

# 16 • Chemical Equilibrium

## PRACTICE TEST

1. Consider the reaction system,  
 $\text{CoO}(s) + \text{H}_2(g) \rightleftharpoons \text{Co}(s) + \text{H}_2\text{O}(g)$   $\frac{[\text{H}_2\text{O}]}{[\text{H}_2]}$
- The equilibrium constant expression is
- a)  $\frac{[\text{CoO}][\text{H}_2]}{[\text{Co}][\text{H}_2\text{O}]}$       d)  $\frac{[\text{H}_2]}{[\text{H}_2\text{O}]}$   
 b)  $\frac{[\text{Co}][\text{H}_2\text{O}]}{[\text{CoO}][\text{H}_2]}$       e)  $\frac{[\text{H}_2\text{O}]}{[\text{H}_2]}$   
 c)  $\frac{[\text{Co}][\text{H}_2\text{O}]}{[\text{H}_2]}$

2. Given the equilibrium,  
 $2\text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{SO}_3(g)$ , if this equilibrium is established by beginning with equal number of moles of  $\text{SO}_2$  and  $\text{O}_2$  in a 1.0 Liter bulb, then the following **must** be true at equilibrium:
- a)  $[\text{SO}_2] = [\text{SO}_3]$       d)  $[\text{SO}_2] < [\text{O}_2]$   
 b)  $2[\text{SO}_2] = 2[\text{SO}_3]$       e)  $[\text{SO}_2] > [\text{O}_2]$   
 c)  $[\text{SO}_2] = [\text{O}_2]$

Questions 3 & 4 refer to the following:

At a given temperature, 0.300 mole  $\text{NO}$ , 0.200 mol  $\text{Cl}_2$  and 0.500 mol  $\text{ClNO}$  were placed in a 25.0 Liter container. The following equilibrium is established:  
 $2\text{ClNO}(g) \rightleftharpoons 2\text{NO}(g) + \text{Cl}_2(g)$

3. At equilibrium, 0.600 mol of  $\text{ClNO}$  was present. The number of **moles** of  $\text{Cl}_2$  present at equilibrium is
- a) 0.050      d) 0.200  
 b) 0.100      e) 0.250  
 c) 0.150

*as soon as even 1 rxn occurs  $[\text{SO}_2] \neq [\text{O}_2]$*

	$2\text{ClNO}$	$\rightleftharpoons$	$2\text{NO} + \text{Cl}_2$	
I	0.500		0.300      0.200	
C	+0.100		-0.100      -0.050	
E	0.600		0.200      0.150	

4. The equilibrium constant,  $K_c$ , is:
- a)  $4.45 \times 10^{-4}$       d) 0.167  
 b)  $6.67 \times 10^{-4}$       e) 1500  
 c) 0.111
- $K_c = \frac{(0.200)^2 (0.150)}{(0.600)^2}$
5. At  $985^\circ\text{C}$ , the equilibrium constant for the reaction,  
 $\text{H}_2(g) + \text{CO}_2(g) \rightleftharpoons \text{H}_2\text{O}(g) + \text{CO}(g)$  is 1.63. What is the equilibrium constant for the reverse reaction?
- a) 1.63      d) 0.613  
 b) 0.815      e) 1.00  
 c) 2.66

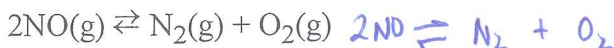
6. What is the relationship between  $K_p$  and  $K_c$  for the reaction,  $2\text{ICl}(g) \rightleftharpoons \text{I}_2(g) + \text{Cl}_2(g)$ ?
- a)  $K_p = K_c(RT)^{-1}$       d)  $K_p = K_c$   
 b)  $K_p = K_c(RT)$       e)  $K_p = K_c(2RT)$   
 c)  $K_p = K_c(RT)^2$

- $\Delta n = 0$   
 $\Delta n = 1 - 2 = -1$
7. For the reaction  $2\text{NO}_2(g) \rightleftharpoons \text{N}_2\text{O}_4(g)$ ,  $K_p$  at  $25^\circ\text{C}$  is 7.3, when all partial pressures are expressed in atmospheres. What is  $K_c$  for this reaction? [ $R=0.0821 \text{ L}\cdot\text{atm}\cdot\text{mol}^{-1}\cdot\text{K}^{-1}$ ]
- a) 4270      d) 179  
 b) 0.0119      e) 2.06  
 c) 0.291

