Practice Test F1 Unit F - General Equilibrium K_p and K_c

Name______________Per____

- ***This is practice - Do NOT cheat yourself of finding out what you are capable of doing. Be sure you follow the testing conditions outlined below.***
- ***For MC, DO NOT USE A CALCULATOR. You may use ONLY a periodic table.***
- ***Try to work at a pace of about 1.5 min per MC question. Time yourself. You know how important it is that you practice working for speed.***

1. When the substances in the exothermic equation below are at equilibrium at pressure P and temperature T, the equilibrium can be shifted to favor the products by

   $MgO_{(s)} + H_{2(g)} \leftrightharpoons Mg_{(s)} + H_2O_{(g)}$

   a. increasing the pressure in the reaction vessel while keeping the temperature constant
   b. increasing the pressure by adding an inert gas such as argon
   c. decreasing the temperature
   d. allowing some hydrogen gas to escape at constant P and T
   e. adding a catalyst

$$2\,A_{(g)} + B_{(g)} \rightleftarrows 2C_{(g)}$$

2. When 0.60 mole of A and 0.75 mole of B are placed in an evacuated 1.00 L flask, the reaction represented above occurs. After the reactants and the product reach equilibrium and the initial temperature is restored, the flask is found to contain 0.30 mole of product C. Based on these results, the equilibrium constant, K_c for the reaction is

   a. 0.60
   b. 0.90
   c. 1.7
   d. 3.4
   e. 6.0

$$2\,CO + O_{2(g)} \leftrightharpoons 2\,CO_{2(g)}$$

3. Given the following data, calculate K for the reaction above.

   $C_{(s)} + \frac{1}{2}O_{2(g)} \leftrightharpoons CO_{(g)}$ $\quad K_1 = 5$

   $C_{(s)} + O_{2(g)} \leftrightharpoons CO_{2(g)}$ $\quad K_2 = 1$

   a. $\frac{1}{25}$
   b. 0.25
   c. 1
   d. 5
   e. 25

4. Which of the following systems would NOT experience a change in the number of moles of the substances present at equilibrium when the volume of the system is changed at constant temperature?

   a. $SO_{(g)} + NO_{(g)} \rightleftarrows SO_{2(g)} + \frac{1}{2}N_{2(g)}$
   b. $O_{2(g)} + 2H_{2(g)} \rightleftarrows 2H_2O_{(g)}$
   c. $N_{2(g)} + 2O_{2(g)} \rightleftarrows 2NO_{2(g)}$
   d. $N_2O_{4(g)} \rightleftarrows 2NO_{2(g)}$
   e. $CH_{4(g)} + 2O_{2(g)} \rightleftarrows CO_{2(g)} + 2H_2O_{(g)}$

5. For the reaction $2W_{(g)} \rightleftarrows 2X_{(g)} + Y_{(g)}$, the equilibrium constant, K_p, is 8×10^3 at 298 K. A mixture of the three gases at 298 is placed in a rigid metal cylinder, and the initial pressures are $P_x = 1$ atm, $P_y = 0.8$ atm, and $P_w = 2$ atm. At the instant of mixing, which of the following is true?

   a. More product will form.
   b. More reactant will form.
   c. The quantities placed in the reaction are equilibrium quantities.
   d. There is not enough information to determine if the reaction is at equilibrium or not.

$$CaCO_{3(s)} \rightleftarrows CaO_{(s)} + CO_{2(g)}$$

6. After the equilibrium represented above is established, some pure $CO_{2(g)}$ is added to the reaction vessel at constant temperature. AFTER equilibrium is reestablished, all of the following will happen EXCEPT

   a. The pressure of CO_2 in the reaction vessel remains constant.
   b. The mass of CaO would be lower than before the CO_2 was added.
   c. The concentration of $CaCO_3$ would be the same as before the CO_2 was added.
   d. K_p for the reaction will increase.
   e. K_c for the reaction will remain the same.

