SL Chemistry Internal Assessment

Efficiency per unit mass of five different catalysts for hydrogen peroxide decomposition

Candidate code: gbc594

Research question

Which catalyst out of Manganese (IV) Oxide, Lead (II, IV) Oxide, Potassium Iodide, Copper (II) Oxide, and Iron (III) Oxide will most rapidly catalyze the decomposition of hydrogen peroxide per unit mass based on the initial rate of reaction measured through the time taken for the catalyzed reaction to produce a specific volume of gas?

Introduction

Background

The "elephant's toothpaste" is one of the first chemical experiments I witnessed as a kid. The reaction happens when high concentrations of hydrogen peroxide rapidly decays into oxygen gas and water in the presence of potassium iodide as a catalyst, where the rapid release of small oxygen bubbles from the solution causes a large volume foam to emerge. The research question originates from the thought that there could be a more effective catalyst than potassium iodide for the "elephant's toothpaste" reaction.


Figure-1 (Burns)

Relevant Theory

Hydrogen peroxide, H_2O_2, is an unstable compound that spontaneously decays into water and oxygen through the equation:

$$2H_2O_2 \rightarrow 2H_2O + O_2 \quad [1]$$
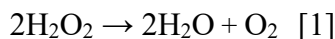
However, the spontaneous decay of hydrogen peroxide is relatively slow, with a half-life of 10-20 hours *(FMC Corporation)* in open air at 298 Kelvin. This makes spectacles such the "elephant's toothpaste" reaction impossible without a catalyst.

Catalysts are chemicals that increase the rate of certain reactions without undergoing any permanent chemical change itself. Illustrated on *Figure-1 (Burns)*, the presence of catalysts offer an alternative pathway to the reaction that requires an overall lower activation energy. With reference to the Maxwell-Boltzmann distribution on *Figure-2 (Polypompholyx)*, this statistically allows a higher proportion of the reactants to have enough energy to react, and hence increases the rate of reaction.

Figure-2 (Polypompholyx)

Most catalysts do undergo chemical change during the reaction they catalyze, but the chemical change is reversed after the reaction has finished. Potassium iodide as the catalyst used in the "elephant's toothpaste" reaction is one such example, as can be seen from the following system of equations describing how potassium iodide catalyzes the decomposition of hydrogen peroxide:

$$KI \rightarrow K^+ + I^- \quad [2]$$

$$I^- + H_2O_2 \rightarrow H_2O + OI^- \quad [3]$$
$$OI^- + H_2O_2 \rightarrow H_2O + O_2 + I^- \quad [4]$$
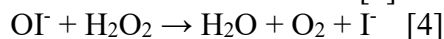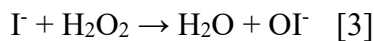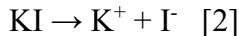
This shows that hydrogen peroxide initially acts as an oxidizing agent that oxidizes the dissociated iodide ions to hypoiodite ions, but it later acts as a reducing agent that reduces hypoiodite ions back to iodide ions. Adding the two conversion equations between iodide and hypoiodite together, the iodide and hypoiodite on each side of the equation cancels out to give just:
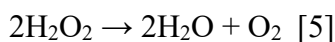
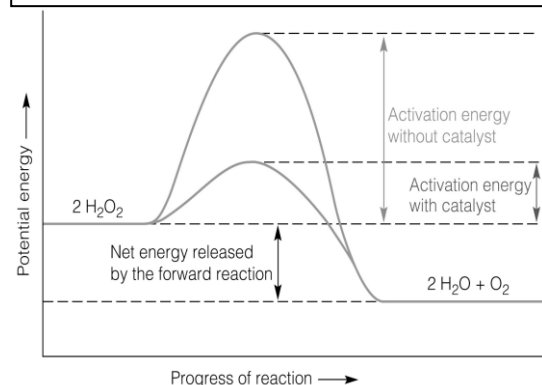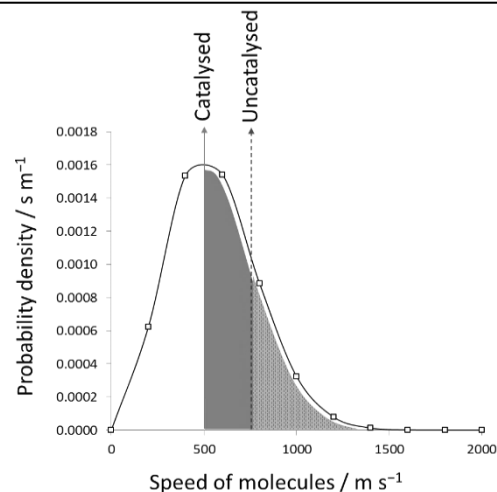$$2H_2O_2 \rightarrow 2H_2O + O_2 \quad [5]$$
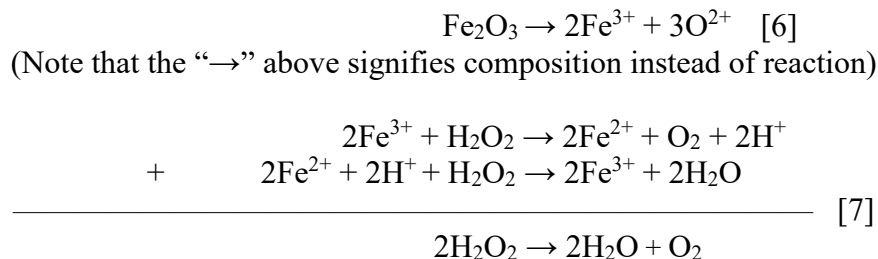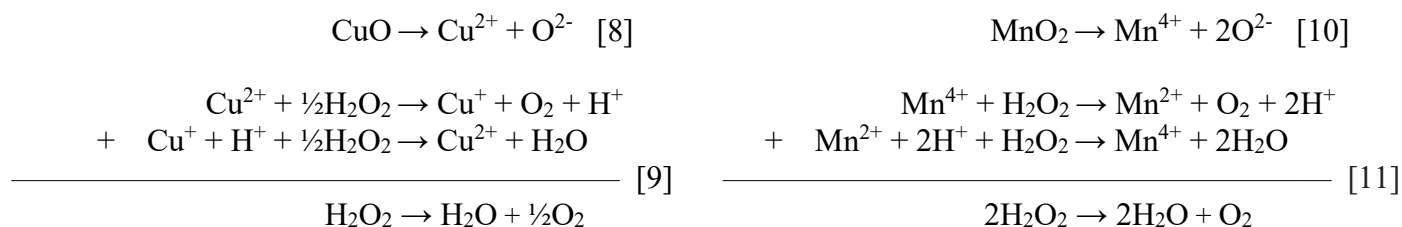
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, which proves that potassium iodide is a catalyst of hydrogen peroxide decomposition, and quantitatively shows that 1 mol of potassium iodide will be able to decompose 2 mol of hydrogen peroxide per catalytic cycle.

