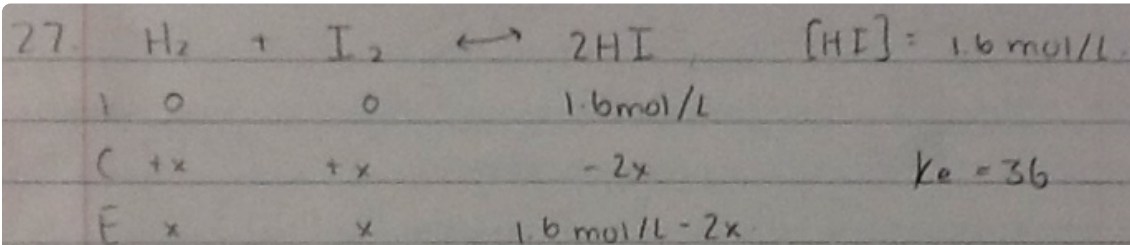


because $2x = 0.4$, $\therefore [\text{H}_2] = [\text{I}_2] = 1.2 \text{ mol/L} - 0.2 \text{ mol/L}$
 $x = 0.2 \text{ mol/L}$ $= 1 \text{ mol/L}$

$$\therefore K_e = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} = \frac{[0.4 \text{ mol/L}]^2}{[1 \text{ mol/L}][1 \text{ mol/L}]} = 0.16$$

\therefore the value of K_e at equilibrium is 0.16



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

$$36 = \frac{(1.6 - 2x)^2}{x^2}$$

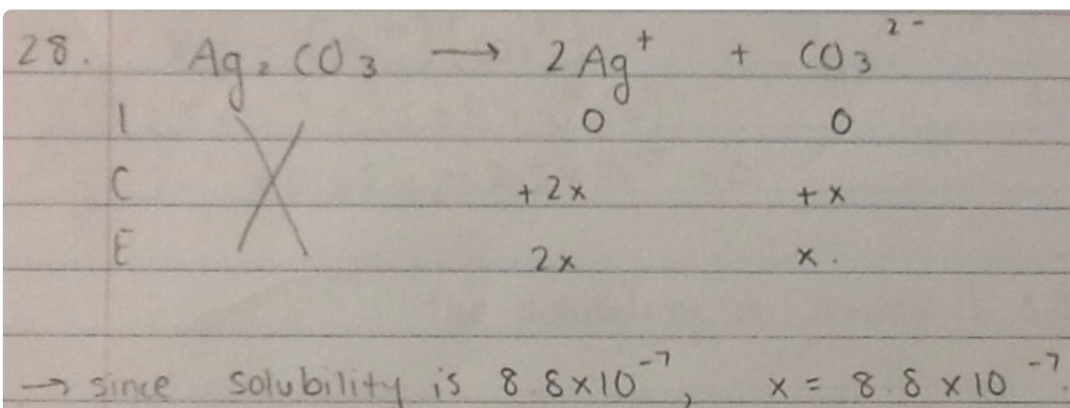
$$36x^2 - 2.56 - 6.4x + 4x^2$$

$$32x^2 + 6.4x - 2.56 = 0$$

$$= 0.2 \text{ mol/L}$$

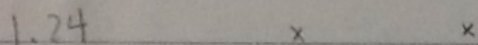
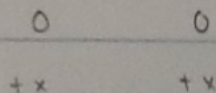
$$\therefore [\text{H}_2] = [\text{I}_2] = 0.2 \text{ mol/L}$$

$$[\text{HI}] = 1.2 \text{ mol/L}$$



$$\begin{aligned}
 K_{sp} &= [\text{Ag}^+]^2 [\text{CO}_3^{2-}] \\
 &= [2x]^2 [x] \\
 &= 4x^3 \\
 &= 4(8.8 \times 10^{-7})^3 \\
 &= 2.7 \times 10^{-15}
 \end{aligned}$$

∴ the K_{sp} value is 2.7×10^{-15}



→ since it is a 1:1 relationship, $[\text{CN}^-] = [\text{H}^+]$

$$K_a = \frac{[\text{H}^+][\text{CN}^-]}{[\text{HCN}]}$$

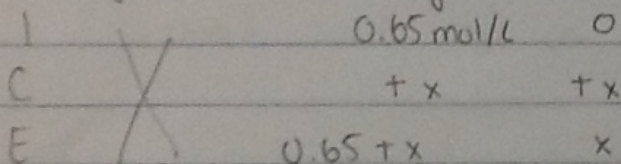
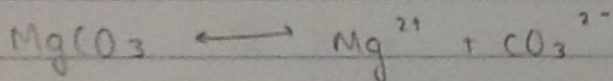
$$6.2 \times 10^{-10} = \frac{[\text{H}^+]^2}{1.24 \text{ mol/L}}$$

$$[\text{H}^+] = 2.77 \times 10^{-5}$$

$$\begin{aligned} \text{pH} &= -\log [\text{H}^+] \\ &= -\log [2.77 \times 10^{-5}] \\ &= 4.56 \end{aligned}$$

∴ the pH of the solution is 4.56.

if $[\text{MgCl}_2] = 0.65 \text{ mol/L}$, $\text{MgCl}_2 \rightleftharpoons \text{Mg}^{2+} + 2\text{Cl}^-$
1:1 relationship ∴ $[\text{Mg}^{2+}] = 0.65 \text{ mol/L}$.



