Science Notebook

40S Chemistry

Reaction Rates

Chemistry4 0S **KineticsAssignment#1-Key**

1. a) $\text{rate} = -\frac{\Delta A}{\Delta t} = -\frac{12.0g - 25.0g}{5.0min - 0.0min} = -\frac{-13.0g}{5.0min} = 2.6 g/min$

   b) $\text{rate} = -\frac{\Delta A}{\Delta t} = -\frac{13.0g - 17.0g}{4.0min - 2.0min} = -\frac{-4.0g}{2.0min} = 2.0 g/min$

2. a) $\text{rate} = -\frac{\Delta[CH_3CHO]}{\Delta t} = -\frac{0.00586 - 0.00667 mol/L}{105s - 42s}$

   $= -\frac{-8.10\times10^{-4} mol/L}{63s} = 1.29\times10^{-5} mol/L\cdot s$

   b) $\text{rate} = -\frac{\Delta[CH_3CHO]}{\Delta t} = -\frac{0.00342 - 0.00505 mol/L}{480s - 190s}$

   $= -\frac{-1.63\times10^{-3} mol/L}{290s} = 5.62\times10^{-6} mol/L\cdot s$

   c) since $\text{rate} = -\frac{\Delta[CH_3CHO]}{\Delta t} = \frac{\Delta[CH_4]}{\Delta t} = \frac{\Delta[CO]}{\Delta t}$, from the stoichiometry the rate of production of CH_4 and CO is 0.10 mol/Ls.

3.

Time (min)	Mass of beaker and contents (g)	Mass loss (CO_2 produced) (g)	
0.0	200.00	0.00	
1.0	199.40	0.60	(200.00–199.40)
2.0	199.00	1.00	(200.00–199.00)
3.0	198.65	1.35	(200.00–198.65)
4.0	198.35	1.65	(200.00–198.35)
5.0	198.10	1.90	(200.00–198.10)
6.0	197.90	2.10	(200.00–197.90)
7.0	197.75	2.25	(200.00–197.75)
8.0	197.65	2.35	(200.00–197.65)
9.0	197.57	2.43	(200.00–197.57)
10.0	197.52	2.48	(200.00–197.52)

b) $\text{rate} = \frac{\Delta mass CO_2}{\Delta t} = \frac{2.48g - 0.0g}{10.0 - 0.0min} = 0.248 g CO_2/min$

c) i) $\text{rate} = \dfrac{\Delta \text{mass } CO_2}{\Delta t} = \dfrac{1.35\text{ g} - 0.0\text{ g}}{3.0 - 0.0\text{ min}} = 0.450\text{ g } CO_2/\text{min}$

ii) $\text{rate} = \dfrac{\Delta \text{mass } CO_2}{\Delta t} = \dfrac{2.48\text{ g} - 2.25\text{ g}}{10.0 - 7.0\text{ min}} = \dfrac{0.23\text{ g}}{3.0\text{ min}} = 0.077\text{ g } CO_2/\text{min}$

4. From the equation coefficients,

$$(0.20\text{ mol/L}\cdot\text{s } N_2O_5)\left(\frac{4\text{ mole } NO_2}{2\text{ mole } N_2O_5}\right) = 0.40\text{ mol/L s } NO_2$$

$$(0.20\text{ mol/L}\cdot\text{s } N_2O_5)\left(\frac{1\text{ mole } O_2}{2\text{ mole } N_2O_5}\right) = 0.10\text{ mol/L s } O_2$$

5. a) $(0.090\text{ mol/L}\cdot\text{s } NH_3)\left(\dfrac{5\text{ mole } O_2}{4\text{ mole } NH_3}\right) = 0.1125\text{ mol/L s } O_2 \approx 0.11\text{ mol/L s } O_2$

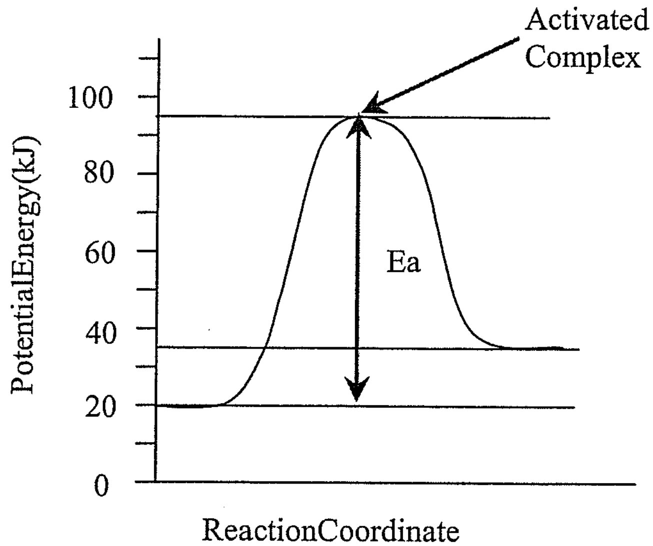
b) $(0.090\text{ mol/L}\cdot\text{s } NH_3)\left(\dfrac{6\text{ mole } H_2O}{4\text{ mole } NH_3}\right) = 0.135\text{ mol/L s } H_2O$

