

Chemistry 212
EXAM 1
January 27, 2004

100 (of 100)

KEY

Name _____

Part 1: Multiple Choice. (1 point each, circle only one answer,

1. Consider the following rate law:

$$\text{Rate} = k[\text{A}]^n[\text{B}]^m$$

How are the exponents n and m determined?

- By using the balanced chemical equation
 - By using the subscripts for the chemical formulas
 - By using the coefficients of the chemical formulas
 - By educated guess
 - By experiment**
2. The average rate of disappearance of ozone in the reaction
- $$2\text{O}_3(\text{g}) \rightarrow 3\text{O}_2(\text{g})$$
- is found to be $9.0 \cdot 10^{-3}$ atm/s over a certain interval of time. What is the rate of appearance of O_2 during this interval?
- $1.3 \cdot 10^{-2}$ atm/s**
 - $9.0 \cdot 10^{-3}$ atm/s
 - $6.0 \cdot 10^{-3}$ atm/s
 - $3.0 \cdot 10^{-5}$ atm/s
 - $2.7 \cdot 10^{-5}$ atm/s

3. Consider the reaction $\text{X} \rightarrow \text{Y} + \text{Z}$
Which of the following is a possible rate law?

- Rate = $k[\text{X}]$**
- Rate = $k[\text{Y}]$
- Rate = $k[\text{Y}][\text{Z}]$
- Rate = $k[\text{X}][\text{Y}]$
- Rate = $k[\text{Z}]$

- 4–5. The reaction $2\text{NOBr} \rightarrow 2\text{NO} + \text{Br}_2$ exhibits the rate law

$$\text{Rate} = k[\text{NOBr}]^2 = - \frac{D[\text{NOBr}]}{Dt}$$

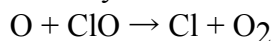
where $k = 1.0 \cdot 10^{-5} \text{ M}^{-1}\text{s}^{-1}$ at 25°C . This reaction is run where the initial concentration of NOBr ($[\text{NOBr}]_0$) is $1.00 \cdot 10^{-1} \text{ M}$.

4. What is one half-life for this experiment?
- $5.0 \cdot 10^{-1} \text{ s}$
 - $6.9 \cdot 10^4 \text{ s}$
 - $1.0 \cdot 10^{-5} \text{ s}$
 - $1.0 \cdot 10^6 \text{ s}$**
 - none of these

5. The $[\text{NO}]$ after 1.00 h has passed is
- $3.5 \cdot 10^{-4} \text{ M}$**
 - $9.9 \cdot 10^{-3} \text{ M}$
 - $9.7 \cdot 10^{-3} \text{ M}$
 - $1.0 \cdot 10^{-3} \text{ M}$
 - none of these

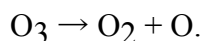
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6. The elementary chemical reaction



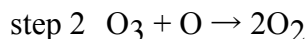
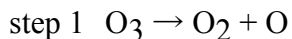
is made pseudo-first order in oxygen atoms by using a large excess of ClO radicals. The rate constant for the reaction is $3.5 \cdot 10^{-11} \text{ cm}^3/\text{molecule} \cdot \text{s}$. If the initial concentration of ClO is $1.0 \cdot 10^{11} \text{ molecules/cm}^3$, how long will it take for the oxygen atoms to decrease to 10.% of their initial concentration?

- 2.4 s
 - 0.017 s
 - $3.2 \cdot 10^{-3} \text{ s}$
 - 0.66 s**
 - 23 s
7. Determine the molecularity of the following elementary reaction:



- unimolecular**
- bimolecular
- termolecular
- tetramolecular
- the molecularity cannot be determined

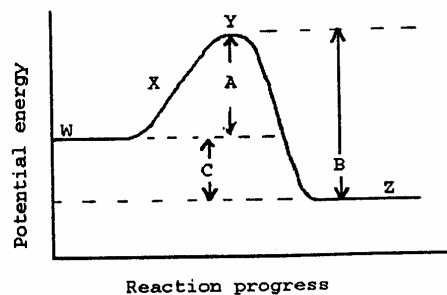
8. The decomposition of ozone may occur through the two-step mechanism shown:



