

# CHEMISTRY 1220

## CHAPTER 15 PRACTICE EXAM



1. Which is a proper description of chemical equilibrium?

- (A) The frequencies of reactant and product collisions are equal.
- (B) The concentrations of products and reactants are identical.
- (C) The kinetic energy of product and reactant molecules are identical.
- (D) Reactant molecules react to form products as fast as product molecules are reacting to form reactants.**
- (E) The formation of reactant molecules to form product molecules has stopped.

• emphasize that Equilibrium is dynamic  
+ reference PhET simulation from class.

2. The value of an equilibrium constant can be used to predict which of the following?

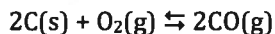
- I. Direction of a reaction
  - II. Extent of a reaction
  - III. Time required to reach equilibrium
- (A) I only   **(B) I and II**   (C) II and III   (D) I and III   (E) I, II and III

(I) Using  $Q$  vs.  $K$  + the magnitude of  $K$   
we can predict the direction

(II) and also the extent of rxn.

~~III~~ Kinetics give time to reach Eq.

3. Consider this reaction



What is the equilibrium expression for this reaction?

(A)  $K_c = \frac{[\text{CO}]}{[\text{C}][\text{O}_2]}$

(B)  $K_c = \frac{[\text{CO}]^2}{[\text{C}]^2[\text{O}_2]}$

(C)  $K_c = \frac{[2\text{CO}]}{[2\text{C}][\text{O}_2]}$

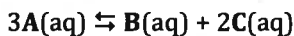
(D)  $K_c = \frac{[\text{CO}]^2}{[\text{O}_2]}$

(E)  $K_c = \frac{[\text{CO}]}{[\text{O}_2]}$

$$K_c = \frac{[\text{CO}]^2}{[\text{O}_2]}$$

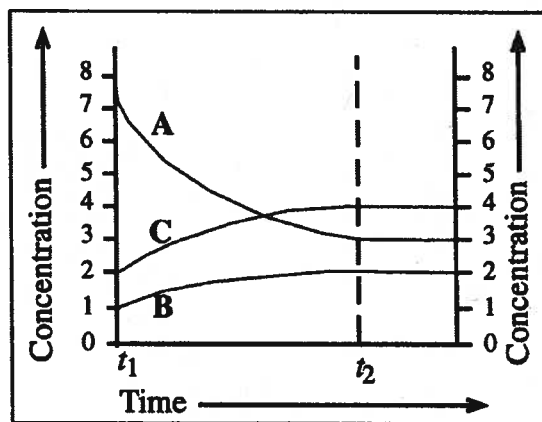
Note: pure liquids + solids do not show up in Eq Expression

4. The graph shows a variation of concentration with time for this reaction at 25 °C.



What is the value of the equilibrium constant at time  $t_2$ ?

- (A) 1.2 (B) 0.84 (C) 0.57 (D) 0.28 (E) 0.22



$$[\text{A}]_{\text{eq}} = 3, [\text{B}]_{\text{eq}} = 2, [\text{C}]_{\text{eq}} = 4$$

$$K = \frac{[\text{B}][\text{C}]^2}{[\text{A}]^3} = \frac{[2][4]^2}{[3]^3} = 1.2$$

5. Carbon monoxide gas reacts with hydrogen gas at elevated temperatures to form methanol according to this equation.



When 0.40 mol of CO and 0.30 mol of H<sub>2</sub> are allowed to reach equilibrium in a 1.0 L container, 0.060 mol of CH<sub>3</sub>OH are formed. What is the value of K<sub>c</sub>?

(A) 0.50

(B) 0.98

(C) 1.7

(D) 3.1

(E) 5.4

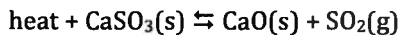
$$K_c = \frac{[\text{CH}_3\text{OH}]}{[\text{CO}][\text{H}_2]^2}$$

	CO(g)	+ 2H <sub>2</sub> (g)	⇌	CH <sub>3</sub> OH(g)
I	0.40	0.30		0
C	-0.060	-2(0.060)		+0.060
E	0.34	0.18		0.060

$$K_c = \frac{[0.060]}{[0.34][0.18]^2}$$

$$K_c = 5.4$$

6. Consider this reaction.



What will cause an increase in the pressure of SO<sub>2</sub>(g) when equilibrium is re-established?

