*JOE PERFECT*

Practice Test Chapter 15
Equilibrium

Part I: _Circle the letter of the best annswer(s). No calculators are allowed on the multiple choice portion of the practice test!_

> **Target 1:** I can write the equilibrium expression for a balanced chemical equation.

1. At equilibrium . . .
   a. all chemical processes have ceased.
   b. the rate of the forward reaction equals that of the reverse reaction.
   c. the rate constant of the forward reaction equals the rate constant of the reverse reaction.
   d. both the rate of the forward reaction equals that of the reverse reaction and the rate constant for the forward reaction equals the rate constant of the reverse reaction.
   e. None of the above.

2. Considering the equation below, which expression is correct for K_{eq}?

   $$2\,SO_2\,(g)\ +\ O_2\,(g)\ \rightleftharpoons\ 2\,SO_3\,(g)$$
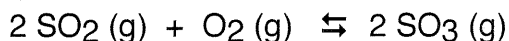
   a. $K_{eq} = \dfrac{[SO_3]^2}{[SO_2]^2\,[O_2]}$

   b. $K_{eq} = \dfrac{[2SO_3]}{[2SO_2]\,[O_2]}$

   c. $K_{eq} = \dfrac{[SO_2]^2\,[O_2]}{[SO_3]^2}$

   d. $K_{eq} = \dfrac{[2SO_3]^2}{[2SO_2]^2\,[O_2]}$

   Always products in numerator!

3. What reaction has the equilibrium expression . . . $[D]^2\,/\,[A]\,[B]\,[C]$?
   a. $2\,D \rightleftharpoons A + B + C$
   b. $2\,D \rightleftharpoons ABC$
   c. $A + B + C \rightleftharpoons 2\,D$
   d. $ABC \rightleftharpoons 2\,D$
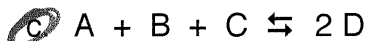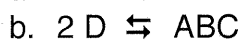
   $$Keq = \dfrac{[D]^2}{[A][B][C]}$$

   products
   reactants

4. Considering the equation below, which expression is correct for Keq?

   $$CuS(s) + O_2(g) \rightleftharpoons Cu(s) + SO_2(g)$$

   a. $Keq = \dfrac{[SO_2][Cu]}{[CuS][O_2]}$

   b. $Keq = \dfrac{[CuS][O_2]}{[Cu][SO_2]}$

   c. $Keq = \dfrac{[SO_2]}{[O_2]}$

   d. $Keq = \dfrac{[O_2]}{[SO_2]}$

   Never include solids or liquids in equilib. expression.

> Target 2: I can numerically evaluate K_C and K_p from a knowledge of the equilibrium concentrations (or pressures) of reactants or products, or from the initial concentrations and the equilibrium concentration of at least one substance.

5. The relationship between the rate constants for the forward and reverse reactions and the equilibrium constant for a process is . . .

   a. $K_{eq} = K_f \cdot K_r$

   b. $K_{eq} = K_f - K_r$

   c. $K_{eq} = K_f + K_r$

   d. $K_{eq} = K_f / K_r$ ⟵ (circled)
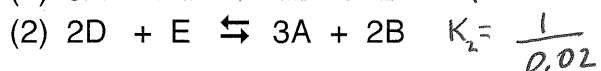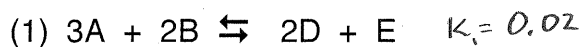
   e. $K_{eq} = 1/(K_f \cdot K_r)$

   EXAMPLE: $A \rightleftharpoons B$

   Forward: rate $= k_f [A]$ ⎫ rates are equal!!
   Reverse: rate $= k_r [B]$ ⎭

   $k_f [A] = k_r [B]$

   $\dfrac{[B]}{[A]} = \dfrac{K_f}{K_r} = K_{eq}$

6. The equilibrium constant for reaction (1) below is 0.02. The value of the equilibrium constant for reaction (2) is _____.

