WORKSHEET: CHEMICAL EQUILIBRIUM

Name Key
Last First

FOR ALL EQUILIBRIUM PROBLEMS, YOU MUST:
1) **Write all equilibrium equations**
2) **Write all equilibrium concentrations**
3) **Write all equilibrium expressions**

SET A:

1) a) What is the equilibrium constant expression for the reaction:

$3\,Fe(s) + 4\,H_2O(g) \rightleftharpoons Fe_3O_4(s) + 4\,H_2(g)$ **Ans:** $[H_2]^4/[H_2O]^4$

b) The equilibrium constant, K_c, for the reaction:

$2\,NOCl(g) \rightleftharpoons 2\,NO(g) + Cl_2(g)$ is 2.4×10^{-7}

What is the equilibrium constant, K_c, for the reaction:

$1/3\ Cl_2(g) + 2/3\ NO(g) \rightleftharpoons 2/3\ NOCl(g)$

$$K = \sqrt[3]{\frac{1}{2.4 \times 10^{-7}}}$$

Ans: 1.6×10^2

c) Given the following equilibrium equations and their corresponding equilibrium constants:

i) $2\,CO_2(g) + H_2O(g) \rightleftharpoons 2\,O_2(g) + CH_2CO(g)$ $K_c = 6.1 \times 10^8$

ii) $CH_4(g) + 2\,O_2(g) \rightleftharpoons CO_2(g) + 2\,H_2O(g)$ $K_c = 1.2 \times 10^{14}$

Find K_c for the reaction: $CH_4(g) + CO_2(g) \rightleftharpoons CH_2CO(g) + H_2O(g)$

Setup:

Add equations (i) and (ii). The sum is equation (iii)

$$K_{overall} = K_{(i)} \cdot K_{(ii)} = (6.1 \times 10^8)(1.2 \times 10^{14}) = 7.3 \times 10^{22}$$

Ans: 7.3×10^{22}

2) A mixture of 9.22 moles of **A**, 10.11 moles of **B**, and 27.83 moles of **C** is placed in a one-liter container at a certain temperature. The reaction is allowed to reach equilibrium. At equilibrium the number of moles of **B** is 18.32. Calculate the equilibrium constant for the reaction:

A (g) + 2 **B** (g) $\rightleftharpoons$ 3 **C** (g

	A	B	C
Initial conc	9.22	10.11	27.83
change in conc.	+ 4.10	+ 8.20	− 12.3
Equilibrium conc.	9.22+4.10 = 13.32	18.32	27.83 − 12.31 = 15.51

$$8.21 \text{ moles "B" reacted} \left(\frac{3 \text{ moles C}}{2 \text{ moles B}}\right) = 12.3 \text{ moles "C" reacted}$$

$$K_c = \frac{[C]^3}{[A][B]^2} = \frac{(15.51)^3}{(13.22)(18.32)^2} = 0.832$$

Ans: 0.832

3) a) At a certain temperature, K_c is 4.13×10^{-2} for the equilibrium:

$$2\ IBr\ (g) \rightleftharpoons I_2\ (g) + Br_2\ (g)$$

Assume that equilibrium is established at the above temperature by adding only IBr (g) to the reaction flask. What are the concentrations of I_2 (g) and Br_2 (g) in equilibrium with 0.0124 moles/liter of IBr(g) ?

Setup:

	$2\ IBr(g)$ $\rightleftharpoons$	$I_2(g)$ +	$Br_2(g)$
...itial conc.	$M_{initial}$	0	0
...e in conc	$-2x$	$+x$	$+x$
...ui. conc.	$M_{initial} - 2x = .0124$	x	x

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

$$\sqrt{4.13 \times 10^{-2}} = \sqrt{\frac{(x)(x)}{(.0124)^2}}$$

$x = 2.52 \times 10^{-3}$

$[H_2] = [I_2] = x = 2.52 \times 10^{-3}$

Ans: 2.52×10^{-3} M

b) What was the **initial** concentration of IBr before equilibrium was established?

Setup:

$M_{Initial} - 2x = .0124$

$M_{Initial} = .0124 + 2x$

$= .0124 + 2(2.52 \times 10^{-3})$

$= .0174$

Ans: 0.0174

4) 0.924 mole of **A** (g) is placed in a 1.00 liter container at 700 ° C, where it is 38.8 % dissociated when equilibrium was established.

$$3\ \mathbf{A}\ (g) \rightleftharpoons 5\ \mathbf{B}\ (g) + 2\ \mathbf{C}\ (g)$$

What is the value of the equilibrium constant, K_c, at the same temperature?

Setup:

	$3A(g)$ $\rightleftharpoons$	$5B(g)$ +	$2C(g)$
...nitial conc.	0.924	0	0
... in conc.	$-.359$	$+.598$	$+0.239$
...uilibrium conc	$0.924 - .359$ react $= .566$	0.598	0.239

$\frac{38.8}{100} \times .924$ mol $= .359$ moles "A" reacted

0.359 moles A $\left(\frac{2 \text{ moles C}}{3 \text{ moles A}}\right) = 0.239$... formed

0.359 mole A $\left(\frac{5 \text{ moles B}}{3 \text{ moles A}}\right) = .598$... formed

Ans: 0.0241

$$K_c = \frac{[B]^5[C]^2}{[A]^3} = \frac{(.598)^5(.239)^2}{(.566)^3} = .0241$$

5) The equilibrium constant for the reaction:

$$2\ NO\ (g) \rightleftharpoons N_2\ (g) + O_2\ (g)$$

is 2.60×10^{-3} at 1100 °C. If 0.820 mole of NO (g) and 0.223 mole **each** of N_2 (g) and O_2 (g) are mixed in a 1.00 liter container at 1100 °C, what are the concentrations of NO (g), N_2(g), and O_2 (g) at equilibrium?

Setup:

$$Q = \frac{(N_2)(Cl_2)}{(NO)^2} = \frac{(.0223)(.0223)}{(.820)^2}$$

$$= 7.40 \times 10^{-2}$$

Q > K ; The reaction is proceeding spontaneously to the left.

$$2NO\ (g) \rightleftharpoons N_2\ (g) + Cl_2\ (g)$$

	2NO (g)	N_2 (g)	Cl_2 (g)
Initial conc	.820	.223	.223
change in conc	+ 2x	−x	−x
Equi. conc.	.820+2x	.223 − x	.223 − x

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

$$\sqrt{2.60 \times 10^{-3}} = \sqrt{\frac{(0.223-x)(.223-x)}{(.820+2x)^2}}$$

$$0.0510 = \frac{(.223-x)}{(.820+2x)}$$

$$x = 0.164\ M$$

$$[O_2] = [N_2] = 0.223 - x = .223 - .164 = .058\ M$$

$$[NO] = 0.820 + 2x = .820 + 2(.164)$$

$$= 1.15\ M$$

Ans: [NO]= 1.15 M
$[N_2]$ = 0.058 M
$[Cl_2]$= 0.058 M

SET B :

1) A mixture of 1.16 mole of **A**, 1.35 mole of **B** and 0.641 mole of **C** is placed in a one-liter container at a certain temperature. The reaction was allowed to reach equilibrium. At equilibrium. the number of moles of **A** is 1.95. Calculate the equilibrium constant, Kc, for the reaction:

$$2\ A\ (g) \rightleftharpoons 2\ B\ (g) + C\ (g)$$

Setup:

	2 A (g)	2 B (g)	C (g)
Initial conc.	1.16	1.35	0.641
Change in conc.	+.79	−0.79	$-0.\frac{79}{2}$
Equi. conc.	1.95	(1.35 − .79) = .56	$(.641 - \frac{.79}{2})$ = .246

#moles A formed = 1.95 − 1.16 = 0.79

$$K_c = \frac{[B]^2[C]}{[A]^2} = \frac{(.56)^2(.246)}{(1.95)^2} = 0.020$$

Ans: 0.020

2) 0.822 mole of SO_3 (g) is placed in a 1.00 liter container at 600 K. 36.7 % of the SO_3 (g) are decomposed when equilibrium is established.

$$2\ SO_3\ (g) \rightleftharpoons 2\ SO_2\ (g) + O_2\ (g)$$

What is the value of the equilibrium constant , K_c, at the same temperature?

Setup:

	$2\ SO_3 (g)$	$2\ SO_2 (g)$	$O_2 (g)$
Initial conc.	0.822	0	0
Change in conc	−0.302	+0.302	+(0.302/2)
Equi conc.	.822 (initial) − .302 (react) = .520	.302 formed	$(\frac{.302}{2})$ formed

$$K = \frac{[SO_2]^2[O_2]}{[SO_3]^2} = \frac{(.302)^2(.151)}{(.520)^2} = 5.05 \times 10^{-2}$$

Ans: 5.05×10^{-2}

3) For the equilibrium:

$$2\ BrCl\ (g) \rightleftharpoons Br_2\ (g) + Cl_2\ (g)$$

at 205 ° C, the equilibrium constant, K_c, is 0.143. If 1.34 moles **each** of Br_2 (g) and Cl_2 (g) are introduced in a container which has a volume of 11.0 liters and allowed to reach equilibrium at 205 ° C, what would be the concentrations of Br_2 (g), Cl_2 (g), and BrCl (g) at equilibrium?

