SAMPLE PROBLEMS CHEMISTRY 12 EQUILIBRIUM

Key

1. Equilibrium is considered to be a "dynamic" process because

   a. it occurs in a closed system.
   b. equilibrium can be achieved from either direction.
   (c.) the forward and reverse reactions continue to occur.
   d. the concentrations of reactants and products are constant.

2. The equilibrium constant expression for the following reaction is

   pure liquid — solid

   $$2\,Hg_{(l)} + O_{2(g)} <====> 2\,HgO_{(s)}$$

   (a.) $K_{eq} = \frac{1}{[O_2]}$

   b. $K_{eq} = [O_2]$

   c. $K_{eq} = \frac{[2\,HgO]}{[O_2][2Hg]}$

   d. $K_{eq} = \frac{[HgO]^2}{[Hg]^2[O_2]}$

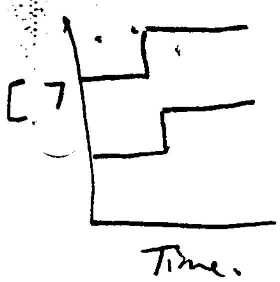
3. In which of the following systems would the tendencies toward minimum enthalpy and maximum entropy be in opposition to each other?

   ΔS →, ← ΔH

   (A) $Br_{2(l)}$ + heat ===> $Br_{2(g)}$
   B. $NaOH_{(s)}$ ===> $Na_{(aq)}^+ + OH_{(aq)}^-$ + heat
   C. $2C_{(g)} + 2H_{2(g)}$ ===> $C_2H_{4(g)}$ ΔH is positive
   D. $K_{(s)} + H_2O_{(l)}$ ===> $K_{(aq)}^+ + OH_{(aq)}^- + \frac{1}{2}H_{2(g)}$ ΔH is negative

4. Which of the following reactions will shift left when pressure is increased and when temperature is decreased?

   a. $N_{2(g)} + 2O_{2(g)}$ + heat <====> $2NO_{2(g)}$
   b. $N_{2(g)} + 3H_{2(g)}$ <====> $2NH_{3(g)}$ + heat
   (c.) $CH_{4(g)} + H_2O_{(g)}$ + heat <====> $CO_{(g)} + 3H_{2(g)}$ fewer gas reactants, endotherm.
   d. $CS_{2(g)} + 4H_{2(g)}$ <====> $CH_{4(g)} + 2H_2S_{(g)}$ + heat



5. Consider the following equilibrium system:

$$FeO_{(s)} + H_{2(g)} <===> Fe_{(s)} + H_2O_{(g)}$$

Which one of the following statements describes the effect that a decrease in volume would have on the position of equilibrium and the $[H_2]$ in the above system?

A. No shift, $[H_2]$ increases.
B. Shift right, $[H_2]$ increases.
C. Shift right, $[H_2]$ decreases.
D. No shift, $[H_2]$ remains constant. ~~both sides have 1 mol of~~ gas

6. Tooth enamel, $Ca_5(PO_4)_3OH$ establishes the following equilibrium:

$$Ca_5(PO_4)_3OH_{(s)} <===> 5Ca_{(aq)}^{2+} + 3PO_{4(aq)}^{3-} + OH_{(aq)}^{-} \downarrow$$

Which one of the following, when added to the above equilibrium system, would result in shift to the right?

A. $H_{(aq)}^{+}$
B. $OH_{(aq)}^{-}$
C. $Ca_{(aq)}^{2+}$
D. $Ca_5(PO_4)_3OH_{(s)}$

7. In which of the following reactions does the tendency towards minimum enthalpy and maximum entropy oppose each other?

a. $3O_{2(g)} \longrightarrow 2O_{3(g)}$	$\leftarrow \Delta S$	$\Delta H = +285$ kJ
b. $1/2N_{2(g)} + O_{2(g)} \longrightarrow NO_{2(g)}$	$\leftarrow \Delta S$	$\Delta H = +34$ kJ
c. $2H_2O_{(g)} \longrightarrow 2H_{2(g)} + O_{2(g)}$	$\Delta S \rightarrow$	$\Delta H = +484$ kJ
d. $P_{4(s)} + 6H_{2(g)} \longrightarrow 4PH_{3(g)}$	$\leftarrow \Delta S$	$\Delta H = +37$ kJ

($\Delta H \leftarrow$)

8. Consider the following equilibrium:

$$2O_{3(g)} <=====> 3O_{2(g)} \downarrow \quad K_{eq} = 55$$

If 0.060 mol of $O_{3(g)}$ and 0.70 mol of $O_{2(g)}$ are introduced into a 1.0 L vessel, the

a. $K_{trial} > K_{eq}$ and the $[O_{2(g)}]$ increases.
b. $K_{trial} < K_{eq}$ and the $[O_{2(g)}]$ increases.
c. $K_{trial} > K_{eq}$ and the $[O_{2(g)}]$ decreases.
d. $K_{trial} < K_{eq}$ and the $[O_{2(g)}]$ decreases.

$$K_{eq} = \frac{[O_2]^3}{[O_3]^2} \qquad K_{trial} = \frac{(0.70)^3}{(0.060)^2} = 95$$

$$K_{trial} > K_{eq}.$$

9. $CO_{2(g)} + H_{2(g)} <===> CO_{(g)} + H_2O_{(g)}$ $\frac{1.00\text{ mol}}{2.00\text{ L}} = 0.500\text{ M}$ $\frac{2.00\text{ mol}}{2.00\text{ L}} = 1.00\text{ M}$

1.00 mole of CO_2 and 2.00 moles of $H_{2(g)}$ are placed into a 2.00 litre container. A equilibrum, the [CO] = 0.31 mol/L. based on this data, the equilibrium [CO_2] is:

(A) 0.19M
B. 0.31M
C. 0.38M
D. 0.69M

	CO_2	H_2	CO	H_2O
I	.500	1.00 M	–	–
C	-.31	-.31 M	+.31 M	+.31 M
E	0.19	0.69	.31 M	.31 M

10. What is "equal" in a chemical reaction that has reached a state of equilibrium?

Forward and Reverse Reaction Rates

11. Consider the following equilibrium:

$$N_{2(g)} + 3H_{2(g)} <====> 2NH_{3(g)} \quad K_{eq} = 626 \text{ at } 200°C$$

At equilibrium, [N_2] is 1.06 mol/L and [H_2] is 0.456 mol/L.
Calculate [NH_3] in the equilibrium mixture.

$$K_{eq} = \frac{[NH_3]^2}{[N_2][H_2]^3} = 626.$$

$$\frac{[NH_3]^2}{(1.06)(0.456)^3} = 626$$

$$\frac{[NH_3]^2}{0.1005} = 626$$

$$[NH_3]^2 = 62.92$$

$$[NH_3] = 7.93\text{ M}$$

Name Key

Class ____________ Date ____________

1. 0.0040 mol of NO and 0.0030 mol of O_2 are introduced into a 1.0 L flask, and the reaction $2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$ occurs. At equilibrium, it is determined that $[NO_2] = 0.0035$ mol/L. What is the value of K_{eq} for the reaction?

