

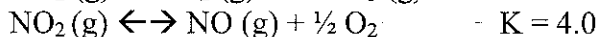
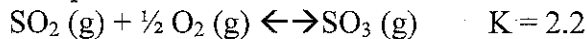
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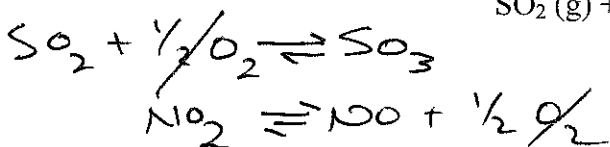
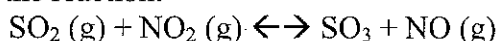
Swimming - Equilibrium Practice Problems

Adding K_{eq}

1. The following two reactions have equilibrium constants listed:



What is the equilibrium constant for the reaction:

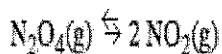


$$K_{eq} = (K_1)(K_2) = (2.2)(4.0) = \boxed{8.8}$$

2.

Calculate the value of K_c for the reaction: $2 \text{N}_2\text{O}(\text{g}) + 3 \text{O}_2(\text{g}) \rightleftharpoons 2 \text{N}_2\text{O}_4(\text{g})$, using the following information.

Equation

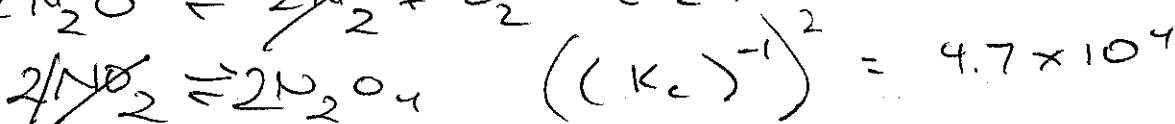
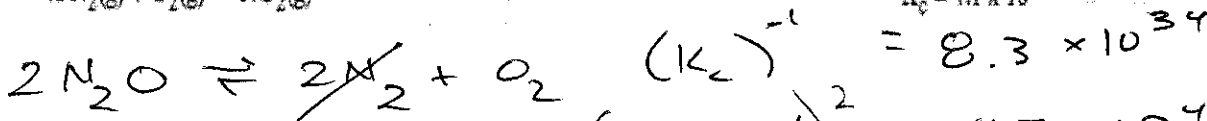


Equilibrium Constant

$$K_c = 1.2 \times 10^{-35}$$

$$K_c = 4.6 \times 10^{-3}$$

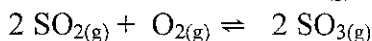
$$K_c = 4.1 \times 10^{-9}$$



$$\boxed{1.1 \times 10^6}$$

ICE Tables

3. Six moles of $\text{SO}_2(\text{g})$ and four moles of $\text{O}_2(\text{g})$ are introduced into a 1.00 L reaction vessel and allowed to react to form $\text{SO}_3(\text{g})$. At equilibrium, the vessel contains four moles of $\text{SO}_3(\text{g})$. Calculate K_{eq} for this reaction



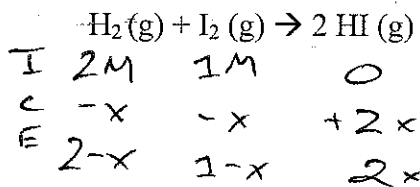
I	6M	4M	0	
C	-2x	-x	+2x	$\frac{4}{2} = \frac{2x}{2}$
E	6-2x	4-x	4M	$2 = x$
	↓	↓	↓	
	2	2	4	

$$K = \frac{(4)^2}{(2)^2(2)} = \frac{16}{8} = \boxed{2}$$

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4. 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 $250^\circ C$.



$$8 - 12x + 4x^2 = 4x^2$$

$$8 - 12x = 0$$

$$8 = 12x$$

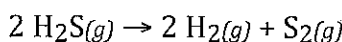
$$\frac{8}{12} = \frac{x}{1}$$

$$4 = \frac{(2x)^2}{(2-x)(1-x)} \Rightarrow 4 = \frac{4x^2}{2-3x+x^2}$$

$$4 = \frac{4x^2}{2-3x+x^2}$$

$$0.67 = x \quad | \quad [HI] = 1.34 M$$

5.