The oxygen atom is considered to be a(n)

- reactant
- product
- catalyst
- reaction intermediate**
- activated complex

The questions 9–12 below refer to the following diagram:

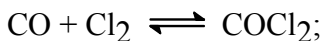


9. Why is this reaction considered to be exothermic?
- Because energy difference B is greater than energy difference C
 - Because energy difference B is greater than energy difference A**
 - Because energy difference A is greater than energy difference C
 - Because energy difference B is greater than energy difference C plus energy difference A
 - Because energy difference A and energy difference C are about equal
10. At what point on the graph is the transition state present?
- point W
 - point X
 - point Y**
 - point Z
 - none of these
11. If the reaction were reversible, would the forward or the reverse reaction have a higher activation energy?
- The diagram shows no indication of any activation energy.
 - The forward and reverse activation energies are equal.
 - The forward activation energy
 - The reverse activation energy**
 - none of these

12. If the reaction were reversible, would the forward or the reverse reaction have a higher activation energy?
- The diagram shows no indication of any activation energy.
 - The forward and reverse activation energies are equal.
 - The forward activation energy
 - The reverse activation energy**
 - none of these
13. What would happen if the kinetic energy of the reactants was not enough to provide the needed activation energy?
- The products would be produced at a lower energy state.
 - The rate of the reaction would tend to increase.
 - The activated complex would convert into products.
 - The reactants would re-form.**
 - The products would form at an unstable energy state.
14. If the half-life of a first-order process is 3.00 minutes, the rate constant for the process is
- 1.50/min.
 - 1.05/min.
 - 4.34/min.
 - 0.405/min.
 - 0.231/min.
15. The reaction
- $$\text{CH}_3\text{CHO}(\text{g}) \rightarrow \text{CH}_4(\text{g}) + \text{CO}(\text{g})$$
- proceeds via the rate expression $\Delta[\text{CO}]/\Delta t = [\text{CH}_3\text{CHO}]^{3/2}$. What is the overall order of the reaction?
- zero-order
 - first-order
 - second-order
 - third-order
 - three-halves-order**
16. The half-life for a first-order reaction at 550 °C is 85 seconds. How long would it take for 23% of the reactant to decompose?
- 0.82 seconds
 - 26 seconds
 - 32 seconds**
 - 44 seconds
 - 180 seconds
17. Calculate the activation energy, E° , for
- $$\text{N}_2\text{O}_5(\text{g}) \rightarrow 2\text{NO}_2(\text{g}) + 1/2 \text{O}_2(\text{g})$$
- given k (at 25 °C) = $3.46 \times 10^{-5}/\text{s}$ and k (at 50 °C) = $1.10 \times 10^{-3}/\text{s}$. $R = 8.3145 \cdot 10^{-3} \text{ kJ/mol}\cdot\text{K}$.
- 231 kJ
 - 111 kJ**
 - 99.3 kJ
 - 76.2 kJ
 - 56.5 kJ
18. The reaction $2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2$ has the following mechanism?
- $$\text{H}_2\text{O}_2 + \text{I}^- \leftrightarrow \text{H}_2\text{O} + \text{IO}^-$$
- $$\text{H}_2\text{O}_2 + \text{IO}^- \leftrightarrow \text{H}_2\text{O} + \text{O}_2 + \text{I}^-$$
- The catalyst in the reaction is:
- H_2O
 - I^-**
 - H_2O_2
 - IO^-
 - there is no catalyst present.
19. The catalyzed pathway in a reaction mechanism has a(n) _____ activation energy and thus causes a(n) _____ reaction rate.
- higher, lower
 - higher, higher
 - lower, higher**
 - lower, steady
 - higher, steady