(A) increasing the reaction temperature

(B) decreasing the volume of the container

(C) adding some more CaSO<sub>3</sub>

(D) removing some of the CaO

(E) all of the above

- ✓ a) since heat is a reactant, ↑ Temp shift (R) which increases P<sub>SO<sub>2</sub></sub>
- ✗ b) decreasing the volume shifts Eq. to the direction (↑ P) in which # mol (g) is smaller ⇒ (L)
- ✗ c) + d) adding/removing solids does not shift Eq.

7. Calculate  $K_c$  at 303 K for  $\text{SO}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightleftharpoons \text{SO}_2\text{Cl}_2(\text{g})$  if  $K_p = 36.5$  at this temperature.

(A) 36.5

(B) 1.47

(C) 908

(D)  $9.19 \times 10^4$

(E)  $1.45 \times 10^{-2}$

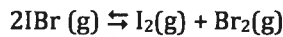
$$K_p = K_c (RT)^{\Delta n}$$

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

$$K_c = \frac{36.5}{(0.08206 \frac{\text{L atm}}{\text{mol K}} \cdot 303 \text{ K})^{-1}}$$

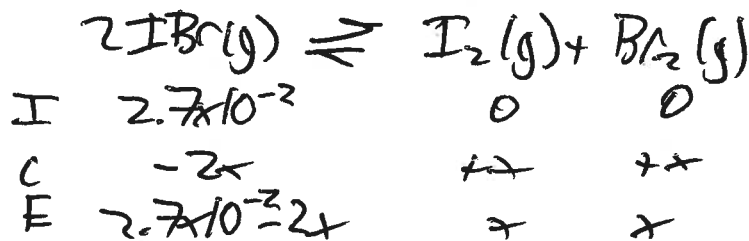
$$K_c = 907.5$$

8. The  $K_p$  for the following equilibrium is  $8.5 \times 10^{-3}$  at  $150^\circ\text{C}$ .



If  $2.7 \times 10^{-2}$  atm of IBr is placed in a 2.0 - L container, what is the partial pressure of this substance after equilibrium is reached?

(A)  $2.7 \times 10^{-2}$  atm (B)  $2.3 \times 10^{-2}$  atm (C)  $2.5 \times 10^{-2}$  atm (D)  $2.1 \times 10^{-3}$  atm (E)  $2.5 \times 10^{-3}$  atm



$$K_p = \frac{P_{\text{I}_2} P_{\text{Br}_2}}{(P_{\text{IBr}})^2}$$

$$\sqrt{\frac{x^2}{(2.7 \times 10^{-2} - 2x)^2}} = \sqrt{8.5 \times 10^{-3}}$$

$$\frac{x}{2.7 \times 10^{-2} - 2x} = 0.092195 (2.7 \times 10^{-2} - 2x)$$

$$x = 0.002489 - 0.184391x$$

$$1.184391x = 0.002489$$

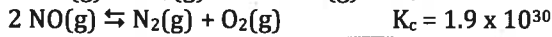
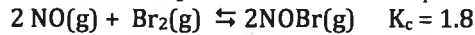
$$x = 2.1 \times 10^{-3}$$

$$[\text{IBr}] = 2.7 \times 10^{-2} - 2(2.1 \times 10^{-3}) = 2.3 \times 10^{-2} \text{ atm}$$

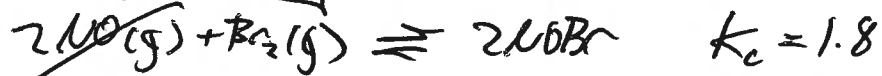
9. Consider the following equilibrium



Calculate the equilibrium constant,  $K_p$  for this reaction given the following (at 298 K):



- (A)  $9.5 \times 10^{-31}$  (B)  $3.9 \times 10^{-32}$  (C)  $3.4 \times 10^{30}$  (D)  $2.3 \times 10^{-29}$  (E)  $3.8 \times 10^{-34}$