c) $(0.090\text{ mol/L}\cdot\text{s } NH_3)\left(\dfrac{4\text{ mole } NO}{4\text{ mole } NH_3}\right) = 0.090\text{ mol/L s } NO$

Chemistry 40S

Kinetics Assignment #2 – Key

1.

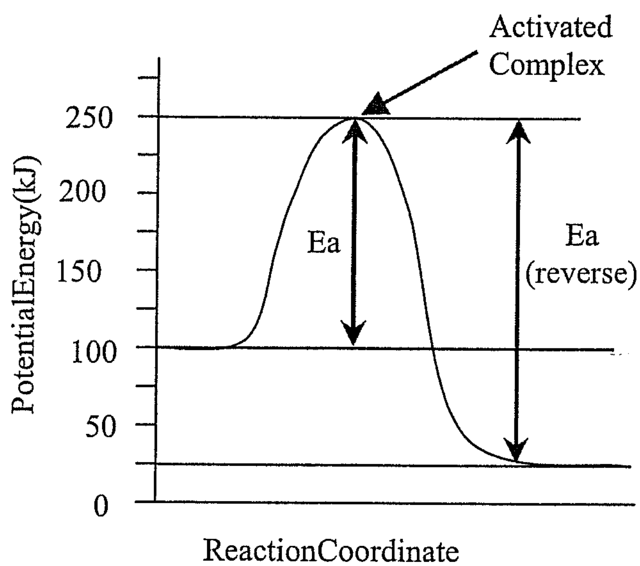


a) $E_a = 95\text{ kJ} - 20\text{ kJ}$
$= 75\text{ kJ}$

b) $\Delta H = H_{product} - H_{reactant}$
$= 35\text{ kJ} - 20\text{ kJ}$
$= 15\text{ kJ}$

c) Since ΔH is positive the reaction is endothermic.

2.



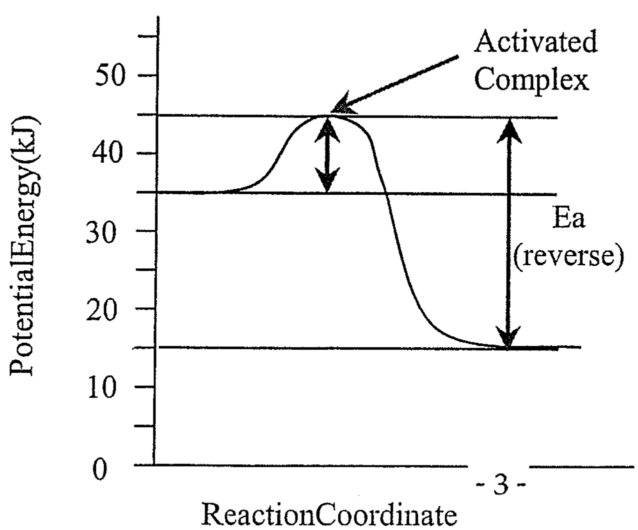
a) $E_a = 250\text{ kJ} - 100\text{ kJ}$
$= 150\text{ kJ}$

b) $\Delta H = H_{product} - H_{reactant}$
$= 25\text{ kJ} - 100\text{ kJ}$
$= -75\text{ kJ}$

c) Since ΔH is negative the reaction is exothermic.

d) For the reverse reaction, products now become the reactants. Therefore,
$E_a = 250\text{ kJ} - 25\text{ kJ}$
$= 225\text{ kJ}$

3.



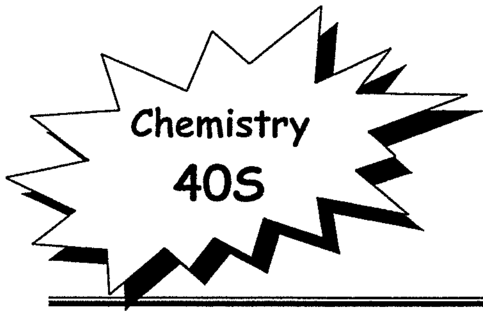
a) $E_a = 45\text{ kJ} - 35\text{ kJ}$
$= 10\text{ kJ}$

b) $\Delta H = H_{product} - H_{reactant}$
$= 15\text{ kJ} - 35\text{ kJ}$
$= -20\text{ kJ}$

c) Since ΔH is negative the reaction is exothermic.

d) $E_a = 45\text{ kJ} - 15\text{ kJ}$
$= 30\text{ kJ}$

Student Name: ______________________	Date: ____________________



Reaction Rates and Chemical Equilibrium

Kinetics Assignment #3 - Key

Answer the following questions in your Chemistry notebook.

