WORKSHEET: CHEMICAL EQUILIBRIUM Name ____________________
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:** $[H_2]^4/[H_2O]^4$

compare eqns

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

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

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

$1/3\ Cl_2\ (g) + 2/3\ NO\ (g) \rightleftharpoons 2/3\ NOCl\ (g)$ $K_c = \boxed{}$

Ans: 1.6×10^2

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

Have:

a) $2\ CO_2\ (g) + H_2O\ (g) \rightleftharpoons 2\ O_2\ (g) + CH_2CO\ (g)$ $Kc = 6.1 \times 10^8$

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

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

Setup:

$K_{overall} = 6.1 \times 10^8 \times 1.2 \times 10^{14}$
$= 7.3 \times 10^{22}$

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
I	9.22	10.11	27.83
C	+4.10	+2x	−3(4.10)
E	(13.32)	(18.32)	15.51

import

Kc calc

Calc "x"

$10.11 + 2x = 18.32$
$x = 4.10$

$$K_c = \frac{[C]^3}{[A][B]^2}$$

$$= \frac{(27.83 - 3x)^3}{(9.22 + x)(18.32)^2}$$

$$= \frac{[27.83 - 3(4.10)]^3}{(9.22 + 4.10)(18.32)^2}$$

Ans: 0.832

$= .832$

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:

	$2IBr$	$\rightleftharpoons$	I_2	+	Br_2
I	$\boxed{Y}$		∅		∅
C	$-2x$		$+x$		$+x$
E	(.0124		x		x

Kc calc

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

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

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

∴ $Y - 2x = .0124$

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

$Y = 2.52 \times 10^{-3} M$

Ans: 2.52×10^{-3} M

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

Setup:

$M_{IBr\ initial} - 2x = 0.0124$

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

$M_{IBr\ initial} = 0.0174M$

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.

change

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

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

Setup:

	$3A$	$\rightleftharpoons$	$5B$	+	$2C$
I	.924		∅		∅
C	$-.329$		~~$+5x$~~ $+5(.120)$		~~$+2x$~~ $+2(.120)$
E	(.566)		(.598)		(.239)

export

Kc calc

$$K_c = \frac{[B]^5 [C]^2}{[A]^3}$$

$$= \frac{(.598)^5 (.239)^2}{(.566)^3}$$

$$= .0241$$

"x" calc

$.924\ M\ A_{ini} \times \frac{38.8}{100}$

$3X = .329\ M$ "A" dissociated

∴ $x = .120$

Ans: 0.0241

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:

	$2NO$	$\rightleftharpoons$	N_2	+	O_2	
I	.820		.223		.223	← only Initial!
C	+2x		−x		−x	A ha!
E	$\boxed{(.820+2x)}$		$\boxed{(.223-x)}$		$\boxed{(.223-x)}$	I need Qc!

Qc calc

• $Q_c = \dfrac{[N_2]_{ini}\ [O_2]_{ini}}{[NO]^2_{ini}}$

$= \dfrac{(.223)\ (.223)}{(.820)^2} = 7.40 \times 10^{-2}$

• $Q_c > K_c$

$7.40 \times 10^{-2} > 2.60 \times 10^{-3}$

∴ net rxn to the left

Kc calc

$K_c = \dfrac{[N_2][O_2]}{[NO]^2}$

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

$x = 0.164M$

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

$[NO] = .820 + 2(.164) = 1.15M$

$[N_2] = [O_2] = .223 - .164 = 0.058M$

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, K_C, for the reaction:

Setup:

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

	2 A	2 B	C
I	1.16	1.35	.641
C	+ 2x	− 2(.40)	− .40
E	(1.95)	(.55)	(.24)

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

$$K_c = \frac{(.55)^2(.24)}{(1.95)^2} = 0.020$$

Calc "x"

$1.16 + 2x = 1.95$

$x = .40$

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	2 SO_2	O_2
I	.822	Ø	Ø
C	− $\cancel{2x}$ .302	+ 2(.151)	+ .151
E	.520	.302	.151

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

Calc "x"

$.822\,M \times \frac{36.7}{100}$

C = .302

$2x = .302$

$x = 0.151$

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:

	$2BrCl$	$\rightleftharpoons$	Br_2	+	Cl_2
I	∅		$\frac{1.34}{11.0}$		$\frac{1.34}{11.0}$
C	$+2x$		$-x$		$-x$
E	$(2x)$		$(.122-x)$		$(.122-x)$

$$K = \frac{[Br_2][Cl_2]}{[BrCl]^2}$$

$$\sqrt{.143} = \sqrt{\frac{(.122-x)(.122-x)}{(2x)^2}}$$

$x = .0693$

$[BrCl] = 2x = 2(.0693) = .139M$

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

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:

$$\boxed{3}N_2\ (g) + \boxed{3}O_2\ (g) \rightleftharpoons \boxed{6}NO\ (g) \quad K_c = \boxed{} \leftarrow \text{calc}$$

P

if the equilibrium constant for the reaction:

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

R

Setup:

$$3\left((1)N_2 + (1)O_2 \rightleftharpoons (2)NO\right) \quad K_c' = \left(\frac{1}{3.5 \times 10^{-6}}\right)^3$$

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:

	$2NO_2$	+	O_2	$\rightleftharpoons$	$2NO_3$
I	$.0500^2$		$.122$		$.300$
C	$+2x$		$+x$		$-2x$
E					

1st $$Q_c = \frac{[NO_3]^2_{ini}}{[NO_2]^2_{ini}[O_2]_{ini}} = \frac{(.300)^2}{(.0500)^2(.122)} = 295$$

2nd $Q_c > K_c$ ∴ Too much Product

295   42.5 → A net rxn to the left until equilibria is re-established!

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:

	$2NO$	$\rightleftharpoons$	N_2	+	O_2
I	$\emptyset$		$\frac{1.45}{6.00}$		$\frac{1.45}{6.00}$
C	$+2x$		$-x$		$-x$
E	$(2x)$		$(.242-x)$		$(.242-x)$

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

$$\sqrt{.185} = \sqrt{\frac{(.242-x)(.242-x)}{(2x)^2}}$$

$$x = .130$$

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

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

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 .

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

"C" calc
calc "x"

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

Setup:

	$2A$	$\rightleftharpoons$	$3B$	+	C
I	1.87		$\emptyset$		$\emptyset$
C	$-2(.405)$		$+3(.405)$		$+.405$
E	(1.06)		(1.21)		$(.405)$

x

$$1.87 \times \frac{43.3}{100} = 2x$$

$$x = .405$$

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

Ans: 0.639

$$K_c = .639$$

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:

$$2\ SO_3\ (g) \rightleftharpoons 2\ SO_2\ (g) + O_2\ (g) \qquad K_c = \boxed{}$$

if the equilibrium constant for the reaction given below is is 2.7 x 10^{-3} ?

Given

$$1/2\ O_2\ (g) + SO_2\ (g) \rightleftharpoons SO_3\ (g) \qquad K_c = 2.7 \times 10^{-3}$$

Setup:

$$K_c = \frac{[SO_3]}{[O_2]^{1/2}[SO_2]} = 2.7 \times 10^{-3}$$

$$K_c = \frac{[SO_2]^2[O_2]}{[SO_3]^2} = \left(\frac{1}{2.7 \times 10^{-3}}\right)^2$$

Ans: 1.4 x 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:

	A	⇌ 3 B	+ 2 C
I	3.31	4.33	5.95
C	−.610	+3(.610)	+2(.610)
E	2.70	6.16	7.17

"x" calc

$4.33 + 3x = 6.16$

$x = .610$

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

$$= 4.5 \times 10^3$$

Ans: 4.5 x 10^3

5. For the reaction:

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

at 900 °C, K_c is 1.40×10^{-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:

Step 1 Calc Q_c & compare to K_c

a) $Q_c = \dfrac{[B]_{ini}\,[C]_{ini}}{[A]^2_{ini}} = \dfrac{(.244)\ (.244)}{(.780)^2}$

$Q_c = .0978$

b) $Q_c > K_c$

$.0978 > 1.40 \times 10^{-3}$

Step 2

	$2A$	$\rightleftharpoons$	B	+	C
I	.780		.244		.244
C	$+2x$		$-x$		$-x$
E	$(.780+2x)$		$(.244-x)$		$(.244-x)$

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

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

$x = .200$

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

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

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