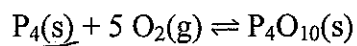
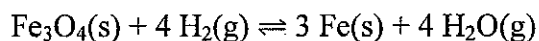


Equilibrium Practice



1. What is the equilibrium expression for this reaction?

- (A) $K_c = [\text{P}_4\text{O}_{10}] / [\text{P}_4] [\text{O}_2]^5$
 (B) $K_c = [\text{P}_4\text{O}_{10}] / 5 [\text{P}_4] [\text{O}_2]$
 (C) $K_c = [\text{O}_2]^5$
 (D) $K_c = 1 / [\text{O}_2]^5$



$\Delta H > 0$ 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 Fe₃O₄(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
 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

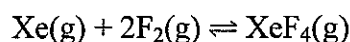
- (A) 1 only
 (B) 2 only
 (C) 1 and 2 only
 (D) neither 1 nor 2

5. What is the relationship between the equilibrium constant (K_c) of a reaction and the rate constants for the forward (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, XeF₄, can be prepared by heating Xe and F₂ together according to this equation.

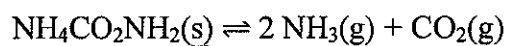


What is the equilibrium expression for this reaction?

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

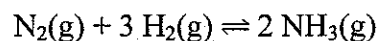
- (A) $K = [\text{XeF}_4] / ([\text{Xe}] [\text{F}_2])$
 (B) $K = [\text{XeF}_4] / (2[\text{Xe}] [\text{F}_2])$
 (C) $K = [\text{XeF}_4] / ([\text{Xe}] [\text{F}_2]^2)$
 (D) $K = ([\text{Xe}] [\text{F}_2]) / [\text{XeF}_4]$

7. What is the equilibrium expression for the decomposition of ammonium carbamate, NH₄CO₂NH₂, that occurs according to this equation:



- (A) $K = [\text{NH}_3][\text{CO}_2]$
 (B) $K = [\text{NH}_3]^2[\text{CO}_2]$
 (C) $K = [\text{NH}_3][\text{CO}_2] / [\text{NH}_4\text{CO}_2\text{NH}_2]$
 (D) $K = [\text{NH}_3]^2[\text{CO}_2] / [\text{NH}_4\text{CO}_2\text{NH}_2]$

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



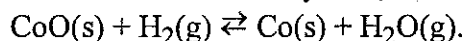
- (A) increasing the volume of the container
 (B) adding N₂
 (C) removing NH₃
 (D) lowering the temperature

$K_c \uparrow$

Ch 13 Chemical Equilibrium

PRACTICE TEST

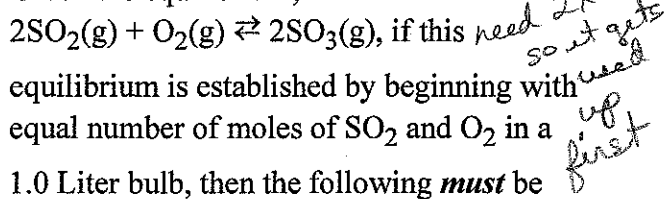
1. Consider the reaction system,



The equilibrium constant expression is

- a) $\frac{[\text{CoO}][\text{H}_2]}{[\text{Co}][\text{H}_2\text{O}]}$ d) $\frac{[\text{H}_2]}{[\text{H}_2\text{O}]} \frac{[\text{Prod}]}{[\text{React}]}$
 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,



- 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 NO,
 0.200 mol Cl_2 and 0.500 mol ClNO were

placed in a 25.0 Liter container. The
 following equilibrium is established:



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

4. The equilibrium constant,
- K_c
- , is:

a) 4.45×10^{-4}

d) 0.167

b) 6.67×10^{-4}

e) 1500

c) 0.111

5. At
- 985°C
- , the equilibrium constant for the reaction,



is 1.63. What is the equilibrium constant for the reverse reaction? $K' = \frac{1}{K_c} = \frac{1}{1.63}$

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,
- $2\text{ICl(g)} \rightleftharpoons \text{I}_2\text{(g)} + \text{Cl}_2\text{(g)}$
- ?

