETERS FOR GROUND AND UNGROUND SAMPLES OF ALUMINA NEUTRAL**

<table>
<thead>
<tr>
<th>Diameter (microns)</th>
<th>Cumulative Percentage Greater than the Stated Diameter</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Ground</td>
</tr>
<tr>
<td>1</td>
<td>100</td>
</tr>
<tr>
<td>2.5</td>
<td>99.9</td>
</tr>
<tr>
<td>5</td>
<td>90.4</td>
</tr>
<tr>
<td>7.5</td>
<td>73.4</td>
</tr>
<tr>
<td>10</td>
<td>55.3</td>
</tr>
<tr>
<td>15</td>
<td>27.2</td>
</tr>
<tr>
<td>20</td>
<td>12.2</td>
</tr>
<tr>
<td>25</td>
<td>1.7</td>
</tr>
<tr>
<td>30</td>
<td>1</td>
</tr>
</tbody>
</table>
Since diffusion is not rate limiting when the heterogeneous catalytic system is stirred as described above, the observed rate constant should be a function of the rate constants for adsorption, desorption, both forward and reverse surface reactions, and the concentrations of catalyst, substrate and product. The rate equation for the surface catalyzed reaction, when diffusion is not rate limiting, is derived in the next section.
VII RATE EQUATION FOR THE HETEROGENEOUS CATALYTIC SYSTEM
VII RATE EQUATION FOR THE HETEROGENEOUS CATALYTIC SYSTEM

A simple heterogeneous catalytic reaction, which is not diffusion controlled, can be represented by the following equation,

\[
S + C \xrightleftharpoons[k_2]{k_1} SC \xrightleftharpoons[k_4]{k_3} PC \xrightleftharpoons[k_1]{k_2} C + P \quad (26)
\]

where \( S, P \) and \( C \) represent substrate, product, and catalyst, respectively, and \( SC \) and \( PC \) are the substrate-catalyst and product-catalyst complexes. \( k_3 \) and \( k_4 \) are the rate constants for the forward and reverse surface catalytic reactions. For adsorption involving hydrogen bonding and dipole-dipole interactions between the hydroxyl groups of glucose and the polar alumina surface, rate constants for adsorption \((k_1)\) and desorption \((k_2)\) of \( \alpha \)-D-glucose (substrate \( S \)) and \( \beta \)-D-glucose (product \( P \)) would be almost the same since the only difference between the two molecules is the configuration at the anomeric carbon.

The rate equation for the reaction in Equation 26 is derived by assuming that a heterogeneous catalytic system, consisting of a powdered solid catalyst dispersed evenly in a solution of a substrate and product, behaves similar to a homogeneous catalytic system. Then, the rate of the reaction, when the catalyst \( C \) and the complexes \( SC \) and \( PC \) are present in steady state, is given by:

\[
\frac{dp}{dt} = \frac{(k_1 k_2 k_3 s - k_1 k_2 k_4 p) c_0}{k_2^2 + k_2 k_4 + k_2 k_3 + k_1 (k_2 + k_3 + k_4) s + k_1 (k_2 + k_3 + k_4) p}
\]

where \( c_0 \) is the total catalyst concentration, and \( s \) and \( p \) are the substrate and product concentrations at time \( t \).
If \( s + p = s_0 \) the concentration of the substrate in solution when the adsorption is complete (\( \approx \) initial concentration of substrate if the amount adsorbed is small) then,

\[
\frac{dp}{dt} = \frac{k_1 k_2 k_3 c_0 s}{k_2^2 + k_2 k_4 + k_2 k_3 + k_1(k_2 + k_3 + k_4)s_0} + \frac{k_1 k_2 k_4 c_0 p}{k_2^2 + k_2 k_4 + k_2 k_3 + k_1(k_2 + k_3 + k_4)s_0}
\]

\[
= X_3 s + X_4 p
\]

where

\[
X_3 = \frac{k_1 k_2 k_3 c_0}{k_2^2 + k_2 k_4 + k_2 k_3 + k_1(k_2 + k_3 + k_4)s_0}
\]

and

\[
X_4 = \frac{k_1 k_2 k_4 c_0}{k_2^2 + k_2 k_4 + k_2 k_3 + k_1(k_2 + k_3 + k_4)s_0}
\]

are constants.

\[
\text{Therefore } \frac{dp}{dt} = X_3 (s_0 - p) - X_4 p \quad (27)
\]

Equation (27) is similar to Equation (8) for homogeneous catalysis discussed in the Introduction and becomes upon integration

\[
X_3 + X_4 = \frac{1}{t} \ln \left( \frac{p e}{p e - p} \right) \quad (28)
\]
where \( p_e \) = equilibrium concentration of the product. Equation (28) can be expressed in terms of optical rotation as,

\[
X_3 + X_4 = \frac{1}{t} \ln \frac{(a_o - a_\infty)}{(a_t - a_\infty)}
\]

where \( a_o \) is the initial rotation, \( a_t \) is the rotation at time \( t \), and \( a_\infty \) is the equilibrium rotation.

\[
\therefore \ln (a_t - a_\infty) = -(X_3 + X_4)t + \ln (a_o - a_\infty) \tag{29}
\]

Hence the system should show first order kinetics with observed rate constant \( k_{obs} = X_3 + X_4 \)

\[
k_{obs} = \frac{k_1 k_2 a_o (k_3 + k_4)}{k_2^2 + k_2 k_4 + k_2 k_3 + k_1 (k_2 + k_3 + k_4) s_o} \tag{30}
\]

A surface reaction involves breakage and formation of chemical bonds, while adsorption and desorption involve formation and breakage of dipole-dipole interactions and hydrogen bonds. Hence the rate constants for adsorption and desorption should be greater than those for the surface reaction, i.e. \( k_1, k_2 >> k_3, k_4 \), and the above equation simplifies to

\[
k_{obs} = \frac{k_1 (k_3 + k_4) a_o}{k_2 + k_1 s_o} \tag{31}
\]

\[
= \frac{(k_3 + k_4) a_o}{1/k + s_o} \tag{32}
\]
where \( K = k_1/k_2 \) is the equilibrium constant for adsorption of glucopyranose on alumina.

From the above discussion it is clear that the observed rate constant for a heterogeneous catalytic system involves the rate constants for adsorption and desorption steps \((k_1 \text{ and } k_2)\), and the total substrate concentration \( (s_0) \) in addition to the rate constants for the catalytic (surface) reaction \((k_3 \text{ and } k_4)\) and the catalyst concentration \( (c_0) \). This is a characteristic difference between a simple homogeneous catalytic system and a heterogeneous or any other system where there is initial complex formation (binding) between substrate and catalyst. Hence any inhibitory, activatory, or deuterium isotope effect, observed on the rate constant during mechanistic studies can result from the corresponding effects on the catalytic reaction or on the adsorption desorption process or both.

From Equation (32) it is also clear that the catalytic constant for the surface reaction \((k_3 + k_4)\) can be determined from the observed rate constant using equilibrium constant \( K \) for adsorption of glucopyranose on alumina. The relation between the equilibrium constant \( K = k_1/k_2 \) and the experimentally determined equilibrium constant is discussed in the next section.

VII.1 Relation between Equilibrium Constant \( K \) for Adsorption of Glucopyranose on Alumina and the Observed Equilibrium Constant

In Equation (26) for the catalytic reaction two surface complexes were considered, viz. \( SC \) and \( PC \) the \( \alpha \)-D-glucopyranose adsorbed on alumina and \( \beta \)-D-glucopyranose adsorbed on alumina, and they were considered to undergo direct interconversion on the surface. As mentioned in the Introduction the homogeneous catalysis of mutarotation occurs via the acyclic
intermediate whose concentration in water is only about 0.003%. Hence it is likely that the heterogeneous process also occurs via an acyclic (or possibly even carbonium ion) intermediate (Y) which may be stabilized by the surface. In order to accomodate the YC complex Equation (26) should be modified to,

\[ S + C \xrightleftharpoons[k_2]{k_1} SC \xrightarrow{PC} YC \xleftarrow{P + C} k_1 \]

(26')

In the adsorption studies mentioned in this thesis the amount adsorbed was measured by determining the initial and final optical rotations of solutions of α,β mixture in DMSO. Hence the amount lost from the solution should be present as SC, PC, and YC complexes, and the observed equilibrium constant \( K \) is equal to

\[ \frac{[SC + PC + YC]}{[S + P][C]} \]

(23')

which was given simply as \( \frac{[GC]}{[C]} \) in Equation (23) with \( GC = SC + PC + YC \). From this discussion it is clear that the actual adsorption process, given as a one step process in Equation (22), occurs via two consecutive steps at least for a small percentage of glucose molecules.

\[ C + S + P \leftrightarrow SC + PC \leftrightarrow YC \]

(22')

In solution the concentration of acyclic form is very small but it is possible that the acyclic form may be partially stabilized by the surface. However, most probably, it is still much less than the concentration of
pyranose forms on the surface. Thus Equation (23'). simplifies to,

\[
K = \frac{[SC + PC + XC]}{[S + P][C]} \approx \frac{[SC + PC]}{[S + P][C]} = \frac{k_1}{k_2}
\]

Hence the observed equilibrium constant should be very close to that for the adsorption of the pyranose forms from solution.

It has been shown in this section that glucose mutarotation according to the mechanism in Equation (26) should show first order kinetics (Equation (29)). However, as described in Section III, the first order plots for glucose mutarotation by alumina show curvature throughout the kinetic runs. The cause of this curvature in the first order plots is investigated in the next section.
VIII STUDY OF THE CURVATURE IN FIRST ORDER PLOTS
VIII STUDY OF THE CURVATURE IN FIRST ORDER PLOTS

As mentioned in Section III, the upward curvature in first order plots could be a result of progressive deactivation of alumina during catalysis or due to the fact that the catalytic reaction is not first order in α-D-glucose.

The mutarotation of α-D-glucose in solution is a well known first order reaction. There are a few examples where 2,3,4,6-tetra-O-acetyl-D-glucose, 2,3,4,6-tetra-O-methyl-D-glucose and D-glucose (in each case both α- and β-anomers) self-catalyzed the mutarotation of the respective α-D-glucose derivatives in pyridine. But, even in such cases, the reaction was first-order with respect to the corresponding α-D-glucose derivative. Hveding et al. have reported that the first-order plots for mutarotation of α-D-glucose in N,N-dimethylformamide, and in N,N-dimethylformamide-water mixtures where the mole fraction of water was less than about 0.7, showed significant deviation from a straight line. They stated that the mutarotation under such conditions is complex in analogy with that found for D-galactose in water. But for dimethyl sulfoxide and dimethyl sulfoxide-water mixtures, Robertson et al. found that the mutarotation is first-order in glucose. Hence it is more likely that alumina is undergoing progressive deactivation during catalysis.

The deactivation of the catalyst during catalysis can arise from strong adsorption on the active sites by the substrate itself (or even by the solvent), referred to as self-poisoning or coking, or it may be due to the formation of a side product which competes with the substrate for the active sites. The latter process is quite unlikely since the side products detected by glc analysis (in Section IV) are negligible and
also, the small percentage of α- and β-D-furanosides detected towards the end of a kinetic run are known to be formed at the beginning of the reaction.

To determine whether the curvature in the kinetic plots is due to deactivation of the catalyst, the catalytic activity of a sample of alumina which had been previously used to catalyze mutarotation was compared with that of a fresh sample of alumina under similar conditions as described below.

VIII.1 Tests for Deactivation of the Catalyst

The catalytic activity of a fresh sample of alumina neutral was determined by stirring a 60 ml solution which is 0.05 M in α-D-glucose and 0.05 M in α,β mixture with 1.6 g of the catalyst. The plot of \(\ln 10^3(\alpha_t - \alpha_0)\) versus time gave the "standard curve" shown in Fig. 18.

It was found that the mutarotation of 60 ml 0.05 M solution of α-D-glucose (contains 0.54 g α-D-glucose) by 1.6 g alumina is complete in about 6 hours. To test for deactivation of alumina during this period, a solution of 0.54 g α-D-glucose in 50 ml DMSO was stirred for 6 hours with 1.6 g alumina. After 6 hours, a 10 ml solution containing 0.54 g α-D-glucose in DMSO was added to the equilibrated solution and the change in optical rotation with time was followed. The results given in Fig. 18 show that the observed rate constant has decreased by about 50% after an initial 6.5 hours of reaction. As mentioned earlier this deactivation can arise from strong adsorption of glucose and/or solvent on the active site.

To determine whether DMSO poisons the catalyst, 1.6 g of alumina was stirred with 50 ml DMSO for 20 hours. A solution of 0.54 g α-D-glucose and 0.54 g α,β mixture in 10 ml DMSO was added and the change in optical rotation with time was followed. The results given in Fig. 18 show that alumina has
Deactivation by DMSO for 20 hrs
Deactivation during mutarotation for 6 hrs
Deactivation during mutarotation for 22 hrs
Deactivation during mutarotation for 70 hrs

Fig. 18 Deactivation of Alumina Neutral by DMSO and by Glucose During Catalysis of Mutarotation
been deactivated slightly by DMSO during the 20 hours, but this deactivation is negligible compared to deactivation during 6 hours of catalytic reaction with glucose. Moreover, alumina that has been used for 6 hours of catalysis underwent very little deactivation over the next 5 hours, and gave a linear first order plot over almost two half-lives, while alumina treated with DMSO gave rise to a "standard" non-linear plot.

To determine whether there is any further deactivation during mutarotation, two more experiments were carried out. In one, 10 ml $\alpha$-D-glucose solution was added after 22 hours of catalysis and in the other, after 70 hours of catalysis. Results in Fig. 18 show that alumina has undergone further deactivation although at a much slower rate, 55% after 22 hours and 70% after 70 hours. The reproducibility of the experiment carried out after 22 hours of catalysis was found to be very good. It is clear from Fig. 18 that alumina deactivated for 22 hours has given rise to a linear first order plot for more than two half-lives. Hence the catalysis of the mutarotation of $\alpha$-D-glucose by alumina is first order in $\alpha$-D-glucose, and the curvature of the first order plots is due to progressive deactivation of the catalyst.

VIII.2 Cause of Deactivation

Since it has been shown that the activity of the catalyst decreases during catalysis, it is of interest to determine the possible cause of deactivation. It has been mentioned earlier that side products may act as a catalyst poison. Alumina is a well known catalyst for dehydration of alcohols$^{14}$, and there are many reactions catalyzed by alumina, where removal of water from the surface increases the activity of the catalyst$^{23}$. Therefore we investigated the effect of water added to DMSO and onto alumina
on its catalytic activity.

Fifty ml of DMSO containing 0.054 ml (= 0.006 mole = the number of moles of D-glucose used for the deactivation studies in Section VIII.1) of distilled water was stirred with 1.6 g of alumina for one hour. A 10 ml solution of 0.54 g α-D-glucose and 0.54 g α,β mixture in DMSO was added and the change in the optical rotation with time was followed. Comparison of results, given in Fig. 19, with the "standard curve" shows that alumina has been only very slightly deactivated compared with deactivation by glucose during 6 hours of catalysis. Since the deactivation can be due to DMSO, the experiment was repeated after pretreating the alumina with DMSO containing water for 22 hours. Results in Fig. 19 show that the alumina has been deactivated further, slightly more than deactivation by DMSO during the same period.

Since water added into DMSO had very little effect on the activity, we investigated the effect of water added directly onto alumina on its activity. Distilled water (0.054 ml) was added onto 1.6 g of alumina held in a 100 ml 3-neck flask. The flask was stoppered and the alumina was mixed slowly, by rotation of the flask, for 60 minutes. A 60 ml solution of 0.54 g α-D-glucose and 0.54 g α,β mixture in DMSO was added and the change in optical rotation with time was followed. The results given in Fig. 19 show that alumina has been only slightly deactivated by direct contact with water and that the extent of deactivation is almost the same as the deactivation caused by water in DMSO over the same period.

These results show conclusively that the deactivation of alumina during mutarotation catalysis is not due to water produced by dehydration of glucose by alumina. As mentioned earlier, the deactivation of the catalyst could arise from the strong adsorption, on the active sites, of the
Fig. 19 The Effect of Water on the Catalytic Activity of Alumina Neutral
reactant itself or due to the formation of a side product which competes
with the substrate for the active sites. That is, the inhibiting substance
could be present on the surface or in solution. However, the results of
product analysis (in Section IV) showed only a small percentage of side-
products. To confirm that deactivation of alumina is due to strong adsorption
of substrate molecules onto the active sites and not due to side products
having an inhibitory effect, which might not have been detected by glc,
the following experiment, similar to that described by Dirkx and Baan 120,
was performed.

Two g of alumina was stirred with a solution of 0.54 g α-D-glucose
in 50 ml DMSO for 24 hours. The slurry was filtered, and the alumina was
washed with 20 ml DMSO as described in the Experimental Section. The "used"
alumina was dried in vacuum at room temperature for 24 hours. Thus 70 ml of
"used" glucose (equilibrated) solution and about 2 g of "used" alumina were
obtained.

Five g of alumina was treated with 100 ml of DMSO for 24 hours. The
slurry was filtered as above and the alumina was dried in vacuum at room
temperature for 24 hours. Thus about 5 g of "fresh" alumina (DMSO has very
little effect on the activity of alumina) was obtained.

The control run was carried out by stirring 1.6 g "fresh" alumina
with a solution of 0.54 g of α-D-glucose and 0.54 g of α,β mixture in 80 ml
DMSO. The first order kinetic plot obtained is given in Fig. 20 and will be
used as a control run to determine whether "used" alumina has been
deactivated and whether "used" glucose solution causes deactivation of
fresh alumina.

The first order kinetic plot when 1.6 g "fresh" alumina was stirred
with, 70 ml "used" glucose solution and 0.54 g of α-D-glucose in 10 ml DMSO
Fig. 20 Tests for the Presence of Inhibitory Products in Solution and on the Surface of the Catalyst

[α-D-Glucose] = 0.05 M
[α,β Mixture] = 0.05 M
[Alumina] = 26.7 mg.ml$^{-1}$
is also shown in Fig. 20. The activity of the alumina is about the same as that of the control and hence, "used" glucose solution does not contain any inhibitors.

However, the kinetic plot when 1.6 g "used" alumina is stirred with a solution of 0.54 g \( \alpha \)-D-glucose and 0.54 g \( \alpha,\beta \) mixture in 80 ml DMSO shows that the activity of the alumina has decreased 42% during 24 hours of catalysis. In addition, it gave rise to a linear first order plot as in Section VIII.1.

For comparison the first order plot obtained when 1.6 g alumina which had not been pre-treated with DMSO but evacuated at room temperature for 24 hours, was stirred with 0.54 g \( \alpha \)-D-glucose and 0.54 g \( \alpha,\beta \) mixture is given in Fig. 20. Its activity is higher than that of alumina which had been pre-treated with DMSO and dried in vacuum. The decrease in activity of the "fresh" alumina could be due to removal of water on the catalyst surface (discussed later in Section XIII) by DMSO, or due to the presence of some DMSO on alumina which would decrease the percentage of alumina in a given weight of the sample.

The results discussed in this section have shown that the glucose mutarotation by alumina is a first order reaction and the curvature in first order plots is due to progressive deactivation of the catalyst caused by strong adsorption of glucose on active sites. Hence the rate constant obtained at \( t = 25 \) min should be the actual initial rate constant after surface adsorption is complete but before much catalyst deactivation has occurred. Therefore, the observed rate constant should be equal to the first order rate constant derived in Section VII (Equation (32)).
\[ k_{\text{obs}} = \frac{(k_3 + k_4) c_0}{1/K + s_o} \] (32)

According to this equation, \( k_{\text{obs}} \) should be directly proportional to the total catalyst concentration \( c_0 \), while a plot of \( 1/k_{\text{obs}} \) versus substrate concentration \( s_o \) should be linear. These relationships are tested in the next two sections.
IX  THE EFFECT OF CATALYST CONCENTRATION ON THE

OBSERVED RATE CONSTANT
XI THE EFFECT OF CATALYST CONCENTRATION ON THE OBSERVED RATE CONSTANT

It was shown in Section VII, that the observed rate constant for the catalytic reaction is given by,

\[ k_{\text{obs}} = \frac{(k_3 + k_4) c_o}{1/k + s_o} \]  \hspace{1cm} (32)

and hence, it is directly proportional to the total catalyst concentration \( c_o \) when the substrate concentration \( s_o \) is constant. As suggested in Section VIII, the rate constant obtained at \( t = 25 \) mins gives the actual rate constant for the first order reaction, and it should be directly proportional to the weight of alumina used.

First order plots were obtained by stirring 60 ml 0.05 M \( \alpha \)-D-glucose solutions with different weights of alumina. The rate constants at \( t = 25 \) mins were obtained and are plotted against the weight of catalyst in Fig. 21. Results clearly show that at constant substrate concentration \( s_o \), the rate constant \( k_{\text{obs}} \) is directly proportional to the weight of alumina used. Further, the slope is equal to \( \frac{k_3 + k_4}{1/K + s_o} \), and can be used to determine the catalytic constant for the surface reaction \( (k_3 + k_4) \) when the equilibrium constant for adsorption \( K \) and the number of reversible adsorption sites on a gram of the catalyst are determined from the adsorption isotherm in Section XV.
[α-D-Glucose] = 0.05 M

Slope = $1.14 \times 10^{-5}$ sec$^{-1}$ mg$^{-1}$ ml

Fig. 21  The Effect of Catalyst Concentration on the Observed Rate Constant
THE EFFECT OF SUBSTRATE CONCENTRATION ON THE OBSERVED RATE CONSTANT
According to Equation (32), at constant catalyst concentration $c_0$, the observed rate constant should decrease with the increase of glucose concentration $s_0$. A graph of $k_{obs}$ versus concentration of glucose should be a curve while a graph of $1/k_{obs}$ versus glucose concentration should be linear with slope $= \frac{1}{(k_3 + k_4)c_0}$ and intercept $= \frac{1}{K(k_3 + k_4)c_0}$

$$k_{obs} = \frac{(k_3 + k_4)c_0}{1/K + s_0} \quad (32)$$

$$1/k_{obs} = \frac{1}{K(k_3 + k_4)c_0} + \frac{s_0}{(k_3 + k_4)c_0} \quad (33)$$

The rate constants were obtained as before using 1.6 g of catalyst and different concentrations of glucose. The rate constants at high glucose concentrations (0.4 and 0.8 M) were corrected for slow homogeneous catalysis. The results given in Figs. 22 and 23 agree well with the predictions of Equations (32) and (33). Again, from the slope of Fig. 23, the catalytic constant $(k_3 + k_4)$ can be determined when the number of adsorption sites is determined from the adsorption isotherm. However, the intercept is too close to zero to be of any use.
Fig. 22 The Effect of $\alpha$-D-Glucose Concentration on the Observed Rate Constant

$[\text{Alumina}] = 26.7 \text{ mg.ml}^{-1}$
Fig. 23  Relation of (Observed Rate Constant)$^{-1}$ to the Concentration of $\alpha$-D-Glucose

Slope = $7.8 \times 10^4$ sec.mole$^{-1}$litre

[Alumina] = 26.7 mg.ml$^{-1}$
XI  REPRODUCIBILITY OF THE RESULTS AND COMPARISON

OF CATALYTIC ACTIVITIES OF DIFFERENT ALUMINAS
XI REPRODUCIBILITY OF THE RESULTS AND COMPARISON
OF CATALYTIC ACTIVITIES OF DIFFERENT ALUMINAS

As mentioned in the Introduction (Section 1.1.3) small amounts of
impurities can change the activity of the catalyst. Small amounts of
water adsorbed on the alumina, when exposed to the atmosphere or from the
solvent during a kinetic run, can change its surface active sites, specially
Lewis acid sites. In fact, reproducibility of the results is one of
the main disadvantages of many heterogeneous catalytic systems when
compared with homogeneous systems.

XI.1 Effect of Drying and Distilling DMSO on the Catalytic Activity

To check the effect of any impurities present in DMSO on the catalytic
activity, catalysis was carried out in undistilled DMSO, in DMSO that has
been dried and distilled as described in the Experimental Section, and
also with dried and distilled DMSO which has been treated with neutral
alumina activity super I (a highly dehydrated form of aluminum oxide, see
Section XI.3). The results in Fig. 24 show that there is very little
difference in activity between the three runs. However, in all kinetic
results described in this thesis, DMSO which has been dried and distilled
over CaH₂ and stored, over neutral alumina activity super I, under nitrogen
was used because of the need for anhydrous DMSO for certain experiments.

XI.2 Reproducibility of Results with Alumina Neutral for TLC

The reproducibility of kinetics with the same batch of alumina neutral
was checked and the following results were obtained. The first order plots
obtained within two consecutive days were superimposable. The reproducibility
Fig. 24  The Effect of Drying and Distilling of DMSO on the Catalytic Activity of Alumina Neutral
after about 3 weeks was still very good as shown in Fig. 25. After one year the activity of alumina decreased only by 4%.

A different batch of alumina neutral showed an activity which is only 20% lower than that of the first batch. These results show that one can obtain excellent reproducibility with Woelm alumina neutral and that different batches show activities which are of the same order.

XI.3 Comparison of Catalytic Activities of Different Neutral Aluminas

It is of interest to compare activities of aluminas for column chromatography with those of the tlc aluminas used in this project. Hence, the first order plots for alumina neutral activity super I (used by Posner in his studies of alumina promoted organic reactions) and alumina neutral activity I were determined and are given in Fig. 25. The activities of alumina neutral activity super I and activity I are only 18% and 22%, respectively, of the first batch of alumina neutral, and the percentage of glucose adsorbed under similar conditions (0.8 g catalyst used with 60 ml 0.05 M glucose solution) was about 8% compared with 10% adsorption with alumina neutral. However, according to the manufacturer (Woelm Pharma) the BET surface area of alumina neutral activity super I is 200 m$^2$/g compared with 140 m$^2$/g for alumina neutral. The nature of the exposed surface and the pore size distribution may be responsible for the decrease in glucose adsorption by alumina activity super I with higher surface area.

On the other hand α-alumina (manufactured by Alfa Products; contains 90% Al$_2$O$_3$ and 9% H$_2$O; surface area 320 m$^2$/g) showed high activity while Puratronic aluminum oxide (distributed by Alfa Products; 99.999% metals

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Catalytic activities of acidic, basic and neutral aluminas for tlc will be compared in Section XII
\[ [\alpha-D-Glucose] = 0.05 \text{ M} \]

\[ [\text{Alumina}] = 13.3 \text{ mg.ml}^{-1} \]

- and \( \times \) change in activity of alumina neutral (first batch) for tlc in 3 weeks
- Change in activity after 1 year
- Activity of the second batch of alumina neutral for tlc

Fig. 25 Reproducibility of the Kinetics with Alumina Neutral and Comparison of the Catalytic Activities of Different Chromatographic Aluminas (Neutral)
Fig. 26  Comparison of the Catalytic Activities of Some Non-Chromatographic Aluminas
basis) had very low activity (Fig. 26). A sample of pure alumina prepared by hydrolysis of aluminum isopropoxide (as described in the Experimental Section) showed high activity, probably because of high surface area (percentage adsorption 28% compared with 14% adsorption by alumina neutral under similar conditions) and the presence of highly active catalytic sites on the surface.

The low activities observed with Puratronic aluminum oxide, alumina neutral activity super I and alumina neutral activity I may be due to the highly dehydroxylated nature of the surface. This is supported by the fact that the activity of alumina activity V prepared by adding 19% water to alumina activity super I is about 2.5 times the activity of the latter (Fig. 25), and also by the studies of the effect of heat on the catalytic activity of tlc aluminas described in the next section.

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* According to the manufacturer, the water lost by alumina activity super I on ignition at 1000°C is 1% by weight.
EFFECT OF DEHYDRATION ON THE CATALYTIC ACTIVITY OF ALUMINA
XII  EFFECT OF DEHYDRATION ON THE CATALYTIC ACTIVITY OF ALUMINA

It was mentioned in the Introduction (Section I.1.3) that on heating a sample of alumina which has water molecules adsorbed on the surface, first, some of the water molecules are desorbed and some react with the surface forming OH groups. As the sample is heated further, acidic and basic hydroxyl groups eliminate water to form Lewis acid and basic sites. When the sample is heated above 300° to 400°C, defect sites, which are catalytically more active, are formed (Section I.1.4). As the temperature is increased further, migration of ions tend to decrease the strength of the acidic and basic sites and at temperatures >1100°C α-alumina, which is normally considered to be catalytically inactive, is formed.

Hence the change in activity per unit surface area as the sample is heated can give information on the nature of the active sites. For example, if only Lewis acid sites or/and basic oxide ions are catalytically active, an increase in surface activity per unit area will be observed, while there should be a decrease in activity if only Brönsted acid sites or/and basic hydroxyl ions are active. On the other hand, because different groups can have different specific activities, if several different types of surface groups are active in the mutarotation it would be difficult to predict the change in activity caused by heating.

