

# **Empirically Determining the Rate of Reaction of the Decomposition of Hydrogen Peroxide is Affected by the Use of Different Catalysts**

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## **Abstract:**

There are a number of ways to increase the rate at which a reaction occurs by changing different factors, one of these factors is the presence of catalysts. Here we report the investigation of whether different catalysts affect the rate at which hydrogen peroxide decomposes in attempts to tackle the purpose of determining how variables affect the reaction rate of the decomposition of hydrogen peroxide. This is why we tested five different catalysts: potassium iodide, potassium chloride, sodium iodide, sodium chloride, and lithium chloride. By determining the rate of reaction from the change in pressure over time, we found that the catalyst whose anion was iodine had an evident change in pressure over the 300 seconds that we observed the reaction. Surprisingly, our results lead us to conclude that the more electron orbitals that the anion of the catalyst has, the more readily it will separate in solution, thus, beginning the catalyzation of the hydrogen peroxide.

## **Introduction:**

Chemical reactions require varying lengths of time for completion depending on the characteristics of the reactants and products as well as the conditions under which the reaction is occurring. Some reactions are essentially instantaneous, while others can take days or even months until the products are formed and the reactants are consumed. The speed at which a chemical reaction takes place is commonly referred to as the reaction rate; the more

encompassing definition of reaction rate is the concentration (amount per volume) of product(s) that is formed in a unit of time, or the concentration of reactant(s) that is used in a unit of time.

The human body provides examples of biological rates of reactions, including the decomposition of hydrogen peroxide ( $\text{H}_2\text{O}_2$ ), which is the toxic byproduct of cellular reactions that if left to build up can kill cells.<sup>1</sup> That is why there is an enzyme, which is a biological catalyst that increases the rate of the chemical reaction without undergoing any change, that works inside of the cells to catalyze the breakdown of hydrogen peroxide to oxygen ( $\text{O}_2$ ) and water ( $\text{H}_2\text{O}$ ).

There is a theory that can be used to predict the rates of reactions, usually of gases, called Collision Theory. Collision Theory is based on the kinetic laws of gases and states that if reactants collide in the correct orientation with a minimum amount of initial energy, then the chemical reaction will occur “successfully.”<sup>2</sup> Thus, according to Collision Theory, the rate at which a chemical reaction occurs is equivalent to the frequency of *effective* collisions.<sup>3</sup> Although it is difficult to control the orientation of collisions, it is still possible to influence the rate of reaction with certain factors, such as temperature, nature of reactants, concentration of reactants, and presence of catalysts.

As a group, we sought to design a set of experiments that would help us determine how the presence of different catalysts would affect the reaction rate of the decomposition of hydrogen peroxide to oxygen and water. Chemical catalysts are used to speed up reactions since they allow for the reaction to occur through an alternate pathway that uses less energy, which

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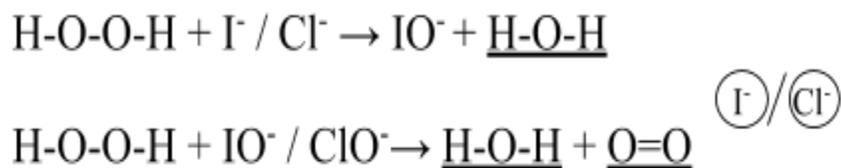
<sup>1</sup>What Can Affect Enzyme Controlled Reactions? Retrieved October 29, 2018, from <https://www.birmingham.ac.uk/undergraduate/preparing-for-university/stem/Biology/stem-legacy-enzymes.aspx>.

<sup>2</sup>Libretexts. (2017, May 22). Collision Theory. Retrieved October 29, 2018, from [https://chem.libretexts.org/Textbook\\_Maps/Physical\\_and\\_Theoretical\\_Chemistry\\_Textbook\\_Maps/Supplemental\\_Modules\\_\(Physical\\_and\\_Theoretical\\_Chemistry\)/Kinetics/Modeling\\_Reaction\\_Kinetics/Collision\\_Theory](https://chem.libretexts.org/Textbook_Maps/Physical_and_Theoretical_Chemistry_Textbook_Maps/Supplemental_Modules_(Physical_and_Theoretical_Chemistry)/Kinetics/Modeling_Reaction_Kinetics/Collision_Theory)

<sup>3</sup>Ibid.

indicatively increases the rate of reaction. We decided to use a constant amount of five different catalysts: potassium iodide (KI), potassium chloride (KCl), lithium chloride (LiCl), sodium iodide (NaI), and sodium chloride (NaCl), which we added to the hydrogen peroxide and then indirectly measured the reaction rate of the decomposition of the hydrogen peroxide using a Vernier Gas Pressure Sensor (Figure 1).