Besides potassium iodide, a number of transition metal oxides are also catalysts for hydrogen peroxide decomposition. Transition metals' ability to change between different oxidation states also allows catalytic cycles with hydrogen peroxide to occur. Take iron (III) oxide for example:
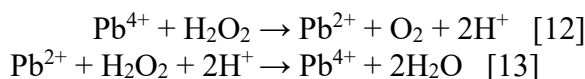
$$Fe_2O_3 \rightarrow 2Fe^{3+} + 3O^{2+} \quad [6]$$
(Note that the "$\rightarrow$" above signifies composition instead of reaction)

$$
\begin{aligned}
& 2Fe^{3+} + H_2O_2 \rightarrow 2Fe^{2+} + O_2 + 2H^+ \\
+ \quad & 2Fe^{2+} + 2H^+ + H_2O_2 \rightarrow 2Fe^{3+} + 2H_2O \\
\hline
& 2H_2O_2 \rightarrow 2H_2O + O_2
\end{aligned}
\quad [7]
$$

As shown, hydrogen peroxide can reduce Fe^{3+} to Fe^{2+} and also oxidize Fe^{2+} back to Fe^{3+}, which creates a catalytic cycle similar to that of potassium iodide, where 1 mol of iron (III) oxide is able to decompose 2 mol of hydrogen peroxide per catalytic cycle.

The catalytic cycle for copper (II) oxide and manganese (IV) oxide largely follow the same logic, which that:

$$CuO \rightarrow Cu^{2+} + O^{2-} \quad [8]$$

$$MnO_2 \rightarrow Mn^{4+} + 2O^{2-} \quad [10]$$

$$
\begin{aligned}
& Cu^{2+} + \tfrac{1}{2}H_2O_2 \rightarrow Cu^+ + O_2 + H^+ \\
+ \quad & Cu^+ + H^+ + \tfrac{1}{2}H_2O_2 \rightarrow Cu^{2+} + H_2O \\
\hline
& H_2O_2 \rightarrow H_2O + \tfrac{1}{2}O_2
\end{aligned}
\quad [9]
$$

$$
\begin{aligned}
& Mn^{4+} + H_2O_2 \rightarrow Mn^{2+} + O_2 + 2H^+ \\
+ \quad & Mn^{2+} + 2H^+ + H_2O_2 \rightarrow Mn^{4+} + 2H_2O \\
\hline
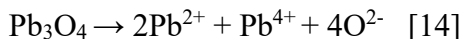& 2H_2O_2 \rightarrow 2H_2O + O_2
\end{aligned}
\quad [11]
$$

1 mol of copper (II) oxide is only able to decompose 1 mol of hydrogen peroxide per catalytic cycle, whereas 1 mol of manganese (IV) oxide will be able to 2 mol of hydrogen peroxide per catalytic cycle.
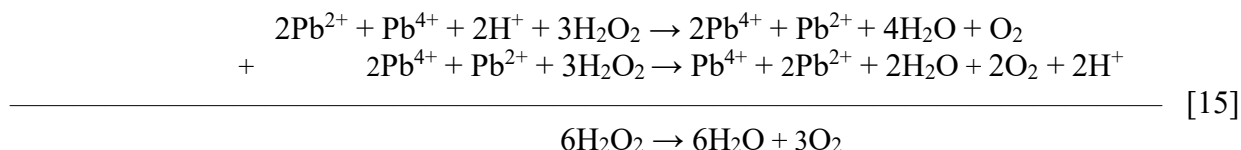
The same logic also applies lead (II, IV) oxide, but we have to realize that lead ions exists as both lead (II) and lead (IV) in this compound. The catalytic cycle between Pb^{2+} and Pb^{4+} are as below:

$$Pb^{4+} + H_2O_2 \rightarrow Pb^{2+} + O_2 + 2H^+ \quad [12]$$
$$Pb^{2+} + H_2O_2 + 2H^+ \rightarrow Pb^{4+} + 2H_2O \quad [13]$$

Given that:
$$Pb_3O_4 \rightarrow 2Pb^{2+} + Pb^{4+} + 4O^{2-} \quad [14]$$

, it can be deduced that the complete catalytic cycle for lead (II, IV) oxide is:

$$
\begin{aligned}
& 2Pb^{2+} + Pb^{4+} + 2H^+ + 3H_2O_2 \rightarrow 2Pb^{4+} + Pb^{2+} + 4H_2O + O_2 \\
+ \quad & 2Pb^{4+} + Pb^{2+} + 3H_2O_2 \rightarrow Pb^{4+} + 2Pb^{2+} + 2H_2O + 2O_2 + 2H^+ \\
\hline
& 6H_2O_2 \rightarrow 6H_2O + 3O_2
\end{aligned}
\quad [15]
$$

Hence, 1 mol of lead (II, IV) oxide will be able to decompose 6 mol of hydrogen peroxide per catalytic cycle.

Variables

Independent Variable	Definition	Catalysts
	Five different catalysts for hydrogen peroxide decomposition.	Manganese (IV) Oxide, MnO_2 Lead (II, IV) Oxide, Pb_3O_4 Potassium Iodide, KI Copper (II) Oxide, CuO Iron (III) Oxide, Fe_2O_3

Dependent Variable	Symbol	Definition	Proximate Variables		
	r_0	The instantaneous initial rate of reaction when the catalyst comes into contact with hydrogen peroxide.	Proximate independent variable	Symbol	Definition
				V	Volume of gas produced by the reacting system (cm^3)
		The instantaneous initial rate will be indicative a catalyst's efficiency at decomposing hydrogen peroxide ($cm^3\ s^{-1}$).	Proximate dependent variable	Symbol	Definition
				t	Time taken for the system to produce V cm^3 of gas (seconds)

Controlled Variables	Symbol	Definition	Controlled Value	Reason for control
	m	Mass of the catalyst used	1.0 ± 0.1 gram	The aim of this experiment is to determine the efficiency of each catalyst per unit mass.
	T_0	Initial temperature of the system	21.0 ± 0.3 °C	Initial temperature influences the initial rate of reaction independent of the catalyst used.
	A	Surface area of the catalyst	All catalysts used are powdered (unable to control quantitatively)	Surface area influences the initial rate of reaction independent of the catalyst used.
	$[H_2O_2]$	Concentration of hydrogen peroxide solution used in the experiment	0.176 ± 0.002 mol dm^{-3} (See page 4 for calculation)	Concertation of hydrogen peroxide as a reactant influences the initial rate of reaction independent of the catalyst used.
	$V[H_2O_2]$	Volume of hydrogen peroxide solution used in the experiment	200.0 ± 0.2 cm^3 (Class B 200 cm^3 volumetric flask)	The volume hydrogen peroxide solution determines the length of time that the reaction is going to persist. Since time is measured as the independent variable, this variable should kept constant.

