



SOLUTIONS

EXERCISE # O-I

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1.  $\text{H}_2\text{O} \rightarrow \text{H}^+ + \text{OH}^-$   $\Delta H = +ve$   
 $K_w$  if  $T$   $K_{eq} = K_w$
2.  $K_w = [\text{H}^+] \times [\text{OH}^-]$   
 for pure water  $[\text{H}^+] = [\text{OH}^-]$   
 $K_w = 10^{-6.7} \times 10^{-6.7} = 10^{-13.4}$
3.  $\text{pH}_i = 3$   $\therefore [\text{H}^+]_i = 10^{-3} \text{ M}$   
 $\text{pH}_f = 6$   $\therefore [\text{H}^+]_f = 10^{-6} \text{ M}$   
 $[\text{H}^+]_i V_i = [\text{H}^+]_f V_f$   
 $10^{-3} \times V_i = 10^{-6} \times V_f$   
 $V_f = 1000 V_i$   
 $C_i = 1000 C_f$
4.  $\text{pH}_i = 2 \Rightarrow [\text{H}^+]_i = 10^{-2} \text{ M}$   
 $\text{pH}_f = 6 \Rightarrow [\text{H}^+]_f = 10^{-6} \text{ M}$   
 $\frac{[\text{H}^+]_i}{[\text{H}^+]_f} = \frac{10^{-2}}{10^{-6}}$   
 $[\text{H}^+]_i = 10^4 [\text{H}^+]_f$
5.  $\text{pH} = 13$ , so  $\text{pOH} = 1 \Rightarrow [\text{OH}^-] = 0.1$   
 $0.1 = \frac{\text{Moles of OH}^-}{0.25}$   
 Moles of  $\text{OH}^- = 0.025$   
 Mass of  $\text{NaOH} = 0.025 \times 40 = 1 \text{ g}$
6.  $\text{pH}$  of acid  $< 7$   
 So  $\text{pH}$  sol should be  $6 - 7$
7.  $[\text{H}^+]_i = 10^{-2}$   $[\text{H}^+]_f = 10^{-3}$   
 initial moles  $= 10^{-2}$  final moles of  $\text{H}^+ = 10^{-3}$   
 Moles of  $\text{H}^+$  should be removed  $= 10^{-2} - 10^{-3} = 0.009$
8.  $[\text{H}^+]_f = \frac{10 \times \frac{1}{200} \times 2 + 40 \times \frac{1}{200} \times 2}{50} = \frac{1}{100} \Rightarrow \text{pH} = 2$
9. **Higher the  $K_a$ , stronger the acid**

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10.  $\text{NH}_3$  is base, so  $[\text{OH}^-]$ ,  $[\text{H}^+]$

12.  $\text{CH}_3\text{COOH} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}^+$

$$0.1(1 - \alpha) \quad 0.1\alpha \quad 0.1\alpha$$

$$10^{-5} = \frac{0.1\alpha \times 0.1\alpha}{0.1(1 - \alpha)}$$

$$10^{-5} = \frac{0.1\alpha^2}{1 - \alpha} \Rightarrow \frac{\alpha^2}{1 - \alpha} = 10^{-4} \Rightarrow \alpha = 10^{-2}$$

13.  $\text{NH}_4\text{OH} \rightarrow \text{NH}_4^+ + \text{OH}^-$

$$c(1 - \alpha) \quad c\alpha \quad c\alpha$$

$$[\text{OH}^-] = c\alpha = 0.02 \times \frac{5}{100} = 10^{-3}$$

$$\text{pOH} = 3 \Rightarrow \text{pH} = 11$$

15.  $[\text{H}^+]_f = \frac{10^{-3} + 10^{-5}}{2} = \frac{101}{2} \times 10^{-5}$

$$\text{pH} = 5 - \log \frac{101}{2} = 3.3$$

16.  $[\text{H}^+] = \sqrt{K_{a1}C_1 + K_{a2}C_2} = \sqrt{3 \times 10^{-4} \times 0.03 + 1 \times 10^{-10} \times 0.1} = 3 \times 10^{-3} \text{ M}$


$$[\text{A}^-] = \frac{3 \times 10^{-4} \times 0.03}{3 \times 10^{-3}} = 3 \times 10^{-3} \text{ M}$$