The Vernier Gas Pressure Sensor did not give us the direct rate of reaction, however, we were able to observe the rate of reaction by using the sensor to measure the amount of oxygen produced. As the reaction progressed, more oxygen was produced and attempted to escape from the flask, which in turn increased the pressure. Thus, we were able to indirectly measure the reaction rate by using the change in pressure (kPa) over time (s) from the best fit line of our graph. We hypothesized that the catalyst whose cation and anion were lower down in the periodic table would make the reaction occur at a faster rate because the valence orbital is farther away from the nucleus, which meant they would be less attracted to each other and more likely to react with other particles in the solution. Based on that, we thought that the order from faster reaction rate to slowest reaction rate would be KI, KCl, NaI, NaCl, and then LiCl because of the distances between the ions.



**Figure 1:** This is the balanced chemical equation of the decomposition of hydrogen peroxide in the presence of a catalyst whose anion is either iodide or chloride.

## Results:

*The Study of How Trials of Different Catalysts Affect the Change in Pressure Over Time For the Decomposition Reaction of Hydrogen Peroxide*

<b>Catalysts Used During Experimentation</b>	<b>Change in Pressure Over Time (kPa/s)</b>
Trial 1 of Potassium Iodide (KI)	0.01072 kPa/s
Trial 2 of Potassium Iodide (KI)	0.009163 kPa/s
Trial 1 of Potassium Iodide (KCl)	0.001844 kPa/s
Trial 2 of Potassium Iodide (KCl)	-0.002660 kPa/s
Trial 1 of Sodium Iodide (NaI)	0.01457 kPa/s
Trial 2 of Sodium Iodide (NaI)	0.007380 kPa/s
Trial 1 of Sodium Chloride (NaCl)	-0.001225 kPa/s
Trial 2 of Sodium Chloride (NaCl)	-0.0005896 kPa/s
Trial 1 of Lithium Chloride (LiCl)	-0.0004169 kPa/s
Trial 2 of Lithium Chloride (LiCl)	-0.00005955 kPa/s

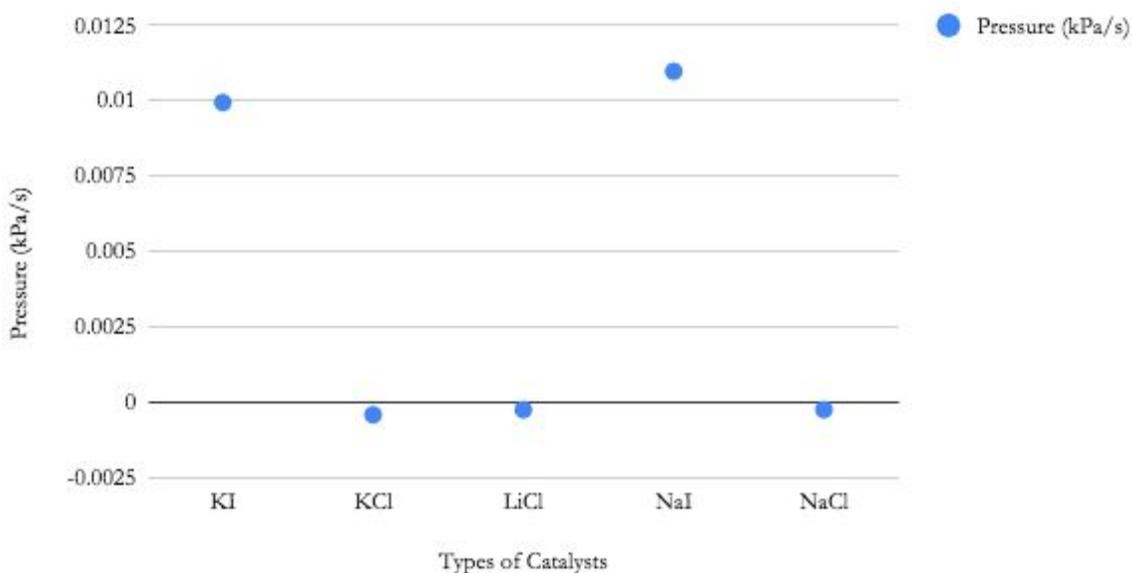
**Figure 2:** This table details the change in pressure over time for the two trials that were conducted for each of the five catalysts (KI, KCl, NaI, NaCl, and LiCl). These values of the change in pressure over time were generated from the slopes of the best-fit lines of each graph made on Logger Pro.