$2NO + O_2 \leftrightarrow 2NO_2$

$K_{eq} = \frac{[NO_2]^2}{[NO]^2[O_2]}$

$K_{eq} = \frac{(0.0035)^2}{(0.0005)^2(0.00125)} = 3.9 \times 10^4$

$= 4 \times 10^4$

	NO	O_2	NO_2
[Initial]	0.0040	0.0030	0
change	−0.0035	−0.00175	+0.0035
[Eq m]	0.0005	0.00125	0.0035

2. 0.020 mol of each of SO_2, O_2, and SO_3 is placed in a 1.0 L flask and allowed to come to equilibrium. The equilibrium $[SO_2]$ is found to be 0.0080 mol/L. What is the value of K_{eq} for the reaction $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$?

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

$K_{eq} = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$

$K_{eq} = \frac{(0.032)^2}{(0.0080)^2(0.014)} = 1.1 \times 10^3$

	2 SO_2	1 O_2	2 SO_3
[Initial]	0.020	0.020	0.020
change	−0.012	+0.0060	+0.012
[Eq m]	0.0080	0.014	0.032

3. A solution initially contains 0.35*M* of A and 0.75*M* of B. A reaction occurs according to the equation $2A + B \rightleftharpoons 3C + D$. At equilibrium, [D] is found to be 0.10*M*. What is the value of K_{eq}?

$2A + B \leftrightarrow 3C + D$

$K_{eq} = \frac{[C]^3[D]}{[A]^2[B]}$

$K_{eq} = \frac{(0.30)^3(0.10)}{(0.15)^2(0.65)} = 0.18$

	A	B	C	D
[Initial]	0.35	0.75	0	0
change	−0.20	−0.10	+0.30	+0.10
[Eq m]	0.15	0.65	0.30	0.10

4. **For the equilibrium reaction $CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$, the K_{eq} value at 690°C is 10.0. A mixture of 0.300 mol of CO, 0.300 mol of H_2O, 0.500 mol of CO_2, and 0.500 mol of H_2 is placed in a 1.0 L flask.**
   a. **Show that the reaction is not at equilibrium.**
   b. **Determine the direction in which the reaction will shift to reach equilibrium.**
   c. **Calculate the equilibrium concentrations of all four species.**

Key.

a) trial product = $\frac{[CO_2][H_2]}{[CO][H_2O]} = \frac{(0.500)(0.500)}{(0.300)(0.300)} = 2.8 \neq 10.0$ ∴ not at eq=

b) trial product < Keq ∴ shifts to products

c) $K_{eq} = \frac{[CO_2][H_2]}{[CO][H_2O]} = 10.0$

	CO	H_2O	CO_2	H_2
[Initial]	0.300	0.300	0.500	0.500
Change	−x	−x	+x	+x
[Eq=]	0.300−x	0.300−x	0.500+x	0.500+x

$K_{eq} = \frac{(0.500+x)^2}{(0.300-x)^2} = 10.0$ √ both sides

x = 0.108 M

$[CO] = [H_2O] = 0.300 - 0.108 = 0.192$ M $[CO_2] = [H_2] = 0.500 + 0.108 = 0.608$ M

5. **The K_{eq} for the reaction $2HI(g) \rightleftharpoons H_2(g) + I_2(g)$ has a value of 1.85×10^{-2} at 425°C. If 0.18 mol of HI is placed in a 2.0 L flask and allowed to come to equilibrium at this temperature, what will be the equilibrium $[I_2]$?**

$K_{eq} = \frac{[H_2][I_2]}{[HI]^2} = 1.85 \times 10^{-2}$

	HI	H_2	I_2
[Initial]	0.18 mol / 2.0 L	0	0
change	−2x	+x	+x
[Eq=]	0.090−2x	x	x

$= \frac{x^2}{(0.090-2x)^2} = 1.85 \times 10^{-2}$

√ both sides

$x = 9.6 \times 10^{-3}$ M

$[I_2] = 9.6 \times 10^{-3}$ M

6. For the reaction $2A(g) \rightleftharpoons 2B(g) + C(s)$, the value of K_{eq} is known to be 6.8×10^2. If 0.42 mol of A is placed in a 3.0 L container and allowed to reach equilibrium, what is the equilibrium [B]?

$K_{eq} = [B]^2 / [A]^2 = 6.8 \times 10^2$

A	B
0.42/3.0	0
−x	+x
0.14−x	x

$K_{eq} = \frac{x^2}{(0.14-x)^2} = 6.8 \times 10^2$ √ both sides

x = .13 M

$[B]_{eq=}$ = 0.13 M

Key

Equilibrium Worksheet No. 1

1. Equilibrium is said to be dynamic. Explain why this is so and give an example of a dynamic equilibrium.

   Both the forward and reverse reactions continue to occur, but at the same rate. eg. Saturated solution, both ions and crystal continue to form.

2. Indicate in each of the following reactions whether the tendency towards maximum entropy favors reactants or products.

   a. $N_{2(g)} + 2O_{2(g)} + \text{heat} <====> 2NO_{2(g)}$ reactants
   b. $N_{2(g)} + 3H_{2(g)} <====> 2NH_{3(g)} + \text{heat}$ reactants
   c. $CH_{4(g)} + H_2O_{(g)} + \text{heat} <====> CO_{(g)} + 3H_{2(g)}$ products
   d. $CS_{2(g)} + 4H_{2(g)} <====> CH_{4(g)} + 2H_2S_{(g)} + \text{heat}$ reactants

3. In each of the following reactions show the direction the reaction must proceed in to attain minimum enthalpy and maximum entropy.

