27 ~~26~~/30

Chemistry 205 Report

Rate of Iodination of Acetone

Name:

Date: 15-03-2013

Partner:

Purpose: 1. To study the kinetics of the reaction of iodine with acetone
2. To determine the order of the reaction with respect to each reactant, and then deduce the overall order.
3. To determine the rate and the rate constant for the reaction at a given temperature.
4. To determine the activation energy of the reaction.

Table 1. Reaction Rate Data @ T=

Mixture	Volume of acetone	Volume of HCl	Volume of H_2O	Volume of I_2	Time for 1st Run	Time for 2nd Run	Av.Time
I	10,00 ml	10.00 ml	20,00 ml	10.00	7,14	7,22	7,18 = 430,8 s.
II	20,00	10,00	10,00	10.00	4,03	4.00	4.015 = 240,9 s
III	10.00	20.00	10,00	10.00	4.15	3.58	3,865 = 231,9 s
IV	20.00	10.00	15.00	5.00	1;45	1,40	1,425 = 85,5 s.

Table 2. Determination of Reaction Orders

Mixture	Initial Conc. of Acetone	Initial Conc. of H^+	Initial Conc. of I_2	rate=$[I_2]_0$/av.time
I	0,800	0,200	1×10^{-3}	$2,32\times10^{-6}$
II	1,60	0,200	0,00100	$4,15\times10^{-6}$
III	0,800	0,4	0,00100	$4,31\times10^{-6}$
IV	1,6	0,2	5×10^{-4}	$5,85\times10^{-6}$

$M_1 \times V_1 = M_2 \times V_2$

$M_1 \Rightarrow \frac{M_2 \times V_2}{V_1} = \frac{4\times10}{50}$

$= \frac{1\times10^{-3}}{\text{Av.Time}}$

Rate Law of the Iodination of Acetone.

rate = k [acetone]^m [I_2]^n [H^+]^p

Sample calculation!

rate I = $k(0,8)^m(1\times10^{-3})^n \times (0,2)^p = 2,32\times10^{-6}$

rate II = $k(1,6)^m(1\times10^{-3})^n(0,2)^p = 4,15\times10^{-6}$

rate III = $k(0,8)^m(1\times10^{-3})^n(0,4)^p = 4,31\times10^{-6}$ Ms^{-1}

rate IV = $k(1,6)^m(5\times10^{-4})^n(0,2)^p = 5,85\times10^{-6}$ unit?

rate II / rate I = $\frac{4,15\times10^{-6}}{2,32\times10^{-6}} = 2 \Leftrightarrow 2^m = 2$ $\boxed{m = 1.}$

rate III / rate I = $\frac{4,31\times10^{-6}}{2,32\times10^{-6}} = 2 \Leftrightarrow 2^p = 2$ $\boxed{p = 1}$

rate IV / rate $\cancel{II}$ = $\frac{5,85 \times 10^{-6}}{4,15 \times 10^{-6}} = 1 \Leftrightarrow 2^m = 1$ $\boxed{n = 0}$

Table 3. Determination of the Rate Constant k

Mixture	I	II	III	IV	Average
k	$1,45 \times 10^{-5}$	$1,30 \times 10^{-5}$	$1,35 \times 10^{-5}$	$1,8 \times 10^{-5}$	$1,475 \times 10^{-5}$

Sample Calculation: rate I $= k(0,8)^m (1\times10^{-3})^n \times (0,2)^p = 2,32 \times 10^{-6}$

$$k = \frac{2,32 \times 10^{-6}}{(0,8)(0,2)} = 1,45 \times 10^{-5} / M \cdot s^{-1}$$

$\frac{1}{2}$

Table 4. Determination of the Activation Energy

Mixture	Time (minutes)	Temp. (°C)	Temp. (K)	1/T (K^{-1})	Rate	k'	ln k'
II	1,35 = 81s	40	313	$3,19 \times 10^{-3}$	$1,23 \times 10^{-5}$	$3,8 \times 10^{-5}$	−10,17
II	2,48 = 148,8	30	303	$3,30 \times 10^{-3}$	$6,7 \times 10^{-6}$	$2,1 \times 10^{-5}$	−10,77
II	6,50 = 390	20	293	$3,41 \times 10^{-3}$	$2,56 \times 10^{-6}$	8×10^{-6}	−11,7
II	18,5 = 1110	10	283	$3,53 \times 10^{-3}$	9×10^{-7}	$2,8 \times 10^{-6}$	−12,78
II	4,015 = 240,9	26	299	$3,34 \times 10^{-3}$	$4,15 \times 10^{-6}$	$1,29 \times 10^{-5}$	−11,25