Method

Chemicals and equipment list

List of chemicals		
Chemical	Quantity	Description
Manganese (IV) Oxide	5 grams	Solid, finely powdered
Lead (II, IV) Oxide	5 grams	Solid, finely powdered
Potassium Iodide	5 grams	Solid, finely powdered
Copper (II) Oxide	5 grams	Solid, finely powdered
Iron (III) Oxide	5 grams	Solid, finely powdered
6% Hydrogen Peroxide solution	500 cm³	Assuming ± 1% uncertainty of concentration
Distilled water	-	-

List of equipment		
Equipment	Quantity	Uncertainty
500 cm³ Conical Flask	1	-
100 cm³ beaker	2	-
Glass vial	5	-
Stop watch	1	± 0.01 seconds
Class B 20.00 cm³ pipette and suction bulb	1	± 0.03 cm³; ± 0.15%
Class B 200.0 cm³ volumetric flask	1	± 0.2 cm³; ± 0.08%
100 cm³ Gas syringe set (smallest division of 5 cm³)	1	± 2.5 cm³
BSM1204 digital electronic balance	1	± 0.0001 gram
TFA Pocket-digitemp S Digital Thermometer	1	± 0.1 °C
Glass stirring rod	1	-
Molar and pestle	1	-
Clamp stand and clamp	1	-

Safety precaution for hazardous chemical

The metal oxides involved in this investigation, especially Lead (II, IV) Oxide, are labelled as very toxic for aquatic life with long lasting effects under hazard statement H410 and harmful if inhaled under hazard statement H332 *(WSH council)*. therefore any liquid contaminated by the compound should be thoroughly filtered before going down the sink, while the powdered solid should be handled with extreme care to prevent it from becoming airborne. Hands are washed thoroughly after handling said chemicals, while rubber gloves and safety goggles are worn at all times during the experiment.

Preparing diluted hydrogen peroxide solution

A test trial with manganese (IV) oxide shows that the reaction would be too rapid for any meaningful measurement to be taken using the requested 6% hydrogen peroxide solution, hence the following steps are taken to prepare diluted hydrogen peroxide solutions on which meaningful measurements can be made.

1. A 20 cm³ aliquot of the 6% hydrogen peroxide solution is taken with the class B pipette and suction bulb, and transferred into the class B 200 cm³ volumetric flask.
2. The volumetric flask is filled with distilled water to the 200 cm³ mark.

The molar concentration of diluted hydrogen peroxide solution prepared by this method is determined as follows:

6% hydrogen peroxide solution translates to 6 grams of hydrogen peroxide per 100 cm³. Original solution: $[H_2O_2] = \left(\frac{60 \text{ g} \pm 1\%}{1 \text{ dm}^3}\right) \div 34 \text{ g mol}^{-1}$ $[H_2O_2] = 1.76 \text{ mol dm}^{-3} \pm 1\%$ During step 1: $n(H_2O_2) = \frac{1.76 \text{ mol} \pm 1\%}{1 \text{ dm}^3} \times (0.02 \text{ dm}^3 \pm 0.15\%)$ $n(H_2O_2) = 0.0352 \text{ mol} \pm 1.15\%$ After step 2: $[H_2O_2] = \frac{0.0352 \text{ mol} \pm 1.15\%}{0.2 \text{ dm}^3 \pm 0.08\%}$ $[H_2O_2] = 0.176 \text{ mol dm}^{-3} \pm 1.23\%$ $[H_2O_2] = 0.176 \pm 0.002 \text{ mol dm}^{-3}$	Molar mass of H_2O_2: 34 g mol⁻¹

Experiment procedure
Repeat two trials of the following step for each of the five catalysts:

1. The catalyst should be grinded with a pestle and molar if it contains visible large pieces.

2. Measure and transfer 1 gram ± 0.002% of the catalyst into a glass vial by using an electronic balance.

3. Transfer the prepared 200 cm^3 of 0.10 ± 0.01 mol dm^{-3} diluted hydrogen peroxide solution into the 500 cm^3 conical flask.

4. Carefully lower the glass vial containing the catalyst into the conical flask by using a glass stirring rod as shown in *Figure-3*, make sure the glass vial rests on the wall of the conical flask so that the catalyst does not come into contact with the solution.

5. Measure and record the temperature of the solution in the conical flask with a digital thermometer, note that the experiment should only proceed when the initial temperature of the solution is 21.0 ± 0.3 °C.

6. Stopper the conical flask with a gas stopper connected to the gas syringe with a delivering tube as shown on *Figure-4*. Check to make sure that the system is airtight.

7. Shake the conical flask to let the glass vial to fall into the hydrogen peroxide solution as demonstrated on *Figure-4*, start timing on the stopwatch the moment that the catalyst is exposed to the solution.

8. Measure and record the time taken for the reacting system to produce 10 cm^3, 20 cm^3, 30 cm^3, 40 cm^3, 50 cm^3, 60 cm^3 of gas.

Figure-3: Setting up the reaction vessel

Glass stirring rod
Thermometer
Conical flask
Glass vial
Diluted hydrogen peroxide solution
≈1 gram powdered catalyst

Figure-4: Experiment setup

Gas Syringe
Delivery tube
Gas stopper
Clamp stand
Conical flask
Glass vial
Diluted hydrogen peroxide solution
≈1 gram powdered catalyst
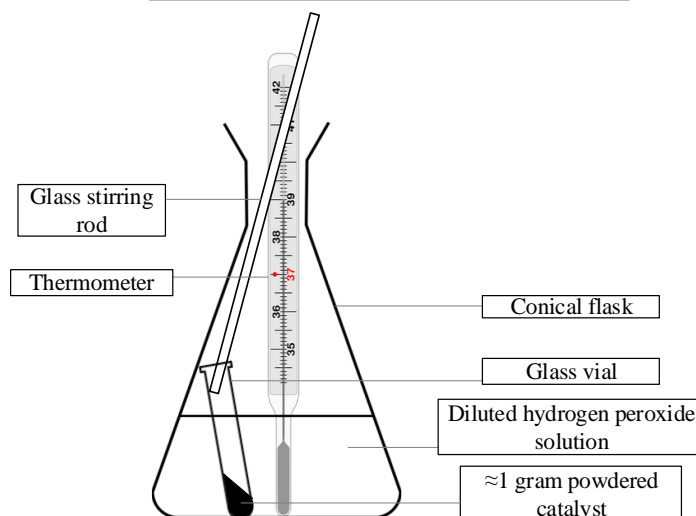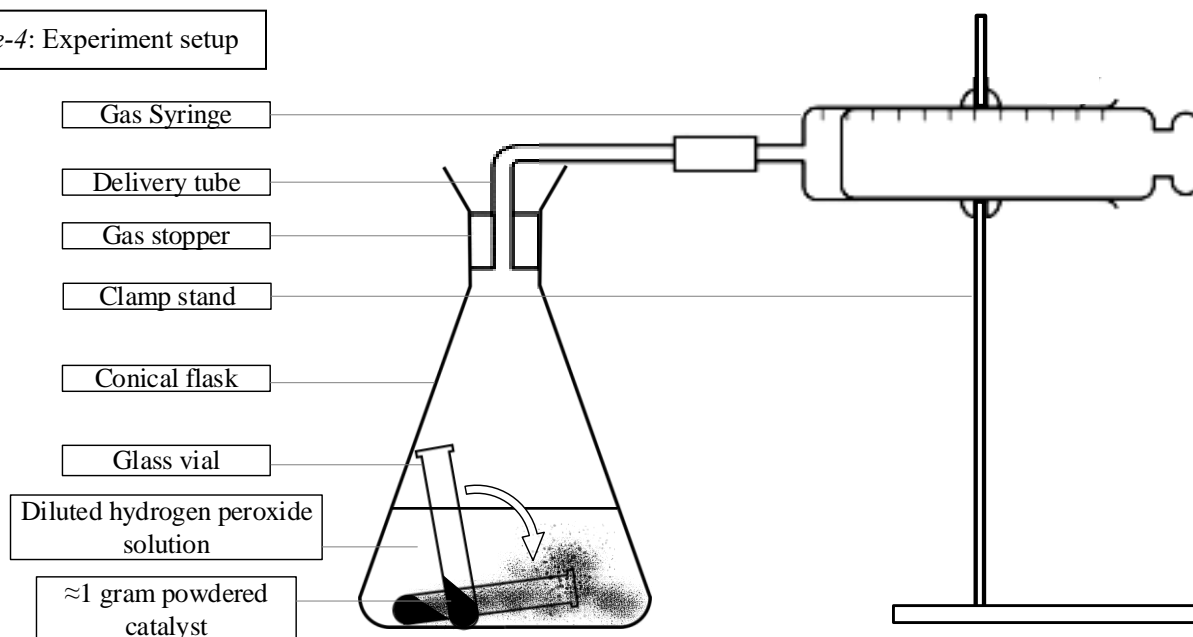
Processing raw results