   a. $3O_{2(g)} \text{---> } 2O_{3(g)}$ $\Delta H = +285$ kJ $\xleftarrow{\Delta H}$ $\xleftarrow{\Delta S}$
   b. $1/2N_{2(g)} + O_{2(g)} \text{---> } NO_{2(g)}$ $\Delta H = +34$ kJ $\xleftarrow{\Delta H}$ $\xleftarrow{\Delta S}$
   c. $2\ H_2O_{(g)} \text{----> } 2H_{2(g)} + O_{2(g)}$ $\Delta H = +484$ kJ $\xleftarrow{\Delta H}$ $\xrightarrow{\Delta S}$
   d. $P_{4(s)} + 6H_{2(g)} \text{---> } 4PH_{3(g)}$ $\Delta H = +37$ kJ $\xleftarrow{\Delta H}$ $\xleftarrow{\Delta S}$

4. In which of the following reactions will the entropy favour the reactants while enthalpy favours the products?

   (a.) $Cl_{2(g)} <=====> Cl_{2(aq)} + 25$ kJ
   b. $P_{4(s)} + 6H_{2(g)} + 37\text{ kJ} <=====> 4PH_{3(g)}$
   c. $PCl_{5(g)} <=====> PCl_{3(g)} + Cl_{2(g)}$; $\Delta H = +92.5$ kJ
   d. $NO_{2(g)} <=====> \frac{1}{2}N_{2(g)} + O_{2(g)}$; $\Delta H = -33.8$ kJ

5. For each of the following reactions determine the direction of the enthalpy drive and the direction of the entropy drive. Then determine which one factor is responsible for the forward reaction.

   a. $\text{heat} + N_2O_{3(g)} <----> NO_{(g)} + NO_{2(g)}$ $\xleftarrow{\Delta H}$ $\xrightarrow{\Delta S}$ entropy
   b. $2H_{2(g)} + 2NO_{(g)} <---> N_{2(g)} + 2H_2O_{(g)} + \text{heat}$ $\xrightarrow{\Delta H}$ $\xleftarrow{\Delta S}$ enthalpy
   c. $CO_{(g)} + 1/2O_{2(g)} <--> CO_{2(g)}$ $\Delta H = -283$ kJ $\xrightarrow{\Delta H}$ $\xleftarrow{\Delta S}$ enthalpy
   d. $I_{2(g)} + Cl_{2(g)} <----> 2ICl_{(g)}$ $\Delta H = 35$ kJ $\xleftarrow{\Delta H}$ $\xrightarrow{\Delta S}$ entropy (if it is [illegible] it must increase [illegible])
   e. $H_2O_{(l)} <-------> H_2O_{(g)}$ $\Delta H = 41$ kJ $\xleftarrow{\Delta H}$ $\xrightarrow{\Delta S}$ entropy

Key.

Equilibrium #2

Describe the effect on the equilibrium when the following changes take place for each reaction.

1. $N_{2(g)} + 3H_{2(g)} \text{<----->} 2NH_{3(g)}$ ▲H = -92 kJ

   (a) increase the $[N_2]$ ⟶ shifts to products
   (b) increase the temperature ⟵ shifts to reactants
   (c) increase the volume ⟵ shifts to reactants
   (d) increase the pressure by changing volume ⟶ shifts to products

2. $2HF_{(g)} \text{<----->} F_{2(g)} + H_{2(g)}$ ▲H = 536 kJ

   (a) decrease the temperature ⟵ shifts to reactants
   (b) decrease the $[H_2]$ ⟶ shifts to products
   (c) decrease the amount of HF at constant volume ⟵ shifts to reactants
   (d) decrease the volume no change (2 mols of gas on each side)
   (e) increase the partial pressure of $H_{2(g)}$ ⟵ shifts to reactants

3. $SnO_{2(s)} + 2CO_{(g)} \text{<----->} Sn_{(s)} + 2CO_{2(g)}$ ▲H = 13 kJ

   (a) increase the temperature ⟶ shifts to products
   (b) increase the $[CO_2]$ ⟵ shifts to reactants
   (c) add a catalyst no change
   (d) add $Kr_{(g)}$ at constant volume no change
   (e) add $Kr_{(g)}$ at constant pressure no change (2 mols of gas on each side)
   (f) add SnO_2 solids do not change the position of equilibrium (more to come on this)

4. $CO_{(g)} + H_2O_{(g)} \text{<----->} CO_{2(g)} + H_{2(g)}$ ▲H = -41 kJ

   (a) add CO_2 ⟵ shifts to reactants
   (b) increase the temperature ⟵ shifts to reactants
   (c) remove some H_2O ⟵ shifts to reactants
   (d) decrease the pressure by changing the volume no change (2 mols of gas on each side)
   (e) add CO ⟶ shifts to products
   (f) add a catalyst no change
   (g) adding a catalyst. no change.

key

Equilibrium #3

Describe how each of the following changes listed below each equation will affect the **amount** of substance that is highlighted.

1. $N_{2(g)} + \mathbf{3H_{2(g)}} \text{<----->} 2NH_{3(g)}$ $\Delta H = -92$ kJ

   (a) increase the $[N_2]$ decreases $[H_2]$
   (b) increase the temperature increases $[H_2]$
   (c) increase the volume increases $[H_2]$
   (d) increase the pressure by changing volume decreases $[H_2]$

2. $2HF_{(g)} \text{<----->} \mathbf{F_{2(g)}} + H_{2(g)}$ $\Delta H = 536$ kJ

   (a) decrease the temperature decreases $[F_2]$
   (b) decrease the $[H_2]$ increases $[F_2]$
   (c) decrease the amount of HF at constant volume decreases $[F_2]$
   (d) decrease the volume no effect (2 mol of gas on each side)
   (e) increase the partial pressure of $H_{2(g)}$ decreases $[F_2]$

3. $SnO_{2(s)} + \mathbf{2CO_{(g)}} \text{<----->} \mathbf{Sn_{(s)}} + 2CO_{2(g)}$ $\Delta H = 13$ kJ

   (a) increase the temperature decreases $[CO]$
   (b) increase the $[CO_2]$ increases $[CO]$
   (c) add a catalyst no effect
   (d) add $Kr_{(g)}$ at constant volume no effect
   (e) add $Kr_{(g)}$ at constant pressure no effect (2 mol of gas on each side)
   (f) add SnO_2 - no effect - solids don't affect eq^m

4. $CO_{(g)} + H_2O_{(g)} \text{<----->} CO_{2(g)} + \mathbf{H_{2(g)}}$ $\Delta H = -41$ kJ

   (a) add CO_2 decrease $[H_2]$
   (b) increase the temperature decrease $[H_2]$
   (c) remove some H_2O decrease $[H_2]$
   (d) decrease the pressure by changing the volume no effect (2 mols of gas on each side)
   (e) add CO increase $[H_2]$
   (f) add a catalyst no effect
   (g) adding a catalyst. no effect

Equilibrium #3b

Describe the effect of the changes on each equilibrium.

