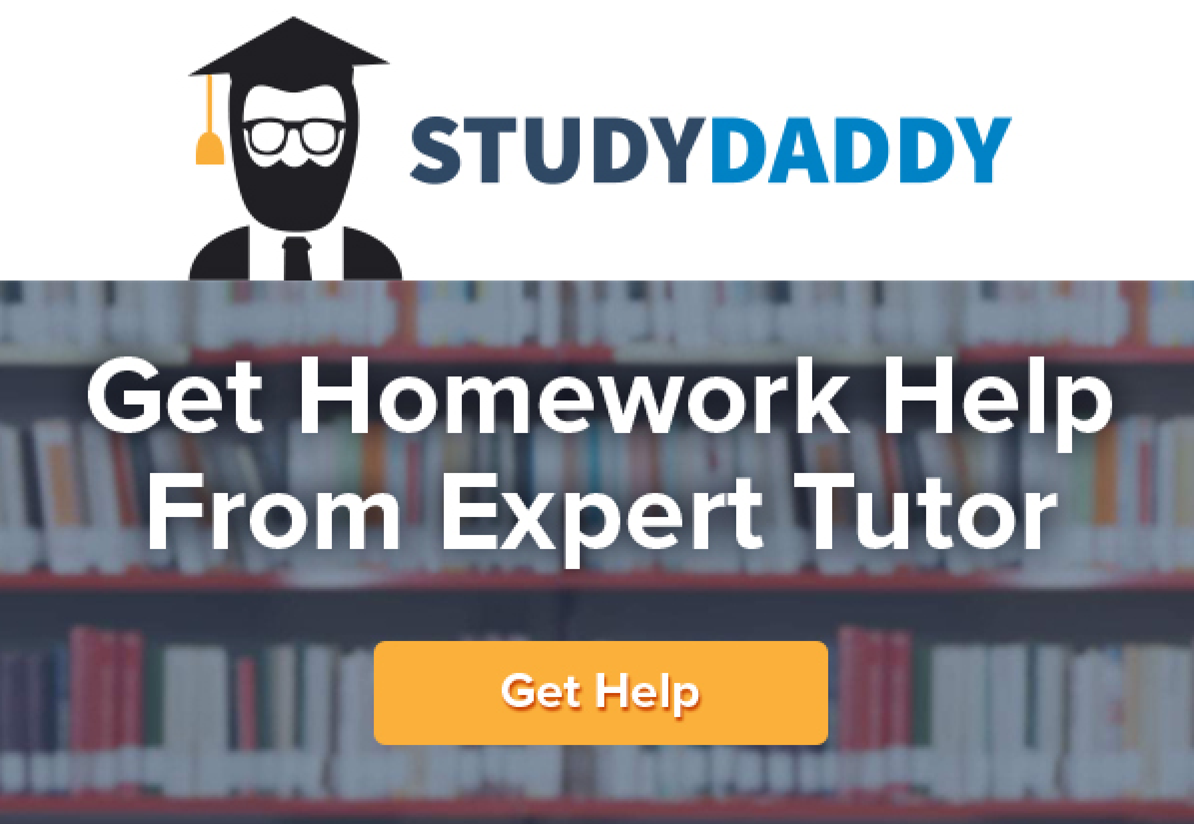
STUDYDADDY
Get Homework Help
From Expert Tutor
Get Help

Name . . . . TA . . . . .Lab Section . . . . Date . . . . . .

EXPERIMENT

Chemistry 0330 Laboratory Report

How Do Certain Factors Affect the Rate of a Chemical Reaction?

1. Present all the data for runs 1–6 in the table below. The rate ratio called for in the table is the inverse of the ratio of reaction times. The rate ratios will be used to find reaction orders and activation energies. Subscript i refers to the individual reactions (do not fill in the shaded areas). Δt_{1av} is the average Δt for mixture 1a and 1b.

temp reactant concent.

	1a	1b	2	3	4	5	6
T (°C)	21.2°C	21.2°C	21.4°C	21.5°C	22.0°C	34.3°C	22.4°C
Δt (sec)	107	109	242	125	40	41	51
Δt_{1av}	108						
$Rate_i/rate_{1av}$ $=\Delta t_{1av}/\Delta t_i$			.45	.86	2.7	2.63	2.12

2. Calculate the concentration of each reactant in reaction mixtures 1a and 1b, and the concentration in the reaction mixture in which it differs (do not fill in the shaded areas).