To evaluate the instantaneous initial rate r_0 of each reacting system, we must first be able to model the rate of reaction r as a function of time t from the raw measurements of t and V. The definition of rate of reaction r measured in terms of the volume of gas produced by a system is:

$$r = \frac{\Delta V}{\Delta t} \quad [16]$$

, hence it can be said that:

$$r = \frac{V_2 - V_1}{t_2 - t_1} \quad [17]$$

, where V_1 and V_2 are two consecutive volumes at which time is measured, while t_1 and t_2 are two consecutive measurements of time at V_1 and V_2. It should be noted that [2] determines r as the average rate of reaction between the time interval $t_2 - t_1$. Hence if we are to assign a specific value of time to r, this value should be the average time t_μ of the interval, where:

$$t_\mu = \frac{t_2 + t_1}{2} \quad [18]$$

All raw data are processed using [17] and [18] to evaluate the behavior of the rate of reaction for each reacting system.

Sample calculation of rate of reaction and average time: *Table-1*, processed data, for $10 \text{ cm}^3 \leq V \leq 20 \text{ cm}^3$			
Rate of reaction r: $\Delta V = 20 \text{ cm}^3 - 10 \text{ cm}^3$ $\Delta V = 10 \text{ cm}^3$ $\Delta t = 61.1\text{s} - 32\text{s}$ $\Delta t = 29.1\text{s}$	$r = \frac{10 \text{ cm}^3}{29.1 \text{ s}}$ $r \approx 0.3 \text{ cm}^3 \text{ s}^{-1}$	Average time t_μ: $t_1 = 32 \text{ s}$ $t_2 = 61.1\ 0.4 \text{ s}$	$t_\mu = \frac{32 \text{ s} + 61.1 \text{ s}}{2}$ $t_\mu \approx 47 \text{ s}$

Treatment of errors and uncertainties

To account for both the uncertainty in repeated trials of measurement and the instrument error in individual measurements at the same time, an aggregate uncertainty is calculated for every trial average as follows:

$$\text{Aggregate error} = \frac{\text{Trial maximum} - \text{Trial minimum}}{2} + \sum \text{Instrument error} \quad [19]$$

, where each measured value's deviation from the average is added to the sum of the instrument error.

Sample calculation of aggregate error: *Table-1*, raw data, Trial average $= 32 \pm 2$ s	
Trial maximum $= 34.11$ s Trial minimum $= 30.62$ s Total instrument error $= \pm 0.01$	Trial average $= \frac{34.11 + 30.62}{2} \pm \left(\frac{34.11 - 30.62}{2} + 0.01 \right)$ seconds Trial average $= 32.365 \pm 1.765$ seconds Trial average $\approx 32 \pm 2$ seconds

When uncertainties are propagated down into calculated values, the sum of the relative uncertainties associated with each value used in the calculation is taken, and then multiplied by the value of the result to give the propagated absolute error.

$$\text{Propagated error} = \sum \frac{\text{Associated error}}{\text{Value used in calculation}} \times \text{Calculated value} \quad [20]$$

Sample calculation of propagated error: *Table-1*, processed data, $r = 0.3 \pm 0.2 \text{ cm}^3 \text{ s}^{-1}$	
$t_1 = 32 \pm 2$ s $t_2 = 61.1 \pm 0.4$ s $V_1 = 10.0 \pm 2.5 \text{ cm}^3$ $V_2 = 20.0 \pm 2.5 \text{ cm}^3$	$r = \frac{(61.1 \pm 0.4) - (32 \pm 2)}{(20.0 \pm 2.5) - (10.0 \pm 2.5)} \text{ cm}^3 \text{ s}^{-1}$ $r = \frac{29.1 \pm 8.25\%}{10 \pm 50\%} \text{ cm}^3 \text{ s}^{-1}$ $r = 0.291 \pm (0.291 \times 58.25\%) \text{ cm}^3 \text{ s}^{-1}$ $r \approx 0.3 \pm 0.2 \text{ cm}^3 \text{ s}^{-1}$

Experiment results

Table-1: Data and observations for **Manganese (IV) Oxide**

Initial conditions and controlled variables, with aggregate error		Raw data				Processed data			
		Volume of gas V (cm^3) $\pm$ 2.5 cm^3 instrument error	Time t (Seconds)			Average time t_μ (Seconds) with propagated error		Rate r (cm^3 s^{-1}) with propagated error	
			Trial 1	Trial 2	Trial average				
			$\pm$ 0.01 seconds instrument error		with aggregate error				
Average initial temperature across trials (°C)	21.1 ±0.2	10.0	34.11	30.62	32 ± 2				
Average mass of catalyst across trials (grams)	1.019 ±0.003	20.0	60.72	61.45	61.1 ± 0.4	47 ± 3		0.3 ± 0.2	
		30.0	96.78	93.29	95 ± 2	78 ± 2		0.29 ± 0.07	
Volume of H$_2$O$_2$ solution (dm^3)	200.0 ±0.6	40.0	133.92	140.63	137 ± 3	116 ± 5		0.24 ± 0.04	
Concentration of H$_2$O$_2$ solution (mol dm^{-3})	0.176 ±0.002	50.0	193.05	188.23	191 ± 2	164 ± 6		0.19 ± 0.03	
		60.0	260.47	256.87	259 ± 2	225 ± 4		0.15 ± 0.02	

Qualitative Observations

- Catalyst is a finely powdered black solid.
- Instant bubbling when the catalyst is released into the prepared H$_2$O$_2$ solution.
- Catalyst has poor solubility, mostly remaining solid in the solution as black suspensions.