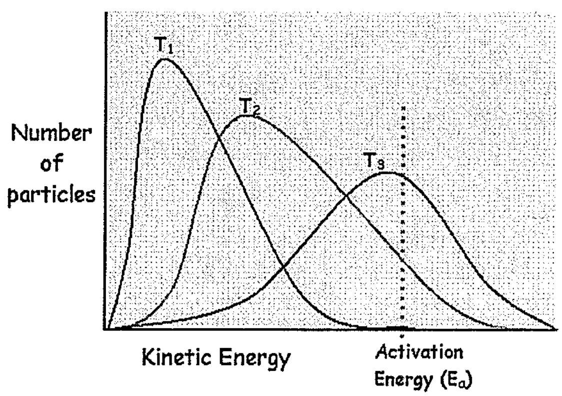
1. In general, what effect does an increase in the concentration of the reactants have on the rate of the reaction? (Explain using the collision theory.)

   Increasing concentration, according to the collision theory, increases the number of particles. An increase in the number of particles results in an increase in the number of collisions. This increase in the number of collisions increases the rate of the reaction.

2. How do changes each of the following factors affect the rate of a chemical reaction? Use diagrams to clarify your explanations.

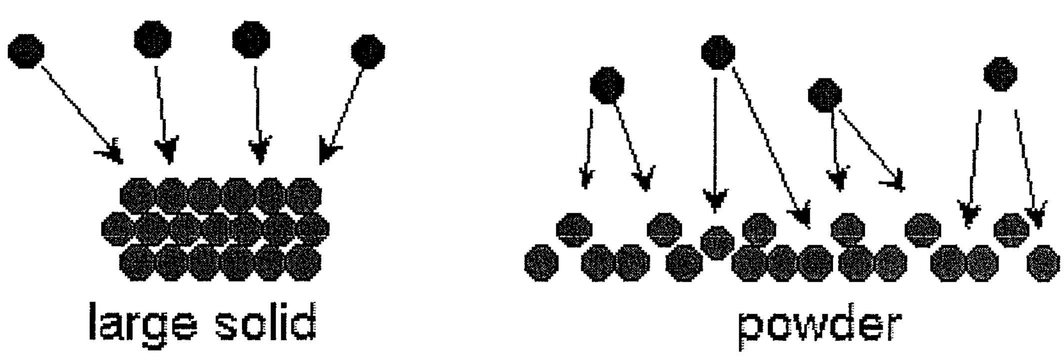
   a) temperature

   An increase in the temperature indicates an increase in the kinetic energy of the system. This increases the reaction rate. A decrease in temperature (kinetic energy) results in a decrease the rate of reaction.



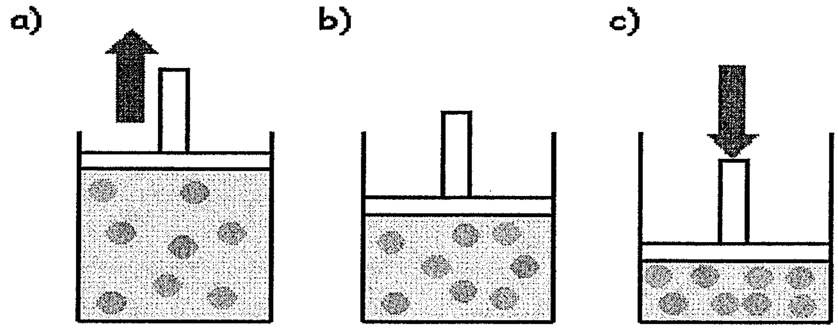
   b) particle size

   Decreasing particle size increases the surface area of the substance. This results in an increase in the number of collisions between particles and an increase in reaction rate.



c) pressure

Increasing pressure has the same effect as increasing the concentration of the particles involved in the reaction. A greater number of collisions results and the reaction rate increases.

