

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

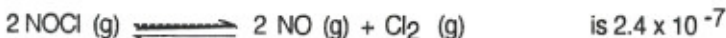
SET A:

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

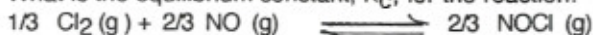


Ans:  $[\text{H}_2]^4 / [\text{H}_2\text{O}]^4$

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



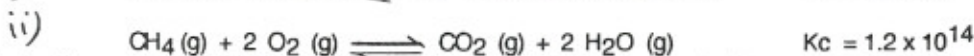
What is the equilibrium constant,  $K_c$ , for the reaction:



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

Ans:  $1.6 \times 10^2$

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



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

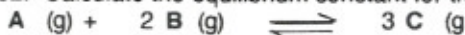
Setup:

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

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

Ans:  $7.3 \times 10^{22}$

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 moles of B is 18.32. Calculate the equilibrium constant for the reaction:



|                   |                      |       |                       |
|-------------------|----------------------|-------|-----------------------|
| Initial conc      | 9.22                 | 10.11 | 27.83                 |
| change in conc.   | +4.10                | +8.20 | -12.3                 |
| Equilibrium conc. | 9.22+4.10<br>= 13.32 | 18.32 | 27.83-12.3<br>= 15.51 |

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

$$K_c = \frac{[\text{C}]^3}{[\text{A}][\text{B}]^2} = \frac{(15.51)^3}{(13.32)(18.32)^2} = 0.832$$

3) a) At a certain temperature,  $K_c$  is  $4.13 \times 10^{-2}$  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  $\text{I}_2$  (g) and  $\text{Br}_2$  (g) in equilibrium with 0.0124 moles/liter of IBr(g) ?

Setup:

|                 |                                    |                      |                          |   |                           |
|-----------------|------------------------------------|----------------------|--------------------------|---|---------------------------|
|                 | $2 \text{ IBr (g)}$                | $\rightleftharpoons$ | $\text{I}_2 \text{ (g)}$ | + | $\text{Br}_2 \text{ (g)}$ |
| initial conc.   | $M_{\text{initial}}$               |                      | 0                        |   | 0                         |
| change in conc. | $-2x$                              |                      | $+x$                     |   | $+x$                      |
| equi. conc.     | $M_{\text{initial}} - 2x = 0.0124$ |                      | $x$                      |   | $x$                       |

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

$$\sqrt{4.13 \times 10^{-2}} = \sqrt{\frac{(x)(x)}{(0.0124)^2}}$$

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

$$[\text{I}_2] = [\text{Br}_2] = x = 2.52 \times 10^{-3}$$

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

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

Setup:

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

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

$$= 0.0124 + 2(2.52 \times 10^{-3})$$

$$= 0.0174$$

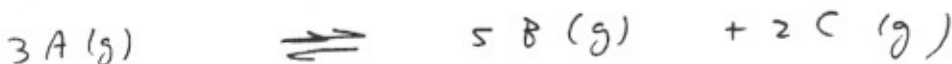
Ans: 0.0174

4) 0.924 mole of A (g) is placed in a 1.00 liter container at  $700^\circ \text{C}$ , 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:



initial conc. 0.924

in conc.  $-0.359$

$+0.598$   $+0.239$

$\frac{38.8}{100} \times 0.924 \text{ mole} = 0.359 \text{ moles "A" reacted}$

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

$0.359 \text{ mole A} \left( \frac{5 \text{ moles B}}{3 \text{ moles A}} \right) = 0.598 \text{ moles B formed}$

equilibrium conc.  $0.924 - 0.359 = 0.566$  react

0.598

0.239

Ans: 0.0241

$$K_c = \frac{[\text{B}]^5 [\text{C}]^2}{[\text{A}]^3} = \frac{(0.598)^5 (0.239)^2}{(0.566)^3} = 0.0241$$

5) The equilibrium constant for the reaction:  
 $2 \text{NO} (g) \rightleftharpoons \text{N}_2 (g) + \text{O}_2 (g)$   
 is  $2.60 \times 10^{-3}$  at  $1100^\circ \text{C}$ . If  $0.820$  mole of  $\text{NO} (g)$  and  $0.223$  mole each of  $\text{N}_2 (g)$  and  $\text{O}_2 (g)$  are mixed in a  $1.00$  liter container at  $1100^\circ \text{C}$ , what are the concentrations of  $\text{NO} (g)$ ,  $\text{N}_2 (g)$ , and  $\text{O}_2 (g)$  at equilibrium?  
 Setup:

$$Q = \frac{(\text{N}_2)(\text{O}_2)}{(\text{NO})^2} = \frac{(0.223)(0.223)}{(0.820)^2} = 7.40 \times 10^{-2}$$

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



|                |            |           |           |
|----------------|------------|-----------|-----------|
| Initial conc   | 0.820      | 0.223     | 0.223     |
| Change in conc | + 2x       | -x        | -x        |
| Equi. conc.    | 0.820 + 2x | 0.223 - x | 0.223 - x |

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

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

$$0.0510 = \frac{(0.223-x)}{(0.820+2x)}$$

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

$$[\text{O}_2] = [\text{N}_2] = 0.223 - x = 0.223 - 0.164 = 0.058 \text{ M}$$

$$[\text{NO}] = 0.820 + 2x = 0.820 + 2(0.164)$$

$$= 1.15 \text{ M}$$

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

**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,  $K_c$ , for the reaction:

Setup:

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

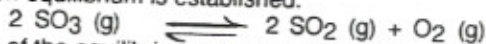
|                 |       |                           |                                       |
|-----------------|-------|---------------------------|---------------------------------------|
| Initial conc.   | 1.16  | 1.35                      | 0.641                                 |
| Change in conc. | +0.79 | -0.79                     | $-\frac{0.79}{2}$                     |
| Equi. conc.     | 1.95  | $(1.35 - 0.79)$<br>= 0.56 | $(0.641 - \frac{0.79}{2})$<br>= 0.246 |

# moles A =  $1.95 - 1.16$   
formed = 0.79

$$K_c = \frac{[B]^2 [C]}{[A]^2} = \frac{(0.56)^2 (0.246)}{(1.95)^2} = 0.020$$

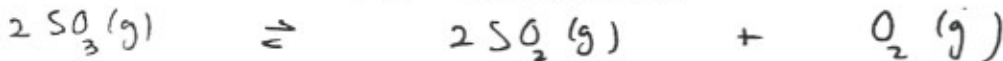
- 2) 0.822 mole of  $SO_3 (g)$  is placed in a 1.00 liter container at 600 K. 36.7% of the  $SO_3 (g)$  are decomposed when equilibrium is established.

Ans: 0.020



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

Setup:



|                 |                                       |              |                            |
|-----------------|---------------------------------------|--------------|----------------------------|
| Initial conc.   | 0.822                                 | 0            | 0                          |
| Change in conc. | -0.302                                | +0.302       | $+\frac{0.302}{2}$         |
| Equi. conc.     | $0.822 - 0.302$<br>initial<br>= 0.520 | 0.302 formed | $(\frac{0.302}{2})$ formed |

$$K = \frac{[SO_2]^2 [O_2]}{[SO_3]^2} = \frac{(0.302)^2 (0.151)}{(0.520)^2} = 5.05 \times 10^{-2}$$



3) For the equilibrium:



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

Setup:

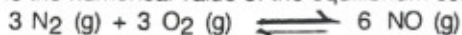
|                       |                     |                      |                                                  |   |                           |
|-----------------------|---------------------|----------------------|--------------------------------------------------|---|---------------------------|
|                       | $2 \text{BrCl (g)}$ | $\rightleftharpoons$ | $\text{Br}_2 \text{ (g)}$                        | + | $\text{Cl}_2 \text{ (g)}$ |
| <u>Initial conc.</u>  | 0                   |                      | $\frac{1.34 \text{ mol}}{11.0 \text{ L}} = .122$ |   | .122                      |
| <u>Change in conc</u> | +2x                 |                      | -x                                               |   | -x                        |
| <u>Equi conc.</u>     | 2x                  |                      | $0.122 - x$                                      |   | $0.122 - x$               |

$$K = \frac{[\text{Br}_2][\text{Cl}_2]}{[\text{BrCl}]^2} \quad ; \quad \sqrt{.143} = \sqrt{\frac{(.122-x)(.122-x)}{(2x)^2}}$$

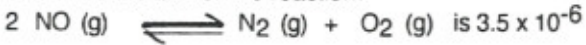
$$x = .0693$$

$[\text{Br}_2] = [\text{Cl}_2] = .122 - x = .122 - .0693 = .0529 \text{ M}$     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:



if the equilibrium constant for the reaction:

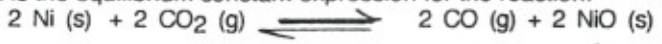


Setup:

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

Ans:  $2.3 \times 10^{16}$

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  $\text{NO}_2 \text{ (g)}$ , 0.122 mole of  $\text{O}_2 \text{ (g)}$  and 0.300 mole of  $\text{NO}_3 \text{ (g)}$  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:

|                             |   |                          |                      |                             |
|-----------------------------|---|--------------------------|----------------------|-----------------------------|
| $2 \text{NO}_2 \text{ (g)}$ | + | $\text{O}_2 \text{ (g)}$ | $\rightleftharpoons$ | $2 \text{NO}_3 \text{ (g)}$ |
| .0500                       |   | .122                     |                      | .300                        |

$$Q = \frac{[\text{NO}_3]^2}{[\text{NO}_2]^2 [\text{O}_2]} = \frac{(0.300)^2}{(.0500)^2 (.122)} = 295$$

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

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:



|                        |     |                          |                          |
|------------------------|-----|--------------------------|--------------------------|
| <u>Initial conc.</u>   | 0   | $\frac{1.45}{6.0} = .24$ | $\frac{1.45}{6.0} = .24$ |
| <u>Change in conc.</u> | +2x | -x                       | -x                       |
| <u>Equi conc</u>       | 2x  | $.242-x$                 | $0.242-x$                |

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

$$\sqrt{0.185} = \sqrt{\frac{(0.242-x)(0.242-x)}{(2x)^2}}$$

$$x = .130$$

$$[\text{N}_2] = [\text{O}_2] = .242 - x = .242 - .130 = .112 \text{ M}$$

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

Ans:

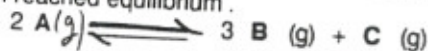
$$[\text{N}_2] = 0.112 \text{ M}$$

$$[\text{O}_2] = 0.112 \text{ M}$$

$$[\text{NO}] = 0.260 \text{ M}$$

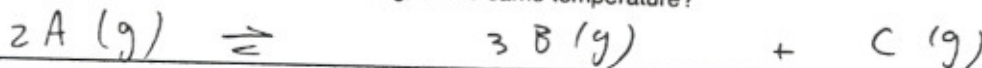
2) 1.87 mole of A (g) are placed in a 1.00 liter container at 700 K. when the reaction reached equilibrium.

43.3 % of A (g) had decomposed



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

Setup:



|                        |                      |       |        |
|------------------------|----------------------|-------|--------|
| <u>Initial conc.</u>   | 1.87                 | 0     | 0      |
| <u>Change in conc.</u> | -.810                | +1.21 | +0.405 |
| <u>Equi conc.</u>      | $1.87 - .810 = 1.06$ | 1.21  | 0.405  |

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

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

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

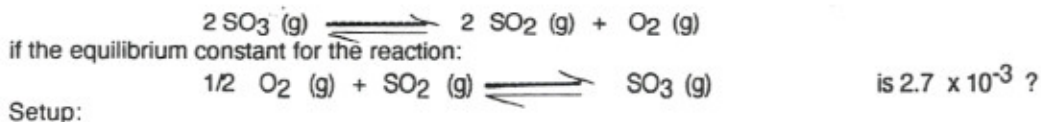
$$K = \frac{[\text{B}]^3 [\text{C}]}{[\text{A}]^2} = \frac{(1.21)^3 (0.405)}{1.06} = 0.639$$

Ans: 0.639

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

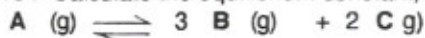
- b) What is the numerical value of the equilibrium constant,  $K_c$ , for the reaction:



$$K = \frac{1}{(2.7 \times 10^{-3})^2} = 1.4 \times 10^5$$

Ans:  $1.4 \times 10^5$

- 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:

|                      |      |        |      |
|----------------------|------|--------|------|
| <u>Initial conc.</u> | 3.31 | 4.33   | 5.95 |
| <u>Equi conc.</u>    | ?    | 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) | + 2C (g)           |
| <u>Initial conc.</u> | 3.31               |                      | 4.33   | 5.95               |
| change in conc.      | -0.61              |                      | +1.83  | +1.22              |
| <u>Equi conc.</u>    | 3.31 - 0.61 = 2.70 |                      | 6.16   | 5.95 + 1.22 = 7.17 |

$$K = \frac{[\text{B}]^3 [\text{C}]^2}{[\text{A}]} = \frac{(6.16)^3 (7.17)^2}{2.7} = 4.5 \times 10^3$$

Ans:  $4.5 \times 10^3$

5) For the reaction:

$2 A (g) \rightleftharpoons B (g) + C (g)$   
 at  $900^\circ C$ ,  $K_c$  is  $1.40 \times 10^{-3}$ . 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  $900^\circ C$ , what are the concentrations of A, B, and C at equilibrium?  
 Setup:

$$Q = \frac{(B)(C)}{(A)^2} = \frac{(0.244)(0.244)}{(0.780)^2} = 0.0978$$

$$Q = 0.0978 > K = 1.4 \times 10^{-3}$$

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



|                |            |           |           |
|----------------|------------|-----------|-----------|
| Initial conc   | 0.780      | 0.244     | 0.244     |
| Change in conc | +2x        | -x        | -x        |
| Equi conc.     | 0.780 + 2x | 0.244 - x | 0.244 - x |

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

$$\sqrt{1.40 \times 10^{-3}} = \sqrt{\frac{(0.244 - x)(0.244 - x)}{(0.780 + 2x)^2}}$$

$$3.74 \times 10^{-2} = \frac{(0.244 - x)}{(0.780 + 2x)}$$

$$x = 0.200$$

$$[B] = [C] = 0.244 - x = 0.244 - 0.200 = 0.044 M$$

$$[A] = 0.780 + 2x = 0.780 + 2(0.200) = 1.180 M$$

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