*The Average of Change in Pressure Over Time for the Trials of Each Catalyst Used During the Decomposition Reaction of Hydrogen Peroxide*

<b>Catalysts Used During Experimentation</b>	<b>Average of the Change in Pressure Over Time (kPa/s)</b>
Potassium Iodide (KI)	0.0099415 kPa/s
Potassium Chloride (KCl)	-0.000408 kPa/s
Sodium Iodide (NaI)	0.010975 kPa.s
Sodium Chloride (NaCl)	-0.000238225 kPa/s
Lithium Chloride (LiCl)	-0.000238225 kPa/s

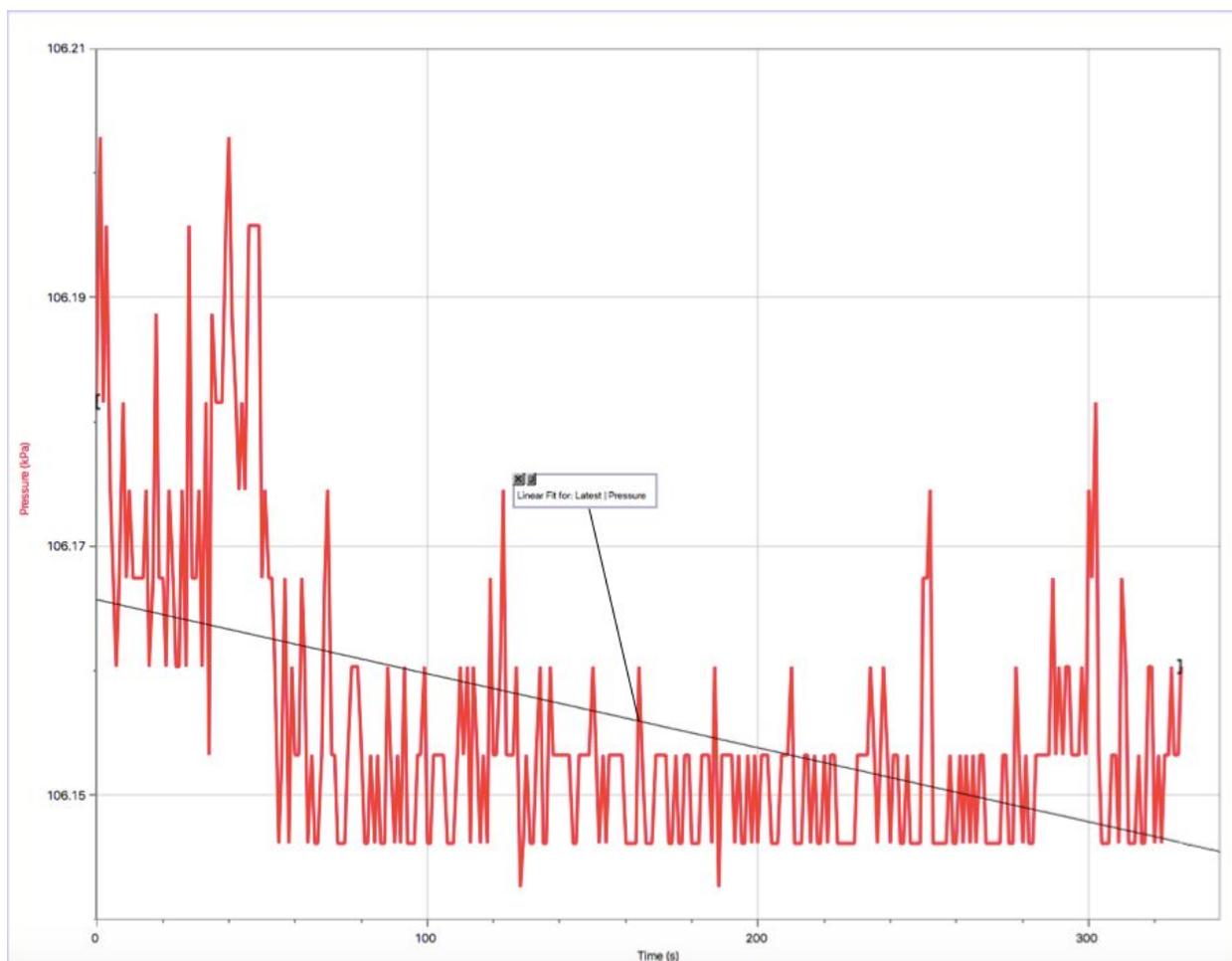
**Figure 3:** This table details the average change in pressure over time for the two trials that were conducted for each of the five catalysts (KI, KCl, NaI, NaCl, and LiCl). These values were calculated by adding the change in pressure over time for the two trials of a respective catalysts, and then dividing it by two in order to receive the average.

