JEE Mains 2019 Chapter wise Question Bank

Ionic Equilibrium - Questions

Q1

20 mL of 0.1 M H_2SO_4 solution is added to 30 mL of 0.2 M NH_4OH solution. The pH of the resultant mixture is: [pK_b of $NH_4OH = 4.7$].

(1) 5.2 (2) 9.0
(3) 5.0 (4) 9.4

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Q2

For the equilibrium

$2H_2O \rightleftharpoons H_3O^+ + OH^-$; the value of ΔG^o at 298 K is approximately:

(1) 100 kJ mol^{-1} (2) –80 kJ mol^{-1}
(3) 80 kJ mol^{-1} (4) –100 kJ mol^{-1}

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Q3

If K_{sp} of Ag_2CO_3 is 8×10^{-12}, the molar solubility of Ag_2CO_3 in 0.1 M $AgNO_3$ is :

(1) 8×10^{-12} M (2) 8×10^{-11} M
(3) 8×10^{-10} M (4) 8×10^{-13} M

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Q4

If solubility product of $Zr_3(PO_4)_4$ is denoted by K_{sp} and its molar solubility is denoted by S, then which of the following relation between S and K_{sp} is correct ?

(1) $S = \left(\frac{K_{sp}}{144}\right)^{1/6}$ (2) $S = \left(\frac{K_{sp}}{6912}\right)^{1/7}$

(3) $S = \left(\frac{K_{sp}}{929}\right)^{1/9}$ (4) $S = \left(\frac{K_{sp}}{216}\right)^{1/7}$

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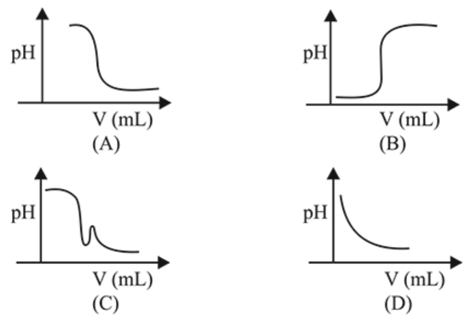
Q5

In an acid base titration, 0.1 M HCl solution was added to the NaOH solution of unknown strength. Which of the following correctly shows the change of pH of the titration mixture in this experiment?



(1) (B) (2) (A) (3) (C) (4) (D)

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Q6

Consider the following statements

(a) The pH of a mixture containing 400 mL of 0.1 M H_2SO_4 and 400 mL of 0.1 M NaOH will be approximately 1.3.
(b) Ionic product of water is temperature dependent.
(c) A monobasic acid with K_a–10^{-5} has a pH = 5. The degree of dissociation of this acid is 50%.
(d) The Le Chatelier's principle is not applicable to common-ion effect.

The correct statements are :

(1) (a), (b), and (d) (2) (a), (b) and (c)
(3) (b) and (c) (4) (a) and (b)

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Q7

The pH of a 0.02 M NH_4Cl solution will be [given $K_b(NH_4OH)=10^{-5}$ and log 2 = 0.301]

(1) 2.65 (2) 4.35 (3) 4.65 (4) 5.35

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Q8

What is the molar solubility of $Al(OH)_3$ in 0.2 M NaOH solution ? Given that, solubility product of $Al(OH)_3$ = 2.4×10^{-24} :

(1) 3×10^{-19} (2) 12×10^{-21}

(3) 3×10^{-22} (4) 12×10^{-23}

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Q9

The molar solubility of $Cd(OH)_2$ is 1.84×10^{-5} M in water. The expected solubility of Cd $(OH)_2$ in a buffer solution of pH = 12 is :

(1) 1.84×10^{-9} M (2) $\frac{2.49}{1.84}\times10^{-9}$ M

(3) 6.23×10^{-11} M (4) 2.49×10^{-10} M

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Ionic Equilibrium - Answers

Q1

(2) m. mol of $H_2SO_4 = 20 \times 0.1 = 2$

m. mol of $NH_4OH = 30 \times 0.2 = 6$

$$H_2SO_4 + 2NH_4OH \rightarrow (NH_4)_2SO_4 + 2H_2O$$

	H_2SO_4	$2NH_4OH$	$(NH_4)_2SO_4$
Initial	2 m mol	6 mmol	0
Final	(2-2) = 0 m mol	$(6-2\times2)$ = 2 m mol	2 m mol