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

$$K_p = 2.47 \times 10^{-31} (0.08206 \cdot 298)^{-1}$$

$$K_p = 3.9 \times 10^{-32}$$

10. Cyclopropane ( $\text{C}_3\text{H}_6$ ) reacts to form propene ( $\text{C}_3\text{H}_6$ ) in the gas phase. The reaction is first order in cyclopropane and has a rate constant of  $5.87 \times 10^{-4}/\text{s}$  at  $485^\circ\text{C}$ . If a 2.3-L reaction vessel initially contains 722 torr of cyclopropane at  $485^\circ\text{C}$ , how long will it take for the partial pressure of cyclopropane to drop to below 105 torr?  $\rightarrow 758\text{K}$

- (A) 13.9 sec (B) 4.13 min (C) 195 min (D) 78.4 min (E) 54.7 min

$$n = \frac{PV}{RT} \quad 722 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.95 \text{ atm}, \quad 105 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.138 \text{ atm}$$

$$\frac{1}{V} = \frac{P}{RT} \quad [\text{C}_3\text{H}_6]_0 = \frac{P}{RT} = \frac{0.95 \text{ atm}}{0.08206 \frac{\text{L atm}}{\text{mol K}} (758\text{K})} = 0.015273 \text{ M}$$

$$[\text{C}_3\text{H}_6]_t = \frac{P}{RT} = \frac{0.138 \text{ atm}}{0.08206 \frac{\text{L atm}}{\text{mol K}} (758\text{K})} = 0.002221 \text{ M}$$

1st order

$$\ln [\text{C}_3\text{H}_6]_t = -kt + \ln [\text{C}_3\text{H}_6]_0$$

$$\ln [0.002221] = -5.87 \times 10^{-4} \text{ s}^{-1} (t) + \ln [0.015273]$$

$$-6.109794 = -5.87 \times 10^{-4} \text{ s}^{-1} (t) + -4.181667$$

$$t = \frac{-1.928125}{-5.87 \times 10^{-4}} = 3,284.7 \text{ sec} \times \frac{1 \text{ min}}{60 \text{ sec}} = 54.7 \text{ min}$$

11. Dinitrogen pentoxide decomposes in the gas phase to form nitrogen dioxide and oxygen gas. The reaction is first order in dinitrogen pentoxide and has a half-life of 2.81 hours at 25°C. If a 1.5-L reaction vessel initially contains 735 torr of N<sub>2</sub>O<sub>5</sub> at 25°C, what partial pressure of O<sub>2</sub> will be present in the vessel after 230 minutes?

(A) 9.18 torr (B) 225 torr (C) 286 torr (D) 449 torr (E) 163 torr

$$t_{1/2} = \frac{0.693}{k} \quad k = \frac{0.693}{2.81 \text{ hr}} = 0.246619 \text{ hr}^{-1}, \quad 25^\circ\text{C} \rightarrow 298\text{K}, \quad 735 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.967 \text{ atm}$$

$$M = \frac{n}{V} = \frac{P}{RT} = \frac{0.967 \text{ atm}}{0.08206 \frac{\text{L atm}}{\text{mol K}} (298\text{K})} = 0.039548 \text{ M} \quad 230 \text{ min} \times \frac{1 \text{ hr}}{60 \text{ min}} = 3.833 \text{ hr}$$

1st order

$$\ln [N_2O_5]_t = -kt - \ln [N_2O_5]_0$$

$$\ln [N_2O_5]_t = -0.246619 \text{ hr}^{-1} (3.833 \text{ hr}) - \ln [0.039548]$$

$$\ln [N_2O_5]_t = -4.175531$$

$$[N_2O_5]_t = e^{-4.175531}$$

$$[N_2O_5]_t = 0.015367 \text{ M}$$



$$\begin{array}{l} I \ 0.039548 \\ C \ -0.024181 \\ F \ 0.015367 \end{array}$$

$$0 \quad 0 \\ +\frac{1}{2}(0.024181) \\ 0.01209$$

$$PV = nRT$$

$$P = \frac{n}{V} RT$$

$$= 0.01209 \text{ M} (0.08206 \frac{\text{L atm}}{\text{mol K}}) (298\text{K}) = 0.295659 \text{ atm} \times \frac{760 \text{ torr}}{1 \text{ atm}} = 225 \text{ torr}$$

