CHEMICAL EQUILIBRIUM WORKSHEET

On the line at the left, write the letter of the description that best matches each term.

_____ 1. Equilibrium position

_____ 2. Law of chemical equilibrium

_____ 3. Reaction quotient

_____ 4. Law of mass action

_____ 5. Equilibrium constant

a. used to determine if a reaction has reached equilibrium
b. depends on the initial concentrations of the substances in a reaction
c. states that every reaction proceeds to an equilibrium state with a specific K_{eq}
d. expresses the relative concentration of reactants and products at equilibrium in terms of an equilibrium constant
e. the ratio of product concentration to reactant concentration at equilibrium

Answer each of the following in the space provided

6. What is the equilibrium expression for the equation $H_2(g) + I_2(g) \text{ <==> } 2HI(g)$?

7. What is the equilibrium expression for the equation $NH_4Cl(s) \text{ <==> } NH_3(g) + HCl(g)$?

8. What is the equilibrium expression for the equation $As_4O_6(s) + 6C(s) \text{ <==> } As_4(g) + 6CO(g)$?

9. What is the equilibrium expression for the equation $SnO_2(s) + 2CO(g) \text{ <==> } Sn(s) + 2CO_2(g)$?

10. What is the equilibrium expression for the equation $CaCO_3(s) \text{ <==> } CaO(s) + CO_2(g)$?

11. For the reaction $2CO(g) \text{ <==> } C(s) + CO_2(g)$, $K_{eq} = 7.7 \times 10^{-15}$. At a particular time, the following concentrations are measured: $[CO]=0.034$ M, $[CO_2] = 3.6\times10^{-17}$ M. Is this reaction at equilibrium? If not which direction will the reaction proceed?

12. For the reaction $N_2O_4(g) \text{ <==> } 2NO_2(g)$, $K_{eq} = 0.2$. At a particular time, the following concentrations are measured: $[N_2O_4]=2.0$ M, $[NO_2] = 0.2$ M. Is this reaction at equilibrium? If not which direction will the reaction proceed?

13. For the reaction $2ICl(g) \text{ <==> } I_2(g) + Cl_2(g)$, $K_{eq} = 0.11$. At a particular time, the following concentrations are measured: $[ICl]=2.5$ M, $[I_2] = 2.0$ M, $[Cl_2]= 1.2$ M. Is this reaction at equilibrium? If not which direction will the reaction proceed?

14. At 340 °C, $K_{eq} = 0.064$ for the reaction $Fe_2O_3(s) + 3H_2(g) \text{ <==> } 2Fe(s) + 3H_2O(g)$ Given that $[H_2]=0.45$ M and $[H_2O]=0.37$ M, find Q and predict how the reaction will proceed.

Match each statement with the appropriate letter. Each letter can be used once, more than once, or not at all.

_____ 15. The equilibrium concentration of products is much greater than that of reactants.

_____ 16. The equilibrium concentration of products is much less than that of reactants

_____ 17. There is a considerable amount of both reactants and products at equilibrium

a. K_{eq} is much greater than 1.
b. K_{eq} is about equal to 1.
c. K_{eq} is much less than 1.

Complete the following charts by writing left, right or none for equilibrium shift, and decreases, increases or remains the same for the concentrations of reactants and products and for the value of K.

$N_2(g) + 3H_2(g) <==> 2NH_3(g) + 22.0$ kcal

Stress	Equilibrium Shift	$[N_2]$	$[H_2]$	$[NH_3]$	K
18. Add N_2	right	-------	decreases	increases	Remains the same
19. Add H_2			-------		
20. Add NH_3				------	
21. Remove N_2		-------			
22. Remove H_2			------		
23. Remove NH_3				------	
24. Increase Temperature					
25. Decrease Temperature					
26. Increase Pressure					
27. Decrease Pressure					

$NaOH(s) <==> Na^+(aq) + OH^-(aq) +$ 10.6 kcal (Remember that pure solids and liquids do not affect equilibrium values)

Stress	Equilibrium Shift	Amount NaOH(s)	$[Na^+]$	$[OH^-]$	K
28. Add NaOH(s)		-------			
29. Add NaCl (adds Na^+)			-------		
30. Add KOH (Adds OH^-)				-------	
31. Add H^+ (Removes OH^-)				-------	
32. Increase Temperature					
33. Decrease Temperature					
34. Increase Pressure					
35. Decrease Pressure					

CHEMICAL EQUILIBRIUM WORKSHEET

On the line at the left, write the letter of the description that best matches each term.

a	1. Equilibrium position	a. used to determine if a reaction has reached equilibrium
c	2. Law of chemical equilibrium	b. depends on the initial concentrations of the substances in a reaction
b	3. Reaction quotient	c. states that every reaction proceeds to an equilibrium state with a specific K_{eq}
d	4. Law of mass action	d. expresses the relative concentration of reactants and products at equilibrium in terms of an equilibrium constant
e	5. Equilibrium constant	e. the ratio of product concentration to reactant concentration at equilibrium