$[NH_4OH]_{left} = 2$ m mol

$[(NH_4)_2SO_4] = 2$ m mol

$[\overset{+}{N}H_4] = 2 \times 2 = 4$ m mol

Total Volume = 30 + 20 = 50 mL

$$pOH = pK_b + \log\left[\frac{Salt}{Base}\right]$$

$$= 4.7 + \log\frac{4/50}{2/50}$$

$$= 4.7 + \log 2 = 5$$

$pH = 14 - pOH$

$pH = 14 - 5 = 9$

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Q2

(3) $\Delta G = \Delta G^\circ + RT\ln Q$

At equilibrium; $\Delta G = 0$ and $Q = K_{eq}$

$\Rightarrow \Delta G^\circ = -2.303\,RT\log K_w$

$= -2.303 \times 8.314 \times 298 \times \log 10^{-14}$

$= 79.9$ kJ/mol ≈ 80 kJ/mol

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Q3

(3) As $AgNO_3$ dissociates completely,

therefore in 0.1 M $AgNO_3$ solution, $[Ag^+] = 0.1$ M

$$\underset{}{AgNO_3} \longrightarrow \underset{0.1}{Ag^+} + \underset{0.1}{NO_3^-}$$

$$Ag_2CO_3 \rightleftharpoons \underset{0.1+2s}{2Ag^+} + \underset{s}{CO_3^{2-}}$$

$$K_{sp} = [Ag^+]^2[CO_3^{2-}]$$

$8 \times 10^{-12} = (0.1 + 2s)^2 \times s$

$0.01\,s = 8 \times 10^{-12}$; $(0.1 + 2s \approx 0.1)$

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

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Q4

(2) $Zr_3(PO_4)_4 \rightleftharpoons \underset{3S}{3Zr^{4+}} + \underset{4S}{4PO_4^{3-}}$

$K_{sp} = [Zr^{4+}]^3[PO_4^{3-}]^4 = (3S)^3(4S)^4$

$K_{sp} = 6912\,S^7$

$$S = \left(\frac{K_{sp}}{6912}\right)^{1/7}$$

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Q5

(2) Graph A and B, both represents the titration curve between strong acid and strong base, i.e., HCl and NaOH but, the pH of NaOH is more than 7 and during the titration it decreases, so graph (A) is correct.

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Q6

(2) $H_2SO_4 + NaOH \rightarrow NaHSO_4 + H_2O$

	H_2SO_4	$NaOH$	$NaHSO_4$	H_2O
Initial mol	0.04	0.04	0	0
mol at eqm.	0	0	0.04	0.04

$NaHSO_4 \rightarrow Na^+ + H^+ + SO_4^{2-}$

$$[H^+] = \frac{0.04}{0.80} = 0.05M; pH = 1.3$$

(b) Ionic product of water increases with increase in temperature because ionisation of water is endothermic.

(c) $HA \rightleftharpoons H^+ + A^-$

	HA	H^+	A^-
Initial	C	0	0
At eqm.	$C(1-\alpha)$	$C\alpha$	$C\alpha$

Given $pH = 5 \Rightarrow -\log(H^+) = 5$

$\therefore [H^+] = 10^{-5}$

As we know,

$$K_a = \frac{C\alpha^2}{1-\alpha}$$

$$10^{-5} = \frac{C\alpha^2}{1-\alpha} = \frac{C\alpha.\alpha}{(1-\alpha)}$$

$$10^{-5} = 10^{-5}\frac{\alpha}{1-\alpha}; \alpha = \frac{1}{2}\ i.e., 50\%$$

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Q7

(4) $pH = 7 - \frac{1}{2}pK_b - \frac{1}{2}\log C$

$$= 7 - \frac{5}{2} - \frac{1}{2}(\log 2\times10^{-2}) = 5.35$$

$pH = 5.35$

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Q8

(3) Let the solubility of $Al(OH)_3$ in 0.2M NaOH solution be s.

Then,

$$\underset{}{Al(OH)_3} \rightleftharpoons \underset{s}{Al^{3+}} + \underset{3s}{3OH^-}$$

and $\underset{0.2M}{NaOH} \rightleftharpoons \underset{0.2M}{Na^+} + \underset{0.2M}{OH^-}$

$[Al^{3+}] = s$ and $[OH^-] = 3s + 0.2 \approx 0.2$

$K_{sp} = 2.4 \times 10^{-24} = [Al^{3+}][OH]^3$

$2.4 \times 10^{-24} = s(0.2)^3$

$$s = \frac{2.4\times10^{-24}}{8\times10^3} = 3\times10^{-22} mol/L$$

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Q9

(4) $Cd(OH)_2 \rightleftharpoons \underset{s}{Cd^{2+}} + \underset{2s}{2OH^-}$

At equilibrium, $K_{sp} = s(2s)^2 = 4s^3$

$\Rightarrow K_{sp} = 4 \times (1.84 \times 10^{-5})^3$

Solubility in buffer solution having pH = 12

$[OH^-] = 10^{-2}$

$Cd(OH)_2 \rightleftharpoons \underset{s}{Cd^{2+}} + \underset{2s'+10^{-2}\approx10^{-2}}{2OH^-}$

$\therefore K_{sp} = 4 \times (1.84 \times 10^{-5})_3 = s'(10^{-2})^2$

$$\Rightarrow s' = \frac{24.9\times10^{-15}}{10^{-4}} = 2.49\times10^{-10} M$$

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