To determine the effect of dehydration, alumina neutral was heated under different conditions and the loss in weight was determined. Conditions used to dehydrate alumina, the loss in weight and the surface areas of the resulting samples (from Section II.1) are tabulated in the first three columns of Table VII. Activities of the samples were determined as described
<table>
<thead>
<tr>
<th>Dehydration Conditions</th>
<th>Percentage Loss in Weight</th>
<th>Surface Area m²/g</th>
<th>Rate Constant $^{*}$ $10^4$ sec$^{-1}$ g$^{-1}$</th>
<th>Rate Constant per Unit Area $10^6$ sec$^{-1}$ m$^{-2}$</th>
<th>Percentage Glucose Adsorbed</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 None</td>
<td>-</td>
<td>140</td>
<td>$1.9 \pm 0.1$</td>
<td>1.4</td>
<td>14</td>
</tr>
<tr>
<td>2 Room Temp (24°C), 0.01 mm press over $\text{P}_2\text{O}_5$ 4 days</td>
<td>2.8</td>
<td>144$^+$</td>
<td>$1.4 \pm 0.2$</td>
<td>1.0</td>
<td>14</td>
</tr>
<tr>
<td>3 150 ± 5°C 0.01 mm press 2 days</td>
<td>5.7</td>
<td>130</td>
<td>$0.9 \pm 0.1$</td>
<td>0.7</td>
<td>14</td>
</tr>
<tr>
<td>4 600 ± 50°C under dry $\text{N}_2$ 4 hours</td>
<td>6.2</td>
<td>116</td>
<td>$0.5 \pm 0.05$</td>
<td>0.43</td>
<td>13</td>
</tr>
<tr>
<td>5 800 ± 50°C under dry $\text{N}_2$ 4 hours</td>
<td>6.2</td>
<td>100</td>
<td>$0.7 \pm 0.1$</td>
<td>0.7</td>
<td>10</td>
</tr>
<tr>
<td>6 1100 ± 50°C under dry $\text{N}_2$ 3 hours</td>
<td>7.2</td>
<td>14</td>
<td>$0.54 \pm 0.02$</td>
<td>3.9</td>
<td>3.5</td>
</tr>
<tr>
<td>7 1100 ± 50°C under dry $\text{N}_2$ 6 hours</td>
<td></td>
<td></td>
<td></td>
<td>1.53 $\pm 0.05$</td>
<td>3.5</td>
</tr>
<tr>
<td>8 1250 ± 50°C under dry $\text{N}_2$ 6 hours</td>
<td>7.9</td>
<td>6.2</td>
<td>$2.20 \pm 0.05$</td>
<td>36</td>
<td>3</td>
</tr>
</tbody>
</table>

$^*$ First order rate constant calculated for 60 ml 0.05 M $\alpha$-D-glucose solution with 1.6 g alumina using natural logarithms. Errors were estimated from the maximum and minimum slopes.

$^+$ Estimated from surface area of alumina neutral and its loss of weight on evacuation.
earlier using 1.6 g samples of the catalyst and 60 ml 0.05 M solutions of α-D-glucose in DMSO (Fig. 27). The final optical rotations (\(\alpha_{\infty}\)) for samples with low activity were determined by adding a few drops of 0.1 N NaOH or n-butylamine, or by stirring the catalyst with an equilibrated glucose solution. The activities per gram and activities per unit area of the catalyst (in the 60 ml 0.05 M glucose solution), and the percentage of glucose adsorbed are given in columns 4, 5, and 6 of Table VII. The change in activity per unit area with temperature of dehydration is plotted in Fig. 28.

Results show that the catalytic activity decreases as the sample is heated up to 600°C. However, the main characteristics of the curved first order plots (initial adsorption followed by deactivation of the catalyst) have been retained. At 1100°C, when alumina has undergone extensive sintering and change in crystal structure to form α-alumina, the catalyst shows a complete change in its catalytic behaviour. With this sintered (1100°C) alumina the amount of glucose adsorbed has decreased as the surface area decreased and yet the catalyst showed activation rather than deactivation. In addition, the first order plots were linear over three half-lives and the catalytic activity increased further on heating to 1250°C. This increase in activity of the catalyst with decrease in surface area causes a rapid increase in specific activity (activity per unit area) as shown in Fig. 28. Such high activity for α-alumina (compared with γ-alumina) is virtually unknown. As mentioned in the Introduction (Section I.1.2), α-alumina is considered to be catalytically less active than γ-alumina and has shown activity towards only a few reactions.

Reproducibility of the results with alumina dehydrated at room temperature, 150°C, 600°C and 800°C was very good. Samples, when stored
Fig. 27 The Effect of Dehydration Temperature of Alumina Neutral on Its Catalytic Activity
Fig. 23 The Effect of Dehydration Temperature of Alumina Neutral on Its Catalytic Activity per Unit Area

[α-D-Glucose] = 0.05 M
under anhydrous conditions, did not show any change in activity over several weeks. By dehydrating alumina neutral under the same conditions new batches of dehydrated alumina with the same activity could be prepared. However, the reproducibility of kinetics with alumina dehydrated at 1100°C was not good. Activity of a single batch decreased with time even when stored under anhydrous conditions. Activity of different batches prepared under the same conditions varied as much as 20%. Mixing a batch in a minimill for several hours did not improve the results. Activity of alumina dehydrated at 1250°C showed much better reproducibility. New batches with almost the same activity could be prepared but it also decreased in activity, over several days, when stored under anhydrous conditions. Reproducibility of kinetic runs with alumina sintered at 1250°C was very good when carried out within a few hours of each other. Hence alumina dehydrated for 6 hours at 1250°C was used, where necessary, in the mechanistic studies described in this thesis.

Absence of deactivation in alumina heated to 1100°C and higher temperatures appears, at first, to be related to the lack of porosity in the sintered alumina (see Section II.2 and Appendix A.2). However, it is unlikely that the presence of pores in low temperature aluminas (dehydrated at 800°C or less) is responsible for deactivation by blocking of the pores by substrate molecules. The dimensions of the pores observed in standard alumina neutral (most between \( r_p = 18 \) and 27 Å) and the pores in aluminas heated to 800°C (most between \( r_p = 22 \) and 31 Å) are both much larger than the dimensions of a glucose molecule (radius of pyranose form \( \approx 4.5 \) Å). It is more likely that the deactivation is caused by the presence of strongly adsorbing sites on the surface which were observed on alumina neutral during adsorption studies described in Section V, and will be referred to again in
Section XIV on Adsorption Isotherms. The presence of catalytic sites which can be converted to inactive or less active sites, when glucose molecules are allowed to interact with low temperature aluminas for a sufficiently long time, is probably due to the presence of reactive functional groups at the catalytic sites causing the glucose molecules to undergo slow but irreversible reaction (e.g. ether formation) with the surface.

XII.1 Cause of High Catalytic Activity of α-Alumina

Examination of Fig. 28 shows that the specific activity of alumina has increased rapidly when the pyrolysis temperature was increased over about 800°C. The Tammann temperature for alumina is about 870°C. Above the Tammann temperature ionic diffusion (volume diffusion) begins to occur at an appreciable rate, causing a rapid increase in the plasticity of the solid. This results in a marked acceleration of the rate of sintering as observed in Table VII. The volume diffusion above the Tammann temperature can also result in a change in the composition of the crystal surface depending on the impurities present in alumina. As mentioned in Section II.5, there are about 0.2% of cationic impurities in alumina neutral used in these studies and hence, a redistribution of ions may be responsible for the high catalytic activity of the α-alumina produced by pyrolysis.

To check whether this increase in activity towards mutarotation on pyrolysis at high temperature is characteristic of all neutral aluminas, the same experiments were performed with a second batch of Woelm alumina neutral, purchased from ICN Pharmaceuticals. This alumina was progressively deactivated*, rather than activated, on further heating above 800°C (Fig. 29).

* However, since α-alumina produced at 1250°C has a low surface area per unit weight the specific activity of the sintered alumina may be higher than that of the original alumina sample.
The same behaviour was observed with alumina neutral for tlc (aluminim oxide neutral Type T) purchased from BDH Chemicals. Examination of the composition of samples, given in the Experimental Section, and the pH's of the 10% slurries in distilled water, given in Table VIII, does not indicate the cause of the increase in activity observed with the first batch of alumina neutral.

**TABLE VIII**

**pH's OF 10% SLURRIES IN WATER**

<table>
<thead>
<tr>
<th>Alumina Sample</th>
<th>Original Sample</th>
<th>Sintered (1250°C, 6 hrs) Sample</th>
</tr>
</thead>
<tbody>
<tr>
<td>Neutral Woelm first batch</td>
<td>7.5</td>
<td>10.1</td>
</tr>
<tr>
<td>Neutral Woelm second batch</td>
<td>7.5</td>
<td>9.3</td>
</tr>
<tr>
<td>Neutral BDH</td>
<td>7.6</td>
<td></td>
</tr>
<tr>
<td>Acidic Woelm</td>
<td>4.5</td>
<td>9.2</td>
</tr>
<tr>
<td>Basic Woelm</td>
<td>9.8</td>
<td>10.3</td>
</tr>
</tbody>
</table>

Therefore, the catalytic activities of both acidic and basic aluminas (Woelm) for tlc, purchased from ICN Pharmaceuticals, were examined to determine the cause of this difference in behaviour. As shown in Fig. 29 the activity of alumina basic is close to that of the first batch of alumina neutral, while that of alumina acidic is only about 30% of the activity of
Fig. 29 Comparison of the Catalytic Activities of Acidic, Basic, and Neutral Aluminas (for TLC) and the Effect of Dehydration at 1250°C for 6 hrs on Their Catalytic Activities
alumina neutral. Both acidic and basic aluminas underwent deactivation during catalysis and adsorbed 13% and 16% of the glucose in solution (1.6 g catalyst used with 60 ml 0.05 M glucose solution). When they were sintered at 1250°C, the acidic and basic aluminas showed a marked difference in catalytic activity. Alumina acidic showed only a slight increase (≈10%) in activity on pyrolysis at 1250°C and it did not undergo deactivation and adsorbed only 3% glucose like other sintered aluminas. However, alumina basic became highly active although it also adsorbed only 3% glucose from solution. In fact its activity was too high to measure accurately at normal concentrations (26.7 mg/ml) of aluminum oxide (Fig. 29). When kinetics was carried out using 2.7 mg/ml of alumina, the catalyst showed an increase in activity with time and the initial rate constant for the same weight of sintered catalyst was about 30 times that for alumina basic for tlc. It is interesting to note that pH's of slurries of all aluminas are higher after sintering at 1250°C. This indicates an increase in surface basicity on pyrolysis but does not explain why some aluminas become activated while others become inactive on sintering.

The increase in activity observed with alumina basic when pyrolysed at 1250°C suggests that the somewhat similar behaviour observed with the first batch of alumina neutral for tlc may be due to the presence of residual basic character in that batch of neutral alumina. Hence it seemed possible to prepare a sample of alumina which would behave similar to the first batch of alumina neutral by treating the second batch of alumina neutral with a base, or by partial neutralization of alumina basic with an acid. Preparation of a batch of alumina which behaves similar to the first batch of alumina neutral by treating alumina basic with an acid is described in Appendix B.
The above results and discussion show that the first batch of alumina neutral used in this work is not a typical alumina neutral but it appears to possess some basic character which is brought out by pyrolysis. In any case, as described above, highly active sintered alumina can be prepared from alumina basic. It was pointed out in the Introduction (Section I.2.3, Table III) that bases are very efficient catalysts of the mutarotation reaction. The fact that high activity is observed with sintered alumina prepared from basic alumina or partially neutralized basic alumina but not with sintered alumina prepared from neutral or acidic alumina suggests that the high activity per unit area of sintered basic aluminas may be due to predominance of basic sites at temperatures above the Tammann temperature. It would increase the catalytic constant for surface reaction \((k_3 + k_4)\) in Equation (32) (this will be discussed further in Section XVIII.1). The observed rate constant would also be affected by changes in Equilibrium constant for adsorption, as will be discussed in Section XIV, where the adsorption isotherms are presented.

XII.2 Some Observations on the Nature of Active Sites

It was pointed out at the beginning of this section that dehydration causes, initially, the loss of water adsorbed on the surface and later, the dehydroxylation of acidic and basic hydroxyl groups. From Fig. 28 it is clear that initial dehydration has caused a decrease in the specific activity of all the aluminas towards glucose mutarotation. This effect of mild thermal treatment is unlike other reactions where catalytic activity is observed only on heating above 300 to 400°C (giving rise to defect sites). Hence it is clear that defect sites (consisting of clusters of vacancies and neighboring oxide ions), even if they possess catalytic activity
towards mutarotation, do not possess a monopoly over the catalytic activity as has been observed with many other reactions. Further, the presence of catalytic activity on aluminas containing a large amount of adsorbed water molecules indicate that Lewis acid sites are not essential for catalytic activity (this will be discussed further in Section XVIII.2.1). These results emphasize the fact that glucose mutarotation differs from all the reactions previously studied on alumina as mentioned in the Introduction. This difference in behavior should necessarily be related to the high sensitivity of the mutarotation reaction to acidic and basic sites on the surface.

It was shown earlier in this section (Table VII) that evacuation of alumina at room temperature, causing the loss of loosely bound water molecules without any change in the nature of the surface, has decreased the catalytic activity per unit area. Hence it is clear that water plays an active role in the catalytic process. The effect of water on the catalytic activity will be investigated in the following section.
XIII EFFECT OF WATER ON THE CATALYTIC ACTIVITY
XIII EFFECT OF WATER ON THE CATALYTIC ACTIVITY

XIII.1 Effect of Water on Alumina Evacuated at Room Temperature

It was pointed out in the previous section that the removal of adsorbed water by evacuation at room temperature under $P_2O_5$ has caused a decrease in the catalytic activity. To check whether the decrease in activity is reversible, 3% water ($\approx$ loss in weight) was added to a sample of evacuated alumina held in an erlenmeyer flask, stoppered, mixed well and was allowed to stand overnight at room temperature. Kinetics was carried out as described before using 0.8 g of the hydrated sample (Fig. 30). It is clear that the addition of water, lost from alumina by mild heating has increased its activity to the original value. Hence, addition and loss of water on $\gamma$-alumina at room temperature is a reversible process.

XIII.2 Effect of Water on Alumina Dried at 150°C

Dehydration at higher temperatures causes some water molecules to react with the surface and surface hydroxyl groups combine eliminating water as described in Section I.1.3. Further, surface areas decrease (see Section II.1) and the crystal structure begins to change (Sections I.1.1 and II.4). Hence the change should be irreversible when rehydration is attempted at room temperature.

Experiments were conducted by adding water as described above, to alumina dried at 150°C, for 3 days. Results given in Fig. 31 show that the activity is increased by addition of water, but the activity reaches only about 56% of the original value (i.e. the activity of alumina neutral) by the addition of 5% water which was originally lost from the alumina. Hence the hydrated 150°C alumina is still less active than standard alumina
Fig. 30  The Effect of Water on Alumina Dried at Room Temperature

$[\alpha-D-Glucose] = 0.05 \text{ M}$

$[Alumina] = 13.3 \text{ mg.ml}^{-1}$
Fig. 31  The Effect of Water on Alumina Dried at 150°C
neutral. Addition of excess water (7%) appears to keep the initial activity constant, but slows down the rate of deactivation (Fig. 31).

XIII.2.1 Cause of the Increase in Activity by Water

There may be many causes for the increase in activity by water added onto dehydrated aluminas. It may be due to an independent catalytic process by water adsorbed on the surface. Water might also act as a catalyst promoter and increase the activity of catalytic sites (i.e. increase $(k_3 + k_4)$) or/and activate catalytically inactive sites (i.e. increase $\theta_0$). It might easily form hydrogen bonds with glucose and thus, might increase the amount of glucose adsorbed on the surface (i.e. increase equilibrium constant $K$).

To determine how water increases the catalytic activity, experiments were carried out by adding water to a known weight of alumina dehydrated at 150°C (for 2 days). Water was added to the alumina in the usual reaction flask, the flask was then stoppered and rotated 2.5 hours to mix the solid with the water. The rate of mutarotation was followed after addition of 60 ml 0.05 M α-D-glucose solution to the reaction flask. The experiment was repeated by adding different amounts of water to the same weight of alumina (to keep the surface area constant). The results are given in Fig. 32 where the initial catalytic activity per unit weight of catalyst is plotted against the percentage (by weight) of water added. It shows that the catalytic activity has increased linearly with the increase of amount of water until about 3% by weight of water has been added. Further addition of water increased the rate only very slowly; the ratio of the slopes in the two regions being about 200 : 1.
Fig. 32  Relation of the Observed Rate Constant to the Amount of Water Added to Alumina Dehydrated at 150°C

[α-D-Glucose] = 0.05 M

[Alumina] = 26.7 mg. ml⁻¹
It was mentioned at the beginning of this section that the increase in activity by water can arise from the increase of the catalytic constant \((k_3 + k_4)\), the concentration of catalytically active sites \(c_o\), and the equilibrium constant for adsorption \(K\), and also from an independent catalytic process by water. As described in Section VII, the observed rate constant is related to \(K\), \(c_o\), and \((k_3 + k_4)\) by Equation (32).

\[
    k_{obs} = \frac{(k_3 + k_4)c_o}{1/K + s_o} \tag{32}
\]

\[
    = \left(\frac{1}{1/K + s_o}\right) \left[(k_3 + k_4)c_o\right] \tag{34}
\]

According to Equation (34), the observed rate constant is a product of two factors; one determined by the equilibrium constant for adsorption and the substrate concentration which is kept constant, and the other is the product of the catalytic constant for the surface reaction and the concentration of the catalytically active sites. Hence, the results can be analyzed in terms of the effect of water on the two factors.

XIII.2.1.1 Effect of Water on the Amount of Glucose Adsorbed

When the effect of water on the rate of reaction was studied (Section XIII.2), no change (within experimental error of about 5%) in the amount of glucose adsorbed was detected (Table IX), using a 0.1 dm path length cell, even though a 60% increase in the rate constant was observed. This means that there is no change in the equilibrium constant for
### TABLE IX

**EFFECT OF WATER ON EQUILIBRIUM OPTICAL ROTATION**

<table>
<thead>
<tr>
<th>Percentage (by weight) Water Added onto 150°C Alumina</th>
<th>Optical rotation (degrees) at $t = \infty$</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>0.124/0.125</td>
</tr>
<tr>
<td>1</td>
<td>0.124</td>
</tr>
<tr>
<td>2</td>
<td>0.124</td>
</tr>
<tr>
<td>3</td>
<td>0.124</td>
</tr>
<tr>
<td>5</td>
<td>0.124</td>
</tr>
<tr>
<td>7</td>
<td>0.124</td>
</tr>
<tr>
<td>15</td>
<td>0.124</td>
</tr>
</tbody>
</table>

adsorption and hence, the factor $(\frac{1}{1/K + \sigma_o})$ should be independent of the concentration of water on the surface. Therefore the observed increase in the rate constant with addition of water onto alumina should be due to an increase in the factor $(k_3 + k_4)\sigma_o$ or due to an independent catalytic process by water on the surface, as discussed below.
XIII.2.1.2. **Effect of Water on the Factor** \((k_3 + k_4)c_o\) **and the Expression for the Observed Rate Constant in the Presence of an Independent Catalytic Process due to Water**

Using Equations (13) and (14), the factor \((k_3 + k_4)c_o\) in Equation (32), can be expressed as a sum of several terms,

\[
(k_3 + k_4)c_o = k_A[A] + k_B[B] + k_{A.B}[A.B] 
\]  

(35)

where \(k_A, k_B,\) and \(k_{A.B}\) are the catalytic constants for acid, base, and bifunctional catalysis, respectively. \([A]\) and \([B]\) are the concentrations of acidic and basic sites, respectively, and \([A.B]\) is the concentration of acid/base pair sites capable of bifunctional catalysis of glucose mutarotation.

If there is an independent catalytic process by water (see Section 1.2.5.1.3.1), then the observed rate constant is given by,

\[
k_{o.b.s}^\text{H_2O} = k_{o.b.s} + k_{H_2O}[H_2O]^n
\]

(36)

where \(k_{o.b.s}\) is the observed rate constant in the absence of water.

On the other hand, if water acts as an acid or a base favouring bifunctional mechanism, then,

\[
(k_3 + k_4)c_o = k_A[A] + k_B[B] + k_{A.B}[A.B] + k'[A \text{ or } B,H_2O]
\]

(37)

In such a case, the presence of water on the alumina surface will change the catalytic constant for the surface reaction.
Another possibility in a surface catalyzed acid/base reaction is that water can act as a medium to transfer protons or hydroxide ions from acidic or basic sites on the surface to the adsorption sites. Thus, adsorption sites which are not catalytically active will become effective mutarotation catalytic sites on the addition of water. In such a case, Equation (35) would still apply with higher values for the concentration of acidic and basic sites.

For equations of type (36) the order (n) with respect to water has been found to be 2 or 3 from studies in non-aqueous solvents as mentioned in the Introduction (Section I.2.5.1.3.1). The observed linear increase in observed rate constant with increase in concentration of water shows that for the surface reaction \( n = 1 \) and hence, independent catalysis by water is not possible. This conclusion will be supported by deuterium isotope effect studies discussed later in Section XIX.

Hence water should increase the observed rate constant either by acting according to Equation (37) as one of the participants of a bifunctional system or by acting as a medium for proton or hydroxide ion transfer. Both of these mechanisms would show a linear increase in the observed rate constant with increase in the concentration of water. Again, the deuterium isotope effect studies discussed in Section XIX will be used to distinguish between the two possibilities.

XIII.3 Effect of Water on the Activity of Alumina Dehydrated at 800°C and 1250°C

Alumina dehydrated at 800°C was very sensitive to water and its activity increased more than 300% (see Section XIX) on treatment with water. Thus, hydration produced a catalyst which is more active than the
original alumina neutral.

However, alumina dehydrated at 1250°C gave irreproducible results on treatment with water. The activity of the catalyst always increased on treatment with water (freshly distilled) but the increase in activity appeared to depend on contact time as well as other unknown factors.

The results discussed in the previous sections have shown that there are two forms of alumina with characteristic properties towards glucose mutarotation. One is γ-alumina (porous) with low activity per unit area which undergoes deactivation during mutarotation and the other is low surface area α-alumina (non-porous) with high activity per unit area which shows constant activity during a kinetic run. This difference in behaviour should at least partly be determined by the surface area available for glucose adsorption and the strength of adsorption. These factors are investigated in the next section where the glucose adsorption isotherms are presented.
XIV ADSORPTION OF GLUCOSE ON ALUMINA SURFACE

PART II ADSORPTION ISOTHERMS

The first theoretical equation which describes the relationship between the amount of gas adsorbed and the equilibrium pressure of the gas at constant temperature was advanced by Langmuir\textsuperscript{123}. The basic assumption of the theory was that adsorption was limited to formation of a unimolecular layer.

Langmuir theory can also be applied to adsorption of non-electrolytes from solution, on the assumption that adsorption is essentially confined to a monolayer next to the surface\textsuperscript{113,124}. The assumption is valid if the solute molecules interact with the surface and not with each other so that bulk solution exists above the monolayer of solute on the surface.

In Section V.1.1 the adsorption of glucose onto the alumina surface was represented by the equation,

\[ G + C \underset{k_2'}{\overset{k_1}{\rightleftharpoons}} GC \]  

(22)

This is an oversimplification because the solvent molecules also get adsorbed on the surface and hence, the adsorption process is best considered as a competition between the solute and solvent molecules (Equation (38)). This leads to the formation of an ideal two dimensional solution of solute and solvent molecules on the surface\textsuperscript{113,124}.

\[
G + DC \overset{k_1}{\underset{k_2'}{\rightleftharpoons}} GC + D
\]

(38)

where $D$ is the solvent DMSO.


'.' Equilibrium constant \( K' = \frac{k_1}{k_2} = \frac{[GC][D]}{[G][DC]} \)

For a dilute solution \([D]\) is a constant and therefore,

\[
\frac{K'}{[D]} = \frac{k_1}{k_2'[D]} = \frac{[GC]}{[G][DC]} = K
\]  \hspace{1cm} (39)

where \( K \) is the observed equilibrium constant discussed in Sections V.1.1 and VII.1. Hence it follows that the first order rate constant \( k_2 \) mentioned in those sections is really a pseudo first order rate constant, for it is equal to the product of a second order rate constant and a concentration term which is a constant. From Equation (39) it also follows that, the greater the strength of adsorption the greater is the value of \([GC]\) for a given concentration of glucose and greater is the magnitude of the equilibrium constant.

If \([GC]_m\) is the maximum concentration of adsorbed glucose at a constant catalyst concentration and at constant temperature then,

\[
[GC]_m = [GC] + [DC]
\]

'.' \([DC] = [GC]_m - [GC]\)

and from Equation (39), \([GC] = K[G][DC] = K[G]([GC]_m - [GC])\)

\[
\therefore \frac{[GC]}{[GC]_m} = K[G] \left(1 - \frac{[GC]}{[GC]_m}\right)
\]
which is the familiar form of a Langmuir equation for adsorption of glucose onto alumina\textsuperscript{113,124}. Hence a plot of the moles of glucose adsorbed per gram of the catalyst versus the equilibrium concentration of glucose should show a linear increase at low glucose concentrations, with a slope proportional to $K$, and should show a constant maximum adsorption at high glucose concentrations. Furthermore, the smaller the value of $K$ (weak adsorption) the higher is the concentration of glucose at which the plateau is reached.

Equation (40) can be transformed to,

$$\frac{[GC]}{[G]} = \frac{K[G]}{[GC]_m} (40)$$

If $K$ is a constant then a graph of $\frac{[GC]}{[G]}$ versus $[GC]$ would be linear with slope $= -K$ and intercept $= K[GC]_m$. Hence, if $[GC]_m$ is known then $K$ can be determined from the intercept as well. Whether or not $K$ is a constant depends on the energy of the sites and on the lateral interactions of adsorbate molecules. In the ideal case all adsorption sites would be energetically homogeneous and adsorbate molecules would not interact with each other on the surface. Under such conditions a linear relationship would be observed for all values of $\theta$, the fraction of surface occupied by the solute. But most surfaces are energetically heterogeneous and there is lateral interaction of adsorbate molecules on the surface\textsuperscript{123}. However, a linear relationship is exhibited by many systems over certain values of $\theta$ when the two factors compensate each other and tend to keep
XIV.1 Adsorption Isotherm for Alumina Neutral

The procedure for obtaining the adsorption isotherm is described in detail in the Experimental Section. Hence only a brief account is given here. An equilibrated solution of α,β mixture in DMSO (60 ml) was stirred with 1.6 g alumina neutral at 25°C until the optical rotation of the filtered slurry was essentially constant. The number of moles of glucose adsorbed onto the alumina and the equilibrium concentration of glucose were calculated from the change in optical rotation of the solution and the final optical rotation, respectively, using the calibration curve described in the Experimental Section. The experiment was repeated with solutions of different concentrations and the results are given in Table X. The adsorption isotherm in Fig. 33 was obtained by plotting the amount of glucose adsorbed per gram catalyst against the equilibrium concentration of glucose.

XIV.2 Adsorption Isotherm for Sintered Alumina

To obtain the adsorption isotherm for alumina heated at 1250°C for 6 hours the same procedure was adopted, but with 3.2 g samples of alumina. The weight of catalyst used was doubled because of lower adsorption with sintered aluminas. The results are given in Table XI and the adsorption isotherm in Fig. 33.