$$N_{2(g)} + 3\,H_{2(g)} \rightleftarrows 2\,NH_{3(g)}$$

7. Which of the following changes alone would cause a decrease in the value of K_{eq} for the exothermic reaction represented above?
   a. Removing NH_3
   b. Increasing the temperature
   c. Adding an inert gas such as argon
   d. Adding more hydrogen gas at constant temperature
   e. Adding a catalyst

$$2SO_{2(g)} + O_{2(g)} \rightleftarrows 2SO_{3(g)} + 198\text{ kJ}$$

8. Consider the equilibrium above. Which of the following changes will increase the concentration of $SO_{2(g)}$?
   a. Increasing the concentration of $O_{2(g)}$
   b. Increasing the pressure in the reactant vessel at constant temperature
   c. Increasing the mass of SO_3 present
   d. Adding a catalyst
   e. Decreasing the temperature

9. Consider the reaction system,
   $NiO_{(s)} + H_{2(g)} \leftrightharpoons Ni_{(s)} + H_2O_{(g)}$
   The equilibrium constant expression is
   a. $K_{eq} = \dfrac{[Ni][H_2O]}{[NiO][H_2]}$
   b. $K_{eq} = \dfrac{[Ni]}{[NiO][H_2]}$
   c. $K_{eq} = \dfrac{1}{[NiO][H_2]}$
   d. $K_{eq} = \dfrac{1}{[H_2]}$
   e. $K_{eq} = \dfrac{[H_2O]}{[H_2]}$

10. Given the equilibrium,
$$2SO_{2(g)} + O_{2(g)} \leftrightharpoons 2SO_{3(g)}$$
    If this equilibrium is established by beginning with equal number of moles of SO_2 and O_2 in an empty 1.0 L bulb, then the following must be true at equilibrium:
    a. $[SO_2] = [SO_3]$
    b. $2[SO_2] = 2[SO_3]$
    c. $[SO_2] = [O_2]$
    d. $[SO_2] < [O_2]$
    e. $[SO_2] > [O_2]$

11. Consider: $N_2O_{4(g)} \leftrightharpoons 2NO_{2(g)}$
    - At 25°C, in a 1.0 L container 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 equilibrium.
    - At 90°C, 0.11 mole of N_2O_4 results in 0.050 mole of N_2O_4 and 0.12 mole of NO_2 at equilibrium. From these data we can conclude

    a. Equilibrium constants, in general, are always larger at higher temperatures.
    b. N_2O_4 molecules react by a first order rate law.
    c. the reaction is endothermic.
    d. NO_2 molecules react twice as fast as N_2O_4 molecules at 90°C.
    e. the equilibrium constant for the reaction above decreases with an increase in temperature.

The next 2 questions refer to the following:

At a given temperature, 0.300 mole NO, 0.160 mol Cl_2 and 0.500 mol ClNO were placed in a 10.0 Liter container. The following equilibrium is established:

$$2ClNO_{(g)} \leftrightharpoons 2NO_{(g)} + Cl_{2(g)}$$

At equilibrium, 0.600 mol of ClNO was present

12. The number of moles of Cl_2 present at equilibrium is
    a. 0.090
    b. 0.050
    c. 0.060
    d. 0.200
    e. 0.110

13. The equilibrium constant, K_c, is closest to:
    a. 10^{-6}
    b. 10^{-3}
    c. 1
    d. 10^{3}
    e. 10^{6}

14. At 985°C, the equilibrium constant for the reaction,
$$H_{2(g)} + CO_{2(g)} \leftrightharpoons H_2O_{(g)} + CO_{(g)}$$
    is 1.5. What is the equilibrium constant for the reaction below?
$$H_2O_{(g)} + CO_{(g)} \leftrightharpoons H_{2(g)} + CO_{2(g)}$$
    a. −1.5
    b. 0.75
    c. 3.0
    d. 0.67
    e. There is no way of knowing what it would be.

15. For the reaction system,

$$N_{2(g)} + 3\ H_{2(g)} \leftrightharpoons 2NH_{3(g)} + \text{heat}$$

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

16. Solid HgO, liquid Hg, and gaseous O_2 are placed in a glass bulb and are allowed to reach equilibrium at a given temperature.

$$43.4\ \text{kJ} + 2HgO_{(s)} \leftrightharpoons 2Hg_{(L)} + O_{2(g)}$$

The mass of HgO in the bulb could be increased by

a. adding more Hg.
b. removing some O_2.
c. reducing the volume of the bulb.
d. increasing the temperature.
e. removing some Hg.