Table-2: Data and observations for **Lead (II, IV) Oxide**

Initial conditions and controlled variables, with aggregate error		Raw data				Processed data			
		Volume of gas V (cm^3) $\pm$ 2.5 cm^3 instrument error	Time t (Seconds)			Average time t_μ (Seconds) with propagated error		Rate r (cm^3 s^{-1}) with propagated error	
			Trial 1	Trial 2	Trial average				
			$\pm$ 0.01 seconds instrument error		with aggregate error				
Average initial temperature across trials (°C)	21.2 ±0.1	10.0	5.82	6.25	6.0 ± 0.2				
Average mass of catalyst across trials (grams)	1.016 ±0.002	20.0	10.27	9.78	10.0 ± 0.3	8.0 ± 0.5		3 ± 1	
		30.0	15.61	13.42	15 ± 1	12 ± 1		1.9 ± 0.6	
Volume of H$_2$O$_2$ solution (dm^3)	200.0 ±0.6	40.0	21.52	22.63	22.1 ± 0.6	18 ± 2		1.7 ± 0.4	
Concentration of H$_2$O$_2$ solution (mol dm^{-3})	0.176 ±0.002	50.0	28.13	30.22	29 ± 1	26 ± 2		1.5 ± 0.3	
		60.0	35.09	37.56	36 ± 1	33 ± 2		1.4 ± 0.2	

Qualitative Observations

- Catalyst is a finely powdered, vividly orange solid.
- Instant rapid bubbling when the catalyst is released into the prepared H$_2$O$_2$ solution.
- Catalyst has poor solubility, mostly remaining solid in the solution as orange suspensions.

Table-3: Data and observations for **Potassium Iodide**, where the system is left **undisturbed**

Initial conditions and controlled variables, with aggregate error		Raw data				Processed data			
		Volume of gas V (cm^3) $\pm$ 2.5 cm^3 instrument error	Time t (Seconds)			Average time t_μ (Seconds) with aggregate error		Rate r (cm^3 s^{-1}) with propagated error	
			Trial 1	Trial 2	Trial average				
			$\pm$ 0.01 seconds instrument error		with propagated error				
Average initial temperature across trials (°C)	20.9 ±0.2	10.0	30.63	32.67	32 ± 1				
Average mass of catalyst across trials (grams)	1.043 ±0.003	20.0	65.85	70.51	68 ± 2	48 ± 3		0.3 ± 0.1	
		30.0	103.59	108.92	106 ± 3	85 ± 5		0.26 ± 0.07	
Volume of H$_2$O$_2$ solution (dm^3)	200.0 ±0.6	40.0	143.13	152.28	148 ± 5	123 ± 7		0.25 ± 0.05	
Concentration of H$_2$O$_2$ solution (mol dm^{-3})	0.176 ±0.002	50.0	185.73	191.56	189 ± 3	164 ± 8		0.23 ± 0.04	
		60.0	230.77	262.81	250 ± 20	210 ± 20		0.22 ± 0.04	

Qualitative Observations

- Catalyst is a powdered crystalline white solid.
- The catalyst is highly soluble, the solid catalyst dissolves instantly when released into the prepared H$_2$O$_2$ solution, followed by bubbling.
- The originally clear and colourless H$_2$O$_2$ solution turns into a clear brown colour after the catalyst is dissolved in it.
- The reactive system is sensitive to motion, where disturbing the system leads to observably more rapid bubbling.

Table-4: Data and Observations for **Copper (II) Oxide**

Initial conditions and controlled variables, with aggregate error

Average initial temperature across trials (°C)	20.8	±0.1
Average mass of catalyst across trials (grams)	0.968	±0.004
Volume of H_2O_2 solution (dm³)	200.0	±0.6
Concentration of H_2O_2 solution (mol dm⁻³)	0.176	±0.002

Raw data / Processed data

Volume of gas V (cm³) ± 2.5 cm³ instrument error	Trial 1 (± 0.01 seconds instrument error)	Trial 2	Trial average with aggregate error	Average time t_μ (Seconds) with propagated error	Rate r (cm³ s⁻¹) with propagated error
10.0	179.93	217.67	200 ± 20		
20.0	444.61	482.19	460 ± 20	330 ± 40	0.04 ± 0.02
30.0	836.17	1004.02	920 ± 80	690 ± 90	0.022 ± 0.007
40.0	1351.71	1572.23	1500 ± 100	1200 ± 200	0.018 ± 0.006
50.0	1942.11	2207.48	2100 ± 100	1800 ± 200	0.016 ± 0.004
60.0	2662.65	2905.25	2800 ± 100	2400 ± 300	0.014 ± 0.003

Qualitative Observations
- Catalyst is a finely powdered black solid.
- Instant rapid bubbling when the catalyst is released into the prepared H_2O_2 solution.
- Catalyst has poor solubility, mostly remaining solid in the solution as black suspensions.
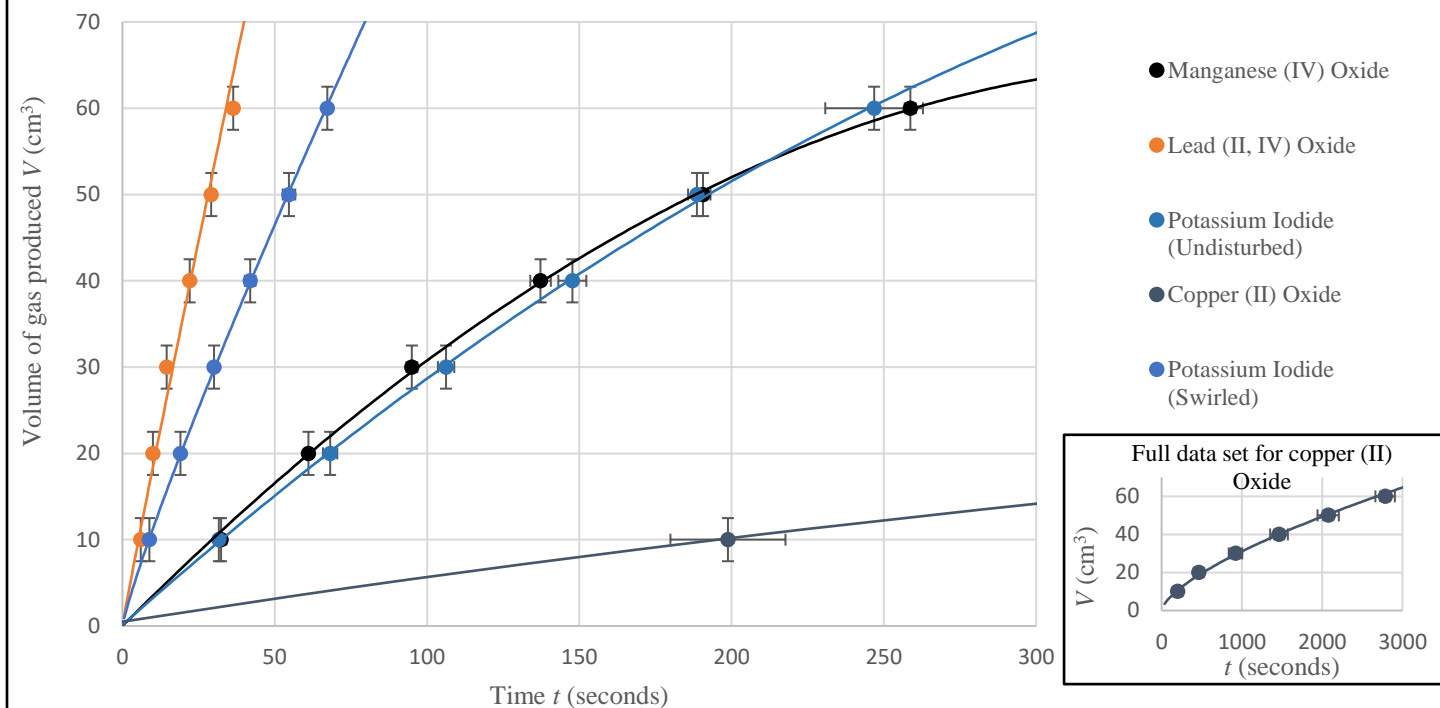
Table-5: Data and observations for **Iron (III) Oxide**