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) $K_p = K_c(RT)^2$

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

$$\Delta n = 0$$

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

7. For the reaction
- $2\text{NO}_2\text{(g)} \rightleftharpoons \text{N}_2\text{O}_4\text{(g)}$
- ,
- K_p
- at

25°C is 7.3, when all partial pressures are expressed in atmospheres. What is K_c for this reaction? [$R = 0.0821 \text{ L}\cdot\text{atm}\cdot\text{mol}^{-1}\cdot\text{K}^{-1}$]

a) 4270

d) 179

b) 0.0119

e) 2.06

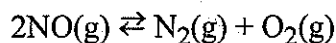
c) 0.291

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

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

$$= 178.6$$

8. 0.200 mol NO is placed in a one liter flask at 2273 K. After equilibrium is attained, 0.0863 mol N₂ and 0.0863 mol O₂ are present. What is K_c for this reaction?



- a) 9.92
b) 3.15
c) 0.0372

- d) 39.7
e) 0.576

$$2\text{NO} \rightleftharpoons \text{N}_2 + \text{O}_2$$

Initial: 0.2, 0, 0
Change: -2(0.0863), +0.0863, +0.0863
Equilibrium: 0.0274, 0.0863, 0.0863

$$K_c = \frac{[\text{N}_2][\text{O}_2]}{[\text{NO}]^2} = \frac{(0.0863)(0.0863)}{(0.0274)^2} = 11.1$$

9. N₂O₄(g) \rightleftharpoons 2 NO₂(g)

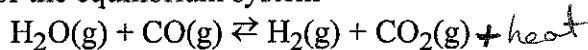
At 25°C, 0.11 mole of N₂O₄ reacts to form 0.10 mol of N₂O₄ and 0.02 mole of NO₂. At

90°C, 0.11 mole of N₂O₄ forms 0.050 mole of N₂O₄ and 0.12 mole of NO₂. From these data we can conclude

- a) N₂O₄ molecules react by a second order rate law.
b) N₂O₄ molecules react by a first order rate law.
c) the reaction is exothermic.
d) N₂O₄ 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



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

K_c equals 0.62 at 1260 K. If 0.10 mole each of H₂O, CO, H₂ and CO₂ (each at 1260 K) were placed in a 1.0-Liter flask at 1260 K, when the system came to equilibrium...

- a) decrease
b) decrease
c) remain constant
d) increase
e) increase

The temperature would	The mass of CO would
decrease	increase
decrease	decrease
remain constant	increase
increase	decrease
increase	increase

$$Q = \frac{(0.10)(0.10)}{(0.10)(0.10)} = 1$$

Q > K
will shift to left (make reactants)

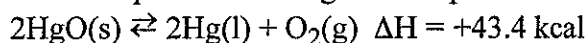
11. For the reaction system,



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

- 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 O₂ 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 effect
b) removing some O₂. 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
2.	D
3.	C
4.	B
5.	D
6.	D

7.	D
8.	A
9.	E
10.	A
11.	D
12.	C

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

Ch 13 Chemical Equilibria**PROBLEM SET #2**

1. Consider the equilibrium:
- $2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2 \text{SO}_3(\text{g})$
- $K_c = 4.36 \text{ M}^{-1}$

Calculate the value of "Q" for a situation in which the concentrations are $[\text{SO}_2] = 2.00 \text{ M}$, $[\text{O}_2] = 1.50 \text{ M}$, and $[\text{SO}_3] = 1.25 \text{ M}$.

$$Q = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2 [\text{O}_2]} = \frac{(1.25 \text{ M})^2}{(2.00 \text{ M})^2 (1.50 \text{ M})} = 0.260$$

Does this mixture shift toward the reactants or products to reach equilibrium? Toward Products $Q < K$ \therefore shifts toward products - numerator increase and denominator decrease

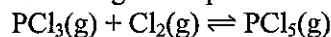
2. Study the discussion in your textbook about converting
- K_c
- and
- K_p
- . Write the
- K_p
- expression for the reaction in question 1 and calculate its value at
- 0°C
- . Remember,
- $R = 0.0821 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K}$
- . Temp = 273

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

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

3. Consider the equilibrium
- $\text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g}) \rightleftharpoons \text{PCl}_5(\text{g})$
- .