17. In a particular balanced equation, the number of moles of gaseous reactants is more than the number of moles of gaseous products. The forward reaction is known to be endothermic. Which of the following is always true of this reaction?

a. An increase in pressure and an increase in temperature lead to a shift toward the products.
b. An increase in pressure and an increase in the concentration of reactants lead to a shift toward the reactants.
c. A decrease in temperature and an increase in the concentration of reactants lead to a shift toward the products.
d. An increase in pressure and the use of a catalyst lead to a shift toward the reactants.
e. A decrease in pressure and an increase in temperature lead to a shift toward the products.

18. For the equilibrium system

$$H_2O_{(g)} + CO_{(g)} \leftrightharpoons H_{2(g)} + CO_{2(g)} + 42\ \text{kJ}$$

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

19. Consider the reaction

$$SO_{2(g)} + \tfrac{1}{2}O_{2(g)} \leftrightharpoons SO_{3(g)} \quad K_c = 49 \text{ at } 1000\ \text{K}$$

What is the value of K_c for the reaction below?

$$2SO_{3(g)} \leftrightharpoons 2SO_{2(g)} + O_{2(g)}$$

a. $\frac{1}{49}$
b. $\frac{1}{7}$
c. $\frac{1}{(49)^2}$
d. 7
e. $(49)^2$

20. Which of the following causes the equilibrium constant, K, to change for any reaction?

a. Adding more reactants
b. Adding more products
c. Increasing the pressure
d. Changing the temperature
e. Adding catalyst

For the next three questions, consider an equilibrium system based on the reaction below.

$$heat + N_2O_{4(g)} \leftrightharpoons 2NO_{2(g)}$$

21. Which occurs when the volume of the system is increased at constant temperature?

	#molecules of NO_2	**total # of molecules of all gases**	**K_p**
a.	increases	increases	remains same
b.	increases	decreases	remains same
c.	increases	remains same	remains same
d.	remains same	decreases	decreases
e.	remains same	increases	decreases

22. Which occurs when a catalyst is added?

	partial pressure of N_2O_4	**total pressure of all gases**	**K_p**
a.	decreases	increases	remains same
b.	increases	decreases	remains same
c.	stays the same	stays the same	remains same
d.	increases	decreases	decreases
e.	increases	increases	decreases

23. Which occurs when the temperature is increased at constant volume?

	#molecules of N_2O_4	**total # of molecules of all gases**	**K_p**
a.	increases	decreases	remains same
b.	decreases	increases	remains same
c.	increases	decreases	increases
d.	decreases	decreases	increases
e.	decreases	increases	increases

24. At a particular temperature, the equilibrium partial pressure of SO_2, O_2, and SO_3 are found to be 1.0 atm, 0.50 atm, and 4.0 atm respectively. Calculate K_p for the following reaction.

$$2SO_{2(g)} + O_{2(g)} \leftrightharpoons 2SO_{3(g)}$$

a. $\frac{0.50}{16}$
b. 4.0
c. 2.0
d. 8.0
e. 32

The next four questions apply to an equilibrium system based on the reversible reaction given below.

$$N_{2(g)} + 3H_{2(g)} \leftrightharpoons 2NH_{3(g)} + 92\text{ kJ}$$

25. When the temperature of such an equilibrium system is increased at constant volume, which property is least affected?
a. density
b. pressure
c. concentration of $NH_{3(g)}$
d. concentration of $N_{2(g)}$
e. average kinetic energy

26. Which observation confirms the fact that equilibrium has been reached in such a system confined in a closed, rigid container?
a. The density remains constant.
b. The odor of ammonia can first be detected.
c. The pressure is decreasing at a constant rate.
d. The partial pressure of hydrogen remains constant.
e. The mass of the system has decreased to a constant value.

