

SL worksheet number 3

1) State the Le Chatelier's Principle (hasn't been covered yet just google it)
 When a system at \rightleftharpoons is stressed, the system will shift to reduce the stress until a new \rightleftharpoons is established

2) Why is a chemical equilibrium considered dynamic?
 because although macroscopically 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

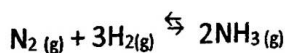
4) At 200 °C the reaction $\text{N}_2(\text{g}) + 3\text{Cl}_2(\text{g}) \rightleftharpoons 2\text{NCl}_3(\text{g})$ with a $K = 15$ (with substance amounts expressed in atm) if a reaction flask initially contains 1 atm of product and 5 atm of each reactant which direction must the reaction proceed to achieve equilibrium?

$$Q = \frac{(P_{\text{NCl}_3})^2}{(P_{\text{N}_2})(P_{\text{Cl}_2})^3} = \frac{(1)^2}{(5)(5)^3} = 0.0016 = Q$$

$Q < K$ OR $K > Q \therefore$ shift to products



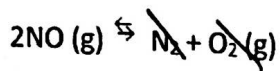
5) What is the reaction quotient Q for the reaction below



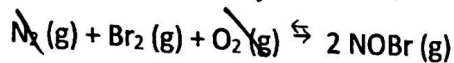
$$Q = \frac{\text{Products}^{\text{coeff.}}}{\text{Reactants}^{\text{coeff.}}} = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3}$$

6) Calculate the value of K_c for the reaction $2 \text{NO}(\text{g}) + \text{Br}_2(\text{g}) \rightleftharpoons 2 \text{NOBr}(\text{g})$

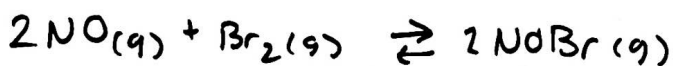
Given the following information



$$K_{c1} = 1 \times 10^{30}$$



$$K_{c2} = 2 \times 10^{-27}$$



nothing is Δ (changed) \therefore

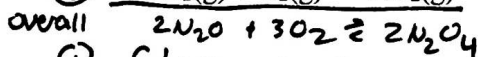
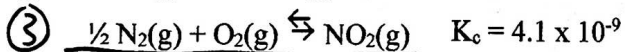
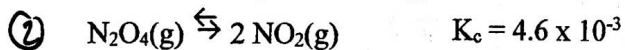
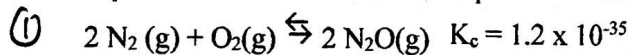
$$K_c = K_{c1} \cdot K_{c2} = (1 \times 10^{30})(2 \times 10^{-27}) = \underline{2 \times 10^3}$$

7) Calculate the value of K_c for the reaction $2 \text{N}_2\text{O}(\text{g}) + 3 \text{O}_2(\text{g}) \rightleftharpoons 2 \text{N}_2\text{O}_4(\text{g})$

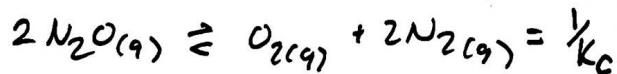
Using the following information

Equation

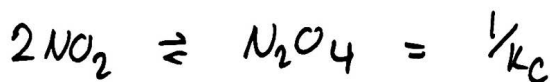
Equilibrium Constant



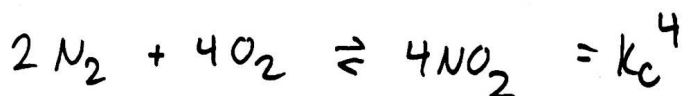
① flip eq. 1



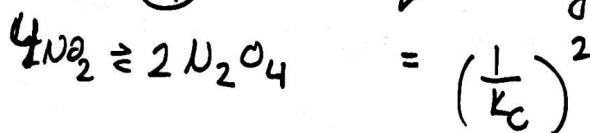
② flip eq. 2.



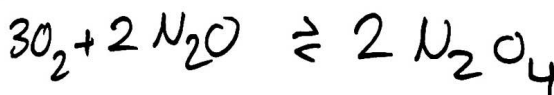
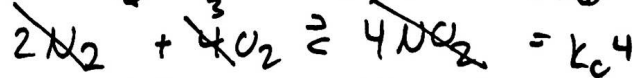
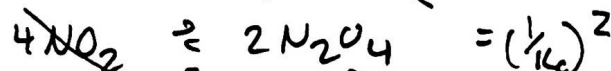
③ times eq. 3 by 4



④ times eq. 2 by 2



\therefore



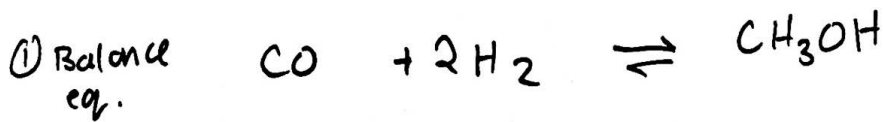
times all new K 's together

$$(8.3 \times 10^{34})(4.7 \times 10^4)(2.8 \times 10^{-34}) =$$

ANSWER = 1.1×10^6

work

- 8) Carbon monoxide reacts with hydrogen to form methanol. All species in this system are gases
 $K_{eq} = 25$ and this reaction releases tons of energy.
- a. If initially there are 2.00 moles of each species in a 2.50 L container in which direction must the system shift to reach equilibrium