Key:
→ denotes shift to products
← denotes shift to reactants
↑ denotes increase
↓ denotes decrease
NC denotes no change

1. $2ICl_{(g)}$ <----> $I_{2(g)}$ + $Cl_{2(g)}$ ▲H = +
   a. increase volume NC
   b. decrease temperature ←
   c. increase partial pressure of $I_{2(g)}$ ←
   d. effect of c. on $Cl_{2(g)}$ concentration ↓

2. $N_{2(g)}$ + $O_{2(g)}$ <----> $2NO_{(g)}$ ▲H = -
   a. decrease volume NC
   b. increase temperature ←
   c. increase partial pressure of $O_{2(g)}$ →
   d. effect of c. on $N_{2(g)}$ concentration ↓

3. $3O_{2(g)}$ <----> $2O_{3(g)}$ ▲H = -
   a. decrease volume →
   b. increase temperature ←
   c. increase partial pressure of $O_{2(g)}$ →
   d. effect of c. on $O_{3(g)}$ concentration ↑

4. $N_2O_{3(g)}$ <----> $NO_{(g)}$ + $NO_{2(g)}$ ▲H = +
   a. increase volume →
   b. decrease temperature ←
   c. increase partial pressure of $NO_{2(g)}$ ←
   d. effect of c. on $NO_{(g)}$ concentration ↓

5. $2H_{2(g)}$ + $2NO_{(g)}$ <----> $N_{2(g)}$ + $2H_2O_{(g)}$ ▲H = -
   a. decrease volume →
   b. increase temperature ←
   c. increase partial pressure of $N_{2(g)}$ ←
   d. effect of c. on $NO_{(g)}$ concentration ↑

6. $2Bi^{3+}_{(aq)}$ + $3H_2S_{(g)}$ <----> $Bi_2S_{3(s)}$ + $6H^+_{(aq)}$ ▲H = -
   a. increase volume ←
   b. increase $[H^+]$ (decrease pH) ←
   c. add more $Bi_2S_{3(s)}$ NC
   d. add NaOH →

7. $CaCO_{3(s)}$ <----> $CaO_{(s)}$ + $CO_{2(g)}$ ▲H = +
   a. decrease volume ←
   b. add $Ar_{(g)}$ at constant **volume** NC
   c. add $Ar_{(g)}$ at constant **pressure** →
   d. add a catalyst NC

8. $CaC_{2(s)}$ + $2H_2O_{(l)}$ <---> $C_2H_{2(g)}$ + $Ca(OH)_{2(s)}$ ▲H = +
a. add more water NC
b. increase pressure by changing volume ←
c. add more $CaC_{2(s)}$ NC
d. heat up the mixture →

9. $C_6H_{6(l)}$ + $Br_{2(l)}$ <---> $C_6H_5Br_{(l)}$ + $HBr_{(g)}$ ▲H = +
a. add more $C_6H_5Br_{(l)}$ ←
b. decrease pressure by changing volume →
c. increase the temperature →
d. add a catalyst NC

10. $Cu_{(s)}$ + $2Ag^+_{(aq)}$ <---> $Cu^{2+}_{(aq)}$ + $2Ag_{(s)}$
a. add more $Cu_{(s)}$ NC
b. add some HCl ←
c. add some $Cu(NO_3)_2$ ←
d. add more $Ag_{(s)}$ NC

11. $4NH_{3(g)}$ + $5O_{2(g)}$ <---> $6H_2O_{(g)}$ + $4NO_{(g)}$ ▲H = +
a. increase the partial pressure of $NO_{(g)}$ ←
b. decrease the temperature ←
c. add $Ne_{(g)}$ at constant pressure →
d. add $Ne_{(g)}$ at constant volume NC

12. $H_{2(g)}$ + $\frac{1}{2}O_{2(g)}$ <---> $H_2O_{(l)}$ ▲H = -
a. add $H_2O_{(l)}$ NC
b. decrease the partial pressure of $O_{2(g)}$ ←
c. increase the temperature ←
d. decrease the volume →

13. $CO_{(g)}$ + $\frac{1}{2}O_{2(g)}$ <---> $CO_{2(g)}$ ▲H = -
a. add $CO_{2(g)}$ ←
b. decrease the partial pressure of $O_{2(g)}$ ←
c. increase the temperature ←
d. decrease the volume →

14. $I_{2(g)}$ + $Cl_{2(g)}$ <---> $2ICl_{(g)}$ ▲H = -
a. add $Cl_{2(g)}$ →
b. decrease the partial pressure of $I_{2(g)}$ ←
c. increase the temperature ←
d. decrease the volume NC

Equilibrium #4

Key.

Write the equilibrium constant expression for each equilibrium.

1. $2ICl_{(g)} \;<\!\!-\!\!-\!\!-\!\!> \; I_{2(g)} + Cl_{2(g)}$

$$K_{eq} = \frac{[I_2][Cl_2]}{[ICl]^2}$$

2. $N_{2(g)} + O_{2(g)} \;<\!\!-\!\!-\!\!-\!\!> \; 2NO_{(g)}$

$$K_{eq} = \frac{[NO]^2}{[N_2][O_2]}$$

3. $3O_{2(g)} \;<\!\!-\!\!-\!\!-\!\!> \; 2O_{3(g)}$

$$K_{eq} = \frac{[O_3]^2}{[O_2]^3}$$

4. $N_2O_{3(g)} \;<\!\!-\!\!-\!\!-\!\!> \; NO_{(g)} + NO_{2(g)}$

$$K_{eq} = \frac{[NO][NO_2]}{[N_2O_3]}$$

5. $2H_{2(g)} + 2NO_{(g)} \;<\!\!-\!\!-\!\!-\!\!> \; N_{2(g)} + 2H_2O_{(g)}$

$$K_{eq} = \frac{[N_2][H_2O]^2}{[H_2]^2[NO]^2}$$

6. $2Bi^{3+}{}_{(aq)} + 3H_2S_{(g)} \;<\!\!-\!\!-\!\!-\!\!> \; Bi_2S_{3(s)} + 6H^+{}_{(aq)}$

