Marks
1

- Explain the following terms or concepts.

Heterogeneous catalysis

Marks
4

- At a certain temperature the following data were collected for the reaction shown.

$$2ICl + H_2 \rightarrow I_2 + 2HCl$$

Experiment	Initial [ICl] (mol L^{-1})	Initial [H_2] (mol L^{-1})	Rate of formation of [I_2] (mol L^{-1} s^{-1})
1	0.10	0.10	0.0015
2	0.20	0.10	0.0030
3	0.10	0.050	0.00075

Determine the rate law for the reaction.

What is the value of the rate constant?

Answer:

THE REMAINDER OF THIS PAGE IS FOR ROUGH WORKING ONLY.

Marks 4

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

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

Experiment	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.

What is the value of the rate constant?

Answer:

- The rate constant of a polymer cross-linking reaction was established as a function of temperature. How can we demonstrate that the kinetics of this reaction follow Arrhenius behaviour? If it does follow Arrhenius behaviour, how can we derive the activation energy for the reaction and the pre-exponential factor *A*? **2**

Marks 4

- The major pollutants emitted by cars, NO(g), CO(g), NO_2(g) and CO_2(g), can react according to the following equation.

$$NO_2(g) + CO(g) \rightarrow NO(g) + CO_2(g)$$

The following rate data were collected at 215 °C.

Experiment	$[NO_2]_0$ (M)	$[CO]_0$ (M)	Initial rate (d[NO_2]/dt, M s^{-1})
1	0.263	0.826	1.44×10^{-5}
2	0.263	0.413	1.44×10^{-5}
3	0.526	0.413	5.76×10^{-5}

Determine the rate law for the reaction.

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

Marks
3

- Explain the following terms or concepts.

a) Lewis base

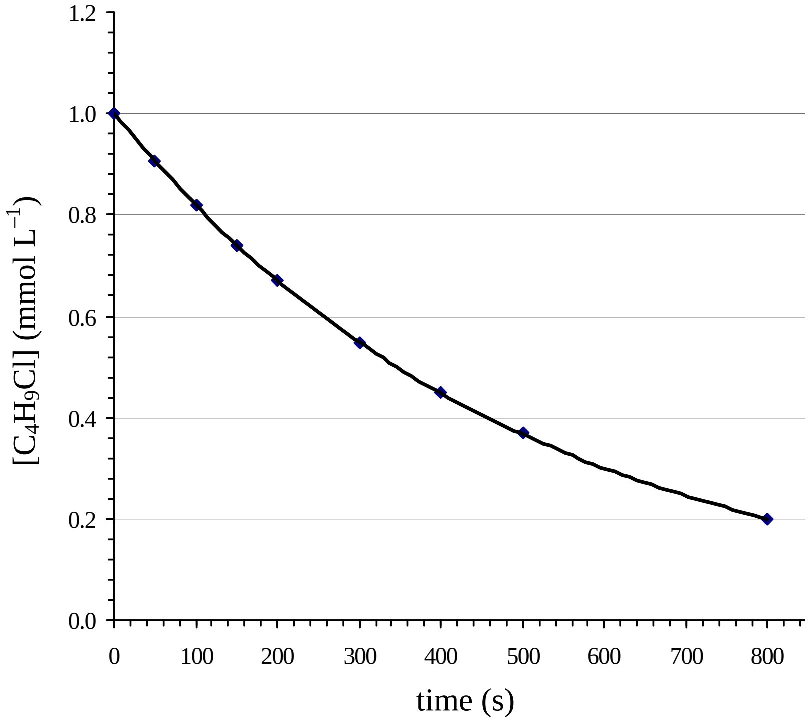
b) Le Châtelier’s principle

c) Heterogeneous catalysis

Marks
4

- The following chart shows the concentration of butyl chloride, C_4H_9Cl, as a function of time when it reacts with water according to the following equation:

$$C_4H_9Cl(aq) + H_2O(l) \rightarrow C_4H_9OH(aq) + H^+(aq) + Cl^-(aq)$$



Determine the instantaneous rate of reaction when $[C_4H_9Cl]$ = 1.0 mmol L^{-1}.

Answer:

Determine the instantaneous rate of reaction when $[C_4H_9Cl]$ = 0.5 mmol L^{-1}.

Answer:

THIS QUESTION CONTINUES ON THE NEXT PAGE

Page Total:

Marks
4

What is the order of the reaction with respect to C_4H_9Cl?

Answer:

How long would be required for the concentration of C_4H_9Cl to reach 0.01 mmol L^{-1}?

Answer:

Marks
6

- The disproportionation of hydrogen peroxide into oxygen and water has an enthalpy of reaction of –98.2 kJ mol^{-1} and an activation barrier of 75 kJ mol^{-1}. Iodide ions act as a catalyst for this reaction, with an activation barrier of 56 kJ mol^{-1}. The enzyme, catalase, is also a catalyst for this reaction, and this pathway has an activation barrier of 23 kJ mol^{-1}. Draw a labelled potential energy diagram for this process both without and with each of the catalysts.