   (1) $3A + 2B \rightleftharpoons 2D + E$ $K_1 = 0.02$
   (2) $2D + E \rightleftharpoons 3A + 2B$ $K_2 = \dfrac{1}{0.02}$

   a. 0.02
   b. 0.04
   c. -0.02
   d. 50 ⟵ (circled)
   e. 3.0×10^3

   $\dfrac{1}{0.02} = \dfrac{1}{2/100} = \dfrac{100}{2} = 50$

7. Which **one** of the following will change the value of an equilibrium constant?

   a. Changing the temperature of the reaction. ⟵ (circled)
   b. Adding other substances that do not react with any of the species involved in the equilibrium.
   c. Varying initial concentration of reactants.
   d. Varying initial concentration of products.
   e. All of these will affect the value of the equilibrium constant.

8. Consider the following equilibrium: $2\,SO_2(g) + O_2(g) \rightleftharpoons 2\,SO_3(g)$

   From which of the following starting conditions would it be impossible for this equilibrium to be achieved?

   a. 1.0 mol $SO_3(g)$ in a 1.0 liter container
   b. 0.25 mol $SO_2(g)$, 0.25 mol $O_2(g)$ in a 1.0 liter container
   c. 0.25 mol $SO_2(g)$, 0.50 mol $O_2(g)$, and 0.10 mol $SO_3(g)$ in a 1.0 liter container
   d. 0.50 mol $O_2(g)$ and 0.50 mol $SO_3(g)$ in a 1.0 liter container
   e. Equilibrium can be achieved from any of these starting conditions ⟵ (circled)

$K_c \approx 1$
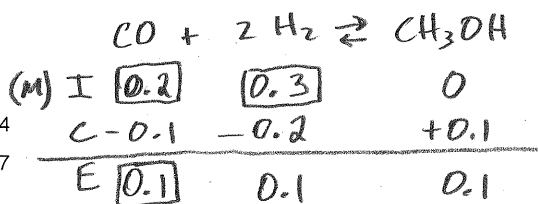
9. The reaction: $X + Y \rightleftharpoons 2\,M$ has a $K_c = 0.89$ at 672 K. At equilibrium_____.
   a. products predominate substantially
   b. reactants predominate substantially
   (c) roughly equal molar amounts of products and reactants are present
   d. only products exist
   e. only reactants exist

If $K \gg 1$ ... products favored.
If $K \ll 1$ ... reactants favored.

10. Consider the following reaction: $CO(g) + 2\,H_2\,(g) \rightleftharpoons CH_3OH\,(g)$
    In an experiment, 0.200 mol of CO and 0.300 mol of H_2 were placed in a 1.00-L reaction vessel. At equilibrium, there were 0.100 mol of CO remaining. Calculate the K_{eq} of the equilibrium.
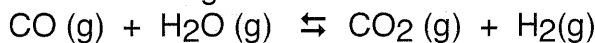   a. 0.0100
    b. 1.00
    (c) 100.
    d. 1.00×10^4
    e. 1.00×10^7

$$CO \; + \; 2\,H_2 \rightleftharpoons CH_3OH$$

(M)	I	$\boxed{0.2}$	$\boxed{0.3}$	0
	C	-0.1	-0.2	$+0.1$
	E	$\boxed{0.1}$	0.1	0.1

$$K = \frac{[CH_3OH]}{[CO][H_2]^2} = \frac{0.1}{(0.1)(0.1)^2}$$

$$K = \frac{1}{(0.1)^2} = \frac{1}{0.01} = \frac{1}{1/100}$$

$$\boxed{= 100}$$

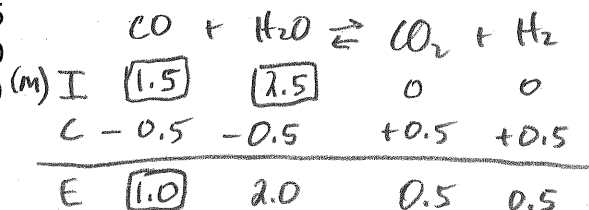11. In the coal-gasification process, carbon monoxide is converted to carbon dioxide via the following reaction:
    $$CO\,(g) + H_2O\,(g) \rightleftharpoons CO_2\,(g) + H_2(g)$$

    convert to MOLARITY before using ICE table.