20. The rate constant k is dependent on
- the concentration of the reactant.
 - the nature of the reactants.
 - the temperature.
 - the order of the reaction.
- none of these
 - one of these
 - two of these**
 - three of these
 - all of these
21. Which of the following statements is true?
- When two opposing processes are proceeding at identical rates, the system is at equilibrium.**
 - Catalysts are an effective means of changing the position of an equilibrium.
 - The concentration of the products equals that of reactants and is constant at equilibrium.
 - An endothermic reaction shifts toward reactants when heat is added to the reaction.
 - None of these statements is true.
22. A chemist prepared a sealed tube with 0.85 atm of PCl_5 at 500 K. The pressure increased as the following reaction occurred. When equilibrium was achieved, the pressure in the tube had increased to 1.25 atm. Calculate K_p .
- $$\text{PCl}_5(\text{g}) \leftrightarrow \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$$
- 0.36**
 - 0.19
 - 0.10
 - 0.047
 - 0.089
23. If $K_c = 0.44$ for the reaction
- $$2\text{NOBr}(\text{g}) \leftrightarrow 2\text{NO}(\text{g}) + \text{Br}_2(\text{g})$$
- at a particular temperature, what is K_c for the following reaction?
- $$\text{NOBr}(\text{g}) \leftrightarrow \text{NO}(\text{g}) + 1/2\text{Br}_2(\text{g})$$
- 0.19
 - 0.22
 - 0.44
 - 0.66**
 - 2.3
24. A flask contains the following system at equilibrium:
- $$\text{Mg}(\text{OH})_2(\text{s}) \leftrightarrow \text{Mg}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq})$$
- Which of the following reagents could be added to increase the solubility of $\text{Mg}(\text{OH})_2$?
- NH_3
 - NaOH
 - HCl**
 - H_2O
 - MgCl_2
25. For the reaction below, what is the expression for K_c ?
- $$2\text{H}_2(\text{g}) + 2\text{FeO}(\text{s}) \leftrightarrow 2\text{Fe}(\text{s}) + 2\text{H}_2\text{O}(\text{g})$$
- $K_c = [\text{Fe}]^2[\text{H}_2\text{O}]^2/[\text{H}_2]^2[\text{FeO}]^2$
 - $K_c = [\text{H}_2\text{O}]^2/[\text{H}_2]^2[\text{FeO}]^2$
 - $K_c = [\text{Fe}]^2/[\text{H}_2]^2[\text{FeO}]^2$
 - $K_c = [\text{H}_2\text{O}]^2/[\text{H}_2]^2$**
 - $K_c = [2\text{Fe}]^2[2\text{H}_2\text{O}]^2/[2\text{H}_2]^2[2\text{FeO}]^2$
26. Consider the reaction $\text{A}(\text{g}) \leftrightarrow 2\text{B}(\text{g})$ where $K_c = 1.5$ at the temperature of the reaction. If 3.0 moles of A and 3.0 moles of B are introduced into a 1.00 liter flask, what change in concentrations (if any) would occur in time?
- [A] increases and [B] increases
 - [A] increases and [B] decreases**
 - [A] decreases and [B] increases
 - [A] decreases and [B] decreases
 - [A] and [B] remain the same

Questions 27–28 refer to the following chemical system:

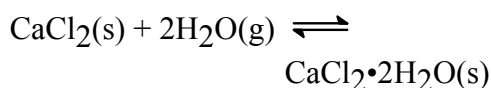


$$K = 4.6 \cdot 10^9 \text{ L/mol.}$$

27. How do the equilibrium concentrations of the reactants compare to the equilibrium concentration of the product?
- They are much smaller.**
 - They are much bigger.
 - They are about the same.
 - They have to be exactly equal.
 - You can't tell from the information given.
28. If the concentration of the product were to double, what would happen to the equilibrium constant?
- It would double its value.
 - It would become half its current value.
 - It would quadruple its value.
 - It would not change its value.**
 - It would depend on the initial conditions of the product.
29. Determine the equilibrium constant for the system $\text{N}_2\text{O}_4 \rightleftharpoons 2\text{NO}_2$ at 25°C . The concentrations are shown here: $[\text{N}_2\text{O}_4] = 4.27 \cdot 10^{-2} \text{ M}$, $[\text{NO}_2] = 1.41 \cdot 10^{-2} \text{ M}$
- 0.33
 - 3.0
 - 0.66
 - 0.05
 - 0.0047**

30. Which of the following is true for a system whose equilibrium constant is relatively small?
- It will take a short time to reach equilibrium.
 - It will take a long time to reach equilibrium.
 - The equilibrium lies to the left.**
 - The equilibrium lies to the right.
 - None of these statements is true.

