South Pasadena • AP Chemistry

Name CALVIN
Period ___ Date ___/___/___

# 16 • Chemical Equilibrium

**PRACTICE TEST**

1. Consider the reaction system,
$\cancel{CoO(s)} + H_2(g) \rightleftarrows \cancel{Co(s)} + H_2O(g)$.

$\frac{[H_2O]}{[H_2]}$

The equilibrium constant expression is

a) $\frac{[CoO][H_2]}{[Co][H_2O]}$ d) $\frac{[H_2]}{[H_2O]}$

b) $\frac{[Co][H_2O]}{[CoO][H_2]}$ e) $\frac{[H_2O]}{[H_2]}$

c) $\frac{[Co][H_2O]}{[H_2]}$

2. Given the equilibrium,
$2SO_2(g) + O_2(g) \rightleftarrows 2SO_3(g)$, if this equilibrium is established by beginning with equal number of moles of $SO_2$ and $O_2$ in a 1.0 Liter bulb, then the following ***must*** be true at equilibrium:

a) $[SO_2] = [SO_3]$ d) $[SO_2] < [O_2]$

b) $2[SO_2] = 2[SO_3]$ e) $[SO_2] > [O_2]$

c) $[SO_2] = [O_2]$

as soon as even 1 RXN occurs $[SO_2] \neq [O_2]$

**Questions 3 & 4** refer to the following:

At a given temperature, 0.300 mole NO, 0.200 mol $Cl_2$ and 0.500 mol ClNO were placed in a 25.0 Liter container. The following equilibrium is established:
$2ClNO(g) \rightleftarrows 2NO(g) + Cl_2(g)$

3. At equilibrium, 0.600 mol of ClNO was present. The number of ***moles*** of $Cl_2$ present at equilibrium is

a) 0.050 d) 0.200

b) 0.100 e) 0.250

c) 0.150

| | $2ClNO$ | $\rightleftarrows$ $2NO$ + | $Cl_2$ |
|---|---|---|---|
| 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

$$K_c = \frac{\left(\frac{0.200}{25}\right)^2\left(\frac{0.150}{25}\right)}{\left(\frac{0.600}{25}\right)^2}$$

5. At 985°C, the equilibrium constant for the reaction,
$H_2(g) + CO_2(g) \rightleftarrows H_2O(g) + CO(g)$
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

$$K' = \frac{1}{K} = \frac{1}{1.63}$$

6. What is the relationship between $K_p$ and $K_c$ for the reaction, $2ICl(g) \rightleftarrows I_2(g) + Cl_2(g)$?

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 $2NO_2(g) \rightleftarrows N_2O_4(g)$, $K_p$ at ~~25°C~~ 298K is 7.3, when all partial pressures are expressed in atmospheres. What is $K_c$ for this reaction? [R=0.0821 L·atm·mol$^{-1}$·K$^{-1}$]

a) 4270 d) 179

b) 0.0119 e) 2.06

c) 0.291

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

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

8. 0.200 mol NO is placed in a one liter flask at 2273 K. After equilibrium is attained, 0.0863 mol $N_2$ and 0.0863 mol $O_2$ are present. What is $K_c$ for this reaction?

$$2NO(g) \rightleftarrows N_2(g) + O_2(g)$$

| | $2NO$ | $\rightleftarrows$ $N_2$ | $+$ $O_2$ |
|---|---|---|---|
| I | .200 | 0 | 0 |
| C | −0.173 | +0.0863 | +0.0863 |
| E | 0.0274 | 0.0863 | 0.0863 |

$K_c = \dfrac{(0.0863)(0.0863)}{(0.0274)^2}$

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

9. $N_2O_4(g) \rightleftarrows 2\ NO_2(g)$

0.10 — 0.02 @ 25°C
.05 — 0.12 @ 90°C

At 25°C, 0.11 mole of $N_2O_4$ reacts to form 0.10 mol of $N_2O_4$ and 0.02 mole of $NO_2$. At 90°C, 0.11 mole of $N_2O_4$ forms 0.050 mole of $N_2O_4$ and 0.12 mole of $NO_2$. From these data we can conclude

a) $N_2O_4$ molecules react by a second order rate law.
b) $N_2O_4$ molecules react by a first order rate law.
c) the reaction is ~~exothermic~~ endo
d) $N_2O_4$ molecules react faster at 25°C than at 90°C.
e) the equilibrium constant for the reaction above increases with an increase ~~in~~ temperature. 90°C than more products at 25°C

10. For the equilibrium system

$$H_2O(g) + CO(g) \rightleftarrows H_2(g) + CO_2(g)\ + \text{heat}$$

$$\Delta H = -42\ kJ/mol$$

$K_c$ equals 0.62 at 1260 K. If 0.10 mole each of $H_2O$, CO, $H_2$ and $CO_2$ (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 |

$Q = \dfrac{(0.10)(0.10)}{(0.10)(0.10)} = 1$ $\quad K_c = 0.62$

$Q > K_c$ shift left

11. For the reaction system,

$$N_2(g) + 3H_2(g) \rightleftarrows 2NH_3(g) + \text{heat}$$

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 $O_2$ are placed in a glass bulb and are allowed to reach equilibrium at a given temperature.

$$\text{heat} + 2HgO(s) \rightleftarrows 2Hg(l) + O_2(g) \quad \Delta H = +43.4\ kcal$$

The mass of HgO in the bulb could be increased by

a) adding more Hg. No change
b) removing some $O_2$.
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 |

Go VIKINGS!

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