Equilibrium Practice

$$P_4(s) + 5\,O_2(g) \rightleftharpoons P_4O_{10}(s)$$

1. What is the equilibrium expression for this reaction?

(A) $K_c = [P_4O_{10}] / [P_4] [O_2]^5$
(B) $K_c = [P_4O_{10}] / 5 [P_4] [O_2]$
(C) $K_c = [O_2]^5$
(D) $K_c = 1 / [O_2]^5$

$$Fe_3O_4(s) + 4\,H_2(g) \rightleftharpoons 3\,Fe(s) + 4\,H_2O(g)$$
$\Delta H > 0$ Endo

2. For this reaction at equilibrium, which changes will increase the quantity of Fe(s)?

1. increasing temperature YES
2. decreasing temperature NO
3. adding $Fe_3O_4(s)$ No

(A) 1 only
(B) 1 and 2 only
(C) 2 and 3 only
(D) 1,2, and 3

3. Which reaction characteristics are changing by the addition of a catalyst to a reaction to a reaction at constant temperature?

1. activation energy
2. equilibrium concentrations NO
ΔH 3. reaction enthalpy NO

(A) 1 only
(B) 3 only
(C) 1 and 2 only
(D) 1, 2, and 3

4. Which reaction characteristics will be affected by a change in temperature?

1. value of equilibrium constant
2. equilibrium concentrations

Both!

(A) 1 only
(B) 2 only
(C) 1 and 2 only
(D) neither 1 nor 2

5. What is the relationship between the equilibrium constant (K_c) of a reaction and the rate constants for the foward (k_f) and backward (k_b) steps? Rate constants

(A) $K_c = k_f\, k_b$
(B) $K_c = k_b / k_f$
(C) $K_c = k_f / k_b$
(D) $K_c = 1 / (k_f k_b)$

6. Xenon tetrafluoride, XeF_4, can be prepared by heating Xe and F_2 together according to this equation.

$$Xe(g) + 2F_2(g) \rightleftharpoons XeF_4(g)$$

What is the equilibrium expression for this reaction?

$\frac{[\text{Products}]}{[\text{Reactants}]}$

(A) $K = [XeF_4] / ([Xe] [F_2])$
(B) $K = [XeF_4] / (2[Xe] [F_2])$
(C) $K = [XeF_4] / ([Xe] [F_2]^2)$
(D) $K = ([Xe] [F_2]) / [XeF_4]$

7. What is the equilibrium expression for the decomposition of ammonium carbamate, $NH_4CO_2NH_2$, that occurs according to this equation:

$$NH_4CO_2NH_2(s) \rightleftharpoons 2\,NH_3(g) + CO_2(g)$$

(A) $K = [NH_3][CO_2]$
(B) $K = [NH_3]^2[CO_2]$
(C) $K = [NH_3][CO_2] / [NH_4CO_2NH_2]$
(D) $K = [NH_3]^2[CO_2] / [NH_4CO_2NH_2]$

8. Which factors will affect both the position of equilibrium and the value of the equilibrium constant for this reaction? The $\Delta H = -92$ kJ

$$N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g)$$

(A) increasing the volume of the container
(B) adding N_2
(C) removing NH_3
(D) lowering the temperature $K_c \uparrow$

AP Chemistry

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Ch 13 Chemical Equilibrium

PRACTICE TEST

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

   The equilibrium constant expression is

   a) $\frac{[CoO][H_2]}{[Co][H_2O]}$
   b) $\frac{[Co][H_2O]}{[CoO][H_2]}$
   c) $\frac{[Co][H_2O]}{[H_2]}$
   d) $\frac{[H_2]}{[H_2O]}$
   e) $\frac{[H_2O]}{[H_2]}$ $\frac{\boxed{Prod}}{\boxed{React}}$

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: need $2 \times SO_2$ so it gets used up first

   a) $[SO_2] = [SO_3]$
   b) $2[SO_2] = 2[SO_3]$
   c) $[SO_2] = [O_2]$
   d) $[SO_2] < [O_2]$
   e) $[SO_2] > [O_2]$

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:

$2ClNO(g) \rightleftarrows 2NO(g) + Cl_2(g)$

ICE	$2ClNO$	$2NO$	Cl_2
I	.5	.3	.2
C	+.1	−.1	−.05
E	=.6	.2	.150

3. At equilibrium, 0.600 mol of ClNO was present. The number of <u>*moles*</u> of Cl_2 present at equilibrium is

   a) 0.050
   b) 0.100
   c) 0.150
   d) 0.200
   e) 0.250

4. The equilibrium constant, K_c, is: $\frac{\left(\frac{.2}{25}\right)^2\left(\frac{.15}{25}\right)}{\left(\frac{.6}{25}\right)^2}$

   a) 4.45×10^{-4}
   b) 6.67×10^{-4}
   c) 0.111
   d) 0.167
   e) 1500

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

   a) 1.63
   b) 0.815
   c) 2.66
   d) 0.613
   e) 1.00

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

   a) $K_p = K_c(RT)^{-1}$
   b) $K_p = K_c(RT)$
   c) $K_p = K_c(RT)^2$
   d) $K_p = K_c$
   e) $K_p = K_c(2RT)$

   $K_p = K_c(RT)^{\Delta n}$ $\Delta n = 0$

$\Delta n = 1 - 2 = -1$

7. For the reaction $2NO_2(g) \rightleftarrows N_2O_4(g)$, K_p at 25°C 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}] +273

   a) 4270
   b) 0.0119
   c) 0.291
   d) 179
   e) 2.06

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

$K_c = \frac{7.3}{[(0.0821)(298)]^{-1}}$

$= 178.6$

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

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

$2NO \rightleftarrows N_2 + O_2$
0.2 0 0
$-2(0.0863)$ $+.0863$ $+.0863$
0.0274 .0863 .0863
$\frac{[N_2][O_2]}{[NO]^2}$

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

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

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 ↑T↑ 90°C, 0.11 mole of N_2O_4 forms 0.050 mole Rxn 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. No → 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.