12. At equilibrium

a) All chemical reactions have ceased.

b)  $K_c = (\text{rate}_{\text{reverse}})(\text{rate}_{\text{forward}})$

c) The rate constant for the forward reaction equals the rate constant of the reverse reaction.

d) The rate of the forward reaction equals the rate of the reverse reaction.

e)  $K_c = (\text{rate}_{\text{reverse}})/(\text{rate}_{\text{forward}})$

= 225 torr

X a) Eq. is dynamic

X b) and e)  $K_c = \frac{[C]_{\text{products}}}{[C]_{\text{reactants}}}$

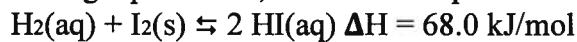
X c) rate forward = rate reverse  $\Rightarrow K_{\text{forward}} \neq K_{\text{reverse}}$

13. An equilibrium constant with a small magnitude indicates that a system favors \_\_\_\_\_ when it reaches equilibrium.

- a) reactants
- b) products
- c) neither reactants nor products
- d) both reactants and products
- e) none of the above

$$K = \frac{[ ]_{\text{products}}}{[ ]_{\text{reactants}}} \leftarrow \text{This is large when } K \text{ is small}$$

14. Considering the following equilibrium, what is the expression for K?



- a)  $[\text{HI}]^2 / [\text{H}_2]$
- b)  $[\text{H}_2][\text{I}_2] / [\text{HI}]^2$
- c)  $([\text{H}_2][\text{I}_2])^{1/2} / [\text{HI}]^2$
- d)  $[\text{HI}] / [\text{H}_2]^{1/2}$
- e)  $[\text{HI}]^2 / [\text{H}_2][\text{I}_2]$

$$K = \frac{[\text{HI}]^2}{[\text{H}_2]}$$

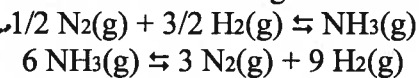
solids + pure liquids are not part of Eq. expression

15. At a given temperature,  $K = 1.3 \times 10^{-2}$  for the reaction  

$$\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightleftharpoons 2 \text{NH}_3(\text{g})$$

Calculate the values of  $K$  for each of the following two reactions:

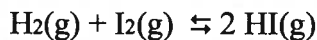
- a)  $2.8 \times 10^4, 4.5 \times 10^{-2}$
- b)  $2.65 \times 10^{-4}, 5.3$
- c)  $0.11, 4.6 \times 10^5$
- d)  $6.8, 7.3 \times 10^3$
- e)  $4.5 \times 10^{-2}, 2.8 \times 10^4$



$$K = (1.3 \times 10^{-2})^{1/2} = 0.114$$

$$K = \left(\frac{1}{1.3 \times 10^{-2}}\right)^3 = 4.55 \times 10^5$$

Questions 16-18. An equilibrium mixture of  $\text{H}_2$ ,  $\text{I}_2$  and  $\text{HI}$  at  $458^\circ\text{C}$  contains 1.34 atm of  $\text{H}_2$ , 1.34 atm of  $\text{I}_2$  and 9.30 atm of  $\text{HI}$ .



16. What is the equilibrium constant ( $K_p$ ) for this reaction?

- a) 31.6
- b) 48.2
- c) 64.7
- d) 78.3
- e) 92.5

$$K_p = \frac{(P_{\text{HI}})^2}{P_{\text{H}_2} \cdot P_{\text{I}_2}} = \frac{(9.30)^2}{1.34(1.34)} = 48.168$$



17. The addition of 0.100 mol of HI is made. This is equivalent to adding 1.20 atm of HI pressure. What is the value of Q?