② Need M not mols

do this for all of them

$$\begin{aligned} [CO] &= \frac{2.0 \text{ mol}}{2.50 \text{ L}} = 0.80 \text{ M} \\ [H_2] &= 0.80 \text{ M} \\ [CH_3OH] &= 0.80 \text{ M} \end{aligned}$$

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

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

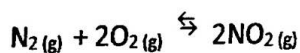
$$= 1.25 \quad K > Q$$

- 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 reactants and products
 will remain unchanged

→ Products

- 9) For the following reaction

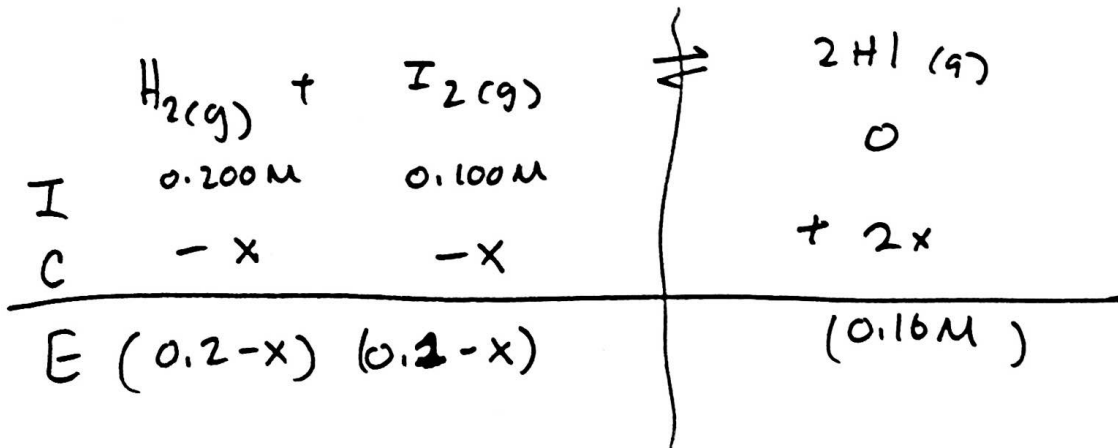


- a. At equilibrium which side is favoured

$$K_{eq} = 4.8 \times 10^{-3} \lll 1$$

∴ reactants are favoured at equilibrium

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 $[\text{HI}] = 0.1600 \text{ M}$. Find the value of K_{eq} .



$$K_{eq} = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$$

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

$K_{eq} = 10.7$

we know $0 + 2x = 0.16 \text{ M}$

\therefore

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

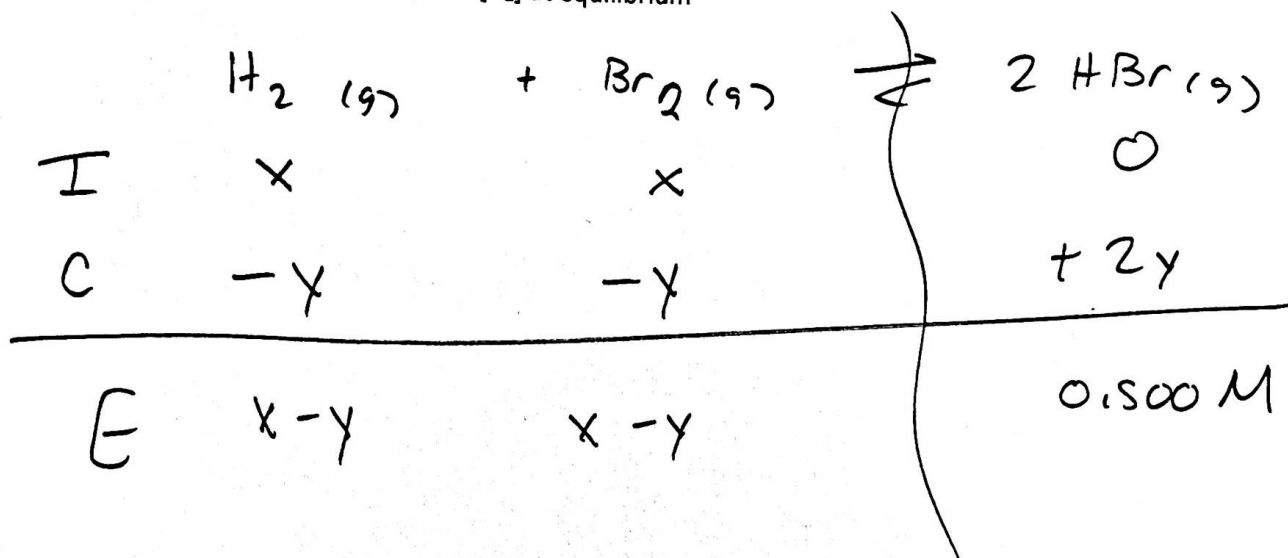
$$x = 0.08 \text{ M}$$

\therefore

$$[\text{H}_2] = 0.2 - 0.08 = 0.12 \text{ M}$$

$$[\text{I}_2] = 0.1 - 0.08 = 0.02 \text{ M}$$

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



Step ① $2y = 0.500 M$

$y = 0.2500 M$

$\therefore [H_2] = x - 0.25$

$[Br_2] = x - 0.25$

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

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

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

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

$1.40x - 0.349 = 0.50$

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

$x = 0.607 M$

$[H_2] = x - y = 0.607 - 0.25$
 at equilibrium $= \underline{0.357 M}$