Initial conditions and controlled variables, with aggregate error

Average initial temperature across trials (°C)	21.3	±0.2
Average mass of catalyst across trials (grams)	0.942	± 0.002
Volume of H_2O_2 solution (dm³)	200.0	±0.6
Concentration of H_2O_2 solution (mol dm⁻³)	0.176	±0.002

Attempts to collect data and qualitative observations
- Catalyst is a finely powdered red solid.
- No visual evidence of any reaction occurring after the catalyst is released into the prepared H_2O_2 solution initially, but tiny bubbles begin to emerge after around 600 seconds.
- Catalyst has poor solubility, mostly remaining solid in the solution as red suspensions.
- The amount of gas produced by the system in 1800 seconds is not measureable by the gas syringe used, no data could be collected on the prepared H_2O_2 solution as a result.
- An attempt was made to use the undiluted 1 mol dm⁻³ H_2O_2 solution with the catalyst, but still yielded no measurable result after 1800 seconds, hence measurement for this catalyst is terminated.

Table-6: Data and observations for **Potassium Iodide**, where the system is constantly **swirled**

Initial conditions and controlled variables, with aggregate error

Average initial temperature across trials (°C)	21.1	± 0.1
Average mass of catalyst across trials (grams)	1.012	±
Volume of H_2O_2 solution (dm³)	200.0	±0.6
Concentration of H_2O_2 solution (mol dm⁻³)	0.176	±0.002

Raw data / Processed data

Volume of gas V (cm³) ± 2.5 cm³ instrument error	Trial 1 (± 0.01 seconds instrument error)	Trial 2	Trial average with aggregate error	Average time t_μ (Seconds) with aggregate error	Rate r (cm³ s⁻¹) with propagated error
10.0	9.15	8.47	8.8 ± 0.4		
20.0	19.34	18.76	19.1 ± 0.3	13.9 ± 0.8	1.0 ± 0.4
30.0	31.17	29.02	30 ± 1	25 ± 1	0.91 ± 0.2
40.0	43.68	40.21	42 ± 2	36 ± 3	0.8 ± 0.2
50.0	56.72	52.64	55 ± 2	48 ± 4	0.8 ± 0.2
60.0	67.33	67.22	67.28 ± 0.08	61 ± 2	0.8 ± 0.1

Qualitative Observations
- Catalyst is a powdered crystalline white solid.
- The catalyst is highly soluble, the solid catalyst dissolves instantly when released into the prepared H_2O_2 solution, followed by bubbling.
- The originally clear and colourless H_2O_2 solution turns into a clear brown colour after the catalyst is dissolved in it.
- Bubbling significantly intensifies after swirling begins.

Figure-5: Relationship between the volume of gas produced V and time t for each catalyst reacting with the prepared H_2O_2 solution

Legend:
- Manganese (IV) Oxide
- Lead (II, IV) Oxide
- Potassium Iodide (Undisturbed)
- Copper (II) Oxide
- Potassium Iodide (Swirled)

Inset: Full data set for copper (II) Oxide

Figure-6: Relationship between the calculated rate of reaction r of each catalyst reacting with the the prepared H_2O_2 solution and time t.

- Manganese (IV) Oxide $\quad r = 0.4287e^{-0.005t}$ $\quad R^2 = 0.9946$
- Lead (II, IV) Oxide $\quad r = 2.6184e^{-0.02t}$ $\quad R^2 = 0.8411$
- Potassium Iodide (Undisturbed) $\quad r = 0.3041e^{-0.002t}$ $\quad R^2 = 0.9952$
- Copper (II) Oxide $\quad r = 0.0344e^{-4E\text{-}04t}$ $\quad R^2 = 0.8101$
- Potassium Iodide (Swirled) $\quad r = 1.019e^{-0.005\,t}$ $\quad R^2 = 0.8995$
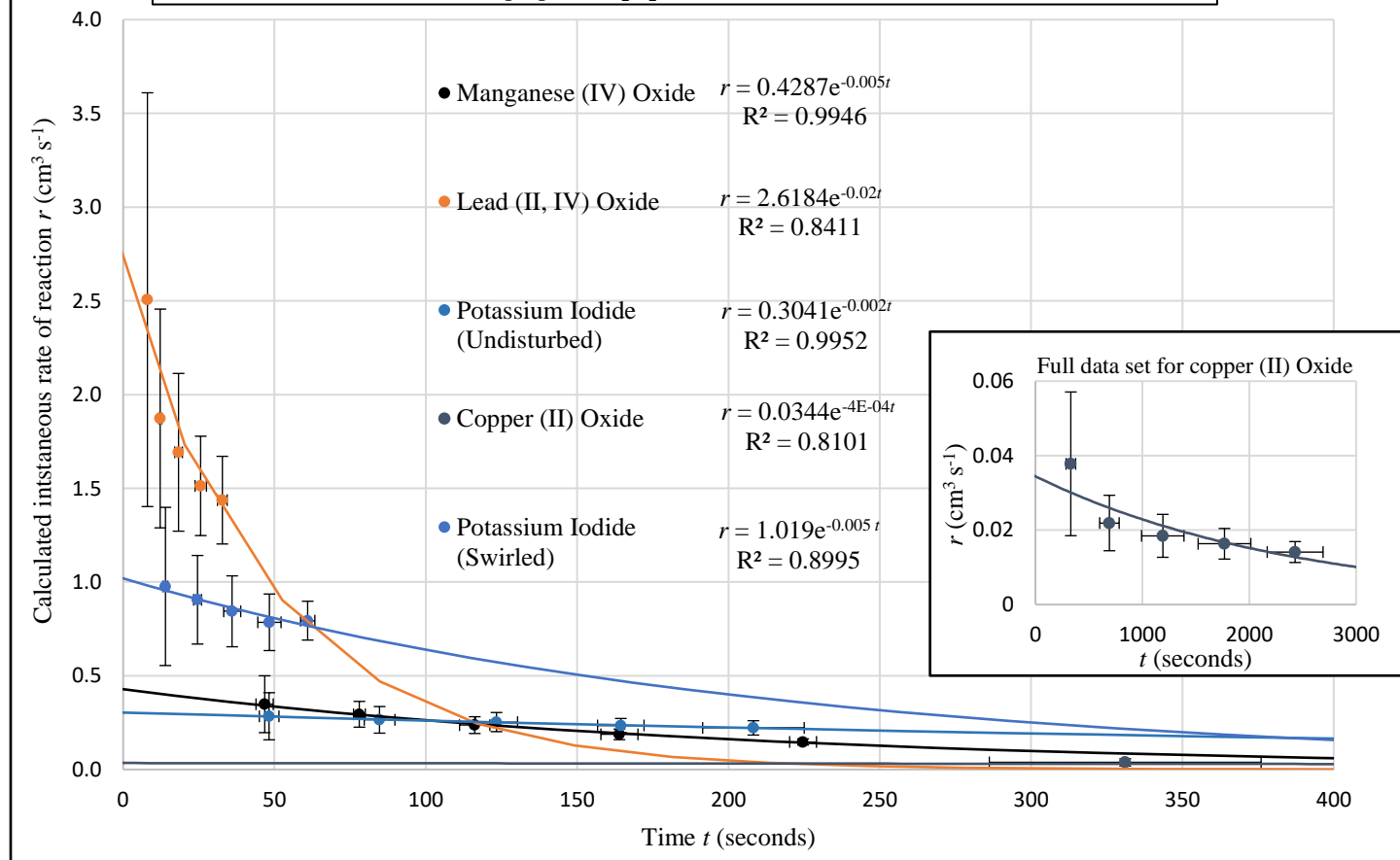
Inset: Full data set for copper (II) Oxide