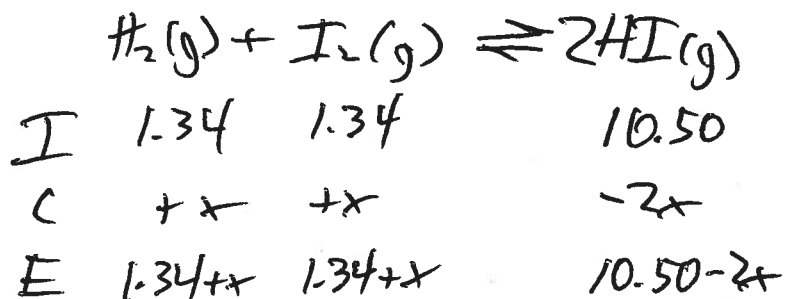
- a) 61.4 b) 81.3 c) 98.4 d) 121 e) 148

$$Q = \frac{(P_{HI})^2}{P_{H_2} \cdot P_{I_2}} = \frac{(10.50)^2}{1.34 \cdot 1.34} = 61.4$$

18. What is the final pressure of  $H_2(g)$ ?

- a) 1.07 atm b) 1.27 atm c) 1.37 atm d) 1.47 atm e) 1.57 atm

$Q > K \Rightarrow E_{eq}$  shifts  $\text{L}$



$$K_p = \frac{(10.50 - 2x)^2}{(1.34 + x)^2} = 48.168$$

$$\frac{10.50 - 2x}{1.34 + x} = 6.94 (1.34 + x)$$

$$10.50 - 2x = 9.3 + 6.94x$$

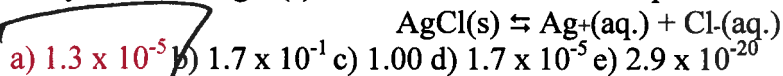
$$1.2 = 8.94x$$

$$x = 0.134$$

$$P_{H_2} = 1.34 + x$$

$$= 1.34 + 0.134 = 1.474$$

19. The equilibrium constant for the following reaction is  $1.7 \times 10^{-10}$ . How many moles of  $\text{AgCl}(s)$  will dissolve in 1.00 L of pure water?



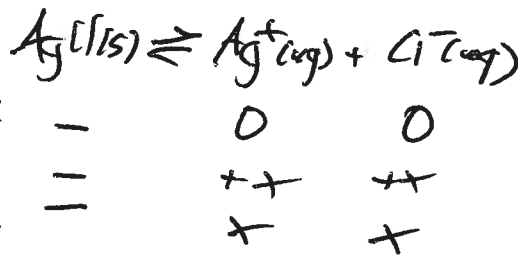
- a)  $1.3 \times 10^{-5}$  b)  $1.7 \times 10^{-1}$  c) 1.00 d)  $1.7 \times 10^{-5}$  e)  $2.9 \times 10^{-20}$

$$K = [\text{Ag}^+][\text{Cl}^-] = 1.7 \times 10^{-10}$$

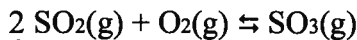
$$x^2 = 1.7 \times 10^{-10}$$

$$x = \frac{1.3 \times 10^{-5} \text{ mol}}{1.00 \text{ L}}$$

$$x = 1.3 \times 10^{-5} \text{ mol}$$



20. At 1000 K, the value  $K_c$  for the reaction below is  $4.08 \times 10^{-3}$ .



You have a mixture of  $8.0 \times 10^{-3} \text{ M SO}_3$ ,  $1.6 \times 10^{-2} \text{ M SO}_2$  and  $5.2 \times 10^{-4} \text{ M O}_2$ . What is the value of  $Q$  and which direction must the mixture shift to reach equilibrium?

- a)  $1.7 \times 10^{-3}$  right shift  
 b)  $6.0 \times 10^4$  right shift  
 c)  $6.0 \times 10^4$  left shift  
 d)  $1.7 \times 10^{-3}$  left shift  
 e) This system is at equilibrium

$$Q = \frac{[\text{SO}_3]}{[\text{SO}_2]^2[\text{O}_2]} = \frac{[8.0 \times 10^{-3}]}{[1.6 \times 10^{-2}]^2 [5.2 \times 10^{-4}]}$$

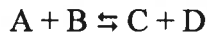
$$Q = 6.0 \times 10^4$$

$$Q > K$$

Eq shifts  $\text{L}$

21. The chemical reaction below has an equilibrium constant of 1.5. What is the final concentration of D at equilibrium if the initial concentrations are [A] = 1.00 M and [B] = 2.00 M?