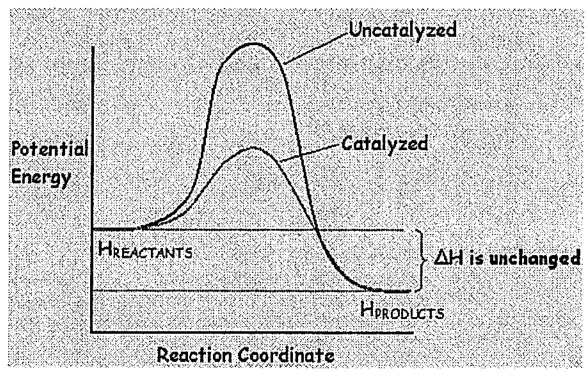


3. How does a catalyst influence the rate of a reaction? How do catalysts make this possible? Illustrate your answer with a specific example. Include a labeled diagram.

Catalysts increase the rate of reaction by lowering the activation barrier (reducing the activation energy required for the reaction to take place).

V_2O_5 is a catalyst for the reaction:

$SO_2(g) + O_2(g) \rightarrow 2\ SO_3(g)$



4. Which equation of the following pairs of equations would occur the fastest under the same conditions. Explain your answers.

a) **$Zn_2+(aq) + S_2-(aq) \rightarrow ZnS(s)$** is faster, since the reactants are all aqueous. The reactants in reaction (ii) are solids. Solid reactants react slowest.

b) **$Cu(s) + 2\ AgNO_3(aq) \rightarrow 2\ Ag+(aq) + Cu(NO_3)_2(aq)$** is the fastest. This reaction involves the transfer of electrons from Cu to Ag+, while the other involves the breaking of covalent bonds.

c) **$Pb(NO_3)_3(aq) + 2\ KI(aq) \rightarrow PbI_2(aq) + 2\ KNO_3(aq)$** is faster. Reactions with aqueous reactants are almost instantaneous. The other reaction involved the breaking of many covalent bonds.

d) **$2\ NO(g) + O_2(g)\ 2\ NO_2(g)$** is a homogeneous reaction of gaseous reactants, while the reaction $2\ Fe(s) + 3\ O_2(g) \rightarrow 2\ Fe_2O_3(s)$ is between heterogeneous reactants. (Consider how slowly iron rusts)

Student Name: ______________________ Date: ______________

Chemistry 40S

Reaction Rates and Chemical Equilibrium

Kinetic Assignment #4 - Key

Answer the following questions in your Chemistry notebook:

1. Sketch a potential energy (reaction coordinate) diagram for the endothermic reaction whose mechanism is:

$$H_2O_2 + I^- \rightarrow H_2O + IO^- \quad \text{(fast)}$$

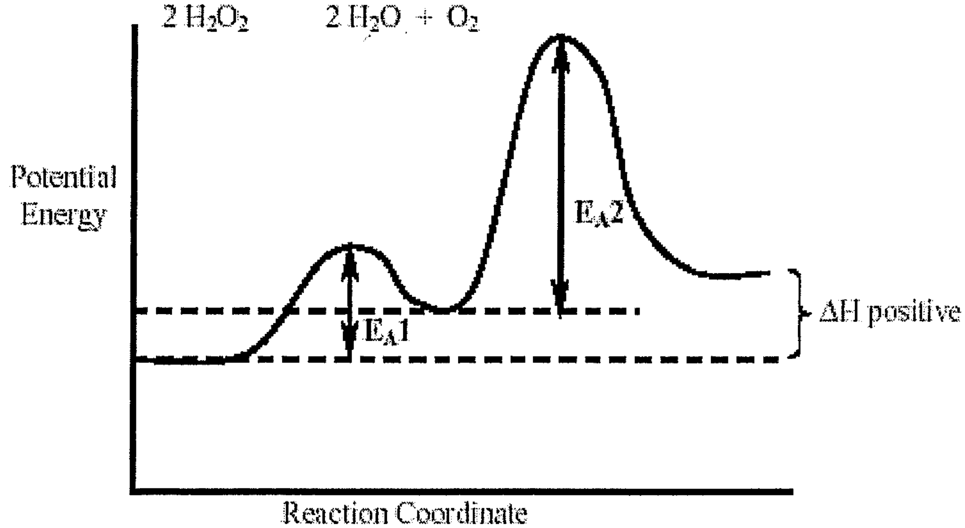
$$H_2O_2 + IO^- \rightarrow H_2O + O_2 + I^- \quad \text{(slow)}$$

Write the balanced net reaction.

$$H_2O_2 + \cancel{I^-} \rightarrow H_2O + \cancel{IO^-} \qquad \text{fast}$$

$$\underline{H_2O_2 + \cancel{IO^-}} \rightarrow \underline{H_2O + O_2 + \cancel{I^-}} \qquad \text{slow}$$

$$2\,H_2O_2 \qquad 2\,H_2O + O_2$$



2. Examine the following reaction mechanism:

$$P + Q \rightarrow I_1 + R \quad \text{(slow)}$$
$$I_1 + P \rightarrow I_2 + W \quad \text{(moderate)}$$
$$I_2 + S \rightarrow T \quad \text{(fast)}$$

a) Write out the net reaction.