10. For the equilibrium system

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

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

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

$Q = \frac{(.10)(.10)}{(.10)(.10)} = 1$ $Q > K$ will shift to left (make reactants)

11. For the reaction system, Exo

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

4 mol gas 2 mol gas

the conditions that would favor maximum conversion of the reactants to products would be

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.

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

The mass of HgO in the bulb could be increased by

a) adding more Hg. No effect
b) removing some O_2. No shift right
c) reducing the volume of the bulb.
d) increasing the temperature. NO →
e) removing some Hg. No effect

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

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

AP Chemistry

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Ch 13 Chemical Equilibria

PROBLEM SET #2

1. Consider the equilibrium: $2\ SO_2(g) + O_2(g) \rightleftarrows 2\ SO_3(g)$ $K_c = 4.36\ \underline{M}^{-1}$
Calculate the value of "Q" for a situation in which the concentrations are $[SO_2] = 2.00\ \underline{M}$, $[O_2] = 1.50\ \underline{M}$, and $[SO_3] = 1.25\ \underline{M}$.

$$Q = \frac{[SO_3]^2}{[SO_2]^2[O_2]} = \frac{(1.25\,M)^2}{(2.00\,M)^2(1.50\,M)} = 0.260$$

Does this mixture shift toward the reactants or products to reach equilibrium? Toward Products

$Q < K$ ∴ shifts toward products – numerator increase and denominator decrease

2. Study the discussion in your textbook about converting K_c and K_p. Write the K_p expression for the reaction in question 1 and calculate its value at 0°C. Remember, R = 0.0821 L·atm/mol·K. Temp = 273

$$K_p = K_c(RT)^{\Delta n} \qquad \Delta n = 2 - 3 = -1$$

$$= (4.36)[(0.0821)(273)]^{-1} = 0.1945$$

3. Consider the equilibrium $PCl_3(g) + Cl_2(g) \rightleftharpoons PCl_5(g)$.
How would the following changes affect the partial pressures of each gas at equilibrium?

	$PCl_3(g)$	$Cl_2(g)$		$PCl_5(g)$
a) addition of PCl_3	↑	↓	→	↑
b) removal of Cl_2	↑	↓	←	↓
c) removal of PCl_5	↓	↓	→	↓
d) decrease in the volume of the container	↓	↓	→	↑
e) addition of He without change in volume	–	–		–

inert gas no effect

4. How will each of the changes in question 3 affect the K_{eq}? (↑=increase; ↓=decrease; —= unchanged)

a ___ b ___ c ___ d ___ e ___ None!

The only thing that changes the K is change in Temp.

5. Indicate how each of the following changes affects the amount of each gas in the system below, for which $\Delta H_{reaction} = +9.9$ kcal. Endo

	$H_2(g)$	$CO_2(g)$		$H_2O(g)$	$CO(g)$	
a) addition of CO_2	↓	↑	→	↑	↑	
b) addition of H_2O	↑	↑	←	↑	↓	
c) addition of a catalyst	–	–		–	–	No shift
d) increase in temperature	↓	↓	→	↑	↑	
e) decrease in the volume of the container	–	–		–	–	Same moles of gas on both sides

6. How will each of the changes in question 5 affect the equilibrium constant? K

a – b – c – d ↑ e –

only changed by changes in Temperature

b/c Endothermic ↑T ↑K

7. Consider the equilibrium: $2N_2O(g) + O_2(g) \rightleftharpoons 4NO(g)$
How will the amount of chemicals at equilibrium be affected by

	$2N_2O(g)$	$+ O_2(g) \rightleftharpoons$		$4NO(g)$
a) adding N_2O	↑	↓	→	↑
b) removing O_2	↑	↓	←	↓
c) increasing the volume of the container	↓	↓	→	↑
d) adding a catalyst	No shift			

8. For the reaction, $4NH_3(g) + 3O_2(g) \rightleftharpoons 2N_2(g) + 6H_2O(l)$
How will the concentration of each chemical be affected by

	$4NH_3(g)$	$+ 3O_2(g) \rightleftharpoons$		$2N_2(g)$	$+ 6H_2O(l)$
a) adding O_2 to the system	↓	↑		↑	—
b) adding N_2 to the system	↑	↑		↑	
c) removing H_2O from the system	No effect				
d) decreasing the volume of the container	↓	↓	→	↑	

9. Consider the equilibrium: $2N_2O(g) + O_2(g) \rightleftharpoons 4NO(g)$
3.00 moles of NO(g) are introduced into a 1.00-Liter evacuated flask. When the system comes to equilibrium, 1.00 mole of $N_2O(g)$ has formed. Determine the equilibrium concentrations of each substance. Calculate the K_c for the reaction based on these data.

	$2\ N_2O$	O_2	$4\ NO$
initial	0	0	3.00 M
change	+ 1.00 M	+ 0.5 M	− 2.00 M
equilibrium	1.00 M	0.5 M	1.00 M

Remember: The "ice" box may be used with moles, molarity, or Liters (for gaseous equilibria)... never grams.

$$K_c = \frac{[NO]^4}{[N_2O]^2[O_2]} = \frac{(1.00M)^4}{(1.00)^2(0.5)} = 2.00M$$