Reaction Rates and Equilibrium Exam Review

1. For the reaction $NO_2 + O_2 \rightarrow 2N_2O_4$, the data shown was collected.

Trial #	Initial $[NO_2]$	Initial $[O_2]$	Initial Rate of formation of N_2O_4
1	0.025 M	0.011 M	3.1×10^{-4} mol/L·s
2	0.025 M	0.022 M	6.2×10^{-4} mol/L·s
3	0.050 M	0.011 M	6.2×10^{-4} mol/L·s

$2^x = 2$ $2^x = 2$

a. Write the rate law expression for the reaction.

$rate = k[NO_2]^1[O_2]^1$

b. Calculate the value of the rate constant. (k) pick a trial & plug in!

$3.1 \times 10^{-4} = k(0.025)(0.011)$

$3.1 \times 10^{-4} = k(2.75 \times 10^{-4})$

$k = 1.1$

c. Use the rate law expression and calculated value of k to compute the initial rate of formation of N_2O_4, if the initial concentrations of NO_2 and O_2 were each 0.030M.

$rate = (1.127...)(0.030)(0.030)$

$rate = 0.0010$ mol/L·s

2. An energy diagram is shown to the right. Answer the following:

a. What is the definition of "activated complex"?

high energy transition state between reactants & products.

b. What is the definition of "activation energy"?

amt. of energy required to begin the rxn.

c. What is the potential energy of the products? 350 KJ

d. What is the potential energy of the reactants? 200 KJ

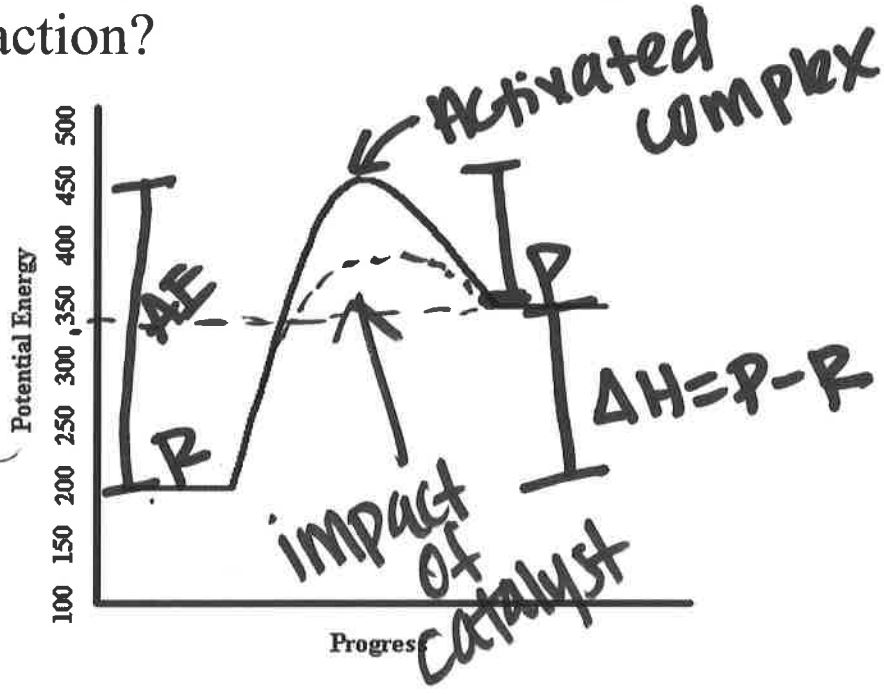
e. What is the activation energy of the forward reaction? → 450-200=250 KJ

f. What is the activation energy of the reverse reaction? ← 450-350=100 KJ

g. What is ΔH for the forward reaction? 350-200=150 KJ, endo bc +ΔH

h. Draw a line to represent the reaction with a catalyst. What does the catalyst do to the rate of the reaction?

Speeds up the reaction by lowering the activation energy.



3. The rate law for the reaction $NO(g) + O_3(g) \rightarrow NO_2(g) + O_2(g)$ is a first-order in NO and O_3, and second order overall. Write the complete rate law for this reaction.

rate = $k[NO]^1[O_3]^1$ or rate = $k[NO][O_3]$

4. Consider the decomposition of N_2O_5 in carbon tetrachloride (CCl_4) at 45°C.

$2\ N_2O_5\ (aq) \leftrightarrow 4\ NO_2\ (g) + O_2\ (g)$ rate = $k[N_2O_5]^1$

The reaction is first-order in N_2O_5, with the specific rate constant 6.08×10^{-4} s^{-1}. Calculate the reaction rate at these conditions. ↑ k

a. $[N_2O_5] = 0.200$ mol/L

rate = $k[N_2O_5]$

= $(6.08 \times 10^{-4})(0.200)$

rate = 1.22×10^{-4} mol/L·s

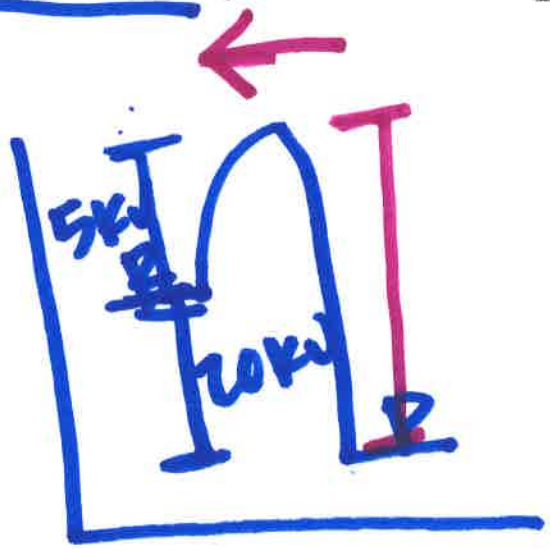
b. $[N_2O_5] = 0.319$ mol/L

rate = $k[N_2O_5]$

= $(6.08 \times 10^{-4}\ 1/s)(0.319$ mol/L)