Answer each of the following in the space provided

6. What is the equilibrium expression for the equation $H_2(g) + I_2(g) \text{ <==> } 2HI(g)$? $K = \frac{[HI]^2}{[H_2][I_2]}$

7. What is the equilibrium expression for the equation $NH_4Cl(s) \text{ <==> } NH_3(g) + HCl(g)$?
$K = [NH_3][HCl]$

8. What is the equilibrium expression for the equation $As_4O_6(s) + 6C(s) \text{ <==> } As_4(g) + 6CO(g)$?
$K = [As_4][CO]^6$

9. What is the equilibrium expression for the equation $SnO_2(s) + 2CO(g) \text{ <==> } Sn(s) + 2CO_2(g)$?
$K = [CO_2]^2 / [CO]^2$

10. What is the equilibrium expression for the equation $CaCO_3(s) \text{ <==> } CaO(s) + CO_2(g)$?
$K = [CO_2]$

11. For the reaction $2CO(g) \text{ <==> } C(s) + CO_2(g)$, $K_{eq} = 7.7 \times 10^{-15}$. At a particular time, the following concentrations are measured: $[CO]=0.034$ M, $[CO_2] = 3.6 \times 10^{-17}$ M. Is this reaction at equilibrium? If not which direction will the reaction proceed?
$Q = [CO_2] / [CO]^2 = 3.6 \times 10^{-17} / (.034)^2 = 3.11 \times 10^{-}$ ND, $K < Q$ left

12. For the reaction $N_2O_4(g) \text{ <==> } 2NO_2(g)$, $K_{eq} = 0.2$. At a particular time, the following concentrations are measured: $[N_2O_4]=2.0$ M, $[NO_2] = 0.2$ M. Is this reaction at equilibrium? If not which direction will the reaction proceed?
$Q = [NO_2]^2 / [N_2O_4] = (.2)^2/2 = 2 \times 10^{-2}$ ND, $K > Q$ right

13. For the reaction $2ICl(g) \text{ <==> } I_2(g) + Cl_2(g)$, $K_{eq} = 0.11$. At a particular time, the following concentrations are measured: $[ICl]=2.5$ M, $[I_2] = 2.0$ M, $[Cl_2] = 1.2$ M. Is this reaction at equilibrium? If not which direction will the reaction proceed?
$Q = [Cl_2][I_2] / [ICl]^2 = (1.2)(2.0)/(2.5)^2 = 0.384$ $K < Q$ left

14. At 340 °C, $K_{eq} = 0.064$ for the reaction $Fe_2O_3(s) + 3H_2(g) \text{ <==> } 2Fe(s) + 3H_2O(g)$ Given that $[H_2]=0.45$ M and $[H_2O]=0.37$ M, find Q and predict how the reaction will proceed.
$Q = [H_2O]^3 / [H_2]^3 = (.37)^3/(.45)^3 = 0.55$ $K < Q$ left

Match each statement with the appropriate letter. Each letter can be used once, more than once, or not at all.

a	15. The equilibrium concentration of products is much greater than that of reactants.	a. K_{eq} is much greater than 1.
c	16. The equilibrium concentration of products is much less than that of reactants	b. K_{eq} is about equal to 1.
b	17. There is a considerable amount of both reactants and products at equilibrium	c. K_{eq} is much less than 1.

$\frac{[Prod]}{[Reactants]}$

Complete the following charts by writing left, right or none for equilibrium shift, and decreases, increases or remains the same for the concentrations of reactants and products and for the value of K.

$N_2(g) + 3H_2(g) <==> 2NH_3(g) + 22.0$ kcal

Stress	Equilibrium Shift	$[N_2]$	$[H_2]$	$[NH_3]$	K
18. Add N_2	right	decrease	decreases	increases	Remains the same
19. Add H_2	→	decrease	-------	increase	" "
20. Add NH_3	←	increase	increase	------	" "
21. Remove N_2	←	-------	increase	decrease	" "
22. Remove H_2	←	↑	------	↓	" "
23. Remove NH_3	→	↓	↓	------	" "
24. Increase Temperature	←	↑	↑	↓	Change
25. Decrease Temperature	→	↓	↓	↑	Change
26. Increase Pressure	→	↓	↓	↑	no change
27. Decrease Pressure	←	↑	↑	↓	no change

$NaOH(s) <==> Na^+(aq) + OH^-(aq) + 10.6$ kcal (Remember that pure solids and liquids do not affect equilibrium values)

Stress	Equilibrium Shift	Amount NaOH(s)	$[Na^+]$	$[OH^-]$	K
28. Add NaOH(s)	→	-------	↑	↑	Same
29. Add NaCl (adds Na^+)	←	↑	-------	↓	Same
30. Add KOH (Adds OH^-)	←	↑	↓	-------	Same
31. Add H^+ (Removes OH^-)	→	↓	↑	-------	Same
32. Increase Temperature	←	↑	↓	↓	Change
33. Decrease Temperature	→	↓	↑	↑	Change
34. Increase Pressure	no shift	—	—	—	—
35. Decrease Pressure	no shift	—	—	—	—