Examine the following reaction mechanism:

P + Q d $\cancel{I_1}$ + R		slow
$\cancel{I_1}$ + P d $\cancel{I_2}$ + W		moderate
$\cancel{I_2}$ + S d T		fast

a) 2 P + Q + S d R + W + T

b) Identify the overall rate of the net reaction.

b) Since the rate determining step is slow, the entire rate is slow.

c) Increasing [P], increases the rate of the net reaction.
Increasing [Q], increases the rate of the net reaction.
Increasing [S], has no effect of the rate.
Explain why this is possible.

This occurs because both P and Q are present in the rate determining step, but S is not. Only reactants in the rate determining step will significantly affect the rate of the reaction.

3. A proposed mechanism for the preparation of the poisonous liquid nitrobenzene ($C_6H_6NO_2$) is:

$$C_6H_6 + NO_2{+} \rightarrow C_6H_6NO_2^+ \quad \text{(slow)}$$
$$H_2SO_4 \rightarrow H^+ + HSO_4^- \quad \text{(very fast)}$$
$$C_6H_6NO_2{+} + HSO_4^- \rightarrow C_6H_6NO_2 + H_2SO_4 \quad \text{(fast)}$$

a) What is the RDS? Why?

$C_6H_6 + NO_2{+} \rightarrow C_6H_6NO_2^+$ It is the slowest step.

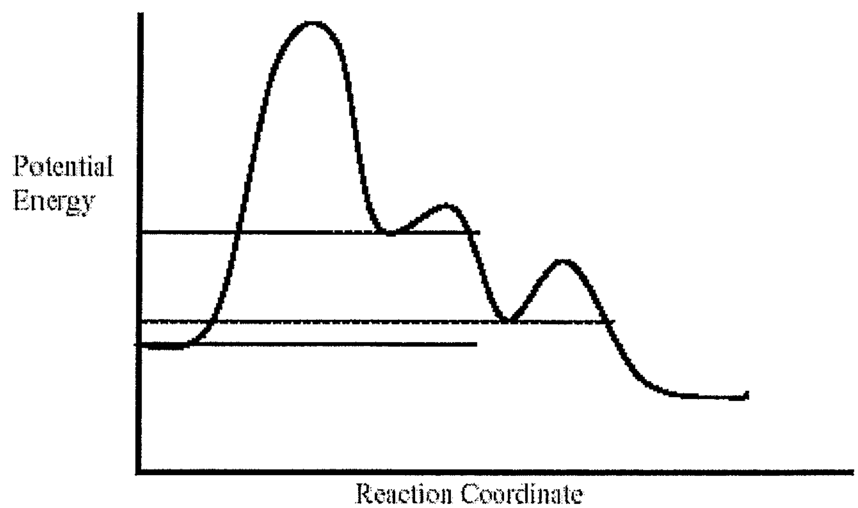
b) What is the net reaction?

$C_6H_6 + NO_2^+ \rightarrow C_6H_6NO_2^+$	slow
H_2SO_4 cl H^+ HSO_4^-	v. fast
$C_6H_6NO_2^+ + HSO_4^- \rightarrow C_6H_6NO_2 + H_2SO_4$	fast
$C_6H_6 + NO_2^+ \rightarrow C_6H_6NO_2$	

c) Without H_2SO_4 this is a very slow reaction. Explain.

H_2SO_4 is introduced in step 2 and is unchanged at the end of the reaction. H_2SO_4 is most likely a catalyst since it speeds up the reaction and is unchanged at the end of the reaction.

d) Sketch a possible EP diagram for this reaction.



Step 1 is slowest, therefore it has the largest activation energy. Step 2 is the fastest, so it should have the smallest activation enrgy. Step 3 is fast, so it should have a small activation energy.

4. The following mechanism is proposed for a reaction:

$$P + Q \rightarrow A \quad \text{(fast)}$$
$$A + 2\,Q \rightarrow X + Y \quad \text{(slow)}$$
$$X \rightarrow Z \quad \text{(very fast)}$$

a) What would happen to the rate as the concentration of Q was increased?

The reaction rate would increase.

b) Write the net reaction.

$$\begin{array}{ll} \mathbf{P + Q \rightarrow A} & \textbf{(fast)} \\ \mathbf{A + 2\,Q \rightarrow X + Y} & \textbf{(slow)} \\ \mathbf{X \rightarrow Z} & \textbf{(very fast)} \\ \hline \mathbf{P + 3\,Q \rightarrow Z + Y} & \end{array}$$

c) What is the RDS? Why?