$$K_{eq} = \frac{[H^+]^6}{[Bi^{3+}]^2[H_2S]^3}$$

7. $CaCO_{3(s)} \;<\!\!-\!\!-\!\!-\!\!> \; CaO_{(s)} + CO_{2(g)}$

$$K_{eq} = [CO_2]$$

8. $CaC_{2(s)} + 2H_2O_{(l)} \;<\!\!-\!\!-\!\!-\!\!> \; C_2H_{2(g)} + Ca(OH)_{2(s)}$

$$K_{eq} = [C_2H_2]$$

9. $C_6H_{6(l)} + Br_{2(l)} \;<\!\!-\!\!-\!\!-\!\!> \; C_6H_5Br_{(l)} + HBr_{(g)}$

$$K_{eq} = \frac{[C_6H_5Br][HBr]}{[C_6H_6][Br_2]}$$

10. $Cu_{(s)} + 2Ag^+{}_{(aq)} \;<\!\!-\!\!-\!\!-\!\!> \; Cu^{2+}{}_{(aq)} + 2Ag_{(s)}$

$$K_{eq} = \frac{[Cu^{2+}]}{[Ag^+]^2}$$

Key.

11. $4NH_{3(g)} + 5O_{2(g)} <\!\!-\!\!-\!\!-\!\!> 6H_2O_{(g)} + 4NO_{(g)}$

$$K_{eq} = \frac{[H_2O]^6[NO]^4}{[NH_3]^4[O_2]^5}$$

12. $H_{2(g)} + ½O_{2(g)} <\!\!-\!\!-\!\!-\!\!> H_2O_{(l)}$

$$K_{eq} = \frac{1}{[H_2][O_2]^{1/2}}$$

13. $CO_{(g)} + ½O_{2(g)} <\!\!-\!\!-\!\!-\!\!> CO_{2(g)}$

$$K_{eq} = \frac{[CO_2]}{[CO][O_2]^{1/2}}$$

14. $I_{2(s)} + Cl_{2(g)} <\!\!-\!\!-\!\!-\!\!> 2ICl_{(g)}$

$$K_{eq} = \frac{[ICl]^2}{[Cl_2]}$$

15. $SnO_{2(s)} + 2CO_{(g)} <\!\!-\!\!-\!\!-\!\!> Sn_{(s)} + 2CO_{2(g)}$

$$K_{eq} = \frac{[CO_2]^2}{[CO]^2}$$

16. $Fe^{3+}{}_{(aq)} + SCN^-{}_{(aq)} <\!\!-\!\!-\!\!-\!\!> FeSCN^{2+}{}_{(aq)}$

$$K_{eq} = \frac{[FeSCN^{2+}]}{[Fe^{3+}][SCN^-]}$$

17. $2CrO_4{}^{2-}{}_{(aq)} + 2H^+{}_{(aq)} <\!\!-\!\!-\!\!-\!\!> Cr_2O_7{}^{2-}{}_{(aq)} + H_2O_{(l)}$

$$K_{eq} = \frac{[Cr_2O_7{}^{2-}]}{[CrO_4{}^{2-}]^2[H^+]^2}$$

18. $Fe^{3+}{}_{(aq)} + 4Cl^-{}_{(aq)} <\!\!-\!\!-\!\!-\!\!> FeCl_4{}^-{}_{(aq)}$

$$K_{eq} = \frac{[FeCl_4{}^-]}{[Fe^{3+}][Cl^-]^4}$$

19. $H_2O_{(l)} <\!\!-\!\!-\!\!-\!\!> H^+{}_{(aq)} + OH^-{}_{(aq)}$

$$K_{eq} = [H^+][OH^-]$$

20. $BaSO_{4(s)} <\!\!-\!\!-\!\!-\!\!> Ba^{2+}{}_{(aq)} + SO_4{}^{2-}{}_{(aq)}$

$$K_{eq} = [Ba^{2+}][SO_4{}^{2-}]$$

Equilibrium Worksheet 5 Name Key

Solve each of the following problems. Show your work in the space provided. Write your final answer on the blank line.

Part A

1. Write an equilibrium expression for the following reaction:

$A_{(g)} + B_{(g)} <=====> AB_{(g)}$

Then calculate the value of K_{eq} given that $[A] = 1.1 \times 10^{-3}M$, $[B] = 4.4M$, and $[AB] = 1.5 \times 10^{-8}M$. Finally, tell whether reactants or products are favored, and why.

$$K_{eq} = \frac{[AB]}{[A][B]} = \frac{(1.5\times10^{-8})}{(1.1\times10^{-3})(4.4)} = 3.1\times10^{-6}$$

Reactants are favored because Keq is {small or less than 1

2. Write an equilibrium expression for the following reaction:

$A_{2(g)} + B_{(s)} <=====> A_2B_{(g)}$

Then calculate the value of K_{eq} given that $[A_2] = 1.9 \times 10^{-3}M$, and $[A_2B] = 1.4 \times 10^{-5}M$. Finally, tell whether reactants or products are favored, and why.

$$K_{eq} = \frac{[A_2B]}{[A_2]} = \frac{1.4\times10^{-5}}{1.9\times10^{-3}} = 7.4\times10^{-3}$$

Reactants are favored because keq is {small or less than 1

3. Write an equilibrium expression for the following reaction:

$2A_{(g)} + B_{(g)} <=====> A_2B_{(g)}$

Then calculate the value of K_{eq} given that $[A] = 1.0 \times 10^{-6}M$, $[B] = 2.2 \times 10^{-4}M$, and $[A_2B] = 6.5 \times 10^{-1}M$. Finally, tell whether reactants or products are favored, and why.

$$K_{eq} = \frac{[A_2B]}{[A]^2[B]} = \frac{6.5\times10^{-1}}{(1.0\times10^{-6})^2(2.2\times10^{-4})} = 3.0\times10^{15}$$

Products are favored because keq is greater than 1

Key.

4. **Write an equilibrium expression for the following reaction:**

$A_{(s)} + B_{2(g)} <====> AB_{(s)} + B_{(g)}$

Then calculate the value of K_{eq} given that $[B_2] = 5.5 \times 10^{-4}$M, and $[B] = 3.9 \times 10^{-7}$M. Finally, tell whether reactants or products are favored, and why.

$K_{eq} = \frac{[B]}{[B_2]} = \frac{3.9 \times 10^{-7}}{5.5 \times 10^{-4}} = 7.1 \times 10^{-4}$

Reactants are favored because Keq is less than 1.

5. Write an equilibrium expression for the following reaction:

$2A_{(g)} + 3B_{(g)} <=====> A_2B_{3(s)}$

Then calculate the value of K_{eq} given that $[A] = 4.6 \times 10^{-3}$M, and $[B] = 1.5 \times 10^{-5}$M. Finally, tell whether reactants or products are favored, and why.