$$[H_2] = 1.33 M \quad | \quad [I_2] = 0.33 M$$

When heated, hydrogen sulfide gas decomposes according to the equation above. A 3.40 g sample of $H_2S(g)$ is introduced into an evacuated rigid 1.25 L container. The sealed container is heated to 483 K, and 3.72×10^{-2} mol of $S_2(g)$ is present at equilibrium.

(a) Write the expression for the equilibrium constant, K_c , for the decomposition reaction represented above.

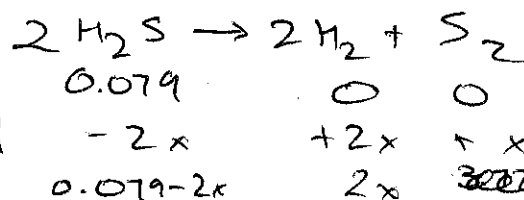
$$K_c = \frac{[S_2][H_2]^2}{[H_2S]^2}$$

(b) Calculate the equilibrium concentration, in $mol L^{-1}$, of the following gases in the container at 483 K.

(i) $H_2(g)$

$$3.40 g \left(\frac{1 mol}{34.076} \right) = 0.0998 mol$$

$$\frac{0.0998 mol}{1.25 L} = 0.079 M$$



(ii) $H_2S(g)$

$$[H_2] = 2 \left(\frac{2.98 \times 10^{-2}}{2.98} \right) = \frac{5.96}{5.96} \times 10^{-2} M$$

$$[H_2S] = 0.079 - 0.0372 = 4.92 \times 10^{-2} M$$

(c) Calculate the value of the equilibrium constant, K_c , for the decomposition reaction at 483 K.

$$K_c = \frac{(2.98 \times 10^{-2})(5.96 \times 10^{-2})^2}{(4.92 \times 10^{-2})^2} = \frac{1.06 \times 10^{-4}}{2.42 \times 10^{-3}} = 4.37 \times 10^{-2}$$

(d) Calculate the partial pressure of $S_2(g)$ in the container at equilibrium at 483 K.

$$P = \frac{nRT}{V} = \frac{(3.72 \times 10^{-2})(0.0821)(483)}{1.25 L} = 1.18 atm$$

(e) For the reaction $H_2(g) + \frac{1}{2} S_2(g) \rightarrow H_2S(g)$ at 483 K, calculate the value of the equilibrium constant, K_c .

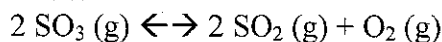
$$\left((K_c)^{-1} \right)^{1/2} = \sqrt{4.78}$$

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Reaction Quotients and Le Chatelier's

6. At 1000 K the value of K_p for the reaction:



is 0.338. Calculate the value for Q_p and predict the direction in which the reaction proceeds toward equilibrium if the initial partial pressures are $P_{\text{SO}_3} = 0.16 \text{ atm}$; $P_{\text{SO}_2} = 0.41 \text{ atm}$; $P_{\text{O}_2} = 2.5 \text{ atm}$.