- a) 0.18
- b) 0.39
- c) 0.58
- d) 0.73
- e) 0.92



I	1.00	2.00	0	0
C	-x	-x	+x	+x
E	1-x	2-x	x	x

$$(1-x)(2-x)$$

$$2-x-2x+x^2$$

$$\frac{x^2}{(1-x)(2-x)} = 1.5$$

$$x^2 = 1.5(x^2 - 3x + 2)$$

$$[D] = x$$

$$x^2 = 1.5x^2 - 4.5x + 3$$

$$-\frac{1}{2}x^2 + 4.5x - 3 = 0$$

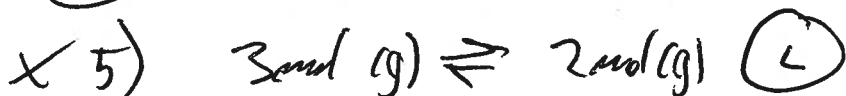
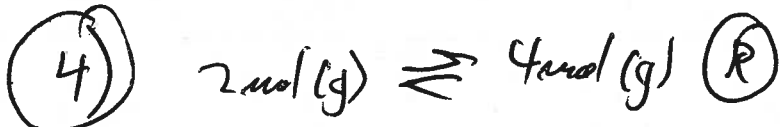
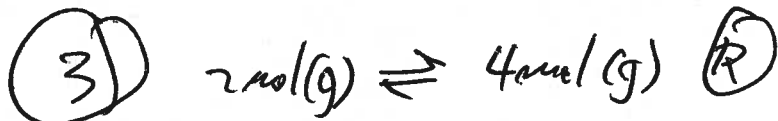
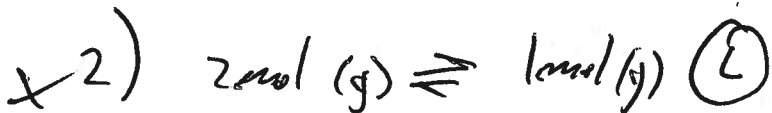
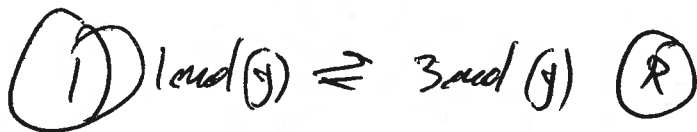
$$x = 0.73 = [D]$$

22. Consider the following reactions at equilibrium. In which cases will the reaction proceed to the right by decreasing the pressure?

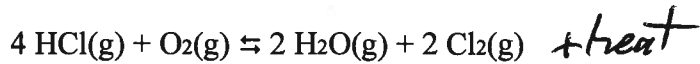
- 1)  $\text{CH}_3\text{OH}(\text{g}) \rightleftharpoons \text{CO}(\text{g}) + 2 \text{H}_2(\text{g})$
- 2)  $2 \text{CO}(\text{g}) \rightleftharpoons \text{CO}_2(\text{g}) + \text{C}(\text{s})$
- 3)  $\text{CH}_4(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightleftharpoons \text{CO}(\text{g}) + 3 \text{H}_2(\text{g})$
- 4)  $2 \text{NH}_3(\text{g}) \rightleftharpoons \text{N}_2(\text{g}) + 3 \text{H}_2(\text{g})$
- 5)  $2 \text{CO}(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2 \text{CO}_2(\text{g})$

- a) 1, 2, 3 b) 2, 3, 4 c) 3, 4, 5 d) 2, 3, 5 e) 1, 3, 4

When pressure on an Eq mixture involving gases ↓, the Eq shifts in the direction where # mol (g) becomes larger



23. For the following exothermic reaction, which condition will shift the equilibrium to the right?



- a) increasing the temperature
- b) increasing the pressure**
- c) adding a catalyst
- d) adding more water
- e) more than one of the above is correct

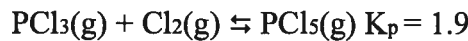
x a) shifts (L)

✓ b) 5 mol(g) ⇌ 4 mol(g) (R)

x c) catalyst does not influence direction of Eq.

x d) ↑ H<sub>2</sub>O (L)

\*24. (5 pts) An equilibrium mixture in a 1.00 liter flask contains 4.5 mol of PCl<sub>3</sub> and 1.5 mol of PCl<sub>5</sub>. What equilibrium concentration of Cl<sub>2</sub> must be present?