    In an experiment, 3.00 mol of CO and 5.00 mol of H_2O were placed in a 2.00-L reaction vessel. At equilibrium, there were 2.00 mol of CO remaining. K_{eq} at the temperature of the experiment is ___.
    (a) 0.125
    b. 0.250
    c. 0.500
    d. 1.00
    e. 2.00

$$CO \; + \; H_2O \rightleftharpoons CO_2 + H_2$$

(M)	I	$\boxed{1.5}$	$\boxed{2.5}$	0	0
	C	-0.5	-0.5	$+0.5$	$+0.5$
	E	$\boxed{1.0}$	2.0	0.5	0.5

$$K = \frac{(0.5)(0.5)}{2} = \frac{0.25}{2}$$

$$K = 0.125$$
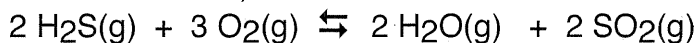

Target 3: I can interconvert K_c and K_p.

12. For the reaction below, $K_p = 2.1 \times 10^4$ at 400 K.
    $$2\,A(g) + B(s) \rightleftharpoons C(g) + D(g)$$

What is the value of K_c for the above reaction at this temperature?
    a. 4.2×10^8 b. 3.5×10^1 c. 1.0×10^9 d. 0.50×10^4 (e) 2.1×10^4

$$K_p = K_c (RT)^{\Delta n} \; ... \; \text{since } \Delta n = 2 - 2 = 0, \; K_p = K_c!$$

NO MATH NEEDED ☺

13. For the following reaction at 25°C, $K_c = 3.0 \times 10^5$. What is K_p? (You can use a calculator for this one!)

$$2\ H_2S(g) + 3\ O_2(g) \rightleftharpoons 2\ H_2O(g) + 2\ SO_2(g)$$

    ⓐ 1.2 x 10⁴
    b. 8.2 x 10⁻⁵
    c. 3.3 x 10⁻⁶
    d. 3.0 x 10⁵
    e. 7.3 x 10⁶

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

Don't forget to change to KELVIN.

$K_p = (3.0 \times 10^5)(0.0821 \cdot 298)^{4-5} = \boxed{1.2 \times 10^4}$

Target 4: I can use the equilibrium constant to calculate equilibrium concentrations.

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14. Consider the following equilibrium at 27°C:

$$NH_4HS(s) \rightleftharpoons NH_3(g) + H_2S(g); \quad K_p = 0.16$$
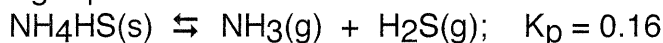
A sample of solid NH_4HS is placed in a closed vessel and allowed to equilibrate. Calculate the equilibrium partial pressure (atm) of ammonia, assuming that some solid NH_4HS remains.

    ⓐ 0.40
    b. 0.16
    c. 0.32
    d. 0.64
    e. 0.0032

$NH_4HS(s) \rightleftharpoons NH_3(g) + H_2S(g)$

(atm) I   0   0
C   +x   +x
E   x   x

$K_p = x^2$

$0.16 = x^2$

$X = 0.4$

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15. Consider the following equilibrium: $2\ AB(g) \rightleftharpoons A_2(g) + B_2(g)$; $K_c = 16$
Assume that the initial concentration of AB is 2.0 M and that there is no A_2 or B_2 originally present. Calculate the equilibrium concentration of B_2.

    a. 0.0034
    b. 0.00049
    c. 1.5
    d. -0.104
    ⓔ 0.89

$2\ AB \rightleftharpoons A_2 + B_2$

(M) I 2   0   0
C −2x +x +x
E 2−2x x x

$16 = \dfrac{x^2}{(2-2x)^2}$ (take $\sqrt{}$ of both sides)

$4 = \dfrac{x}{2-2x}$; $8-8x = x$

$8 = 9x$

$X = 8/9$

↳ choice "E" is only choice reasonable.

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16. Consider the following reaction:

$$CuS(s) + O_2(g) \rightleftharpoons Cu(s) + SO_2(g)$$

A reaction mixture initially contains 3.0 M O_2. Determine the equilibrium concentration of O_2 if K_c for the reaction at this temperature is 2.0.

    a. 0.10 M
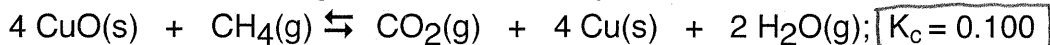   b. 0.20 M
    c. 0.50 M
    ⓓ 1.0 M
    e. 2.0 M

