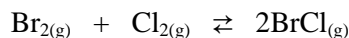
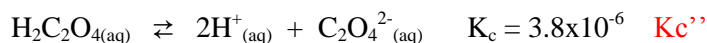
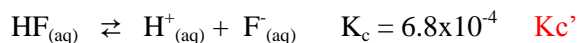


1. In an analysis of the following reaction at 100 °C, the equilibrium concentrations were found to be the following: $[\text{Br}_2] = 2.3 \times 10^{-3} \text{ M}$, $[\text{Cl}_2] = 1.2 \times 10^{-2} \text{ M}$, and $[\text{BrCl}] = 1.4 \times 10^{-2} \text{ M}$. Write the equilibrium expression and calculate the equilibrium constant for this reaction at 100 °C. (10 pts)

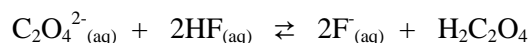


$$K_c = \frac{[\text{BrCl}]^2}{[\text{Br}_2][\text{Cl}_2]} = \frac{(1.4 \times 10^{-2})^2}{(2.3 \times 10^{-3})(1.2 \times 10^{-2})} = 7.1$$

2. Given the following information:



Determine the value of K_c for the following reaction: (10 pts)



$$K_c = (K_c')^2 \left(\frac{1}{K_c''} \right) = (6.8 \times 10^{-4})^2 \left(\frac{1}{3.8 \times 10^{-6}} \right) = 0.12$$

3. Calculate K_p for the following reaction at 35 °C: (10 pts)



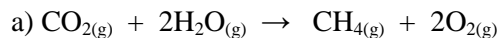
$$K_p = K_c(0.0821 \times T)^{\Delta n}$$

$$T = 273 + 35 = 308$$

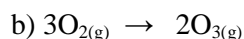
$$\Delta n = 2 - 3 = -1$$

$$K_p = (6.5 \times 10^4)(0.0821 \times 308)^{-1} = 2.6 \times 10^3$$

4. Write the possible rate expressions for each of the following: (10 pts)



$$\text{rate} = -\frac{\Delta[\text{CO}_2]}{\Delta t} = -\frac{1}{2} \frac{\Delta[\text{H}_2\text{O}]}{\Delta t} = \frac{\Delta[\text{CH}_4]}{\Delta t} = \frac{1}{2} \frac{\Delta[\text{O}_2]}{\Delta t}$$



$$\text{rate} = -\frac{1}{3} \frac{\Delta[\text{O}_2]}{\Delta t} = \frac{1}{2} \frac{\Delta[\text{O}_3]}{\Delta t}$$

5. Write the balanced equation that corresponds to the following rate expressions: (5 pts)

$$\text{rate} = -\frac{1}{2} \frac{\Delta[\text{N}_2\text{O}_5]}{\Delta t} = \frac{1}{2} \frac{\Delta[\text{N}_2]}{\Delta t} = \frac{1}{5} \frac{\Delta[\text{O}_2]}{\Delta t}$$



6. Consider the following reaction:



From the following data, use the method of initial rates to determine the rate law and rate constant (including units) for this reaction. Clearly indicate the overall order for the reaction. (15 pts)

trial	$[\text{ClO}_2]$ (M)	$[\text{OH}^-]$ (M)	Initial Rate (M/s)
1	0.060	0.030	0.0248
2	0.020	0.030	0.0083
3	0.020	0.090	0.0247

Dependence on $[\text{ClO}_2]$

$$\frac{0.0248}{0.0083} = 3 \frac{(0.060)^x}{(0.020)^x} = 3^x$$

$$x = 1$$

Dependence on $[\text{OH}^-]$

$$\frac{0.0247}{0.0083} = 3 \frac{(0.090)^y}{(0.030)^y} = 3^y$$

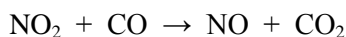
$$y = 1$$

$$\text{rate} = k[\text{ClO}_2][\text{OH}^-]$$

$$k = \frac{\text{rate}}{[\text{ClO}_2][\text{OH}^-]} = \frac{0.0248 \text{ M/s}}{(0.060 \text{ M})(0.030 \text{ M})} = 13.8 \text{ M}^{-1}\text{s}^{-1}$$

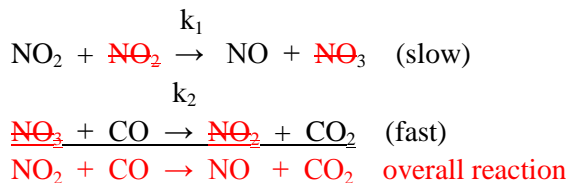
2nd order overall

7. The overall reaction between nitrogen dioxide and carbon monoxide is shown below:



The experimentally observed rate law is: Rate = $k[\text{NO}_2]^2$

The proposed mechanism for this reaction is shown below:



Determine if this is an acceptable mechanism for this reaction. Is there an intermediate present in the reaction? (10 pts)

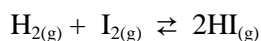
Rate is equal to the slow step.

$$\text{Rate} = k_1[\text{NO}_2][\text{NO}_2] = k_1[\text{NO}_2]^2$$

This is equal to the experimentally observed rate law, therefore this is an acceptable mechanism.

The intermediate is NO_3 .

8. A mixture of 5.75 atm of H_2 and 5.75 atm of I_2 is contained in a 1.0 L vessel at 430 °C. The equilibrium constant, K_p , for this reaction is 54.3. Determine the equilibrium partial pressures of H_2 , I_2 , and HI . (15 pts)



I	5.75	5.75	0
C	-x	-x	+2x
E	5.75-x	5.75-x	2x

$$K_p = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} = 54.3 = \frac{(2x)^2}{(5.75 - x)^2} \quad \text{square root both sides yields} \quad 7.37 = \frac{2x}{5.75 - x}$$

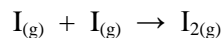
$$\text{Solve for } x: \quad 42.4 - 7.37x = 2x$$

$$x = 4.53$$

$$[\text{H}_2] = [\text{I}_2] = 5.75 \text{ atm} - 4.53 \text{ atm} = 1.22 \text{ atm}$$

$$[\text{HI}] = 2(4.53 \text{ atm}) = 9.06 \text{ atm}$$

9. Iodine atoms combine to form molecular iodine in the gas phase:



This reaction has a rate constant of $7.0 \times 10^9 \text{ M}^{-1}\text{s}^{-1}$ at 23°C . If the initial concentration of I is 0.086 M , calculate the concentration after 2.0 minutes. Also, calculate the half-life for this reaction. (15 pts)

This is a second-order rate constant.

Use the Integrated rate law for a second-order reaction:

$$\frac{1}{[I]_t} = kt + \frac{1}{[I]_0} = (7.0 \times 10^9 \text{ M}^{-1}\text{s}^{-1})(120 \text{ s}) + \frac{1}{0.086 \text{ M}} = 8.4 \times 10^{11} \text{ M}^{-1}$$

$$[I]_t = \frac{1}{8.4 \times 10^{11} \text{ M}^{-1}} = 1.2 \times 10^{-12} \text{ M}$$

$$t_{1/2} = \frac{1}{k[I]_0} = 2 \times 10^{-9} \text{ s}$$

10. Extra Credit. For each reaction, predict the direction the equilibrium would shift for the following reactions: (5 pts)



Increased pressure as volume decreases.

Reaction will shift from right to left



Endothermic reaction = energy is a reactant.

Reaction will shift from left to right