- a) 0.065 mol
- b) 0.11 mol
- c) 0.18 mol
- d) 0.24 mol
- e) 0.37 mol

I	4.5	0	1.5
C	++	++	-x
E	4.5+x	x	1.5-x

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

$$K_c = \frac{1.9}{(0.08206 \cdot 298)^{-1}}$$

$$K_c = 46.46$$

$$K_c = \frac{1.5-x}{(4.5+x)(x)} = 46.46$$

$$1.5-x = 46.46(4.5x+x^2)$$

$$1.5-x = 209.08x + 46.46x^2$$

$$46.46x^2 + 210.08x - 1.5 = 0$$

$$x = 0.0071 \text{ M} = [\text{Cl}_2]$$

25. A mixture of 7.50 g of H<sub>2</sub> and 440. g of Br<sub>2</sub> is heated in a 5.00 L vessel at 600 K. These substances react as shown below. 4.50 g of hydrogen are present at equilibrium. Calculate K.

(MW: H<sub>2</sub> = 2.00, Br<sub>2</sub> = 160.00; HBr = 80.0)



- a) 0.42
- b)  $4.8 \times 10^{-5}$
- c)  $3.9 \times 10^{-4}$
- d) 3.2
- e) 0.099

I 0.743 0.550 0

C -0.297 -0.297 +2(0.297)

E ~~0.446~~ 0.253 0.594  
D. 446M

$$7.50 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.02 \text{ g}} = \frac{3.71 \text{ mol}}{5.00 \text{ L}}$$

$$= 0.743 \text{ M}$$

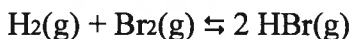
$$440.0 \text{ g Br}_2 \times \frac{1 \text{ mol Br}_2}{160.0 \text{ g}} = \frac{2.75 \text{ mol}}{5.00 \text{ L}}$$

$$= 0.550 \text{ M}$$

$$4.50 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.02 \text{ g}} = \frac{2.23 \text{ mol}}{5.0 \text{ L}} = 0.446 \text{ M}$$

$$K = \frac{[\text{HBr}]^2}{[\text{H}_2][\text{Br}_2]} = \frac{[0.594]^2}{[0.446][0.253]} = 3.13$$

26. The equilibrium constant for the following reaction is 256 at a given temperature. If 1.75 mol of H<sub>2</sub> and 1.75 mol of Br<sub>2</sub> are placed in a 1.00 L container at this temperature, what will be the equilibrium concentration of HBr?



- a) 4.18 M
- b) 3.11 M
- c) 2.44 M
- d) 1.82 M
- e) 1.14 M

I 1.75 1.75 0

C -x -x +2x

E 1.75-x 1.75-x 2x

$$K = \frac{[2x]^2}{[1.75-x]^2} = \sqrt{256}$$

$$\frac{2x}{1.75-x} = 16$$

$$2x = 16(1.75-x)$$

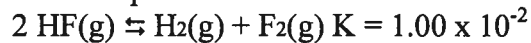
$$2x = 28 - 16x$$

$$18x = 28 \quad x = 1.556$$

$$[\text{HBr}] = 2x$$

$$= 3.11 \text{ M}$$

27. If 1.00 mole of HF(g), 0.250 mole of H<sub>2</sub>(g) and 0.075 mole of F<sub>2</sub>(g) are mixed in a 2.00 L flask, what is the reaction quotient and the net direction to reach equilibrium?



