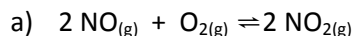


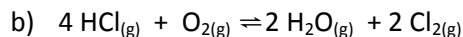
Equilibrium Worksheet **SOLUTIONS**

Complete the following questions on a separate piece of paper.

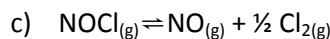
1. Write the equilibrium expression,  $K_{eq}$ , for each of the following reactions:



$$K_{eq} = \frac{[\text{NO}_2]^2}{[\text{NO}]^2[\text{O}_2]}$$



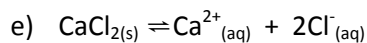
$$K_{eq} = \frac{[\text{H}_2\text{O}]^2[\text{Cl}_2]^2}{[\text{HCl}]^4[\text{O}_2]}$$



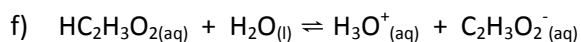
$$K_{eq} = \frac{[\text{NO}][\text{Cl}_2]^{\frac{1}{2}}}{[\text{NOCl}]}$$



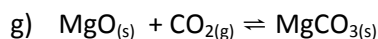
$$K_{eq} = \frac{[\text{FeSCN}^{2+}]}{[\text{Fe}^{3+}][\text{SCN}^{-}]}$$



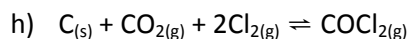
$$K_{eq} = [\text{Ca}^{2+}][\text{Cl}^{-}]^2$$



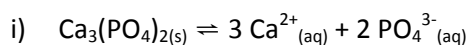
$$K_{eq} = \frac{[\text{H}_3\text{O}^{+}][\text{C}_2\text{H}_3\text{O}_2^{-}]}{[\text{HC}_2\text{H}_3\text{O}_2]}$$



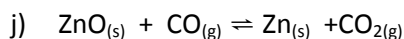
$$K_{eq} = \frac{1}{[\text{CO}_2]}$$



$$K_{eq} = \frac{[\text{COCl}_2]}{[\text{CO}_2][\text{Cl}_2]^2}$$



$$K_{eq} = [\text{Ca}^{2+}]^3[\text{PO}_4^{3-}]^2$$

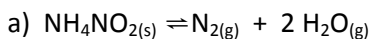


$$K_{eq} = \frac{[\text{CO}_2]}{[\text{CO}]}$$

2. The equilibrium constant for the equilibrium  $\text{CO}_{(g)} + \text{H}_2\text{O}_{(g)} \rightleftharpoons \text{CO}_{2(g)} + \text{H}_{2(g)}$  is 302 at 600K. What is the value of the equilibrium constant for the reverse reaction at the same temperature?

$$K_{eq \text{ reverse}} = \frac{1}{K_{eq}} = \frac{1}{302} = 3.31 \times 10^{-3}$$

3. Classify the following equilibria as heterogeneous or homogeneous, and write an equilibrium expression for each.

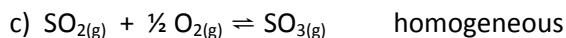


heterogeneous

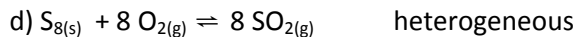
$$K_{eq} = [\text{N}_2][\text{H}_2\text{O}]^2$$



$$K_{eq} = [\text{H}_2\text{O}_{(g)}]$$



$$K_{eq} = \frac{[\text{SO}_3]}{[\text{SO}_2][\text{O}_2]^{1/2}}$$



$$K_{eq} = \frac{[\text{SO}_2]^8}{[\text{O}_2]^8}$$

4. At the equilibrium point in the decomposition of phosphorus pentachloride to chlorine and phosphorus trichloride, the following concentrations are obtained: 0.010 mol/L  $\text{PCl}_5$ , 0.15 mol/L  $\text{PCl}_3$  and 0.37 mol/L  $\text{Cl}_2$ . Determine the  $K_{eq}$  for the reaction.



$$C = 0.001 \text{ M} \quad 0.15 \text{ M} \quad 0.37 \text{ M}$$

$$K_{eq} = \frac{[\text{PCl}_3][\text{Cl}_2]}{[\text{PCl}_5]}$$

$$K_{eq} = \frac{(0.15)(0.37)}{(0.010)} \\ = 5.55$$

Therefore the  $K_{eq}$  value for the reaction is 5.55.

5. The colourless gas dinitrogen tetroxide decomposes to the brown coloured air pollutant nitrogen dioxide and exists in equilibrium. A 0.125 mol sample of dinitrogen tetroxide is introduced into a 1.00 L container and allowed to decompose at a given temperature. When equilibrium is reached, the concentration of the dinitrogen tetroxide is 0.0750 mol/L. What is the value of  $K_{eq}$  for this reaction?