The adsorption isotherms show that the amount of glucose adsorbed onto alumina increases with increase of concentration of glucose in solution.
<table>
<thead>
<tr>
<th>Initial Concentration of Glucose (M)</th>
<th>Optical Rotation (Degrees)</th>
<th>Equilibrium Concentration (M)</th>
<th>Change in Optical Rotation (Degrees)</th>
<th>Moles (x 10^4) Glucose Adsorbed per Gram Catalyst</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.0025</td>
<td>0.077</td>
<td>0.008</td>
<td>0.00029</td>
<td>0.070</td>
</tr>
<tr>
<td>0.005</td>
<td>0.146</td>
<td>0.049</td>
<td>0.0017</td>
<td>0.100</td>
</tr>
<tr>
<td>0.01</td>
<td>0.291</td>
<td>0.171</td>
<td>0.0059</td>
<td>0.120</td>
</tr>
<tr>
<td>0.025</td>
<td>0.720</td>
<td>0.571</td>
<td>0.0198</td>
<td>0.150</td>
</tr>
<tr>
<td>0.05</td>
<td>1.454</td>
<td>1.274</td>
<td>0.044</td>
<td>0.180</td>
</tr>
<tr>
<td>0.10</td>
<td>(a) 2.910</td>
<td>2.700</td>
<td>0.094</td>
<td>0.210</td>
</tr>
<tr>
<td></td>
<td>(b) 2.908</td>
<td>2.695</td>
<td>0.213</td>
<td></td>
</tr>
<tr>
<td></td>
<td>(c) 2.938</td>
<td>2.733</td>
<td>0.205</td>
<td></td>
</tr>
<tr>
<td>0.25</td>
<td>7.280</td>
<td>7.250</td>
<td>0.25</td>
<td>0.230</td>
</tr>
<tr>
<td>0.50</td>
<td>14.450</td>
<td>14.220</td>
<td>0.50</td>
<td>0.230</td>
</tr>
<tr>
<td>1.0</td>
<td>(a) 28.780</td>
<td>28.560</td>
<td>1.0</td>
<td>0.220</td>
</tr>
<tr>
<td></td>
<td>(b) 28.788</td>
<td>28.562</td>
<td>0.226</td>
<td></td>
</tr>
<tr>
<td>1.75</td>
<td>50.340</td>
<td>50.110</td>
<td>1.75</td>
<td>0.230</td>
</tr>
</tbody>
</table>
Fig. 33 Isotherms for Adsorption of Glucose on (a) Alumina Neutral and (b) Alumina Sintered at 1250°C for 6 hours.
<table>
<thead>
<tr>
<th>Initial Concentration of Glucose (M)</th>
<th>Optical Rotation (Degrees)</th>
<th>Equilibrium Concentration (M)</th>
<th>Change in Optical Rotation (Degrees)</th>
<th>Moles (x 10^4) Glucose Adsorbed per Gram of Catalyst</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Initial</td>
<td>Final</td>
<td></td>
<td></td>
</tr>
<tr>
<td>0.0005</td>
<td>0.015</td>
<td>0.006</td>
<td>0.00021</td>
<td>0.009</td>
</tr>
<tr>
<td>0.001</td>
<td>0.027</td>
<td>0.015</td>
<td>0.00052</td>
<td>0.012</td>
</tr>
<tr>
<td>0.005</td>
<td>0.139</td>
<td>0.122</td>
<td>0.0042</td>
<td>0.017</td>
</tr>
<tr>
<td>0.01</td>
<td>0.275</td>
<td>0.255</td>
<td>0.0088</td>
<td>0.020</td>
</tr>
<tr>
<td>0.24</td>
<td>0.696</td>
<td>0.673</td>
<td>0.0232</td>
<td>0.023</td>
</tr>
<tr>
<td>0.05</td>
<td>1.400</td>
<td>1.372</td>
<td>0.0473</td>
<td>0.028</td>
</tr>
<tr>
<td>0.10</td>
<td>2.768</td>
<td>2.732</td>
<td>0.0943</td>
<td>0.036</td>
</tr>
<tr>
<td>0.24</td>
<td>7.000</td>
<td>6.951</td>
<td>0.240</td>
<td>0.049</td>
</tr>
<tr>
<td>0.5</td>
<td>13.902</td>
<td>13.842</td>
<td>0.488</td>
<td>0.060</td>
</tr>
<tr>
<td>1.0</td>
<td>27.716</td>
<td>27.644</td>
<td>0.954</td>
<td>0.072</td>
</tr>
<tr>
<td>2.02</td>
<td>58.420</td>
<td>(a) 58.360</td>
<td>2.01</td>
<td>0.07 ± 0.01</td>
</tr>
</tbody>
</table>

(a) 58.340
until a plateau is reached. This increase should be due to reversible adsorption of glucose onto the alumina surface. The presence of reversible adsorption sites, discussed earlier in Section V, can be confirmed by releasing the adsorbed glucose into the solution. These tests for reversibility of adsorption are described in the Experimental Section. Further, from the plateau's of the adsorption isotherms, surface areas of the alumina samples can be estimated as described in the next section.

XIV.3 Maximum Amount of Glucose Adsorbed and the Surface Areas of Samples

From the adsorption isotherms the maximum amount of glucose adsorbed on alumina neutral is \(3.0 \times 10^{-4}\) mole/g, and \(0.45 \times 10^{-4}\) mole/g for alumina heated at 1250°C for 6 hours. Since both isotherms show Langmuir type adsorption with saturation of adsorption at high concentration of glucose, the plateaus should correspond to completion of monolayer of glucose on the alumina surface. Hence the maximum amounts of glucose adsorbed can be used to estimate the area of the surface available for glucose adsorption. Using atomic models the average van der Waal's radius of the D-glucopyranose molecule was found to be about 4 Å. The glucose molecule is most likely adsorbed with the plane of the pyranose ring parallel to the surface, since it allows the maximum interaction of the hydroxyl groups with the surface. Hence the area occupied by a glucose molecule adsorbed on the alumina surface is about \(50\) Å\(^2\), and \(3.0 \times 10^{-4}\) mole of glucose molecules would occupy 90 m\(^2\). On a gram of sintered alumina 0.45 \(\times\) 10\(^{-4}\) mole glucose molecules should occupy 13 m\(^2\).

It is clear that the area occupied by glucose molecules is only 62% of the area covered by a monolayer of nitrogen adsorbed on the same weight of alumina neutral. One possible reason for the difference is the presence
of pores less than 4 Å in radius into which only the nitrogen molecules (average van der Waal's radius ≈ 2 Å) can enter. But it is not likely that much of the surface area of the sample is present in very small pores, since it can be shown that >75% of the surface area available for nitrogen adsorption is present in pores with radii >13 Å. As mentioned before, most of the pores in alumina neutral have radii between 18 Å and 27 Å.

Therefore it seems more likely that there are distinct sites on the γ-alumina surface for adsorption of glucose, which are determined by the surface structure of the catalyst. In other words, glucose molecules cannot occupy all the surface area available for nitrogen molecules and a relatively large fraction of γ-alumina surface is left vacant. The "effective area" of a glucose molecule adsorbed on the γ-alumina surface would seem to be about 80 Å².

With the assumptions given above in the calculation of surface areas, the surface area apparently occupied by glucose molecules adsorbed on sintered alumina is higher than the area determined by adsorption of nitrogen (viz. 6.2 m²/g; see Section II.1). Since the adsorption isotherms do not indicate the presence of multilayer adsorption, the higher area observed may indicate presence of an acyclic intermediate on the surface or of glucopyranose molecules adsorbed with plane of the ring not parallel to the surface. Of course, such details cannot be determined within the many assumptions involved in the area/g calculation; however, it is interesting that area determinations are at least similar with use of nitrogen and of glucose.

XIV.4 Strength of Adsorption of Glucose

As mentioned in Section XIV, the greater the strength of adsorption the lower is the value of [G] at which plateau is reached. Comparison of
adsorption isotherms given in Fig. 33 shows that on alumina neutral the monolayer is complete when \([G] \approx 0.25\) M while a glucose concentration of 1 M is needed to complete the monolayer on sintered alumina. This indicates that at least part of the adsorption sites on sintered alumina are weaker than the weakest sites on alumina neutral. The strengths of adsorption of glucose by the two aluminas can be compared by determining the equilibrium constants for adsorption using Equation (41).

XIV.4.1 Determination of Equilibrium constant \(K\) for Adsorption of Glucose on Alumina Neutral

It was mentioned in Section XIV that the equilibrium constant \(K\) for adsorption of glucose onto alumina can be determined by plotting \(\frac{[GC]}{[G]}\) versus \([GC]\). Since the Langmuir adsorption applies to reversible adsorption, the concentration of glucose adsorbed reversibly was determined by subtracting the amount adsorbed on irreversible sites (determined in Section V.2) from the total amount of glucose on the surface (Table XII). The values of \(\frac{[GC]}{[G]}\) and \([GC]\) are also shown in the same Table and they are plotted in Fig. 34. This plot shows two linear regions*; one with a high \(K\), and the other with low \(K\) (plotted again Fig. 35). This indicates the presence of two types of sites with different strengths of adsorption each giving rise to a linear plot. Equations that explain such behaviour can be derived as follows.

If adsorption on sites of Type 1 occurs independent of adsorption on sites of Type 2, then Equation (41) can be applied to each type.

* The data do not fit Temkin (i.e. linear, \([GC]\) versus \(\ln[G]\) plot) or Freundlich (i.e. linear, \(\ln[G]C\) versus \(\ln[G]\) plot) isotherms.\(^{113,114}\).
### TABLE XII

Data for the Langmuir plot of $[GC]/[G]$ versus $[GC]$ mole litre$^{-1}$ for adsorption of glucose on alumina neutral at 25.0°C

<table>
<thead>
<tr>
<th>Equilibrium Glucose Concentration M</th>
<th>Moles ($10^{-4}$) Glucose Adsorbed per Gram Catalyst</th>
<th>Moles ($10^{-4}$) Glucose Adsorbed on Reversible Sites per Gram Catalyst</th>
<th>$[GC] 	imes 10^4$ mole litre$^{-1}$ (for 1 g dispersed in 60 ml)</th>
<th>$[GC]$</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.00029</td>
<td>0.91</td>
<td>0.21</td>
<td>3.5 ± 0.3</td>
<td>1.1 ± 0.1</td>
</tr>
<tr>
<td>0.0017</td>
<td>1.29</td>
<td>0.59</td>
<td>9.9 ± 0.3</td>
<td>0.58 ± 0.02</td>
</tr>
<tr>
<td>0.0059</td>
<td>1.55</td>
<td>0.85</td>
<td>14.2 ± 0.3</td>
<td>0.24 ± 0.01</td>
</tr>
<tr>
<td>0.0198</td>
<td>1.94</td>
<td>1.24</td>
<td>20.7 ± 0.4</td>
<td>0.104 ± 0.002</td>
</tr>
<tr>
<td>0.044</td>
<td>2.33</td>
<td>1.63</td>
<td>27.2 ± 0.4</td>
<td>0.061 ± 0.001</td>
</tr>
<tr>
<td>0.094</td>
<td>2.70</td>
<td>2.00</td>
<td>33.3 ± 0.5</td>
<td>0.037 ± 0.0004</td>
</tr>
<tr>
<td>0.25</td>
<td>2.98</td>
<td>2.28</td>
<td>38.0 ± 0.5</td>
<td>0.0152 ± 0.0002</td>
</tr>
</tbody>
</table>
Fig. 34 The Langmuir Plot for Adsorption of Glucose on Alumina Neutral
Fig. 35  The Langmuir Plot for Adsorption of Glucose on Alumina Neutral at High Concentrations of Glucose
\[
\frac{[GC]_1}{[G]} + K_1[GC]_1 = K_1[GC]_{m_1} \tag{42}
\]

and
\[
\frac{[GC]_2}{[G]} + K_2[GC]_2 = K_2[GC]_{m_2} \tag{43}
\]

where, \( K_1 \) and \( K_2 \) are the equilibrium constants for adsorption on sites of Types 1 and 2 respectively.

\([GC]_1\) and \([GC]_2\) are the concentrations of glucose adsorbed on sites of Types 1 and 2, respectively, when the equilibrium concentration of glucose is \([G]\), and \([GC]_{m_1}\) and \([GC]_{m_2}\) are the maximum concentrations of glucose on sites of Types 1 and 2 respectively.

Adding Equations (42) and (43) we get,
\[
\frac{[GC]_1 + [GC]_2}{[G]} + K_1[GC]_1 + K_2[GC]_2 = K_1[GC]_{m_1} + K_2[GC]_{m_2} \tag{44}
\]

If \( K_1 >> K_2 \), and \([GC]_{m_1}\) and \([GC]_{m_2}\) are of the same order of magnitude, then, \([GC]_1 >> [GC]_2\) at low concentrations of glucose \([G]\), and Equation (44) reduces to,
\[
\frac{[GC]_1}{[G]} + K_1[GC]_1 = K_1[GC]_{m_1} \tag{42}
\]

Hence a plot of \(\frac{[GC]_1}{[G]} = \frac{[GC]_1}{[G]}\) versus \([GC]_1 + [GC]_2 \approx [GC]_1\), at low concentrations of glucose should be linear with slope = \(-K_1\) and intercept = \(K_1[GC]_{m_1}\). Therefore the strength of adsorption on sites of Type 1 and the number of sites of Type 1 can be determined.

When the glucose concentration is high, \([GC]_1 = [GC]_{m_1}\) and Equation
(44) becomes

\[
\frac{[GC]_m_1 + [GC]_2}{[G]} + K_1[GC]_m_1 + K_2[GC]_2 = K_1[GC]_m_1 + K_2[GC]_m_2
\]

\[
\frac{[GC]_m_1 + [GC]_2}{[G]} + K_2[GC]_2 = K_2[GC]_m_2
\]  \hspace{1cm} (45)

Adding \( K_2[GC]_m_1 \) to both sides of Equation (45) we get,

\[
\frac{[GC]_m_1 + [GC]_2}{[G]} + K_2([GC]_2 + [GC]_m_1) = K_2([GC]_m_2 + [GC]_m_1)
\]  \hspace{1cm} (46)

Hence a graph of \( \frac{[GC]_m_1 + [GC]_2}{[G]} \) versus \( ([GC]_2 + [GC]_m_1) \) should be a straight line with slope = \(-K_2\) and intercept = \( K_2([GC]_m_2 + [GC]_m_1) \).

Therefore the strength of adsorption on sites of Type 2 and the total number of sites can be determined.

Using this treatment the experimental plots in Figs. 34 and 35 will be analyzed below.

From the region showing strong adsorption the slope = \( K_1 = (8.2 \pm 1.2) \times 10^2 \) litre mole\(^{-1}\) and the concentration of sites responsible for strong adsorption is \( \frac{\text{intercept}}{\text{slope}} = (17 \pm 2) \times 10^{-4} \) mole litre\(^{-1}\).

From the slope of Fig. 35, \( K_2 = 44 \pm 4 \) litre mole\(^{-1}\), and using the intercept and concentration of all reversible sites (from Table XII) \( K_2 = 47 \pm 4 \) litre mole\(^{-1}\). Hence \( K_2 \) (average) = \( 46 \pm 3 \) litre mole\(^{-1}\). Further, the concentration of weak adsorption sites = concentration of all reversible adsorption sites - concentration of strong adsorption sites = \( (21 \pm 2) \times 10^{-4} \) litre mole\(^{-1}\).
mole litre$^{-1}$.

Note that according to Equation (46) the concentration of all reversible adsorption sites $= [GC]_1 + [GC]_2$ on one gram of catalyst dispersed in 0.06 litre solution, is given by \( \frac{\text{intercept}}{\text{slope}} \) of the plot in Fig. 35 and equals $(41 \pm 5) 	imes 10^{-4}$ mole litre$^{-1}$. Within experimental error this agrees with the experimentally observed maximum reversible adsorption $= (38 \pm 0.5) 	imes 10^{-4}$ mole litre$^{-1}$ (Table XII), and verifies the applicability of the Equation (46) to the observed Langmuir plot.

Therefore, the results from adsorption studies have given evidence for three types of adsorption sites on alumina neutral. There are $0.7 \times 10^{-4}$ mole ($\approx 23\%$) irreversible adsorption sites, $1.0 \times 10^{-4}$ mole ($\approx 33\%$) strong reversible adsorption sites and $1.3 \times 10^{-4}$ mole ($\approx 43\%$) weak reversible adsorption sites on a gram of alumina neutral.

Out of these three types of adsorption sites, irreversible sites cannot possess any catalytic activity while both types of reversible adsorption sites have the potential to be catalytic sites for the mutarotation reaction. Their catalytic activity is determined in Section XV.

### XIV.4.2 Determination of the Equilibrium Constant for Adsorption of Glucose onto Alumina Sintered at 1250°C

Similar to the case of glucose on standard alumina neutral as described in Section XIV.4.1, the equilibrium constant for adsorption of glucose onto sintered alumina can also be determined from the Langmuir plot of $\frac{[GC]}{[G]}$ versus $[GC]$. Data for the plot are given in Table XIII

* Since 1 g of catalyst is found in 0.06 litre the amount of strong adsorption sites on 1 g catalyst is $(17 \times 10^{-4}) \times 0.06 = 1.0 \times 10^{-4}$ mole.
TABLE XIII

DATA FOR THE LANGMUIR PLOT OF $[GC]/[G]$ VERSUS $[GC]$ MOLE LITRE$^{-1}$ FOR ADSORPTION OF GLUCOSE (AT 25.0°C) ON ALUMINA SINTERED AT 1250°C

<table>
<thead>
<tr>
<th>Equilibrium Concentration of Glucose M</th>
<th>Moles ($x 10^5$) Glucose Adsorbed per Gram Catalyst</th>
<th>$[GC] \times 10^4$ mole litre$^{-1}$ (for 1 g dispersed in 60 ml)</th>
<th>$[GC]/[G]$</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.00021</td>
<td>0.6</td>
<td>1.0</td>
<td>0.48</td>
</tr>
<tr>
<td>0.00052</td>
<td>0.8</td>
<td>1.3</td>
<td>0.26</td>
</tr>
<tr>
<td>0.0042</td>
<td>1.1</td>
<td>1.8</td>
<td>0.044</td>
</tr>
<tr>
<td>0.0088</td>
<td>1.3</td>
<td>2.2</td>
<td>0.025</td>
</tr>
<tr>
<td>0.0232</td>
<td>1.5</td>
<td>2.5</td>
<td>0.011</td>
</tr>
<tr>
<td>0.0473</td>
<td>1.8</td>
<td>3.0</td>
<td>0.0063</td>
</tr>
<tr>
<td>0.0943</td>
<td>2.3</td>
<td>3.8</td>
<td>0.0041</td>
</tr>
<tr>
<td>0.240</td>
<td>3.2</td>
<td>5.3</td>
<td>0.0022</td>
</tr>
<tr>
<td>0.488</td>
<td>3.9</td>
<td>6.5</td>
<td>0.0013</td>
</tr>
<tr>
<td>0.954</td>
<td>4.6</td>
<td>7.7</td>
<td>0.0008</td>
</tr>
</tbody>
</table>
and the plot in Fig. 36. Note that the amount of irreversible sites (if any) was not determined and hence the quantitative data obtained from the plot may contain a small error.

The plot in Fig. 36 is very similar to that observed for alumina neutral which, as shown in Section XIV.4.1, has two types of reversible adsorption sites. The strong adsorption sites on sintered alumina have an equilibrium constant of about 6000 litre mole\(^{-1}\) and there are \(1.2 \times 10^{-5}\) moles of strong adsorption sites on a gram of catalyst. The equilibrium constant for adsorption on weak sites is 13 litre mole\(^{-1}\) and there are \(3.4 \times 10^{-5}\) moles of weak adsorption sites on a gram of sintered alumina.

The results discussed in this section have shown that both alumina neutral and alumina sintered at 1250°C contain two types of reversible adsorption sites. The catalytic activities of both strong and weak reversible adsorption sites on alumina neutral are determined in the next section.
Fig. 36 The Langmuir Plot for Adsorption of Glucose on Alumina Sintered at 1250°C for 6 hours
CATALYTIC ACTIVITY OF REVERSIBLE ADSORPTION SITES
AND THE ACTIVE SITE DENSITY OF ALUMINA NEUTRAL
For a reversible catalytic process given by equation

\[
S + C \rightleftharpoons SC \rightleftharpoons PC \rightleftharpoons P + C
\]  (26)

the initial rate of the reaction (after adsorption is complete) is given by,

\[
\frac{d[P]}{dt} \approx k_2 [PC] \approx k_3 [SC], \text{ since } [P] \approx 0 \text{ and } k_2 \gg k_3, k_4.
\]

Hence the initial rate of the reaction is directly proportional to the concentration of \(\alpha\)-D-glucose, \([SC]\), on the surface. The amount of substrate-catalyst complex formed for a given \(\alpha\)-D-glucose concentration is given by the adsorption isotherm. Therefore, if all reversible adsorption sites are equally active, then a plot of initial rate of reaction versus concentration of glucose in solution, and a plot of moles of glucose reversibly adsorbed versus the concentration of glucose in solution, should be superimposable.

To determine the initial rate of the reaction at different glucose concentrations, plots of optical rotation versus time were made using the data from Section X. The slope at \(t = 25\) mins gave the rate of the reaction in degrees sec\(^{-1}\) (rates at high glucose concentrations were corrected for slow homogeneous reaction). Using optical rotations of equimolar solutions of \(\alpha\)-D-glucose and \(\beta\)-D-glucose it can be shown that the optical rotation, measured at 365 nm with a 0.1 dm path length cell, decreases by 1° when 0.012 mole \(\alpha\)-D-glucose in 60 ml DMSO is converted to 0.012 mole \(\beta\)-D-glucose. Hence the number of moles reacting per second in the 60 ml solution is

\[
\frac{d[P]}{dt}
\]
given by, slope \((\text{degrees sec}^{-1}) \times 0.012 \text{ (mole degree}^{-1})\) and are recorded in Table XIV. The concentration of \(\alpha\)-D-glucose in solution when adsorption is complete was determined from the adsorption isotherm.

The plots of initial rate versus concentration of \(\alpha\)-D-glucose and the adsorption isotherm for reversible adsorption (from Table XII) are shown in Fig. 37. They show that when adsorption is about 37% complete the initial rate is only 7% of its maximum value. This indicates that sites responsible for initial adsorption (strong adsorption sites) are not as active as weak adsorption sites.

In order to compare the activities of the two types of sites, their adsorption isotherms were constructed using the equilibrium constants \((K_1\) and \(K_2)\) and the maximum concentrations of glucose adsorbed on the two sites, \([GC]_{m_1}\) and \([GC]_{m_2}\). They are shown in Fig. 38 together with the plot of initial rate versus concentration of glucose. The values for initial rate have been multiplied by 0.0303 to coincide the maxima of the initial rate and the adsorption isotherm due to weak adsorption sites.

From Fig. 38 it is clear that there is no relationship between initial rate of the reaction and the adsorption isotherm due to strong adsorption sites. For example, when adsorption on those sites is 75% complete the initial rate is only 7% of its maximum value. Hence the strong adsorption sites are either catalytically inactive or their activity is relatively very slow.

The Fig. 38 also shows that within experimental error the adsorption isotherm due to weak sites and the initial rate plot are superimposable. Note that the initial rates observed at low \(\alpha\)-D-glucose concentrations are less than what is expected from the adsorption isotherm. This is because the assumption that the concentration of product \(\beta\)-D-glucose is negligible would hold well at high \(\alpha\)-D-glucose concentrations but not at low


<table>
<thead>
<tr>
<th>Initial Concentration of α-D-Glucose (M)</th>
<th>Concentration of α-D-Glucose after Adsorption (M)</th>
<th>(observed) in degrees/sec x 10^6</th>
<th>Initial Rate in 60 ml solution mole sec^{-1} x 10^8</th>
<th>(scaled) x 0.303 mole sec^{-1} x 10^8</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.01</td>
<td>0.006</td>
<td>2.5</td>
<td>3.0</td>
<td>0.09</td>
</tr>
<tr>
<td>0.02</td>
<td>0.015</td>
<td>10.3</td>
<td>12.4</td>
<td>0.38</td>
</tr>
<tr>
<td>0.03</td>
<td>0.025</td>
<td>15.5</td>
<td>18.6</td>
<td>0.57</td>
</tr>
<tr>
<td>0.05</td>
<td>0.044</td>
<td>22.2</td>
<td>26.6</td>
<td>0.81</td>
</tr>
<tr>
<td>0.10</td>
<td>0.094</td>
<td>29.8</td>
<td>35.8</td>
<td>1.09</td>
</tr>
<tr>
<td>0.20</td>
<td>0.20</td>
<td>33.0</td>
<td>39.6</td>
<td>1.20</td>
</tr>
<tr>
<td>0.40</td>
<td>0.40</td>
<td>35.2</td>
<td>42.2</td>
<td>1.28</td>
</tr>
<tr>
<td>0.80</td>
<td>0.80</td>
<td>35.0</td>
<td>42.0</td>
<td>1.27</td>
</tr>
</tbody>
</table>
Fig. 37  Adsorption Isotherm for Reversible Adsorption of Glucose on Alumina Neutral (left and ●) and Plot of Initial Rate on Alumina Neutral Versus Concentration of Glucose (right and ■).
Amount of Glucose (x 10^4 moles) Adsorbed per Gram of Alumina Neutral, and Initial Rate (x 0.303 x 10^8 mole.sec^-1) on 1.6 g of Catalyst in 60 ml α-D-Glucose Solution

Rate Versus Glucose Concentration.

and (b) Weak Adsorption Sites on Alumina Neutral with (c) the Plot of Initial Rate on 1.6 g of Catalyst in 16 mL of Glucose Solution for Adsorption of Glucose on (a) Strong

Fig. 38 Comparison of the Theoretical Isotherms for Adsorption of Glucose on (a) Strong.
concentrations. At low $\alpha$-D-glucose concentrations the relative amount of $\beta$-D-glucose formed at $t = 25$ mins can be appreciable and hence, the reverse reaction also would occur at the same time. Therefore the observed initial rate would be slower than what is expected when $[\beta$-D-glucose] $\simeq 0$.

The above results lead to the conclusion that the catalytic activity of alumina is due to the weak adsorption sites and the strong adsorption sites have either no or negligible catalytic activity. However, it is impossible to prove that all the weak adsorption sites are catalytically active, because the same result (superimposable adsorption isotherm and initial rate plot) can be produced even when a fraction of the weak adsorption sites are catalytically active. That is, two types of sites on the alumina surface with the same values for $K$ but only one being catalytically active will also produce the same result. It may be argued that the probability of existence of two types of sites having different chemical functional groups but with the same strength of interaction (related to heat of adsorption) with the reactant molecules is very small. Validity of such reasoning can be verified only by determining the active site density by other more reliable methods (see below). However, it is safe to conclude at this point that the number of weak adsorption sites gives the upper limit of the number of active sites on the alumina surface.

This upper limit for the number of catalytic sites on alumina neutral is therefore $(1.3 \pm 0.1)10^{-4}$ mole/gram $= (7.8 \pm 0.8)10^{19}$ sites/gram $= (5.4 \pm 0.6)10^{13}$ sites/cm$^2$.

XV.1 Comparison with Other Methods of Determining Active Site Density

The method used above to determine the number of active sites per gram can be applied to other catalytic systems when there is an experimentally
observed relation between the adsorption isotherm and the plot of initial rate versus substrate concentration. If there is no observable relation, one might be able to divide the observed adsorption isotherm into component isotherms and compare these with the initial rate plot. In the case of glucose on alumina this is possible only because the individual isotherms are widely different with different equilibrium constants. In other cases this method will not be applicable.

The most direct method for determining the number of active sites, which was most successfully applied by Kokes\textsuperscript{33,125}, involves using the reactant itself as a probe molecule (just as in the method described in this thesis). They were able to show that the surface compounds observed by infrared spectroscopy are intermediates of the catalytic reaction and not only their precursors. The number of intermediates could be determined from the adsorption isotherm and their formation was followed by infrared spectroscopy. This is an ideal method which is limited to systems where there is an appreciable concentration of surface intermediates formed which can be detected without any interference by other adsorbed species.

Another rather direct approach involves detailed analysis of absolute rates for the estimation of the active site density from pre-exponential factors\textsuperscript{126}. From the transition state theory it has been shown that rate constant of a surface reaction, assuming the rate determining step to be unimolecular change of adsorbed reactant, is

\[ v = A \exp\left(-\frac{E}{RT}\right) = C_A \left(\frac{kT}{h}\right) \exp\left(\frac{\Delta S}{R}\right) \exp\left(-\frac{E}{RT}\right) \]  

(47)

where \( v \) is the rate constant in molecules reacting per unit surface area per second when the surface is fully covered, \( C_A \) is the concentration of
catalytically active sites, $k$ and $\hbar$ are Boltzmann and Planck constants respectively, and $\Delta S$ is the entropy of activation. If $\Delta S$ is known or if it is assumed to be approximately zero, because both reactant and activated complex are adsorbed species, then the site density can be determined when experimental values of $v$ and activation energy $E$ are known. The validity of the transition state method has been demonstrated by agreement (within an order of magnitude) of the site density obtained using Equation (47) with that obtained by other quite reliable means.$^{126}$

An error is introduced when $E$ is determined from an Arrhenius plot since this method neglects the occurrence of $T$ in the pre-exponential factor. If $\Delta S$ is not known and is considered to be zero, because both reactant and activated complex are adsorbed species, again an error will be introduced in the value of $C_A$ estimated.

Finally there is a more indirect method called specific poisoning of a catalyst surface which is very commonly used by catalytic chemists. A probe molecule other than the reactant molecule is preadsorbed as a poison and its effect on the catalytic activity is studied. From the number of poisoning molecules necessary to bring the activity to zero, the upper limit of the active site density is determined. It is assumed that the reaction of the inhibitor with the surface is irreversible, that it combines only with catalytically active sites, and that there is one inhibitor molecule per active site. Because of these assumptions the upper limit of active site density is obtained. For example, by poisoning the acid sites on $\gamma$-alumina by alkali, Pines and Haag$^{32}$ showed that the upper limit of the total number of sites (all capable of dehydrating 1-butanol) and the number of more highly reactive sites (also capable of isomerizing cyclohexene) are $10 \times 10^{13}$ and $0.8 \times 10^{13}$ sites/cm$^2$ respectively. Rosynek
and Hightower have shown by the poisoning effect of CO₂, that the upper limit of sites capable of exchanging hydrogen atoms on olefins is 3-8 x 10^{12} /cm².

From the above discussion it is clear that the determination of active sites by relating the initial rate to the adsorption isotherm is better than the poisoning method at determining the upper limit. The adsorption isotherm versus rate method is able to eliminate sites that are not catalytically active but which may still possess acid and base functional groups capable of combining with a reactive inhibitor molecule.

