- Briefly explain how a catalyst works.

Marks
2

Marks

- Given the following experimental data, find the rate law and the rate constant for the following reaction: **3**

$$NO(g) + NO_2(g) + O_2(g) \rightarrow N_2O_5(g)$$

Run	[NO(g)] / M	[NO_2(g)] / M	[O_2(g)] / M	Rate / M s^{-1}
1	0.10	0.10	0.10	2.1×10^{-2}
2	0.20	0.10	0.10	4.2×10^{-2}
3	0.20	0.30	0.20	1.26×10^{-1}
4	0.10	0.10	0.20	2.1×10^{-2}

Rate = k =

- The rate constant for a reaction is 5.0×10^{-3} s^{-1} at 215 °C and 1.2×10^{-1} s^{-1} at 452 °C. What is the activation energy of the reaction in kJ mol^{-1}? **3**

Answer:

What is the rate constant for this reaction at 100 °C?

Answer:

Marks 5

- The following data were obtained for the iodide-catalysed decomposition of hydrogen peroxide, H_2O_2.

Experiment	$[I^-]$(M)	$[H_2O_2]$ (M)	Initial rate(M s^{-1})
1	0.375	0	0
2	0.375	0.235	0.000324
3	0.375	0.470	0.000657
4	0.375	0.705	0.001024
5	0.375	0.940	0.001487
6	0	0.948	0
7	0.050	0.948	0.00045
8	0.100	0.948	0.00095
9	0.150	0.948	0.00140
10	0.200	0.948	0.00193

Determine the rate law from these data.

Use the data from Experiment 10 to calculate the rate constant for this reaction.

$k =$

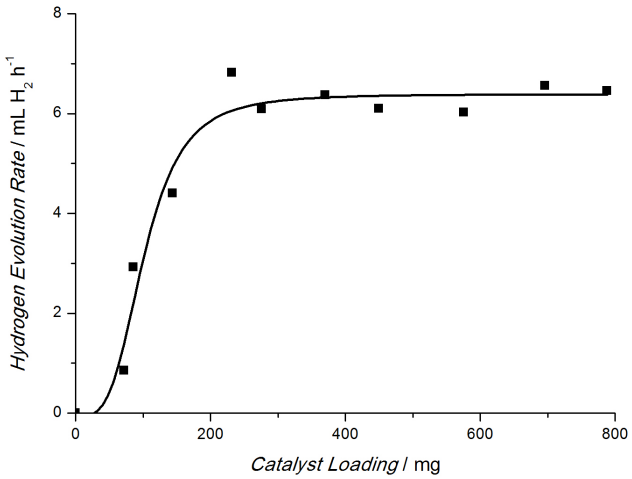
Iodide ion is used as a catalyst in this reaction. What is the role of a catalyst in a chemical reaction?

Marks
2

- When irradiated with visible light, CdS can catalyse the production of H_2 from water.

$$H_2O + \text{light} \xrightarrow{CdS} H_2 + \tfrac{1}{2}O_2$$

The rate of H_2 production from 80 mL of water at constant illumination varies with the amount of catalyst present (*i.e.* CdS loading) as shown below.



Why does the rate of H_2 production as a function of catalyst loading plateau?

THE REMAINDER OF THIS PAGE IS FOR ROUGH WORKING ONLY.

Marks 3

- Consider the reaction $A(g) + B(g) + C(g) \rightarrow D(g)$ for which the following data were obtained at 25 °C.

Experiment	Initial [A] (mol L^{-1})	Initial [B] (mol L^{-1})	Initial [C] (mol L^{-1})	Initial rate (mol L^{-1} s^{-1})
1	0.0500	0.0500	0.1000	6.25×10^{-3}
2	0.1000	0.0500	0.1000	1.25×10^{-2}
3	0.1000	0.1000	0.1000	5.00×10^{-2}
4	0.0500	0.0500	0.2000	6.25×10^{-3}

Write the rate law and calculate the value of the rate constant.

THE REMAINDER OF THIS PAGE IS FOR ROUGH WORKING ONLY.