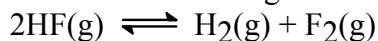
31. Consider the reaction:



The equilibrium constant for the reaction as written is

- $K = \frac{[\text{CaCl}_2 \cdot 2\text{H}_2\text{O}]}{[\text{CaCl}_2][\text{H}_2\text{O}]^2}$
 - $K = \frac{1}{[\text{H}_2\text{O}]^2}$
 - $K = \frac{1}{2[\text{H}_2\text{O}]}$
 - $K = [\text{H}_2\text{O}]^2$
 - $K = \frac{[\text{CaCl}_2 \cdot 2\text{H}_2\text{O}]}{[\text{H}_2\text{O}]^2}$
32. Equilibrium is reached in chemical reactions when:
- the rates of the forward and reverse reactions become equal.**
 - the concentrations of reactants and products become equal.
 - the temperature shows a sharp rise.
 - all chemical reactions stop.
 - the forward reaction stops.

33. Consider the following reaction:

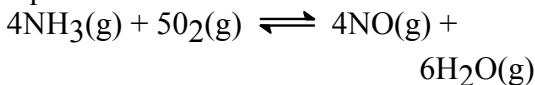


$$(K = 1.00 \cdot 10^{-2})$$

Given 1.00 mole of HF(g), 0.500 mole of H₂(g), and 0.750 mole of F₂(g) are mixed in a 5.00-L flask, determine the reaction quotient, Q , and the net direction to achieve equilibrium.

- $Q = 0.150$; the equilibrium shifts to the right.
 - $Q = 0.375$; the equilibrium shifts to the left.**
 - $Q = 0.150$; the equilibrium shifts to the left.
 - $Q = 0.375$; the equilibrium shifts to the right.
 - $Q = 0.150$; the system is at equilibrium.
34. For a certain reaction at 25.0°C, the value of K is $1.2 \cdot 10^{-3}$. At 50.0°C the value of K is $3.4 \cdot 10^{-1}$. This means that the reaction is
- exothermic.
 - endothermic.**
 - never favorable.
 - More information is needed.
 - None of these (a-d)

Questions 35 and 36 refer to the equilibrium shown here:

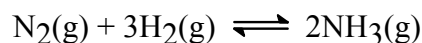


35. What would happen to the system if oxygen were added?
- More ammonia would be produced.
 - More oxygen would be produced.
 - The equilibrium would shift to the right.**
 - The equilibrium would shift to the left.
 - Nothing would happen.

36. What would happen to the system if the pressure were decreased?

- Nothing would happen.
- More oxygen would be produced.
- The water vapor would become liquid water.
- The ammonia concentration would increase.
- The NO concentration would increase.**

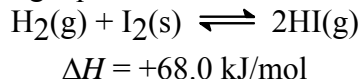
37. Ammonia is prepared industrially by the reaction:



For the reaction, $\Delta H^\circ = -92.2$ kJ and K (at 25°C) = $4.0 \cdot 10^8$. When the temperature of the reaction is increased to 500°C, which of the following is true?

- K for the reaction will be larger at 500°C than at 25°C.
- At equilibrium, more NH₃ is present at 500°C than at 25°C.
- Product formation (at equilibrium) is not favored as the temperature is raised.**
- The reaction of N₂ with H₂ to form ammonia is endothermic.
- None of these is true.