$CuS(s) + O_2(g) \rightleftharpoons Cu(s) + SO_2(g)$

(M) I   3   0
C   −x   +x
E   3−x   x

$K = \dfrac{[SO_2]}{[O_2]} = \dfrac{x}{3-x} = 2$; $6-2x = x$; $6 = 3x$; $X = 2$; $[O_2] = 3-x$ $[O_2] = 3-2 = 1$

17. Consider the following reaction and its equilibrium constant:

$4 CuO(s) + CH_4(g) \rightleftharpoons CO_2(g) + 4 Cu(s) + 2 H_2O(g);$ $K_c = 0.100$

A reaction mixture contains 2.0 M CH_4, 0.50 M CO_2 and 0.80 M H_2O. Which of the following statements is TRUE concerning this system?

a. The reaction will shift in the direction of reactants.
b. The equilibrium constant will increase.
c. The reaction quotient will increase.
d. The reaction will shift in the direction of products.
e. The system is at equilibrium.

$$Q = \frac{(0.5)(0.8)^2}{2} = \frac{(0.25)(0.64)}{2}$$

$$Q = 0.16$$

Since $K < Q$ ... shift left!

18. Consider the following reaction:

$I_2(g) + Br_2(g) \rightleftharpoons 2 IBr(g)$ $K_c = 1.1 \times 10^2$ $\Delta H = 40.88$ kJ

A 3.0 liter vessel was found to contain 3.0 moles of I_2, 6.0 moles of Br_2 and 3.0 moles of IBr. Is the system at equilibrium? If not, which reaction will be favored in order for the reaction to achieve equilibrium?

see work below

a.   Yes, the reaction is at equilibrium.
b.   No, the endothermic reaction will be favored in order to reach equilibrium.
c.   No, the exothermic reaction will be favored in order to reach equilibrium.

19. Consider the following endothermic reaction:

HEAT + $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$
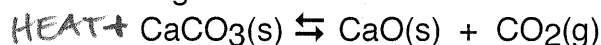
Which of the following would shift the equilibrium position to form more CO_2 gas?

convert to molarity!

a. Increasing the system temperature.
b. Decreasing the system temperature.
c. Increasing the system pressure.
d. Decreasing the volume of the system.
e. Adding neon gas to the system.

#18

$$Q = \frac{[IBr]_0^2}{[I_2][Br_2]_0}$$

$$Q = \frac{1^2}{1 \cdot 2} = 0.5$$

$K > Q$ ... shifts right
forward rxn. is ENDO!

20. In which of the following reactions would increasing pressure at constant temperature not change the concentration of reactants and products.?

    a. $N_2(g) + 3 H_2(g) \leftrightarrows 2 NH_3(g)$

    b. $N_2H_4(g) \leftrightarrows 2 NO_2(g)$

    c. $N_2(g) + 2 O_2(g) \leftrightarrows 2 NO_2(g)$

    d. $2 N_2(g) + O_2(g) \leftrightarrows 2 N_2O(g)$

    e. $N_2(g) + O_2(g) \leftrightarrows 2 NO(g)$ Because you have same # moles of gas as reactants & products!

21. Consider the following reaction at equilibrium:

HEAT + $2 NH_3(g) \leftrightarrows N_2(g) + 3 H_2(g)$; $\Delta H^\circ = +92.4$ kJ
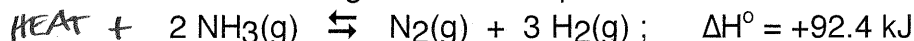
This reaction will shift to the right with _____.

    a. increasing both temperature and pressure

    b. increasing temperature and decreasing pressure

    c. decreasing both temperature and pressure

    d. decreasing temperature and increasing pressure

    e. the addition of extra N_2 to the container

22. Consider the following reaction at equilibrium:

    $2 SO_2(g) + O_2(g) \leftrightarrows 2 SO_3(g)$ $\Delta H^\circ = -99$ kJ
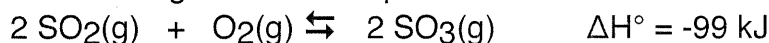
Le Châtelier's Principle predicts that an increase in temperature will result in _____.

    a. a decrease in the partial pressure of SO_2

    b. a decrease in the partial pressure of SO_3

    c. an increase in K_{eq}

    d. the partial pressure of O_2 will decrease

    e. no changes in equilibrium partial pressures

Target 7: I can describe the effect of a catalyst on a system as it approaches equilibrium.