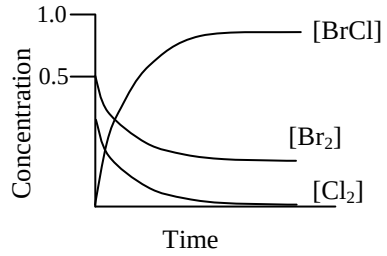
Marks 4

- In the reaction of Cl_2 with Br_2 in CCl_4 solution, BrCl forms according to the equation:

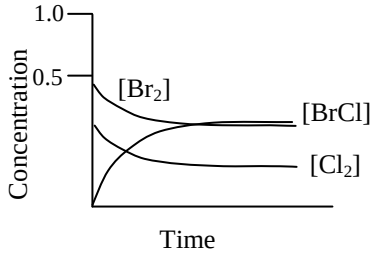
$$Br_2 + Cl_2 \rightarrow 2BrCl \qquad K_c = 2$$

With initial concentrations of $[Br_2] = 0.6$ M, $[Cl_2] = 0.4$ M and $[BrCl] = 0.0$ M, which of the following concentration versus time graphs represents this reaction? Explain qualitatively why you rejected each of the other three graphs.

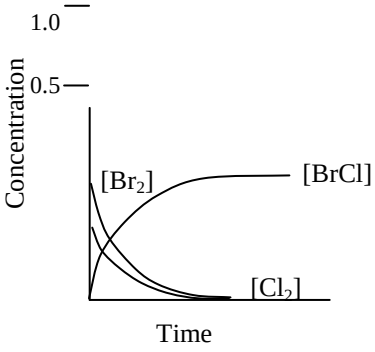
(a)



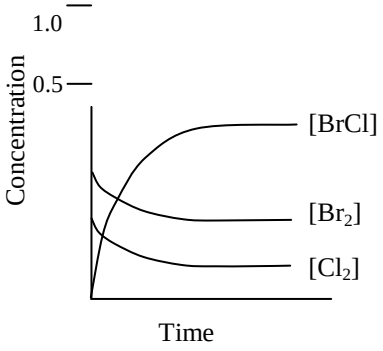
(b)



(c)



(d)



Marks 3

- Hydrogenation of NO to N_2 and water is a potential means of reducing smog-forming NO_x gases:

$$2H_2(g) + 2NO(g) \rightarrow N_2(g) + 2H_2O(g)$$

The initial rates of this reaction at constant temperature were determined at the following combination of initial pressures (P_0).

Experiment	P_0 H_2 (kPa)	P_0 NO (kPa)	Rate (kPa s^{-1})
1	53.3	40.0	0.137
2	53.3	20.3	0.033
3	38.5	53.3	0.213
4	19.6	53.3	0.105

What is the order of the reaction? Show all working.

Answer:

What is the value of the rate constant?

Answer:

Marks 6

- The following reaction is run from 4 different starting positions.

$$H_2SeO_3 + 6I^- + 4H^+ \rightarrow Se + 2I_3^- + 3H_2O$$

Experiment Number	Initial $[H_2SeO_3]$ (mol L^{-1})	Initial $[I^-]$ (mol L^{-1})	Initial $[H^+]$ (mol L^{-1})	Initial rate of increase of $[I_3^-]$ (mol L^{-1} s^{-1})
1	0.100	0.100	0.100	1.000
2	0.100	0.075	0.100	0.422
3	0.075	0.100	0.100	0.750
4	0.100	0.075	0.075	0.237

Determine the rate law for the reaction.

Rate law:

Calculate the value of the rate constant.

Answer:

Page Total:

Marks 3

- Peroxydisulfate and iodide ions react according to the following equation.

$$S_2O_8^{2-}(aq) + 3I^-(aq) \rightarrow 2SO_4^{2-}(aq) + I_3^-(aq)$$

The following rate data were collected at room temperature.

Experiment	$[S_2O_8^{2-}(aq)]_0$ (M)	$[I^-(aq)]_0$ (M)	Initial rate (mol L^{-1} s^{-1})
1	0.080	0.034	2.2×10^{-4}
2	0.080	0.017	1.1×10^{-4}
3	0.160	0.017	2.2×10^{-4}

Determine the rate law for the reaction.

Calculate the value of the rate constant at room temperature.