27. The equilibrium system described by the reaction, held at a constant temperature, is subjected to the addition of $N_{2(g)}$, and the system proceeds to a new equilibrium position. Which of the following statements is *false*?
a. Heat will be released to the environment.
b. The $[N_2]$ will end up smaller than the $[N_2]$ before the addition.
c. The $[H_2]$ will end up smaller than the $[H_2]$ before the addition.
d. The $[NH_3]$ will end up larger than the $[NH_3]$ before the addition.
e. K_c will remain constant.

28. How are the rates of the opposing reactions and the value of the equilibrium constant, K_p, affected when heat is added to this system?
a. Bothe the forward and reverse reaction rates remain unchanged and K_p decreases.
b. The forward reaction will be faster initially and K_p decreases.
c. The forward reaction will be faster initially and K_p increases.
d. The reverse reaction will be faster initially and K_p decreases.
e. The reverse reaction will be faster initially and K_p increases.

The next five questions refer to the following reaction and the information provided.

$$4H_{2(g)} + CS_{2(g)} \leftrightharpoons CH_{4(g)} + 2H_2S_{(g)}$$

A mixture of 2.50 mol $H_{2(g)}$, 1.50 mol $CS_{2(g)}$, 1.5 mol $CH_{4(g)}$ and 2.00 mol $H_2S_{(g)}$ is placed in a 5.0 L rigid reaction vessel, and the system reaches equilibrium according to the equation above. When the equilibrium is achieved, the concentration of $CH_{4(g)}$ has become 0.25 mol L^{-1}.

29. Changes in concentration occur as this system approaches equilibrium. Which expression gives the best comparison of the changes in those concentrations shown in the ratio to the right? $\dfrac{\Delta[H_2S]}{\Delta[CS_2]}$

    a. $\dfrac{+2}{+1}$
    b. $\dfrac{+2}{-1}$
    c. $\dfrac{-2}{+1}$
    d. $\dfrac{-1}{+1}$
    e. $\dfrac{-1}{+2}$

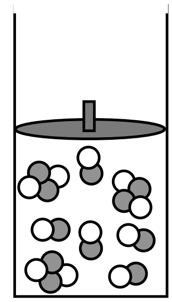
30. What is the change in the number of moles of $H_2S_{(g)}$ present as the system moves from its original state to the equilibrium described?
    a. −2.50
    b. 1.25
    c. −0.50
    d. −0.25
    e. −0.10

31. What is the number of moles of $CS_{2(g)}$ at equilibrium?
    a. 0.25
    b. 0.35
    c. 0.75
    d. 1.25
    e. 1.75

32. What is the concentration in moles per liter of $H_{2(g)}$ at equilibrium?
    a. 0.50
    b. 0.70
    c. 1.0
    d. 2.0
    e. 3.5

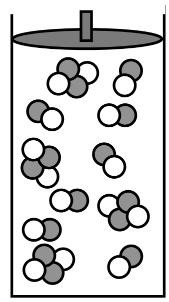
$$2\,NO_{(g)} \leftrightharpoons N_2O_{2(g)}$$

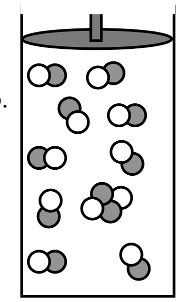
33. Consider an equilibrium mixture of $NO_{(g)}$ and $N_2O_{2(g)}$ at some temperature. The following diagram above shows a representation of an equilibrium mixture at some temperature.

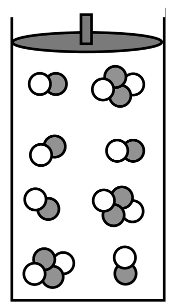


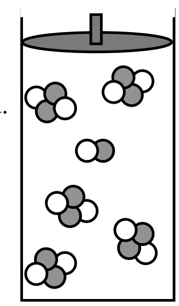
The piston is drawn up, and equilibrium will be re-established.

Which of the following is a possible representation of the molecules in the container after equilibrium is re-established?

a.


b.


c.


d.


