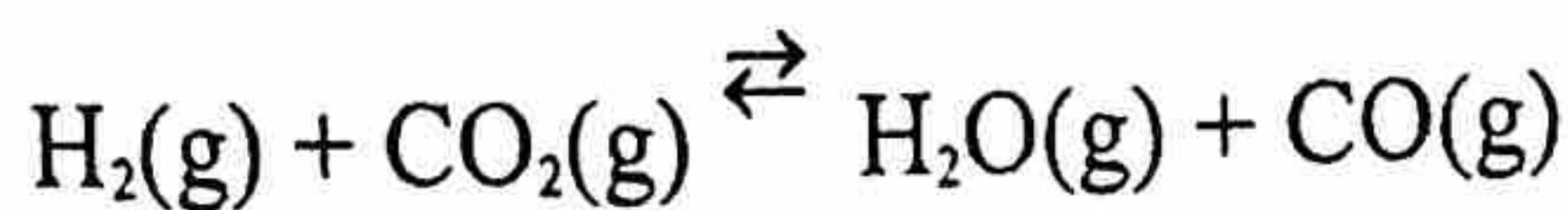


# 13 • Chemical Equilibrium

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.



When  $\text{H}_2(\text{g})$  is mixed with  $\text{CO}_2(\text{g})$  at 2,000 K, equilibrium is achieved according to the equation above. In one experiment, the following equilibrium concentrations were measured.

$$[\text{H}_2] = 0.20 \text{ mol/L}$$

$$[\text{CO}_2] = 0.30 \text{ mol/L}$$

$$[\text{H}_2\text{O}] = [\text{CO}] = 0.55 \text{ mol/L}$$

- What is the mole fraction of  $\text{CO}(\text{g})$  in the equilibrium mixture?
- Using the equilibrium concentrations given above, calculate the value of  $K_c$ , the equilibrium constant for the reaction.
- Determine  $K_p$ , in terms of  $K_c$  for this system.
- When the system is cooled from 2,000 K to a lower temperature, 30.0 percent of the  $\text{CO}(\text{g})$  is converted back to  $\text{CO}_2(\text{g})$ . Calculate the value of  $K_c$  at this lower temperature.
- In a different experiment, 0.50 mole of  $\text{H}_2(\text{g})$  is mixed with 0.50 mole of  $\text{CO}_2(\text{g})$  in a 3.0-liter reaction vessel at 2,000 K. Calculate the equilibrium concentration, in moles per liter, of  $\text{CO}(\text{g})$  at this temperature.

a) Total moles =  $.2 + .3 + .55 + .55 = 1.6 \text{ mol}$

$.55 / 1.6 = .34$

b)  $K_c = \frac{[\text{H}_2\text{O}][\text{CO}]}{[\text{H}_2][\text{CO}_2]} = \frac{(0.55)(0.55)}{(0.20)(0.30)} = 5.04$

c)  $K_p = K_c (RT)^{\Delta n} = 5.04 (0.08206 \cdot 2000)^0$   
 $\Delta n = 0$ ;  $K_p = K_c$

d)  $(0.55)(0.3) = 0.165$ ,  $0.55 - 0.165 = 0.385$   
 $(0.55 - 0.165) = 0.385$   
 $0.2 + 0.165 = 0.365$ ,  $0.3 + 0.165 = 0.465$   
 $K_c = \frac{(0.385)(0.385)}{(0.365)(0.465)} = 0.87$

e)  $\text{H}_2 + \text{CO}_2 \rightleftharpoons \text{H}_2\text{O} + \text{CO}$

I	0.5	0.5	0	0
C	-x	-x	+x	+x
E	0.5-x	0.5-x	x	x

$5.04 = \frac{x^2}{(0.5-x)(0.5-x)}$   
 $0.25 - 0.5x - 0.5x + x^2$   
 $x^2 - x + 0.25$

$0.346 / 3L = 0.12 \text{ M}$   
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$5.04x^2 - 5.04x + 1.26 = x^2$   
 $4.04x^2 - 5.04x + 1.26 = 0$   
 $x = 0.346 \text{ mol}$

# 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  $[\text{SO}_3] = 1.25 \text{ M}$ .

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

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  $0^\circ\text{C}$ . Remember,  $R = 0.0821 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K}$ .

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

$$K_p = (4.36) (0.08206) (273)^{2-3} \quad K_p = 0.195$$

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?

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

	$\text{PCl}_3(\text{g})$	$\text{Cl}_2(\text{g})$	$\text{PCl}_5(\text{g})$
a) addition of $\text{PCl}_3$	↑	↓	↑
b) removal of $\text{Cl}_2$	↑	↓	↓
c) removal of $\text{PCl}_5$	↓	↓	↓
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

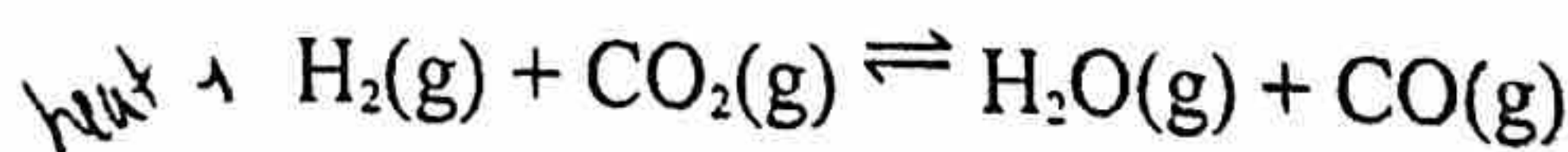
4. How will each of the changes in question 3 affect the  $K_{eq}$ ? (↑=increase; ↓=decrease; —=unchanged)

a —      b —      c —      d —      e —

only  $K_{mp}$  can alter  $K_{eq}$

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

endothermic



a) addition of $\text{CO}_2$	↓	↑	↑	↑
b) addition of $\text{H}_2\text{O}$	↑	↑	↑	↓
c) addition of a catalyst	—	—	—	—
d) increase in temperature	↓	↓	↑	↑
e) decrease in the volume of the container	—	—	—	—

