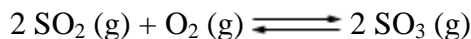


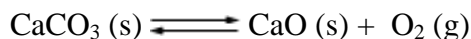
Equilibrium Constant - Practice Problems for Assignment 5

1. Consider the following reaction



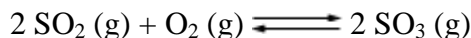
Write the equilibrium expression, K_c .

2. Consider the following reaction



Write the equilibrium expression, K_c .

3. Consider the following reaction



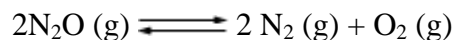
Write the equilibrium expression, K_p .

4. Consider the following reaction



Write the equilibrium expression, K_p .

5. Consider the decomposition of nitrous oxide, laughing gas,



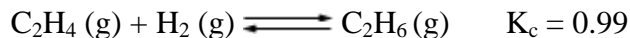
At 25°C, K_c is 7.3×10^{34} .

- (a) Based on the information given, what can you say about the rate of decomposition of the reaction?
(b) Based on the information given, does nitrous oxide have a tendency to decompose into nitrogen and oxygen?
(c) What is the value of K_p for the reaction at 25°C?
6. Consider the following reaction



Calculate the value of the equilibrium constant, K_c , for the above system, if 0.1908 moles of CO_2 , 0.0908 moles of H_2 , 0.0092 moles of CO , and 0.0092 moles of H_2O vapour were present in a 2.00 L reaction vessel at equilibrium.

7. Consider the following reaction



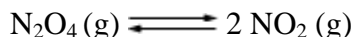
What is the concentration for each substance at equilibrium if the initial concentration of ethene, $\text{C}_2\text{H}_4(\text{g})$, is 0.335 M and that of hydrogen is 0.526 M?

8. Consider the following reaction



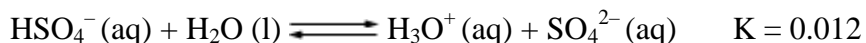
Determine the value of the equilibrium constant, K_c , for the reaction. Initially, a mixture of 0.100 M NO, 0.050 M H_2 , 0.100 M H_2O was allowed to reach equilibrium (initially there was no N_2). At equilibrium the concentration of NO was found to be 0.062 M.

9. Consider the following reaction



A reaction flask is charged with 3.00 atm of dinitrogen tetroxide gas and 2.00 atm of nitrogen dioxide gas. At 25°C, the gases are allowed to reach equilibrium. The pressure of the nitrogen dioxide was found to have decreased by 0.952 atm. Estimate the value of K_p for this system.

10. Consider the following reaction. The initial concentrations are $[\text{HSO}_4^-] = 0.50 \text{ M}$, $[\text{H}_3\text{O}^+] = 0.020 \text{ M}$, $[\text{SO}_4^{2-}] = 0.060 \text{ M}$.



- (a) Which way would the reaction shift to reach equilibrium?
(b) What are the equilibrium concentrations of the products and reactants.

Answers:

$$1. K_c = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2[\text{O}_2]}$$

$$2. K_c = [\text{O}_2]$$

$$3. K_p = \frac{P_{\text{SO}_3}^2}{P_{\text{SO}_2}^2 P_{\text{O}_2}}$$

$$4. K_p = \frac{P_{\text{H}_2} P_{\text{CO}}}{P_{\text{H}_2\text{O}}}$$

5. (a) Based on the information given, you cannot predict the rate of decomposition of nitrous oxide.
(b) From the value of the K_{eq} , nitrous oxide has a strong tendency to decompose into nitrogen and oxygen.
(c) $K_p = 1.8 \times 10^{36}$

6. $[\text{CO}_2] = 0.1908 \text{ mol CO}_2/2.00 \text{ L} = 0.0954 \text{ M}$
 $[\text{H}_2] = 0.0454 \text{ M}$
 $[\text{CO}] = 0.0046 \text{ M}$
 $[\text{H}_2\text{O}] = 0.0046 \text{ M}$

$$K = \frac{(0.0046)(0.0046)}{(0.0954)(0.0454)} = 0.0049 \text{ or } 4.9 \times 10^{-3}$$

7.

	C_2H_4	H_2	C_2H_6
[I]	0.335	0.526	0
[C]	-x	-x	+x
[E]	0.335 - x	0.526 - x	+x

$$K = \frac{x}{(0.335 - x)(0.526 - x)} = 0.0995 \text{ or } 1.77^*$$

* $x=1.77$ is not possible because the concentration of C_2H_4 will result in a negative value.

$$[\text{C}_2\text{H}_4] = 0.236 \text{ M}$$

$$[\text{H}_2] = 0.526 - x = 0.526 - 0.0995 = 0.427 \text{ M}$$

$$[\text{C}_2\text{H}_6] = 0.0995 \text{ M}$$

8.

	NO	H ₂	N ₂	H ₂ O
[I]	0.100	0.0500	0	0.100
[C]	-2x	-2x	+x	+2x
[E]	0.062			
From ICE table	2x = 0.038			

Therefore, substitute for x and calculate [E] for each species:

	NO	H ₂	N ₂	H ₂ O
[I]	0.100	0.0500	0	0.100
[C]	-0.038	-0.038	+0.019	+0.038
[E]	0.062	0.012	0.019	0.138

$$K = \frac{(0.019)(0.138)^2}{(0.062)^2(0.012)^2} = 6.5 \times 10^2$$

9.

	N ₂ O ₄	NO ₂
[I]	3.00	2.00
[C]	+x	-2x = -0.952
[E]		
From ICE table	x = 0.952/2	

Therefore, substitute for x and calculate [E] for each species:

	N ₂ O ₄	NO ₂
[I]	3.00	2.00
[C]	+0.476	-0.952
[E]	3.476	1.048

$$K = \frac{(1.048)^2}{(3.476)} = 0.316$$

10. (a) Use the trial K_{eq} , Q , to determine the reaction direction.

$$Q = \frac{(0.020)(0.060)}{(0.50)} = 0.0024$$

$Q < K_{eq}$, therefore, equilibrium will shift to the right to produce more products.

(b)

	HSO_4^-	H_3O^+	SO_4^{2-}
[I]	0.50	0.020	0.060
[C]	-x	+x	+x
[E]	0.50 - x	0.020 + x	0.060 + x

$$K = \frac{(0.020 + x)(0.060 + x)}{(0.50 - x)}$$

To solve, need to use the quadratic equation

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$x = 0.0372 \text{ or } -0.129 *$$

For $x = 0.0372$,

$$[\text{HSO}_4^-] = 0.46 \text{ M}; [\text{H}_3\text{O}^+] = 0.057 \text{ M}; [\text{SO}_4^{2-}] = 0.097 \text{ M}$$

* For $x = -0.129$,

$$[\text{HSO}_4^-] = 0.63 \text{ M}; [\text{H}_3\text{O}^+] = -0.109 \text{ M}; [\text{SO}_4^{2-}] = -0.069 \text{ M}$$

it yields negative concentrations.

Therefore, the correct equilibrium concentrations are:

$$[\text{HSO}_4^-] = 0.46 \text{ M}; [\text{H}_3\text{O}^+] = 0.057 \text{ M}; [\text{SO}_4^{2-}] = 0.097 \text{ M}$$