consumed

add volume

Reagent	Mixture #	Stock Conc (M)	Vol Taken (mL)	Final Vol (mL)	Final Conc (M)
$S_2O_3^{2-}$	1-6	0.00625	1.50	10	9.38×10^{-4}
H_2O_2	1	1.0	0.50	10	.05
H_2O_2	6	1.0	1.00	10	.1
KI	1	0.05	1.50	10	.0075
KI	2	0.05	0.75	10	.00375
$CH_3CO_2^-$	1	0.0375	1.50	10	.005625
CH_3CO_2H	1	0.0375	1.50	10	.005625
CH_3CO_2H	*Added to 1 to make 3*	0.3	3.00	10	.09
CH_3CO_2H	3			10	.095625
Mo(VI)	4	0.001	0.50	10	5×10^{-5}

3. The dissociation constant of acetic acid (CH_3CO_2H) is 1.8×10^{-5}.

$C_{stock}V_{stock} = C_{final}V_{total}$

a. What is the [H+] in mixture 1? Show calculations below.

$$K_a = \frac{[H^+][Ac^-]}{[HAc]}$$

$$1.8\times10^{-5} = \frac{[H^+](.005625)}{(.005625)} = \frac{[H^+][CH_3CO_2^-]}{[CH_3CO_2H]}$$

$$1.1\times10^{-5} = [H^+]$$

$[H^+]$ = 1.8×10^{-5} M

b. What is the [H+] in mixture 3?[1] Show calculations below.

$$1.8\times10^{-5} = \frac{[H^+](.005625)}{(.095625)} = \frac{[H^+][CH_3CO_2^-]}{[CH_3CO_2H]}$$

$$[H^+] = 3.06\times10^{-4}$$

$[H^+]$ = 3.06×10^{-4}

4. Using the rate equation, rate $=k[H_2O_2]^a [I^-]^b [H^+]^c$, the reaction orders a, b, and c can be obtained from the rates at two different concentrations of each reactant. (Refer to the ratio of rates in question 1.) Answers should be rounded to the nearest integer. Show your calculations.

a. What is the reaction order a in H_2O_2?

$$a = \frac{\log(2.12)}{\log\left(\frac{.1}{.05}\right)} = 1.08 \approx 1$$

1

b. What is the reaction order b in I^-?

$$b = \frac{\log(.45)}{\log\left(\frac{.00375}{.0075}\right)} = 1.15 \approx 1$$

1

c. What is the reaction order c in H^+? (Refer to [H+] in question 3)

$$c = \frac{\log(.86)}{\log\left(\frac{3.06\times10^{-4}}{1.8\times10^{-5}}\right)} = -.05 \approx 0$$

0

d. What is the overall reaction order?

2, 2nd order

$$1+1+0=2$$

e. Write the rate law for the reaction of H_2O_2 with I^- under these conditions.

$$\text{rate} = k[H_2O_2][I^-]$$

5. Rate (mol $L^{-1}s^{-1}$) = $-\frac{d[H_2O_2]}{dt}$ and $-\frac{d[H_2O_2]}{dt} \cong -\frac{\Delta[H_2O_2]}{\Delta t} = -\frac{1}{2}\frac{\Delta[S_2O_3^{2-}]}{\Delta t}$

a. Knowing that all the thiosulfate is consumed at the end point when the starch-I_2 complex first appears, calculate

---

[1] For reaction 3, $[CH_3CO_2^-]$ remains the same as for reaction 1, but $[CH_3CO_2H]$ includes both the CH_3CO_2H present in reaction 1 plus the extra CH_3CO_2H added.

the rate in mol $L^{-1}s^{-1}$ for mixture 1. (Refer to the concentrations listed in question 2.) Show calculations.

$$-\frac{1}{2}\frac{\Delta[S_2O_3^{2-}]_{mix1}}{\Delta t_{lav}} = \frac{1}{2}\,\frac{9.38\times10^{-4}\,mol/L}{108\,sec} = 4.34\times10^{-6}\,\frac{mol}{L\cdot s}$$

Rate $4.34\times10^{-6}\,mol\cdot L^{-1}\cdot s^{-1}$

b. From the rate law, compute the rate constant k for the conditions in reaction 1. Be sure to include units. (Refer to the concentrations listed in question 2.) Show calculations.

$rate = k[H_2O_2][I^-]$

$4.34\times10^{-6}\,mol\cdot L^{-1}\cdot s^{-1} = k(.05\,mol\,L^{-1})(.0075\,mol\cdot L^{-1})$

$k = .0012\,mol^{-1}\cdot L\cdot s^{-1}$

k $.01157\,L\cdot mol^{-1}\,s^{-1}$

c. Compare your value of k with the reported value of 0.0115 L $mol^{-1}s^{-1}$ at 25°C. Calculate the percent error. Show calculations.