rate = 1.94×10^{-4} mol/L·s

5. For the reaction A + B ↔ C, the activation energy of the forward reaction is 5 kJ and the total energy change is -20 kJ. What is the activation energy of the reverse reaction? *(hint: draw a picture)*



AE: 5kJ
ΔH: P-R = -20kJ
exo

Activation energy for reverse rxn: 25kJ

6. For the following reaction: $Zn\ (s) + 2\ H^+\ (aq) \rightarrow Zn^{2+}\ (aq) + H_2\ (g)$ **The reaction is exothermic.**

+heat
0(g)
1(g)
look at (g) only!

Indicate the direction of the equilibrium shift.

a. add H^+ →
b. remove Zn^{2+} →
c. decrease temp →
d. decrease pressure →
e. add Zn (s) →
f. add catalyst n/c

endo: heat is a reactant
exo: heat is a product

7. A mixture at equilibrium at 827°C contains 0.552 mol CO_2, 0.552 mol H_2, 0.448 mol CO, 0.448 mol H_2O. What is the value of K_{eq}? CO_2 (g) + H_2 (g) ↔ CO (g) + H_2O (g)

$K_{eq} = \frac{[P]}{[R]}$

$K_{eq} = \frac{[CO][H_2O]}{[CO_2][H_2]} = \frac{(0.448)(0.448)}{(0.552)(0.552)}$

$= 0.200704$

$= \frac{0.200704}{0.304704}$

$K_{eq} = 0.659$

Would more products or reactants be present at equilibrium?

reactants b/c $K_{eq} < 1$.

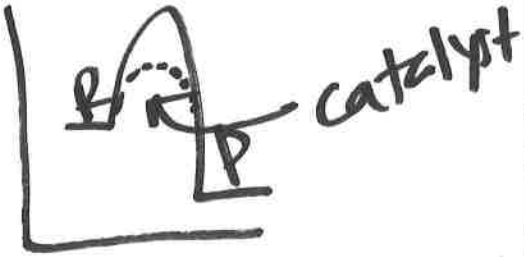
8. List 3 ways that reaction rates can generally be increased:

a) add a catalyst b) increase concentration c) increase temp

d) decrease particle size (inc. surface area)

9. A catalyst speeds up reaction rates by lowering the activation energy.

Biological catalysts are called enzymes.



10. It has been determined that the rate law for the following reaction is third-order overall and first-order with respect to $HgCl_2$. Write the rate law for this reaction:

$$2\,HgCl_2 + Na_2C_2O_4 \leftrightarrow 2\,NaCl + 2\,CO_2 + Hg_2Cl_2$$

$rate = k[HgCl_2]^1[Na_2C_2O_4]^2$

11. Write the equilibrium constant expression for the following reaction:

$2\ N_2O_5(g) \leftrightarrow 4\ NO_2(g) + O_2(g)$

Calculate K_{eq} for the reaction if the equilibrium concentrations are:

$[N_2O_5] = 0.50$ M, $[NO_2] = 0.80$ M, and $[O_2] = 0.20$ M.

$$K_{eq} = \frac{[P]}{[R]} \qquad K_{eq} = \frac{[NO_2]^4[O_2]}{[N_2O_5]^2} = \frac{(0.80)^4(0.20)}{(0.50)^2}$$

$$= \frac{0.08192}{0.25}$$

$K_{eq} = 0.33$ more react. b/c $K_{eq} < 1$

12. For the following reaction, calculate which direction the reaction will shift when these stresses are applied:

$2\ NO_2(g) \leftrightarrow N_2O_4(g) + 57.2\ kJ$

a. The temperature is increased ← ____

b. The pressure is lowered ← ____

look at gas only!

c. NO_2 is added → ____

d. N_2O_4 is added ← ____

LeChat.

ΔS

13. Is entropy (disorder) increasing or decreasing in these reactions or processes?

a. $2\ XeO_3\ (s) \rightarrow 2\ Xe\ (g) + 3\ O_2\ (g)$ inc.

b. $NaCl\ (s) \rightarrow Na^+\ (aq) + Cl^-\ (aq)$ inc.

c. Loose stamps → Stamps in an album dec.

d. Messy room → clean room dec.

e. Ice → water inc.

s ℓ

inc : +ΔS

dec : −ΔS

14. Which combination of factors will always give a spontaneous reaction?

a. +ΔH, −ΔS

(b.) −ΔH, +ΔS

c. −ΔH, −ΔS

d. +ΔH, −ΔS

↑ happens on its own (−ΔG)

high disorder, +ΔS | exo, −ΔH

15. Which combination(s) in question #14 will never give a spontaneous reaction? A & D

−ΔS, +ΔH