

- 24 (d)
 Concentration of HCl solution = 1×10^{-2}
 \therefore Millimoles of HCl solution = $200 \times 1 \times 10^{-2}$
 $= 2$

Similarly, millimoles of NaOH solution
 $= 300 \times 1 \times 10^{-2}$
 $= 3$

Concentration of the resultant solution

$$= \frac{3-2}{300+200}$$

$$= \frac{1}{500} = 0.2 \times 10^{-2}$$

$$[\text{OH}^-] = 0.2 \times 10^{-2}$$

$$[\text{H}^+] = \frac{10^{-14}}{0.2 \times 10^{-2}} = 5 \times 10^{-12}$$

$$\text{pH} = -\log[\text{H}^+]$$

$$\text{pH} = -\log[5 \times 10^{-12}]$$

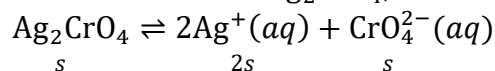
$$\text{pH} = 11.3$$

- 25 (b)
 Dissociation constant
 $\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$; $[\text{H}^+] = [\text{OH}^-] = 1 \times 10^{-7} \text{ M}$
 And $[\text{H}_2\text{O}] = 1 \text{ g/mL} = 1000 \text{ gL}^{-1}$
 $= \frac{1000}{18} \text{ mol L}^{-1} = 55.56 \text{ M}$
 $K = \frac{[\text{H}^+][\text{OH}^-]}{[\text{H}_2\text{O}]} = \frac{10^{-14}}{55.6}$
 $K_w = 1 \times 10^{-14}$

So, $K_w = 55.6 \times K$

- 26 (a)
 The species which can accept as well as donate H^+ can act both as an acid and a base
 $\text{HSO}_4^- + \text{H}^+ \rightleftharpoons \text{H}_2\text{SO}_4$
 base
 $\text{HSO}_4^- \rightleftharpoons \text{SO}_4^{2-} + \text{H}^+$
 acid

- 27 (a)
 For saturated solution of Ag_2CrO_4 , if solubility is 's' mol L^{-1} . Then



$$K_{\text{sp}} = (2s)^2(s) = 4s^3$$

$$K_{\text{sp}} = 3.2 \times 10^{-11} \text{ (given)}$$

$$\therefore 3.2 \times 10^{-11} = 4s^3$$

$$s^3 = \frac{3.2 \times 10^{-11}}{4} = 8 \times 10^{-12}$$

$$\therefore s = \sqrt[3]{8 \times 10^{-12}} = 2 \times 10^{-4} \text{ M}$$

- 28 (b)
 The salt with lower solubility product (AgI) will precipitate first from the solution by adding KI.
- 29 (c)

K_c is a characteristic constant for the given reaction

30

(a)

The value of equilibrium constant is independent of volume of container.

∴ Value of equilibrium constant will remain same (300) if volume of reaction flask is tripled.

31

(c)



Let the solubility of BaCl_2 is x mol/L

$$\begin{aligned} \therefore K_{sp} &= [\text{Ba}^{2+}][\text{Cl}^-]^2 \\ &= (x) \times (2x)^2 \\ &= x \times 4x^2 = 4x^3 \end{aligned}$$

$$\begin{aligned} \text{or solubility of BaCl}_2 &= \frac{(\text{solubility product of BaCl}_2)^{1/3}}{4} \\ &= \frac{(4 \times 10^{-9})^{1/3}}{4} \\ &= 10^{-3} \text{ mol/L} \end{aligned}$$

32

(c)

$$K_p = K_c (RT)^{\Delta n_g}$$

Here, $\Delta n_g = -1$ and $K_c = 26$

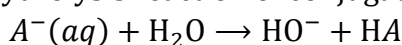
$$R = 0.0812, T = 250 + 273 = 523 \text{ K}$$

$$K_p = 26 \times (0.0812 \times 523)^{-1} = 0.605 = 0.61$$

33

(c)