$K_{eq} = \frac{1}{[A]^2[B]^3} = \frac{1}{(4.6\times10^{-3})^2(1.5\times10^{-5})^3} = 1.4 \times 10^{19}$

Products are favored because keq is greater than 1.

Part B

6. Write an equilibrium expression for the following reaction:

$AB_{2(s)}$ + energy $<=====> A_{(s)} + B_{2(g)}$

Then calculate the value of K_{eq} given that $[B_2] = 1.3 \times 10^{-9}$M. Finally, predict the effect of increased temperature on the value of K_{eq} and explain your answer.

$K_{eq} = [B_2] = 1.3 \times 10^{-9}$

Increasing temperature will shift the eqm to the products, increasing $[B_2]$ and increasing Keq because this reaction is endothermic

key

7. Write an equilibrium expression for the following reaction:

$2A_{(g)} + B_{(g)} + \text{energy} <====> A_2B_{(g)}$

Then calculate the value of K_{eq} given that [A] = 1.6 x 10^{-2}M, [B] = 1.4 x 10^{-4}M, and $[A_2B]$ = 3.6 x 10^{-1} M. Finally, predict the effect of decreased temperature on the value of K_{eq} and explain your answer.

$$K_{eq} = \frac{[A_2B]}{[A]^2[B]} = \frac{3.6\times10^{-1}M}{(1.6\times10^{-2}M)^2(1.4\times10^{-4}M)} = 1.0\times10^{7}$$

This is an endothermic reaction, decreasing the temperature shifts the equilibrium towards the reactants and decreases Keq.

Part C

8. Write an equilibrium expression for the following reaction:

$A_{2(g)} + B_{(s)} <====> A_{(g)} + AB_{(s)}$

Then calculate the concentration of $A_{(g)}$ given that K_{eq} = 1.5 x 10^{-3}M, and $[A_2]$ = 2.5 x 10^{-4} M. Finally, predict the effect of adding some $A_{2(g)}$ on the values for [A], and explain your answer.

$$K_{eq} = \frac{[A]}{[A_2]} = 1.5\times10^{-3} = \frac{[A]}{2.5\times10^{-4}M} \qquad [A] = 3.8\times10^{-7}M$$

Adding some $A_{2(g)}$ will increase the value for [A] but will not change the value of Keq.

9. Write an equilibrium expression for the following reaction:

$2A_{(g)} + B_{2(g)} <====> A_2B_{(g)} + B_{(g)}$

Then calculate the concentration of A_2B, given that K_{eq} = 7.1 x 10^{+} M, [A] = 1.9 x 10^{-2}M, $[B_2]$ = 4.1 x 10^{-3}M, and [B] = 8.4 x 10^{-3} M. Finally, predict the effect of adding some $A_{(g)}$ on the values for $[B_2]$, $[A_2B]$, and [B], and explain your answer.

$$K_{eq} = \frac{[A_2B][B]}{[A]^2[B_2]} = 7.1\times10^{+} = \frac{[A_2B](8.4\times10^{-3}M)}{(1.9\times10^{-2}M)^2(4.1\times10^{-3}M)}$$

$[A_2B] = 1.2\times10^{1} \sim 1.3\times10^{1}$

Adding $A_{(g)}$ will shift the equilibrium to the products; $[B_2]$ will decrease, $[A_2B]$ and [B] will increase. Keq will not change.

Key

Equilibrium #6

A. Reactant/Product Concentration Graph

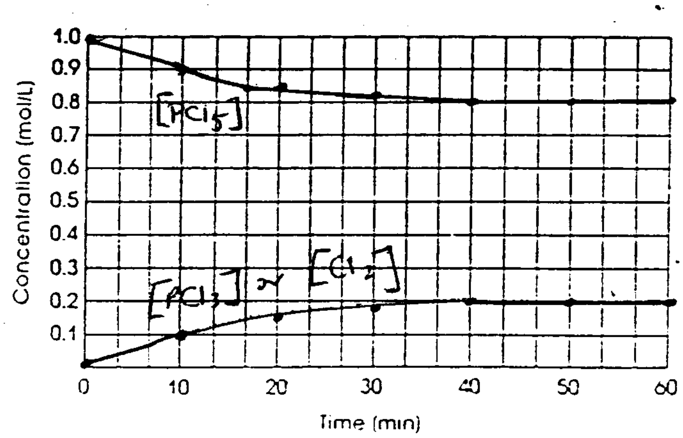
The decomposition of phosphorus pentachloride, PCl_5, into phosphorus trichloride and chlorine gas is a reversible reaction that reaches a state of chemical equilibrium. The equilibrium reaction is

$$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$$

One mole of PCl_5 is placed in a sealed one-liter container and heated to 250 °C. The concentrations of reactant and products are measured at ten-minute intervals. The following data are collected.

TIME (MINUTES)	CONCENTRATIONS (MOLES PER LITER)		
	$[PCl_5]$	$[PCl_3]$	$[Cl_2]$
0	1.00	0.00	0.00
10	0.90	0.10	0.10
20	0.85	0.15	0.15
30	0.82	0.18	0.18
40	0.80	0.20	0.20
50	0.80	0.20	0.20
60	0.80	0.20	0.20

Plot this data on the grid provided. Answer the questions that follow.



1. How long did it take for the reaction to reach equilibrium?

   40 minutes

2. What is the equilibrium concentration of each reaction component?

   $[PCl_5]$ = 0.80 M $[PCl_3]$ = $[Cl_2]$ = 0.20 M

As each PCl_5 particle reacts it forms 1 PCl_3 and 1 Cl_2 particle

4. **Predict the concentration of each component at 70 minutes.**

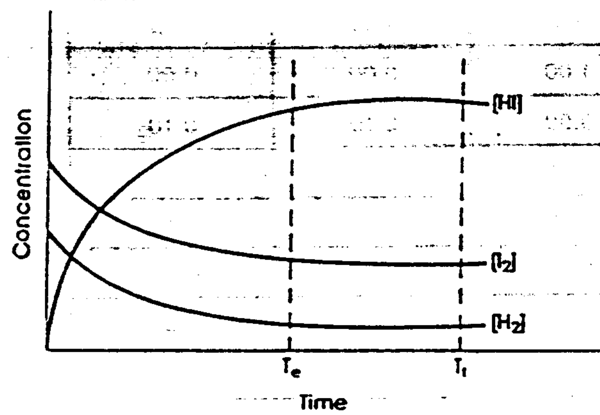
If conditions are unchanged $[PCl_5] = 0.80M$ $[PCl_3]$ & $[Cl_2] = 0.20M$

B. Predicting Reactant/Product Concentrations

The synthesis of hydrogen iodide is a reversible exothermic reaction that proceeds as follows.