$\%\ error = \frac{.01157 - .0115}{.0115}\times100 = .61\%$

% error $.61\%$

6. By what factor did the catalyst in mixture 4 enhance the reaction rate? (Refer to rate ratio in question 1.) Briefly explain the effect of the catalyst on the rate of the reaction.

2.7, a catalyst lowers the activation energy required for a reaction to proceed w/o the catalyst itself getting consumed. as a result, the rate of reaction speeds up (in this case the catalyst sped the reaction up by a factor of 2.7).

7. From the ratio of reaction rates for mixtures 5 and 1 in question 1, calculate the energy of activation E_a in kJ mol^{-1}.

$T_1 = 21.2 + 273 = 294.2\,K$

$T_5 = 34.3 + 273 = 307.3\,K$

$\Delta \ln k = \ln(2.63) = .97$

$\ln k_5 - \ln k_1 = \Delta \ln k = -\frac{E_a}{R}\left(\frac{1}{T_5} - \frac{1}{T_1}\right)$

$.97 = \frac{-E_a}{8.314\,J/mol\cdot K}\left(\frac{1}{307.3K} - \frac{1}{294.2K}\right)$

$E_a = 55656\,J/mol \cdot \frac{1\,kJ}{1000\,J} = 55.7\,kJ/mol$

E_a $55.7\,kJ/mol$

8. Is the rate law experimentally determined in 4e consistent with the one-step mechanism shown below?

$$H_2O_2 + 2\,I^- + 2\,H^+ \rightarrow I_2 + 2\,H_2O$$

Explain why or why not.

no, the experimentally determined rate law from 4e is not consistent w/ the one step mechanism. in this given mechanism, since it's single step, one can determine it's rate law from the stoichiometric coefficients of the reactants (H_2O_2, I^- + H^+) → $rate = k[H_2O_2][I^-]^2[H^+]^2$ which doesn't equal the rate law from 4e : $rate = k[H_2O_2][I^-]$

9. A possible mechanism is given by

$$H_2O_2 + I^- \rightarrow HOI + OH^-$$
$$H^+ + OH^- \rightleftharpoons H_2O$$
$$HOI + H^+ + I^- \rightarrow I_2 + H_2O$$

Show that this mechanism gives the correct stoichiometric reaction. If this mechanism is correct, which step is rate determining according to your experimental rate law? Explain your reasoning.

$H_2O_2 + I^- \rightarrow \cancel{HOI} + \cancel{OH^-}$

$H^+ + \cancel{OH^-} \leftrightarrow H_2O$

$\cancel{HOI} + H^+ + I^- \rightarrow I_2 + H_2O$

overall stoichiometric reaction

$\rightarrow 2H_2O_2 + 2H^+ + 2I^- \rightarrow 2H_2O$

the rate law for a stoichiometric reaction is given by the slowest, elementary reaction ~~step~~ according to my experimental rate law, the 1st elementary reaction step is consistent with it as determined by the stoich coefficients of the reactants the reactants consumed in the 1st step (both stoich coefficient for H_2O_2 + I^- equal 1, which is consistent w/ the orders i calculated for both those molecules (1st order) in my experimental rate law: $rate = k[H_2O_2][I^-]$

10. Given the following errors, briefly explain the effect on the given component of the reaction.

a. Not enough starch was added to the reaction mixture and only a slight blue color appeared. What effect would this have on the order of the reactants?

this would have no effect on the order of the reactants because it serves as an indicator of when the I_2 formed quickly reacts w/ the starch, create blue-color complex observable to us. the order of reactants depends on the start of the reaction and the exact time the solution starts to change color. the complete reaction between I_2 formed + starch, forming a deep blue complex doesn't affect this time and doesn't influence reaction rate + the order of reactants involved.

b. The acetic acid concentration was not correctly stated on the label. What effect would this have on the order of the acetic acid?

c. The concentration of the sodium thiosulfate was twice as concentrated as intended. What effect would this have on the time of each reaction?

the time of each reaction would be twice as long because rate of the reaction is determined by $\frac{1}{2}\Delta[S_2O_3^{2-}]/\Delta t$. since the concentration of S_2O_3 doesn't affect the actual rate of reaction (invoved in faster reaction) and reaction rate depends on the amt of I_2 produced in primary reaction (unchanged so same rate). to get the same rate as before, when S_2O_3 is doubled, time will be doubled.

d. The temperature in the laboratory was 10°C higher than the typical room temperature. What effect would this have on the time of each reaction?

the time of each reaction would decrease because an increase in temp, increases the heat absorbed by particles within the reaction. consequently, the average kinetic energy of the particles increases and greater proportion of them have the minimum amt of energy to collide w/ each other + move faster, which increases the rate of reaction and decreases the time for the reaction to succesfully proceed.