Questions 38 and 39 refer to the following equilibrium:



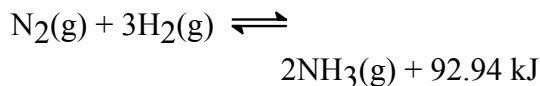
38. The proper K_{eq} expression is:

- $\frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]}$
- $\frac{\sqrt{([\text{H}_2][\text{I}_2])}}{[\text{HI}]^2}$
- $\frac{[\text{HI}]}{\sqrt{([\text{H}_2])}}$
- $\frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$
- $\frac{[\text{HI}]^2}{[\text{H}_2]}$

39. Which of the following statements about the equilibrium is false?

- If the system is heated, the right side is favored.
- This is a heterogeneous equilibrium.
- If the pressure on the system is increased by changing the volume, the left side is favored.
- Adding more $\text{H}_2(\text{g})$ increases the equilibrium constant.**
- Removing HI as it forms forces the equilibrium to the right.

40. Consider the following system at equilibrium:



Which of the following changes will shift the equilibrium to the right?

- increasing the temperature
- decreasing the temperature
- increasing the volume
- decreasing the volume
- removing some NH_3
- adding some NH_3
- removing some N_2
- adding some N_2

- I, IV, VI, VII
- II, III, V, VIII
- I, VI, VIII
- I, III, V, VII
- II, IV, V, VIII**

Show your work Problems (i.e. a number is not an acceptable answer, I must see your calculations for full credit!)

41. The rate constant for a reaction at 40.0°C is exactly three times that at 20.0°C . Calculate the Arrhenius energy of activation for the reaction. ($R = 8.3145 \text{ J/mol}\cdot\text{K}$) (10 points)

$$\ln(k_2/k_1) = E_a/R(1/T_1 - 1/T_2); \text{ Since the rate triples, the } k_2 = 3k_1.$$

$$\ln(3.000) = E_a/8.3145 \text{ J/mol}\cdot\text{K}(1/293 \text{ K} - 1/313 \text{ K}), E_a = 4.33 \cdot 10^4 \text{ J/mol} = 41 \text{ kJ/mol}$$

42. The decomposition of $\text{NO}_2(\text{g})$ occurs by the following bimolecular elementary reaction:



The rate constant at 273 K is $2.3 \cdot 10^{-12}$ L/mol·s, and the activation energy is 111 kJ/mol. How long will it take for the concentration of $\text{NO}_2(\text{g})$ to decrease from an initial partial pressure of 2.5 atm to 1.5 atm at 500. K? Assume ideal gas behavior. (10 points)

We need the value of k at 500. K.

$$\ln(k_2/k_1) = E_a/R(1/T_1 - 1/T_2)$$

$$\ln(k_2/2.3 \times 10^{-12} \text{ L/mol}\cdot\text{s}) = (1.11 \times 10^5 \text{ J/mol})/8.3145 \text{ J/K}\cdot\text{mol}(1/273 \text{ K} - 1/500 \text{ K})$$

$$= 22.2$$

$$k_2/2.3 \times 10^{-12} = e^{22.2}, k_2 = 1.0 \times 10^{-2} \text{ L/mol}\cdot\text{s}$$

Since the decomposition reaction is an elementary reaction, then the rate law can be written using the coefficients in the balanced equation. For this reaction:

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

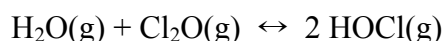
To solve for the time, we must use the integrated rate law for second-order kinetics. The major problem now is converting units so they match.

Rearranging the ideal gas law gives $n/V = P/RT$. Substituting P/RT for concentration units in the second-order integrated rate law equation:

$$\frac{1}{[\text{NO}_2]} = kt + \frac{1}{[\text{NO}_2]_0}, \frac{1}{P/RT} = kt + \frac{1}{P_0/RT}, \frac{RT}{P} \frac{RT}{P_0} = kt, t = \frac{RT}{k} \left(\frac{P_0 - P}{P \times P_0} \right)$$

$$t = \frac{(0.08206 \text{ L}\cdot\text{atm/K}\cdot\text{mol})(500 \text{ K})}{1.0 \times 10^{-2} \text{ L/mol}\cdot\text{s}} \times \left(\frac{2.5 \text{ atm} - 1.5 \text{ atm}}{1.5 \text{ atm} \times 2.5 \text{ atm}} \right) = 1.1 \times 10^3 \text{ s}$$

