Answers to exam questions for Chapter 10 Solubility equilibria

Question 1

a $K_s = [Pb^{2+}][CrO_4^{2-}]$ A

b $[Pb^{2+}][CrO_4^{2-}] = 1.8 \times 10^{-14}$

$[Pb^{2+}] = [CrO_4^{2-}]$ (concentrations are equal, see equation)

$[CrO_4^{2-}]^2 = 1.8 \times 10^{-14}$ (square root both sides)

$[CrO_4^{2-}] = 1.3 \times 10^{-7}$ mol L^{-1}

A = correct method used with minor error, M = correct value with unit

c i Amount of Pb^{2+} present: $n(Pb^{2+}) = 0.100$ mol $L^{-1} \times 30 \times 10^{-3}$ L

Amount of CrO_4^{2-} present: $n(CrO_4^{2-}) = 0.100$ mol $L^{-1} \times 20 \times 10^{-3}$ L

Hence excess Pb^{2+} present: $n(Pb^{2+}) = 0.100$ mol $L^{-1} \times 10 \times 10^{-3}$ L (Pb^{2+}: CrO_4^{2-} react 1:1)

= 0.00100 mol

$V_{tot} = 30 \times 10^{-3}$ L + 20×10^{-3} L $V(Pb^{2+}) + V(CrO_4^{2-})\}$

= 0.050 L

$$[Pb^{2+}] = \frac{0.00100 \text{ mol}}{0.050 \text{ L}}$$

= 0.020 mol L^{-1}

A = correct method used with minor error, M = correct value with unit

ii $$[CrO_4^{2-}] = \frac{Ks}{[Pb^{2+}]}$$ (rearranging equation from **a**)

$$= \frac{1.8 \times 10^{-14}}{0.020 \text{ mol L}^{-1}}$$ (substitute known values)

$= 9.0 \times 10^{-13}$ mol L^{-1}

A = correct method used with minor error, M = correct value with unit

Question 2

Ionic Product = $[Ag^+][Cl^-]$

= 0.0018 mol L^{-1} × 0.00020 mol L^{-1}

= 3.6×10^{-7}

Ionic Product > K_s (1.6×10^{-10}), so a precipitate will form (compare IP with K_s)

A = correct value for IP, M = plus correct conclusion

Question 3

a $K_s = [Mg^{2+}][OH^-]^2$ A

b Mg^{2+} A (1 mole of $Mg(OH)_2$ contains 1 mole of Mg^{2+} and 2 moles of OH^-)

c $K_s = [Mg^{2+}][OH^-]^2$

Let the solubility of $Mg(OH)_2$ be s

$[Mg^{2+}] = s$ $[OH^-] = 2s$ (concentrations are 1:2, see equation)

$K_s = 4s^3$ and K_s for $Mg(OH)_2 = 1.8 \times 10^{-11}$ (cube root both sides after dividing by 4)

$$s = \sqrt[3]{\frac{1.8 \times 10^{-11}}{4}}$$

$= 1.7 \times 10^{-4}$ mol L^{-1} (round to 2 sig.fig. and don't forget the units)

A = correct method used with minor error, M = correct throughout

d $[OH^-] = 2\,[Mg^{2+}]$

$= 3.4 \times 10^{-4}$ mol L^{-1}

$$[H_3O^+] = \frac{1 \times 10^{-14}}{OH^-}$$ ($K_w = [H_3O^+][OH^-]$)

$= 2.92 \times 10^{-11}$ mol L^{-1}

$pH = -\log_{10}[H_3O^+]$

$= 10.5$

$= 11$ (round to 2 sig.fig.)

e **i** the same A (solid $Mg(OH)_2$ present – saturated)

ii The solution is still saturated with Mg^{2+} because more solid $Mg(OH)_2$ will dissolve to keep the concentration of Mg^{2+} unchanged. M

f **i** $Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$ A

ii Lower, because beaker A contains less OH^-, (pH of 0.10 mol L^{-1} KOH = 13). (KOH strong base)

A = correct answer, M = correct reasoning

iii Lower, because in beaker B the precipitation of $Mg(OH)_2$ has removed some of the Mg^{2+} ions from the solution. (common ion effect)

A = correct answer, M = correct reasoning