The results discussed in this section have shown that the weak adsorption sites on alumina neutral possess catalytic activity while the strong adsorption sites are either catalytically inactive or their activity is relatively very low.* The possible reasons for the inactivity or very low activity of strong adsorption sites on alumina neutral are discussed below.

XV.2 Catalytic Inactivity of Strong Adsorption Sites

The observed rate constant \( k_{\text{obs}} \) (Equation 32) may be divided into two terms, one due to weak adsorption sites and the other due to strong adsorption sites. However, the observation that there are two distinct types of reversible adsorption sites on sintered alumina, as well as on alumina neutral, suggests that the method for determining the number of active sites on a catalyst described earlier may be applicable here.
The catalytic inactivity of reversible adsorption sites (whose concentration $c^S_o$ is appreciable) can be due to two reasons. Either the term $(k_3 + k_4)^S$ is zero or very small, or the term $1/K_1$ is very large. However, the results in Section XIV.4.1 showed that the term $1/K_1$ for strong adsorption sites is smaller than the term $1/K_2$ for weak adsorption sites and this factor actually helps to increase the activity. Therefore the reason for the presence of catalytic activity on weak adsorption sites and the absence of activity on strong adsorption sites is that the term $(k_3 + k_4)^w$ is relatively high for weak sites while the term $(k_3 + k_4)^S \approx 0$ for strong adsorption sites. That is, the strong adsorption sites are inactive because the catalytic constant for the surface reaction $(k_3 + k_4)^S$ is very small (probably because the acidic and/or basic groups necessary for catalysis are either absent or not in the proper position for catalysis to occur) and not because the adsorption is unfavorable. The catalytic constant for surface reaction on weak adsorption sites $(k_3 + k_4)^w$ and other kinetic parameters of the surface reaction are determined in the next section.
XVI KINETIC PARAMETERS OF THE CATALYTIC SYSTEM
The equilibrium constant for adsorption on the weak sites on alumina neutral and the upper limit of the catalytic site density can be used to determine various kinetic parameters such as the catalytic constant for the surface reaction in Equation 26, the turnover number of a catalytic site, and the catalytic constant for the overall reaction.

\[ S + C \xrightleftharpoons[k_2^{-1}]^{k_1} SC \xrightleftharpoons[k_4^{-1}]^{k_3} PC \xrightleftharpoons[k_1^{-1}]^{k_2} P + C \]  

(26)

XVI.1 Determination of the Catalytic Constant for the Surface Reaction \((k_3 + k_4)\)

(a) From the plot of \(k_{ob} \) versus concentration of catalyst in Fig. 21 (Section IX) the experimentally obtained slope = \(1.14 \times 10^{-5} \text{ sec}^{-1} \text{ g}^{-1} \text{ litre}\).

From the active site density determination, one gram of catalyst in one litre has \(1.3 \times 10^{-4}\) mole catalytic sites.

\[ \text{Slope} = \frac{1.14 \times 10^{-5}}{1.3 \times 10^{-4}} = 0.088 \text{ sec}^{-1} \text{ mole}^{-1} \text{ litre} \]

Since \(k_{ob} = \frac{(k_3 + k_4)c_o}{1/K + s_o}\), the slope of the plot in Fig. 21, =

\[ \frac{k_3 + k_4}{1/K + s_o} = 0.088 \text{ sec}^{-1} \text{ mole}^{-1} \text{ litre}. \]

From the Langmuir plot in Fig. 35 the equilibrium constant for the
catalytic sites \( K = 46 \text{ litre mole}^{-1} \). Therefore \( 1/K = 0.022 \text{ mole litre}^{-1} \) and the total substrate concentration \( s_0 = 0.050 \text{ mole litre}^{-1} \).

\[
\frac{(k_3 + k_4)}{0.072 \text{ mole litre}^{-1}} = 0.088 \text{ sec}^{-1} \text{ mole}^{-1} \text{ litre}
\]

and \( (k_3 + k_4) = 6.3 \times 10^{-3} \text{ sec}^{-1} \)

\[\approx 6 \times 10^{-3} \text{ sec}^{-1}\]

(b) The catalytic constant can also be determined from the slope of the plot of \( \frac{1}{k_{\text{obs}}} \) versus concentration of glucose \( (s_0) \) given in Fig. 23 (Section X).

Again, since

\[
\frac{1}{k_{\text{obs}}} = \frac{s_0}{(k_3 + k_4)s_0} + \frac{1}{k(k_3 + k_4)s_0}
\]

Slope = \( \frac{1}{(k_3 + k_4)s_0} \)

where \( s_0 \) the catalyst concentration = 26.7 mg/ml = 35 \times 10^{-4} \text{ mole catalytic sites/litre}.

\[
\therefore \frac{1}{(k_3 + k_4)s_0} = \frac{1}{(k_3 + k_4)35 \times 10^{-4}} = 7.8 \times 10^{4} \text{ sec litre mole}^{-1},
\]

from the plot in Fig. 23.

\[
\therefore (k_3 + k_4) = 3.7 \times 10^{-3} \text{ sec}^{-1}
\]

\[\approx 4 \times 10^{-3} \text{ sec}^{-1}\]
The two values for \((k_3 + k_4)\) obtained by the two different plots are quite close to each other and gives \((k_3 + k_4)_{\text{average}} = 5 \times 10^{-3} \text{ sec}^{-1}\).

The term \((k_3 + k_4)\) is the catalytic constant for interconversion of \(\alpha-\) and \(\beta-D\)-glucopyranose on the two dimensional homogeneous medium offered by the alumina surface (Equation 26). Hence it may be called the catalytic constant for "homogeneous reaction" on alumina surface, in comparison with the catalytic constant for homogeneous reaction in solution. The catalytic constant for homogeneous catalysis by water at \(25.0^\circ\text{C}\) is \(4 \times 10^{-4} \text{ sec}^{-1}\) (see Table III) and hence, it follows that this alumina surface offers a better medium for glucose mutarotation than water.

**XVI.2 Catalytic Constant for Heterogeneous Catalysis of Glucose Mutarotation**

From Section XVI.1 the observed rate constant when the catalyst concentration is 1 mole litre\(^{-1}\) (i.e. the catalytic constant \(k_{\text{cat}}\)) is 0.088 sec\(^{-1}\) mole\(^{-1}\) litre and may be compared with other catalytic constants given in Table III for homogeneous catalysis of glucose mutarotation. It is interesting to note that it is about 100 times more active than a weak base like acetate in water, about 9 times more active than strong acids in water, and a little more active than 2-hydroxypyridine in benzene. However, it is important to note that the observed high catalytic constant for alumina is a result not only of the medium (i.e. \(k_3 + k_4\)) but also of relatively strong substrate-catalyst binding (i.e. \(K_2\)) at low substrate concentrations (concentration of glucose = 0.05 M). In spite of these favorable conditions, alumina neutral is \(2 \times 10^4\) times less active than nature's catalyst, enzyme mutarotase (concentration of glucose = 0.11 M).
XVI.3 Turnover Number of a Catalytic Site

The turnover number (also called catalytic center activity) is defined as the number of molecules of substrate reacting per second per active site. Using the upper limit of active site density and the initial rate when the surface is completely covered, the minimum value for the turnover number can be calculated as follows.

Initial rate when the surface of 1.6 g alumina is completely covered = $0.42 \times 10^{-6}$ moles sec$^{-1}$ (from Table XIV).

= $2.5 \times 10^{17}$ molecules sec$^{-1}$

.'. Rate on one gram of catalyst = $1.6 \times 10^{17}$ molecules sec$^{-1}$ gram$^{-1}$

.'. Number of molecules reacting on one site in one second

= $2 \times 10^{-3}$ molecules site$^{-1}$ second$^{-1}$

It may be compared with turnover numbers for some other surface catalyzed reactions at 25°C, $N_{25}$, (calculated using activation energies) given in Table XV. It is clear that alumina is remarkably efficient in catalyzing the mutarotation reaction compared to most other reactions that have been studied so far. The turnover number for o-p conversion is higher than that for glucose mutarotation, but it occurs only on activated aluminas. The high catalytic activity of standard alumina neutral, which has not been activated, towards glucose mutarotation should be related to the high sensitivity of the reaction towards weak acidic and basic sites on alumina. This confirms the advantage of using mutarotation reaction as a probe for weak acidic and basic sites on the alumina surface.

The results discussed in the last two sections have shown that only a part (maximum ≈ 40%) of the adsorption sites on alumina neutral possess
<table>
<thead>
<tr>
<th>Reaction</th>
<th>Catalyst</th>
<th>Turnover Number N&lt;sub&gt;25&lt;/sub&gt; molecules site&lt;sup&gt;-1&lt;/sup&gt; sec&lt;sup&gt;-1&lt;/sup&gt;</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cyclohexanol dehydration</td>
<td>Al&lt;sub&gt;2&lt;/sub&gt;O&lt;sub&gt;3&lt;/sub&gt;</td>
<td>9 x 10&lt;sup&gt;-7&lt;/sup&gt;</td>
</tr>
<tr>
<td>n-Propanol dehydration</td>
<td>Al&lt;sub&gt;2&lt;/sub&gt;O&lt;sub&gt;3&lt;/sub&gt;</td>
<td>2 x 10&lt;sup&gt;-6&lt;/sup&gt;</td>
</tr>
<tr>
<td>Ethanol dehydration</td>
<td>SiO&lt;sub&gt;2&lt;/sub&gt;-Al&lt;sub&gt;2&lt;/sub&gt;O&lt;sub&gt;3&lt;/sub&gt;</td>
<td>6 x 10&lt;sup&gt;-10&lt;/sup&gt;</td>
</tr>
<tr>
<td>HCO&lt;sub&gt;2&lt;/sub&gt;H decomposition</td>
<td>Al&lt;sub&gt;2&lt;/sub&gt;O&lt;sub&gt;3&lt;/sub&gt;</td>
<td>8 x 10&lt;sup&gt;-4&lt;/sup&gt;</td>
</tr>
<tr>
<td>Cyclohexane dehydrogenation</td>
<td>Pt-Al&lt;sub&gt;2&lt;/sub&gt;O&lt;sub&gt;3&lt;/sub&gt;</td>
<td>7 x 10&lt;sup&gt;-5&lt;/sup&gt;</td>
</tr>
<tr>
<td>tert-Butylbenzene cracking</td>
<td>SiO&lt;sub&gt;2&lt;/sub&gt;-Al&lt;sub&gt;2&lt;/sub&gt;O&lt;sub&gt;3&lt;/sub&gt;, cegelled</td>
<td>6 x 10&lt;sup&gt;-4&lt;/sup&gt;</td>
</tr>
<tr>
<td>tert-Butylbenzene cracking</td>
<td>SiO&lt;sub&gt;2&lt;/sub&gt;-Al&lt;sub&gt;2&lt;/sub&gt;O&lt;sub&gt;3&lt;/sub&gt; (prepared by dialysis)</td>
<td>16</td>
</tr>
<tr>
<td>o-p H&lt;sub&gt;2&lt;/sub&gt; conversion</td>
<td>Al&lt;sub&gt;2&lt;/sub&gt;O&lt;sub&gt;3&lt;/sub&gt;</td>
<td>70*</td>
</tr>
</tbody>
</table>

* Turnover number at -196°C
catalytic activity. Irreversible adsorption sites cannot give rise to catalytic activity even if they contain acidic and basic functional groups, and reversible adsorption on sites where the acidic and/or basic functional groups necessary for catalysis are either absent or not in proper position for catalysis to occur will not cause mutarotation. Hence less than 40% of the adsorption sites show reversible adsorption and also possess appropriate functional groups (acidic and/or basic sites) which catalyze mutarotation. The nature of adsorption and the nature of acid/base functional groups present at the active sites are investigated in the next two sections.
XVII  NATURE OF ADSORPTION ON ACTIVE SITES
The adsorption of glucose on alumina should occur by the interaction of the polar (hydroxyl) groups on glucose molecule with the polar groups (Al$^{3+}$, O$^{2-}$, OH, OH$^+$) on the surface. It was observed in Section VII.2 that excess water added onto alumina neutral had no effect on the rate constant for mutarotation. This shows that the glucose molecule is undergoing a strong specific adsorption on the alumina surface. To understand the specificity of adsorption by catalytic sites, the inhibitory effect of simple organic molecules with different functional groups was studied.

(a) Benzene

Benzene is a simple organic molecule with a π electron cloud (a weak base) which can interact with electron deficient centres on the surface. Benzene is known to exchange its hydrogen with D$_2$, at 25°C, on activated alumina. However, when a 60 ml solution, 0.05 M in α-D-glucose and 0.05 M in benzene, was stirred with 1.6 g alumina neutral no effect on the rate was observed.

(b) Naphthalene

When the above experiment was performed with naphthalene (which has a more reactive π-electron cloud), instead of benzene, again, no change in activity was observed.

These results indicate that glucose molecules on active sites are probably not adsorbed on strong Lewis acid sites. This result is not surprising since any strong Lewis acid sites on the surface should already be occupied by water molecules present on alumina neutral.
(c) **Methanol**

Since water had no inhibitory effect on the reaction (as mentioned above) the effect of the simplest alcohol methanol was studied. When a 60 ml solution, 0.03 M in α-D-glucose and 0.15 M in methanol, was stirred with 1.6 g alumina neutral no change in rate was observed. Increasing the concentration of methanol to 0.4 M (so that \( \frac{[\text{MeOH}]}{[\text{glucose}]} \approx 13 \)) had very little effect on the rate.

These experiments show that simple hydroxy compounds cannot compete with glucose for its active sites even when used in excess. Hence the effect of the following polyhydroxy compounds on the rate of mutarotation was studied.

(d) **Methyl α-D-Glucoside**

When a 60 ml 0.05 M solution of methyl α-D-glucoside was stirred with 1.6 g alumina neutral, about 8.5% adsorption (complete in 20-30 mins) was observed at 25°C. However, there was no hydrolysis of the glucoside, even when the temperature was raised to 100°C. This suggests that the mutarotation of glucose on alumina does not occur via a carbonium ion intermediate. The decrease in adsorption compared to α-D-glucose (adsorption 14% under similar conditions) should at least partly be due to the presence of the methyl group which affects the interaction of the surface with one side of the pyranose ring.

It may be expected that this polyhydroxy compound should be adsorbed on almost every site available for adsorption of glucose. Not surprisingly, it was found to be an effective inhibitor of glucose mutarotation. For example, a 60 ml solution 0.05 M in α-D-glucose and 0.01 M in the glucoside decreased the activity (compared with the activity in the absence of
inhibitor) of 1.6 g alumina neutral by 17%. When the concentration of the glucoside was increased to 0.05 M the catalytic activity dropped to 28% of the initial value.

(e) i-Inositol

i-Inositol is one of the nine isomers of hexahydroxycyclohexane and is structurally similar to D-glucose. Both have the same molecular formula and comparison of structures VIII and X shows that carbons 1, 2, 3 and 4 of α-D-glucose and 1, 2, 3 and 4 of i-inositol have the same configuration. Further, inositol has one hydroxyl group more than glucose, and does not possess a methyl group (as in methyl glucoside) which restricts adsorption. Hence it is expected to be a very efficient inhibitor of glucose mutarotation. In fact, when a 60 ml solution 0.05 M in α-D-glucose and 0.05 M in i-inositol, was stirred with 1.6 g alumina neutral the catalytic activity was decreased about 67%. It also decreased the percentage glucose adsorbed on the surface from 14% to 6%.

The results discussed above show that there are relatively specific sites for adsorption of glucose on alumina surfaces. These sites adsorb polyhydroxy compounds strongly, possibly through the cooperative placement of several hydroxyl groups. Simple monohydroxy groups cannot compete for those sites even when used in excess.
A possible intermediate of the mutarotation reaction (as described in the Introduction) is the acyclic form of glucose which is a polyhydroxy aldehyde VII (see Section 1.2). The efficiency of mutarotation will be determined by the stabilization of this species (more particularly the transition state leading to it) by the active sites on the alumina surface. To determine whether the catalytic sites can interact with an aldehyde the following inhibitors were investigated.

(f) **Hexanal**

Hexanal is similar to the acyclic intermediate from glucose except for the absence of hydroxyl groups. It has only one polar group unlike the polyhydroxy compounds used before and its interaction with the catalyst depends on the presence of sites that can interact with an aldehyde functional group containing electropositive carbon and the electron rich oxygen. Hence it is a probe for sites that can stabilize the acyclic form of α-D-glucose.

When a 60 ml solution, 0.05 M in α-D-glucose and 0.01 M in hexanal, was stirred with 1.6 g alumina neutral a 17% decrease in activity was observed. When the concentration of hexanal was increased to 0.05 M, catalytic activity decreased by 26% of the original value, and the percentage glucose adsorbed on alumina decreased from 14% to 12%. When the concentration of hexanal was increased further to 0.25 M the catalytic activity decreased by 95% and the amount of glucose adsorbed on alumina was 6%.

These results show that hexanal is as effective as methyl α-D-glucoside in inhibiting glucose mutarotation on alumina and also show that the catalytically active sites on alumina can interact with an aldehyde functional group. However, any specific interaction of the hydroxyl groups of glucose adsorbed on the alumina surface with the aldehyde functional group of hexanal
may also contribute to the inhibitory effect of hexanal.

(g) **DL-Glyceraldehyde**

Glyceraldehyde is a dihydroxy aldehyde and should be a more effective inhibitor than hexanal since it has two hydroxyl groups which also can interact with the surface. Further, it has structural similarities to the acyclic intermediate that may be formed during mutarotation.

When a 60 ml solution, 0.05 M in α-D-glucose and 0.05 M in DL-glyceraldehyde, was stirred with 1.6 g alumina neutral the activity was about 50% less than that observed in the absence of the inhibitor. The percentage glucose adsorbed also decreased from 13% to 6%. Glyceraldehyde is therefore little less effective than inositol as an inhibitor.

All the results obtained with the inhibitors are tabulated in Table XVI. Comparison of these results leads to the following conclusions.

(i) Equimolar amounts of inositol and glucose can decrease the rate by 67% and the amount of glucose adsorbed by about the same percentage (≈60%). Hence inositol is adsorbed on alumina more effectively than glucose itself. This is not surprising since it has one hydroxyl group more than glucose.

(ii) On the other hand, glyceraldehyde and glucose are able to compete equally well for adsorption and active sites of alumina. Again, the percentage decrease in activity and adsorption are almost equal. This indicates that inositol and glyceraldehyde have no preference for adsorption on inactive over catalytically active sites on the surface. (iii) However, the more interesting inhibitor is hexanal. It is less efficient than inositol and glyceraldehyde mole to mole, since it has only one polar group. But, the results in Table XVI show, that the percentage decrease in activity is
# TABLE XVI

**EFFECT OF INHIBITORS ON THE CATALYTIC ACTIVITY AND ON THE AMOUNT OF GLUCOSE ADSORBED**

<table>
<thead>
<tr>
<th>Inhibitor</th>
<th>Concentration of Inhibitor M</th>
<th>Concentration of Glucose M</th>
<th>Percentage Inhibition</th>
<th>Percentage Decrease in Glucose Adsorbed</th>
</tr>
</thead>
<tbody>
<tr>
<td>Benzene</td>
<td>0.05</td>
<td>0.05</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Naphthalene</td>
<td>0.05</td>
<td>0.05</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Methanol</td>
<td>(i) 0.15</td>
<td>0.03</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td></td>
<td>(ii) 0.40</td>
<td>0.03</td>
<td>≈ 0</td>
<td>0</td>
</tr>
<tr>
<td>Methyl α-D-Glucoside</td>
<td>(i) 0.01</td>
<td>0.05</td>
<td>17</td>
<td></td>
</tr>
<tr>
<td></td>
<td>(ii) 0.05</td>
<td>0.05</td>
<td>28</td>
<td></td>
</tr>
<tr>
<td>i-Inositol</td>
<td>0.05</td>
<td>0.05</td>
<td>67</td>
<td>60</td>
</tr>
<tr>
<td>Hexanal</td>
<td>(i) 0.01</td>
<td>0.05</td>
<td>17</td>
<td>5</td>
</tr>
<tr>
<td></td>
<td>(ii) 0.05</td>
<td>0.05</td>
<td>26</td>
<td>14</td>
</tr>
<tr>
<td></td>
<td>(iii) 0.25</td>
<td>0.05</td>
<td>95</td>
<td>60</td>
</tr>
<tr>
<td>DL-Glyceraldehyde</td>
<td>0.05</td>
<td>0.05</td>
<td>50</td>
<td>54</td>
</tr>
</tbody>
</table>
always higher than the percentage decrease in adsorption. This indicates that hexanal is a more selective adsorbent and prefers catalytically active sites to other adsorption sites. Hence it may be concluded that catalytically active sites are specific for adsorption of polyhydroxy compounds and can, in addition, interact with an aldehyde group better than the other adsorption sites.

XVII.1 Effect on Catalytic Activity of Pretreatment of Alumina with Inhibitors

It was observed during the study of the curvature in the first order plots (Section VIII), that pretreatment of alumina with glucose deactivated the catalyst and the deactivated catalyst then gave rise to linear first order plots. The results showed that the curvature arose from the progressive deactivation of the catalyst by glucose. It was concluded that deactivation was caused by irreversible adsorption of glucose, for example by ether formation, on the active sites.

The inhibitors mentioned above, with different functional groups which adsorb on the active sites, can be used in a study of the functional groups involved in the deactivation process. Experiments similar to those used in the study of curvature in first order plots were carried out by pretreating alumina neutral with inhibitors instead of glucose.

(a) Methyl α-D-Glucoside

As mentioned above, methyl α-D-glucoside is not hydrolyzed by alumina and hence, it has only hydroxyl groups that can combine with functional groups on the surface. Fifty ml DMSO containing 0.582 g glucoside was stirred with 1.6 g alumina neutral. After six hours a 10 ml solution of
0.54 g α-D-glucose in DMSO was added (so that the resulting 60 ml solution is 0.05 M in α-D-glucose and 0.05 M in methyl α-D-glucoside) and the change in optical rotation with time was followed. The results given in Fig. 39 show that pretreatment with methyl glucoside for six hours has deactivated the catalyst by 35% (compared with about 50% by glucose) and the first order plot obtained is linear. Pretreatment for 24 hours did not deactivate the catalyst any further. (Note that with glucose the deactivation increased with pretreatment time although at a very slow rate Fig. 18).

These results show that methyl glucoside is able to deactivate active sites which rapidly get deactivated during mutarotation but it is much less effective than glucose itself.

(b) Hexanal

The same experiments were performed with hexanal. It was found to be less effective than even methyl glucoside in deactivating the catalyst. There was about 26% deactivation after pretreatment for 6 hours and about 33% after 24 hours pretreatment. However, the first order plots still showed curvature (Fig. 40) indicating that at least some of the rapidly deactivating sites have not been deactivated by hexanal.

These results show that deactivation of the catalyst can be caused by compounds with hydroxyl groups as well as with an aldehyde group. However, hydroxyl groups most effectively cause the initial rapid deactivation observed with glucose and any further deactivation due to hydroxyl groups is negligible. (Similar behavior was shown by methanol. When 1.6 g alumina dried at 150°C was pretreated directly with 1.1 ml (0.03 mole) methanol for 150 mins, a catalyst with activity the same as the initial activity of the original catalyst was obtained. However, the methanol treated catalyst did
Fig. 39  The Effect of Pretreatment of Alumina Neutral with Methyl α-D-Glucoside on Its Catalytic Activity

\[ [\alpha-D-Glucose] = 0.05 \text{ M} \]

\[ [\text{Alumina}] = 26.7 \text{ mg.ml}^{-1} \]

- Catalytic activity in the absence of methyl glucoside
- Catalytic activity in the presence of methyl glucoside (0.05 M)
- Catalytic activity after pretreatment of alumina neutral with methyl glucoside for 6 hrs
- Catalytic activity after pretreatment of alumina neutral with methyl glucoside for 24 hrs
[α-D-Glucose] = 0.05 M

[Alumina] = 26.7 mg.ml\(^{-1}\)

- Catalytic activity in the absence of hexanal
- Catalytic activity in the presence of hexanal (0.05 M)
- Catalytic activity after pretreatment of alumina neutral with hexanal for 6 hrs
- Catalytic activity after pretreatment of alumina neutral with hexanal for 24 hrs

Fig. 40  The Effect of Pretreatment of Alumina Neutral with Hexanal on Its Catalytic Activity
not get progressively deactivated during a run). Some further deactivation possibly may be caused by aldehyde functional groups formed during the mutarotation reaction.

In summary, the results discussed in this section have shown that the catalytically active sites are specific for adsorption of polyhydroxy compounds. About 35% of the activity can be lost by progressive deactivation of the catalyst by reaction of hydroxyl groups with surface functional groups. The active sites can also interact with aldehyde groups and aldehyde groups may be involved in reactions that cause progressive catalyst deactivation at a much slower rate. Hence, the adsorption and stabilization of functional groups which are essential for surface catalysis of mutarotation also lead to side reactions which cause permanent deactivation of some of the active sites of alumina neutral. These catalytic sites, in addition to possessing adsorption sites for glucose and any intermediate formed, should also possess acid/base functional groups to catalyze mutarotation.
NATURE OF ACID/BASE FUNCTIONAL GROUPS ON ACTIVE SITES
XVIII. NATURE OF ACID/BASE FUNCTIONAL GROUPS ON ACTIVE SITES

It was mentioned in the Introduction that the alumina surface contains many acidic and basic functional groups (viz. $\text{Al}^{3+}$, $\cdot\delta^+$, $\cdot\delta^-$, $\cdot\delta^-$, $\cdot\delta^-$) which are all potential catalysts for glucose mutarotation. It was also mentioned that the use of acid and base sensitive substances as inhibitors has given evidence for the participation of these sites in catalytic reactions. The upper limit of active site density had also been obtained using those inhibitors (see also Section XV.1). To study the nature of sites on alumina involved in glucose mutarotation the same inhibitors were used as described below.

XVIII.1 Test for Basic Sites - Inhibition by Carbon Dioxide

As mentioned in the Introduction, $\text{CO}_2$ reacts with basic functional groups ($\cdot\delta^-$ and $\cdot\delta^-$) on the alumina surface. To test for the effect of $\text{CO}_2$ on alumina neutral, it was first evacuated at room temperature (for 2 days) to remove excess water adsorbed on the surface. Dry $\text{CO}_2$ was passed through a portion of the evacuated sample packed in a glass column for 24 hours. The control sample was prepared by passing dry nitrogen through another portion of the evacuated sample for 24 hours. The activities were determined by stirring 1.6 g of each sample with 60 ml 0.05 M $\alpha$-D-glucose solution. The results showed that alumina neutral has been deactivated 10% by treatment with $\text{CO}_2$ (Table XVII).

The procedure was repeated with aluminas pyrolyzed at 800°C and at 1250°C. However, with these samples the controls were not treated with dry nitrogen since they are already free of adsorbed water molecules. The results showed that alumina pyrolyzed at 800°C has been deactivated 27% and
alumina pyrolyzed at 1250°C has been deactivated 85% by CO₂ (Table XVII).

TABLE XVII

EFFECT OF CO₂ ON DIFFERENT ALUMINAS

<table>
<thead>
<tr>
<th>Alumina</th>
<th>Percentage Deactivation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Alumina Neutral</td>
<td>10 ± 5</td>
</tr>
<tr>
<td>Pyrolyzed at 800°C</td>
<td>27 ± 7</td>
</tr>
<tr>
<td>Pyrolyzed at 1250°C</td>
<td>85 ± 5</td>
</tr>
</tbody>
</table>

* Errors were estimated from maximum and minimum slopes

The partial inhibition of catalytic activity of dehydrated aluminas by CO₂ should be due to the formation of carbonate ions (which are much less active than oxide ions) on the surface. Hence, it should be possible to decompose the carbonate ions by heating the CO₂ treated sintered alumina to a low temperature, and get back the original activity. It was done by heating the CO₂ treated sintered sample, at 650°C for 3 hrs, which is a "mild" heat treatment compared with the original dehydration at 1250°C for 6 hrs. The results in Fig. 41 show that the deactivated sample has regained its original activity after "mild" heat treatment. However, attempts to detect by micro-analysis an increase of carbon on sintered alumina after it has been treated with CO₂, were not successful probably
\[ [\alpha-D-Glucose] = 0.05 \text{ M} \]
\[ [\text{Alumina}] = 26.7 \text{ mg.ml}^{-1} \]

Catalytic activity of alumina sintered at 1250°C

Catalytic activity of alumina sintered at 1250°C and treated with carbon dioxide

Catalytic activity of alumina sintered at 1250°C, treated with carbon dioxide, and heated at 650°C for 3 hrs.

Fig. 41 The Effect of Carbon Dioxide on the Catalytic Activity of Alumina Sintered at 1250°C for 6 hours
because of the low sensitivity of the technique.

The results of inhibition by CO₂ given in Table XVII show that the fraction of activity due to basic sites has increased with increase of dehydration of the alumina surface. According to the model of the alumina surface discussed in the Introduction, basic hydroxide ions are removed and oxide ions are formed as the surface is dehydrated. This suggests that the basic catalytic sites on the alumina surface should be oxide ions. Whether they are individual oxide ions or defect sites cannot be determined without further tests.