**Investigating the Average Change in Pressure Over Time for Two Trials of Each of the Five Different Catalysts**



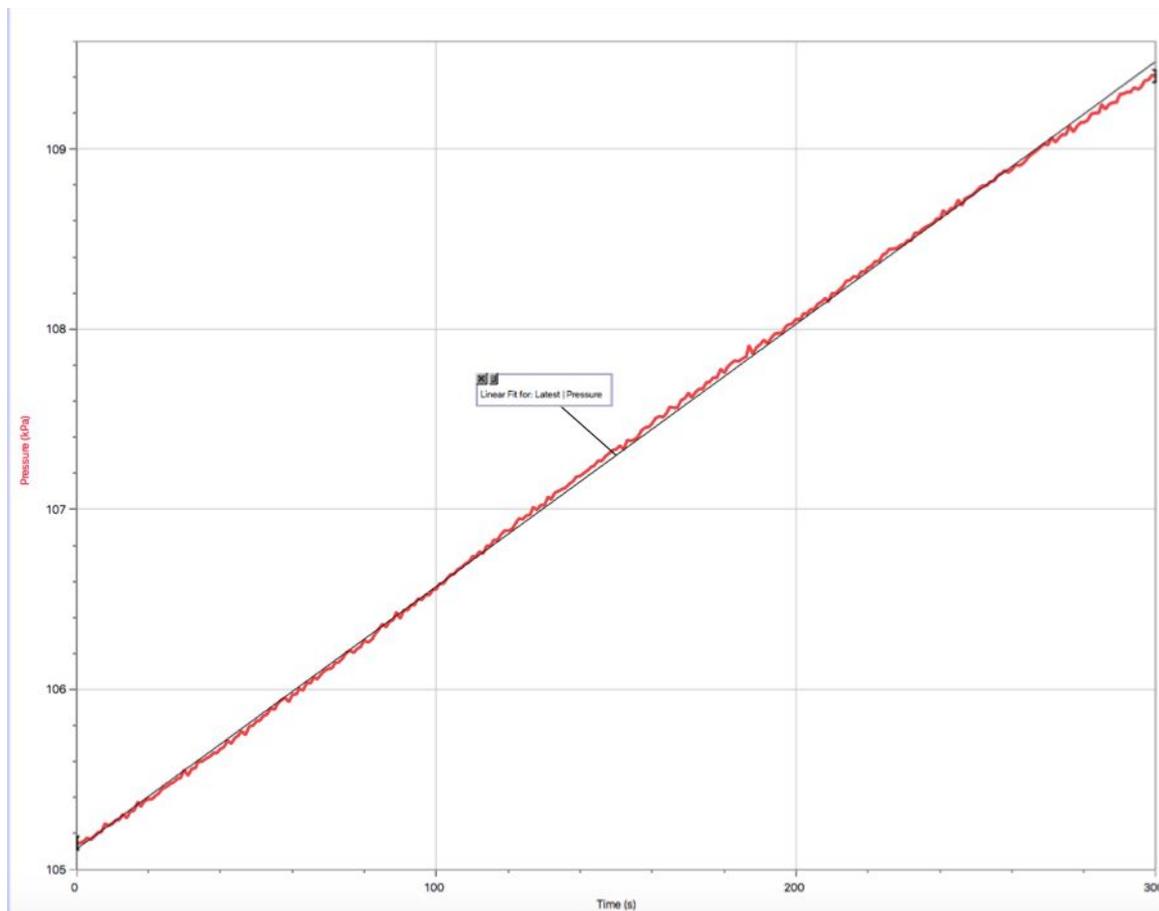
**Figure 4:** This graph details the average change in pressure over time for the two trials that were conducted for each of the five catalysts (KI, KCl, NaI, NaCl, and LiCl). These values were calculated by adding the change in pressure over time for the two trials of a respective catalysts, and then dividing it by two in order to receive the average.