Unit F - General Equilibrium K_p and K_c

1. c The negative ΔH tells us this is an exothermic reaction, and decreasing the temperature will cause the reaction to shift to the product side to replace the heat. Pressure changes will not cause any shift since there are the same number of moles of gas on both sides of the equation. Adding inert gas would not have any effect as it would increase the total pressure but not change the partial pressure of any of the individual gases. Allowing H_2 to escape will decrease a reactant and cause the equilibrium to shift to replace H_2 by shifting left. Adding a catalyst will only make both the forward and reverse reactions occur faster, but will not shift the equilibrium position.

2. c It would be wise to make a RICE box for this problem, resulting in the equilibrium values shown in the E row. It is a 1 L container, so moles = molarity. Then plug these values into the equilibrium expression as shown. $K = \frac{(.3)^2}{(0.3)^2(0.6)}$ which reduced to $\frac{10}{6}$ and without a calculator, it would be best to scan the answers and see that c is the obvious choice.

R	2A +	B ⇆	2C
I	0.6	0.75	
C			
E			0.30

3. a $2\,CO + O_{2(g)} \leftrightharpoons 2\,CO_{2(g)}$

   This is a Hess' Law problem, the first reaction must be reversed and double the stoichiometry − this means that the K value must be inversed and squared, which can then be added to the second reaction which should be doubled.

   $$\frac{1}{(K_1)^2} \times (K_2)^2 = \frac{1}{(5)^2} \times (1)^2 = \frac{1}{25}$$

4. e Each of the reactions represent gaseous equilibria. When the volume of a system changes, at a constant equilibrium, the pressure changes inversely. LeChatelier's principle will cause the system to shift to relieve the stress. When the number of gaseous moles are the same on both sides, there is no shift in the equilibrium position.

5. a insert the initial values into the equilibrium expression $Q = \frac{(1)^2(0.8)}{(2)^2}$ and solve for Q = 0.2 thus K_p > Q and thus the reaction must proceed in the forward direction. It is likely that the ΔS value is positive since more particles form, and ΔG for this reaction will be negative since the forward reaction occurs.

6. d After the initial equilibrium was established and more CO_2 gas is added, the reaction will shift in favor of more reactants to use up the extra carbon dioxide. While the equilibrium position changes, the K_p and K_c values do not change. The quantities of the solid constituents change, but their concentrations do not change.Since $K_p = (CO_2)$, the pressure would have to return to the same value to reestablish equilibrium. This of course would require that there is enough CaO in the container to use up all of the introduced CO_2. If there were not enough CaO, then equilibrium would not be able to be re-established again.

7. b The exothermic reaction will shift towards products when the temperature is increased which results in a decrease in the K_{eq} value. Decreasing the temperature would cause a shift to the right which would increase K_{eq}. Adding an inert gas or a catalyst will cause no shift in the equilibrium position, and adding more hydrogen gas will cause a positional shift to the right, but no change in the Keq value.

8. c Increasing the mass of SO_3 will cause a shift to the left. Adding a catalyst will cause no shift, and the other three options all cause shift to the right.

9. e Remember that solids are not included in the equilibrium expression.

10. d We have know information in this problem to tell us where the equilibrium position will land. But since there are only reactants started with, we know that the reaction must proceed in the forward direction, and stoichiometry will prevail, thus we can know that we will lose 2 moles of SO_2 for every one mole of O_2 which means SO_2 will deplete faster resulting in less SO_2 than O_2 present at equilibrium, wherever it may land.

11. c The information in this problem tells us that at a higher temperature, the equilibrium position is further to the right, or in other words, more product than at cooler temps. This indicates that the reaction is endothermic as the "stress" of the energy will drive the reaction toward products in an "attempt" to use up some of the added heat.

12. e Work a RICE box, I think you will find that there will be 0.11 mole of Cl_2 at equilibrium.

R	2ClNO ⇆	2NO +	Cl_2
I	0.5	0.3	0.16
C	+0.1		-0.05
E	0.6		

13. b First you must change to moles to molarity, then without a calculator, you should be able to come close to 1×10^{-3}

14. d Since the reaction has been reversed, the equilibrium constant must be inverted. The inverse of 1.5 which is $^3/_2$ is equal to ⅔ or 0.67

15. d All the substances are gases, thus low temp will favor the reaction toward the energy and increasing pressure will favor the product side with less molecules than the reactant side.