A + 2 Q → X + Y It is the slowest step.

d) Why might step 2 be slower?

Possible reasons include:

Bonds difficult to break

High E_A

Others?

5. Draw the 'reaction progress' (reaction coordinate) diagram to illustrate the following reaction mechanism. Write the net reaction for the mechanism.

. ΔH = -25 kJ/mol for the net reaction

$A + B \rightarrow I_1 + C$ (very slow, endothermic)

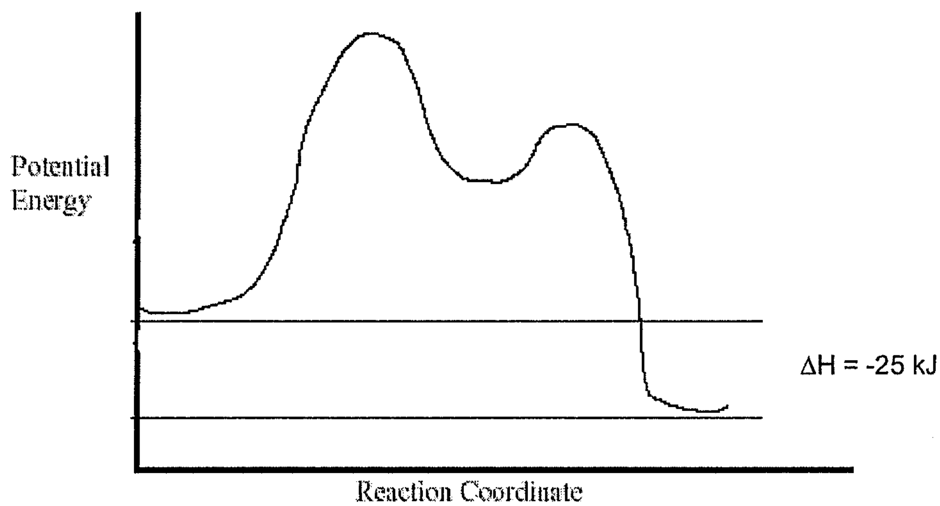
$I_1 + A \rightarrow C + D$ (fast, exothermic)

Solution:

$\mathbf{A + B \rightarrow I_1 + C}$ **(very slow, endothermic)**

$\mathbf{I_1 + A \rightarrow C + D}$ **(fast, exothermic)**

$\mathbf{2A + B \rightarrow 2\,C + D}$



Kinetics Assign: #5

Rate Laws - Chapter 17

EXAMPLE 1: Rate Law = $K[A]^a$

Problem #6: $NH_4^+ (aq) + NO_2^- (aq) \rightarrow N_2(g) + 2H_2O(l)$

Rate = $K[A]^1[B]^1$

NO_2^- = first order
NH_4^+ = first order

Second order overall 1+1 = 2

Mr. Hart Rate of Reaction Problems

1. $A(g) + 2B(g) \rightarrow C(g)$

a) Rate = $K[A][B]^2$

b) If [A] is doubled, The rate doubles (2x)

c) If [B] is Tripled, The rate increases by a factor of 9 (9x)

d) If the container is compressed to 1/4 the volume, the rate is 64x greater.

2. $2A(g) + 3B(g) \rightarrow 2C(g) + D(g)$

a) Rate = $K[A]^2[B]^3$

b) If [A] is cut in half, the reaction is 0.25 x as fast
If [B] is 4x greater, the reaction is 64 x faster.
0.25 x 64 = 16x greater

c) If [A] is 5x greater and [B] is 0.5 x greater
Rate = $[5]^2[0.5]^3$ = [25][0.125]
= 3.125 x greater

3. $2NO_{(g)} \rightarrow N_2O_{2(g)}$

a) $Rate = K[NO]^2$

b) $Rate = 0.50\ M/min$ $\quad Rate = K[NO]^2$

$[NO] = 2.02\ M$

$K = ?$

$$K = \frac{Rate}{[NO]^2}$$

$$= \frac{0.50\ M/min}{[2.02\ M]^2} \qquad \frac{\cancel{Mol/L \times min}}{\cancel{[Mol/L][Mol/L]}}$$

$$= \frac{0.50\ M/Min}{4.0804\ M^2}$$

$$= 1.2253 \times 10^{-1}\ L/\cancel{mol}\ 1/M \cdot M$$

$$= 0.123$$

c) $[NO] = 1.00 \times 10^{-2}\ M$

Rate = ? $\quad Rate = K[NO]^2$

$$= 0.123\ [1.00 \times 10^{-2}\ M]^2$$

$$= 1.23 \times 10^{-5}\ M/Min$$

d) [NO] is doubled $\quad [2]^2 = 4X$ increase

Rate = ?

