SL worksheet number 3

1) State the Le Chatelier's Principle (hasn't been covered yet just google it)

when a system at ⇌ is stressed, the system will shift to reduce the stress until a new ⇌ is established

2) Why is a chemical equilibrium considered dynamic?

because although macroscopidly the reaction seems to have stopped microscopically it has NOT, collisions still occur and system is still in motion, hence dynamic

3) How can the value of the equilibrium constant expression be changed

Δ in Temp of system

4) At 200 *C the reaction N_2 (g) + $3Cl_2$ (g) ⇆ 2 NCl_3 (g) with a K = 15 (with substance amounts expressed in atm) if a reaction flask initially contains 1 atm of product and 5 atm of each reactant which direction must the reaction proceed to achieve equilibrium?

$$Q = \frac{(P_{NCl_3})^2}{(P_{N_2})(P_{Cl_2})^3} = \frac{(1)^2}{(5)(5)^3} = 0.0016 = Q$$

$Q < K$ OR $K > Q$ ∴ shift to products $\longrightarrow$

5) What is the reaction quotient Q for the reaction below

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

$$Q = \frac{\text{Products}^{\text{coeff.}}}{\text{Reagents}^{\text{coeff.}}} = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

6) Calculate the value of K_c for the reaction $2\ NO\ (g) + Br_2\ (g) \leftrightarrows 2\ NOBr$ (g)

Given the following information

$2NO\ (g) \leftrightarrows \cancel{N_2} + \cancel{O_2}(g)$ — $Kc1 = 1 \times 10^{30}$

$\cancel{N_2}(g) + Br_2\ (g) + \cancel{O_2}(g) \leftrightarrows 2\ NOBr\ (g)$ — $Kc2 = 2 \times 10^{-27}$

$2NO_{(g)} + Br_{2(g)} \rightleftarrows 2NOBr_{(g)}$

nothing is Δ (changed) $\therefore$

$K_c = Kc_1 \cdot Kc_2 = (1\times10^{30})(2\times10^{-27}) = 2\times10^{3}$

7) Calculate the value of Kc for the reaction $2\ N_2O(g) + 3\ O_2(g) \leftrightarrows 2\ N_2O_4(g)$

Using the following information

	Equation	Equilibrium Constant
1	$2\ N_2\ (g) + O_2(g) \leftrightarrows 2\ N_2O(g)$	$K_c = 1.2 \times 10^{-35}$
2	$N_2O_4(g) \leftrightarrows 2\ NO_2(g)$	$K_c = 4.6 \times 10^{-3}$
3	$\frac{1}{2} N_2(g) + O_2(g) \leftrightarrows NO_2(g)$	$K_c = 4.1 \times 10^{-9}$
overall	$2N_2O + 3O_2 \rightleftarrows 2N_2O_4$	

1 flip eq. 1

$2N_2O_{(g)} \rightleftarrows O_{2(g)} + 2N_{2(g)} = 1/K_c$

2 flip eq 2.

$2NO_2 \rightleftarrows N_2O_4 = 1/K_c$

3 times eq 3 by 4

$2N_2 + 4O_2 \rightleftarrows 4NO_2 = K_c^4$

4 times eq 2 by 2

$4NO_2 \rightleftarrows 2N_2O_4 = \left(\frac{1}{K_c}\right)^2$

$\therefore$

$2N_2O \rightleftarrows \cancel{O_2} + \cancel{2N_2} = 1/K_c$

$\cancel{4NO_2} \rightleftarrows 2N_2O_4 = (1/K_c)^2$

$\cancel{2N_2} + \cancel{4}^{3}O_2 \rightleftarrows \cancel{4NO_2} = K_c^4$

$3O_2 + 2N_2O \rightleftarrows 2N_2O_4$

times all new K's together

$(8.3\times10^{34})(4.7\times10^{4})(2.8\times10^{-34}) =$

ANSWER $= 1.1\times10^{6}$ work

8) Carbon monoxide reacts with hydrogen to form methanol. All species in this system are gases Keq = 25 and this reaction releases tones of energy.

a. If initially there are 2.00 moles of each species in a 2.500 L container in which direction must the system shift to reach equilibrium

① Balance eq. $CO + 2H_2 \rightleftharpoons CH_3OH$

② Need M not mols

∴

do this for all of them

$[CO] = \frac{2.0 mol}{2.50} L = 0.80 M$

$[H_2] = 0.80 M$

$[CH_3OH] = 0.80 M$

$Q = \frac{[CH_3OH]}{[CO][H_2]^2}$

$= \frac{(0.80)}{(0.80)(0.80)^2}$

$= 1.25$ $K > Q$

→ Products

b. Describe what will happen in terms of rates and in terms of equilibrium shifts if a catalyst is added to a system already at equilibrium

cat. will increase both forward + reverse rates but EQUALLY. [ ] of reactents and products will remain unchanged

9) For the following reaction

$N_{2(g)} + 2O_{2(g)} \leftrightarrows 2NO_{2(g)}$

a. At equilibrium which side is favoured

$Keq = 4.8 \times 10^{-3} << 1$

∴ Reactents are favoured at equilibrium

$\frac{0.400 mol}{2 L} = 0.2M$ $\quad$ 0.1M

10) 0.400 mol of hydrogen gas and 0.200 mol of iodine gas were placed in a 2.00 L flask and allowed to reach equilibrium. At equilibrium {HI] = 0.1600 M. Find the value of Keq.

	$H_{2(g)}$ +	$I_{2(g)}$	$\rightleftharpoons$ $2HI_{(g)}$
I	0.200 M	0.100 M	0
C	$-x$	$-x$	$+2x$
E	$(0.2-x)$	$(0.1-x)$	(0.16M)

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

$$= \frac{(0.16)^2}{[0.12][0.02]}$$

$$\boxed{K_{eq} = 10.7}$$

we know $0 + 2x = 0.16M$

$\therefore$

$x = \frac{0.16}{2} M$

$x = 0.08M$

$\therefore$

$[H_2] = 0.2 - 0.08 = 0.12M$

$[I_2] = 0.1 - 0.08 = 0.02M$

11) Equal moles of hydrogen and bromine gases are placed in a flask at equilibrium [HBr] = 0.500 mol/L, if Keq = 1.95, find [H_2] at equilibrium

	$H_{2\,(g)}$	+	$Br_{2\,(g)}$	$\rightleftharpoons$	$2\,HBr_{(g)}$
I	x		x		0
C	$-y$		$-y$		$+2y$
E	$x-y$		$x-y$		0.500 M

Step ① $2y = 0.500M$

$y = 0.2500M$

$\therefore$ $[H_2] = x - 0.25$

$[Br_2] = x - 0.25$

$$Keq = \frac{[HBr]^2}{[H_2][Br_2]}$$

$$\sqrt{1.95} = \sqrt{\frac{(0.500)^2}{(x-0.25)^2}}$$

$$1.40 = \frac{0.500}{x-0.25}$$

$(x-0.25)(1.40) = 0.500$

$1.40x - 0.349 = 0.50$

$$\frac{1.40x}{1.40} = \frac{0.50 + 0.349}{1.40}$$

$x = 0.607M$

$[H_2]$ at equilibrium $= x - y = 0.607 - 0.25$

$= \underline{0.357M}$