Chem 104
Test 2 Practice Problems

1. a. $K_c = \frac{[CO]^2}{[CO_2]}$ Omit solid C(s)

b. $Q = \frac{(0.750)^2}{0.500} = 1.12_5 < K_c$

$\Rightarrow$ System must shift right (make more product) to reach equilibrium.

c. $\Delta n = 2 - 1 = 1 \Rightarrow K_p = K_c RT$

$K_p = (1.603\ \cancel{mol}/L)(0.08206\ L \cdot atm/\cancel{K} \cdot \cancel{mol})(1273\ \cancel{K})$
$= 167.4_5$ atm

d. $K_p = \frac{P_{CO}^2}{P_{CO_2}} = 167.4_5\ atm = \frac{P_{CO}^2}{0.100\ atm}$

$P_{CO}^2 = 16.74_5\ atm^2 \Rightarrow P_{CO} = 4.092$ atm

e. Shift left

f. Forward reaction is endothermic, so higher temperature favors product CO(g) formation.

g. K_c will increase at higher temperature.

2\. a. $k = 0.693/t_{1/2} = 0.693/21.8\ min = 0.0318\ min^{-1}$

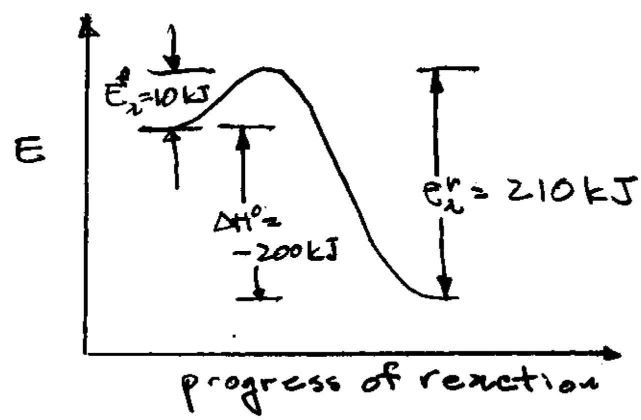
b. elapsed half-lives $= n = t/t_{1/2} = 54.5\ min/21.8\ min$
$= 2.5$

$[N_2O_5] = (0.500\ mol)(1/2)^{2.5} = (0.500\ mol)(0.177)$
$= 0.0884\ mol$

c. Plot $\ln[N_2O_5]$ vs. time. The straight line for this will have slope $= -k$.

d. Plot $\ln k$ vs. $1/T$. The straight line will have slope $= -E_a/R$. The y intercept will give $\ln A$. [To do this properly, the student should obtain k values at more than two temperatures.]

3\.



4. $$H_2O(g) + CO(g) \rightleftharpoons H_2(g) + CO_2(g) \quad K_c = 1.30$$

From the given initial concentrations, calculate Q to see in which direction the reaction runs to reach equilibrium:

$$Q = \frac{[H_2][CO_2]}{[H_2O][CO]} = \frac{(0.100)^2}{(1.20)^2} = 6.9 \times 10^{-3} << K_c$$

The reaction must go right to reach equilibrium.

$$\begin{array}{cccc} H_2O(g) & + \; CO(g) & \rightleftharpoons H_2(g) & + \; CO_2(g) \\ 1.20 - x & 1.20 - x & 0.100 + x & 0.100 + x \end{array}$$

$$K_c = 1.30 = \frac{(0.100 + x)^2}{(1.20 - x)^2} \Rightarrow 1.14 = \frac{0.100 + x}{1.20 - x}$$

$$1.36_8 - 1.14x = 0.100 + x$$
$$2.14x = 1.26_8$$
$$x = 0.592_6 = 0.593$$

$$[H_2] = [CO_2] = 0.100 + 0.593 = 0.693 \text{ mol/L}$$
$$[H_2O] = [CO] = 1.20 - 0.593 = 0.60_7 \text{ mol/L} = 0.61 \text{ mol/L}$$