$$Q = \frac{(0.41)^2(2.5)}{(0.16)^2} = \frac{0.42025}{0.0256} = 16.4$$

$Q > K \therefore$ Shift left

7. At 100 °C the equilibrium constant for the reaction:



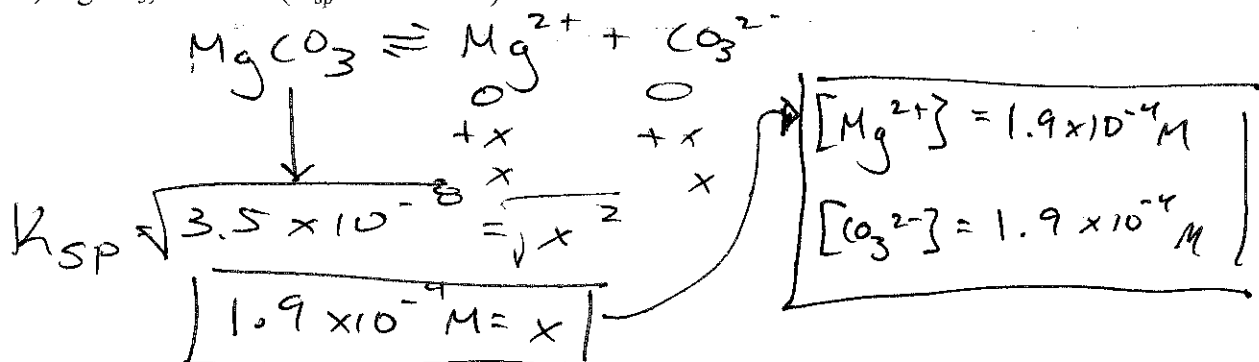
has the value $K_c = 2.19 \times 10^{-10}$. Are the following mixtures of COCl_2 , CO , and Cl_2 at 100 °C at equilibrium? $[\text{COCl}_2] = 2.00 \times 10^{-3} \text{ M}$; $[\text{CO}] = 3.3 \times 10^{-6} \text{ M}$; $[\text{Cl}_2] = 6.62 \times 10^{-6} \text{ M}$

$$Q = \frac{(3.3 \times 10^{-6})(6.62 \times 10^{-6})}{(2.00 \times 10^{-3})} = \frac{2.1846 \times 10^{-11}}{2 \times 10^{-3}} = 1.09 \times 10^{-8}$$

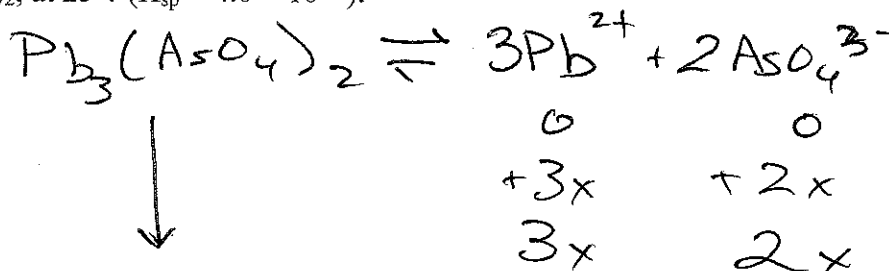
$Q > K \therefore$ Shift left

K_{sp}

8. Calculate the molar solubilities and equilibrium concentrations of Mg^{2+} and CO_3^{2-} in a saturated solution of magnesium carbonate, MgCO_3 , at 25°C? ($K_{sp} = 3.5 \times 10^{-8}$).



9. Calculate the molar solubilities and equilibrium concentrations of Pb^{2+} and AsO_4^{3-} in a saturated solution of lead (II) arsenate, $\text{Pb}_3(\text{AsO}_4)_2$, at 25°C? ($K_{sp} = 4.0 \times 10^{-36}$).



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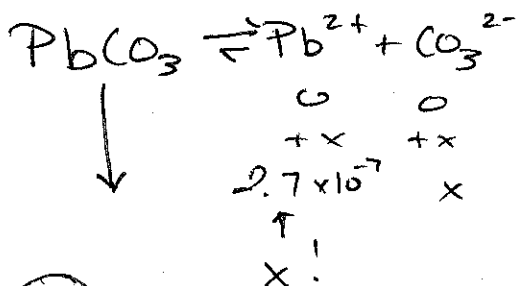
$$K_{sp} = (3x)^3 (2x)^2 = 41 \times 10^{-36}$$

$$4 \times 10^{-36} = 27x^3 \cdot 4x^2 = 108x^5$$

$$\sqrt[5]{3.7 \times 10^{-38}} = \sqrt[5]{x^5} \rightarrow \begin{cases} [Pb^{2+}] = 6.5 \times 10^{-8} M \\ [AsO_3^{3-}] = 1.1 \times 10^{-7} M \end{cases}$$

$$\boxed{3.6 \times 10^{-8} M = x}$$

10. A sample of lead (II) carbonate, $PbCO_3$, is added to water at $25^\circ C$. At equilibrium the concentration of lead (II) ions is $2.7 \times 10^{-7} M$. What is K_{sp} for $PbCO_3$?