(same # of mols of gas on both sides)

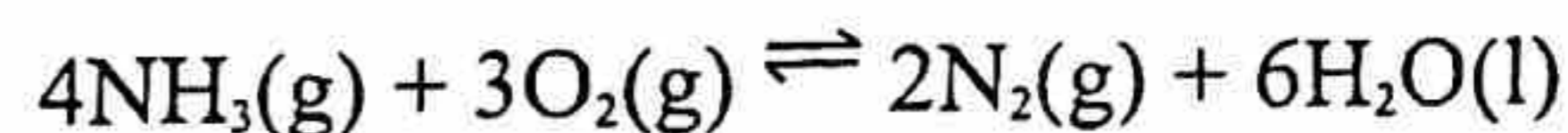
6. How will each of the changes in question 5 affect the equilibrium constant?

a —      b —      c —      d ↑      e —

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

	$2\text{N}_2\text{O}(\text{g})$	$+\text{O}_2(\text{g})$	$\rightleftharpoons$	$4\text{NO}(\text{g})$
a) adding $\text{N}_2\text{O}$	$\uparrow$	$\downarrow$		$\uparrow$
b) removing $\text{O}_2$	$\uparrow$	$\downarrow$		$\downarrow$
c) increasing the volume of the container	$\downarrow$	$\downarrow$		$\uparrow$
d) adding a catalyst	$-$	$-$		$-$

8. For the reaction,  
 How will the concentration of each chemical be affected by



a) adding $\text{O}_2$ to the system	$\downarrow$	$\uparrow$	$\uparrow$	$-$
b) adding $\text{N}_2$ to the system	$\uparrow$	$\uparrow$	$\uparrow$	$-$
c) removing $\text{H}_2\text{O}$ from the system	$-$	$-$	$-$	$-$
d) decreasing the volume of the container	$\downarrow$	$\downarrow$	$\uparrow$	$-$

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.

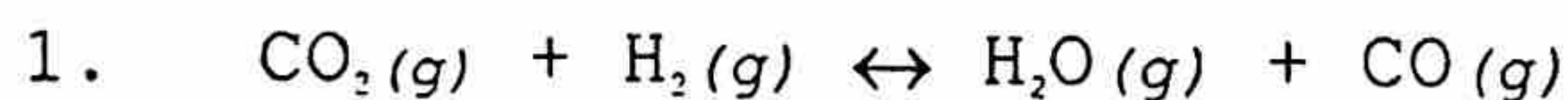
	$2\text{N}_2\text{O}$	$\text{O}_2$	$4\text{NO}$
initial	0	0	3.00
change	$+2x$	$+x$	$-4x$
equilibrium	1.00	.5	1.00

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

$$x = 0.5$$

$$K_c = \frac{[\text{NO}]^4}{[\text{N}_2\text{O}]^2 [\text{O}_2]} = \frac{[1]^4}{[1]^2 [0.5]} \quad (2)$$

Previous AP Problems using ICE



When  $\text{H}_2(g)$  is mixed with  $\text{CO}_2(g)$  at 2,000 K, equilibrium is achieved according to the equation above. In one experiment, the following equilibrium concentrations were measured.

$$\begin{aligned} [\text{H}_2] &= 0.20 \text{ mol/L} \\ [\text{CO}_2] &= 0.30 \text{ mol/L} \\ [\text{H}_2\text{O}] &= [\text{CO}] = 0.55 \text{ mol/L} \end{aligned}$$

- (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  $\text{H}_2(g)$  is mixed with 0.50 mole of  $\text{CO}_2(g)$  in a 3.0-liter reaction vessel at 2,000 K. Calculate the equilibrium concentration, in moles per liter, of  $\text{CO}(g)$  at this temperature.

Same as page 13



When heated, hydrogen sulfide gas decomposes according to the equation above. A 3.40 g sample of  $\text{H}_2\text{S}(g)$  is introduced into an evacuated rigid 1.25 L container. The sealed container is heated to 483 K, and  $3.72 \times 10^{-2}$  mol of  $\text{S}_2(g)$  is present at equilibrium.

$$\frac{3.40 \text{ g H}_2\text{S}}{34.08 \text{ g/mol}} = .0997$$

- (a) Write the expression for the equilibrium constant,  $K_c$ , for the decomposition reaction represented above.
- (b) Calculate the equilibrium concentration, in  $\text{mol}\cdot\text{L}^{-1}$ , of the following gases in the container at 483 K.
  - (i)  $\text{H}_2(g)$
  - (ii)  $\text{H}_2\text{S}(g)$
- (c) Calculate the value of the equilibrium constant,  $K_c$ , for the decomposition reaction at 483 K.

a) 
$$K_c = \frac{[\text{S}_2][\text{H}_2]^2}{[\text{H}_2\text{S}]^2}$$

c) 
$$K_c = \frac{(0.0298)(0.0595)^2}{(0.0202)^2}$$



$K_c = 2.59$

I	.0997 mol	0	0
C	-2x	+2x	+x

E	.0253 / 1.25L	.0744 / 1.25L	$3.72 \times 10^{-2} / 1.25L$
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$[\text{H}_2\text{S}] = .0202 \text{ M}$        $[\text{S}_2] = .0298 \text{ M}$

$[\text{H}_2] = .0595 \text{ M}$

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