• AP Chemistry

Name ____________________

Period ___ Date ___/___/___

13 • Chemical Equilibrium

AP Problem -- 1995

CLEARLY SHOW THE METHOD USED AND THE STEPS INVOLVED IN ARRIVING AT YOUR ANSWERS. It is to your advantage to do this, since you may obtain partial credit if you do and you will receive little or no credit if you do not. Attention should be paid to significant figures.

$$H_2(g) + CO_2(g) \rightleftarrows H_2O(g) + CO(g)$$

When $H_2(g)$ is mixed with $CO_2(g)$ at 2,000 K, equilibrium is achieved according to the equation above. In one experiment, the following equilibrium concentrations were measured.

$$[H_2] = 0.20 \text{ mol/L}$$
$$[CO_2] = 0.30 \text{ mol/L}$$
$$[H_2O] = [CO] = 0.55 \text{ mol/L}$$

(a) What is the mole fraction of CO(g) in the equilibrium mixture?

(b) Using the equilibrium concentrations given above, calculate the value of K_c, the equilibrium constant for the reaction.

(c) Determine K_p, in terms of K_c for this system.

(d) When the system is cooled from 2,000 K to a lower temperature, 30.0 percent of the CO(g) is converted back to $CO_2(g)$. Calculate the value of K_c at this lower temperature.

(e) In a different experiment, 0.50 mole of $H_2(g)$ is mixed with 0.50 mole of $CO_2(g)$ in a 3.0-liter reaction vessel at 2,000 K. Calculate the equilibrium concentration, in moles per liter, of CO(g) at this temperature.

a) Total moles = $.2 + .3 + .55 + .55 = 1.6$ mol

$.55 / 1.6 = \boxed{.34}$

b) $Kc = \frac{[H_2O][CO]}{[H_2][CO_2]} = \frac{(.55)(.55)}{(.20)(.30)} = \boxed{5.04}$

c) $Kp = kc(RT)^{\Delta n} \cdot 5.04(.08206 \cdot 2000)^0$

$\Delta n = 0 ; \boxed{Kp = kc}$

d) $(.55)\cdot(.3) = .165 \qquad .55 - .165 = .385$

$(.55 - .165) = .385$

$.2 + .165 = .365 \qquad Kc = \frac{(.385)(.385)}{(.365)(.465)} \rightarrow \boxed{.87}$

$.3 + .165 = .465$

e) $H_2 + CO_2 \rightleftarrows H_2O + CO$

	H_2	CO_2	H_2O	CO
I	.5	.5	0	0
C	-x	-x	+x	+x
E	.5-x	.5-x	x	x

$5.04 = \frac{x^2}{(.5-x)(.5-x)}$

$.25 - .5x - .5x + x^2$

$x^2 - x + .25$

$5.04x^2 - 5.04x + 1.26 = x^2$

$4.04x^2 - 5.04x + 1.26 = 0$

$x = .346$ mol

$.346 / 3L = \boxed{.12 M}$

13 [redacted]

• AP Chemistry

Name ____________________
Period ___ Date ___/___/___

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 $[SO_3] = 1.25\ \underline{M}$.

$$Q = \frac{[SO_3]^2}{[SO_2]^2[O_2]} = \frac{(1.25)^2}{(2.00)^2(1.50)} = .260 \qquad Q < K$$

Does this mixture shift toward the reactants or products to reach equilibrium? Shift to the right

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.

$$K_p = K_c(RT)^{\Delta n}$$
$$K_p = (4.36)\left((0.08206)(273)\right)^{2-3}$$
$$\boxed{K_p = .195}$$

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?

* this question is asking how ~~the amount~~ the partial pressures would differ based on where they were at equilibrium, not the change that needs to occur to get back to 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	—	—	—

* the addition of an inert gas increases the total pressure, but has no effect on the concentrations or partial pressures of the reactants or products

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

a — b — c — d — e —

only temp can alter Keq

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.

endothermic

ame
#3, but
r [ ]

	heat + $H_2(g)$ +	$CO_2(g)$ ⇌	$H_2O(g)$ +	$CO(g)$
a) addition of CO_2	↓	↑	↑	↑
b) addition of H_2O	↑	↑	↑	↓
c) addition of a catalyst	—	—	—	—
d) increase in temperature	↓	↓	↑	↑
e) decrease in the volume of the container	—	—	—	—

(same # of moles of gas on both sides)

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

a — b — c — d ↑ e —

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

	$2N_2O(g)$	$O_2(g)$	$4NO(g)$
a) adding N_2O	↑	↓	↑
b) removing O_2	↑	↓	↓
c) increasing the volume of the container	↓	↓	↑
d) adding a catalyst	—	—	—

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

	$4NH_3(g)$	$3O_2(g)$	$2N_2(g)$	$6H_2O(l)$
a) adding O_2 to the system	↓	↑	↑	—
b) adding N_2 to the system	↑	↑	↑	—
c) removing H_2O from the system	—	—	—	—
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
change	+2x	+x	-4x
equilibrium	1.00	.5	1.00

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

$x = .5$

$$K_c = \frac{[NO]^4}{[N_2O]^2[O_2]} = \frac{[1]^4}{[1]^2[.5]}$$ ②

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Previous AP Problems using ICE

1. $CO_2(g) + H_2(g) \leftrightarrow H_2O(g) + CO(g)$

When $H_2(g)$ is mixed with $CO_2(g)$ at 2,000 K, equilibrium is achieved according to the equation above. In one experiment, the following equilibrium concentrations were measured.

$$[H_2] = 0.20 \text{ mol/L}$$
$$[CO_2] = 0.30 \text{ mol/L}$$
$$[H_2O] = [CO] = 0.55 \text{ mol/L}$$

(a) Using the equilibrium concentrations given above, calculate the value of K_c, the equilibrium constant for the reaction.

(b) Determine K_p in terms of K_c for this system.

(c) In a different experiment, 0.50 mole of $H_2(g)$ is mixed with 0.50 mole of $CO_2(g)$ in a 3.0-liter reaction vessel at 2,000 K. Calculate the equilibrium concentration, in moles per liter, of $CO(g)$ at this temperature.

Same as page 13

2. $2\ H_2S(g) \rightarrow 2\ H_2(g) + S_2(g)$

When heated, hydrogen sulfide gas decomposes according to the equation above. A 3.40 g sample of $H_2S(g)$ is introduced into an evacuated rigid 1.25 L container. The sealed container is heated to 483 K, and 3.72×10^{-2} mol of $S_2(g)$ is present at equilibrium.

$$3.40 \text{g } H_2S \left| \frac{1 \text{ mol}}{34.086 \text{g}} \right. = .0997$$

(a) Write the expression for the equilibrium constant, K_c, for the decomposition reaction represented above.

(b) Calculate the equilibrium concentration, in mol·L^{-1}, of the following gases in the container at 483 K.

(i) $H_2(g)$

(ii) $H_2S(g)$

(c) Calculate the value of the equilibrium constant, K_c, for the decomposition reaction at 483 K.

a) $$K_c = \frac{[S_2][H_2]^2}{[H_2S]^2}$$

c) $$K_c = \frac{(.298)(.0595)^2}{(.0202)^2}$$

$$\boxed{K_c = 2.59}$$

b) $2 H_2S \rightleftharpoons 2 H_2 + S_2$

I	.0997 mol	0	0
C	$-2x$	$+2x$	$+y$
E	$.0253/_{1.25L}$	$.0744/_{1.25L}$	$3.72\times10^{-2}/_{1.25L}$

$[H_2S] = .0202 M$ $[S_2] = .0298 M$

$[H_2] = .0595 M$