$K_p = K_c (RT)^{\Delta n}$

$K_c = \frac{K_p}{(RT)^{\Delta n}} = \frac{7.3}{(0.0821)^{-1} (298)^{-1}} = 178.6$

8. 0.200 mol NO is placed in a one liter flask at 2273 K. After equilibrium is attained, 0.0863 mol N<sub>2</sub> and 0.0863 mol O<sub>2</sub> are present. What is K<sub>c</sub> for this reaction?



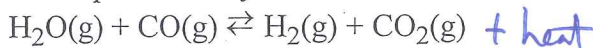
- a) 9.92      d) 39.7  
 b) 3.15      e) 0.576  
 c) 0.0372

*Handwritten:* I .200 0 0  
 C -0.173 +0.0863 +0.0863  
 E 0.0274 0.0863 0.0863  
 $K_c = \frac{(0.0863)(0.0863)}{(0.0274)^2}$

9. At 25°C, 0.11 mole of N<sub>2</sub>O<sub>4</sub> reacts to form 0.10 mol of N<sub>2</sub>O<sub>4</sub> and 0.02 mole of NO<sub>2</sub>. At 90°C, 0.11 mole of N<sub>2</sub>O<sub>4</sub> forms 0.050 mole of N<sub>2</sub>O<sub>4</sub> and 0.12 mole of NO<sub>2</sub>. From these data we can conclude

- a) N<sub>2</sub>O<sub>4</sub> molecules react by a second order rate law.  
 b) N<sub>2</sub>O<sub>4</sub> molecules react by a first order rate law.  
 c) the reaction is ~~exo~~<sup>endo</sup>thermic.  
 d) N<sub>2</sub>O<sub>4</sub> molecules react faster at 25°C than at 90°C.  
 e) the equilibrium constant for the reaction above increases with an increase in temperature. *more products at 25°C*

10. For the equilibrium system



$$\Delta H = -42 \text{ kJ/mol}$$

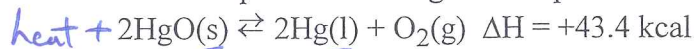
K<sub>c</sub> equals 0.62 at 1260 K. If 0.10 mole each of H<sub>2</sub>O, CO, H<sub>2</sub> and CO<sub>2</sub> (each at 1260 K) were placed in a 1.0-Liter flask at 1260 K, when the system came to equilibrium...

	The temperature would	The mass of CO would
a)	decrease	increase
b)	decrease	decrease
c)	remain constant	increase
d)	increase	decrease
e)	increase	increase

*Handwritten:*  $Q = \frac{(0.10)(0.10)}{(0.10)(0.10)} = 1$        $K_c = 0.62$   
 $Q > K_c$  shift left

11. For the reaction system,  
 $\text{N}_2(g) + 3\text{H}_2(g) \rightleftharpoons 2\text{NH}_3(g) + \text{heat}$   
 the conditions that would favor maximum conversion of the reactants to products would be *fewer moles of gas on product side*
- a) high temperature and high pressure  
 b) high temperature, pressure unimportant  
 c) high temperature and low pressure  
 d) low temperature and high pressure  
 e) low temperature and low pressure

12. Solid HgO, liquid Hg, and gaseous O<sub>2</sub> are placed in a glass bulb and are allowed to reach equilibrium at a given temperature.



The mass of HgO in the bulb could be increased by

- a) adding more Hg. *No change*  
 b) removing some O<sub>2</sub>. *→*  
 c) reducing the volume of the bulb.  
 d) increasing the temperature. *→*  
 e) removing some Hg. *No change*

**Answers:** (Please use CAPITAL letters)

1.	E	7.	D
2.	D	8.	A
3.	C	9.	E
4.	B	10.	A
5.	D	11.	D
6.	D	12.	C

*Handwritten:* GO VIKINGS!