16. c Beware of the solids and liquids since as long as some is present, changing their quantities does not effect equilibrium. Thus a and e have no effect on the equilibrium position. Options b and d will cause the position to shift to the right.

17. a Since the forward reaction is endothermic, raising temp will favor the forward reaction, and since the number of moles of reactants is more than the number of moles of gaseous products, an increase of pressure due to reduced volume will also cause a shift to the right.

18. a Calculate that Q = 1 and since K_c, < Q, thus the reaction will proceed to the left to achieve equilibrium.

19. c Reversing the reaction requires the inverse of the K, and doubling the coefficients requires squaring of K, thus option c will take care of both.

20. d The equilibrium is constant (even when adding a catalyst) except when the temperature is changed, and K will be different. Whether it is larger or smaller depends on if the reaction is exothermic or endothermic, and if the temperature is increased or decreased.

21. a When the volume is increased, the reaction will shift to the side of more gas molecules.

22. c Addition of a catalyst does not upset equilibrium, it simply would get you to equilibrium faster, if if already there, it would simply speed the rate of the forward and reverse reactions equally , not affecting the equilibrium position.

23. e Remember that K is only constant at a particular temp and any change in temp will cause a shift that will affect the magnitude of K. Thus an increase in volume will cause a shift in the equilibrium shift toward more molecules which is to the right which decreases the N_2O_4 and will increase the value of K

24. e $K_p = \dfrac{(4)^2}{(1)^2(0.5)}$ this simplifies to 32

25. a The change in temperature will cause a shift left because the reaction is exothermic. The density of the system will not change because the mass of the system will remain constant.

26. d In a closed rigid container, the density will remain constant. Presence of any NH_3 would give an ammonia smell, though in a closed container, you would have a hard time smelling it! At equilibrium, the pressure would be constant. Like (a), in a closed system, the mass will be constant always. When the pressure of any and all of the gases remain constant, equilibrium has been achieved.

27. b The stress of adding a reactant will cause a shift right to try to eliminate some of the reactant added, yet the amount will never go below the original amount, a new equilibrium position will have been reached with less N_2, and more NH_3 , thus (b) is the false statement. In the process of shifting to the right, energy will be produced, and in order to maintain a constant temperature, heat must be lost to the surroundings.

28. d Any time a system is heated, both the forward and reverse reactions will increase in speed, however, since the reaction is exothermic, adding heat will initially cause the reverse reaction to speed up more than the forward reaction, until eventually both forward and reverse reactions end up at the same speed as each other (although at a faster speed than the system before any heat was added) and new equilibrium position is achieved with a new, lower K_p value.

R	$4H_{2(g)}$ +	$CS_{2(g)}$ $\rightleftharpoons$	$CH_{4(g)}$ +	$2H_2S_{(g)}$
I	2.5	1.5	1.5	2
C	+1.0	+0.25	−0.25	−0.50
E	3.5	1.75	1.25	

29. c Since at equilibrium, CH_4 is less than the starting quantity, thus the reaction must have shifted left, for a loss of H_2S and a gain of CS_2, and the stoichiometry is 2:1.

30. c In this question, it is important to pay attention to moles and molarity, and the fact that the reaction is done in a 5 L container. The equilibrium concentration for CH_4 0.25 M × 5 L is actually 1.25 mol, this means that there was a loss of 0.25 mole of to reach equilibrium. Fill in the Rice Box shown to the right. The quantities known from the information in the problem are shaded, the rest must be deduced calculated.

31. e Use the rice box to the right

32. b Use the rice box shown, but don't forget to change the moles to Molarity. $\dfrac{3.5mol}{5L} = 0.70M$

33. b The enlarged container would cause a lower partial pressure of both gases, thus the equilibrium to shift towards the left which will make more molecules to bring the system back to equilibrium at a new equilibrium position. Thus option (b) is the only one in which there are fewer N_2O_4 molecules.