The hydrolysis reaction of conjugate base of acid is



$$K_h = \frac{K_w}{K_a} = \frac{10^{-14}}{10^{-4}} = 10^{-10}$$

Since, degree of hydrolysis is negligible;

$$[\text{OH}^-] = \sqrt{K_h C} = 10^{-6}. p[\text{OH}] = 6$$

and $\text{pH} = 14 - 6 = 8$

34

(a)



$$\begin{array}{ccc} 1 & 0 & 0 \end{array}$$

$$\begin{array}{ccc} p & p & \text{At equilibrium} \end{array}$$

Given,

$$p + p = 100 \text{ atm}$$

$$2p = 100 \text{ atm}$$

$$p = 50 \text{ atm}$$

$$K_p = p_{\text{NH}_3} \cdot p_{\text{H}_2\text{S}}$$

$$= 50 \times 50 = 2500 \text{ atm}^2$$

35

(c)

$$K_p = \frac{p_{\text{CO}}^2}{p_{\text{CO}_2}} = \frac{4 \times 4}{2} = 8$$

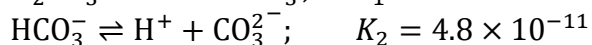
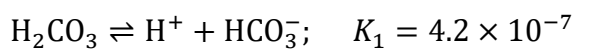
36

(c)

The equilibrium constant does not change when concentration of reactant is changed as the concentration of product also get changed accordingly.

37

(c)



$$K_1 \gg K_2$$

$$\therefore [\text{H}^+] = [\text{HCO}_3^-]$$

$$K_2 = \frac{[\text{H}^+][\text{CO}_3^{2-}]}{[\text{HCO}_3^-]}$$

$$\text{So, } [\text{CO}_3^{2-}] = K_2 = 4.8 \times 10^{-11}$$

38 (d)

Buffer capacity of an acidic buffer is maximum when the ratio of HA to A^- is unity.

$$\text{Since, pH of acidic buffer} = \text{p}K_a + \log \frac{[A^-]}{[HA]}$$

$$\text{For maximum buffer capacity, } [A^-] = [HA]$$

$$\therefore \text{pH} = \text{p}K_a$$

39 (a)

(i) BF_3 and BCl_3 are better Lewis acid than BMe_3 and B_2H_6 due to $-I$ -effect of Cl and F.

(ii) Between BCl_3 and BF_3 , BCl_3 is better Lewis acid due to lesser extent of $p\pi - p\pi$ back bonding in it.

40 (d)

Number of milliequivalents of

$$\text{NaOH} = 10 \times 0.1 = 1$$

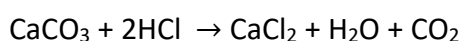
Number of milliequivalents of

$$\text{H}_2\text{SO}_4 = 10 \times 0.05 = 0.5$$

Milliequivalent of H_2SO_4 will neutralise NaOH so, that $(1-0.5)=0.5$ millieq. Of NaOH is left unneutralised.

Hence, pH of the resulting solution will be greater than seven ($\text{pH} > 7$).

41) a)



$$1 \text{ mol CaCO}_3 (100 \text{ g CaCO}_3) = 2 \text{ mol HCl} (73 \text{ g HCl})$$

$$\therefore \text{for } 40 \text{ g CaCO}_3 = x \text{ g HCl}$$

$$\Rightarrow x = \frac{40 \times 73}{100} = 29.2 \text{ g of HCl}$$

HCl is in excess.

$\therefore \text{CaCO}_3$ is limiting reagent

42) b)

43) c)

$$\text{CaCl}_2 \text{ Solution} = 25\% \frac{W}{W}$$

$$\text{Mass (CaCl}_2) = \frac{25 \times 10}{100} = 2.5 \text{ g}$$

$$n_{\text{CaCl}_2} = \frac{2.5}{111} = 0.0225$$

$$n_{\text{H}_2\text{O}} = \frac{20}{18} = 1.111$$

$$x_{\text{CaCl}_2} = \frac{0.0225}{1.111 + 0.0225} = 0.0168$$

44) c

Molarity of acetic acid = 2.05 mol

∴ number of moles of acetic acid = 2.05 mol

Mass of acetic acid in 1 L solution = 2.05 x 60 = 123 g

$$\text{Density of solution} = \frac{\text{mass of solution}}{\text{volume}}$$

$$\Rightarrow \text{mass} = 1.02 \times 1000 = \underline{1020} \text{ g}$$

Mass of solvent = 1020 – 123 = 897 g = 0.897 kg

$$\Rightarrow \text{molality (m)} = \frac{n}{\text{mass of solvent (kg)}}$$

$$= \frac{2.05}{0.897} = 2.28 \text{ mol kg}^{-1}$$

45) a

concentration of BaCl₂ = 0.045 M

0.045 M = 0.045 mol

∴ mass of BaCl₂ = 0.045 x 208 g

∴ 15 g of BaCl₂ are present in

$$= \frac{1000 \times 15}{0.045 \times 208} \text{ ml}$$

$$= 1602.6 \text{ ml}$$