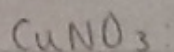
$$K_{sp} = [\text{Mg}^{2+}][\text{CO}_3^{2-}]$$

$$2.5 \times 10^{-5} = (0.65 + x)(x)$$

$$x^2 + 0.65x - 2.5 \times 10^{-5} = 0$$

$$\begin{aligned} x &= \frac{-0.65 \pm \sqrt{(0.65)^2 + 4(2.5 \times 10^{-5})}}{2} \\ &= \frac{-0.65 \pm 0.65008}{2} \\ &= 3.85 \times 10^{-5} \end{aligned}$$

∴ the solubility of MgCO_3 is 3.85×10^{-5} .



$$v = 0.145 \text{ L}$$

$$c = 0.0078 \text{ mol/L}$$

$$n = 1.131 \times 10^{-3} \text{ mol}$$

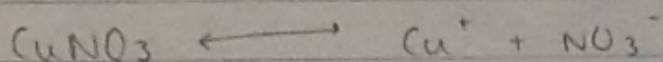


$$v = 0.195 \text{ L}$$

$$c = 1.48 \times 10^{-3} \text{ mol/L}$$

$$n = 2.886 \times 10^{-4} \text{ mol}$$

→ when mixed together, volume becomes 0.34 L.



1

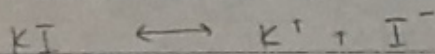
1

$$\text{Cu}^+ : n = 1.131 \times 10^{-3} \text{ mol}$$

$$v = 0.34 \text{ L}$$

$$c = \frac{n}{v}$$

$$[\text{Cu}^+] = 3.326 \times 10^{-3} \text{ mol/L}$$



1

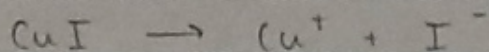
1

$$\text{I}^- : n = 2.886 \times 10^{-4} \text{ mol}$$

$$v = 0.34 \text{ L}$$

$$c = \frac{n}{v}$$

$$[\text{I}^-] = 8.488 \times 10^{-4} \text{ mol/L}$$



$$Q = [\text{Cu}^+][\text{I}^-]$$

$$= [3.326 \times 10^{-3} \text{ mol/L}] [8.488 \times 10^{-4} \text{ mol/L}]$$

$$= 2.82 \times 10^{-6}$$

$$\text{since } 2.82 \times 10^{-6} > 1 \times 10^{-12},$$

$$Q > K_{sp}$$

∴ there will be a precipitate formed.

32. * let acetic acid = HA

HA

KOH

$$V = 0.028 \text{ L}$$

$$V = 0.021 \text{ L}$$

$$C = 0.36 \text{ mol/L}$$

$$C = 0.43 \text{ mol/L}$$

$$n = 0.01008 \text{ mol}$$

$$n = 0.00903 \text{ mol}$$

$\therefore 1.05 \times 10^{-3} \text{ mol}$ of HA will be left over

Since KOH is limiting reagent, $[A^-] = 0.00903$

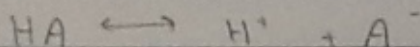
\rightarrow new volume = 0.049 L

$$[HA] = \frac{1.05 \times 10^{-3} \text{ mol}}{0.049 \text{ L}}$$

$$[A^-] = \frac{0.00903 \text{ mol}}{0.049 \text{ L}}$$

$$= 0.1843 \text{ mol/L}$$

$$= 0.0214 \text{ mol/L}$$



$$I \quad 0.0214 \text{ mol/L} \quad 0 \quad 0.1843 \text{ mol/L}$$

$$C \quad -x \quad +x \quad +x$$

$$E \quad 0.0214 - x \quad x \quad 0.1843 \text{ mol/L}$$

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

Since K_a is more than $1000 \times$ smaller, we can assume existing concentrations won't be effected.

$$1.8 \times 10^{-5} = \frac{x(0.1843)}{(0.0214)}$$

$$x = 2.09 \times 10^{-6}$$

$$pH = -\log [H^+]$$

$$= -\log [2.09 \times 10^{-6}]$$

$$= 5.68$$

the pH of the solution is 5.68.