From the deactivating effects of carbon dioxide it is also clear that the high activity per unit area for sintered alumina (reported in Section XII) is because the active sites are predominantly basic sites. The observed increase in activity per unit area as alumina was pyrolyzed at temperatures >600°C, appears to be due to a large increase in the number of basic sites.

Since excess CO₂ has not been able to deactivate any of the aluminas completely, it appears that there are catalytic sites on alumina which are not poisoned by CO₂. Since the size of the CO₂ molecule is smaller than that of the glucopyranose molecule, the inhibitor (CO₂) should be able to reach all active sites accessible to the substrate (glucose). Hence the sites which are not poisoned by CO₂ may be acidic sites or, less likely, they are weak basic sites which are too weak to react with CO₂.

XVIII.2 Tests for Acidic Sites on Alumina Neutral

XVIII.2.1 Effect of Pyridine

As discussed in the Introduction, pyridine reacts with Lewis acid
sites and relatively strong Brönsted acid sites. Also, the size of the pyridine molecule is comparable to that of the glucopyranose molecule and hence, all the active sites should be sterically accessible to pyridine.

Tests of possible inhibition by pyridine was carried out by stirring 1.6 g alumina neutral with 10 ml DMSO containing 0.04 ml (0.5 mM) pyridine. After 15 mins, a solution of 0.54 g α-D-glucose in 50 ml DMSO was added (so that the ratio of \( \frac{\text{moles glucose}}{\text{moles pyridine}} = \frac{6}{1} \)) and the mutarotation was followed as usual. The results showed that the inhibitory effect of pyridine was negligible (±2-3%). The slight variation in rate was within experimental error. Homogeneous catalysis by pyridine not adsorbed on the alumina was also shown to be absent.

The procedure was repeated with 0.24 ml (3 mM) of pyridine so that the ratio \( \frac{\text{moles glucose}}{\text{moles pyridine}} = 1 \). Again, negligible (3-4%) inhibitory effect was observed. Further, there was almost no change in the amount of glucose adsorbed and homogeneous catalysis was, again, negligible. Since pyridine reacts with Lewis acid sites and relatively strong Brönsted acid sites, such sites cannot be involved in the catalytic mutarotation of aluminas.

In order to study the effect of direct treatment of alumina by a potential inhibitor, 0.24 ml (3 mM) pyridine was added onto 1.6 g alumina, the flask was sealed and the sample was mixed for 15 mins. The inhibitory effect of direct treatment was again found to be negligible (±4%) and there was almost no change in the amount of glucose adsorbed.

Since all the above experiments showed negligible effect on the catalytic activity, the effect of prolonged treatment with pyridine was investigated by stirring 1.6 g alumina neutral with 10 ml DMSO containing 0.24 ml pyridine for 24 hrs. In this case only about a 10% decrease in activity and about a 5% decrease in glucose adsorption was observed (Note
that about 90% of activity of alumina neutral is due to acidic sites).

These results show that Lewis acid sites on alumina neutral are not catalytically active and any catalytically active Brønsted acid sites are too weak to react with pyridine. Therefore, the inhibitory effect of a still stronger base was investigated next.

XVIII.2.2 Effect of n-Butylamine

As mentioned in the Introduction, n-butylamine is a stronger base than pyridine and should be sterically able to reach all the sites on an alumina surface which are active towards glucose mutarotation.

Tests of the inhibition by n-butylamine was carried out by stirring 1.6 g alumina neutral with 10 ml DMSO containing 0.03 ml (3 x 10^{-4} mole) n-butylamine. After 150 mins, 50 ml DMSO containing 0.54 g α-D-glucose (so that \( \frac{\text{moles glucose}}{\text{moles n-butylamine}} = \frac{10}{1} \)) was added and kinetics was followed. The samples removed from the reaction flask showed a rapid decrease in optical rotation with time, while the filtered slurries also continued to show a slow decrease in optical rotation (i.e. in the absence of alumina). Hence there was concurrent homogeneous and heterogeneous catalysis, as is shown in the first order plots given in Fig. 42. The plot (a) was obtained by measuring the optical rotations of samples as soon as they were filtered and hence, is the result of simultaneous homogeneous and heterogeneous catalysis. The slope at \( t = 25 \text{ mins} \) equal to \( 3.83 \times 10^{-4} \text{ sec}^{-1} \), is the overall (homogeneous + heterogeneous) rate constant. The first order plots (b) and (c) for homogeneous catalysis were obtained by following the change in optical rotation of filtered slurries with time, and are linear and parallel as expected. From plots (b) and (c) the average rate constant for homogeneous catalysis by the butylamine in solution is \( 0.28 \times 10^{-4} \text{ sec}^{-1} \).
Fig. 42 The Effect of n-Butylamine on the Catalytic Activity of Alumina Neutral
Therefore the rate constant for heterogeneous catalysis is $3.55 \times 10^{-4} \text{ sec}^{-1}$ 
$\approx 3.6 \times 10^{-4} \text{ sec}^{-1}$. Under similar conditions, but in the absence of any 
additive, 1.6 g alumina showed an initial rate constant of $2.5 \times 10^{-4} \text{ sec}^{-1}$. 
Hence, n-butylamine has increased the catalytic activity of alumina neutral 
by about 44%.

It was mentioned in the Introduction that n-butylamine reacts with Lewis acid sites and also Brönsted acid sites. It was shown above by the 
effect of pyridine that Lewis acid sites are not active in glucose mutarotation. Further, the reaction of n-butylamine with a Lewis acid site will 
produce an inactive site and hence should decrease the catalytic activity. 
Hence, the increase in activity by n-butylamine should arise from its reaction 
with Brönsted acid sites and the catalytic activity of the resulting site. 
The reaction of n-butylamine with a Brönsted acid site will produce a more 
active basic site and therefore, as observed, there should be an increase 
in the catalytic activity.

\[
\begin{align*}
\text{Acidic (Type III)} \\
\text{hydroxyl group}
\end{align*}
\]

Hence, this increase in activity on treatment with n-butylamine shows 
that there are catalytically active Brönsted acid sites which are too weak
to react with pyridine. They are, however, reactive with n-butylamine and become more catalytically active. It was mentioned in the Introduction (page 20) that Pearson using a sensitive n.m.r. method showed the presence of protonated species when pyridine was adsorbed on alumina at 0°C. Further, as mentioned in the Introduction (page 35), the catalytic activity of Brönsted acid sites on alumina had been detected only at relatively high temperatures (≈200°C). With use of glucose it has been possible to detect, at 25°C, the catalytic activity of Brönsted acid sites that are too weak to react with pyridine, and hence not detectable by the n.m.r. method of Pearson. This appears to be the only example of a reaction catalyzed by weak Brönsted acid sites on alumina at 25°C. Therefore it follows that glucose mutarotation is more sensitive to Brönsted acid sites than any of the other reactions that have been studied on alumina so far. Further, glucose mutarotation is a more sensitive probe for weak Brönsted acid sites on alumina (or any other solid catalyst) than any of the spectroscopic methods used so far.

It is very likely that the weak Brönsted acid sites on alumina neutral are involved in the 90% of its activity which is not affected by CO₂. It was observed in Section XVI.2 that the catalytic constant for alumina neutral is about 10 times greater than that for strong Brönsted acids in water or DMSO (i.e. homogeneous systems). Generally, the catalytic activity of weak acids (like acetic acid) in water is only about one hundredth that for strong Brönsted acids. Hence, it seems especially notable, that the activity of weak Brönsted acid sites on the alumina surface is greater than that of even strong Brönsted acids in water. The high catalytic activity of weak sites when present on the alumina surface may be due to two factors, (1) Strong specific adsorption of glucose molecules on the alumina surface with the acidic group in the correct orientation to protonate the ring
oxygen of glucose and/or (ii) Stabilization of the transition state leading to the acyclic form by the functional groups on the alumina surface.

It was observed earlier that active sites on alumina can stabilize an aldehyde functional group while, as mentioned in the Introduction, there appears to be very little interaction (e.g. no oxygen atom exchange) between solvent (water) and the acyclic form of glucose during mutarotation. In these respects the alumina surface resembles enzymes in its catalytic properties. Whatever the reason for it may be, it may be concluded that the alumina surface offers a better medium for glucose mutarotation than pure water.

XVIII.2.2.1 Determination of the Amount of n-Butylamine on the Surface

It was observed in Fig. 42 that there was also homogeneous catalysis of glucose mutarotation in the heterogeneous system when alumina was treated with n-butylamine. The slope of the first order plot for homogeneous catalysis should be proportional to the concentration of n-butylamine in solution since catalysis by solvent is negligible. A comparison of the rate constant for a homogeneous reaction catalyzed by n-butylamine (without previous addition of alumina) with the rate constant for a homogeneous component of a reaction system with alumina present would give a measure of the n-butylamine concentration in solution. This in turn gives the n-butylamine adsorbed on the surface of alumina.

The rate constant for homogeneous catalysis by n-butylamine, in the absence of alumina, was determined using a 60 ml solution 0.05 M in α-D-glucose and 0.005 M in n-butylamine (Fig. 42 plot (d)). It was observed that filtration of the solution had no effect on the rate of catalysis. This shows that the rate constant for homogeneous catalysis determined
(from plots (b) and (c)) using filtered slurries is equal to the rate constant for homogeneous catalysis in the slurry. The ratio,
\[
\frac{\text{rate constant for homogeneous catalysis in the presence of alumina}}{\text{rate constant for homogeneous catalysis in the absence of alumina}} \approx \frac{0.28 \times 10^{-4} \text{ sec}^{-1}}{0.93 \times 10^{-4} \text{ sec}^{-1}} = \frac{1}{3.3}
\]
Therefore, only \( \frac{1}{3.3} \) of the n-butylamine added is in solution, and \( \frac{2.3}{3.3} \times 3 \times 10^{-4} = 2.2 \times 10^{-4} \) mole n-butylamine has reacted with the surface. This may be used to determine the number of acid sites/cm\(^2\) on the surface of alumina neutral as shown below.

The amount of n-butylamine that has reacted with 1.6 g alumina neutral equals \( 2.2 \times 10^{-4} \text{ mole} = 13 \times 10^{19} \text{ molecules} \).

Therefore the number of molecules that has reacted with a unit area (1 cm\(^2\)) of the sample equals \( \frac{13 \times 10^{19} \text{ molecules}}{140 \times 1.6 \times 10^{4} \text{ cm}^2} = 5.8 \times 10^{13} \text{ molecules cm}^{-2} \) and may be considered as the active site density on alumina neutral (Note that only about 10% of the activity of alumina neutral is due to basic sites and since, in general, the activity of a basic site is several orders of magnitude greater than that of an acid site, the number of acid sites on alumina neutral \( \cong \) the number of active sites). However, it is not possible to state whether it gives the upper limit of catalytic site density since there may be catalytically active Brönsted acid sites which are too weak to react with n-butylamine. However, it is important to note that the site density determined using n-butylamine equals, within experimental error, the site density determined using adsorption isotherms and the initial rate data (\( = 5.4 \times 10^{13} \text{ sites cm}^{-2} \)) in Section XV.

XVIII.2.3 Tetramethylammonium Hydroxide

Since there may be catalytically active Brönsted acid sites which are too weak to react with n-butylamine, the effect of a still stronger base, \( \text{N}^+\text{(Me)}_4^-\text{OH} \), was investigated. Solutions of the hydroxide in DMSO were
prepared and standardized by titrating against HCl as described in the
Experimental Section.

The effect of $\text{N(Me)}_4^+\text{OH}$ was determined by treating alumina neutral
with 10 ml of DMSO containing a known amount of the base. After 150 mins,
50 ml DMSO containing 0.54 g $\alpha$-D-glucose was added and the rate of muta-
rotation was followed. The experiment was repeated with increasing amounts
of the base. For initial experiments 1.6 g of the catalyst was used and
the rate constant for heterogeneous catalysis increased rapidly with increase
in the concentration of base. Therefore, as the amount of base was increased
the weight of the catalyst used was progressively decreased (e.g. 0.32 g,
0.08 g) so that the heterogeneous catalysis could be followed conveniently
and the presence of any homogeneous catalysis could also be detected. The
reproducibility of the results was good and independent of the pretreatment
time at relatively high base concentrations. With increasing base the
homogeneous catalysis was first detected (in the filtered slurry) when
0.08 g alumina neutral was treated with $2.0 \times 10^{-5}$ mole base (Fig. 43, plots
(b) and (c)). No homogeneous catalysis was observed when 0.08 g alumina
was treated with $1.5 \times 10^{-5}$ mole base and much more rapid homogeneous
catalysis was observed when the same weight of the catalyst was treated
with $2.5 \times 10^{-5}$ mole base. The rate constant for homogeneous catalysis by
$0.5 \times 10^{-5}$ mole base in the absence of alumina was determined using 60 ml
solution containing 0.54 g $\alpha$-D-glucose and $0.5 \times 10^{-5}$ mole $\text{N(Me)}_4^+\text{OH}$
(Fig. 43, plot (d)). The results were analyzed, as in the case of n-butyl-
amine treatment, and the amount of base that reacted with 0.08 g alumina
neutral was determined to be $\approx 2.0 \times 10^{-5}$ mole.

The amount of base that reacts with the surface was also determined
by back titrations as described in the Experimental Section. The amount
The Effect of Tetramethylammonium Hydroxide on the Catalytic Activity of Alumina Neutral

Fig. 43

(a) Catalytic activity of 0.08 g alumina neutral treated with $2 \times 10^{-5}$ mole base

(b) and (c) Homogeneous Catalysis in filtered slurries

(d) Homogeneous catalysis by $0.5 \times 10^{-5}$ mole base

60 ml 0.05 M $\alpha$-D-Glucose

$\ln 10^3 (a_e - a_g)$

Time (min)
of base that reacts with 1.6 g alumina neutral was found to be $5.0 \times 10^{-4}$ mole and it was independent of the contact time and also the amount of excess base used.

Therefore, the average amount of base that reacts with the surface of 1.6 g alumina neutral equals $(4.5 \pm 0.5) \times 10^{-4}$ mole and may be used to determine the upper limit of acidic active site density ($\sim$active site density) on the surface. The upper limit of active site density of alumina neutral using the data from $\text{N}$(Me)$_4$ OH treatment turns out to be $(1.2 \pm 0.1) \times 10^{14}$ sites/cm$^2$, which is about double the upper limit obtained using n-butylamine or adsorption isotherms and initial rate data. For comparison, the reported total lattice site density on alumina surface range from $0.9 \times 10^{15}$ to $1.5 \times 10^{15}$ sites cm$^{-2}$. However, it is not surprising that the treatment of alumina with OH ions has given a higher value for the number of acid sites, since it should react with all strong and weak, Brönsted and Lewis acid sites on the surface whether they are catalytically active or not.

It is interesting to note that the catalytic activity of alumina neutral has increased by a factor of $10^2$ on treatment with $\text{N}$(Me)$_4$ OH. Note that half as much n-butylamine reacted with the surface and increased the catalytic activity only 44%. This difference in the increase in activity can be due to many factors. For example, OH ions should have converted all weak Brönsted acid sites (which are too weak to react with n-butylamine and too weak to catalyze mutarotation) to strong conjugate basic sites with very high activity. Thus, new active sites with very high activity could have been formed. Further, the two types of ion pairs formed on the surface by the two bases, viz. $\text{O}^-$ $\text{N}$(Me)$_4$ and $\text{O}^-\text{NH}_3-(\text{CH}_2)_3\text{CH}_3$ could have different activities towards mutarotation due to the differences in the accompanying cations. For example, steric effects of the four carbon
chain on n-butylamine and the decrease in the basicity of the anion by hydrogen bonding to \(\text{NH}_3^-\) can decrease the catalytic activity of the ion pairs formed from n-butylamine.

XVIII.2.4 Deactivation of Brönsted Acid Sites - Effect of CH\(_2\)N\(_2\)

During the studies described above on the effect of bases on catalytic activity, evidence has been gathered for the presence of catalytic activity due to weak Brönsted acid sites. Since the reaction of a Brönsted acid site with a base produces the conjugate base with higher catalytic activity towards mutarotation, the catalyst produced by treatment with a relatively strong base always had a higher activity than the original catalyst. On the other hand, to deactivate the Brönsted acid sites, (without converting them to basic sites) methylation by treating alumina with diazomethane in ether solution was attempted. However, the diazomethane decomposed on the alumina surface and surprisingly producing a catalyst with slightly higher activity than the control samples. This may be due to breaking up of the surface by diazomethane decomposition; however, this alumina sample was not investigated further.

XVIII.3 Deactivation of Brönsted Acid and Basic Sites - Effect of \((\text{CH}_3)_2\text{SiCl}_2\)

Dichloromethylsilane should react with both hydroxyl groups and the basic sites on the alumina surface producing a new surface without any acidic or basic groups (Equations 48, 49, 50).\(^{132}\)

Silylation of the surface of alumina neutral was carried out as described in the Experimental Section. The silylated aluminum oxide had no catalytic activity but adsorbed about the same amount of glucose over
several hours.

\[ \text{CH}_3\text{SiCl}_2 + \text{OH} \rightarrow \text{SiCH}_3 + 2 \text{HCl} \]  

(48)

\[ \text{CH}_3\text{SiCl}_2 + \text{OH} \rightarrow \text{SiClCH}_3 + \text{HCl} \]  

(49)

\[ \text{CH}_3\text{SiCl}_2 + \text{O}^- \rightarrow \text{SiOCH}_3 + \text{HCl} \]  

(50)
In summary, the specific poisoning of the catalyst surface has given evidence for the presence of basic (probably oxide ions) and weak Brönsted acid sites on the surface. It was mentioned in the Introduction that acids and bases show different deuterium isotope effects for mutarotation in solution. Further, the study of deuterium isotope effect has given information on the mechanism (for example, specific or general acid/base catalysis, bifunctional catalysis) in solution. Therefore, deuterium isotope effect for glucose mutarotation by alumina was investigated in the next section.
DEUTERIUM ISOTOPE EFFECT ON GLUCOSE MUTAROTATION BY ALUMINA
XIX.1 Deuterium Isotope Effect for Catalysis by 800°C Alumina

Alumina dehydrated at 800°C was chosen for detailed studies of glucose mutarotation because it has advantages over other aluminas. For example, in comparison to alumina sintered at 1250°C, it gives better reproducibility even in the presence of water (see Sections XII and XIII.3). Relative to alumina neutral it also has less hydroxyl groups on the surface and hence, it is easy to swamp out the hydroxyl protons with D₂O. The deuterium isotope effect for catalysis by 800°C alumina was studied both in the presence of water (or D₂O) and under anhydrous conditions as described below.

XIX.1.1 Deuterium Isotope Effect under Anhydrous Conditions

Deuterated 800°C alumina and a control sample of 800°C alumina were prepared as described in the Experimental Section. Sixty ml 0.05 M solution of glucose-O-D (i.e. O-deuterated α-D-glucose) in DMSO was stirred with 0.8 g of deuterated 800°C alumina and the kinetics was followed as usual. However, the first measurement of optical rotation was made only after 25 mins to allow sufficient time for lumpy particles of alumina (see Experimental Section) in the sample to break down. The kinetics was also followed for mutarotation of 60 ml 0.05 M solution of glucose-O-H (standard α-D-glucose) in DMSO by 0.8 g of the control sample of 800°C alumina. The two first order plots are shown in Fig. 44, and it is clear that O-deuteration has decreased the observed rate constant. The deuterium isotope effect, from the slopes between t = 30 and 40 mins, is \( \frac{k_H}{k_D} = 1.4 \pm 0.1 \).
[α-D-Glucose] = 0.05 M

[Alumina] = 13.3 mg.ml⁻¹

Fig. 44 Deuterium Isotope Effect with 800°C Alumina Under Anhydrous Conditions
XIX.1.2 Deuterium Isotope Effect in the Presence of Water/Deuterium Oxide

To study the deuterium isotope effect in the presence of water, hydrated 800° alumina was made by adding 7% water to a sample of 800° alumina and 7% D₂O was added to another sample of 800° alumina (Note that there was >6% loss in weight on heating alumina neutral to 800°C). The samples were then mixed well in a minimill. To swamp out the rapidly exchangeable hydroxyl protons on glucose it was decided to add 10 times as many deuterons (as added D₂O). Since 60 ml 0.05 M glucose solution contains 0.003 mole glucose, there are 0.015 mole of exchangeable protons in the solution. Therefore 0.15 mole deuterons are required and they are present in 1.35 ml D₂O.

To 60 ml 0.05 M glucose solution in DMSO, 1.35 ml D₂O was added and was stirred with 0.8 g of 800° alumina treated with D₂O (as described above). The kinetics was followed as usual and the first order plot is given in Fig. 45. The control run was carried out by adding 1.35 ml water to 60 ml 0.05 M glucose solution and stirring with 0.8 g of 800° alumina treated with water. In both the runs no homogeneous catalysis was detected in the filtered slurries (Note that the mole fraction water = 0.09 and homogeneous catalysis due to water is observed only when mole fraction of water is >0.765). The two plots in Fig. 45 again show that $\frac{k_H}{k_D}$ in the presence of H₂O/D₂O is 1.4 ± 0.1.

To check the reproducibility of the results, the experiment was repeated with 2.0 ml H₂O or D₂O (i.e. with 15 times the exchangeable protons in glucose) instead of 1.35 ml. The results in Fig. 45 again show that $\frac{k_H}{k_D} = 1.5 ± 0.1$. Further, no homogeneous catalysis was observed in the filtered slurries.

The experiment was again repeated, this time adding all the D₂O or H₂O to the DMSO. Since 0.8 g of hydrated 800° alumina contains 7% water
Fig. 45  Deuterium Isotope Effect with 800°C Alumina in the Presence of H₂O/D₂O

With 0.75 g anhydrous 800°C alumina and 1.40 ml H₂O (○) or D₂O (●)

With 0.80 g hydrated 800°C alumina and 1.35 ml H₂O (○) or D₂O (●)

With 0.80 g hydrated 800°C alumina and 2.0 ml H₂O (□) or D₂O (■)

ln 10³ (α₁/α₀)

Time (min)
or D\textsubscript{2}O (≈0.05 g water or D\textsubscript{2}O), 0.75 g (= 0.80 - 0.05) anhydrous 800° alumina was used and 1.40 ml (= 1.35 + 0.05) D\textsubscript{2}O or water was added to the 60 ml 0.05 M glucose solution. The results in Fig. 45 again show that \( \frac{k_{H}}{k_{D}} = 1.4 \pm 0.1 \).

These results show that the same normal isotope effect has been observed for glucose mutarotation by 800° alumina in the presence and in the absence of water. It is also clear from Fig. 45 that addition of 7% water to 800° alumina has increased its activity ≈340%. In fact the activity of water treated 800° alumina is greater than that of the original alumina neutral. This should probably be due to the presence of a greater percentage of basic sites on the 800° alumina.

XIX.1.3 Interpretation of the Observed Isotope Effect

Before these results can be interpreted and compared with the known values for the isotope effect by homogeneous catalysts, it is important to determine the individual steps in the catalytic reaction that give rise to the observed isotope effect.

As mentioned in the Introduction, the observed rate constant for homogeneous catalysis (Equation 51) is \( k_{obs} = (k_{1} + k_{2}) c \)

\[
S + C \xrightleftharpoons[k_{2}]^{k_{1}} P + C
\]

where \( c \) is the catalyst concentration, and \( k_{1} \) and \( k_{2} \) are the rate constants for forward and reverse reactions. Hence the observed deuterium isotope effect,
\[ \frac{k_H}{k_D} = \frac{(k_1^H + k_2^H)}{(k_2^D + k_2^D)} \]  

and is the result of the isotope effect on rate constants \( k_1 \) and \( k_2 \). It was shown in Section VII that the observed rate constant for catalysis by alumina (Equation 26)

\[ S + C \underset{k_2}{\xleftrightarrow{}} SC \xrightarrow{k_3} PC \xrightarrow{k_4} P + C \]  

is

\[ k_{obs} = \frac{k_1 k_2 (k_3 + k_4) c_o}{k_2^2 + k_2 k_4 + k_2 k_3 + k_1 (k_2 + k_3 + k_4) s_o} \]  

where \( c_o \) = total catalyst concentration and \( s_o \) = total substrate concentration. Since \( k_1 \) and \( k_2 \), the rate constants for adsorption and desorption are expected to be greater than \( k_3 \) and \( k_4 \), the rate constants for forward and reverse surface reactions (see Section VII), the Equation 30 simplifies to

\[ k_{obs} \approx \frac{(k_3 + k_4) c_o}{k_2/k_1 + s_o} \]  

Hence, the observed deuterium isotope effect reflects the deuterium isotope effect on all four steps (both, adsorption and desorption steps and surface catalytic steps) unlike the homogeneous reaction. However, it is obvious that for comparison of the mechanisms of homogeneous and heterogeneous reactions one should compare the isotope effects for the steps that involve transformation of substrate to product (i.e. \( (k_3 + k_4) \) for the
heterogeneous process in Equation 26 with \((k_1 + k_2)\) for the homogeneous reaction in Equation 51).

In order to carry out this comparison it is convenient to consider first, the simplified Equation 31 which contains the equilibrium constant for adsorption on catalytic sites. According to Equation 31, the observed deuterium isotope effect is the result of isotope effect on (i) the catalytic constant for the surface reaction \((k_3 + k_4)\) and (ii) on the equilibrium constant for adsorption on catalytic sites \((k_1/k_2)\). The observed final optical rotations were slightly higher when D\(_2\)O was used instead of water (e.g. 0.131° when 1.35 ml water was used and 0.132° when 1.35 ml D\(_2\)O was used, and 0.130° when 1.40 ml water was used and 0.131° when 1.40 ml D\(_2\)O was used) but the difference is not significant because the precision of the measurements is ±0.001°. Further, more accurate measurement of optical rotation, for example, with a 1 dm path length cell at 25.0°C may not be very useful because we need to determine the isotope effect on adsorption at active sites and not the isotope effect on total adsorption.

From the above discussion it is clear that it is difficult to determine whether there is an isotope effect on the equilibrium constant for adsorption on catalytic sites. A more direct approach to the interpretation of observed isotope effect is to determine whether there is any isotope effect on the rate constants \(k_1\) and \(k_2\) by studying the adsorption/desorption process in the absence of the catalytic process. If there is no isotope effect on the rate constants \(k_1\) and \(k_2\), then the observed isotope effect should be due to isotope effect on the catalytic constant for the surface reaction \((k_3 + k_4)\) and can be directly compared with the isotope effect observed for homogeneous reactions. On the other hand, if there is an isotope effect on adsorption and desorption, then the interpretation of isotope
effect results would be difficult.

XIX.1.3.1 Deuterium Isotope Effect on the Rate of Adsorption of Methyl α-D-Glucoside

The study of isotope effect on the rate of adsorption was carried out using methyl α-D-glucoside. This compound has essentially the same set of hydroxyl groups for adsorption on alumina surface and is very similar to glucose adsorption as represented by the first step in Equation 26. Methyl glucoside deactivated alumina neutral towards glucose mutarotation (Section XVII) and hence, it should get adsorbed on the catalytically active sites. Further, it does not undergo mutarotation and therefore any complications due to a surface reaction and adsorption of intermediates is eliminated (see below).

A 100 ml solution, 0.05 M in methyl α-D-glucoside was made by dissolving 0.97 g of the glucoside in DMSO, adding 2.3 ml H₂O (so that ratio of exchangeable protons in H₂O : exchangeable protons in glucoside ≈ 13) and making up to the mark with DMSO. Sixty ml of the solution was stirred with 1.6 g 800° alumina and the change in optical rotation with time was measured (Fig. 46) as described in the Experimental Section. The procedure was repeated using a solution containing 2.3 ml D₂O instead of water. The results in Fig. 46 clearly show that there is no deuterium isotope effect on the rate of adsorption of methyl-α-D-glucoside.

The rate of adsorption of the methyl glucoside from DMSO solution, in the presence of water, is given by

\[
\text{Rate}_{(H_2O)} = k_a^H [\text{Methyl glucoside-OH}] [\text{Catalyst}] - k_d^H [\text{Methyl glucoside-OH-Catalyst Complex}]
\] (53)
Fig. 46  Deuterium Isotope Effect on the Rate of Adsorption of Methyl α-D-Glucoside on 800°C Alumina
and the rate in the presence of D$_2$O is,

$$\text{Rate}_{(D_2O)} = k^D_a [\text{Methyl glucoside - OD}] [\text{Catalyst}] - k^D_d [\text{Methyl glucoside - OD - Catalyst Complex}]$$ (54)

where $k_a$ is the rate constant for adsorption of methyl glucoside onto alumina surface and $k_d$ is the rate constant for desorption. Since the two plots in Fig. 46 coincide at all times it follows that Equation 53 equals Equation 54 at all concentrations of glucoside, catalyst, and glucoside-catalyst complex. Hence $k^H_a = k^D_a$ and $k^H_d = k^D_d$ and there is no deuterium isotope effect on adsorption and desorption of glucoside on alumina surface. Since adsorption of methyl glucoside and adsorption of glucose onto alumina surface should both involve hydroxyl groups it may be concluded that there should not be any deuterium isotope effect for adsorption of glucose (i.e. $k^H_1$ should equal $k^D_1$ and $k^H_2$ should equal $k^D_2$).