	$\text{N}_2\text{O}_{4(g)} \rightleftharpoons 2 \text{NO}_{2(g)}$	
MR	1	2
I	0.125	0
C	-x	+2x
E	0.125-x =0.0750	2x

$$x = 0.050$$

$$K_{eq} = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]}$$

$$K_{eq} = \frac{(0.10)^2}{(0.075)} \\ = 0.13$$

Therefore the  $K_{eq}$  value for the reaction is 0.13

6. Phosphorus pentachloride decomposes to phosphorus trichloride and chlorine gas. A 1.10 mol/L sample of  $\text{PCl}_5$  was placed into a vessel. At equilibrium it was determined that the concentration was 0.33 mol/L. Calculate the  $K_{eq}$  for this reaction.

	$\text{PCl}_{5(g)} \rightleftharpoons \text{PCl}_{3(g)} + \text{Cl}_{2(g)}$		
MR	1	1	1
I	1.10	0	0
C	-x	+x	+x
E	1.10-x =0.33	x	x

$$x = 0.77$$

$$K_{eq} = \frac{[\text{PCl}_3][\text{Cl}_2]}{[\text{PCl}_5]}$$

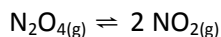
$$K_{eq} = \frac{(0.77)(0.77)}{(0.33)} \\ = 1.80$$

Therefore the  $K_{eq}$  value for the reaction is 1.80

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7. Gaseous dinitrogen tetroxide is placed in a flask and allowed to decompose to nitrogen dioxide and reach equilibrium at 100°C. At 100°C, the value of  $K_{eq}$  is 0.212. If the concentration of dinitrogen tetroxide at equilibrium is 0.155 mol/L, what is the concentration of nitrogen dioxide at equilibrium?



$$K_{eq} = \frac{[NO_2]^2}{[N_2O_4]}$$

$$0.212 = \frac{[NO_2]^2}{(0.155)}$$

$$[NO_2] = 0.181M$$

8. A 0.921 mol sample of dinitrogen tetroxide is placed in a 1.00 L vessel and heated to 100°C. At equilibrium it is found that 20.7 % of the dinitrogen tetroxide has decomposed to nitrogen dioxide. Calculate the  $K_{eq}$  for this reaction.

$N_2O_{4(g)} \rightleftharpoons 2 NO_{2(g)}$		
MR	1	2
I	0.921	0
C	-x	+2x
E	0.921-x	2x

$$K_{eq} = \frac{[NO_2]^2}{[N_2O_4]}$$

$$K_{eq} = \frac{(0.382)^2}{(0.730)}$$

$$K_{eq} = 0.200$$

At equilibrium, 20.7% of  $N_2O_4$  is gone

$x = 20.7\%$  of 0.921

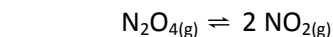
$x = (0.207)(0.921)$

$x = 0.191$

$[N_2O_4] = 0.730 M$   $[NO_2] = 0.382 M$

Therefore the equilibrium constant is 0.200.

9. At 245°C, the equilibrium concentration of dinitrogen tetroxide gas is  $6.38 \times 10^{-3}$  mol/L and the total gas concentration is  $1.23 \times 10^{-2}$  mol/L. Determine the  $K_{eq}$  for the decomposition of dinitrogen tetroxide gas to nitrogen dioxide gas at this temperature.



$$[N_2O_4] = 6.38 \times 10^{-3} M$$

$$[gas] = 1.23 \times 10^{-2} M$$

$$[NO_2] = 1.23 \times 10^{-2} M - 6.38 \times 10^{-3} M$$

$$= 5.92 \times 10^{-3} M$$

$$K_{eq} = \frac{[NO_2]^2}{[N_2O_4]}$$

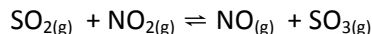
$$K_{eq} = \frac{(5.92 \times 10^{-3})^2}{(6.38 \times 10^{-3})}$$

$$K_{eq} = 5.49 \times 10^{-3}$$

Therefore the equilibrium constant is  $5.49 \times 10^{-3}$ .

Name: \_\_\_\_\_

10. The following reaction has  $K_{eq}$  value of 85.0 at 460°C:



If a mixture of sulfur dioxide and nitrogen dioxide is prepared, each with an initial concentration of 0.100 mol/L, calculate the equilibrium concentrations of nitrogen dioxide and nitrogen monoxide at this temperature.

	$SO_{2(g)}$	$NO_{2(g)}$	$NO_{(g)}$	$SO_{3(g)}$	$K_{eq} = 85.0$
MR	1	1	1	1	
I	0.100	0.100	0	0	
C	-x	-x	+x	+x	
E	0.100-x	0.100-x	x	x	

$0.100/K_{eq} < 1000$   
approx will not  
work

$$K_{eq} = \frac{[NO][SO_3]}{[SO_2][NO_2]}$$

$$85 = \frac{(x)(x)}{(0.100-x)(0.100-x)}$$

$$9.22 = \frac{x}{(0.100-x)}$$

$$0.922 = 10.22x$$

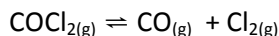
$$x = 0.0902$$

$$[NO_2] = 9.8 \times 10^{-3} \text{ M}$$

$$[NO] = 0.0902 \text{ M}$$

Therefore the concentration of nitrogen dioxide is  $9.8 \times 10^{-3} \text{ M}$  and nitrogen monoxide is 0.0902 M.

11. At 100°C the reaction below has an equilibrium constant,  $K_{eq}$ , value of  $2.2 \times 10^{-10}$ . If 1.00 mol of phosgene,  $COCl_2$ , is placed in a 10.0 L flask, calculate the concentration of carbon monoxide at equilibrium.



	$COCl_{2(g)}$	$CO_{(g)}$	$Cl_{2(g)}$	$K_{eq} = 2.2 \times 10^{-10}$
MR	1	1	1	
I	0.100	0	0	
C	-x	+x	+x	
E	0.100-x	x	x	

$0.100/K_{eq} > 1000$   
approx will work

$$K_{eq} = \frac{[CO][Cl_2]}{[COCl_2]}$$

$$2.2 \times 10^{-10} = \frac{(x)(x)}{0.100-x}$$

$$2.2 \times 10^{-10} = \frac{x^2}{0.100}$$

$$x = 4.7 \times 10^{-6}$$

Therefore  $[CO] = 4.7 \times 10^{-6} \text{ M}$

12. Six moles of  $SO_{2(g)}$  and four moles of  $O_{2(g)}$  are introduced into a 1.00 L reaction vessel and allowed to react to form  $SO_{3(g)}$ . At equilibrium, the vessel contains four moles of  $SO_{3(g)}$ . Calculate  $K_{eq}$  for this reaction.

	$2 SO_{2(g)}$	$O_{2(g)}$	$2 SO_{3(g)}$
MR	2	1	2
I	6.00	4.00	0
C	-2x	-x	+2x
E	6.00-2x	4.00-x	2x=4.00

At equilibrium  $[SO_3] = 4.00 \text{ M}$

Therefore  $x = 2.00 \text{ M}$

$[SO_2] = 2.00 \text{ M}$

$$[O_2] = 2.00 \text{ M}$$

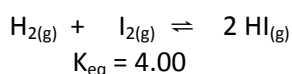
$$K_{eq} = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$$

$$K_{eq} = \frac{(4.00)^2}{(2.00)^2(2.00)}$$

$$K_{eq} = 2.00$$

Name: \_\_\_\_\_

13. Hydrogen and iodine gases react to form hydrogen iodide gas. If 6.00 mol of  $H_2$  and 3.00 mol of  $I_2$  are placed in a 3.00 L vessel and allowed to come to equilibrium at  $250^\circ C$  calculate the equilibrium concentrations of all species. The  $K_{eq}$  for the reaction is 4.00 at  $250^\circ C$ .



MR	1	1	2
I	2.00	1.00	0
C	-x	-x	+2x
E	2.00-x	1.00-x	2x

$1.00/K_{eq} < 1000$   
approx will not work

$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]}$$

$$4.00 = \frac{(2x)^2}{(2.00-x)(1.00-x)}$$

$$4.00(2.00 - 3.00x + x^2) = 4x^2$$

$$8.00 - 12.00x + 4x^2 = 4x^2$$

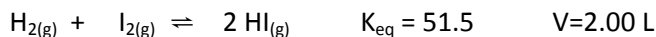
$$8.00 = 12.00x$$

$$x = 0.667M$$

Therefore  $[H_2] = 1.33 M$ ,  $[I_2] = 0.330 M$  and  $[HI] = 1.33 M$ .

14. At  $375^\circ C$ , the equilibrium constant for the reaction between hydrogen and iodine gases to make hydrogen iodide gas is 51.5. A sample of hydrogen iodide was placed into a 2.00 L vessel and it was found that at equilibrium 0.218 mol of  $H_2$  gas was present.

- a) How many moles of HI were originally placed into the flask?



MR	1	1	2
I	0	0	y
C	+x	+x	-2x
E	x	x	y-2x

At equilibrium  $[H_2] = 0.109 M = x$

Therefore at equilibrium  $[I_2] = 0.109 M$  and  $[HI] = y - 0.218$

$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]}$$

$$51.5 = \frac{(y - 0.218)^2}{(0.109)(0.109)}$$

$$7.18 = \frac{y - 0.218}{0.109}$$

$$0.783 = y - 0.218$$

$$y = 1.00$$

$$n_{HI} = (2.00L)(1.00 M)$$

$$= 2.00 \text{ mol}$$

Therefore 2.00 moles of hydrogen iodide were initially added.

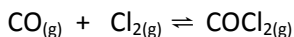
- b) What are the equilibrium concentrations of  $I_2$  and HI?

$$[I_2] = x \quad [HI] = 1.00 - 0.218$$

$$[I_2] = 0.109 M \quad [HI] = 0.782 M$$

Name: \_\_\_\_\_

15. Carbon monoxide and chlorine gases react to form phosgene gas ( $\text{COCl}_2$ ). At equilibrium, there is 0.62 mol of CO, 0.28 mol of  $\text{Cl}_2$  and 1.56 mol of  $\text{COCl}_2$  in a 2.00 L container. How many moles of  $\text{Cl}_2$  must be added to reduce the concentration of CO to 0.25 mol/L? [0.22 mol]



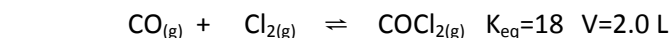
A) Use the equilibrium concentrations given to calculate the  $K_{eq}$  value.

$$K_{eq} = \frac{[\text{COCl}_2]}{[\text{CO}][\text{Cl}_2]}$$

$$K_{eq} = \frac{(0.78)}{(0.31)(0.14)}$$

$$K_{eq} = 18$$

B) A given amount of chlorine gas in mol/L "y" is added to the reaction and equilibrium is re-established at the SAME temperature.



MR	1	1	1
I	0.31	0.14 + y	0.78
C	-x	-x	+x
E	0.31-x	0.14+y-x	0.78+x

$0.31/K_{eq} < 1000$   
approx will not  
work

$$K_{eq} = \frac{[\text{COCl}_2]}{[\text{CO}][\text{Cl}_2]}$$

$$18 = \frac{(0.84)}{(0.25)(0.08 + y)}$$

$$0.36 + 4.5y = 0.84$$

$$4.5y = 0.48$$

$$y = 0.11$$

$$[\text{Cl}_2] \text{ added} = 0.11 \text{ mol/L}$$

At equilibrium  $[\text{CO}] = 0.25 \text{ M}$

$$0.31 - x = 0.25$$

$$x = 0.06 \text{ M}$$

$$\text{Therefore } [\text{Cl}_2] = 0.14 + y - 0.06$$

$$[\text{COCl}_2] = 0.78 + 0.06$$

$$= 0.08 + y$$

$$= 0.84$$

Therefore the moles of  $\text{Cl}_2$  added is 0.22 mol.

16. At  $350^\circ\text{C}$  the equilibrium constant for the reaction  $\text{CaCO}_{3(s)} \rightleftharpoons \text{CaO}_{(s)} + \text{CO}_{2(g)}$  was found to be 0.15.

a) What is the equilibrium concentration of carbon dioxide at this temperature?



$$K_{eq} = [\text{CO}_2]$$

$$[\text{CO}_2] = 0.15 \text{ M}$$

b) If 100.0 g of solid calcium carbonate was placed in a 10.0 L vessel and heated to the same temperature could equilibrium be reached? Justify your answer with calculations.

Equilibrium will be reached with these conditions.

100.0 g of  $\text{CaCO}_3$  in 10.0 L

moles of  $\text{CaCO}_3 = 1.00 \text{ mol}$

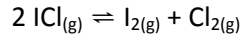
Since 1 mol of  $\text{CaCO}_3$  produces 1 mol of  $\text{CO}_2$ , the maximum amount of carbon dioxide that can be produced is 1.00 mol giving a  $[\text{CO}_2]$  of 0.10 M.

Since the concentration of  $\text{CO}_2$  when equilibrium is reached at this temperature is 0.15 M and the maximum  $[\text{CO}_2]$  in this case is 0.10 M equilibrium will not be established.

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17. The equilibrium constant for the reaction below is 0.11. Calculate all equilibrium concentrations if 0.33 mol of iodine chloride gas is placed in a 1.00 L vessel and allowed to come to equilibrium.



	$2 \text{ICl}_{(g)} \rightleftharpoons \text{I}_{2(g)} + \text{Cl}_{2(g)} \quad K_{eq}=0.11$		
MR	2	1	1
I	0.33	0	0
C	-2x	+x	+x
E	0.33-2x	x	x

$0.33/K_{eq} < 1000$   
approx will not  
work

$$K_{eq} = \frac{[\text{I}_2][\text{Cl}_2]}{[\text{ICl}_2]^2}$$

$$0.11 = \frac{(x)(x)}{(0.33-2x)^2}$$

$$0.33 = \frac{x}{0.33-2x}$$

$$0.11 - 0.66x = x$$

$$0.11 = 1.66x$$

$$x = 0.066$$

Therefore  $[\text{ICl}] = 0.20 \text{ M}$ ,  $[\text{I}_2] = 0.066 \text{ M}$  and  $[\text{Cl}_2] = 0.066 \text{ M}$ .

18. At 1200°C, the value of  $K_{eq}$  for the formation of hydrogen chloride gas from its elements is  $2.51 \times 10^4$ . Determine the equilibrium concentrations of all species if 0.250 mol of both chlorine gas and hydrogen gas are placed in a 1.00 L vessel and allowed to reach equilibrium.

	$\text{H}_{2(g)} + \text{Cl}_{2(g)} \rightleftharpoons 2 \text{HCl}_{(g)} \quad K_{eq}=2.51 \times 10^4$		
MR	1	1	2
I	0.250	0.250	0
C	-x	-x	+2x
E	0.250-x	0.250-x	2x

$0.250/K_{eq} < 1000$   
approx will not  
work

$$K_{eq} = \frac{[\text{HCl}]^2}{[\text{H}_2][\text{Cl}_2]}$$

$$2.51 \times 10^4 = \frac{(2x)^2}{(0.250-x)(0.250-x)}$$

$$158 = \frac{2x}{0.250-x}$$

$$39.5 - 158x = 2x$$

$$39.5 = 160x$$

$$x = 0.247$$

Therefore  $[\text{H}_2] = 3.00 \times 10^{-3} \text{ M}$ ,  $[\text{Cl}_2] = 3.00 \times 10^{-3} \text{ M}$  and  $[\text{HCl}] = 0.494 \text{ M}$

19. Nitrogen gas reacts with oxygen gas to form nitrogen monoxide gas. The  $K_{eq}$  for this reaction at 1530°C is  $1.20 \times 10^{-4}$ . Determine the equilibrium concentrations of all species when 1.00 mol/L of nitrogen and oxygen is placed in a vessel and allowed to reach equilibrium.

	$\text{N}_{2(g)} + \text{O}_{2(g)} \rightleftharpoons 2 \text{NO}_{(g)} \quad K_{eq} = 1.20 \times 10^{-4}$		
MR	1	1	2
I	1.0	1.0	0
C	-x	-x	+2x
E	1.0-x	1.0-x	2x

$1.0/K_{eq} > 1000$   
approx will work

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

$$1.20 \times 10^{-4} = \frac{(2x)^2}{(1.0-x)(1.0-x)}$$

$$1.20 \times 10^{-4} = 4x^2$$

$$x = 5.48 \times 10^{-3}$$

Therefore  $[\text{N}_2] = 0.995 \text{ M}$ ,  $[\text{O}_2] = 0.995 \text{ M}$  and  $[\text{NO}] = 0.0110 \text{ M}$ .

20. At 2000°C, the  $K_{eq}$  for the decomposition of carbon dioxide gas to carbon monoxide and oxygen gases is  $6.40 \times 10^{-7}$ . Determine the oxygen concentration at equilibrium when 1.00 mol of  $CO_2$  is placed in a 1.00 L vessel.

$$2 CO_{2(g)} \rightleftharpoons 2 CO_{(g)} + O_{2(g)}$$

MR	2	2	1
I	1.00	0	0
C	-2x	+2x	+x
E	1.00-2x	2x	x

$1.00/K_{eq} > 1000$   
approx will work

$$K_{eq} = \frac{[CO]^2[O_2]}{[CO_2]^2}$$

$$6.40 \times 10^{-7} = \frac{(2x)^2(x)}{(1.00-2x)^2}$$

$$6.40 \times 10^{-7} = 4x^3$$

$$x = 5.43 \times 10^{-3}$$

Therefore the  $[O_2] = 5.43 \times 10^{-3} M$ .

21. The dissociation of ammonia gas to nitrogen and hydrogen gases has a  $K_{eq}$  value of  $2.63 \times 10^{-9}$  at 27°C. If 2.00 mol of ammonia is placed in a 2.00 L vessel and allowed to reach equilibrium, what is the concentration of hydrogen and nitrogen.

$$2 NH_{3(g)} \rightleftharpoons N_{2(g)} + 3 H_{2(g)} \quad K_{eq} = 2.63 \times 10^{-9}$$

MR	2	1	3
I	1.00	0	0
C	-2x	+x	+3x
E	1.00-2x	x	3x

$1.00/K_{eq} > 1000$   
approx will work

$$K_{eq} = \frac{[N_2][H_2]^3}{[NH_3]^2}$$

$$2.63 \times 10^{-9} = \frac{(x)(3x)^3}{(1.00-2x)^2}$$

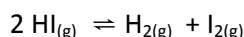
$$2.63 \times 10^{-9} = 27x^4$$

$$9.74 \times 10^{-11} = x^4$$

$$x = 3.14 \times 10^{-3}$$

Therefore  $[N_2] = 3.14 \times 10^{-3} M$  and  $[H_2] = 9.42 \times 10^{-3} M$ .

22. The following reaction takes place in a 1.00L vessel at 500°C.



Equilibrium concentrations were found to be 1.76 mol/L HI, 0.200 mol/L  $H_2$  and 0.200 mol/L  $I_2$ . If an additional 0.500 mol of hydrogen iodide gas is introduced at the same temperature, what the new concentrations of all gases once equilibrium has been reestablished?

- A) Use the given concentrations to calculate  $K_{eq}$

$$K_{eq} = \frac{[H_2][I_2]}{[HI]^2}$$

$$K_{eq} = \frac{(0.20)(0.20)}{(1.76)^2}$$

$$K_{eq} = 1.30 \times 10^{-2}$$

- B) After the addition of an 0.500 mol of  $HI_{(g)}$ .

$$2 HI_{(g)} \rightleftharpoons H_{2(g)} + I_{2(g)} \quad K_{eq} = 1.30 \times 10^{-2}$$

MR	2	1	1
I	2.26	0.200	0.200
C	-2x	+x	+x
E	2.26-2x	0.200+x	0.200+x

$0.200/K_{eq} < 1000$   
approx will not work

Therefore  $[HI] = 2.18 M$ ,  $[H_2] = 0.241 M$  and  $[I_2] = 0.241 M$ .

$$K_{eq} = \frac{[H_2][I_2]}{[HI]^2}$$

$$1.29 \times 10^{-2} = \frac{(0.200+x)(0.200+x)}{(2.26-2x)^2}$$

$$0.114 = \frac{0.200+x}{2.26-2x}$$

$$0.258 - 0.228x = 0.200 + x$$

$$0.0580 = 1.228x$$

$$x = 0.0472$$



Name: \_\_\_\_\_

23. The dissociation of ammonia to nitrogen and hydrogen gases at 400.0°C has a  $K_{eq}$  value of 1.92. If 0.500 mol of ammonia is placed in a 500.0 mL container, determine the equilibrium concentrations of all gases.

$$2 \text{NH}_{3(g)} \rightleftharpoons \text{N}_{2(g)} + 3 \text{H}_{2(g)} \quad K_{eq}=1.92$$

MR	2	1	3
I	1.00	0	0
C	-2x	+x	+3x
E	1.00-2x	x	3x

$1.00/K_{eq} < 1000$   
approx will not work

If  $x = -0.849$  the concentrations of  $\text{N}_2$  and  $\text{H}_2$  will be negative which is not possible.

$$K_{eq} = \frac{[\text{N}_2][\text{H}_2]^3}{[\text{NH}_3]^2}$$

$$1.92 = \frac{(x)(3x)^3}{(1.00 - 2x)^2}$$

$$1.92 = \frac{27x^4}{(1.00 - 2x)^2}$$

$$1.39 = \frac{5.20}{1.00 - 2x}$$

$$1.39 - 2.78x - 5.20x^2 = 0$$

$$a = -5.20, b = -2.78, c = 1.39$$

$$x_1 = -0.849, x_2 = 0.315$$

$$\therefore x = 0.315$$

Therefore  $[\text{NH}_3] = 0.37 \text{ M}$ ,  $[\text{N}_2] = 0.315 \text{ M}$  and  $[\text{H}_2] = 0.945 \text{ M}$ .

24. The equilibrium constant,  $K_{eq}$  for the reaction below is 85.0 at 460°C. If a mixture is prepared where the initial concentration of sulfur dioxide is 1.00 mol/L and nitrogen dioxide is 2.00 mol/L calculate the equilibrium concentration of nitrogen monoxide and nitrogen dioxide.

$$\text{SO}_{2(g)} + \text{NO}_{2(g)} \rightleftharpoons \text{NO}_{(g)} + \text{SO}_{3(g)} \quad K_{eq}=85.0$$

MR	1	1	1	1
I	1.00	2.00	0	0
C	-x	-x	+x	+x
E	1.00-x	2.00-x	x	x

$1.00/K_{eq} < 1000$   
approx will not work

$$K_{eq} = \frac{[\text{NO}][\text{SO}_3]}{[\text{SO}_2][\text{NO}_2]}$$

$$85.0 = \frac{(x)(x)}{(1.00 - x)(2.00 - x)}$$

$$85.0(2.00 - 3.00x + x^2) = x^2$$

$$170 - 255x + 84.0x^2 = 0$$

$$a = 84.0, b = -255, c = 170$$

$$x_1 = 2.05, x_2 = 0.989$$

$$\therefore x = 0.989$$

Therefore  $[\text{NO}] = 0.989 \text{ M}$  and  $[\text{NO}_2] = 1.01 \text{ M}$ .

25. Hydrogen reacts with iodine vapour to produce hydrogen iodide vapour. The value of  $K_{eq}$  is 49.0. Calculate the concentration of all species at equilibrium, if there was 2.00 mol of hydrogen and 0.500 mol of iodine in a one vessel initially.

$$\text{H}_{2(g)} + \text{I}_{2(g)} \rightleftharpoons 2 \text{HI}_{(g)} \quad K_{eq} = 49.0$$

MR	1	1	2
I	2.00	0.500	0
C	-x	-x	+2x
E	2.00-x	0.500-x	2x

$2.00/K_{eq} < 1000$   
approx will not work

$$K_{eq} = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$$

$$49.0 = \frac{(2x)^2}{(2.00 - x)(0.500 - x)}$$

$$49.0(1.00 - 2.50x + x^2) = 4x^2$$

$$49.0 - 122.5x + 45x^2 = 0$$

$$a = 45, b = -122.5, c = 49$$

$$x_1 = 2.24, x_2 = 0.487$$

$$\therefore x = 0.487$$

Therefore  $[\text{H}_2] = 1.51 \text{ M}$ ,  $[\text{I}_2] = 0.0130 \text{ M}$  and  $[\text{HI}] = 0.974 \text{ M}$ .

26. In a 1 L vessel, 3.0 g of phosphorus pentachloride is heated. Only, 30 % of the phosphorus pentachloride dissociates into phosphorus trichloride and chlorine gas. What is the density of the equilibrium mixture?
- 3.0 g/L
  - The Law of Conservation of Mass states that in a closed system mass is not lost during a chemical reaction therefore the density of the equilibrium **MIXTURE** will be the same as the density of the starting materials. ☺
27. Phosphorus pentachloride decomposes into phosphorus trichloride and chlorine gas. If 13.9 % of a 1.0 mol/L concentration of phosphorus pentachloride decomposes, find the value of  $K_{eq}$ .

$$PCl_{5(g)} \rightleftharpoons PCl_{3(g)} + Cl_{2(g)}$$

MR	1	1	1
I	1.0	0	0
C	-x	+x	+x
E	1.0-x	x	x

At equilibrium, 13.9% of  $PCl_5$  is decomposed.

$$\therefore x = 13.9\% \text{ of } 1.00$$

$$x = 0.139$$

$$\therefore [PCl_5] = 0.861M$$

$$[PCl_3] = 0.139M$$

$$[Cl_2] = 0.139M$$

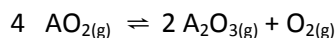
$$K_{eq} = \frac{[PCl_3][Cl_2]}{[PCl_5]}$$

$$K_{eq} = \frac{(0.139)(0.139)}{(0.861)}$$

$$K_{eq} = 2.24 \times 10^{-2}$$

Therefore the  $K_{eq}$  value is  $2.24 \times 10^{-2}$ .

28. The substance  $AO_2$  is 10.0 % molar dissociated according to the following reaction.



Find the equilibrium concentration of each species if 2.0 mol of  $A_2O_3$  and 1.0 mol of  $O_2$  are initially present in a 1.0 L volume.

$$4 AO_{2(g)} \rightleftharpoons 2 A_2O_{3(g)} + O_{2(g)}$$

MR	4	2	1
I	0	2.0	1.0
C	+4x	-2x	-x
E	4x	2.0-2x	1.0-x

At equilibrium,  $AO_2$  is 10% dissociated.

$$\therefore [A_2O_3]_{Equil} = 10\% \text{ of } 2.0$$

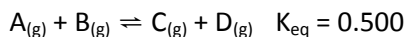
$$[A_2O_3]_{Equil} = 0.20$$

$$\therefore 0.20 = 2.0 - x$$

$$x = 0.90$$

Therefore at equilibrium  $[AO_2]=3.6 M$ ,  $[A_2O_3]=0.20 M$  and  $[O_2]=0.10 M$ .

29. If 0.70 mol of A, 0.50 mol of B, 0.40 mol of C and 0.90 mol of D were placed in 1.0 L container and allowed to come to equilibrium according to the following equation.



Calculate the equilibrium concentrations of all species.

	$A_{(g)}$	$+$	$B_{(g)}$	$\rightleftharpoons$	$C_{(g)}$	$+$	$D_{(g)}$	$K_{eq} = 0.500$
MR	1		1		1		1	
I	0.700		0.500		0.400		0.900	
C	-x		-x		+x		+x	
E	0.700-x		0.500-x		0.400+x		0.900+x	

$0.400/K_{eq} < 1000$   
approx will not work

$$K_{eq} = \frac{[C][D]}{[A][B]}$$

$$0.500 = \frac{(0.400 + x)(0.900 + x)}{(0.700 - x)(0.500 - x)}$$

$$0.50 = \frac{(0.360 + 1.30x + x^2)}{(0.350 - 1.20x + x^2)}$$

$$0.500(0.350 - 1.20x + x^2) = (0.360 + 1.30x + x^2)$$

$$0.185 + 1.90x + 0.500x^2 = 0$$

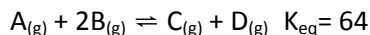
$$a = 0.500, b = 1.9, c = 0.185$$

$$x_1 = -0.100, x_2 = -3.70$$

$$\therefore x = -0.100$$

Therefore at equilibrium  $[A]=0.80$  M,  $[B]=0.60$  M,  $[C]=0.30$  M and  $[D]=0.80$  M.

30. An equimolar concentration of A and B plus 2.00 mol of both C and D were placed in a 1.00 L vessel.



Given that the equilibrium concentration of A is 1.00 M, calculate:

- a) The equilibrium concentrations of all species.

	$A_{(g)}$	$+$	$2B_{(g)}$	$\rightleftharpoons$	$C_{(g)}$	$+$	$D_{(g)}$	$K_{eq} = 64$
MR	1		2		1		1	
I	y		y		2.00		2.00	
C	-x		-2x		+x		+x	
E	y-x		y-2x		2.00+x		2.00+x	

At equilibrium  $[A] = 1.00$

$$\therefore y - x = 1.00$$

$$y = 1.00 + x$$

$$[B] = y - 2x$$

$$[B] = (1.00 + x) - 2x$$

$$[B] = 1.00 - x$$

Therefore at equilibrium  $[A]=1.00$  M,  $[B]=0.334$  M,  $[C]=2.67$  M and  $[D]=2.67$  M.

- b) The initial concentrations of A and B.

$$y = 1.00 + x$$

$$y = 1.00 + 0.667$$

$$y = 1.67$$

Therefore the initial concentrations of A and B is 1.67 M.

$$K_{eq} = \frac{[C][D]}{[A][B]^2}$$

$$64 = \frac{(2.00 + x)(2.00 + x)}{(1.00)(1.00 - x)^2}$$

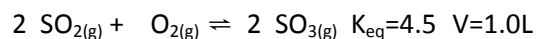
$$8 = \frac{2.00 + x}{1.00 - x}$$

$$8 - 8x = 2.00 + x$$

$$6 = 9x$$

$$x = 0.667$$

31. The  $K_{eq}$  value for the formation of sulphur trioxide gas from sulphur dioxide and oxygen gas is 4.5. If the equilibrium concentration of oxygen was 2.0 M, determine the number of moles of sulfur trioxide originally present in a 1.0 L vessel. [16 mol]



MR	2	1	2
I	0	0	y
C	+2x	+x	-2x
E	2x	x	y-2x

At equilibrium  $[\text{O}_2]=2.0 \text{ M}$

Therefore  $x=2.0 \text{ M}$

$$K_{eq} = \frac{[\text{SO}_3]^2}{[\text{SO}_2][\text{O}_2]}$$

$$4.5 = \frac{(y - 4.0)^2}{(4)^2 (2)}$$

$$(4.5)(32) = (y - 4.0)^2$$

$$144 = (y - 4.0)^2$$

$$12 = y - 4$$

$$y = 16$$

Therefore 16 moles of sulphur trioxide were originally present in the vessel.

If you need more you can complete the following questions from your textbook: 438 # 8, 9, page 472 # 5, 6, page 480 # 9, 10, page 481 # 1-8.