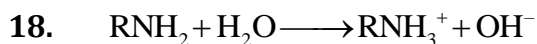
$$[\text{B}^-] = \frac{1 \times 10^{-10}}{3 \times 10^{-3}} = 3.33 \times 10^{-9} \text{ M}$$

17.  $[\text{OH}^-] = \sqrt{K_{a1}C_1 + K_{a2}C_2} = \sqrt{6.4 \times 10^{-5} \times \frac{0.1}{2} + 1.8 \times 10^{-5} \times \frac{4}{45} \times \frac{1}{2}} = \sqrt{\frac{8}{2} \times 10^{-6}} = 2 \times 10^{-3} \text{ M}$

$$\text{pOH} = 3 - \log 2$$

$$\text{pH} = 11 + \log 2 = 11.3$$

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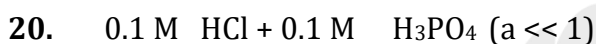
$$0.01-x \qquad \qquad \qquad x \qquad \qquad x$$

$$2 \times 10^{-6} = \frac{x - (x + 10^{-4})}{0.01 - x}$$

$$x^2 + 10^{-4}x - 2 \times 10^{-8} = 0$$

$$x = 10^{-4}$$

$$[\text{OH}^-] = 2 \times 10^{-4}$$



$$\begin{array}{ccc} 0.1(1 - \alpha) & 0.1 + 0.1\alpha & 0.1\alpha \\ & \gg 0.1 & \end{array}$$

$$[\text{H}^+] = 0.1 \Rightarrow \text{pH} = 1$$

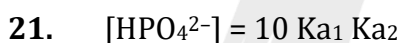
$$K_{a1} = \frac{0.1 \times 0.1\alpha}{0.1} \Rightarrow a = 10 K_{a1}$$



$$\begin{array}{ccc} 0.1\alpha & 0.1 & \alpha_2 \end{array}$$

$$K_{a2} = \frac{0.1 \times [\text{HPO}_4^{2-}]}{0.1\alpha}$$

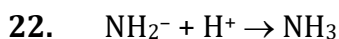
$$[\text{HPO}_4^{2-}] = \alpha K_2 = 10 K_{a1} K_{a2}$$




$$\begin{array}{ccc} 10 K_{a1} K_{a2} - x & x + 0.1 & x \\ & \gg 0.1 & \end{array}$$

$$K_{a3} = \frac{0.1 \times x}{10 K_{a1} K_{a2} - x} = \frac{0.1 \times x}{10 K_{a1} K_{a2}}$$

$$[\text{PO}_4^{3-}] = x = 100 K_{a1} K_{a2} K_{a3}$$



Base                      C.A.

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24. Bronsted acid  $\rightarrow$  which can give  $H^+$

24.  $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$   
 Acid      Base              C.A      C.B.

25.  $CH_3COOH + OH^- \rightleftharpoons CH_3COO^- + H_2O$

$$K_{eq} = \frac{1}{K_h} = \frac{K_a}{K_w} = \frac{1.8 \times 10^{-5}}{10^{-14}} = 1.8 \times 10^9$$

26.  $K_a HF \times K_b F^- = K_w$

$$pK_a = 14 - 10.83 = 3.17$$

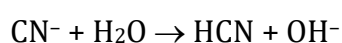
$$K_a = 10^{-3.17} = 6.75 \times 10^{-4}$$

28. At  $85^\circ C$  neutral pH  $< 7$

29.  $[H^+] = \frac{0.01}{100} = 10^{-3} M \Rightarrow pH = 3$


30. (A) 0.1 M NaCl  $\Rightarrow pH = 7$   
 (B) 0.1 M  $NH_4Cl$   $\Rightarrow pH < 7$   
 (C) 0.1 M  $CH_3COONa$   $\Rightarrow pH > 7$   
 (D) 0.1 M HCl  $\Rightarrow pH = 1$

31.  $h = \sqrt{\frac{K_w}{K_a} \times C} = \sqrt{\frac{10^{-14} \times 1/80}{1.3 \times 10^{-9}}} = \sqrt{\frac{10^{-6}}{8 \times 1.3}}$



