

CHM 101 Mock Exam #4

1. Select the correct equilibrium constant, K_{eq} , expression for $3 \text{H}_2(\text{g}) + \text{N}_2(\text{g}) \rightleftharpoons 2 \text{NH}_3(\text{g})$.

A. $[\text{NH}_3]^2 / ([\text{H}_2][\text{N}_2])$
B. $[\text{NH}_3]^2 / ([\text{H}_2]^3[\text{N}_2])$
C. $[\text{NH}_3]^2[\text{H}_2]^3[\text{N}_2]$
D. $[\text{NH}_3] / ([\text{H}_2][\text{N}_2])$
E. $[\text{NH}_3]^2 / ([\text{H}_2]^3[\text{N}_2]^2)$

For $a \text{A} + b \text{B} \rightleftharpoons c \text{C} + d \text{D}$, $K_{eq} = ([\text{C}]^c[\text{D}]^d)/([\text{A}]^a[\text{B}]^b)$

2. Consider $2 \text{HI}(\text{g}) \rightleftharpoons \text{H}_2(\text{g}) + \text{I}_2(\text{g})$ at a specific temperature. If $[\text{HI}] = 1.17 \text{ M}$, $[\text{H}_2] = 1.37 \text{ M}$, and $[\text{I}_2] = 0.100 \text{ M}$ at equilibrium, what is the value of the equilibrium constant, K_{eq} , under these conditions?

A. 0.100
B. 1.17
C. 1.26
D. 8.54
E. 10.0

For $2 \text{HI} \rightleftharpoons \text{H}_2 + \text{I}_2$, $K_{eq} = ([\text{H}_2][\text{I}_2]) / [\text{HI}]^2 = [(1.37)(0.100)] / (1.17)^2 = 0.100$

3. Given below are equilibrium constants for different chemical reactions. Which one corresponds to a reaction that *least* favors the formation of products (i.e., which one is the most reactants-favored reaction)?

A. 0.025
B. 1.0
C. 5.2
D. 2.4×10^5
E. 1.8×10^{-6}

K (or K_{eq}) equals to 1 (not zero) means that the amounts of reactants and products at equilibrium are equal. Therefore, if K is much greater than 1 ($K \gg 1$), it means that there are more products than reactants at equilibrium (products are favored). $K \ll 1$ means that there are more reactants than products (reactants are favored). 0.0000018 is therefore the best answer to this question because even 0.025, which is less than 1, does not indicate a K value that is as reactants-favored as 0.0000018. 240000, on the other hand, is the most products-favored equilibrium constant as opposed to 5.2.

4. Consider the following exothermic reaction initially at equilibrium: $\text{NO}(\text{g}) + \text{Cl}_2(\text{g}) \rightleftharpoons \text{NOCl}_2(\text{g})$. What change (shifting) will be observed if the amount of $\text{NO}(\text{g})$ in the reaction mixture were to increase?

A. Right shifting; the concentrations of NO and Cl_2 will increase
B. No shifting; the concentration of NOCl_2 will remain unchanged
C. Right shifting; the concentration of NOCl_2 will increase
D. No shifting; the concentrations of NO and Cl_2 will remain unchanged
E. Left shifting; the concentration of NOCl_2 will decrease

5. Consider the following exothermic reaction initially at equilibrium: $\text{NO}(\text{g}) + \text{Cl}_2(\text{g}) \rightleftharpoons \text{NOCl}_2(\text{g})$. What change (shifting) will be observed if the temperature of the reaction mixture were to decrease?

A. Right shifting; the concentrations of NO and Cl_2 will increase
B. No shifting; the concentration of NOCl_2 will remain unchanged
C. Right shifting; the concentration of NOCl_2 will increase
D. No shifting; the concentrations of NO and Cl_2 will remain unchanged
E. Left shifting; the concentration of NOCl_2 will decrease

Both Questions 4 and 5 are about **Le Chatelier's principle**. They have the same answer because in both cases, the reaction $\text{NO} + \text{Cl}_2 \rightleftharpoons \text{NOCl}_2$ is subject to stresses that would require a "right shift" meaning "producing more products" to offset the impacts of the stresses. In Question 4, the stress is an increase in the amount of a reactant (too much reactant). In Question 5, however, the stress is a decrease in

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