Calculate the factor by which the reaction speeds up due to the presence of each of these two catalysts at a temperature of 37 °C. Assume that the pre-exponential Arrhenius factor remains constant.

- A proposed kinetic model for the reaction of NO(g) with Br_2(g) to form NOBr(g) is as follows. **3**

Step 1 $$NO(g) + NO(g) \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} N_2O_2(g)$$

Step 2 $$N_2O_2(g) + Br_2(g) \xrightarrow{k_2} 2NOBr(g)$$

If Step 2 is assumed to be very slow compared to the equilibrium of Step 1, derive the overall rate equation you would expect to see for this mechanism.

Marks

- Draw the potential energy diagram for an endothermic reaction. Indicate on the diagram the activation energy for both the forward and reverse reaction, and the enthalpy of reaction. 3

- Consider the reaction: $NO_2(g) + CO(g) \rightarrow NO(g) + CO_2(g)$ 4

  The experimentally determined rate equation is: Rate = $k[NO_2(g)]^2$

  Show the rate expression is consistent with the following mechanism:

Step 1	$2NO_2(g) \rightleftharpoons N_2O_4(g)$	fast
Step 2	$N_2O_4(g) \rightarrow NO(g) + NO_3(g)$	slow
Step 3	$NO_3(g) + CO(g) \rightarrow NO_2(g) + CO_2(g)$	fast

Marks
3

- Nitric oxide reacts with ozone according to the following equation.

$$NO(g) + O_3(g) \rightarrow NO_2(g) + O_2(g)$$

The following rate data were collected at a specified temperature.

Trial	Initial[NO] (M)	Initial $[O_3]$ (M)	Initial rate of reaction (M s^{-1})
1	2.1×10^{-6}	2.1×10^{-6}	1.6×10^{-5}
2	6.3×10^{-6}	2.1×10^{-6}	4.8×10^{-5}
3	6.3×10^{-6}	4.2×10^{-6}	9.6×10^{-5}

What is the experimental rate law for the reaction?

What is the value of the rate constant of this reaction?

Answer:

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 °C.

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 °C.

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
5

- The major pollutants NO(g), CO(g), NO_2(g) and CO_2(g), which are emitted by cars, can react according to the following equation.

$$NO_2(g) + CO(g) \rightarrow NO(g) + CO_2(g)$$

The following rate data were collected at 225 °C.

Experiment	$[NO_2]_0$ (M)	$[CO]_0$ (M)	Initial rate (d$[NO_2]$/dt, M s^{-1})
1	0.263	0.826	1.44×10^{-5}
2	0.263	0.413	1.44×10^{-5}
3	0.526	0.413	5.76×10^{-5}

Determine the rate law for the reaction.

Calculate the value of the rate constant at 225 °C.

Answer:

Calculate the rate of appearance of CO_2 when $[NO_2]$ = [CO] = 0.500 M.

Answer:

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

Marks
4

- Consider the results of the following set of experiments studying the rate of the reaction of nitric oxide with hydrogen at 1280 °C.

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

Experiment #	[NO] / M	$[H_2]$ / M	Initial Rate / M s^{-1}
1	5.0×10^{-3}	2.0×10^{-3}	1.3×10^{-5}
2	1.0×10^{-2}	2.0×10^{-3}	5.2×10^{-5}
3	1.0×10^{-2}	4.0×10^{-3}	1.0×10^{-4}

Write the rate law expression.

Rate =

Calculate the rate constant, *k*. Include units in your answer.

k =

What is the rate of the reaction when [NO] is 1.2×10^{-2} M and $[H_2]$ is 6.0×10^{-3} M?

Rate =

- What is the value of the equilibrium constant for the following reaction at 298 K? **2**

$$2Fe^{3+}(aq) + Sn(s) \rightleftharpoons Sn^{2+}(aq) + 2Fe^{2+}(aq)$$

The reduction half cell reactions and E^0 values are:

$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$ $E^0 = +0.77$ V
$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$ $E^0 = -0.14$ V

In the reaction, Sn is being oxidized and so the overall cell potential is:

$E^0 = (+0.77) - (-0.14) = +0.91$ V

The reaction involves 2 electrons so, using $E^0 = \frac{RT}{nF}\ln K$:

$$\ln K = E^0 \times \frac{nF}{RT} = (+0.91) \times \left(\frac{2 \times 96485}{8.314 \times 298}\right) = 70.9$$

$K = e^{70.9} = 6.05 \times 10^{30}$

Answer: **6.05×10^{30}**

Marks
4

- Consider the results of the following set of experiments studying the rate of the chemical reaction: $2A + B \rightarrow 3C + D$

Experiment #	initial [A] / M	initial [B] / M	Rate / M hr^{-1}
1	0.240	0.120	2.00
2	0.120	0.120	0.500
3	0.240	0.060	1.00

Write the rate law expression.

Rate =

Calculate the rate constant, *k*, with units.

k =

What is the rate of the reaction when [A] is 0.0140 M and [B] is 1.35 M?

Rate =