Figure-6 shows that the relationship between the rates of reaction r with time t for each of the reacting system can be accurately modelled as an exponential decay, where exponential trend lines agree with all data points within the acceptable margin of error. The equation for exponential decay takes the format of:

$$y = a\,e^{-k\,x} \quad [21]$$

, where a and k and real constants. Following this equation, the value of the constant a will be the y-intercept of the trend line, hence the a value of each equation on *Figure-6* will represent the initial rate r_0 of each reacting system. All values of r_0 derived from *Figure-6* are tabulated on *Table-7*. The absolute uncertainty of the initial rate Δr_0 is evaluated by taking the average of the relative uncertainty of all calculated r values plotted for a specific reacting system and multiplied by the derived value of r_0, where:
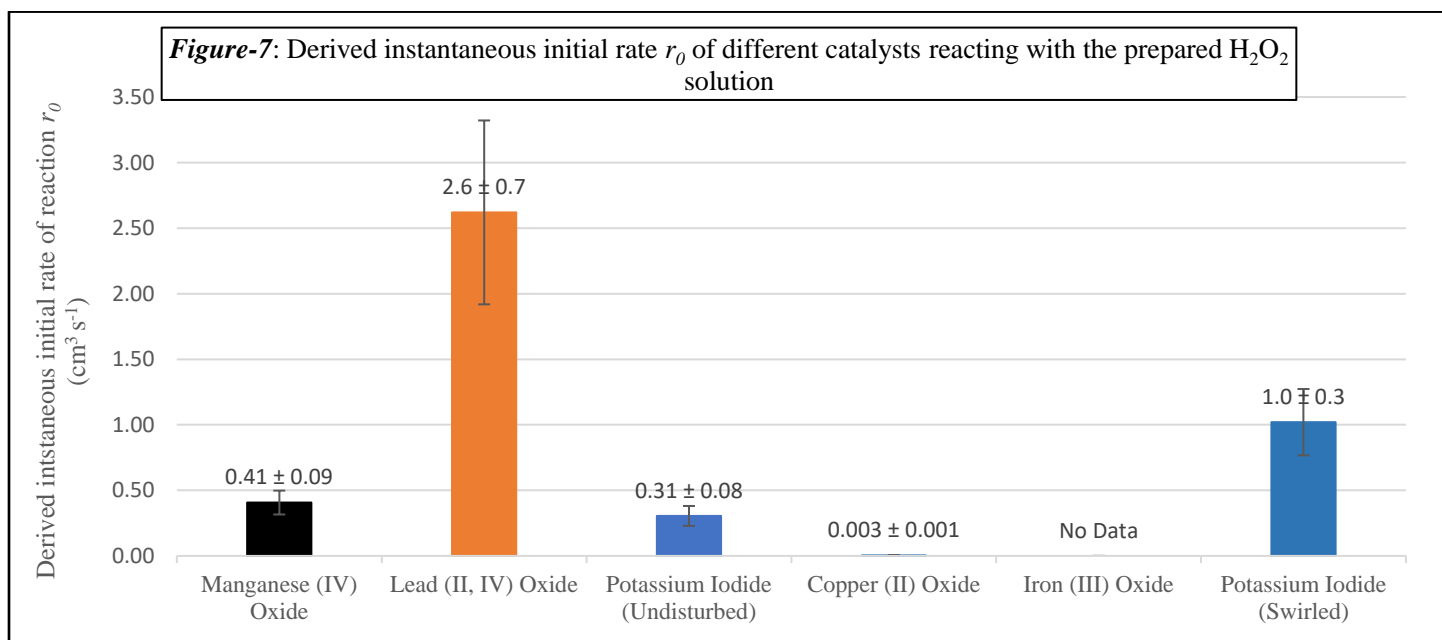
$$\Delta r_0 = \frac{\sum(\Delta r / r)}{n} \times r_0 \quad [22]$$

, n is the total number of r values involved in the calculation.

Sample calculation of uncertainty for r_0: *Table-7*, Lead (II, IV) Oxide, $r_0 = 2.6 \pm 0.7$ cm^3 s^{-1}	
All calculated r values for Lead (II, IV) Oxide, data from *Table-2*:	r_0 derived from *Figure-6*: 2.6 cm^3 s^{-1}

Rate r (cm^3 s^{-1}) with propagated error		
3	$\pm$	1
1.9	$\pm$	0.6
1.7	$\pm$	0.4
1.5	$\pm$	0.3
1.4	$\pm$	0.2

$$\Delta r_0 = \frac{\frac{1}{3} + \frac{0.6}{2.9} + \frac{0.4}{1.7} + \frac{0.3}{1.5} + \frac{0.2}{1.4}}{5} \times 2.6$$

$$\Delta r_0 \approx 0.7 \text{ cm}^3 \text{ s}^{-1}$$

$$r_0 = 2.6 \pm 0.7 \text{ cm}^3 \text{ s}^{-1}$$

Table-7: Derived instantaneous initial rate r_0 of different catalysts reacting with the prepared H_2O_2 solution, with the relevant properties of each catalyst attached.

