81. (c) $BaSO_4 \rightleftharpoons Ba^{2+} + SO_4^{2-}$

$K_{sp} = S^2 \Rightarrow S = \sqrt{K_{sp}}; K_{sp} = [Ba^{2+}] \times [SO_4^{2-}]$

$\mathbf{4 \times 10^{-10} = [1 \times 10^{-4}] \times [SO_4^{2-}]}$

$[SO_4^{2-}] = \frac{4\times10^{-10}}{1\times10^{-4}} = 4 \times 10^{-6}.$

82. (c) $AB_2 \rightleftharpoons \underset{(S)}{A^{2+}} + \underset{(2S)^2}{2B^-}$

$K_{sp} = 4S^3$

$S = \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{4\times10^{-12}}{4}} = 1 \times 10^{-4} gm.mol/litre$

83. (d) $[B] = \frac{K_{sp}AB}{[A]} = \frac{1\times10^{-8}}{10^{-3}} = 1 \times 10^{-5} M$

Where ionic product > K_{sp}, ppt formed

∴8 should be more then $10^{-5} M$.

84. (c) $NaCl_{(s)} \rightleftharpoons Na^+_{(aq)} + Cl^-_{(aq)}$

$HCl \rightleftharpoons H^+ + Cl^-$. The increase in $[Cl^-]$ brings in an increase in $[Na^+][Cl^-]$ which will lead for backward reaction because

$K_{sp}(NaCl) = [Na^+][Cl^-]$

means Ionic product $\geq K_{sp}$

85. (d) $BaSO_4 \rightleftharpoons \underset{(S)}{Ba^{++}} + \underset{(S)}{SO_4^{--}}$

$K_{sp} = S^2; S = \sqrt{K_{sp}} = \sqrt{1.3 \times 10^{-9}}$

$= 3.6 \times 10^{-5} mol/litre$

86. (c) 3.2×10^{-10}

Explanation (Simple & Clear):

AgCl is sparingly soluble (only dissolves a little in water).

Therefore, its Ksp value is very small.

Standard value of Ksp of AgCl at 25°C is:

Ksp = $[Ag^+] \times [Cl^-] = 3.2 \times 10^{-10}$

This value is experimentally determined and generally memorized for exams.

87. (b) Alkaline,

$$CH_3COONa + H_2O \rightleftharpoons \underset{\text{Weakacid}}{CH_3COOH} + \underset{\text{Strongbase}}{NaOH}$$

88. (c) Because it is a strong base.

89. (b) For pure water $[H^+] = [OH^-]$, $\therefore K_w = 10^{-12}$s

90. (a) $MX_2 \rightleftharpoons \underset{(S)}{M^{2+}} + \underset{(2S)^2}{2X^-}$; $4S^3 = 4 \times (0.5 \times 10^{-4})^3$

$= \mathbf{5 \times 10^{-13}}$

91. (a) Solubility coefficient $= [Pb^{2+}][Cl^-]^2$

92. (a) Solubility of $Al(OH)_3$ is lesser than $Zn(OH)_2$.

93. (c) $NaCl_{(s)} \rightleftharpoons Na^+{}_{(aq)} + Cl^-{}_{(aq)}$

$HCl \rightleftharpoons H^+ + Cl^-$

The increase in $[Cl^-]$ brings in an increase in $[Na^+][Cl^-]$ which will lead for backward reaction because $K_{sp}NaCl = [Na^+][Cl^-]$.

94. (c) Common ion effect.

95. (a) $\underset{S}{CaF_2} \rightleftharpoons \underset{S}{Ca^{++}} + \underset{(2S)^2}{2F^-}$

$$K_{sp} = 4S^3$$

$$S = \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{3.2\times10^{-11}}{4}} = 2\times10^{-4} m/l.$$

96. (d) In aqueous solution following euilibrium is exist. $H_2S \rightleftharpoons H^+ + HS^-$

While adding the dilute HCl solution

$(HCl \rightleftharpoons H^+ + Cl^-)$ equilibrium is shift to the left side in $H_2S \rightleftharpoons H^+ + HS^-$

97. (d) $\underset{K_{sp}}{M_2X_3} \rightleftharpoons \underset{(2y)^2}{2M^{+++}} + \underset{(3y)^3}{3X^{--}}$

Solubility product $K_{sp} = 108y^5 mol\frac{d}{m^3}$

98. (b) Solubility is directly proportional to the K_{sp}.

99. (b) $\underset{S}{PbCl_2} \rightleftharpoons \underset{S}{Pb^{++}} + \underset{(2S)^2}{2Cl^-}$

$$K_{sp} = S\times(2S)^2 = 4S^3$$

$$S = \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{1.5\times10^{-4}}{4}} = 3.34\times10^{-2}.$$

100. (b) BCl_3

Explanation

A Lewis acid is a species that can accept an electron pair.

Why BCl_3 is a Lewis acid

Boron has only 3 valence electrons and forms three covalent bonds with chlorine, giving only 6 electrons around boron.

Therefore, BCl_3 is electron-deficient and readily accepts an electron pair → acts as a Lewis acid.