*The Graph of the Change in Pressure Over Time When Lithium Chloride (Trial 2) Catalyzes the Decomposition of Hydrogen Peroxide*



**Figure 5:** This is our pressure over time graph for Trial 2 of lithium chloride, and from the best-fit line it is evident that there is a negative slope. This was the smallest negative slope of change in pressure over time (kPa/s) in comparison to all of the other values.

*The Graph of the Change in Pressure Over Time When Sodium Iodide (Trial 1) Catalyzes the Decomposition of Hydrogen Peroxide*



**Figure 6:** This is our pressure over time graph for Trial 1 of sodium iodide, and from the best-fit line it is evident that there is a positive slope. Out of all of the other values, this graph had the greatest slope of change in pressure over time (kPA/s).

### **Discussion:<sup>4</sup>**

Once we completed our experiment using the five different catalysts to determine whether they affected the reaction rate of the decomposition of hydrogen peroxide, we assessed our graphs that Logger Pro had generated based on the increase or decrease of pressure from the oxygen gas produced over the 300 seconds that we observed the reaction. We utilized Logger Pro to create a line of best-fit on each of our graphs, so that we could use their respective slopes

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<sup>4</sup>Furman, O. (October 29, 2018). Peer Review of Discussion Section.

to discern the change in pressure over the time that we ran the experiment (essentially the reaction rate). We noticed that in our trials that used sodium chloride, lithium chloride, and the second trial of potassium chloride that the values of the change in pressure over time was negative because over the course of the 300 seconds the pressure decreased over time, which can be seen in the data table labeled Figure 2. Out of all of the trials conducted for each of the catalysts, the second trial of lithium chloride had the most negative change in pressure over time as the slope of this graph was  $-0.00005955$  kPa/s. Figure 5 is the graph that represents the second trial of lithium chloride, and it is a clear representation of the inconsistency in our data for the catalysts that had negative slopes. We presumed that these negative slopes were caused from when we applied pressure on the two hole stopper when we did not think that the stopper was in correctly. Therefore, when we stopped applying pressure midway through the 300 seconds, this would cause the pressure to decrease, and we suppose that this is a primary source of error since it caused us to have negative slopes for some of our catalysts.

The average change in pressure over time for both of the trials of potassium chloride, sodium chloride, and lithium chloride were, in fact, negative as seen in Figure 3. We deduced that the change in pressure over time for the three catalysts, potassium chloride, sodium chloride, and lithium chloride, whose anion was chlorine did not display a measurable change in pressure over time as show in our graph in Figure 4. This lead us to assume that since that overall change in pressure that we observed for these three catalysts were negligible. On the other hand, the catalysts whose anion was iodine had greater changes in pressure over time than that of the other three catalysts. As highlighted in both Figure 3 and Figure 4, potassium iodide and sodium iodide were the only catalysts that had a positive slope

over the span of the 300 seconds. The values of the average change in pressure over time for potassium iodide and sodium iodide was 0.0099415 kPa/s and 0.010975 kPa/s respectively, and out of all the trials conducted for each of the catalysts, trial one of sodium iodide had the greatest increase in pressure over time as the slope of this graph seen in Figure 6 was 0.01457 kPa/s.