Answer:

THE REMAINDER OF THIS PAGE IS FOR ROUGH WORKING ONLY.

Marks 5

- Nitric oxide, a noxious pollutant, and hydrogen react to give nitrous oxide and water according to the following equation.

$$2NO(g) + H_2(g) \rightarrow N_2O(g) + H_2O(g)$$

The following rate data were collected at 225 ^oC.

Experiment	$[NO]_0$ (M)	$[H_2]_0$ (M)	Initial rate (d[NO]/dt, M s^{-1})
1	6.4×10^{-3}	2.2×10^{-3}	2.6×10^{-5}
2	1.3×10^{-2}	2.2×10^{-3}	1.0×10^{-4}
3	6.4×10^{-3}	4.4×10^{-3}	5.1×10^{-5}

Determine the rate law for the reaction.

Calculate the value of the rate constant at 225 ^oC.

Answer:

Calculate the rate of appearance of N_2O when $[NO] = [H_2] = 6.6 \times 10^{-3}$ M.

Answer:

Suggest a possible mechanism for the reaction based on the form of the rate law. Explain your answer.

Marks 2

- The following data were obtained for the reaction between gaseous nitric oxide and chlorine at –10 °C:

$$2NO(g) + Cl_2(g) \rightarrow 2NOCl(g)$$

Experiment Number	Initial P_{NO} (atm)	Initial P_{Cl_2} (atm)	Initial Reaction Rate (atm s^{-1})
1	2.16	2.16	0.065
2	2.16	4.32	0.130
3	4.32	4.32	0.518

Derive an expression for the rate law for this reaction and calculate the value of the rate constant.

Rate law:

Rate constant:

THIS QUESTION CONTINUES ON THE NEXT PAGE

Marks 4

The mechanism for this reaction has been postulated to be that below.

$2NO(g) \rightleftharpoons N_2O_2(g)$ fast

$N_2O_2(g) + Cl_2 \rightarrow 2NOCl(g)$ slow

Work out the rate law expected for this mechanism and hence show that it is consistent with the experimental rate law and the chemical equation.

The reaction is exothermic. Draw the potential energy *vs* reaction coordinate diagram for this mechanism, labelling all species that can be isolated.

Marks 4

- The following data were obtained for the reaction between gaseous nitric oxide and chlorine at –10 °C.

$$2NO(g) + Cl_2(g) \rightarrow 2NOCl(g)$$

Experiment Number	Initial [NO] (mol L^{-1})	Initial [Cl_2] (mol L^{-1})	Initial Reaction Rate (mol L^{-1} min^{-1})
1	0.10	0.10	0.18
2	0.10	0.20	0.36
3	0.20	0.20	1.44

Deduce the rate law for this reaction and calculate the value of the rate constant.

RATE LAW	RATE CONSTANT
Answer:	Answer:

Marks 3

- Hydrogenation of nitric oxide to nitrogen and water is a potential means of reducing smog-forming NO_x gases:

$$2NO(g) + 2H_2(g) \rightarrow N_2(g) + 2H_2O(g)$$

The initial rates of this reaction at constant temperature were determined at the following combination of initial pressures (P_0).

Run	P_0 (H_2) / kPa	P_0 (NO) / kPa	Rate / kPa s^{-1}
1	53.3	40.0	0.137
2	53.3	20.3	0.033
3	38.5	53.3	0.213
4	19.6	53.3	0.105

Derive an expression for the rate law for this reaction.

Answer:

Calculate the value of the rate constant.

Answer:

What is the order of the reaction?

- Briefly describe two factors that determine whether a collision between two molecules will lead to a chemical reaction.

1

Marks 4

- The following initial rate data have been obtained for the gas phase reaction of nitrogen dioxide, $NO_2(g)$, and ozone, $O_3(g)$, at 300 K.

$$2NO_2(g) + O_3(g) \rightarrow N_2O_5(g) + O_2(g)$$

$[NO_2(g)]$ M	$[O_3(g)]$ M	Rate M s^{-1}
0.65	0.80	2.61×10^4
1.10	0.80	4.40×10^4
1.10	1.60	8.80×10^4

What is the order of this reaction with respect to each reagent?

What is the rate constant of the reaction?

Answer: