

WORKSHEET: CHEMICAL EQUILIBRIUM

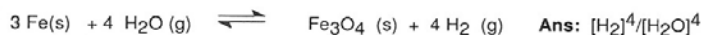
Name _____
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FOR ALL EQUILIBRIUM PROBLEMS, YOU MUST:

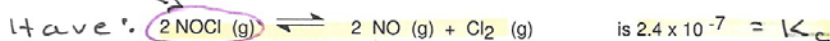
- 1) Write all equilibrium equations
- 2) Write all equilibrium concentrations
- 3) Write all equilibrium expressions

SET A:

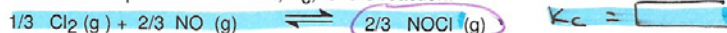
1. a) What is the equilibrium constant expression for the reaction:



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

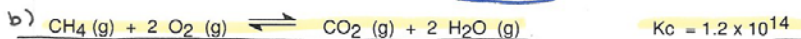
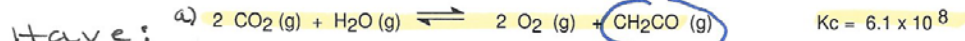


What is the equilibrium constant, K_c , for the reaction:



Ans: 1.6×10^2

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



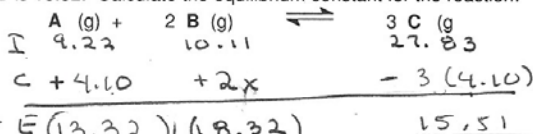
Find K_c for the reaction: $\text{CH}_4 \text{ (g)} + \text{CO}_2 \text{ (g)} \rightleftharpoons \text{CH}_2\text{CO (g)} + \text{H}_2\text{O (g)}$

Setup:

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

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:



import
 K_c calc

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

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

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

$$= 0.832$$

Calc "x"

$$10.11 + 2x = 18.32$$

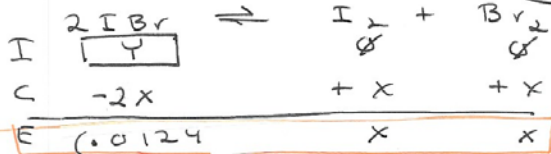
$$x = 4.10$$

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



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:



K_c calc

$$K_c = \frac{[\text{I}_2][\text{Br}_2]}{[\text{IBr}]^2}$$

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

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

$$\begin{aligned} \therefore Y - 2x &= .0124 \\ Y - 2(2.52 \times 10^{-3}) &= .0124 \\ Y &= 2.52 \times 10^{-3} \text{ M} \end{aligned}$$

Ans: $2.52 \times 10^{-3} \text{ M}$

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

Setup:

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

$$M_{\text{IBr initial}} = 0.0174 \text{ M}$$

$$M_{\text{IBr initial}} - 2(2.52 \times 10^{-3}) = 0.0124$$

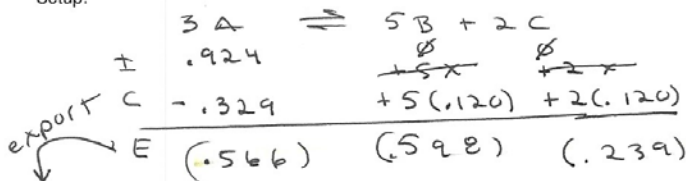
Ans: 0.0174

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.



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

Setup:



K_c calc

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

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

$$= 0.0241$$

Ans: 0.0241

change

"x" calc

$$.924 \text{ M A}_{\text{ini}} \times \frac{38.8}{100}$$

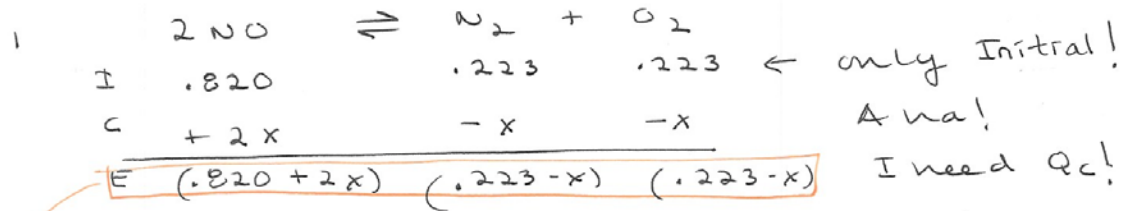
$$3x = .329 \text{ M "A" dissociated}$$

$$\therefore x = .120$$

5. The equilibrium constant for the reaction:



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_c calc

$$\begin{aligned}
 Q_c &= \frac{[\text{N}_2]_{\text{ini}} [\text{O}_2]_{\text{ini}}}{[\text{NO}]_{\text{ini}}^2} \\
 &= \frac{(0.223)(0.223)}{(0.820)^2} = 7.40 \times 10^{-2}
 \end{aligned}$$

$$\begin{aligned}
 Q_c &> K_c \\
 7.40 \times 10^{-2} &> 2.60 \times 10^{-3} \\
 \therefore \text{net rxn to the left}
 \end{aligned}$$

K_c calc

$$K_c = \frac{[\text{N}_2][\text{O}_2]}{[\text{NO}]^2}$$

$$\sqrt{2.60 \times 10^{-3}} = \sqrt{\frac{(0.223 - x)(0.223 - x)}{(0.820 + 2x)^2}}$$

$$x = 0.164 \text{ M}$$

Ans: $[\text{NO}] = 1.15 \text{ M}$
 $[\text{N}_2] = 0.058 \text{ M}$
 $[\text{O}_2] = 0.058 \text{ M}$

$$[\text{NO}] = .820 + 2(.164) = 1.15 \text{ M}$$

$$[\text{N}_2] = [\text{O}_2] = .223 - .164 = 0.058 \text{ M}$$

(3)

3. For the equilibrium:



at 205 °C, the equilibrium constant, K_c , is 0.143. If 1.34 moles each of $\text{Br}_2 (g)$ and $\text{Cl}_2 (g)$ 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 $\text{Br}_2 (g)$, $\text{Cl}_2 (g)$, and $\text{BrCl} (g)$ at equilibrium?

Setup:

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

$$K_c = \frac{[\text{Br}_2][\text{Cl}_2]}{[\text{BrCl}]^2}$$

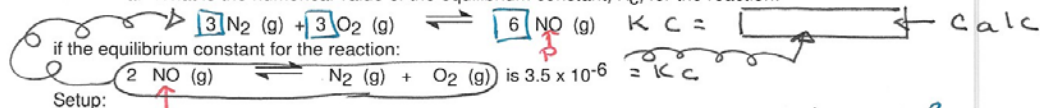
$$0.143 = \frac{(.122 - x)(.122 - x)}{(2x)^2}$$

$$x = .0693$$

$[\text{BrCl}] = 2x = 2(.0693) = .139 \text{ M}$
 $[\text{Br}_2] = [\text{Cl}_2] = .122 - .0693 = .0529 \text{ M}$

Ans: $[\text{Cl}_2] = 0.0529 \text{ M}$
 $[\text{Br}_2] = 0.0529 \text{ M}$
 $[\text{BrCl}] = 0.139 \text{ M}$

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



Setup:

$$3 (2 \text{NO} \rightleftharpoons \text{N}_2 + \text{O}_2) \quad K_c' = \left(\frac{1}{3.5 \times 10^{-6}} \right)^3$$

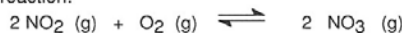
Ans: 2.3×10^{16}

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



Ans: $K = [\text{CO}]^2 / [\text{CO}_2]^2$

5. For the reaction:



at 923 °C, K_c is 42.5. If 0.0500 mole of $\text{NO}_2 (g)$, 0.122 mole of $\text{O}_2 (g)$ and 0.300 mole of $\text{NO}_3 (g)$ 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	$.0500$		$.122$		$.300$
C	$+ 2x$		$+ x$		$- 2x$
E					

1st $Q_c = \frac{[\text{NO}_3]_{\text{ini}}^2}{[\text{NO}_2]_{\text{ini}}^2 [\text{O}_2]_{\text{ini}}} = \frac{(.300)^2}{(.0500)^2 (.122)} = 295$

2nd $Q_c > K_c$
 $295 > 42.5$
 \therefore Too much Product
 \rightarrow A net rxn to the left until equilibrium is re-established.

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

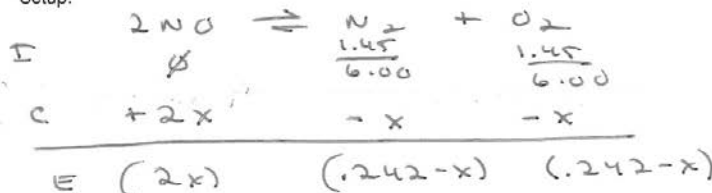
SET C:

1. For the equilibrium:



at 300 K, the equilibrium constant, K_c , is 0.185. If 1.45 moles each of $\text{N}_2 (g)$ and $\text{O}_2 (g)$ 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 $\text{N}_2 (g)$, $\text{O}_2 (g)$, and $\text{NO} (g)$ at equilibrium?

Setup:



$$K = \frac{[\text{N}_2][\text{O}_2]}{[\text{NO}]^2}$$

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

$$x = .130$$

$$[\text{N}_2] = [\text{O}_2] = .242 - x$$

$$= .242 - .130$$

$$= .112 \text{ M}$$

$$[\text{NO}] = 2x = 2(.130)$$

$$= .260 \text{ M}$$

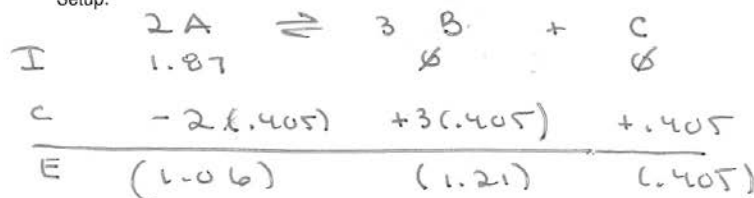
Ans: $[\text{N}_2] = 0.112 \text{ M}$
 $[\text{O}_2] = 0.112 \text{ M}$
 $[\text{NO}] = 0.260 \text{ M}$

2. 1.87 mole of $\text{A} (g)$ are placed in a 1.00 liter container at 700 K. decomposed when the reaction reached equilibrium.



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

Setup:



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

$$= \frac{(1.21)^3 (.405)}{(1.06)^2}$$

$$K_c = .639$$

Ans: 0.639

43.3% of $\text{A} (g)$ had

"C" calc
Calc "x"

$$x$$

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

$$x = .405$$

3. a. What is the equilibrium constant expression for the reaction:
 $2 \text{Fe (s)} + 2 \text{NO}_2 \text{ (g)} \rightleftharpoons 2 \text{NO (g)} + 2 \text{FeO (s)}$

Ans: $K = \frac{[\text{NO}]^2}{[\text{NO}_2]^2}$

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



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

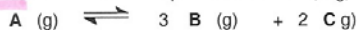
Given Setup: $\frac{1}{2} \text{O}_2 \text{ (g)} + \text{SO}_2 \text{ (g)} \rightleftharpoons \text{SO}_3 \text{ (g)}$ $K_c = 2.7 \times 10^{-3}$

$K_c = \frac{[\text{SO}_3]}{[\text{O}_2]^{1/2} [\text{SO}_2]} = 2.7 \times 10^{-3}$

$K_c = \frac{[\text{SO}_2]^2 [\text{O}_2]}{[\text{SO}_3]^2} = \left(\frac{1}{2.7 \times 10^{-3}} \right)^2$

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:



Setup:

	A	\rightleftharpoons	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

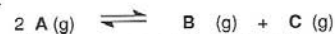
$4.33 + 3x = 6.16$
$x = .610$

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

$= 4.5 \times 10^3$

Ans: 4.5×10^3

5. For the reaction:



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

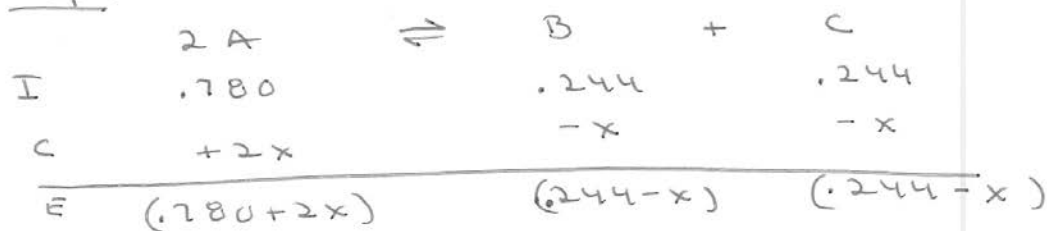
$$a) Q_c = \frac{[\text{B}]_{\text{ini}} [\text{C}]_{\text{ini}}}{[\text{A}]_{\text{ini}}^2} = \frac{(0.244) (0.244)}{(0.780)^2}$$

$$Q_c = .0978$$

$$b) Q_c > K_c$$

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

Step 2



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

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

$$x = .200$$

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

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

Ans: $[\text{A}] = 1.180 \text{ M}$
 $[\text{B}] = 0.044 \text{ M}$
 $[\text{C}] = 0.044 \text{ M}$