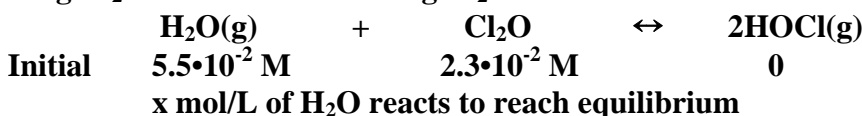
43. At 25 °C, $K = 0.090$ for the reaction



Calculate the concentrations of all species at equilibrium when 1.0 g H_2O and 2.0 g Cl_2O are mixed in a 1.0 L flask. (10 points)

$$1.0 \text{ g H}_2\text{O}/1.0 \text{ L} \times 1 \text{ mol}/18.02 \text{ g H}_2\text{O} = 5.5 \cdot 10^{-2} \text{ mol/L}$$

$$2.0 \text{ g Cl}_2\text{O}/1.0 \text{ L} \times 1 \text{ mol}/86.90 \text{ g Cl}_2\text{O} = 2.3 \cdot 10^{-2} \text{ mol/L}$$



$$K = 0.090 = (2x)^2 / [(5.5 \cdot 10^{-2} - x)(2.3 \cdot 10^{-2} - x)] \quad \text{or}$$

$$1.14 \cdot 10^{-4} - 7.02 \cdot 10^{-3}x + 0.090x^2 = 4x^2$$

$$3.91x^2 + 7.02 \cdot 10^{-3}x - 1.14 \cdot 10^{-4} = 0 \quad (\text{Remember to carry extra significant figures})$$

Solving using the quadratic formula:

$$x = (-7.02 \cdot 10^{-3} \pm (4.93 \cdot 10^{-5} + 1.78 \cdot 10^{-3})^{1/2}) / 7.82 = 4.6 \cdot 10^{-3} \text{ or } -6.4 \cdot 10^{-3}$$

A negative answer makes no physical sense; we can't have less than nothing.

$$\text{So } x = 4.6 \cdot 10^{-3} \text{ M.}$$

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44. The rate of the reaction between hemoglobin (Hb) and carbon monoxide (CO) was studied at 20 °C. The following data were collected with all concentration units in $\mu\text{mol/L}$. (A hemoglobin concentration of 2.21 $\mu\text{mol/l}$ is equal to $2.21 \cdot 10^{-6}$ mol/L)

$[\text{Hb}]_0$ ($\mu\text{mol/L}$)	$[\text{CO}]_0$ ($\mu\text{mol/L}$)	Initial Rate ($\mu\text{mol/L}\cdot\text{s}$)
2.21	1.00	0.619
4.42	1.00	1.24
4.42	3.00	3.71

- a. Determine the orders of this reaction with respect to Hb and CO. (4 points)

Rate = $k[\text{Hb}]^x[\text{CO}]^y$. Comparing the first two experiments, [CO] is unchanged, [Hb] doubles, and the rate doubles. Therefore, the reaction is first order in Hb. Comparing the second and third experiments, [Hb] is unchanged, [CO] triples, and the rates triples. Therefore, $y = 1$, and the reaction is first order in CO.

- b. Calculate the rate law. (2 points)

$$\text{Rate} = k[\text{Hb}][\text{CO}]$$

- c. Calculate the value of the rate constant. (2 points)

From the first experiment:

$$0.619 \mu\text{mol/L}\cdot\text{s} = k(2.21 \mu\text{mol/L})(1.00 \mu\text{mol/L}), k = 0.280 \text{ l}/\mu\text{mol}\cdot\text{s}$$

The second and third experiments give similar values for k, so

$$k_{\text{mean}} = 0.280 \text{ L}/\mu\text{mol}\cdot\text{s}$$

- d. What would be the initial rate for an experiment with $[\text{Hb}]_0 = 3.36 \mu\text{mol/L}$ and $[\text{CO}]_0 = 2.40 \mu\text{mol/L}$? (2 points)

$$\begin{aligned} \text{Rate} &= k[\text{Hb}][\text{CO}] = (0.280 \text{ L}/\mu\text{mol}\cdot\text{s}) \times (3.36 \mu\text{mol/L}) \times (2.40 \mu\text{mol/L}) \\ &= 2.26 \mu\text{mol/L}\cdot\text{s} \end{aligned}$$