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$

In a laboratory experiment, hydrogen gas and iodine gas are placed in a sealed reaction flask. The gases react to produce hydrogen iodide until equilibrium is established. The concentrations of reactants and product are plotted in the graph that follows. Assume equilibrium is reached at point T_e.



The following changes are introduced at time T_f. Determine which situation best describes how the graph would be changed to the right of T_f. Write your answer in the space provided.

1. Concentration of I_2 is increased the HI curve rises, the I_2 curve rises, the H_2 curve drops
   a. The HI curve rises, the I_2 curve drops, the H_2 curve remains the same.
   b. The HI curve drops, the I_2 curve rises, the H_2 curve drops.
   c. The HI curve rises, the I_2 curve rises, the H_2 curve drops.

2. Temperature of the system is increased the H_2 and I_2 curves rise, the HI curve drops
   a. The curves for all three components rise.
   b. The H_2 and I_2 curves rise, the HI curve drops.
   c. The H_2 and I_2 curves drop, the HI curve rises.

3. Pressure of the system is increased All curves rise (depends on method used to increase pressure)
   a. All curves rise
   b. The H_2 and I_2 curves drop, the HI curve rises.
   c. All curves remain the same.

4. If a catalyst is present at the initial introduction of reactants, how would the graph differ from the one shown?

   the time taken to reach equilibrium would be less

C. Applying Le Chatelier's Principle

The oxidation of ammonia is a reversible exothermic reaction that proceeds as follows:

$$4NH_3(g) + 5O_2(g) \rightleftarrows 4NO(g) + 6H_2O(g)$$

Le Chatelier's principle allows us to predict the changes that occur in an equilibrium reaction to compensate for any stress that is placed upon the system. For each situation described in the table, use the symbol ↑ to show an increase in concentration or the symbol ↓ to show that a decrease in concentration is expected.

COMPONENT	STRESS	EQUILIBRIUM CONCENTRATIONS			
		$[NH_3]$	$[O_2]$	$[NO]$	$[H_2O]$
NH_3	addition	1 ↑	↓	↑	↑
	removal	2 ↓	↑	↓	↓
O_2	addition	3 ↓	↑	↑	↑
	removal	4 ↑	↓	↓	↓
NO	addition	5 ↑	↑	↑	↓
	removal	6 ↓	↓	↓	↑
H_2O	addition	7 ↑	↑	↓	↑
	removal	8 ↓	↓	↑	↓
Increase in temperature		9 ↑	↑	↓	↓
Decrease in temperature		10 ↓	↓	↑	↑
Increase in pressure		11 ↑	↑	↓	↓
Decrease in pressure		12 ↓	↓	↑	↑
Addition of a catalyst		13 —	—	—	—

Equilibrium #7

1. $CO_{(g)} + H_2O_{(g)} \longleftrightarrow CO_{2(g)} + H_{2(g)}$

$$K_{eq} = \frac{[CO_2][H_2]}{[CO][H_2O]} = 4.0$$

	CO	H_2O	CO_2	H_2
[Initial]	0.10	0.10	0.20	0.20
change	+0.10	+0.10	+x −0.10	−0.10
[Eq≅]	0.20	0.20	0.10 + x	0.10

$$K_{eq} = \frac{(0.10 + x)(0.10)}{(0.20)(0.20)} = 4.0$$

$x = 1.5\ M$ # of moles = $1.5\ M \times 2.0\ L$ = 3.0 mols CO_2 added.

2. $2NO_{(g)} + O_{2(g)} \leftrightarrow 2NO_{2(g)}$

$$K_{eq} = \frac{[NO_2]^2}{[NO]^2\,[O_2]}$$

	NO	O_2	NO_2
[Initial]	0.30/5.0	0	0.80/5.0
change	+0.040	+0.020	−0.040
[Eq≅]	0.10	0.020	0.12

$$K_{eq} = \frac{(0.12)^2}{(0.10)^2(0.020)} = 72.$$

3. $2NO_{(g)} + Cl_{2(g)} \longleftrightarrow 2NOCl_{(g)}$

$$K_{eq} = \frac{[NOCl]^2}{[NO]^2\,[Cl_2]}$$

	NO	Cl_2	NOCl
[Initial]	2.00/5.0	2.00/5.0	0
change	−.12	−.06	+.12
[Eq≅]	.28	.34	.12

$$K_{eq} = \frac{(.12)^2}{(.28)^2(.34)} = 0.54$$

4. $A_{2(g)} + B_{2(g)} \longleftrightarrow 2AB_{(g)}$

$$K_{eq} = \frac{[AB]^2}{[A_2][B_2]}$$

	A_2	B_2	AB
[Initial]	0.200	0.200	0.400
change	+ x	+ x	+0.100 − 2x
[Eq≅]	0.200 + x	0.200 + x	0.500 − 2x

$$K_{eq} = \frac{(0.400)^2}{(0.200)(0.200)} = 4.00$$

$$K_{eq} = \frac{(0.500 - 2x)^2}{(0.200 + x)^2} = 4.00$$ take square root of both sides.

$$\frac{0.500 - 2x}{0.200 + x} = 2.00 \qquad x = 0.0250M$$

new $[AB] = 0.500 - 2(0.0250) =$ 0.450M

Key.