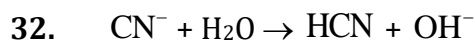
$$\frac{1}{80}(1-h) \quad h/80 \quad h/80$$

$$\frac{10^{-14}}{1.3 \times 10^{-9}} = \frac{h^2}{80(1-h)} \quad \text{p} \quad h^2 = \frac{8}{1.3} \times 10^{-4}$$

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$$h = 2.48 \times 10^{-2}$$

$$\% h = 2.48$$



$$0.01(1-h) \quad 0.01h \quad 0.01h$$

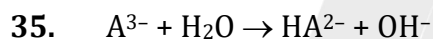
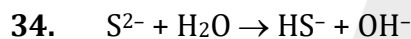
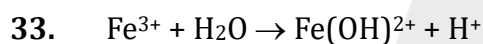
$$\frac{10^{-14}}{1.4 \times 10^{-9}} = \frac{0.01h^2}{1-h}$$

$$h^2 = \frac{10}{1.4} \times 10^{-4}$$

$$h = 2.67 \times 10^{-2}$$

$$\% h = 2.67\%$$

$$K_{b_3} > K_{b_2} > K_{b_1}$$



$$K_{h1} = \frac{K_w}{K_{a3}}$$

$$\text{pH} = \frac{1}{2} (\text{p}K_w + \text{p}K_{a3} + \log C)$$

36.  $\frac{h}{1-h} = \sqrt{Kh}$  (for salt of WA + WB)


h does not depend on conc.

37.  $\text{pH} = \frac{1}{2} (\text{p}K_w + \text{p}K_a - \text{p}K_b) = \frac{1}{2} \times (14 + 4.8 - 4.78) = 7.01$

38. Amphiprotic species @ which can give as well as take  $\text{H}^+$  ion.

41. For 1 M NaCl & 1 M HCl solution

$$\Rightarrow \text{pH} < 7$$

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42.  $pOH = pK_b = 4.74$   
 $pH = 9.26$

43.  $[OH^-] = K_b \times \frac{[NH_3]}{[NH_4^+]} = 1.8 \times 10^{-5} \times \frac{0.05}{0.001} = 9.0 \times 10^{-4}$

44.  $pOH = 5 - \log 1.85 + \log \frac{500 \times 0.5}{300 \times 0.3}$   
 $= 5 + \log \frac{25}{16.2} = 5.188$   
 $pH = 14 - 5.188 = 8.812$

45.  $pH = pK_a + \log \frac{[CH_3COO^-]}{[CH_3COOH]}$   
 $= 4.82 + \log \frac{3}{2} = 5$


46.  $pH = 5 + \log \frac{10 \times 1}{50 \times 2} = 5 - 1 = 4$

47.  $pH = 10 - \log 5 + \log \frac{5 \times V_{mL}}{10 \times 2} = 9$   
 $\log \frac{5V_{mL}}{20} = \log 0.5$   
 $V_{mL} = 2 \text{ mL}$

48.  $4.74 = 5 - \log 1.34 + \log \frac{x}{0.02}$   
 $x = 1.5 \times 10^{-2} \text{ mol}$

49.  $4 = 3.7 + \log \frac{0.1 \times V_{mL}}{50 \times 0.05}$   
 $2 = \frac{0.01 \times V_{mL}}{50 \times 0.05} \Rightarrow V_{mL} = 50 \text{ mL}$

50. If  $[NH_4OH]^-$ ,  $[OH^-]$ ,  $pH^-$

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51.  $\text{pH} = \text{pK}_a + \log \frac{[\text{C.B.}]}{[\text{Acid}]}$

$$\text{pH} - \text{pK}_a = \log \frac{[\text{C.B.}]}{[\text{Acid}]} = 5 = \log 10^5$$

$$[\text{C.B.}] = 10^5 \times [\text{Acid}]$$



$$10 \text{ mmol} \quad x \text{ mmol} \quad 10 \text{ mmol} \quad -$$

$$10-x \quad 0 \quad 10+x \quad -$$

$$\text{pH} = \text{pK}_a + \log \frac{[\text{A}^-]}{[\text{HA}]}$$

$$5.5 = 5 + \log \frac{10+x}{10-x}$$

$$3.16 = \frac{10+x}{10-x} \Rightarrow x = 5.2$$

$$\text{Mass of NaOH added} = 5.2 \times 40 \times 10^{-3} \text{g} = 2.08 \times 10^{-1} \text{g}$$