$$K_{sp} = x^2 = (2.7 \times 10^{-7})^2$$

$$\boxed{K_{sp} = 7.3 \times 10^{-14}}$$

11. A 1.00 L solution of saturated lead (II) iodide at $25^\circ C$ contains 0.54 g of PbI_2 . Calculate the solubility product constant of this solution.

$$\frac{0.54 g}{2} \times \left(\frac{1 mol}{461 g} \right) = 0.00117 mol PbI_2 = x$$



$$K_{sp} = (x)(2x)^2 = 4x^3$$

$$K_{sp} = 4(0.00117)^3 = \boxed{6.4 \times 10^{-9}}$$

Ways to Impact Solubility

12. Identify which of the following compounds would be more soluble in an acidic solution when compared to a neutral solution. NOTE - There can be more than one possible answer

(a) $Mn(OH)_2$ - more soluble

(b) $AgCN$ - more soluble

c. $AuCl_3$

d. $ZnSO_4$

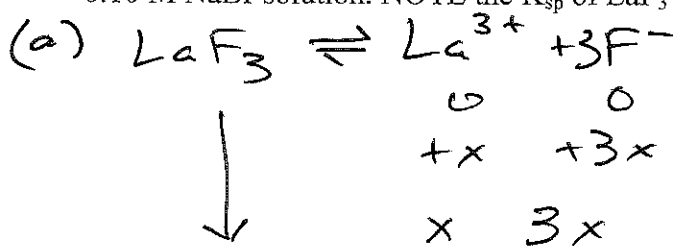
(e) $Ba_3(PO_4)_2$ - more soluble

add H^+ to anion and it is a strong acid (can't form this)

13. Calculate the solubility of $Mn(OH)_2$ in grams per liter when buffered at a pH of (a) 7.0, (b) 9.5, and (c) 11.8. Note the K_{sp} ($Mn(OH)_2$) = 1.6×10^{-13} .

(on back)

14. Calculate the molar solubility of LaF_3 in grams per liter in (a) pure water, (b) 0.010 M KF solution, and (c) 0.10 M NaBr solution. NOTE the K_{sp} of LaF_3 is 2×10^{-19}

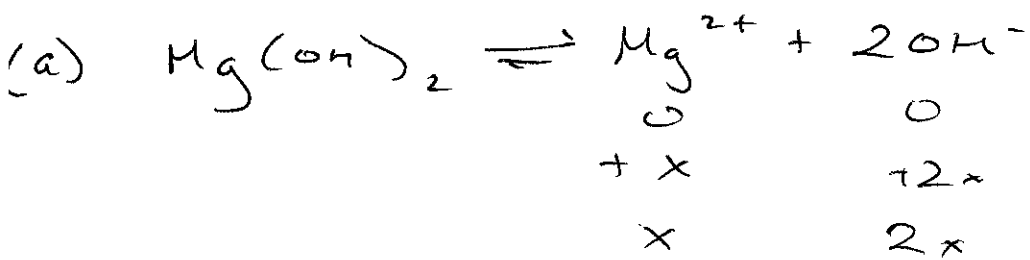


$$2 \times 10^{-19} = (x)(3x)^3$$

$$2 \times 10^{-19} = 27x^4$$

$$\sqrt[4]{7.4 \times 10^{-21}} = \sqrt[4]{x^4}$$

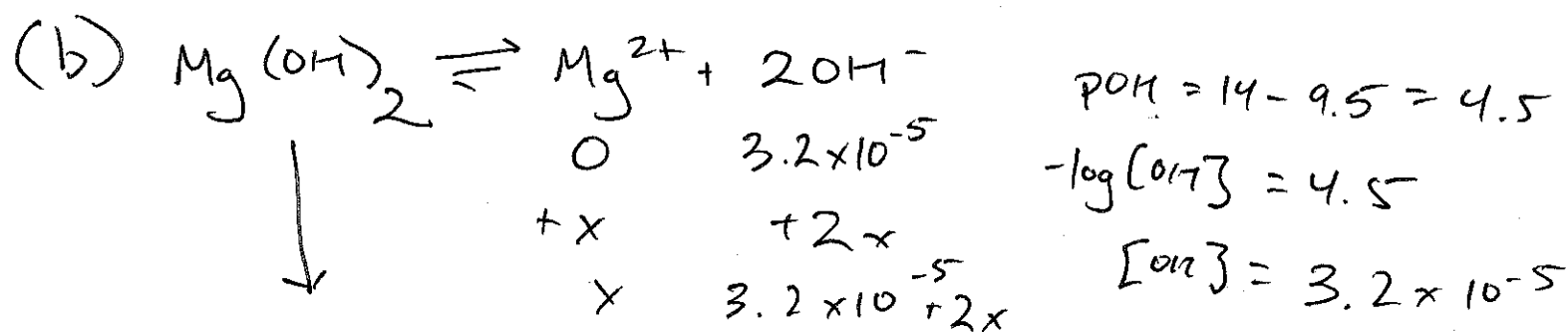
$$\boxed{9.2 \times 10^{-6} = x}$$



$$K_{SP} = 4x^3 = 1.6 \times 10^{-13}$$

$$3 \sqrt{x^3} = \sqrt[3]{4 \times 10^{-14}}$$

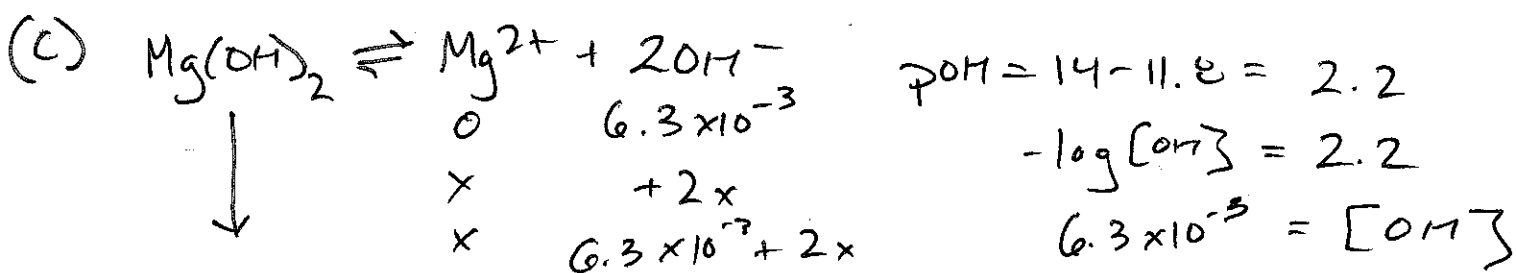
$$x = \frac{4.47 \times 10^{-4} \text{ mol}}{\text{L}}$$