Setup:

	$2BrCl\ (g)$ $\rightleftharpoons$	$Br_2\ (g)$ +	$Cl_2\ (g)$
Initial conc.	0	$\frac{1.34\ mole}{11.0\ L} = .122$	.122
Change in conc	$+2x$	$-x$	$-x$
Equi conc.	$2x$	$0.122 - x$	$.122 - x$

$$K = \frac{[Br_2][Cl_2]}{[BrCl]^2} \quad ; \quad \sqrt{.143} = \sqrt{\frac{(.122-x)(.122-x)}{(2x)^2}}$$

$$x = .0693$$

$[Br_2] = [Cl_2] = .122 - x = .122 - .0693 = .0529M$

$BrCl = 2x = 2(.0693) = .139\ M$

Ans: $[Cl_2]$= 0.0529 M
$[Br_2]$= 0.0529 M
[BrCl] = 0.139 M

4) a) What is the numerical value of the equilibrium constant, K_c, for the reaction:

$$3\ N_2\ (g) + 3\ O_2\ (g) \rightleftharpoons 6\ NO\ (g)$$

if the equilibrium constant for the reaction:

$$2\ NO\ (g) \rightleftharpoons N_2\ (g) + O_2\ (g) \quad \text{is } 3.5 \times 10^{-6}$$

Setup:

$$K = \left(\frac{1}{(3.5 \times 10^{-6})}\right)^3 = (0.286 \times 10^6)^3 = 2.3 \times 10^{16}$$

Ans: 2.3×10^{16}

b) What is the equilibrium constant expression for the reaction:

$$2\ Ni\ (s) + 2\ CO_2\ (g) \rightleftharpoons 2\ CO\ (g) + 2\ NiO\ (s)$$

Ans: $K = [CO]^2/[CO2]^2$

5) For the reaction:

$$2\ NO_2\ (g) + O_2\ (g) \rightleftharpoons 2\ NO_3\ (g)$$

at 923 ° C , K_c is 42.5. If 0.0500 mole of NO_2 (g), 0.122 mole of O_2 (g) and 0.300 mole of NO_3 (g) are mixed in a 1.00 liter container at 923 ° C,in what direction will the reaction proceed?(Show your calculation to prove that your answer is not a guess.)

Setup:

$$2\ NO_2\ (g) + O_2\ (g) \rightleftharpoons 2\ NO_3\ (g)$$

.0500   .122   .300

$$Q = \frac{(NO_3)^2}{(NO_2)^2(O_2)} = \frac{(0.300)^2}{(.0500)^2(.122)} = 295$$

$Q > K$; The reaction will proceed spontaneously to the left.

Answer: $Q_c > Kc$ (The reaction will proceed spontaneously to the left)

SET C :

1) For the equilibrium:

$$2\ NO\ (g) \rightleftharpoons N_2\ (g) + O_2\ (g)$$

at 300 K, the equilibrium constant, K_c, is 0.185. If 1.45 moles **each** of N_2 (g) and O_2 (g) are introduced in a container that has a volume of 6.00 liters and allowed to reach equilibrium at 300 K, what are the concentrations of N_2 (g), O_2(g), and NO (g) at equilibrium?

Setup:

	2NO (g) ⇌	N_2 (g) +	O_2 (g)
Initial conc.	0	$\frac{1.45}{6.0} = .24$	$\frac{1.45}{6.0} = .24$
change in conc.	$+2x$	$-x$	$-x$
Equi conc	$2x$	$.242-x$	$0.242-x$

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

$$\sqrt{.185} = \sqrt{\frac{(0.242-x)(0.242-x)}{(2x)^2}}$$

$x = .130$

$[N_2] = [O_2] = .242 - x = .242 - .130 = .112\ M$

$[NO] = 2x = 2(.130) = .260\ M$

Ans: $[N_2]$= 0.112 M
$[O_2]$= 0.112 M
[NO]= 0.260 M

2) 1.87 mole of **A** (g) are placed in a 1.00 liter container at 700 K. 43.3 % of **A** (g) had decomposed when the reaction reached equilibrium.

$$2\ A(g) \rightleftharpoons 3\ \mathbf{B}\ (g) + \mathbf{C}\ (g)$$

What is the value of the equilibrium constant, K_c, at the same temperature?