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23. The effect of a catalyst on an equilibrium is to _____.

    a. increase the rate of the forward reaction only

    b. increase the equilibrium constant so that products are favored

    c. slow the reverse reaction only

    d. increase the rate at which equilibrium is achieved without changing the composition of the equilibrium mixture

    e. shift the equilibrium to the right

Part II: Solve each of the following problems on separate sheets of paper. Show all of your work and label your answer with appropriate units. Please clearly indicate your final answer by circling it!

1.   The compound butane, C_4H_{10}, occurs in two isomeric forms, *n*-butane and isobutane (2-methyl propane). Both compounds exist as gases at 25°C and 1.0 atm.

(a) Draw the structural formula of each of the isomers (include all atoms). Clearly label each structure.

(b) On the basis of molecular structure, identify the isomer that has the higher boiling point. Justify your answer.

The two isomers exist in equilibrium as represented by the equation below.

n-butane(g) $\leftrightarrows$ isobutane(g) $K_C = 2.5$ at 25°C
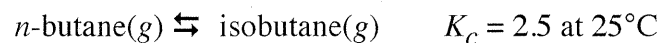
Suppose that a 0.010 mol sample of pure *n*-butane is placed in an evacuated 1.0 L rigid container at 25°C.

(c) Write the expression for the equilibrium constant, K_C, for the reaction.

(d) Calculate the initial pressure in the container when the *n*-butane is first introduced (before the reaction starts).

(e) The *n*-butane reacts until equilibrium has been established at 25°C.
   (i)   Calculate the total pressure in the container at equilibrium. Justify your answer.
   (ii) Calculate the molar concentration of each species at equilibrium.
   (iii) If the volume of the system is reduced to half of its original volume, what will be the new concentration of *n*-butane after equilibrium has been reestablished at 25°C? Justify your answer.

Suppose that in another experiment a 0.010 mol sample of pure isobutane is placed in an evacuated 1.0 L rigid container and allowed to come to equilibrium at 25°C.

(f) Calculate the molar concentration of each species after equilibrium has been established.

2. Consider the following reaction:

$$C(s) + CO_2(g) \rightleftharpoons 2\,CO(g)$$
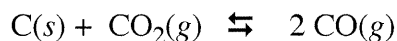
Solid carbon and carbon dioxide gas at 1,160 K were placed in a rigid 2.00 L container, and the reaction represented above occurred. As the reaction proceeded, the total pressure in the container was monitored. When equilibrium was reached, there was still some C(s) remaining in the container. Results are recorded in the table below.

Time (hours)	Total pressure of the gases in the container at 1,160 K
0.0	5.00
2.0	6.26
4.0	7.09
6.0	7.75
8.0	8.37
10.0	8.37

(a)   Write the expression for the equilibrium constant, Kp , for the reaction.

(b)   Calculate the number of moles of $CO_2(g)$ initially placed in the container. (Assume that the volume of the solid carbon is negligible.)

(c)   For the reaction mixture at equilibrium at 1,160 K, the partial pressure of the $CO_2(g)$ is 1.63 atm. Calculate:
       (i) the partial pressure of $CO(g)$, and
       (ii) the value of the equilibrium constant, Kp .

(d)   If a suitable solid catalyst were placed in the reaction vessel, would the final total pressure of the gases at equilibrium be greater than, less than, or equal to the final total pressure of the gases at equilibrium without the catalyst? Justify your answer. (Assume that the volume of the solid catalyst is negligible.)

In another experiment involving the same reaction, a rigid 2.00 L container initially contains 10.0 g of C(s), plus $CO(g)$ and $CO_2(g)$, each at a partial pressure of 2.00 atm at 1,160 K.

(e) Predict whether the partial pressure of $CO_2(g)$ will increase, decrease, or remain the same as this system approaches equilibrium. Justify your prediction with a calculation.

3. Answer the following questions regarding the decomposition of arsenic pentafluoride, $AsF_5(g)$.

(a) A 55.8 g sample of $AsF_5(g)$ is introduced into an evacuated 10.5 L container at 105°C.