$$K_{SP} = (x)(3.2 \times 10^{-5} + 2x)^2 = 1.6 \times 10^{-13}$$

$$x(1 \times 10^{-9}) = \frac{1.6 \times 10^{-13}}{1 \times 10^{-9}}$$

$$\boxed{x = 1.6 \times 10^{-4} \text{ M}}$$



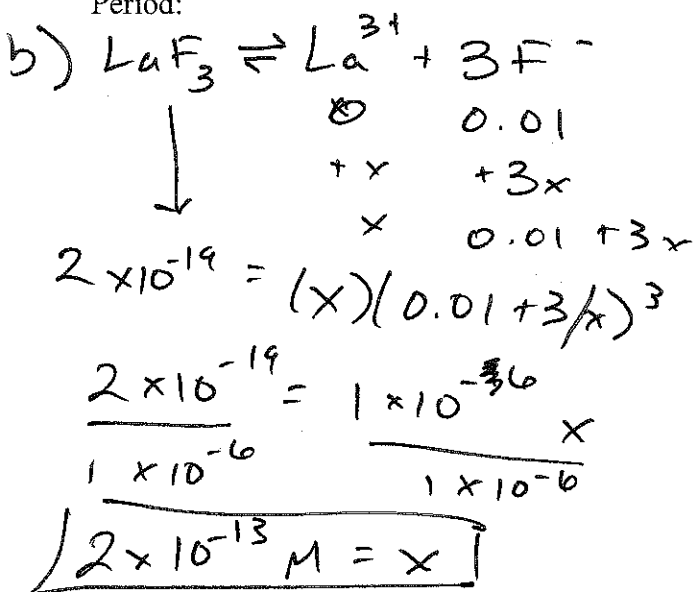
$$1.6 \times 10^{-13} = (x)(6.3 \times 10^{-3} + 2x)^2$$

$$\frac{1.6 \times 10^{-13}}{3.98 \times 10^{-5}} = \frac{x(3.98 \times 10^{-5})}{3.98 \times 10^{-5}}$$

$$\boxed{x = 4.0 \times 10^{-9} \text{ M}}$$

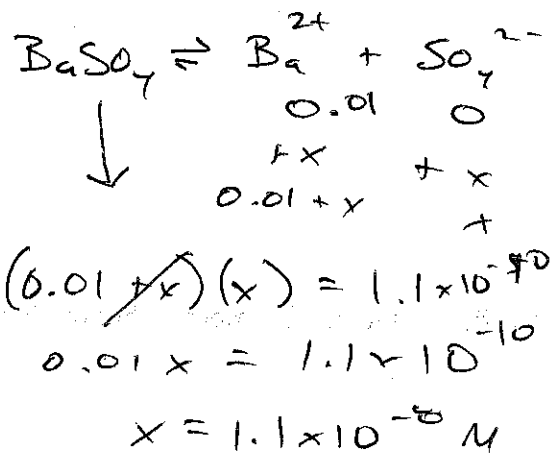
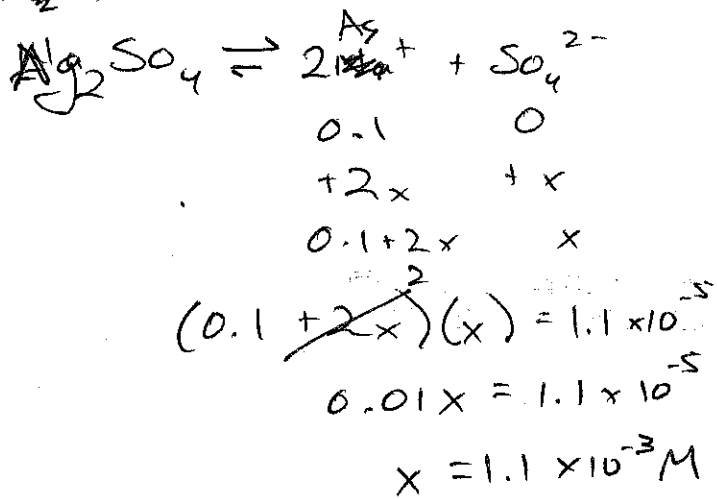
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(c)
 Br and Na
 are spectator ions,
 so x would
 be the same
 as (a) ($9.3 \times 10^{-4} \text{ M}$)

15. In a solution containing 0.010 M Ba^{2+} and 0.10 M Ag^+ , which solid will precipitate first when Na_2SO_4 is added to the solution? Justify your reasoning with calculations. NOTE the $K_{sp}(\text{BaSO}_4) = 1.1 \times 10^{-10}$ and $K_{sp}(\text{Ag}_2\text{SO}_4) = 1.1 \times 10^{-5}$.



Since a lower concentration of $[\text{SO}_4^{2-}]$ is required to reach equilibrium, BaSO_4 will ppt. as soon as $[\text{SO}_4^{2-}]$ exceeds the calculated amount.