XIX.1.3.2 Deuterium Isotope Effect on the Rate of Adsorption of Glucose

It was mentioned earlier, that methyl α-D-glucoside was chosen to study the deuterium isotope effect on adsorption by hydroxyl groups since it does not undergo mutarotation on an alumina surface and hence, any complications due to surface reactions and adsorption of intermediates do not occur. At this stage it is of interest to determine how such additional factors would affect deuterium isotope effect on adsorption by studying the isotope effect on the adsorption of glucose onto alumina as described below.

A 100 ml of 0.05 M equilibrated glucose solution was prepared using 20 ml 0.25 M equilibrated glucose solution, 2.7 ml H$_2$O (so that the ratio of exchangeable protons in water to exchangeable protons in glucose $\approx 12$) and making up to the mark with
DMSO. Sixty ml of the solution was stirred with 1.6 g 800° alumina and the rate of change of optical rotation with time was followed as before (Fig. 47). Another 100 ml, 0.05 M equilibrated glucose solution was prepared using 2.7 ml D_2O (instead of water) and the rate of adsorption was again measured. The results in Fig. 47 show that there is a small normal deuterium isotope effect for adsorption of glucose onto alumina. Similar normal isotope effect was also observed when the rate of adsorption was measured with an equilibrated 0.25 M solution. This observation of a normal isotope effect for glucose adsorption may be due to an additional surface reaction undergone by glucose as explained below.

It was mentioned in Section VII.1 that the catalyst-glucose complexes on alumina surface should consist not only of glucopyranose molecules adsorbed via hydroxyl groups but also a small amount of adsorbed intermediates (e.g. the acyclic form). Hence the adsorption process was represented by

\[
\begin{align*}
\text{Glucopyranose + Catalyst} & \rightarrow_{k_1} \text{Glucopyranose-Catalyst Complex} \\
& \rightarrow_{k_2} \text{Glucopyranose-Catalyst Complex} \\
& \rightarrow_{k_3'} \text{Intermediate-Catalyst Complex} \\
& \rightarrow_{k_4'} \text{Intermediate-Catalyst Complex}
\end{align*}
\]

Thus, initial rapid adsorption leads to formation of a glucopyranose-catalyst complex adsorbed via hydroxyl groups (similar to methyl glucoside-catalyst complex in Equation 53 or 54). However, slow conversion of glucopyranose molecules to the intermediate species by step 3 will decrease the concentration of glucopyranose-catalyst complex. To maintain the equilibrium of the first step, more glucose will get adsorbed onto the surface at a rate equal to the rate of the surface reaction by step 3. Hence any deuterium isotope effect for the rate constant \(k_3'\) will be reflected in the rate of adsorption of glucose from solution. Hence a normal isotope effect for
Deuterium Isotope Effect on the Rate of Adsorption of Glucose on 800°C Alumina

[Glucose] = 0.05 M
[Alumina] = 26.7 mg.ml$^{-1}$
will result in a normal isotope effect for adsorption of glucose, and vice versa. This explains the observed normal isotope effect for glucose adsorption although deuterium substitution did not affect the rate of adsorption of methyl glucoside onto alumina.

From the above studies it is reasonable to conclude that for adsorption of glucopyranose onto alumina, $k_1^H = k_1^D$ and $k_2^H = k_2^D$. This conclusion is supported by the fact that no deuterium isotope effect has been observed for dehydration of alcohols on alumina when O-deuterated alcohols were used\textsuperscript{49}. Note that adsorption of alcohols onto alumina also occurs via hydroxyl groups.

Since there is no isotope effect on the rate constants for adsorption and desorption of glucose it follows from Equation 30 that the observed isotope effect for glucose mutarotation $\frac{k_{obs}^H}{k_{obs}^D} = \frac{k_3^H + k_4^H}{k_3^D + k_4^D}$ and hence is the result of deuterium isotope effect on rate constants $k_3$ and $k_4$ for forward and reverse surface reactions (c.f. Equation 52 obtained for homogeneous reaction). Hence the observed isotope effect should be directly comparable to those obtained for mutarotation in solution.

XIX.1.3.3 Mechanistic Conclusions

From the discussion of deuterium isotope effect on glucose mutarotation given in the Introduction (Section I.2.4), the observed normal isotope effect of about 1.4 indicates that there is general acid/base catalysis by active sites on alumina surface. Further, the observed isotope effect lies between those observed for bases (2.4 for acetate, 2.9 for pyridine etc.) and strong acids ($k_H/k_D = 1.37$) in solution and rules out any bifunctional concerted catalysis by 800° alumina. It was mentioned in the Introduction that bifunctional catalysis is observed in non-polar media.
which cannot stabilize an ionic transition state. The polar alumina surface should be able to stabilize any ionic transition states formed during mutarotation and hence, catalysis by alumina surface occurs by a stepwise mechanism. As suggested by experiments with methyl glucoside (Section XVII (d)), the normal isotope effect also rules out carbonium ion mechanism for alumina catalysis of glucose mutarotation. The similarity of the observed isotope effect to those observed for homogeneous catalysis also suggests that surface catalysis also occurs by the same mechanism (via an acyclic intermediate). The ability of the catalytic sites to stabilize an aldehyde functional group, and hence the transition state leading to it, was observed in Section XVII.

XIX.1.3.3.1 Effect of Water on Catalytic Activity

As described earlier in this section, the same deuterium isotope effect (within experimental error) was observed in the presence and in the absence of water while the activity of 800° alumina has increased about 340% on the addition of 7% water onto the surface (Note that the activity of hydrated 800° alumina is greater than that of the original alumina neutral probably because of the higher percentage of basic sites). It was mentioned in the Introduction that the deuterium isotope effect for catalysis by water is about 3.9 and is due to concerted mechanism involving two or three water molecules in the transition state. Hence, if the increase in activity of 800° alumina on treatment with water is due to a separate mechanism involving several water molecules, then there should be an increase in the observed deuterium isotope effect. The absence of an independent mechanism involving several water molecules was also indicated by the observed linear relationship between the activity of water-treated 150° alumina and the amount of water added (see Section XII.2.1).
Further, any bifunctional catalysis involving a water molecule and an acidic or basic site on the surface also should increase the isotope effect observed with hydrated alumina. Therefore, the increase in activity observed when 800° alumina was treated with $\text{H}_2\text{O}$ or $\text{D}_2\text{O}$ should be due to an increase in the number of active sites by water molecules acting as a medium to transfer protons from acidic sites or to basic sites from the adsorbed glucose molecules. In other words, on hydrated alumina, many potential catalytic sites are not active because the acidic and basic functional groups on the surface and the adsorbed glucose molecules are not in proper orientation for proton transfer to take place. Adsorption of water onto the surface bridges the gap and helps to transfer protons between functional groups and glucose molecules (Note, however, as observed in Section XV, that the surface of a hydrated alumina, like alumina neutral, still contains reversible adsorption sites which are not made active by adsorbed water).

It was mentioned in Section XIII.2.1.1 that treatment with water did not increase the amount of glucose adsorbed on 150° alumina. A small (~10%) increase in the amount of glucose adsorbed was observed on treating 800° alumina with water. However, the contribution of the increased glucose adsorption to about 340% increase in activity probably is negligible.

XIX.2  Deuterium Isotope Effect for Catalysis by Sintered Alumina

The deuterium isotope effect with alumina sintered at 1250°C was obtained only under anhydrous conditions, since as mentioned in Section XIII.3, the reproducibility of the results in the presence of water is not good. Deuterated sintered alumina was not prepared since alumina dried at >900°C contains <1% of the surface covered with hydroxyl groups. Since reproducible
results with alumina sintered at 1250°C can be obtained only when the kinetic runs are carried out within a short time (see Section XII), all the necessary experiments for isotope effect determinations were carried out within a few hours as described below.

A batch (10 g) of alumina sintered at 1250°C for 6 hrs was prepared and was mixed well in a minimill. Sixty ml 0.05 M solution of glucose-O-H in DMSO was stirred with 1.6 g sintered alumina and the kinetics was followed as usual. The procedure was repeated about an hour after completing the first run and the same results were obtained. A third kinetic run was carried out immediately afterwards with 60 ml 0.05 M glucose-O-D in DMSO and 1.6 g sintered alumina and an isotope effect of 1.92 ± 0.06 was obtained (Fig. 48). Another kinetic run, using 60 ml 0.05 M glucose-O-H and 1.6 g sintered alumina, was carried out soon afterwards and gave a plot superimposable with the first two plots and showed that the catalyst had not deactivated over the 10 hour period.

XIX.3 Deuterium Isotope Effect with Alumina Neutral

A sample of catalyst free of excess water adsorbed on it was prepared by evacuating 3 g of alumina neutral for 24 hours. It was then mixed well in a minimill. Two ml water (ratio exchangeable protons in water exchangeable protons in glucose ≈ 15) was added to 60 ml 0.05 M solution of α-D-glucose in DMSO and was stirred with 0.8 g of the prepared sample. The kinetics was followed as usual. The procedure was repeated with 2 ml D₂O instead of water. In both cases no homogeneous catalysis was observed in the filtered slurries. From the first order plots the deuterium isotope effect was determined to be 1.3 ± 0.1.
Fig. 48  Deuterium Isotope Effect with Alumina Sintered at 1250°C for 6 hours
XIX.4 **Relationship between the Observed Deuterium Isotope Effect and the Percentage of Activity due to Basic Sites.**

Examination of the values for the deuterium isotope effect observed with different alumina samples shows (Table XVIII) that the isotope effect increases with the increase of percentage activity due to basic sites as determined by the effect of CO₂ on activity. A plot of deuterium isotope effect versus percentage deactivation by CO₂ in Fig. 49 shows that there is almost linear correlation between the two.

**TABLE XVIII**

**DATA FOR THE RELATIONSHIP BETWEEN THE OBSERVED DEUTERIUM ISOTOPE EFFECT AND PERCENTAGE OF ACTIVITY DUE TO BASIC SITES**

<table>
<thead>
<tr>
<th>Sample of Alumina</th>
<th>Deuterium Isotope Effect $k_H/k_D$</th>
<th>Percentage Deactivation by Carbon Dioxide</th>
</tr>
</thead>
<tbody>
<tr>
<td>Alumina Neutral</td>
<td>1.3 ± 0.1</td>
<td>10 ± 5</td>
</tr>
<tr>
<td>800°C Alumina</td>
<td>1.43 ± 0.05</td>
<td>27 ± 7</td>
</tr>
<tr>
<td>1250°C Alumina</td>
<td>1.92 ± 0.06</td>
<td>85 ± 5</td>
</tr>
</tbody>
</table>

To test this relationship a new sample of alumina which gets deactivated 53 ± 7% by CO₂ was prepared by heating a different batch of alumina neutral at 800°C for 4 hrs. According to the correlation in Fig. 49 it should show a deuterium isotope effect of about 1.65. The isotope effect observed using $H_2O/D_2O$ was 1.7 ± 0.1. Hence it follows that the percentage of catalytic activity towards mutarotation by the basic sites on alumina
Fig. 49  Correlation of the Observed Deuterium Isotope Effect with the Percentage Deactivation by Carbon Dioxide
can be predicted from the deuterium isotope effect studies described earlier in this section.

From the correlation diagram it is also clear that the isotope effect due to a sample of alumina which undergoes complete deactivation by CO\(_2\) would be about 2.1 while the isotope effect when CO\(_2\) has no effect on activity is about 1.2. Hence the correlation diagram indicates that the two types of sites observed during the study of the effect of inhibitors in Section XVIII show different isotope effects; 2.1 due to the basic sites and 1.2 due to the acidic sites.

It was mentioned earlier in this section that the observed deuterium isotope effect indicates the absence of bifunctional catalysis by acidic and basic sites on the alumina surface. The absence of a maximum in the correlation diagram when both acidic and basic sites are present on the alumina surface also indicates absence of bifunctional catalysis. Hence it follows that the acidic and basic sites act independently by a consecutive mechanism in catalyzing the mutarotation of glucose molecules adsorbed on the surface. Therefore the correlation diagram should reflect the independent behaviour of the two types of sites.

XIX.4.1 Derivation of a Relationship Between the Observed Isotope Effect and the Percentage Activity due to Basic Sites

A theoretical relationship between the observed isotope effect and the fraction of the activity due to basic sites, when the acidic and basic sites act independently, can be derived as follows.
Let $A$ = rate constant (in the presence of OH) when all sites are acidic sites

$B$ = rate constant (in the presence of OH) when all sites are basic sites

$f_A$ = fraction of sites that are acidic

$f_B$ = fraction of sites that are basic

Since the isotope effects due to acidic and basic sites are 1.2 and 2.1, respectively, the rate constant due to acidic sites in the presence of -OD is $A/1.2$ and due to basic sites is $B/2.1$.

Therefore, the observed rate constants are

$$k_H = A f_A + B f_B$$

and

$$k_D = \frac{A}{1.2} f_A + \frac{B}{2.1} f_B$$

if the acidic and basic sites act independently. Hence the observed deuterium isotope effect is given by

$$\frac{k_H}{k_D} = \frac{\frac{A}{1.2} f_A + \frac{B}{2.1} f_B}{A f_A + B f_B}$$

and the fraction of activity due to basic sites is given by

$$F_B = \frac{B f_B}{A f_A + B f_B}$$

Using Equations 55 and 56, a theoretical correlation diagram can be
constructed as described below.

Arbitrary values for rate constants $A$ and $B$ (e.g. $A = 1$ and $B = 1$
or $A = 1$ and $B = 100$) are substituted in Equations 55 and 56. For different
values of $F_B$, corresponding values of $f_A$ and $f_B$ are calculated from
Equation 56 (Note that $f_A + f_B = 1$). These values for $f_A$ and $f_B$ are then
substituted in Equation 55 and the corresponding values for the isotope
effect are determined. It can be shown that the values of $k_H/k_D$ obtained
for different values of $F_B$ are independent of the numbers used for rate
constants $A$ and $B$.

The theoretical and experimental plots are shown in Fig. 50. It is
clear that they are superimposable within experimental error (error bars
for the theoretical plot are not shown in Fig. 50). This proves that the
two sites are acting independently and that there is no bifunctional
catalysis.
Fig. 50 Comparison of the Experimental and Theoretical Correlation Plots
This thesis describes the kinetics and mechanism of a simple organic reaction in solution catalyzed by a solid held in suspension. The reaction studied is glucose mutarotation and the solvent is dimethyl sulfoxide. The catalyst used is Woelm alumina neutral (for thin layer chromatography), a porous solid with a BET surface area of 140 m²/g (X-ray crystallography showed that it is the γ-form). All reactions were carried out at 25.0°C.

As "standard conditions" the kinetics of the reaction was carried out by stirring a 60 ml 0.05 M solution of α-D-glucose containing 1.6 g alumina, with an overhead stirrer such that the rate of the reaction was independent of the rate of stirring (Fig. 16). Under such conditions, the observed rate is not controlled by 'external' diffusion of substrate molecules from the bulk solution to the surface of the alumina particles. A decrease of particle size (so as to decrease the pore length) by grinding did not affect the observed rate indicating that the rate of the reaction is also not controlled by 'internal' diffusion of substrate molecules from the surface of alumina particles to active sites within the pores.

Under such conditions, the first order kinetic plots obtained were curved (Fig. 7) unlike the well known linear first order plots observed for homogeneous catalysis. There was an initial rapid decrease in optical rotation during the first 10 to 15 minutes. Using an equilibrated glucose solution it was shown that this initial decrease in optical rotation is due to adsorption of glucose on the alumina surface which is complete in about 15 minutes (Fig. 15). Analysis of reaction products, to determine whether any other (non first order) reaction is occurring on the surface, showed the presence of 40.3% α-D-glucopyranose, 57.5% β-D-glucopyranose and <1% side
products in a sample removed towards the end of a kinetic run. Hence the curvature in the first order plots could mean either the surface catalyzed mutarotation is not first order in glucose or that alumina is undergoing progressive deactivation during catalysis.

Tests for deactivation of the catalyst were carried out with alumina that had been previously used to catalyze glucose mutarotation. Studies with 'used' alumina showed that the catalyst has undergone deactivation during the previous catalytic process. It was also shown that solvent DMSO or water (that may be formed by dehydration of glucose) did not deactivate the catalyst. Six and half hours of mutarotation decreased catalytic activity about 50% and after 70 hours of catalysis the activity had decreased about 70%. Partially deactivated catalysts gave linear first order plots over two or three-half lives. This shows that glucose mutarotation on alumina surface is a first order process. Further studies showed that deactivation of the catalyst is not due to inhibitory effect of any side products in solution but due to strong adsorption (possibly by ether formation) of glucose on active sites.

These results therefore showed that the initial decrease in optical rotation is due to surface adsorption while the curvature in the rest of the plot is due to progressive deactivation of the catalyst. Hence the rate constant for the catalytic reaction was determined by measuring the slope at \( t = 25 \) mins when the adsorption is complete but before much deactivation has occurred.

Dehydration of alumina neutral at low temperatures caused a decrease in specific activity (activity per unit area). But as the temperature was increased above \( \approx 600^\circ C \) the specific activity began to increase and the \( \alpha \)-alumina formed at \( 1250^\circ C \) (surface area = \( 6.2 \text{ m}^2/\text{g} \); percentage glucose
adsorbed by 1.6 g of α-alumina from 60 ml 0.05 M glucose solution = 3% compared to 14% adsorbed by alumina neutral under the same conditions) showed a specific activity which is 26 times that of standard alumina neutral (Fig. 28). Such high activity for α-alumina (compared with γ-alumina) is virtually unknown. Further, this catalyst did not get deactivated during the catalysis and always produced linear first order plots over three half-lives (Fig. 27).

The observed decrease in activity on mild thermal treatment is unlike other reactions catalyzed by alumina where catalytic activity is observed only on heating the catalyst above 300 to 400°C (giving rise to defect sites). This showed that defect sites (consisting of clusters of vacancies and neighboring oxide ions), even if they possess catalytic activity towards mutarotation, do not possess a monopoly over the catalytic activity as they do with many other observed reactions. The observed higher catalytic activity of the hydrated alumina surface towards glucose mutarotation compared with alumina dehydrated by mild thermal treatment (forming defect sites) should be due to the high sensitivity of the mutarotation reaction towards weak acidic and basic sites on the hydrated surface. It was also shown that pyrolysis of alumina neutral at temperatures greater than the Tammann temperature brings out its basic character which gives rise to the observed high catalytic activity of α-alumina.

Rehydration of the dehydrated catalyst increased its activity without increasing the amount of glucose adsorbed. It was shown that the increase of activity is due to the increase of the number of active sites by adsorbed water possibly acting as a medium to transfer protons and hydroxide groups from the acidic and basic sites on the surface to the adsorption sites.
Adsorption studies using equilibrated solutions of glucose showed, that in addition to the reversible adsorption sites, that there are irreversible adsorption sites on the surface of alumina neutral. The amount of irreversible adsorption sites was determined to be \((0.70 \pm 0.02) \times 10^{-4}\) mole/g.

The adsorption isotherms for alumina neutral and alumina sintered at 1250°C showed initial strong adsorption and saturation of adsorption at high glucose concentrations indicating that there is only monolayer formation on the alumina surface (Fig. 33). The maximum adsorption for alumina neutral corresponds to about 60% surface coverage. This suggests that the surface has fairly specific sites for adsorption of glucose. The Langmuir plot obtained from the data for reversible adsorption on alumina neutral showed two linear regions (Fig. 34). It was shown that such a plot can arise from the presence of two types of reversible adsorption sites on the surface; one showing strong adsorption (equilibrium constant for adsorption \(K_1 = (8.2 \pm 1.2) \times 10^2\) litre mole\(^{-1}\) and there are \((1.0 \pm 0.1) \times 10^{-4}\) mole strong adsorption sites/g) and the other showing weak adsorption (equilibrium constant for adsorption \(K_2 = 44 \pm 3\) litre mole\(^{-1}\) and there are \((1.4 \pm 0.2) \times 10^{-4}\) mole weak adsorption sites/g). The comparison of a plot of initial rate versus concentration of glucose with theoretical adsorption isotherms for weak and strong adsorption sites showed (Fig. 38) that only the weak adsorption sites are catalytically active. Thus the active site density on standard alumina neutral was determined to be \((5.4 \pm 0.6) \times 10^{13}\) sites/cm\(^2\). The turnover number of a catalytic site was determined to be \(2 \times 10^{-3}\) molecules/site/sec. It is one of the highest turnover numbers obtained for a reaction on an alumina surface.

It was also shown that the determination of the active site density by comparison of adsorption isotherms with an initial rate plot is a more
direct method than the standard method using inhibitors. Further, it can
distinguish sites that might contain acidic and basic sites but which are
not active (because of irreversible adsorption or improper orientation of
functional groups), from the active sites and give a more complete picture
of the surface sites than other methods of determining active site density.

The surface catalyzed reaction can be represented by Equation 26.

\[ S + C \xrightleftharpoons[k_1]{k_2} SC \xrightarrow[k_3]{k_4} PC \xrightleftharpoons[k_2]{k_1} P + C \] (26)

It was shown that such a system would show first order kinetics with
observed rate constant

\[ k_{\text{obs}} = \frac{(k_3 + k_4) c_0}{k_2/k_1 + s_o} \] (32)

(where \( c_0 \) = total catalyst concentration and \( s_o \) = total glucose concentration) when \( k_1, k_2 \gg k_3, k_4 \). As predicted from Equation 32, the observed rate
constant was found to be directly proportional to the concentration of
catalyst (Fig. 21), while the plot of \( 1/k_{\text{obs}} \) versus concentration of substrate
was linear with a small positive intercept (Fig. 22). Using the slopes of
the two plots and the equilibrium constant for adsorption on catalytic
sites the catalytic constant for the surface reaction \( (k_3 + k_4) \) was
determined to be \( 5 \times 10^{-3} \text{ sec}^{-1} \). Comparison with the catalytic constant
for homogeneous reaction in pure water \( (= 4 \times 10^{-4} \text{ sec}^{-1}) \) showed that the
alumina surface offers a better medium for mutarotation than water.

Comparison of the catalytic constant for overall reaction \( (= k_{\text{obs}} \text{ per unit }
catalyst concentration) with the catalytic constants for homogeneous
catalysts (in Table III) showed that the catalytic sites on alumina neutral are 10 times more active than strong acids in water and a little more active than the bifunctional catalyst 2-hydroxypyridine in benzene.

The nature of the active sites was investigated using "neutral" compounds (to determine the nature of adsorption on the active sites) and acidic and basic compounds (to determine the nature of functional groups involved in catalysis). Simple monohydroxy compounds (e.g. water and methanol) had no effect on catalytic activity or the adsorption of glucose on alumina neutral even when used in excess. However, polyhydroxy compounds (e.g. methyl α-D-glucoside and i-inositol) were efficient inhibitors of glucose mutarotation. For example, equimolar amounts of glucose and inositol decreased the catalytic activity of alumina neutral by 67% and the amount of glucose adsorbed by 60%. This indicated that inositol had no preference for catalytically active over inactive sites. Hexanal was used as a probe for sites that can interact with the acyclic intermediate that may be formed during mutarotation. Unlike polyhydroxy compounds, the percentage deactivation by hexanal was always greater than the percentage decrease in adsorption (Table XVI). This indicated that the active sites on alumina can adsorb and stabilize the acyclic intermediate (more particularly the transition state leading to it) better than the other adsorption sites.

Using methyl α-D-glucoside and hexanal it was shown that rapid deactivation of the catalyst occurs due to interaction of hydroxyl groups (of the glucoside) with the alumina surface, while further, but slow, deactivation occurs by interaction of the aldehyde functional group (of hexanal) with active sites. Hence it appears that interaction of the alumina surface with functional groups, which is essential for adsorption of glucose and stabilization of transition state, also leads to permanent
deactivation of some of the active sites.

Treatment with carbon dioxide deactivated alumina neutral as well as all aluminas prepared by dehydrating alumina neutral. Alumina sintered at 1250°C was deactivated 85% by carbon dioxide. However, the CO$_2$ treated sample regained the original activity on heating to a mild (≈650°C) temperature. Further, the greater the extent of dehydration of the sample the greater was the percentage deactivation (Table XVIII). Since carbon dioxide may be expected to react with basic anionic sites, this deactivation indicates the presence of catalytically active oxide ions at the active sites.

Pyridine had no inhibitory effect on alumina neutral even when used in excess. This showed that Lewis acid sites on alumina neutral are not catalytically active. Stronger bases such as n-butylamine and tetramethylammonium hydroxide increased the catalytic activity indicating that weak Brönsted acid sites on alumina are catalytically active (and are converted to more active anionic sites by reaction with strong bases). Mutarotation of glucose by alumina is probably the only reaction catalyzed by weak Brönsted acid sites on alumina at moderate temperatures. This shows that glucose mutarotation is a more sensitive probe for Brönsted acid sites on solids than any of the techniques used before. Approximately 90% of the catalytic activity of standard alumina neutral is due to these weak Brönsted acid sites. The observed high catalytic activity of these weak acid sites may be due to (i) strong specific adsorption of glucopyranose molecules at the active sites with the Brönsted acid site on the surface at the proper position to protonate the ring oxygen atom and/or (ii) stabilization of the transition state by the polar alumina surface, as suggested by the experiments with hexanal.

Alumina neutral and other aluminas prepared by dehydrating alumina
neutral showed normal isotope effects which are similar to those observed for homogeneous catalysts in solution. There was no isotope effect on the adsorption-desorption process and hence, the observed isotope effect is due to the catalytic reaction on the surface. Therefore, the observed normal isotope effects indicate that glucose mutarotation on alumina surface is a general acid-base catalyzed reaction, and proceeds by a consecutive mechanism, via the acylic intermediate. The observed isotope effect increased linearly with the increase of percentage activity due to basic sites as determined by CO$_2$ inhibition (Fig. 49). The acidic sites seem to show an isotope effect of $k_H/k_D = 1.2$ and basic sites an isotope effect of $k_H/k_D = 2.1$. The observed isotope effect agreed closely with that calculated assuming that the two sites are acting independently without any bifunctional concerted catalysis (Fig. 50).

In conclusion, the results discussed in this thesis have not only established, for the first time, the rate constants for an alumina catalyzed organic reaction in solution but have also shown that the main features of the surface catalyzed glucose mutarotation (for example linear or non-linear first order plots, amount of glucose adsorbed, effect of water on the catalytic activity, deuterium isotope effects) which is highly sensitive to weak acid and base sites on the surface, may be used to characterize different aluminas.
XXI EXPERIMENTAL
XXI  EXPERIMENTAL

XXI.1 General Methods

Optical rotations were determined on a Perkin-Elmer 241 MC Polarimeter at $\lambda = 365$ nm. The measurements were made at 365 nm, instead of the usual 589 nm, because of the greater sensitivity (sensitivity increased by a factor of 3) at the shorter wavelength. Specific rotations ($[\alpha]_{365}^{24}$) were calculated using the equation,

$$[\alpha]_{365}^{24} = \frac{\text{observed rotation (degrees)}}{\text{path length (dm)} \times \text{concentration (g/ml)}}$$

Gas Liquid Chromatography (glc) was carried out in a Perkin-Elmer 900 Gas Chromatograph equipped with a flame ionization detector using helium as carrier gas. Proton nuclear magnetic resonance spectra were recorded on a Varian XL-100 spectrometer by staff members of the NMR laboratory, the University of British Columbia. Samples were prepared as 10% solutions in DMSO-$d_6$ and had 1% tetramethylsilane (TMS) as an internal standard. Electron Micrographs of alumina samples were obtained on an Elec Corporation Autoscan Scanning Electron Microscope at 20 KV by Ms. Sally Finora, University of British Columbia. pH's of slurries of aluminas in water were determined with a Radiometer pH Meter 26 equipped with glass and calomel electrodes. The melting point of methyl $\alpha$-D-glucoside was determined with a Thomas Hoover Unimelt Capillary Melting Point Apparatus using an open tube capillary and is not corrected. Elemental micro-analysis were performed by Mr. Peter Borda, University of British Columbia. Batches of dehydrated aluminas and $\alpha,\beta$ mixtures of glucose were mixed on a Fischer Minimill.

All glassware used in catalytic studies were first cleaned in a
chromic acid bath. Then the glassware, stirrer, syringes and Sweeny syringe filter were washed in an ultrasonic cleaner. The equipment was then washed in tap water, rinsed with distilled water followed by acetone (reagent grade) and dried in an oven at \(\approx 120^\circ C\). After about 3 hours the equipment was cooled in a stream of dry nitrogen.