How would the following changes affect the partial pressures of each gas at equilibrium?

a) addition of PCl_3

\uparrow	\downarrow	\rightarrow	\uparrow
------------	--------------	---------------	------------

b) removal of Cl_2

\uparrow	\downarrow	\leftarrow	\downarrow
------------	--------------	--------------	--------------

c) removal of PCl_5

\downarrow	\downarrow	\rightarrow	\downarrow
--------------	--------------	---------------	--------------

d) decrease in the volume of the container

\downarrow	\downarrow	\rightarrow	\uparrow
--------------	--------------	---------------	------------

e) addition of He without change in volume

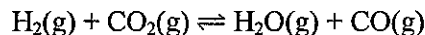
$-$	$-$	$-$	$-$
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inert gas no effect

4. How will each of the changes in question 3 affect the
- K_{eq}
- ? (
- \uparrow
- =increase;
- \downarrow
- =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
- $\Delta H_{\text{reaction}} = +9.9 \text{ kcal}$
- .
- Endo

a) addition of CO_2

\downarrow	\uparrow	\rightarrow	\uparrow	\uparrow
--------------	------------	---------------	------------	------------

b) addition of H_2O

\uparrow	\uparrow	\leftarrow	\uparrow	\downarrow
------------	------------	--------------	------------	--------------

c) addition of a catalyst

$-$	$-$	$-$	$-$	$-$
-----	-----	-----	-----	-----

d) increase in temperature

\downarrow	\downarrow	\rightarrow	\uparrow	\uparrow
--------------	--------------	---------------	------------	------------

e) decrease in the volume of the container

$-$	$-$	$-$	$-$	$-$
-----	-----	-----	-----	-----

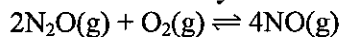
No shiftsame moles of gas on both sides

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

a $-$ b $-$ c $-$ d \uparrow e $-$ only changed by changes in Temperature
b/c Endothermic $\uparrow T \uparrow K$

7. Consider the equilibrium: $2\text{N}_2\text{O}(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 4\text{NO}(\text{g})$

How will the amount of chemicals at equilibrium be affected by



a) adding N_2O



b) removing O_2



c) increasing the volume of the container

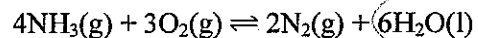


d) adding a catalyst

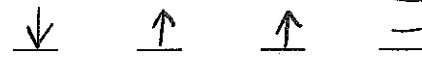


8. For the reaction,

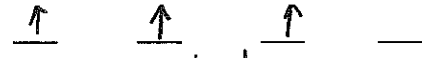
How will the concentration of each chemical be affected by



a) adding O_2 to the system



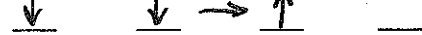
b) adding N_2 to the system



c) removing H_2O from the system



d) decreasing the volume of the container



9. Consider the equilibrium: $2\text{N}_2\text{O}(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 4\text{NO}(\text{g})$

3.00 moles of $\text{NO}(\text{g})$ are introduced into a 1.00-Liter evacuated flask. When the system comes to equilibrium, 1.00 mole of $\text{N}_2\text{O}(\text{g})$ has formed. Determine the equilibrium concentrations of each substance. Calculate the K_c for the reaction based on these data.

	$2\text{N}_2\text{O}$	O_2	4NO
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.

$$K_c = \frac{[\text{NO}]^4}{[\text{N}_2\text{O}]^2 [\text{O}_2]} = \frac{(1.00\text{M})^4}{(1.00)^2 (0.5)} = 2.00\text{M}$$