53. On moderate dilution pH of buffer solution does not change.


54.  $\text{Buffer capacity} = \frac{0.02}{0.05} = 0.4$

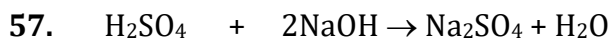
56.  $\text{Moles of OH}^- = \frac{8}{40} = \frac{1}{5}$

$$\text{Moles of H}^+ = \frac{4.9}{98} \times 2 = \frac{1}{10}$$

$$[\text{OH}^-] = \frac{0.1}{1} = 0.1$$

$$\text{pOH} = 1 \Rightarrow \text{pH} = 13$$

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0.5 mmol      1 mmol

0                  0                  1 mmol

Solution will be neutral so pH = 7



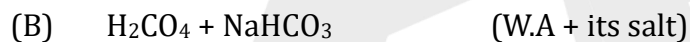
M Moles of  $\text{OH}^-$  = 5

$$[\text{H}^+]_f = \frac{10}{100} = \frac{1}{10} \quad \text{p} \quad \text{pH}_f = 1$$



mmi      4          2                  0          0

mmf      2          0                  2          2



mm :      16                  8                  0                  0

mmf      8                  0                  8                  -


$$\text{pH} = 3.7 + \log \frac{8}{8} = 3.7$$

$$\text{pOH} = 10.3$$

62.  $\text{pH} = 4.74 + \log \frac{0.6}{0.4} = 4.916$

63.  $\text{pH} = \text{pKa} = 3.7$



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64.  $\text{pH} = 5 + \log \frac{25}{75}$  (25% neutralization)

$$= 5 + \log \frac{1}{3}$$

$\text{pH} = 5$  (50% neutralization)

$\text{pH} = 5 + \log \frac{75}{25}$  (75% neutralization)

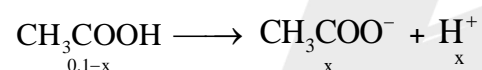
$$= 5 - \log 3$$



mmi	2	2	0	0
-----	---	---	---	---

mmf	0	0	2	-
-----	---	---	---	---

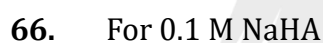
$$[\text{CH}_3\text{COOH}] = \frac{2}{20} = 0.1$$



$$1.8 \times 10^{-5} = \frac{x^2}{0.1-x}$$

$$[\text{H}^+] = x = \sqrt{1.8} \times 10^{-3}$$

$$\text{pH} = 2.88$$

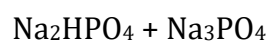
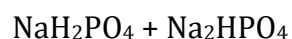



$$\text{pH} = \frac{7+11}{2} = 9$$

for 0.1 M  $\text{H}_2\text{A}$   $\text{pH} < 7$

for 0.1 M  $\text{Na}_2\text{A}$   $\text{pH} > 7$

for 0.1 M  $\text{NaHA} + 0.1 \text{ M Na}_2\text{A} \Rightarrow \text{pH} = \text{pK}_{a2} = 11$



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$$\text{mmi} \quad 9 \quad 5 \quad 0 \quad 0$$

$$\text{mmf} \quad 4 \quad 0 \quad 5 \quad -$$

$$\text{pH} = 3 + \log \frac{5}{4} = 3.1$$

69.

$$\text{pH} = \text{pK}_{\text{In}} + \log \frac{[\text{In}^-]}{[\text{HIn}]}$$

70.  $\text{pH} = \text{pK}_{\text{In}} \pm 1$

71. pH range of methyl red  
 $\Rightarrow 4.2 - 6.3$

72. pH at equivalence point is  $(5.5 - 11)$

73. For HCl Vs  $\text{NH}_4\text{OH} \Rightarrow \text{pH range } (3 - 8.5)$

74. For NaOH Vs  $\text{H}_2\text{C}_2\text{O}_4$  titration  
 pH range is  $(11 - 5.5)$

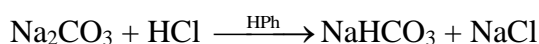
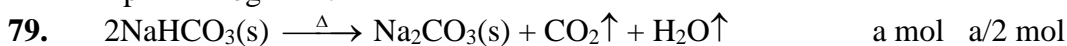
75. Oxalic acid Vs  $\text{KMnO}_4$  titration is a redox titration.