45. The rate law for the decomposition of phosphine (PH₃) is

$$\text{Rate} = -\Delta[\text{PH}_3]/\Delta t = k[\text{PH}_3]$$

It takes 120. s for 1.00 M PH₃ to decrease to 0.250 M. How much time is required for 2.00 M PH₃ to decrease to a concentration of 0.350 M. (10 points)

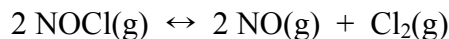
For a first order reaction, the integrated rate law is: $\ln([A]/[A]_0) = -kt$. Solving for k:

$$\ln((0.250 \text{ mol/L})/(1.00 \text{ mol/L})) = -k \times 120 \text{ s}, k = 0.0116 \text{ s}^{-1}$$

$$\ln((0.350 \text{ mol/L})/(2.00 \text{ mol/L})) = -0.0116 \text{ s}^{-1} \times t$$

$$t = 150 \text{ s}$$

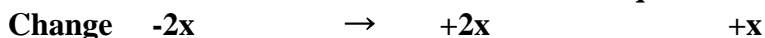
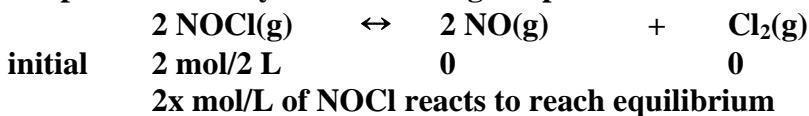
46. At 35°C, $K = 1.6 \cdot 10^{-5}$ for the reaction



Calculate the concentrations of all species at equilibrium for each of the following original mixtures.

- a. 2.0 mol pure NOCl in a 2.0 L flask. (5 points)

The reaction must proceed to products to reach equilibrium since only reactants are present initially. Summarizing the problem in a table:



$$K = 1.6 \cdot 10^{-5} = [\text{NO}]^2[\text{Cl}_2]/[\text{NOCl}]^2 = (2x)^2(x)/(1.0 - 2x)^2$$

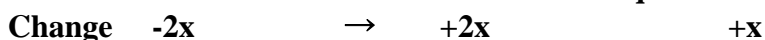
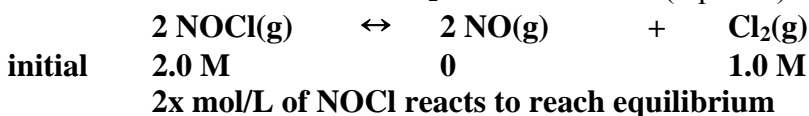
If we assume that $1.0 - 2x \approx 1.0$ (from the size of L, we know that not much reaction will occur so x is small), then:

$$1.6 \cdot 10^{-5} = 4x^3/1.0^2, x = 1.6 \cdot 10^{-2}$$

$$[\text{NO}] = 2x = 0.032 \text{ M}; [\text{Cl}_2] = x = 0.016 \text{ M}; [\text{NOCl}] = 1.0 - 2x = 0.97 \text{ M} \approx 1.0 \text{ M}$$

Assumption is valid.

- b. 2.0 mol NOCl and 1.0 mol Cl₂ in a 1.0 L flask. (5 points)



$$K = 1.6 \cdot 10^{-5} = [\text{NO}]^2[\text{Cl}_2]/[\text{NOCl}]^2 = (2x)^2(1.0 + x)/(2.0 - 2x)^2 \approx 4x^2/4.0 \text{ (assuming that } x \ll 1.0, \text{ as explained above.)}$$

Solving: $x = 4.0 \cdot 10^{-3}$, Assumption good.

$$[\text{Cl}_2] = 1.0 + x = 1.0 \text{ M}; [\text{NO}] = 2(4.0 \cdot 10^{-3}) = 8.0 \cdot 10^{-3}; [\text{NOCl}] = 2.0 \text{ M}$$