Equilibrium # 7

5. $C_2 + D_2 \longleftrightarrow 2CD$

$$K_{eq} = \frac{[CD]^2}{[C_2][D_2]} = 9.0 \times 10^{-2}$$

	C_2	D_2	CD
[Initial]	0.24	0.24	0
change	− x	− x	+ 2x
[Eq m]	0.24 − x	0.24 − x	2x

$$K_{eq} = \frac{(2x)^2}{(0.24 - x)^2} = 9.0 \times 10^{-2}$$

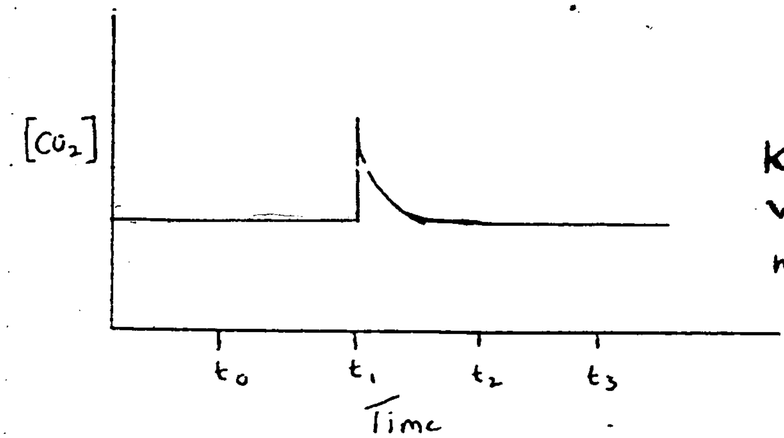
take the square root of both sides

$$\frac{2x}{0.24 - x} = 0.30 \qquad x = 0.0313$$

$[C_2]$ at eq m = 0.24 − 0.0313 = **0.21 M**

Equilibrium #8 #1–5

1.



$K_{eq} = [CO_2]$

V↓ P↑ . [change]

no effect on Keq, so $[CO_2]$ same

2. $CO_{(g)} + H_2O_{(g)} \longleftrightarrow CO_{2(g)} + H_{2(g)}$

	CO	H_2O	CO_2	H_2
I	0.30	0.30	0.90	0.90
C	−x	−x	+x	+x
E	0.30−x	0.30−x	0.90+x	0.90+x

$$K_{eq} = \frac{[CO_2][H_2]}{[CO][H_2O]} = 16$$

$$\frac{(0.90+x)^2}{(0.30-x)^2} = 16 \quad \sqrt{\ } \text{ both sides}$$

$x = 0.060$ $\quad [CO_2]_{Eq} = 0.90 + 0.060 = \underline{\underline{0.96\ M}}$

3. When the temperature is increased Keq will decrease. Increasing the temperature favors the endothermic, or reverse, reaction which decreases the [products] and increases [reactants] so keq is lower.

4. a) $[CH_3OH]$ increases when volume is decreased at constant temperature the eqm shifts to favor the side with fewer gas particles.
b) [CO] is unchanged when a catalyst is added, as adding a catalyst affects the forward and reverse rates equally and does not change the position of the equilibrium.

5. $Fe^{3+}_{(aq)} + SCN^-_{(aq)} \longleftrightarrow FeSCN^{2+}_{(aq)}$

	Fe^{3+}	SCN^-	$FeSCN^{2+}$
I	2.6×10^{-3}	1.00	0
C	-2.8×10^{-4}	-2.8×10^{-4}	$+2.8\times10^{-4}$
E	2.32×10^{-3}	1.00	2.8×10^{-4}

$$K_{eq} = \frac{[FeSCN^{2+}]}{[Fe^{3+}][SCN^-]}$$

$$K_{eq} = \frac{2.8\times10^{-4}}{(2.32\times10^{-3})(1.00)} = \underline{0.121}$$

$= \underline{\underline{0.12}}$

Equilibrium #8 #6-10 key.

6. $2NO_{(g)} + Br_{2(g)} \longleftrightarrow 2NOBr_{(g)}$

$$K_{eq} = \frac{[NOBr]^2}{[NO]^2[Br_2]} = 1.0 \times 10^2$$

	NO	Br_2	NOBr
I	0	0	X
C	+0.0800	+0.0400	−0.0800
E	0.0800	0.0400	X − 0.080

$$K_{eq} = \frac{(X - 0.0800)^2}{(0.0800)^2(0.0400)} = 1.0 \times 10^2$$

X = 0.24 M # mols NOBr = 0.24 M × 2.00 L = **0.48 moles**

7. If oxygen is added there is a net increase in O_2, NO and H_2O. there is a net decrease in N_2H_4.

8. $H_{2(g)} + I_{2(g)} \longleftrightarrow 2HI_{(g)}$

$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]}$$

	H_2	I_2	HI
I	0.200	0.200	0
C	−0.160	−0.160	+0.320
E	0.040	0.040	0.320

$$K_{eq} = \frac{(0.320)^2}{(0.040)^2} = \underline{64}$$

9. $2CH_{4(g)} \longleftrightarrow C_2H_{2(g)} + 3H_{2(g)}$

$$K_{eq} = \frac{[C_2H_2][H_2]^3}{[CH_4]^2}$$

	CH_4	C_2H_2	H_2
I	0.180	0	0
C	−0.160	+0.0800	+0.240
E	0.020	0.0800	0.240

$$K_{eq} = \frac{(0.0800)(0.240)^3}{(0.020)^2} = \underline{2.76}$$

10. $2NO_{(g)} + O_{2(g)} \longleftrightarrow 2NO_{2(g)}$

(a) $$K_{eq} = \frac{[NO_2]^2}{[NO]^2[O_2]}$$

(b) K_{eq} is greater than 1 ∴ when $[O_2] = 1.0\,M$ the value of $[NO_2]^2$ is greater than the value of $[NO]^2$ ∴ $[NO_2] > [NO]$

Worksheet No. 8

11.

	$2HI_{(g)} \longleftrightarrow$	$H_{2(g)}$ +	$I_{2(g)}$
R	$2HI_{(g)}$	$H_{2(g)}$	$I_{2(g)}$
I	4.00 M	0	0
C	$-2x$	$+x$	$+x$
E	$4.00-2x$	x	x

$K_{eq} = 81.0$

$$K_{eq} = \frac{[H_2][I_2]}{[HI]^2}$$

$$81.0 = \frac{x^2}{(4.00-2x)^2} \qquad x = 1.89$$

$[HI] = 4.00 - 2(1.89) = 0.21\ M$

12. When HCl is added the H^+ ion from the HCl reacts with the OH^- in the equilibrium to produce water. The $[OH^-]$ decreases which shifts the equilibrium to the products to make more OH^- and more $Cr_2O_7^{2-}$ which is orange.

13. $2NO_{(g)} + O_{2(g)} \longleftrightarrow 2NO_{2(g)}$

$$[NO] = \frac{0.044\ mol}{2.00\ L} = 0.022\ M$$

$$[O_2] = \frac{0.100\ mol}{2.00\ L} = 0.050\ M$$

$$[NO_2] = \frac{7.88\ mol}{2.00\ L} = 3.94\ M$$

$$K_{eq} = \frac{[NO_2]^2}{[NO]^2[O_2]}$$

$$K_{eq} = \frac{(3.94)^2}{(0.022)^2(0.050)} = 6.4 \times 10^5$$