XXI.2 Materials

Aluminum oxide neutral for thin layer chromatography (without any binder) manufactured by Woelm Pharma (distributed by ICN Pharmaceuticals), was used as the catalyst throughout this work. Other aluminas used to compare activities and the names of distributors or manufacturers are given in the text. All the aluminas used in this thesis do not contain any binder. Eastman (Reagent ACS) DMSO was dried and distilled as described below (Section XXI.3.1). Anhydrous \(\alpha\)-D-glucose (Reagent ACS, Matheson, Coleman and Bell) and anhydrous \(\beta\)-D-glucose (ICN Pharmaceuticals) were dried \textit{in vacuo} over \(\mathrm{P}_2\mathrm{O}_5\) at 60°C for 24 hours. All anhydrous compounds were stored in a dessicator over \(\mathrm{P}_2\mathrm{O}_5\). Freshly distilled water was used to hydrate alumina samples and for isotope effect studies. Deuterium oxide (Gold label, Aldrich) containing 99.8 atom % deuterium was used. Methyl \(\alpha\)-D-glucoside (BDH Chemicals) was recrystallized from five parts of methanol and dried \textit{in vacuo} at 60°C mp 163-164°C (lit. 127 164-165°C). Benzene, methanol, pyridine, n-butylamine (all reagent grade), naphthalene (recrystallized, Eastman), \(\alpha\)-inositol (Sigma), DL-glyceraldehyde (Aldrich) and, hexamethyldisilazane, trimethylchlorosilane and dichlorodimethylsilane obtained from Alfa Products were used without any further purification. A purified (by means of an alumina column to remove small amount of trimer) sample of hexanal (Aldrich) was a generous gift of Dr. V. Gujral.
XXI.3 Preparations

XXI.3.1 Preparation of Dry, Distilled DMSO

One litre of Eastman Reagent ACS Grade DMSO was dried overnight over 10 g of calcium hydride and was refluxed at reduced pressure for one hour. It was then distilled (at 38-40°C) under reduced pressure and the center fraction (≈800 ml) was collected over 5 g of alumina neutral activity super I (Woelm Pharma). The vacuum was broken by introducing dry nitrogen into the distillation apparatus. The flask containing distilled DMSO was stored under dry nitrogen.

XXI.3.2 Preparation of Alumina from Aluminum Isopropoxide

Two hundred grams of aluminum isopropoxide (98+%, Alfa Products) was distilled (at ≈110°C) under reduced pressure and middle fraction (≈150 g) was collected. The latter was distilled again under reduced pressure and middle fraction (≈100 g) was collected. This pure ion-free aluminum isopropoxide was slowly added with stirring to 700 mls of distilled water. After all the isopropoxide had been added the slurry was stirred for another six hours. The slurry was filtered under suction and the aluminum hydroxide was washed with excess water. It was then dried at 120°C for three days. A portion (20 g) was further dried at 600°C under nitrogen for four hours to give a sample of pure alumina.

XXI.3.3 Preparation of Alumina Neutral Activity V from Alumina Neutral Activity Super I

Fifteen grams of alumina neutral activity super I (on the Brockmann scale) taken in an erlenmeyer flask, was treated with 2.85 ml
(19% by weight of alumina) of distilled water. The flask was stoppered, shaken well and was allowed to stand at room temperature (≈24°C) for 2 hours.

XXI.3.4 Preparation of α,β Mixture

The α,β mixture of equilibrium composition used in adsorption (Section V) and the deactivation (Section VIII.1) studies was prepared by mixing together α-D-glucose and β-D-glucose in the ratio 1.00 : 1.67 (equilibrium constant in DMSO at 24.7°C = 1.67)\(^6\). It was then ground well with a pestle and mortar, mixed for an hour on a minimill and dried in vacuo at 60°C over P\(_2\)O\(_5\).

XXI.3.5 Preparation of a Completely Equilibrated Solution of α,β Mixture in DMSO

For the study of the adsorption isotherms (Section XIV.1) and the isotope effect on the rate of adsorption (Section XIX.3.2), completely equilibrated 1.0 M and 2.0 M solutions of α,β mixture in DMSO were made as described below.

In order to prepare 500 ml of 1.0 M solution, 33.7 g of α-D-glucose and 56.3 g of β-D-glucose were dissolved in about 400 ml of DMSO, the solution was transferred into a 500 ml volumetric flask and the volume was made up with DMSO. Addition of a drop of 0.1 N NaOH to about 20 ml of the solution increased its optical rotation showing that the solution was not completely equilibrated. Hence the solution was transferred to a 1 litre erlenmeyer flask containing 10 g of sintered (1250°C) alumina and 10 g of alumina neutral, and the slurry was stirred with a magnetic stirbar under dry nitrogen at 25.0°C. After a few (2 or 3) days the alumina was allowed
to settle, a sample of the slurry was removed by a Pasteur pipette and filtered through a Swinny syringe filter. The optical rotation of the clear solution was determined using a 1 dm path length cell thermostatted at 25.0°C. The stirring was continued and the optical rotation of another sample was measured after a few more days. This process of stirring and measuring optical rotations of filtered slurries was continued until no further change in the optical rotation was observed (generally in about 7 to 10 days; in some cases complete equilibration was achieved in a shorter time by using 15 to 20 g of sintered alumina). The alumina was allowed to settle and the supernatant was filtered under suction to give a clear solution. Complete equilibration of the solution was again determined by adding a drop of 0.1 N NaOH to about 20 ml of the solution (no change in optical rotation was observed). This completely equilibrated solution was stored under nitrogen at 25.0°C.

The same procedure was adopted for the preparation of the 2.0 M equilibrated solution. Solutions of lower concentrations were made by progressive dilution of more concentrated solutions with DMSO.

XXI.3.6 Preparation of Dehydrated Aluminas

XXI.3.6.1 Dehydration at Room Temperature

Dehydration of alumina neutral at room temperature (24°C) was carried out in a drying pistol over P₂O₅ at 0.01 mm Hg pressure. After drying the sample for 4 days (P₂O₅ was changed once), the vacuum in the system was broken by introducing dry nitrogen into the drying pistol. The alumina sample was then mixed well on a minimill.
XXI.3.6.2 Dehydration at 150°C

Dehydration of alumina neutral at 150°C was carried out in an Ace Instatherm drying apparatus (manufactured by Ace Glass Inc., Vineland, N.J.) at 0.01 mm Hg pressure. After drying the sample for two days the apparatus was allowed to cool to room temperature and the vacuum was broken by introducing dry nitrogen. The sample was then mixed well on a minimill.

XXI.3.6.3 Dehydration at 600°C and Higher Temperatures

Dehydration of alumina neutral at 600°C and higher temperatures was carried out in an Electric Hi-Speed Furnace Type G-05-PT (power 3500 Watts, manufactured by Hevi Duty Electric Co., Milwaukee, Wisconsin). A long quartz tube was placed in the furnace and one end of the tube projected out through a hole in the rear wall of the furnace. Alumina sample was taken in an Alumina Combustion Boat (Grade AD-99, purchased from Coors Porcelain Company, Golden, Colorado) and introduced into the quartz tube. Dry nitrogen (Linde, U.S.P., 99% N₂) was introduced into the furnace through the end of the tube which projected out of the furnace. The furnace was switched on and was heated at maximum rate until the required temperature was reached. After the sample had been dried at the required temperature (e.g. 600°C) for a particular length of time (e.g. 4 hours), the furnace was switched off and was allowed to cool to room temperature overnight under nitrogen. The dried alumina sample was then mixed on a minimill.

XXI.3.7 Preparation of Carbon Dioxide Treated Alumina

About 4 g of an alumina sample was packed into a chromatographic separation tube with a disc to support column packing. By means of a gas inlet tube connected to the top of the separation tube, dry carbon dioxide
(Linde, U.S.P., 99% CO₂) was passed through the alumina sample. The flow of carbon dioxide was monitored by a gas bubbler connected to the bottom of the separation tube by a vinyl tube. After the treatment was complete (24 hours) the sample was mixed on a minimill. The percentage carbon, determined by micro-analysis, on sintered (1250°C) alumina is 0.05% and on CO₂ treated sintered alumina is 0.04%.

XXI.3.8 Preparation of a Standard Solution of Tetramethylammonium Hydroxide Pentahydrate and Determination of the Number of Acid Sites on Alumina Neutral

XXI.3.8.1 Preparation and Standardization of a Solution of Tetramethylammonium Hydroxide Pentahydrate

A known weight (0.055 g) of the hydroxide (Eastman) taken in a 50 ml beaker was dissolved in 10 ml of DMSO with slight warming. The small amount of the undissolved material (possibly carbonate) was removed by filtration, through a Buchner glass funnel with a fritted disc (fine), into an erlenmeyer flask. The clear solution of the hydroxide was titrated against a standard (0.0122 M) solution of hydrochloric acid using phenolphthalein as indicator. The amount of acid needed for neutralization was 16.8 ml. The procedure was repeated with a new sample (same weight) of the hydroxide and it required 16.0 ml of the acid for neutralization.

Hence the average number of moles of the acid used = \( \frac{16.4 \times 0.0122}{1000} \) = 2.0 \times 10^{-4} moles. Therefore \( \frac{0.055}{181} = 3.0 \times 10^{-4} \) mole of the tetramethylammonium hydroxide pentahydrate contains 2.0 \times 10^{-4} mole (i.e. 66%) of the base.

During the above operations dry nitrogen was passed over the beaker,
the funnel and the flask and they were covered with parafilm to minimize exposure of the solutions to the atmosphere.

XXI.3.8.2 Determination of the Number of Acid Sites on Alumina Neutral by Treatment with Excess Tetramethylammonium Hydroxide

XXI.3.8.2.1 Amount of the Hydroxide Required

From the results of the treatment of alumina neutral with tetramethylammonium hydroxide discussed in Section XVIII.2.3, 0.08 g of the alumina is neutralized by $2 \times 10^{-5}$ mole of the base. Hence 1.6 g of alumina neutral would need $4 \times 10^{-4}$ mole of the base for neutralization. Further, 10 ml of 0.0122 M hydrochloric acid solution needs $1.2 \times 10^{-4}$ mole of the base for neutralization. Therefore, if 1.6 g of alumina neutral is treated with $5.2 \times 10^{-4}$ (=$4 \times 10^{-4} + 1.2 \times 10^{-4}$) mole of the hydroxide, 10 ml of the standard acid would be required for neutralization of unreacted base. From the results of standardization experiment above, $5.2 \times 10^{-4}$ mole of the hydroxide is found in $\approx 0.16$ g of tetramethylammonium hydroxide pentahydrate.

XXI.3.8.2.2 Treatment of Alumina Neutral with Excess Tetramethylammonium Hydroxide

A known weight (0.165 g) of the hydroxide was dissolved in 10 ml of DMSO and filtered as described above. The clear solution was added to 1.6 g of alumina neutral taken in an erlenmeyer flask. The flask was then sealed and stirred for 120 mins. The slurry was filtered through a Buchner glass funnel with a fritted disc (fine) and the alumina was washed with 2 ml of DMSO in three portions. The combined filtrate was
titrated against 0.0122 M hydrochloric acid solution using phenolphthalein as indicator and 8.0 ml of the acid was needed for colour change to occur. The procedure was repeated with the same weights of the hydroxide and alumina, and the filtrate required 8.2 ml of the acid for neutralization.

From the results of the standardization experiment described above, 0.165 g of the tetramethylammonium hydroxide pentahydrate contains 6.0 x 10^{-4} mole of the hydroxide. The number of moles of hydroxide left after treating it with 1.6 g of alumina = \frac{8.1 \times 0.0122}{1000} = 0.99 \times 10^{-4} mole. Therefore the number of moles of hydroxide that has reacted with 1.6 g of alumina neutral = 5 \times 10^{-4} mole.

The experiment was repeated using a greater amount of the hydroxide (0.216 g of the hydroxide dissolved in 20 ml of DMSO) but treating it with 1.6 g of alumina neutral for one hour. The unreacted hydroxide needed 24.9 ml of the 0.0122 M acid for neutralization. Again, from the results of the standardization experiment, 0.216 g of the hydroxide contains 7.9 x 10^{-4} mole of hydroxide ions. The number of moles of hydroxide ions left after treating it with 1.6 g of alumina neutral = \frac{24.9 \times 0.0122}{1000} = 3.0 \times 10^{-4} mole. Therefore the number of moles of hydroxide that has reacted with 1.6 g of alumina neutral = 5 \times 10^{-4}, which is the same as the result obtained by the previous experiment. From this result the number of acid sites on alumina was calculated as described in Section XVIII.2.3.
XXI.3.9 Preparation of Diazomethane and Treatment of Alumina Samples with Diazomethane

XXI.3.9.1 Amount of Diazomethane Required

It was observed in Section XII that further dehydration of a sample of alumina neutral, that has been dehydrated at room temperature, causes about 5% loss in weight. Hence a 4 g sample can lose 0.2 g of water (on further dehydration) or contains 0.022 mole of hydroxyl groups on the surface. If all the hydroxyl groups are acidic a minimum of 0.022 mole of diazomethane is required to methylate the surface.

The amount of diazomethane required may also be estimated from the number of acid groups determined by treatment of alumina neutral with tetramethylammonium hydroxide in Section XXI.3.8.2.2 above. Since 1.6 g of alumina reacted with $5 \times 10^{-4}$ mole of the hydroxide, 4 g of the alumina sample would need $1.2 \times 10^{-3}$ mole of the base for neutralization. The higher number of moles of surface hydroxyl groups determined by dehydration studies should be due to the presence of basic hydroxyl groups and strongly adsorbed water molecules on the surface. These strongly adsorbed water molecules may also react with diazomethane if the reaction is catalyzed by Lewis acid sites on the surface. Therefore, in order to allow for decomposition of diazomethane, it was decided to use about 3 times (i.e. 0.07 mole) the minimum estimated by dehydration studies.

XXI.3.9.2 Preparation of Diazomethane

The procedure outlined by de Boer and Backer\textsuperscript{131} was employed to prepare an alcohol-free ethereal solution of diazomethane. The yield of diazomethane prepared by this method is 70-78\%\textsuperscript{131(a)}. Therefore to
prepare 0.07 mole of diazomethane 0.1 mole (21.5 g) of Diazald* was used. About 200 ml of ethereal** diazomethane solution was prepared by the method of de Boer and Backer and was dried over anhydrous sodium sulphate (Reagent ACS; from Mallinckrodt) for one hour. The ethereal solution was then decanted into two flasks. About 150 ml was added into one and was used to treat alumina samples with diazomethane and the rest of the diazomethane was allowed to decompose at room temperature and was used to prepare the control samples.

XXI.3.9.3 Treatment of Alumina Samples with Diazomethane

Slurries, of alumina neutral evacuated at room temperature and alumina sintered at 1250°C, in diethyl ether (4 g of each in 30 ml of ether) were allowed to cool in an ice-salt mixture. They were stirred with Teflon-coated magnetic stirbars and the diazomethane solution (prepared in Section XXI.3.9.2 above) was added slowly with a Pasteur pipette. The slurry of sintered alumina turned yellow when a few mLs of the diazomethane solution was added while the slurry of alumina neutral needed 8-10 mLs to turn yellow. Decomposition of diazomethane (i.e. evolution of bubbles) was observed on the surface of alumina neutral. More diazomethane was added to both the slurries and they were left overnight at room temperature in a fume hood. The diazomethane solution was stored in a refrigerator. After about 12 hours the slurry containing sintered alumina still had a light yellow colour while that of alumina neutral was colourless. More diazomethane was added to both the slurries and as the colour in the

*Diazald (N-Methyl-N-nitroso-p-tolunesulphonamide) is available from Aldrich Chemical Company.

**Anhydrous Ethyl Ether (Reagent ACS) from Mallinckrodt was used.
slurry of alumina neutral disappeared more diazomethane was added. The procedure was continued for 6 hours and both the slurries were again allowed to stand overnight. After about 12 hours the addition of diazomethane was continued until all the 150 mls had been added. The slurry of sintered alumina was decanted and the supernatant (total volume ≈50 mls) was collected. The alumina was washed with three 50 ml portions of ether. Ether remaining with the diazomethane treated sintered alumina was evaporated in a water bath and the sample was dried in vacuo at room temperature (≈23°C) for 4 hours. The same procedure was adopted to obtain diazomethane treated alumina neutral from its slurry.

XXI.3.9.4 Preparation of the Control Samples

The supernatants collected from the two slurries (in Section XXI.3.9.3 above) were combined with the 50 ml of the diazomethane solution that was left to decompose at room temperature. It was used to treat slurries of alumina neutral evacuated at room temperature and sintered alumina in ether. From the two slurries the control samples were obtained using the procedure described above.

XXI.3.10 Preparation of Dichlorodimethylsilane $^{132}((CH_3)_2SiCl_2)$ Treated Alumina

XXI.3.10.1 Amount of Dichlorodimethylsilane needed

The minimum amount of $(CH_3)_2SiCl_2$ needed can be determined by two methods; (i) from the number of oxygens on the surface determined by model studies and (ii) from the number of hydroxyl groups determined by dehydration studies, as discussed below.
From the literature, the average number of oxygens (present as oxide ions and hydroxyl groups) on the surface of alumina is $10^{15} \text{ cm}^{-2} = 10^{19} \text{ m}^{-2}$. Therefore the number of moles of oxygen on the surface of 3 g of alumina neutral dehydrated at room temperature is

$$\frac{144 \times 10^{19} \times 3}{6 \times 10^{23}} = 7 \times 10^{-3} \text{ mole}.$$ 

Hence $7 \times 10^{-3}$ mole of $(\text{CH}_3)_2\text{SiCl}_2$ should completely silylate the alumina surface.

From the dehydration studies described in Section XII, about 5% of the weight of alumina neutral dehydrated at room temperature can be lost as water. Therefore the number of moles of hydroxyl groups on the surface of a 3 g sample is

$$\frac{5 \times 3 \times 2}{100 \times 18} = 0.017 \text{ mole}.$$ 

The higher value determined by dehydration studies may be due to the presence of strongly adsorbed water molecules on the surface of the sample.

Therefore the minimum amount of $(\text{CH}_3)_2\text{SiCl}_2$ necessary to completely silylate 3 g of alumina is about 0.02 mole. However, in order to allow for side reactions with any moisture in the solvent and the surface of the glass vessel, 0.08 mole of $(\text{CH}_3)_2\text{SiCl}_2$ was used as described below.

XXI.3.10.2 Treatment of Alumina with $(\text{CH}_3)_2\text{SiCl}_2$

Into 100 ml of Fischer ACS Toluene (dried and distilled over lithium aluminum hydride) was added 10 g (0.08 mole) of $(\text{CH}_3)_2\text{SiCl}_2$ and 3 g of alumina neutral that had been evacuated at room temperature ($24^\circ\text{C}$). The slurry was stirred for 20 hours under nitrogen and was filtered using a Buchner glass funnel with a fritted disc (fine). Dry nitrogen was passed over the funnel to minimize exposure of the alumina surface to water vapour in the atmosphere. The alumina in the funnel was washed three times with 10 ml portions of methanol. Another 10 ml of methanol was added and the alumina in the funnel was broken up with a spatula. The methanol also
was filtered through and a fifth 10 ml portion of methanol was added. The filtration was continued until the sample was dry. The alumina was transferred into a drying pistol and was dried in vacuo at room temperature (24°C).

XXI.3.11 Preparation of Deuterated 800° Alumina and Control Sample of Alumina

XXI.3.11.1 Amount of Replaceable Hydrogen on Surface of Alumina

It was observed in Section XII that about 8% of the weight of alumina neutral can be lost on dehydration. Hence the amount of replaceable hydrogen on 4 g of alumina neutral is \( \frac{4 \times 8 \times 2}{100 \times 18} = 0.036 \) mole. Five ml of \( \text{D}_2\text{O} \) contains \( \frac{5 \times 1.1 \times 2}{20} = 0.56 \) mole of deuterons. Hence the deuterons in 5 ml of \( \text{D}_2\text{O} \) should replace most of the protons on the surface of a 4 g sample of the alumina. To replace almost all the protons on the surface with deuterons it was decided to treat the 4 g sample with three 5 ml portions of \( \text{D}_2\text{O} \).

XXI.3.11.2 Preparation of Deuterated 800° Alumina

Four g of alumina neutral was mixed well with 5 ml of \( \text{D}_2\text{O} \) in a Buchner glass funnel with a fritted disc (fine). The funnel was covered with a piece of rubber sheet and was allowed to stand for 5 mins. The \( \text{D}_2\text{O} \) was removed by suction and another 5 ml portion of \( \text{D}_2\text{O} \) was added to the funnel and mixed well with the alumina. After 5 mins it was filtered and the alumina was treated with the third 5 ml portion of \( \text{D}_2\text{O} \). The procedure was repeated and most of the \( \text{D}_2\text{O} \) was removed by applying suction for one hour. The alumina sample was broken up by a spatula and dried at 800°C.
under nitrogen for 4 hours. The dried sample had lumpy particles (formed by sticking together of alumina particles) and was mixed well on a minimill which broke down the large lumpy particles. The smaller lumpy particles remaining tended to block a Pasteur pipette when samples were removed during early stages of a kinetic run. However, they broke down quickly (within about 30 mins) as the slurry was rapidly stirred and hence, the first optical rotation measurement during kinetic runs involving such aluminas was taken at t ≈ 30 mins.

XXI.3.11.3 Preparation of the Control Sample of 800° Alumina

The same procedure was followed except that the same quantity of distilled water was used instead of D₂O.

XXI.3.12 Preparation of O-Deuterated α-D-Glucose

An industrial method for preparation of anhydrous α-D-glucose was adopted to prepare O-deuterated α-D-glucose as described below.

The preparation was carried out in a 50 ml two neck round bottom flask with one neck connected to a dropping funnel and the other neck connected to a receiving flask through a distillation head and a condenser. Twelve grams of α-D-glucose (contains 0.33 mole of exchangeable protons) was taken in the flask and 50 ml of D₂O (contains 5.5 moles of exchangeable deuterons) was added to the dropping funnel. Fifteen ml of the D₂O was added to the flask and the glucose was dissolved by stirring with a magnetic stirbar with slight warming on a water bath. Most of the solvent was distilled off at about 40°C under reduced pressure leaving a thick syrup in the flask. The water bath was removed and another 15 ml of D₂O was added from the dropping funnel and the syrup was dissolved in D₂O. The
D₂O was distilled off as before and the procedure was repeated with a third portion (10 ml) of D₂O. After distilling off the third portion of D₂O, more D₂O was added from the dropping funnel until the total volume was about 20 ml (therefore the solution was 60% in glucose). The dropping funnel and the distillation head were removed and dry nitrogen was blown over the solution with the temperature of the water bath raised to 65°C. The solution was seeded with a few crystals of α-D-glucose (from National Bureau of Standards) and growth of the crystals began within the first day. The crystallization was allowed to proceed until almost all the solvent had evaporated (in about 3 days). The solid mass was broken up in a glove bag under nitrogen, ground with a pestle and mortar and dried at 60°C, in vacuo, over P₂O₅. This sample of O-deuterated α-D-glucose was used to seed the 60% glucose solutions obtained during the preparation of additional batches of O-deuterated α-D-glucose.

The specific rotation of the O-deuterated glucose sample [α]$_{24}^{25}$ = 328°, and may be compared with the specific rotation of α-D-glucose from National Bureau of Standards = 333°. This shows that O-deuterated glucose prepared by the above procedure is almost completely the α-form. The proton nmr spectra of an O-deuterated glucose sample (i.e. α-glucose-OD), anhydrous α-D-glucose (i.e. α-glucose-OH), and anhydrous β-D-glucose (i.e. β-glucose-OH) in DMSO-d$_6$ are shown in Figs. 51, 52 and 53 respectively. The absence of the OH resonance at δ = 4.86 in Fig. 51 clearly shows that the five hydroxyl protons present in normal glucose (Figs. 52 and 53) have been replaced by deuterons.

O-deuterated α-D-glucose could also be prepared by the above procedure but without seeding with α-D-glucose. However, crystallization started only after blowing nitrogen over the sample for about 2 days and
Fig. 51  Proton NMR Spectrum of O-deuterated α-D-Glucose in DMSO-d$_6$
Fig. 52  Proton NMR Spectrum of α-D-Glucose in DMSO-d$_6$
Fig. 53  Proton NMR Spectrum of β-D-Glucose in DMSO-d$_6$
and hence, complete crystallization by this process took about 2 to 3 days longer. Further, the O-deuterated glucose prepared without seeding contained about 10% of $\beta$-D-glucose.

XXI.4 Kinetic and Adsorption Methods

XXI.4.1 General Procedure for Following Kinetics of the Surface Catalyzed Reaction

Sixty ml of a $\alpha$-D-glucose solution was pipetted into a 200 ml three neck round bottom flask placed in a water bath maintained at 25.0°C. One of the side necks was connected to a nitrogen inlet tube and the other side neck, which was generally left stoppered, was used to add the catalyst and to remove samples of the slurry. A dispersion stirrer connected to an overhead Fischer Dyna-Mix passed through a Teflon stopper fixed to the central neck.

After the solution in the flask had reached thermal equilibrium with the water bath (in 15-20 mins) the flask was charged with a preweighed sample of the catalyst and immediately the stirrer and timer were started. Approximately 2 ml samples of the slurry were removed periodically with a Pasteur pipette and filtered through a Swinny syringe filter. Whatman No. 50 filter paper was used in all filtrations and produced a clear filtrate (first few drops of the filtrate which tended to be slightly cloudy were discarded) free of any particles. Whatman No. 42 filter paper could also be used to filter slurries of sintered alumina samples.

Optical rotations (at 365 nm) of the filtered slurries were measured using a 0.1 dm path length cell at room temperature (23-24°C). The infinity
values of optical rotations ($\alpha_\infty$) for catalysis by highly active aluminas (e.g. alumina neutral, alumina dehydrated at 150°C, sintered aluminas) or water treated catalysts were determined after stirring the slurry overnight. For less active catalysts (e.g. 800° alumina) the infinity values ($\alpha_\infty$) were obtained by adding a homogeneous catalyst (e.g. few drops of n-butylamine or 0.1 N sodium hydroxide) or by stirring the catalyst with an equilibrated glucose solution of the same concentration. Infinity value of the optical rotation for homogeneous reaction was determined by adding a few drops of n-butylamine or 0.1 N sodium hydroxide to about 20 ml of the original glucose solution.

XXI.4.2 Rate of Adsorption of Glucose on Alumina (Section V.3)

A 60 ml 0.05 M solution of $\alpha,\beta$ mixture in DMSO taken in the 200 ml flask described above was allowed to come to thermal equilibrium with a water bath at 25.0°C. It was then charged with a preweighed sample of alumina and immediately the stirrer and timer were started. Samples of the slurry were removed every 2 mins, filtered as described before and the optical rotations were determined with a 0.1 dm path length cell at room temperature. In order to filter many samples in a short time several syringes and Swinny syringe filters were used.

XXI.4.3 Determination of Adsorption Isotherms

For the study of adsorption isotherms, the completely equilibrated solutions of $\alpha,\beta$ mixtures in DMSO, prepared according to the procedure given in Section XXI.3.5, were used. The amount of glucose adsorbed by alumina neutral at a particular equilibrium concentration of glucose was determined as described below.
Sixty ml of an equilibrated glucose solution was stirred with a 1.6 g sample of alumina neutral under nitrogen at 25.0°C. Optical rotations of filtered slurries (see Section XXI.4.1) were measured periodically, using a 1 dm path length cell thermostatted at 25.0°C, until there was no further decrease. Results obtained with a 0.05 M solution are given in Table XIX.

**TABLE XIX**

DATA FOR ADSORPTION OF GLUCOSE BY 1.6 G OF ALUMINA NEUTRAL FROM A 60 ML 0.05 M SOLUTION OF α,β MIXTURE IN DMSO AT 25.0°C

<table>
<thead>
<tr>
<th>Time (mins)</th>
<th>Optical Rotation (α°) at 25.0°C</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>1.454</td>
</tr>
<tr>
<td>60</td>
<td>1.285</td>
</tr>
<tr>
<td>120</td>
<td>1.278</td>
</tr>
<tr>
<td>200</td>
<td>1.274</td>
</tr>
<tr>
<td>225</td>
<td>1.274</td>
</tr>
<tr>
<td>300</td>
<td>1.273</td>
</tr>
</tbody>
</table>

A very slow decrease in optical rotation was observed as slurries (specially with solutions of low glucose concentrations) were continuously stirred and could be due to side reactions occurring on the surface. The essentially constant optical rotation after 3 to 4 hours of stirring was taken as the equilibrium value. Further, the effect of filtration on the optical rotation of a solution was found to be negligible. The procedure was repeated with solutions of different glucose concentrations. When solutions
of high glucose concentrations (e.g. 1.5 M, 2.0 M) were used it was observed that air bubbles accumulated in filtered slurries. Hence, in order to remove the bubbles, the filtered slurries were allowed to stand for several hours before optical rotations were determined. The initial and final (equilibrium) optical rotations of the solutions used are given in Columns 2 and 3 of Table X (Section XIV.1). In order to determine the reproducibility of the results, the experiment was repeated at two concentrations and the results are included in Table X.