e) $[NO] \times \frac{1}{3}$ $\quad [\frac{1}{3}]^2 = \frac{1}{9} X$ increase

Rate of Reaction Problems - Mr Hart

(4) $H_2 + I_2 \rightarrow 2HI$

a) Rate = $K[H_2][I]_2$

b) $(H_2) \times 2$ $[I_2] \times 0.5$ $2 \times 0.5 = 1$ ∴ Rate is the same.

c) HI produced @ rate of 0.020 M/min

$[H_2] = 2.50 \times 10^{-2}$ M

$[I_2] = 5.0 \times 10^{-2}$ M

k = ?

Rate = $K[H_2][I_2]$

$$k = \frac{\text{Rate}}{[H_2][I_2]}$$

$$= \frac{0.020\ M/L\text{-min}}{(2.50 \times 10^{-2} M)(5.0 \times 10^{-2} M)}$$

$$= \frac{0.020 M/L\text{-min}}{0.00125\ M/L^2}$$

$$= 16\ \cancel{M}/\cancel{L}\text{-min}$$

M·min

d) Rate = ?

$[H_2] = 0.10M$

$[I_2] = 0.10M$

Rate = $K[H_2][I_2]$

$= 16\ M/L\text{-min}\ [0.10M][0.10M]$

$= 0.16\ M/L\text{-M}$

5. $2A + C \rightarrow D + E$

a) $Rate = k[A]^2[C]$

b) $[A] = (\frac{1}{2})^2$ $\quad \frac{1}{4} \times \frac{1}{2} = \frac{1}{8} X$ * because concentration is halved for both reactants.

$[C] = \frac{1}{2}$

c) Rate = 8.00 M/min

$[A] = 1.00\ M$ $\quad Rate = k[A]^2[C]$

$[C] = 2.00\ M$ $\quad k = \frac{Rate}{[A]^2[C]}$

$k = ?$

$\frac{\frac{M}{min}}{(M^2)M}$ $\quad \frac{\frac{M}{min}}{M^3} = \frac{\frac{1}{min}}{M^2}$ $\quad \frac{1}{M^2 \cdot min}$

$k = \frac{8.00\ M/min}{[1.00M]^2[2.00M]}$

$k = 4.00\ M/L\text{-}min$ $\cancel{M/L}\ m,$ or $1/M^2 \cdot min$

d) $[A] = 2.00\ M$ $\quad Rate = k[A]^2[C]$

$[C] = 3.00\ M$ $\quad = 4.00\ M/L\text{-}min\ [2.00M]^2[3.00M]$

Rate = ? $\quad = 48.0\ M/min$

6. $2B \rightarrow C + D$

a) $Rate = k[B]^2$

b) $[B] = 1.50\ M$ $\quad Rate = k[B]^2$

rate = 3.00 moles/L-s $\quad k = \frac{Rate}{[B]^2}$

$k = ?$

$= \frac{3.00\ moles/L\text{-}s}{[1.50M]^2}$

$= \frac{3.00\ mole/L\text{-}s}{2.25\ mol^2/L^2}$

$= 1.33\ M/\cancel{L\text{-}min}$ $= \frac{M}{min}$

$\frac{mol}{L \cdot S} \rightarrow \frac{\cancel{mol}}{\cancel{L} \cdot S} \times \frac{L^{2}}{mol^{2}}$ $\quad \frac{L}{mol \cdot S}$

$\frac{mol^2}{L^2}$

$2A + B \rightarrow C + D$

a) $Rate = K[A]^2[B]$

b) $[A] = 6 \times$ greater $\qquad 36 \times \frac{1}{3} = 12 \times$ greater

$[B] = \frac{1}{3} \times$ greater.