Catalysts	Derived instantaneous initial rate r_0 (cm^3 s^{-1}), with average error
Manganese (IV) Oxide	0.41 $\pm$ 0.09
Lead (II, IV) Oxide	2.6 $\pm$ 0.7
Potassium Iodide (undisturbed)	0.31 $\pm$ 0.08
Copper (II) Oxide	0.003 $\pm$ 0.001
Iron (III) Oxide	No data / Too small to measure given the procedure followed
Potassium Iodide (swirled)	1.0 $\pm$ 0.3

Figure-7: Derived instantaneous initial rate r_0 of different catalysts reacting with the prepared H_2O_2 solution

Conclusion

The experiment suggests that lead (II, IV) oxide will most rapidly catalyze the decomposition of hydrogen peroxide per unit mass out of all five tested catalysts, where it has the highest derived instantaneous initial rate of 2.6 ± 0.7 cm^3 s^{-1}. Potassium iodide has the second highest initial instantaneous rate of 1.0 ± 0.3 cm^3 s^{-1} when swirled, but the rate drops to the second lowest at 0.31 ± 0.08 cm^3 s^{-1} when the system is left undisturbed. The margin of error for swirled potassium iodide does not overlap with the margin of error for lead (II, IV) oxide, which means that the results of this experiment definitively shows that lead (II, IV) oxide will most rapidly catalyze the decomposition of hydrogen peroxide out of the five tested catalysts.

An explanation for why lead (II, IV) oxide could most rapidly catalyze the decomposition of hydrogen peroxide is that lead (II, IV) oxide contains lead in both of its most common oxidation states, which allows the both reduction and oxidation of lead to happen simultaneously in one catalytic cycle as seen in the system of equations [15]. Since the forward and backward transition between lead (II) and lead (IV) will have the exact opposite energy profile, the presence of lead in both oxidation states in lead (II, IV) will cause hydrogen peroxide to more readily decompose through the reaction pathway involving the catalyst as each phase of the catalytic cycle will overall have a lower activation energy due to the oxidation and reduction of lead happening simultaneously.

The high solubility of potassium iodide gives it the highest contact surface area with the reactant out of all five tested catalysts, which explains why it can relatively more rapidly catalyze the decomposition of hydrogen peroxide per unit mass compared to the remaining three tested catalysts. The reason why the rate of hydrogen peroxide decomposition catalyzed by potassium iodide increases when the system is swirled is most likely the result of a systematic error, which that the hydrogen peroxide and potassium iodide solution might not be well-mixed during experiment. Under standard conditions, potassium iodide with a density of 3.12 g cm^{-3} *(NCBI)* is more than twice as dense as hydrogen peroxide at 1.44 g cm^{-3} *(NCBI)*, hence it is possible that the aqueous hydrogen peroxide and aqueous potassium iodide in the undisturbed system would be separated by a density gradient, which greatly reduces the contact surface area between the catalyst and the reactants and results in a lower rate of reaction. Measures that can be taken to eliminate this systematic error will be discussed under evaluation.

Out of the remaining three tested catalysts, manganese (IV) oxide will most rapidly catalyze the decomposition of hydrogen peroxide per unit mass, with an instantaneous initial rate of 0.41 ± 0.09 cm^3 s^{-1}, followed by copper (II) oxide with 0.003 ± 0.001 cm^3 s^{-1}. The decomposition of hydrogen peroxide catalyzed by iron (III) oxide was the least rapid out of all five tested catalysts, as the rate of reaction was too slow for any meaningful data to be collected from the reacting system given the procedure followed. An improvement to the method that enables meaningful data to be collected data for iron (III) oxide is detailed under evaluation.

Although no literature of similar experiments can be found through secondary sources, data from an experiment investigating how rapidly transition metal chlorides can catalyze the decomposition of hydrogen peroxide per unit mass was available. Since that transition metal ions are responsible for the catalysis of hydrogen peroxide decomposition, while the chloride and oxide ions are just spectators, one can expect similar results in terms of how rapidly a transition metal oxide and chloride can catalyze the decomposition of hydrogen peroxide when compared to the chlorides and oxides of other transition metals. Despite lead chloride and potassium iodide are not tested in the chloride experiment, the experiment concludes that manganese (II) chloride will most rapidly catalyze the decomposition of hydrogen peroxide per unit mass, followed by copper (II) chloride, while iron (II) chloride is will least rapidly catalyze the decomposition of hydrogen peroxide per unit mass compared to the other two transition metals *(Lee, et al.)*. This strongly agrees with the results of this investigation for manganese (IV) oxide, copper (II) oxide, and iron (III) oxide, which suggests that the method of this experiment is largely valid, while the data collected from this experiment is mostly accurate.

Evaluation

Limitations and validity of method

Limitation	Impact and evidence	Resolution
The dissolved potassium iodide solution is not well-mixed.	The rate of hydrogen peroxide decomposition increased when the system is swirled.	Stir all reacting systems with an magnetic stirrer at a constant rate of rotation in a container with a constant dimension when the catalyzed reaction begins.
The swirling of the system catalyzed by potassium iodide was done by hand.	The rate of swirling was not constant and not quantitatively controlled. This aspect of the experiment is not reproducible, hence leading to a loss of reliability.	
The 5 cm^3 volume graduation of the gas syringe used was too big.	No meaningful data could be collected for iron (III) oxide, while there is a loss of precision due to large instrument errors.	Use a more accurate gas syringe. Alternatively the experiment could also be done using gravimetric methods, where the change in mass is measured against time on an electronic balance.
Only two repeated trials of measurement are conducted for each reacting system.	The error bars for the initial rate of the system catalyzed by lead (II, IV) oxide is relatively large (2.6 ± 0.7 cm^3 s^{-1}), accounting for approximately 27% of the calculated value.	Conduct at least three trials of measurements for each reacting system, so that in concordant measurements can be rejected and re-measured, which will greatly improve the accuracy and reliability of this experiment.
Surface area of the transition metal oxide catalysts are not quantitatively controlled.	This aspect of the experiment is not reproducible, hence leading to a loss of reliability.	Instead of using powdered transition metal oxides, use transition metal chips that can be cut into regular shapes with calculable surface area.
The experiment takes place over two weeks, which is significantly longer than the 10-20 hour half-life of the spontaneous decomposition of hydrogen peroxide in open air under standard conditions. *(FMC Corporation)*	This means the loss of control over the controlled variable [H$_2$O$_2$], leading to inaccurate measurements. However, the impact of this should be minimal, as the container for hydrogen peroxide is always sealed air-tight after use to maintain internal pressure. Since the decomposition of hydrogen peroxide produces oxygen gas, an air-tight container with a stable internal pressure will sufficiently slow down the rate of decomposition.	To further maintain the constant concentration of hydrogen peroxide over time, the hydrogen peroxide used should not only be stored in an airtight container, but also placed in a fridge in the absence of ultra violet light to further slowdown the rate of spontaneous decomposition.

Despite various limitations stated above, the method of this experiment is quite easily reproducible given a similar set of apparatus and material. Since the experiment did provide sufficient data that enables a meaningful conclusion to be drawn, the method of this investigation is overall valid.

Scope for further investigation

The presence of different types of catalyst is not the only variable that can change the rate of hydrogen peroxide decomposition. Further investigations can explore the rate of hydrogen peroxide decomposition under different temperatures and pressures, or the relationship between the rate of hydrogen peroxide decomposition and the intensity of ultraviolet light that the system is exposed to.

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