Check:

$$Q = \frac{(0.693)^2}{(0.607)^2} = 1.30 = K_c \Rightarrow OK$$

5. $H_2(g) + I_2(g) \rightleftharpoons 2HI(g) \quad K_c = 54.8$

$0.500 - x \quad 0.500 - x \quad 0.500 + 2x$

$$K_c = 54.8 = \frac{(0.500+2x)^2}{(0.500-x)^2} \Rightarrow 7.40 = \frac{0.500+2x}{0.500-x}$$

$$3.70 - 7.40x = 0.500 + 2x$$
$$9.40x = 3.20$$
$$x = 0.340_4$$

$[H_2] = [I_2] = 0.500 - 0.340 = 0.160$ mol/L

$[HI] = 0.500 + (2)(0.340_4) = 1.180_8$ mol/L

Check:

$$Q = \frac{(1.180_8)^2}{(0.160)^2} = 54.4_6 \approx K_c \Rightarrow \text{OK}$$

6. a. $rate_2 = k_2[NOCl_2][NO]$

b. $NOCl_2$

c. $K_c = \frac{[NOCl_2]}{[NO][Cl_2]}$

d. $[NOCl_2] = K_c[NO][Cl_2]$

$$Rate = rate_2 = k_2\{K_c[NO][Cl_2]\}[NO]$$
$$= k_{obs}[NO]^2[Cl_2]$$

e. Yes

7. Determine the rate law and calculate the value of the rate constant (with the appropriate units) for the reaction

$$A + B \rightarrow \text{products}$$

given the following data:

Exp.	[A]	[B]	Rate, M/s
#1	0.125	0.125	1.04 x 10^{-4}
#2	0.375	0.125	9.36 x 10^{-4}
#3	0.375	0.250	9.36 x 10^{-4}

$\Rightarrow \text{Rate} \propto [A]^2$

$\Rightarrow \text{Rate} \propto [B]^0$

$\Rightarrow \text{Rate} = k[A]^2 \Rightarrow k = (1.04 \times 10^{-4}\,M/s)/(0.125\,M^2) = 6.66 \times 10^{-3}\,M^{-1}\cdot s^{-1}$

8. Determine the rate law and calculate the value of the rate constant (with the appropriate units) for the reaction

$$A_2 + B + C \rightarrow AB + AC$$

given the following data:

Exp.	$[A_2]$, M	[B], M	[C], M	Rate, $M\cdot s^{-1}$
#1	0.125	0.111	0.702	1.07 x 10^{-3}
#2	0.500	0.111	0.702	2.14 x 10^{-3}
#3	0.125	0.444	0.702	4.28 x 10^{-3}
#4	0.125	0.444	0.351	4.28 x 10^{-3}

$\Rightarrow \text{Rate} \propto [A_2]^{1/2}$

$\Rightarrow \text{Rate} \propto [B]$

$\Rightarrow \text{Rate} \propto [C]^0$

$\Rightarrow \text{Rate} = k[A_2]^{1/2}[B] \Rightarrow k = (1.07 \times 10^{-3}\,M\cdot s^{-1})/(0.125\,M)^{1/2}(0.111\,M) = 2.73 \times 10^{-2}\,M^{-1/2}\cdot s^{-1}$

9. Determine the rate law and calculate the value of the rate constant (with the appropriate units) for the reaction

$$A + B + C \rightarrow \text{products}$$

given the following data:

Exp.	[A], M	[B], M	[C], M	Rate, $M\cdot s^{-1}$
#1	0.128	0.384	0.702	3.56 x 10^{-3}
#2	0.384	0.384	0.702	1.07 x 10^{-2}
#3	0.128	0.128	0.702	3.56 x 10^{-3}
#4	0.128	0.128	0.351	8.90 x 10^{4}

$\Rightarrow \text{Rate} \propto [A]$

$\Rightarrow \text{Rate} \propto [B]^0$

$\Rightarrow \text{Rate} \propto [C]^2$

$$\Rightarrow \text{Rate} = k[A][C]^2 \Rightarrow k = \frac{3.56 \times 10^{-3}\,M\cdot s^{-1}}{(0.128\,M)(0.702\,M)^2}$$

$$= 5.64 \times 10^{-2}\,M^{-2}\cdot s^{-1}$$