(i) What is the initial molar concentration of $AsF_5(g)$ in the container?

(ii) What is the initial pressure, in atmospheres, of the $AsF_5(g)$ in the container? At 105°C, $AsF_5(g)$ decomposes into $AsF_3(g)$ and $F_2(g)$ according to the following chemical equation.

$$AsF_5(g) \leftrightarrows AsF_3(g) + F_2(g)$$

(b) In terms of molar concentrations, write the equilibrium-constant expression for the decomposition of $AsF_5(g)$.

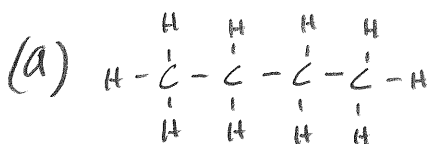(c) When equilibrium is established, 27.7 percent of the original number of moles of $AsF_5(g)$ has decomposed.

(i) Calculate the molar concentration of $AsF_5(g)$ at equilibrium.

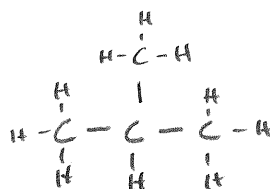(ii) Using molar concentrations, calculate the value of the equilibrium constant, Keq, at 105°C.

(d) Calculate the mole fraction of $F_2(g)$ in the container at equilibrium.

① 2010 FORM B

(a)

n-butane

2-methyl propane

➤ At this point in the course, you are not expected to be able to draw the structures of organic compounds... but soon, you will. ☺
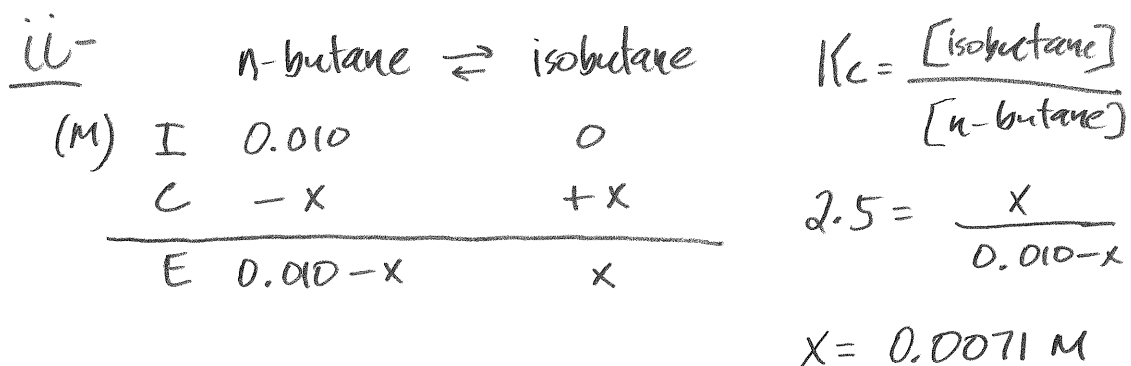
(b) n-butane will boil at a higher temperature as it has some IM forces between the molecules. Specifically, n-butane has more London dispersion forces than 2-methyl propane due to its linear structure. The linear molecules can approach one another more closely and therefore increase the # of instantaneous dipoles.

(c) $K_c = \dfrac{[\text{ISOBUTANE}]}{[\text{n-BUTANE}]}$

(d) $P = \dfrac{nRT}{V} = \dfrac{(0.010 \text{ mol})(0.0821 \frac{L \cdot atm}{mol \cdot K})(298 \text{ K})}{1.0 \text{ L}}$

$P = 0.24 \text{ atm}$

(e)   i- The total pressure remains the same, 0.24 atm. As the reaction proceeds, the number of molecules in the container is unchanged. One molecule of gas is produced for every one molecule consumed.

ii- $n\text{-butane} \rightleftharpoons \text{isobutane}$ $K_c = \dfrac{[\text{isobutane}]}{[n\text{-butane}]}$

(M)	n-butane	isobutane
I	0.010	0
C	$-x$	$+x$
E	$0.010 - x$	x

$$2.5 = \frac{x}{0.010 - x}$$

$$x = 0.0071\ M$$

$[\text{ISOBUTANE}] = x = \boxed{0.0071\ M}$ ← 2 sig. fig's

$[n\text{-BUTANE}] = 0.010 - x = 0.010 - 0.0071 = \boxed{0.003\ M}$ ← 1 sig. fig

iii- $[\text{BUTANE}] = 2 \times 0.003\ M = \boxed{0.006\ M}$

Since the volume is halved, the concentration will double! The pressure will be doubled but not shift the equilibrium position.