Sample calculation: °C → K (+273) Rate $= \frac{[I_2]_0}{\Delta t} = \frac{1\times10^{-3}}{81} = 1,23 \times 10^{-5}$

$\frac{1}{T} = \frac{1}{313} = 3,19 \times 10^{-3}$

$k' = \frac{Rate}{(1,6)(0,2)} = \frac{1,23 \times 10^{-5}}{1,6 \times 0,2} = 3,8 \times 10^{-5}$

Attach the graph with the report!

Calculate (show your work):

Slope = $\frac{-12,78-(-10,17)}{3,53\times10^{-3} - 3,19\times10^{-3}} = \frac{-2,61}{3,4\times10^{-4}} = -7676$

Ea =

slope $= \frac{-Ea}{R}$

Ea = slope × 8,314

= 7676 × 8,314 = 63 818 J

Questions

1. Consider the following reaction:

   $$3A + 2B \rightarrow C + 3D$$

   a) Write the rate law of the reaction: given that its first order in A and second order in B.

   $rate = k[A][B]^2$

   b) Express the rate of the reaction as a function of the rate of disappearance/formation of each species participating in the reaction.

   $$rate = -\frac{1}{3}\frac{\Delta[A]}{\Delta t} = -\frac{1}{2}\frac{\Delta[B]}{\Delta t} = \frac{\Delta[C]}{\Delta t} = \frac{1}{3}\frac{\Delta[D]}{\Delta t}$$

2. What volumes of the following stock solutions are necessary to prepare 50.0 mL of the following reaction mixture: 0.800 M acetone; 0.200 M HCl; 0.00100 M iodine.

   n before = n after dilution dilution

   Stock solutions:

   4.00 M acetone: $V_{acetone} = \frac{0{,}8 \times 50}{4} = 10{,}00\ ml$

   1.00 M HCl: $V_{HCl} = \frac{0{,}2 \times 50}{1} = 10{,}00\ ml$

   0.00500 M I_2: $V_{I2} = \frac{0{,}001 \times 50}{0{,}005} = 10{,}00\ ml$

   Distilled water: $V_{water} = 20{,}00\ ml$.

3. The following data was collected for the reaction between hydrogen and nitric oxide at 700 °C:

   $$2H_{2(g)} + 2NO_{(g)} \rightarrow 2H_2O_{(g)} + N_{2(g)}$$

Experiment	$[H_2]$	[NO]	Initial rate (M/s)
1	0.010	0.025	2.4×10^{-6}
2	0.0050	0.025	1.2×10^{-6}
3	0.010	0.0125	0.60×10^{-6}

   a) Determine the rate law of the reaction

   b) What is the order of the reaction?

   c) Calculate the rate constant.

a) $rate\ law = k[H_2]^x[NO]^y$

b) $rate\ I = k[2H_2]^x[NO]^y = 2{,}4 \times 10^{-6}$

$rate\ II = k[H_2]^x[NO]^y = 1{,}2 \times 10^{-6}$

$$\frac{rate\ I}{rate\ II} = \frac{k[2H]^x[NO]^y}{k[H_2]^x[NO]^y} = \frac{[2H]^x}{[H]^x} = 2^x = \frac{2{,}4 \times 10^{-6}}{1{,}2 \times 10^{-6}} = 2.$$

$\Leftrightarrow 2^x = 2$

$\Leftrightarrow \boxed{x = 1}$

$rate\ III = k[2H_2]^x[\frac{1}{2}NO]^y = 0{,}60 \times 10^{-6}$

$$\frac{rate\ III}{rate\ I} = \frac{k[2H_2]^x[\frac{1}{2}NO]^y}{k[2H]^x[NO]^y} = \left(\frac{1}{2}\right)^y = \frac{0{,}60 \times 10^{-6}}{2{,}4 \times 10^{-6}} = \frac{1}{4} \longrightarrow$$