Unit 5 Major Quiz Review

Note: mol L^{-1} means M

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

For the gas-phase reaction represented above, the following experimental data were obtained.

Experiment	Initial [A] (mol L^{-1}) M	Initial [B] (mol L^{-1}) M	Initial Reaction Rate (mol L^{-1} s^{-1}) M/s
1	0.033	0.034	6.67×10^{-4}
2	0.034	0.137	1.08×10^{-2}
3	0.136	0.136	1.07×10^{-2}
4	0.202	0.233	?

Use expts 2+3

(a) Determine the order of the reaction with respect to reactant A. Justify your answer.

$$\frac{\text{rate of 3}}{\text{rate of 2}} = \left(\frac{1.07\times10^{-2}}{1.08\times10^{-2}}\right) = \frac{\cancel{k}(.136)^a\, \cancel{0.137}^b}{\cancel{k}(.034)^a\, \cancel{0.136}^b}$$

$$1 = 4^a \text{ so } \boxed{a=0}$$

(b) Determine the order of the reaction with respect to reactant B. Justify your answer. Use expts 1+2

$$\frac{\text{rate of 2}}{\text{rate of 1}} = \frac{1.08\times10^{-2}}{6.67\times10^{-4}} = \frac{\cancel{k}(0.137)^b\,(\cancel{.034})^a}{\cancel{k}(0.034)^b\,(\cancel{.033})^a}$$

$$16 = 4^b \text{ so } \boxed{b=2}$$

(c) Write the rate law for the overall reaction. $\underline{\text{rate} = k[A]^0[B]^2}$ or rate $= k[B]^2$

(d) Determine the value of the rate constant, k, for the reaction. Include units with your answer. Use any expt.

expt. 1

$$\text{rate} = k[B]^2$$

$$6.67\times10^{-4}\,\frac{M}{s} = k\,(.034\,M)^2$$

$$\frac{6.67\times10^{-4}\,\frac{M}{s}}{(.034\,M)^2} = k = .58\,M^{-1}s^{-1}$$

$$\frac{\frac{M}{s}}{M^2} = \frac{\cancel{M}}{s}\times\frac{1}{M^{\cancel{2}}} = \frac{1}{Ms} = M^{-1}s^{-1}$$

(e) Calculate the initial reaction rate for experiment 4.

expt. 4

$$\text{rate} = k[B]^2$$

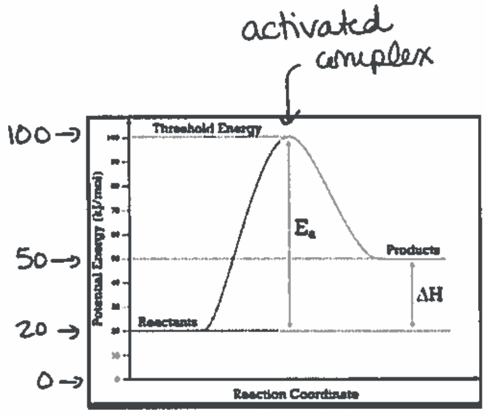
$$\text{rate} = .58\,M^{-1}s^{-1}\,(.233\,M)^2 = .031\,M/s$$

2. a) What two factors are required for a successful molecular collision to occur?

1) correct / sufficient energy (activation energy) 2) correct orientation

b) What are the four factors that can change the rate of a reaction? How does each factor affect the rate of successful collisions?

1) ↑ concentration of reactant – more particles = more collisions occur
2) ↑ temp – more energy means a greater fraction of molecules have the activation energy required for the reaction to start
3) ↑ surface area – particles surround each other more easily – increases chance of correct orientation for collision
4) catalyst – lowers activation energy so more particles have the energy for collision



3. The following questions refer to the diagram.

a) Is this reaction endothermic or exothermic?

b) How much is the activation energy for this reaction? $100 - 20 = 80$ KJ

c) How much is the energy of the activated complex?

100 KJ

4.

	$2\ H_2S(g) \leftrightarrow 2\ H_2(g) + S_2(g)$		
R	$2\ H_2S(g)$	$2\ H_2(g)$	$S_2(g)$
I	.0800 M	0	0
C	−.0596 M	+.0596	+.0298
E	.0204 M	.0596 M	$=(3.72\times10^{-2})/1.25 = .0298$ M

$[H_2S]$ init. $= \frac{.100}{1.25} = .0800$

A 0.100 mol sample of H2S*(g)* is introduced into an evacuated rigid 1.25 L container. The sealed container is heated to 483 K, and 3.72×10^{-2} mol of $S_2(g)$ is present at equilibrium.

(a) Write the expression for the equilibrium constant, ~~Kc, for~~ the decomposition reaction represented above.

$$K_c = \frac{[H_2]^2[S_2]}{[H_2S]^2}$$

(b) Calculate the equilibrium concentration, in mol L-1, of the following gases in the container at 483 K.

(i) $H_2(g)$ $\quad .0298 \times 2 = .0596$ M

(ii) $H_2S(g)$ $\quad .0800$ M $- .0596$ M $= .0204$ M

(c) Calculate the value of the equilibrium constant, *Kc*, for the decomposition reaction at 483 K.

sub. values from b)

$$K_c = \frac{(.0596)^2(.0298)}{(.0204)^2} = .254 \quad \text{(favors reactants)}$$

See packet problems

~~(d) In a different experiment (but the same temperature), 0.150 M of H2S is present initially and then goes to equilibrium. Calculate the equilibrium concentrations of all three species.~~

Yes, it's a cubic... I will make sure the quiz will feature a quadratic!

R	$2H_2S$	$\leftrightarrow 2H_2$	$+ S_2$
I	0.150 M	0	0
C	−2x	+2x	+x
E	.150−2x	2x	x

$$K_c = \frac{[H_2]^2[S_2]}{[H_2S]^2} = \frac{(2x)^2(x)}{(.150-2x)^2} = .254$$

$$\frac{4x^3}{.0225 - .600x + 4x^2} = .254$$

$$4x^3 = .254(.0225 - .600x + 4x^2) = .005715 - .1524x + 1.016x^2$$

5. $\text{heat} + NH_4Cl(s) \leftrightarrow NH_3(g) + HCl(g)$ $\quad \Delta H = +42.1$ kilocalories endothermic

For this system at equilibrium and at 600 K, state whether the concentration of $NH_3(g)$ will have increased, decreased, or remained the same when equilibrium is reestablished. Justify each answer with a one-or-two sentence explanation.

shift?

none (a) A small quantity of NH_4Cl is added. Solids do not change or affect equilibrium position

R (b) The temperature of the system is increased. System will use up excess heat by adding to products

R (c) The volume of the system is increased. (P↓) System needs to form more particles to fill up the space created by the change

L (d) A quantity of gaseous HCl is added. This reduces the excess HCl you added by forming more reactant.

L (e) An inert gas is added. (P↑) → This means the system must shift to reduce the # of moles of gas in the container since space has been reduced (by adding the inert gas)

none (f) A catalyst is added. → This increases the rate at which equilibrium is reached – not the position