76. For WA Vs SB titration  
 pH range is  $5.5 - 11$ . So suitable indicator is Hph.

78.  $\text{pH}_i = \text{pK}_a + \log \frac{20}{80}$

$$\text{pH}_f = \text{pK}_a + \log \frac{80}{20}$$

$$\Delta \text{pH} = 2 \log 4 = 1.2$$




$$\frac{a}{2} = 0.1 \times 0.5$$

$$a = 0.1 \text{ mol}$$

$$w = 0.1 \times 84 = 8.4 \text{ gm}$$

$$\% \text{ purity} = \frac{8.4}{10} \times 100 = 84\%$$

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81.  $K_{sp} = (28)^2 \times (3s)^3 = 108 s^5$

82.  $K_{sp} = s^2 = 64 \times 10^{-6}$

$$s = 8 \times 10^{-3} \text{ M}$$

$$= 8 \times 10^{-3} \text{ mol/L}$$

$$= 8 \text{ mol/m}^3$$

83.  $s = \frac{14.3 \times 10^{-4}}{143} = 10^{-5} \text{ M}$

$$K_{sp} = s^2 = 10^{-10} \text{ M}^2$$

84.  $s = \frac{7}{100} = 7 \times 10^{-2} \text{ M}$

$$K_{sp} = s^2 = 4.9 \times 10^{-3}$$

85. (A)  $K_{sp} = s^2$

$$s = \sqrt{80} \times 10^{-19} \text{ M}$$

(B)  $K_{sp} = s^2$

$$s = \sqrt{7} \times 10^{-8} \text{ M}$$

(C)  $K_{sp} = 108 s^5 = 1 \times 10^{-72}$

$$s = \left( \frac{1000}{108} \right)^{1/5} \times 10^{-15} \text{ M}$$

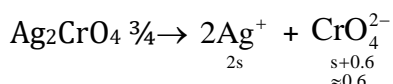
(D)  $K_{sp} = 27 s^4$

$$s = \left( \frac{180}{27} \right) \times 10^{-5} \text{ M}$$

86.  $4s^3 = 8.64 \times 10^{-13}$

$$s^3 = 216 \times 10^{-15}$$

$$s = 6 \times 10^{-5} \text{ M (in water)}$$




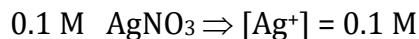
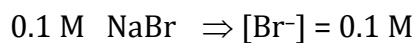
$$(2s)^2 \times 0.6 = 8.64 \times 10^{-13}$$

$$4s^2 = 14.4 \times 10^{-13}$$

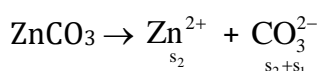
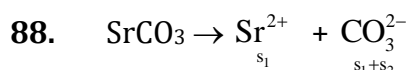
$$s^2 = 3.6 \times 10^{-13}$$

$$s = 6 \times 10^{-7} \text{ M}$$

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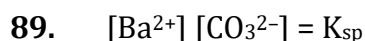


In case of  $0.1 \text{ M CaBr}_2$ , molarity of common ion is maximum so. There maximum common ion effect.