Based on the values of the average change in pressure for the five catalysts, we concluded that the catalysts with the anion, iodine, possessed the greatest change in pressure over time. We expected that the compound whose distance between the cation and anion was greatest to have the fastest reaction rate, which is why we predicted that potassium iodide and potassium chloride would have the fastest reaction rates, however, this was evidently not supported by our data. Instead, we observed that the number of orbital rings that the anion had affected the change in pressure over time. The anion of the three catalysts that had a negative slope was chlorine, which had fewer valence orbitals than iodine, meaning that the attraction between the valence electrons and nucleus was greater, so that the time for the catalysts to separate into the solution was consequently more. Since iodine is further down in the periodic table than chlorine, it has more orbital rings and is more readily able to separate from its cation, which explains why the catalysts whose anion was iodine did, in fact, have a greater change in pressure over time.

Ultimately, our hypothesis was not supported as we thought it would have been as we thought that the placement of *both* the anion and the cation and how far down they were on the periodic table would affect the change in pressure over time. However, after experimentation, we concluded that the placement of the anion is the significant factor in affecting the change of pressure over time, or the rate of reaction, of the decomposition of hydrogen peroxide. The attraction between the nucleus and valence electrons of the anion of the catalyst determines the

rate at which the pressure will increase over time because the lesser the attraction the faster the reaction rate. Our data values of the decomposition of hydrogen peroxide when using catalysts with iodine supported this as the reaction rate was much faster since it more readily separates from its cation than chlorine does. In retrospect, this makes much more sense than what we had originally postulated because during the decomposition of hydrogen peroxide the anion is the mechanism that takes the reactants of the hydrogen peroxide under a lower energy pathway, which leads us to believe that the cation doesn't play a role in catalyzing the reaction.

Because we were under the assumption that the distance between the cation and anion affected the rate at which hydrogen peroxide decomposed, there wasn't much variation in the anion of our catalysts. Based off of what we concluded during this investigation, in future experiments, it would be valuable to further investigate the role that the anion of the catalyst has on the rate of the decomposition of hydrogen peroxide by using catalysts whose anion was bromine, asinine, or transition metals. This would allow us to determine whether what we deduced about the anion of the catalysts was, in fact, supported. Besides this, in future experiments, it would also be beneficial to extend the time in which we observed the reaction because the span of 300 seconds did not exhibit the scale of increase in pressure over time that we expected.

### **Conclusion:**

After investigating how the five different catalysts, potassium iodide, potassium chloride, sodium iodide, sodium chloride, and lithium chloride, influence the rate of reaction of the decomposition of hydrogen peroxide, we observed that the catalysts whose anion was iodine had a negative change in pressure over time, and that the catalysts whose anion was chlorine had a positive change in pressure over time. From this, we concluded that the distance between the

valence orbital and the nucleus of the catalyst's anion determines how readily it will separate and catalyze the reactants, thus, affecting how fast or slow that the reaction will occur.

### **Experimental Methods:**

Began by creating 20 mL solution of each catalyst with a molarity of 0.5 using 0.1 mole of each catalyst in solid form. These solutions were left out to equilibrate. Used Dr. Lurain's original procedure on the Decomposition of Hydrogen Peroxide with some adjustments to better tend to the purpose of this experiment.<sup>5</sup> Mixed 10 mL of 3% hydrogen peroxide with 45 mL of distilled water in a 125 mL Erlenmeyer flask and measured the temperature. Followed steps 2b through 2d of Dr. Lurain's procedure, and then measured out 10 mL of the 0.5 M KI solution and took its temperature.<sup>6</sup> Transferred the KI solution into the flask, sealed the flask tightly with the two hole stopper and surrounded it with parafilm, and turned on the stirrer to Speed 3. Followed steps 3b and 3c of Dr. Lurain's procedure, and then repeated this entire process for the other four catalysts.<sup>7</sup>

### **Acknowledgements:**

I would like to acknowledge Dr. Lurain for guiding my lab group in the right direction during our experimentations. Also, I would like to thank my entire advanced experimental chemistry class for helping my group modify our procedure as well as our analysis during group meetings.

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<sup>5</sup>Lurain, A. (2018, October 16). *Decomposition of Hydrogen Peroxide: A reaction rates case study*. New York, Brooklyn.

<sup>6</sup>Ibid.

<sup>7</sup>Ibid.

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