SCH4U
Chemical Systems and Equilibrium

ICE Practice Problems

#1 – Relatively easy, no ICE table required because eq'm concentrations are given
For the reaction $CH_{4(g)} + H_2O_{(g)} \leftrightarrow CO_{(g)} + 3H_{2\ (g)}$ @ 1500° C an equilibrium mixture of these gases was found to have the following concentrations [CO] = 0.300M, [H_2] = 0.800M and [CH_4] = 0.400M. K_c @ 1500° C = 5.67. Determine the equilibrium concentration of H_2O in this mixture.

#2 – Requires an ICE table because you do not know the equilibrium concentrations– no product is yet formed
For the reaction $CO_{(g)} + H_2O_{(g)} \leftrightarrow CO_{2(g)} + H_{2(g)}$ calculate the equilibrium concentrations of all species if 1.000 mol of **each reactant** is mixed in a 1.000L flask. Kc = 5.10 at the temperature of this reaction.

M

		$CO_{(g)}$	+	$H_2O_{(g)}$	↔	$CO_{2(g)}$	+	$H_{2(g)}$
[Initial]	I	1.000		1.000		0		0
[Change in]	C							
[Equilibrium]	E							

#3 – Requires an ICE table because you do not know the equilibrium concentrations. The initial concentrations must be calculated– no product is yet formed
For the reaction $H_{2(g)} + F_{2(g)} \leftrightarrow 2HF_{(g)}$ calculate the equilibrium concentrations of all species if 3.000 mol of **each reactant** was added 1.500L flask. K_c at the temperature of the reaction is 1.15×10^2.

M

		$H_{2(g)}$	+	$F_2(g)$	↔	$2HF_{(g)}$
[Initial]	I	2.00		2.00		0
[Change in]	C	$-x$		$-x$		$+2x$
[Equilibrium]	E	$2.00-x$		$2.00-x$		$2x$

$K = \frac{[HF]^2}{[H_2][F_2]}$

$115 = \frac{(2x)^2}{(2-x)(2-x)}$ $\quad 10.72 = \frac{2x}{2-x}$ $\quad 21.45 - 10.72x = 2x$ $\quad x = 1.69$ $\quad 21.45 = 12.72x$

$[H_2] = [I_2] = 2.00 - 1.69 = 0.31$

$[HF] = 2(1.69) = 3.38$

#4 – Requires an ICE table because you do not know the equilibrium concentrations. Initial concentrations of reactants are given.

0.200mol of H_2, 0.200mol of I_2, and 0.200mol of HI were placed in a 1.00 L flask and allowed to come to equilibrium. The K_c value of the reaction at this temperature is 49.5. Determine the equilibrium concentrations of all species.

M

		$H_{2(g)}$	+	$I_{2(g)}$	↔	$2HI_{(g)}$
[Initial]	I	.200		.200		.200
[Change in]	C	$-x$		$-x$		$+x$
[Equilibrium]	E	$.2-x$		$.2-x$		$.2+x$

$Q = \frac{(.2)^2}{(.2)^2} = 1$ $\quad K > Q \rightarrow$

$K = \frac{[HI]^2}{[H_2][I_2]}$ $\quad 49.5 = \frac{(.2+x)^2}{(.2-x)(.2-x)}$ $\quad 7.04 = \frac{.2+x}{.2-x}$ $\quad 1.41 - 7.04x = .2 + x$ $\quad 1.21 = 8.04x$ $\quad x = .15$

$[H_2] = [I_2] = .2 - .15 = .05$

$[HI] = 2(.15) = .30$

#5 – Requires an ICE table because you do not know the equilibrium concentrations. Initial concentrations of reactants must be calculated and no product is yet formed

For the reaction $H_{2(g)} + F_{2(g)} \rightarrow 2HF_{(g)}$ calculate the equilibrium concentrations of each species if 3.000 mol of H_2 and 6.000mol of F_2 are mixed in a 3.000L flask. K_c at this temperature is 1.15×10^2. $K = [HF]^2/[H_2][F_2]$

		$H_{2(g)}$	+	$F_{2(g)}$	↔	$2HF_{(g)}$
[Initial]	I	1.0		2.0		0
[Change in]	C	$-x$		$-x$		$2x$
[Equilibrium]	E	$1.0-x$		$2.0-x$		$2x$

$115 = \frac{(2x)^2}{(1-x)(2-x)}$

$(2 - x - 2x + x^2) = 4x^2$

$230 - 115x - 230x + 115x^2 = 4x^2$

$111x^2 - 345x + 230 = 0$

$x = 2.14$ or 0.97

$[H_2] = 1.0 - .97 = 0.3$

$[F_2] = 2.0 - .97 = 1.03$

$[HF] = 2(.97) = 1.94$