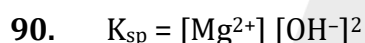
$$\frac{(s_1 + s_2)s_1}{(s_1 + s_2)s_2} = \frac{10^{-10}}{1.5 \times 10^{-11}} = \frac{100}{15}$$

$$\frac{s_1}{s_2} = \frac{20}{3}$$



$$5.1 \times 10^{-9} = [\text{Ba}^{2+}] \times 10^{-4}$$

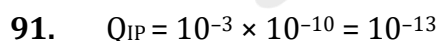
$$[\text{Ba}^{2+}] = 5.1 \times 10^{-5} \text{ M}$$



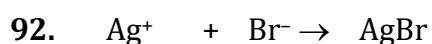
$$1 \times 10^{-11} = 0.1 \times [\text{OH}^-]^2$$

$$[\text{OH}^-] = 10^{-5} \text{ M}$$

$$\text{pOH} = 5 \text{ } \& \text{ pH} = 9$$



$$Q_{IP} < K_{sp} \text{ (no ppt)}$$

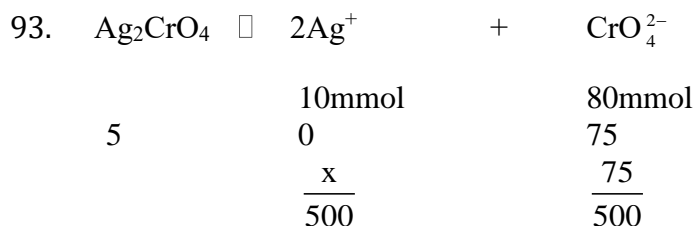


$$\text{Ci} \quad 10^{-7} \text{ M} \quad 10^{-7} \text{ M} \quad 0$$

$$Q_{ip} = 10^{-7} \times 10^{-7} = 10^{-14}, Q_{ip} < K_{eq},$$

So there will be no precipitation

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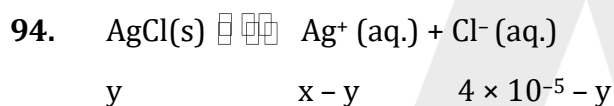


$$\text{then } \left(\frac{x}{500}\right)^2 \left(\frac{75}{500}\right) = 15 \times 10^{-12}$$

$$\frac{x}{500} = 10^{-5}$$

$$x = 5 \times 10^{-3}$$

$$[\text{Ag}^+] = \frac{5 \times 10^{-3}}{500} = 10^{-5}$$



$$4 \times 10^{-5} - y = 10^{-5}$$

$$y = 3 \times 10^{-5}$$

$$(x - y)(4 \times 10^{-5} - y) = 10^{-10}$$

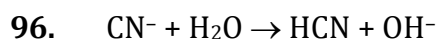
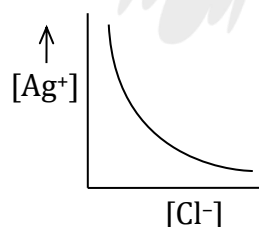
$$(x - y)(10^{-5}) = 10^{-10}$$

$$x - y = 10^{-5}$$

$$x - 3 \times 10^{-5} = 10^{-5}$$


$$x = 4 \times 10^{-5}$$

95.  $K_{sp} = [\text{Ag}^+][\text{Cl}^-]$



Solubility if  $[\text{OH}^-]$ ,  $[\text{H}^+]$ ,  $\text{pH}$

97. Solubility of  $\text{Ag}_2\text{CO}_3$  will be max. in  $\text{NH}_3$  due to complex formation.

Link to View Video Solution:  [Click Here](#)



$$x \quad 0.8 \quad 0.1$$



$$0.1 - x \quad x \quad 0.8 + x$$

$$\sim 0.1 \quad \sim 0.8$$

$$\frac{1}{K_{f_2}} = \frac{1}{10^4} = \frac{(0.8)x}{0.1} \Rightarrow x = 1.25 \times 10^{-5}$$



$$S \quad S-x$$



$$S-x \quad x \quad x$$

$$10^{-4} = \frac{K_w}{K_a} = \frac{x^2}{S-x} \Rightarrow \frac{10^{-10}}{S-x} \Rightarrow (S-x) = \frac{10^{-10}}{10^{-4}} = 10^{-6}$$

$$\Rightarrow S = x + 10^{-6} \text{ \{but } x = 10^{-5}\} \Rightarrow S = (1.1 \times 10^{-5})$$

$$K_{sp} = [\text{Ag}^+(aq)][\text{A}^-(aq)] \Rightarrow 1.1 \times 10^{-5} \times 10^{-6}$$

$$K_{sp} = 1.1 \times 10^{-11}$$