Setup:

	2 A (g) ⇌	3 B (g) +	C (g)
Initial conc.	1.87	0	0
change in conc.	$-.810$	$+1.21$	$+.405$
Equi conc.	$1.87 - .810 = 1.06$	1.21	0.405

$\frac{43.3}{100} \times 1.87$ mole A = .810 mole A reacted

0.810 mole "A" reacting $\left(\frac{3\text{ moles B}}{2\text{ moles A}}\right)$ = 1.21 moles B formed

0.810 mole "A" reacting $\left(\frac{1\text{ mole C}}{2\text{ moles A}}\right)$ = .405 mole C formed

$$K = \frac{[B]^3[C]}{[A]^2} = \frac{(1.21)^3(.405)}{1.06} = 0.639$$

Ans: 0.639

3) a) What is the equilibrium constant expression for the reaction:
$2\ Fe(s) + 2\ NO_2(g) \rightleftharpoons 2\ NO(g) + 2\ FeO(s)$

Ans: $K = [NO]^2/[NO_2]^2$

b) What is the numerical value of the equilibrium constant, K_c, for the reaction:

$$2\ SO_3(g) \rightleftharpoons 2\ SO_2(g) + O_2(g)$$

if the equilibrium constant for the reaction:

$$1/2\ O_2(g) + SO_2(g) \rightleftharpoons SO_3(g)$$ is 2.7×10^{-3} ?

Setup:

$$K = \frac{1}{(2.7\times10^{-3})^2} = 1.4\times10^5$$

Ans: 1.4×10^5

4) A mixture of 3.31 moles of **A**, 4.33 moles of **B**, and 5.95 moles of **C** is placed in a one-liter container at a certain temperature. The reaction was allowed to reach equilibrium. At equilibrium, the number of moles of **B** is 6.16 . Calculate the equilibrium constant, K_c, for the reaction:

$$\mathbf{A}\ (g) \rightleftharpoons 3\ \mathbf{B}\ (g) + 2\ \mathbf{C}\ g)$$

Setup:

Initial conc.	3.31	4.33	5.95
Equi conc.	?	6.16 M	?

$$1.83 \text{ moles B formed} \left(\frac{2 \text{ moles C}}{3 \text{ moles B}}\right) = 1.22 \text{ moles C formed}$$

$$1.83 \text{ moles B formed} \left(\frac{1 \text{ mole A}}{3 \text{ moles B}}\right) = 0.61 \text{ mole A react}$$

	A (g) $\rightleftharpoons$	3B (g)	+ 2 C (g)
Initial conc.	3.31	4.33	5.95
change in conc	− 0.61	+ 1.83	+1.22
Equi conc	3.31 − .61 = 2.70	6.16	5.95 + 1.22 =

$$K = \frac{[B]^3[C]^2}{[A]} = \frac{(6.16)^3(7.17)^2}{2.7} = 4.5\times10^3$$

Ans: 4.5×10^3

5) For the reaction:

$$2\,A\,(g) \rightleftharpoons B\,(g) + C\,(g)$$

at 900 °C, K_c is 1.40 x 10^{-3}. If 0.780 mole of **A** (g) and 0.244 mole **each** of **B** (g) and **C** (g) are mixed in a 1.00 liter container at 900 °C, what are the concentrations of **A**, **B**, and **C** at equilibrium?

Setup:

$$Q = \frac{(B)(C)}{(A)^2} = \frac{(.244)(.244)}{(0.780)^2} = 0.0978$$

$$\underset{.0978}{Q} > \underset{1.4 \times 10^{-3}}{K}$$

The reaction will proceed spontaneously to the left in order to reach equilibrium

$$2A\,(g) \rightleftharpoons B\,(g) + C\,(g)$$

	2A (g)	B (g)	C (g)
Initial conc	.780	.244	.244
change in conc	+2x	−x	−x
Equi conc.	.780 +2x	.244 − x	.244 − x

$$K = \frac{[B][C]}{[A]^2}$$

$$\sqrt{1.40 \times 10^{-3}} = \sqrt{\frac{(0.244 - x)(.244 - x)}{(.780 + 2x)^2}}$$

$$3.74 \times 10^{-2} = \frac{(.244 - x)}{(.780 + 2x)}$$

$x = 0.200$

$[B] = [C] = .244 - x = .244 - .200 = .044\ M$

$[A] = .780 + 2x = .780 + 2(.200) = 1.180\ M$

Ans: [A]= 1.180 M
[B]= 0.044 M
[C]=0.044 M