As shown in Fig. 54 a plot of optical rotation versus the concentration of solutions of α, β mixture is linear, and passes through the origin, over the range of concentrations used in this experiment. Hence the optical rotation of a solution is directly proportional to its concentration and the change in optical rotation of a solution is directly proportional to the change in its concentration. The slope of the plot = $3.47 \times 10^{-2}$ mole litre$^{-1}$ degree$^{-1}$, is the change in concentration of a solution when its optical rotation changes by one degree. If the difference between the initial and final optical rotations $\alpha_{\text{initial}} - \alpha_{\text{equil.}}$, equals $\alpha$ degrees then the change in concentration of the solution = $3.47 \times 10^{-2} \times \alpha$ mole litre$^{-1}$. Hence the number of moles of glucose adsorbed by a gram of catalyst dispersed in 60 ml solution is $\frac{3.47 \times 10^{-2} \times \alpha \times 0.060}{1.6}$ and is given in Column 6 of Table X. Equilibrium (final) concentrations of the glucose solutions were determined using the equilibrium optical rotations and the slope of the plot in Fig. 54.

The same procedure was used to determine the adsorption isotherm for alumina sintered at 1250°C for 6 hours. However, as discussed in Section XIV.2, 3.2 g samples of the catalyst were used in order to increase the sensitivity of measurements.
Fig. 54  Relation of the Optical Rotation to the Concentration of $\alpha, \beta$ Mixture in DMSO
XXI.4.4 Tests for Reversibility of Adsorption

Tests for reversibility of adsorption on alumina neutral were carried out by diluting a glucose solution in equilibrium with glucose adsorbed on a sample of alumina, in order to release some of the adsorbed glucose.

(a) A 60 ml 0.05 M (concentration used in most kinetic studies) solution of $\alpha,\beta$ mixture in DMSO was stirred with 3.2 g of alumina neutral until the optical rotation was constant (in about 3 hours). Samples were not removed from this slurry. However, the equilibrium optical rotation (1.116°) and concentration (0.0385 M) were determined by another experiment. Sixty ml of DMSO was added to the system and the stirring was continued until there was no further change in optical rotation ($\alpha = 0.569°$; concentration = 0.0196 M). If there was no glucose released on dilution then the optical rotation should be $\frac{1.116°}{2} = 0.558°$. The increase in optical rotation by 0.01° indicates that $4.5 \times 10^{-5}$ mole of glucose has been released from the surface to the solution.

(b) The same experiment was performed with a 0.005 M solution. The results indicate that even at this very low concentration there is reversible adsorption because $1 \times 10^{-5}$ mole of glucose was released on dilution.

XXI.4.5 Deuterium Isotope Effect on the Rate of Adsorption

Solutions of methyl $\alpha$-D-glucoside in DMSO containing H$_2$O or D$_2$O were prepared as described in Section XIX.1.3.1. Sixty ml of the solution was taken in a 200 ml flask placed in a water bath maintained at 25.0°C (see Section XXI.4.1). After thermal equilibrium has been reached, the flask was charged with 1.6 g of 800° alumina and the rate of adsorption
was followed as described in Section XXI.4.2. Several syringes and Swinny syringe filters were used to facilitate filtration of several samples in a short time. Further, in order to determine optical rotations with high sensitivity a 1 dm path length cell thermostatted at 25.0°C was used.

XXI.5 Analytical Methods

XXI.5.1 Product Analysis

Analysis of products of glucose mutarotation by alumina was carried out by preparing trimethylsilyl ethers of glucose in the filtered slurry according to the procedure of Sweeley et al. To 1 ml of the filtered slurry obtained near the end of a kinetic run was added 0.2 ml of hexamethyldisilazane, 0.1 ml of trimethylchlorosilane and 1 ml of anhydrous pyridine (kept over KOH pellets). The mixture was shaken for 30 mins and was allowed to stand at room temperature for 6 hrs. It was then concentrated in vacuo, extracted into hexane and analyzed by gas chromatography using a stainless steel column 1/8" x 6' packed with 8% SE 30 on AW Chromosorb W 60/80 at 215°C. Two main peaks and a small peak were observed. The three peaks A, B and C had relative retention times 0.82, 1.00 and 1.35 which accounted for 2.2%, 40.3% and 57.5% of the area. In order to identify these peaks, trimethylsilyl ethers of α-D-glucopyranose and β-D-glucopyranose were prepared and analyzed by gas chromatography as described below.

Ten mg of anhydrous α-D-glucopyranose was dissolved in 1 ml of DMSO and was silylated using the procedure described above. It was analyzed by gas chromatography under the same conditions and produced a large peak with
retention time equal to that of peak B above (relative retention time 1.00) and area ≈ 96% of the total area, and a small peak (area 4%) with relative retention time 1.35. This shows that the peak B observed during product analysis is due to α-D-glucopyranose. In addition to these two peaks, a very small peak with area ≈ 0.5% of the total area and relative retention time 0.81 was also observed (see below).

The same procedure was adopted to prepare the trimethylsilyl ether from anhydrous β-D-glucopyranose and was analyzed by gas chromatography under the same conditions. A large peak (area 97% of the total area) with a retention time equal to peak C above (relative retention time 1.35) and a small peak (area 3%) with relative retention time 1.00 were observed. This shows that the peak C observed during product analysis is due to β-D-glucopyranose. In addition to these two peaks a small peak with area ≈ 0.5% of the total area and relative retention time 0.82 was also observed.

As mentioned in Section IV the small peak with relative retention time 0.82 observed during analysis of the products and also during analysis of α- and β-D-glucopyranoses may be due to a small amount of trimethylsilyl ethers of α- and β-D-glucofuranoses present in DMSO solution. In addition to the three peaks A, B and C, a smaller peak with area less than 1% of the total area and relative retention time 1.10 was observed during the analysis of products, but it was not identified.

XXI.5.2 Analysis of Alumina Samples for Trace Metal Impurities

Alumina samples were tested for trace metal impurities by Can Test Ltd., Vancouver, British Columbia. According to the test report "samples were digested using a combination of acids (HF, HNO$_3$, HCl, HClO$_4$). A trace metal scan was performed using an Inductively Coupled Argon Plasma
Spectrograph". The impurities detected in different alumina samples obtained from Woelm Pharma and the Puratronic alumina sample are given in Table XX. The Puratronic sample was also analyzed in order to compare the results obtained by Can Test Ltd., with Certificate of Analysis of the Puratronic Sample provided by Johnson Matthey Chemicals Ltd., England, where only Na (2 ppm), Ca, Mg, Si and Ag (each <1 ppm) had been detected. The reasons for the discrepancies in the two analytical test reports are not clear.

XXI.5.3 Crystalline Structure of Aluminas

The crystalline structures of alumina neutral and alumina sintered at 1250°C were determined from X-ray powder photographs obtained using a Phillips powder camera of 57 mm radius and nickel filtered CuKα radiation. The samples were ground with pestle and mortar and were sealed in 0.5 mm OD Lindermann glass capillaries. Each sample was exposed to nickel filtered CuKα radiation for 18 hours. The density of diffraction lines on powder photographs was measured on a densitometer. The relative intensities of the peaks on the densitometer plot were determined by cutting them out and weighing. The diffraction lines on the photograph were measured with a travelling microscope and were converted to θ (the angle of reflection) values. The d spacings were then obtained from the θ values using the relation

\[ d = \frac{\lambda}{2 \sin \theta} \]

where \( \lambda \) is the wavelength of CuKα radiation (1.5418 Å). The results for alumina neutral are given in Table XXI together with data from the literature on relative intensities of diffraction lines with same d spacings and the forms of alumina giving rise to those lines.
<table>
<thead>
<tr>
<th>Metal</th>
<th>Neutral* Batch 1 Sintered at 1250°C, 6 hrs</th>
<th>Neutral* Batch 2 Sintered at 1250°C, 6 hrs</th>
<th>Basic* Sintered at 1250°C, 6 hrs</th>
<th>Acidic*</th>
<th>Puratronic</th>
</tr>
</thead>
<tbody>
<tr>
<td>Barium</td>
<td>1.5</td>
<td>1.4</td>
<td>3.4</td>
<td>4.4</td>
<td>1.2</td>
</tr>
<tr>
<td>Calcium</td>
<td>1050</td>
<td>1080</td>
<td>1180</td>
<td>95</td>
<td>138</td>
</tr>
<tr>
<td>Copper</td>
<td>&lt;1.5</td>
<td>26.4</td>
<td>18.1</td>
<td>&lt;1.5</td>
<td>5.6</td>
</tr>
<tr>
<td>Iron</td>
<td>262</td>
<td>102</td>
<td>189</td>
<td>104</td>
<td>199</td>
</tr>
<tr>
<td>Magnesium</td>
<td>216</td>
<td>214</td>
<td>142</td>
<td>137</td>
<td>2.8</td>
</tr>
<tr>
<td>Manganese</td>
<td>1.8</td>
<td>1.1</td>
<td>1.6</td>
<td>1.2</td>
<td>2.0</td>
</tr>
<tr>
<td>Molybdenum</td>
<td>&lt;4</td>
<td>&lt;4</td>
<td>&lt;4</td>
<td>8.6</td>
<td>&lt;4</td>
</tr>
<tr>
<td>Sodium</td>
<td>229</td>
<td>254</td>
<td>276</td>
<td>237</td>
<td>1190</td>
</tr>
<tr>
<td>Strontium</td>
<td>5.6</td>
<td>8.5</td>
<td>11.3</td>
<td>16.9</td>
<td>&lt;0.1</td>
</tr>
<tr>
<td>Tin</td>
<td>14.1</td>
<td>5.8</td>
<td>16.0</td>
<td>4.7</td>
<td>&lt;3</td>
</tr>
<tr>
<td>Titanium</td>
<td>30.5</td>
<td>140</td>
<td>36.2</td>
<td>30.5</td>
<td>29.0</td>
</tr>
<tr>
<td>Vanadium</td>
<td>2.3</td>
<td>3.5</td>
<td>1.5</td>
<td>2.2</td>
<td>1.7</td>
</tr>
<tr>
<td>Zinc</td>
<td>11.1</td>
<td>11.5</td>
<td>&lt;1.5</td>
<td>18.6</td>
<td>17.8</td>
</tr>
</tbody>
</table>

**Expressed in parts per million by weight

* Aluminas for thin layer chromatography manufactured by Woelm Pharma
<table>
<thead>
<tr>
<th>d Spacing Å</th>
<th>Observed</th>
<th>From Literature</th>
<th>Form of Alumina</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Relative Intensity</td>
<td>Relative Intensity</td>
<td></td>
</tr>
<tr>
<td>4.54</td>
<td>broad 11</td>
<td>12</td>
<td>γ</td>
</tr>
<tr>
<td>2.42</td>
<td>38</td>
<td>60</td>
<td>γ</td>
</tr>
<tr>
<td>2.29</td>
<td>8</td>
<td>33</td>
<td>γ</td>
</tr>
<tr>
<td>2.12</td>
<td>50</td>
<td>30</td>
<td>κ, strong</td>
</tr>
<tr>
<td>1.99</td>
<td>doublet 60</td>
<td>65</td>
<td>γ</td>
</tr>
<tr>
<td>1.89</td>
<td>20</td>
<td>10</td>
<td>κ</td>
</tr>
<tr>
<td>1.395</td>
<td>doublet 100</td>
<td>100</td>
<td>γ</td>
</tr>
</tbody>
</table>
A doublet of high intensity at 1.40 Å, a doublet at 1.98 Å and a broad band at 4.6 Å are characteristic of \( \gamma \)-alumina\(^{17} \). However, a few of the lines observed (for example at 1.89 Å, 2.12 Å) could have resulted from the presence of a small amount of other forms of alumina (e.g. \( \kappa \)- and/or \( \chi \)-alumina). The absence of \( \eta \)-form is indicated by the absence of a sharp band at 4.6 Å. These results indicate that alumina neutral consist mainly of the \( \gamma \)-form with some \( \kappa \)- (and \( \chi \)-) aluminas. The presence of forms of alumina belonging to \( \beta \)-series (see Section I.1.2) may be due to the presence of a relatively large amount of calcium in alumina neutral (Table XX).

Results of X-ray diffraction studies of alumina neutral sintered at 1250°C for 6 hours are given in Table XXII. All the observed lines were sharp and the results agree well with the literature\(^{17} \) values for \( \alpha \)-alumina. This shows that \( \gamma \)-alumina has been converted completely to the \( \alpha \)-form on sintering at 1250°C.
<table>
<thead>
<tr>
<th>d Spacing Å</th>
<th>Relative Intensity</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Observed*</td>
</tr>
<tr>
<td>3.479</td>
<td>69</td>
</tr>
<tr>
<td>2.556</td>
<td>91</td>
</tr>
<tr>
<td>2.382</td>
<td>42</td>
</tr>
<tr>
<td>2.170</td>
<td>1.5</td>
</tr>
<tr>
<td>2.087</td>
<td>100</td>
</tr>
<tr>
<td>1.964</td>
<td>2.0</td>
</tr>
<tr>
<td>1.743</td>
<td>44</td>
</tr>
<tr>
<td>1.604</td>
<td>89</td>
</tr>
<tr>
<td>1.548</td>
<td>3.6</td>
</tr>
<tr>
<td>1.515</td>
<td>15</td>
</tr>
<tr>
<td></td>
<td></td>
</tr>
<tr>
<td>1.408</td>
<td>33</td>
</tr>
<tr>
<td>1.374</td>
<td>54</td>
</tr>
<tr>
<td>1.340</td>
<td>2.1</td>
</tr>
<tr>
<td>1.277</td>
<td>2.8</td>
</tr>
<tr>
<td>1.242</td>
<td>27</td>
</tr>
<tr>
<td>1.193</td>
<td>10</td>
</tr>
<tr>
<td>1.151</td>
<td>7</td>
</tr>
<tr>
<td>1.128</td>
<td>11</td>
</tr>
</tbody>
</table>

* All lines were sharp
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   (b) P. Fink, Z. Chem., _7_, 324 (1967).
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73. F. Urech, Ber., 15, 2130 (1882).


   (b) R.W. Maatman, **J. Catal.**, 19, 64 (1970).


APPENDIX A

CHARACTERIZATION OF ALUMINAS

A.1 Determination of BET Surface Areas

A.1.1 Theory

Surface areas of the aluminas were determined using the two parameter BET equation \(^{123,136,137}\),

\[
\frac{1}{w(\frac{1}{x} - 1)} = \frac{1}{w_m c} + \frac{c - 1}{w_m c} x
\]

where \(w_m\) = weight of vapour required to cover the surface by a complete monolayer, \(w\) = weight of vapour adsorbed when adsorbate relative pressure \(x = p/p_o\) (where \(p\) = adsorbate partial pressure, and \(p_o\) = adsorbate saturated vapour pressure), and parameter \(c\) is a constant for a great majority of vapour adsorption isotherms from \(x = 0.05\) to about \(0.35^{123,137}\).

Therefore from a plot of \(\frac{1}{w(\frac{1}{x} - 1)}\) versus \(x\) in this range of relative pressures the weight of adsorbate necessary to form a complete monolayer \(w_m\) is given by \(\frac{1}{\text{slope} + \text{intercept}}\), and the surface area of the sample is given by \(\frac{w_m x (6.02 \times 10^{23}) \times A_{cs}}{M}\), where \(M\) is the molecular weight of the adsorbate and \(A_{cs}\) is the 'cross-sectional' area of the adsorbate molecule (for nitrogen \(A_{cs} = 16.2 \text{Å}^2\) per molecule \(^{113}\)).

A.1.2 Experimental Procedure

To determine the quantity of vapour adsorbed \((w)\) on a sample of alumina at different relative pressures \((x)\), a Quanta-Sorb instrument
(manufactured by Quanta-Chrome Corp., N.Y.) was used. Nitrogen gas was chosen as adsorbate. Using the Quanta-Sorb, nitrogen gas was adsorbed on a preweighed sample of alumina, kept at liquid nitrogen temperature, from a flowing mixture of nitrogen ($x = 0.15$) and an inert non-adsorbable carrier gas (helium)$^{137,138}$. The amount of nitrogen adsorbed was monitored by measuring the change in thermal conductivity of the gas mixture. A change in the thermal conductivity produced a peak in the recorder, and it was calibrated with a known volume of nitrogen. By comparing the areas of the adsorption and calibration signals the weight of nitrogen adsorbed ($w$) at $x = 0.15$ was determined. The procedure was repeated at higher relative pressures of upto $x = 0.35$.

From the plots of $\frac{1}{w(\frac{1}{x} - 1)}$ versus $x$ for different samples of alumina their BET surface areas were determined as described above. They are tabulated in Table IV in Section II.1.

A.2 Determination of Pore Size Distributions

A.2.1 Theory

It has been observed that adsorption of vapours on a solid most often gives rise to multilayer adsorption. However, if the solid is porous, in addition to multilayer adsorption, capillary condensation also takes place and adsorption—desorption isotherms often exhibit hysteresis$^{113-116}$. That is, in a certain region of the isotherm, which is determined by the range of pore radii present, the adsorption isotherm does not follow the desorption isotherm. In addition, from the shape of hysteresis loop the shape of the pores can be determined using de Boer's classification,$^{139,140}$ and from the desorption isotherm the pore size distribution can be determined, using
Zsigmondy's capillary condensation theory\textsuperscript{113,141} and Kelvin equation\textsuperscript{113}, as outlined below.

According to the capillary condensation theory, adsorbed vapour is condensed to the ordinary liquid condition in the pores of the adsorbent. The Kelvin equation for adsorption of nitrogen\textsuperscript{138},

\[
r_k = \frac{-4.146}{\log p/p_0} \text{ Å}
\]
gives the maximum radius of the pore $r_k$ where capillary condensation of nitrogen occurs at a relative pressure of nitrogen $= p/p_0$. The Kelvin radius $r_k$ of the pore is not the actual pore radius because some adsorption has taken place on the wall of the pore prior to the occurrence of condensation in the pore, or during desorption an adsorbed layer remains on the wall after evaporation has occurred. Hence the actual pore radius $r_p$ is given by

\[
r_p = r_k + t
\]

where $t$ is the thickness of the adsorbed layer at the relative pressure $p/p_0$. Values for $t$ at different relative pressures of nitrogen have been determined from studies involving non-porous solids\textsuperscript{142}.

If the pores are assumed to be cylindrical and if the relative pressure is changed from $p_2/p_0$ to $p_1/p_0$ then pores between radii $r_2$ and $r_1$ will empty ($p_2 > p_1$, $r_2 > r_1$). When $p_2$ is lowered to $p_1$ the thickness of the adsorbed film on previously emptied pores changes from $t_2$ to $t_1$. As a result, nitrogen gas gets desorbed from the surface and from the volume of gas desorbed ($V_{\text{gas}}$ cm$^3$), corresponding volume of liquid ($V_{\text{liq}}$ cm$^3$) is
determined. The relation between the volume of pores \( V_p \) having radii between \( r_2 \) and \( r_1 \), and \( V_{\text{liq}} \) is given by\(^{138,139}\),

\[
V_p = \left( \frac{\overline{r}_p}{\overline{r}_k} \right)^2 \left[ V_{\text{liq}} - \Delta t \Sigma A \right] \text{ cm}^3
\]  \hspace{1cm} (57)

where \( \overline{r}_p \) and \( \overline{r}_k \) (in Å) are the average pore and Kelvin radii in the range \( r_2 \) to \( r_1 \), \( \Delta t = (t_2 - t_1) \) Å, and \( \Sigma A \) is the total area of adsorbed film remaining in the previously empty pores and the pores with radii between \( r_2 \) and \( r_1 \) after evaporation out of the pores have occurred. For cylindrical pores, the area of the pores having radii between \( r_2 \) and \( r_1 \) is given by

\[
A = \frac{2 V_{\text{liq}}}{\overline{r}_p} \times 10^4 \text{ m}^2, \text{ with } V_{\text{liq}} \text{ in cm}^3 \text{ and } \overline{r}_p \text{ in Å}. 
\]

Therefore from \( V_{\text{liq}} \), the volume of liquid nitrogen desorbed when the relative pressure of nitrogen is decreased from \( p_2/p_0 \) to \( p_1/p_0 \), the volume of pores \( V_p \) with radii between \( r_2 \) and \( r_1 \) can be determined using Equation (57).

A.2.2 Experimental Procedure

The adsorption isotherms for standard alumina neutral, alumina dried at 800°C and alumina sintered at 1250°C were determined by continuing the procedure used to obtain the BET plot in Appendix A.1, from \( p/p_0 = 0.35 \) to 0.94.

To obtain the desorption isotherm\(^ {113,138}\) for alumina neutral, pure adsorbate (i.e. nitrogen) was allowed to flow through the sample immersed in liquid nitrogen. Then gas mixture, at the required relative pressure (0.94), was allowed to pass through the sample cell while still immersed in the liquid nitrogen. When the recorder indicated a straight line the liquid nitrogen was removed and the desorption signal was monitored. It was calibrated using a known volume of nitrogen and gave a point on the
desorption isotherm corresponding to the relative pressure of nitrogen = 0.94.

The procedure was repeated at progressively lower relative pressures of nitrogen. Thus the desorption isotherm for alumina neutral was obtained.

The desorption isotherms for 800° alumina and alumina sintered at 1250°C were also obtained by repeating the procedure described above.

A.2.3 Results and Discussion

The adsorption and desorption isotherms for alumina neutral, 800° alumina and alumina sintered at 1250°C are shown in Figs. 55, 56 and 57, respectively, and it is clear that all three samples show multilayer adsorption. The sintered sample shows Type II adsorption (according to the classification by Brunauer, Deming, Deming and Teller\(^\text{143}\)), where only multilayer adsorption takes place, while both alumina neutral and 800° alumina sample show Type IV isotherm where both multilayer adsorption and capillary condensation occur. This shows that alumina neutral and 800° alumina are both porous while sintering at 1250°C has removed the porous texture of the catalyst. The presence of pores in alumina neutral and 800° alumina has also given rise to hysteresis of adsorption and desorption isotherms (Figs. 55 and 56), while the absence of pores in sintered alumina is indicated by the observation that there is no hysteresis of adsorption and desorption isotherms (Fig. 57). In addition, the shapes of hysteresis loops in Figs. 55 and 56 may be taken to indicate that the pores are cylindrical\(^\text{139,140}\).

Using the theory outlined above (Appendix A.2.1), the pore volumes for different intervals of relative pressure were determined. The plots of \(\Delta V_p / \Delta r_p\) (ratio of pore volume in each interval to the change in pore radius) versus \(\bar{r}_p\) (the mean pore radius for the same interval) for alumina neutral
Weight (x 10^3 g) of Nitrogen Adsorbed per Gram of Alumina
Fig. 56  Adsorption (×) and Desorption (○) Isotherms of Nitrogen 800°C Alumina
Fig. 57  Adsorption (×) and Desorption (○) Isotherms of Nitrogen on Alumina Sintered
at 1250°C for 6 hours
and 800° alumina are shown in Fig. 3 (Section II.2).

A.3 Determination of Particle Size Distributions

A.3.1 Theory

The particle size distributions of alumina samples were determined by the electrolyte sensing zone method using an Electrozone Celloscope. This instrument determines the number and size of particles suspended in an electrically conductive liquid by application of a resistance principle. The principle consists of forcing the suspension to flow through a small aperture having an immersed electrode on each side as shown in Fig. 58. As each particle passes through the aperture, it replaces its own volume of electrolyte within the aperture, momentarily changing the electrical resistance between the electrodes. To keep the current constant, the Electrozone Celloscope produces a voltage pulse of short duration having a
magnitude proportional to the particle volume. The series of pulses generated by the particles is electrically amplified, scaled, and counted. Measurement precision with 1% on diameter basis has been commonly experienced. Particle density does not affect response but, where gross particle porosity exists, the pores aligned with the aperture axis may provide a degree of electrical conductivity with proportionately lesser pulse height. The particle shape and structure have little effect on the response.

During the experimental process, agitation (short of air bubble inclusion) and chemical dispersants are used to maintain a uniform suspension, avoid flocculation of the particles, or prevent adherence to the sample container. Low particle concentrations (10 to 100 ppm by volume) are used to avoid the passage of two particles through the orifice at the same time (coincidence particle problem).

A.3.2 Experimental Procedure

Electrozone Celloscope Model 112 LTS/ADC, connected to a PDP 8/m minicomputer with 8K capacity and to a telequipment oscilloscope 551B and an ASR33 teletype, was used.

Alumina samples were mixed for 3 hours each on a minimill. Random spatula cuts of the samples were taken and placed in a blender with a 10% sodium hexametaphosphate (Calgon) solution, which was stirred vigorously for 30 secs. Distilled water was added to reduce the phosphate concentration to 4%. To ensure representative sampling, a scoopful of the mixture was taken while stirring rapidly. The sample was then diluted with a 0.75% sodium chloride - 0.5% sodium tetrapyrophosphate electrolyte to proper levels (particle concentration = 50 ppm by volume) for testing on the
Electrozone Celloscope. The blank electrolyte was first checked for background count (none) and then the sample was analyzed using an orifice tube with an orifice diameter = 150 microns.

Plots of the number of counts (scaled) versus particle diameter in centimicrons, for standard alumina neutral and alumina sintered at 1250°C are given in Fig. 4 and the characteristics of the plots in Table V in Section II.3.1.

Determination of the crystalline structure and the trace impurities present in aluminas are described in the Experimental Section XXI.
APPENDIX B

PREPARATION OF A BATCH OF ALUMINA WHICH ACTIVATES ON PYROLYSIS AT HIGH TEMPERATURES, SIMILAR TO THE FIRST BATCH OF ALUMINA NEUTRAL

It was observed in Section XII that, the first batch of Woelm alumina neutral was activated rapidly on pyrolysis at temperatures above 800°C (Figs. 27 and 28). However, the second batch of Woelm alumina neutral and alumina neutral purchased from BDH Chemicals were progressively deactivated on further heating above 800°C (Fig. 29). It was also observed that, Woelm alumina basic on pyrolysis at 1250°C produced a highly active catalyst (Fig. 29). This suggested that the somewhat similar behaviour observed with the first batch of alumina neutral may be due to the presence of residual basic character in that batch of neutral alumina. Therefore, preparation of a batch of alumina which behaves similar to the first batch of alumina neutral was attempted by treatment of the second batch of alumina neutral with a base and also by partial neutralization of alumina basic with an acid, as described below.

B.1 Treatment of the Second Batch of Alumina Neutral with a Base

A stirred slurry of alumina neutral (second batch) was treated with 0.1 N NH₄OH until about 20% of acid sites (determined in Section XVIII.2.3) were neutralized. The alumina was filtered under suction and pyrolyzed at 1250°C for 6 hours. However, the sintered sample thus obtained was found to possess very low catalytic activity. The procedure was repeated using 0.1 N NaOH as the base and again, a catalyst with very low activity was obtained.
B.2 Partial Neutralization of Alumina Basic for TLC

During the attempts to produce highly active 1250° aluminas it was noted that, when alumina basic for TLC was heated at 1250°C, the quartz tube used to hold the boat containing alumina vitrified during the process. Such vitrification was not observed when other aluminas were heated at the same temperature. Examination of compositions of aluminas given in Experimental Section XXI shows that the main difference between alumina basic and other chromatographic aluminas is the presence of a large amount of sodium in the former. The vitrification observed on sintering may be due to vapourization of sodium oxide (which sublimes at 1275°C) and/or hydroxide (boiling point 1390°C) from the basic alumina at 1250°C. Hence removal of the excess (alkali) metal ions from alumina basic by leaching with an acid may ultimately enable production of a less basic alumina which behaves like the first batch of alumina neutral.

Removal of metal ions was carried out by stirring 20 ml of 0.1 M hydrochloric acid solution containing 6 g of alumina basic for 12 hours. The slurry was filtered on a Buchner glass funnel with a fritted disc (medium) and the alumina was washed with excess distilled water for several hours. The water coming through the funnel was very slightly acidic to litmus. The alumina sample in the funnel was washed with more water until the water coming through was no longer acidic. The alumina was allowed to dry under suction, and was then heated at 1250°C under nitrogen for 3 hours (no vitrification of the quartz tube was observed). The sintered sample showed high catalytic activity (but less than that of sintered alumina basic) with linear first order plots as shown in Fig. 59.

A second batch of acid washed alumina basic was prepared and was dried *in vacuo* for two days. It was divided into three parts. One was
Fig. 59  Catalytic Activities of Different Aluminas Prepared from Acid Treated Basic Alumina
heated at 800°C under nitrogen for 4 hours, and another part was heated at 1250°C under nitrogen for 6 hours. First order plots obtained with the three samples are also shown in Fig. 59, and it is clear that on pyrolysis, this acid treated alumina basic behaves similar to that of the first batch of alumina neutral. The small differences in the observed catalytic activities are to be expected since it is related to the amount of acid used and its contact time.

Another batch was prepared by titrating a slurry of basic alumina with 0.1 M hydrochloric acid until the pH was about 7.5, the pH of a slurry of alumina neutral in distilled water (see Table VIII). It was filtered, washed and heated at 1250°C under nitrogen for 6 hours. The sintered alumina again showed a linear first order plot and high activity, although a little less than that of the two previous batches.