(f) The concentrations of both isomers would be the same as calculated in part e, ii above! Since temperature is the same, the position of equilibrium is not affected.

②
2008 FRQ

(a) $K_p = \dfrac{(P_{CO})^2}{P_{CO_2}}$

(b) $PV = nRT$; $n = \dfrac{PV}{RT} = \dfrac{(5.00\text{ atm})(2.00\text{ L})}{(0.0821\frac{L \cdot atm}{mol \cdot K})(1160\text{ K})} = \boxed{0.105\text{ moles}}$

(c) i— $P_{TOTAL} = P_{CO_2} + P_{CO}$

$P_{CO} = P_{TOTAL} - P_{CO_2} = 8.37\text{ atm} - 1.63\text{ atm} = \boxed{6.74\text{ atm}}$

ii— $K_p = \dfrac{(P_{CO})^2}{P_{CO_2}} = \dfrac{(6.74\text{ atm})^2}{1.63\text{ atm}} = \boxed{27.9}$

(d) The total pressures at equilibrium with & without the catalyst would be the SAME! Although a catalyst would cause the equilibrium to be attained quicker, it will not affect the extent of the reaction. K_p is not affected!

(e) $Q = \dfrac{(P_{CO})^2}{P_{CO_2}} = \dfrac{(2.00\text{ atm})^2}{2.00\text{ atm}} = 2.00$

$K = 27.9$ (from part c)
$Q = 2.00$

Since $K > Q$, the reaction will shift RIGHT to attain a new equilibrium position. Shifting right will DECREASE the P_{CO_2} !!
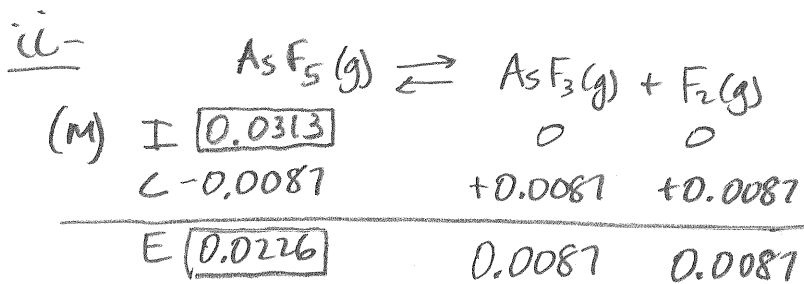
③

(a) i- $X \frac{mol}{L} = \left| \frac{55.8 \text{ g } AsF_5}{10.5 L} \right| \frac{1 \text{ mol } AsF_5}{169.9 \text{ g } AsF_5} \right| = \boxed{0.0313 \text{ M}}$

ii- $P = \frac{nRT}{V} = \frac{(55.8 g / 169.9 g/mol)(0.0821 \frac{L \cdot atm}{mol \cdot k})(378 \text{ K})}{10.5 L} = \boxed{0.969 \text{ atm}}$

(b) $K_c = \frac{[AsF_3][F_2]}{[AsF_5]}$

(c) i- $100.0\% - 27.7\% = 72.3\%$

$[AsF_5] = (0.723)(0.0313 \text{ m}) = \boxed{0.0226 \text{ M}}$

ii-

	$AsF_5 (g)$	⇌	$AsF_3 (g)$	+	$F_2 (g)$
(M) I	0.0313		0		0
C	-0.0087		+0.0087		+0.0087
E	0.0226		0.0087		0.0087

$K_{eq} = \frac{(0.0087)^2}{0.0226} = \boxed{0.0033}$

(d) $mol \ AsF_5 = M \times L = (0.0226 M)(10.5 L) = 0.237$ moles AsF_5

$mol \ F_2 = mol \ AsF_3 = M \times L = (0.0087 M)(10.5 L) = 0.091$ moles

$X_{F_2} = \frac{0.091}{0.091 + 0.091 + 0.237} = \boxed{0.217}$