- a) Q = 0.019, the system is at equilibrium.  
 b) Q = 0.038, the equilibrium shifts to the right.  
**c) Q = 0.019, the equilibrium shifts to the left.**  
 d) Q = 0.038, the equilibrium shifts to the left.  
 e) Q = 0.019, the equilibrium shifts to the right.

$$Q = \frac{[\text{H}_2][\text{F}_2]}{[\text{HF}]^2}$$

$$[\text{HF}] = \frac{1.00 \text{ mol}}{2.00 \text{ L}} = 0.500$$

$$[\text{H}_2] = \frac{0.250 \text{ mol}}{2.00 \text{ L}} = 0.125$$

$$[\text{F}_2] = \frac{0.075 \text{ mol}}{2.00 \text{ L}} = 0.0375$$

$$Q = \frac{(0.125)(0.0375)}{[0.500]^2} = 0.01875$$

~~Q > K~~  $Q < K$  Eq shift (L)

28. At 900 K, the reaction C<sub>2</sub>H<sub>6</sub>(g) ⇌ C<sub>2</sub>H<sub>4</sub>(g) + H<sub>2</sub>(g) has K<sub>p</sub> = 0.0503. If 2.00 atm of C<sub>2</sub>H<sub>6</sub>(g) is placed in a container with a metal catalyst at this temperature, what is the equilibrium pressure of H<sub>2</sub>(g) in atm?

- a) 0.017 atm  
 b) 1.17 atm  
 c) 1.84 atm  
**d) 0.293 atm**  
 e) 0.76 atm



I	2.00 atm	0	0
C	-x	+x	+x
E	2.00 - x	x	x

$$K_p = \frac{x^2}{2.00 - x} = 0.0503$$

$$x^2 = 0.0503(2.00 - x)$$

$$x^2 = 0.1006 - 0.0503x$$

$$x^2 + 0.0503x - 0.1006 = 0$$

$$x = 0.293$$

$$P_{\text{H}_2} = x = 0.293$$

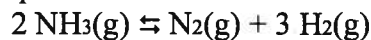
29. Consider the reactions given below. In which case(s) is product formation favored by decreasing the pressure and increasing temperature?

- 1)  $\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightleftharpoons 2 \text{NH}_3(\text{g}) \Delta H = -91.8 \text{ kJ}$
- 2)  $\text{Ni}(\text{CO})_4(\text{g}) \rightleftharpoons \text{Ni}(\text{s}) + 4 \text{CO}(\text{g}) \Delta H = 160 \text{ kJ}$
- 3)  $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2 \text{NO}_2(\text{g}) \Delta H = 58.0 \text{ kJ}$
- 4)  $2 \text{NO}(\text{g}) \rightleftharpoons \text{N}_2(\text{g}) + \text{O}_2(\text{g}) \Delta H = -181 \text{ kJ}$
- 5)  $2 \text{CO}(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2 \text{CO}_2(\text{g}) \Delta H = -566 \text{ kJ}$

a) 4 b) 2, 4 **c) 2, 3** d) 2, 4, 5 e) 2, 3, 4

- 1)  $4 \text{ mol}(\text{g}) \rightleftharpoons 2 \text{ mol}(\text{g})$   $\downarrow P$  (L) exothermic  $\uparrow \text{Temp}$  (L) ✗
- 2)  $1 \text{ mol}(\text{g}) \rightleftharpoons 4 \text{ mol}(\text{g})$   $\downarrow P$  (R) endothermic  $\uparrow \text{Temp}$  (R) ✓
- 3)  $1 \text{ mol}(\text{g}) \rightleftharpoons 2 \text{ mol}(\text{g})$   $\downarrow P$  (R) endothermic  $\uparrow \text{Temp}$  (R) ✓
- 4)  $2 \text{ mol}(\text{g}) \rightleftharpoons 2 \text{ mol}(\text{g})$   $\downarrow P$  - exothermic  $\uparrow \text{Temp}$  (L) ✗
- 5)  $3 \text{ mol}(\text{g}) \rightleftharpoons 2 \text{ mol}(\text{g})$   $\downarrow P$  (L) exothermic  $\uparrow \text{Temp}$  (L) ✗

30. Consider the following reaction at equilibrium. Le Chatelier's principle predicts that adding  $\text{N}_2(\text{g})$  to the system at equilibrium will result in



- a) a decrease in the concentration of  $\text{NH}_3(\text{g})$  ✗
- b) an increase in the value of the equilibrium constant ✗
- c) a decrease in the concentration of  $\text{H}_2(\text{g})$**
- d) removal of all the  $\text{H}_2(\text{g})$  ✗
- e) more than one of the above options. ✗