c) $[A] = 2.25\,M$ $\qquad Rate = K[A]^2[B]$

$[B] = 0.500\,M$ $\qquad K = \frac{Rate}{[A]^2[B]}$

$Rate = 4.00\,M/s$

$$= \frac{4.00\,M/s}{[2.25\,M]^2[0.500\,M]}$$

$$= \frac{4.00\,M/s}{2.53125\,M^3}$$

$$= 1.58\ M^2/L\text{-}s \quad \frac{1}{M^2 \cdot s}$$

d) $[A] = 5.50\,M$

$[B] = 0.400\,M$ $\qquad Rate = K[A]^2[B]$

$Rate = ?$

$$= 1.58\ M/L\text{-}s\,[5.50\,M]^2[0.400\,M]$$

$$= 19.1\ M/s$$

8. $2A \rightarrow C$ $\qquad$ a) $Rate = K[A]^2$

b) $~~Rate~~ = \frac{Rate}{[A]^2}$ $\qquad Rate = 20.0\,M/s$

$[A] = 15.0\,moles / 2.50\,L$

$= 6.00\,M$

$$= \frac{20.0\,M/s}{[6.00\,M]^2}$$

$$= 0.556\ \frac{1}{M \cdot s}$$

c) $[A] = 2.00\,M$ in 400 ml

$[A] = 5.00\,M$ $\qquad Rate = k[A]^2$

$$= 5.56 \times 10^{-1}\,[5.00\,M]^2$$

$$= 13.9\ M/s$$

9. $2A + 3B \rightarrow C + D$ $\qquad$ $A + 2B \rightarrow E$

$E + A + B \rightarrow C + D$

a) $Rate = K[A][B]^2$

b) $[A] = 4x$ $\qquad$ $[4][\frac{1}{3}]^2 = 0.445$ X greater

$[B] = \frac{1}{3}x$

c) $K = ?$ $\qquad$ $Rate = K[A][B]^2$

rate = 8.00 M/min $\qquad$ $K = \frac{Rate}{[A][B]^2}$

$[A] = 0.500 M$

$[B] = 1.20 M$ $\qquad$ $= \frac{8.00 \, M/min}{[0.500 M][1.20 M]^2}$

$= \frac{8.00 \, M/min}{0.72 \, M}$

$= 11.1 \, 1/M^2 \cdot min$

d) $[A] = 2.00 M$ $\qquad$ $Rate = K[A][B]^2$

$[B] = 1.80 M$ $\qquad$ $= 11.1 \, [2.00 M][1.80 M]^2$

$= 71.9 \, M/Min = 72.0 \, M/min$

10. $2A + B \rightarrow$ Products $\qquad$ $R = K[A]^2[B]$

a) simple reaction because rate law [] are same as reaction equation

b) $= (3)^2(3)$

$= 9(3)$

= 27 Xtimes greater

Student Name: KEY Date: ____________

Chemistry 40S

Reaction Rate
Kinetics Assignment #6

Answer each of these questions on loose-leaf paper.

1. Assume that NO(g) and H_2(g) react according to the rate law:

$$\text{Rate} = k[No]^2\,[H_2]$$

How does the rate change if:

a) The concentration of H_2 doubles?

Reaction is first order in $[H_2]$ ∴ doubling $[H_2]$ doubles the reaction rate.

b) The volume of the enclosing vessel is suddenly halved?

Halving the volume doubles the [ ].

∴ $(2)^2 \times (2) = 4 \times 2 = 8\times$ greater rate

c) The temperature decreases?

Temp = average KE.

decreasing temperature decreases reaction rate.

2. The reaction $CH_3COOH_3 + I_2 \rightarrow CH_3COCH_2I + HI$ is run under carefully controlled conditions in the presence of an excess of acid. The following data were obtained:

	$[CH_3COOH_3]$ Mol/L	$[I_2]$ Mol/L	Rate (mol/Ls)	
Trial 1	0.100	0.100	1.16×10^{-7}	
Trial 2	0.0500 2x	0.100 1x	5.79×10^{-8}	2x
Trial 3	0.0500 1x	0.500 0.2x	5.79×10^{-8} 1x	

Write the rate law for the reaction.

$[CH_3COOH_3]$ is 1st order.
$[I_2]$ is zero order

$\text{Rate} = k[CH_3COOH_3]$

Reaction Rate Worksheet Key #7

1. $2A + B \rightarrow C + E$

a) $Rate = k[B]$

b) $k = \dfrac{Rate}{[B]}$

$k = \dfrac{1.4\ M/min}{2.5M}$

$k = 0.56\ {}^{1}/_{min}$

2. a) $Rate = k\cancel{[A]}[B]$

b) $k = \dfrac{Rate}{\cancel{[A]}[B]}$

$k = \dfrac{0.20\ mol/L \cdot min}{\cancel{(2.0\ mol/L)}(1.50\ mol/L)}$

$k = \overset{7.0}{\cancel{0.67}}\ {}^{1}/_{\cancel{M} \cdot min}$

?

3. a) $Rate = k[A]^2$

b) $k = \dfrac{Rate}{[A]^2}$

$k = \dfrac{0.11 \times 10^{-2}\ M/s}{(0.50\ M)^2}$

$k = 4.4 \times 10^{-3}\ {}^{1}/_{M \cdot s}$