

61. (b) $[H^+] = 2 \times 10^{-2} M$

$$\therefore pH = -\log[2 \times 10^{-2}];$$

pH = 1.7 i.e. in between 1 and 2.

62. (c) Basic

Word-friendly explanation (fast):

Molar mass of $Na_2CO_3 \cdot H_2O \approx 124 \text{ g mol}^{-1}$.

Moles added = $0.62 / 124 = 0.005 \text{ mol}$ (this gives $0.005 \text{ mol CO}_3^{2-}$).

$(NH_4)_2SO_4$ given as 0.1 N in 100 mL:

For acid–base purposes (2 NH_4^+ per formula unit) molarity = $N/2 = 0.05 \text{ M}$.

Moles of $(NH_4)_2SO_4 = 0.05 \times 0.100 \text{ L} = 0.005 \text{ mol} \rightarrow$ gives 0.01 mol NH_4^+ .

Stoichiometry: each CO_3^{2-} can accept 2 H^+ (or neutralize 2 NH_4^+).

Here CO_3^{2-} moles = 0.005 and NH_4^+ moles = $0.01 \Rightarrow$ exactly enough NH_4^+ to neutralize CO_3^{2-} .

Net reaction converts $CO_3^{2-} + 2 \text{ NH}_4^+ \rightarrow \text{H}_2\text{CO}_3 + 2 \text{ NH}_3$.

H_2CO_3 partly decomposes to $\text{CO}_2 + \text{H}_2\text{O}$, while NH_3 (a weak base) remains in solution.

Amount of NH_3 produced $\approx 0.01 \text{ mol}$ in $\approx 0.1 \text{ L} \rightarrow [\text{NH}_3] \approx 0.1 \text{ M}$, which makes the solution basic ($pH > 7$).

63. (b) $pH = 4, (H^+) = 10^{-pH} = 10^{-4} M$

64. (b) 11, 3, 7

Explanation (word-friendly):

NaOH (10^{-3} M): strong base $\rightarrow [\text{OH}^-] = 10^{-3} \text{ M}$.

$pOH = 3 \rightarrow pH = 14 - 3 = 11$.

HCl (10^{-3} M): strong acid $\rightarrow [\text{H}^+] = 10^{-3} \text{ M}$.

$pH = 3$.

NaCl (10^{-3} M): neutral salt (from strong acid + strong base) \rightarrow solution is essentially neutral ($pH \approx 7$) at this dilution (neglecting very small ionic effects).





$$[\text{OH}^-] = 10^{-5} \text{M}; [\text{H}^+][\text{OH}^-] = 10^{-14}$$

$$[\text{H}^+] = \frac{10^{-14}}{10^{-5}}; [\text{H}^+] = 10^{-9} \text{M}; \text{pH} = 9.$$



Explanation (Word-friendly):

Let the acid be:

H_2A (dibasic acid)

It gives 2 H^+ ions when fully ionized.

Concentration of acid = 0.05 M

So, total H^+ ions produced = $2 \times 0.05 = 0.10 \text{ M}$

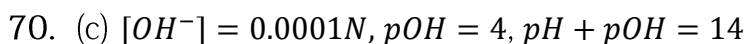
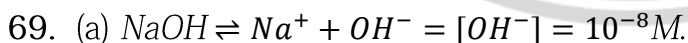
Now,

$$\text{pH} = -\log (\text{H}^+)$$

$$\text{pH} = -\log (0.10) = 1$$

67. (b) $\text{pH} = \text{pK}_a + \log \frac{[\text{Salt}]}{[\text{Acid}]}; \text{pH} = 4.75 + \log \frac{0.1}{0.1}$

$$\text{pH} = 4.75 + \log 1; \text{pH} = 4.75$$



$$\text{pH} = 14 - \text{pOH} = 14 - 4 = 10$$



$$[\text{OH}^-] = 0.001 \text{M} = 1 \times 10^{-3} \text{M}$$

$$[\text{H}^+] \times [\text{OH}^-] = 1 \times 10^{-14}$$



$$[H^+] = \frac{1 \times 10^{-14}}{[OH^-]}$$

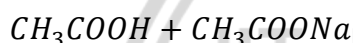
$$[H^+] = \frac{1 \times 10^{-14}}{1 \times 10^{-3}} = 1 \times 10^{-14} \times 10^3$$

$$[H^+] = 10^{-11} M$$

$$pH = 11$$

72. (a) An acid buffer solution consists of solution of weak acid with strong base of its salt.

73. (b) An acid buffer solution consists of a weak acid and its salt with strong base. *i.e.*



74. (a) $pOH = pK_b + \log \frac{[salt]}{[base]}$

$$= 5 + \log \frac{0.02}{0.2} = 5 + \log \frac{1}{10} = 5 + (-1) = 4$$

$$pH = 14 - pOH = 14 - 4 = 10$$

75. (b) $[Salt] = 0.1 M$, $[Acid] = 0.1 M$

$$K_a = 1.8 \times 10^{-5}; \quad pH = -\log K_a + \log \frac{[Salt]}{[Acid]}$$

$$= -\log 1.8 \times 10^{-5} + \log \frac{0.1}{0.1} = -\log 1.8 \times 10^{-5}$$

$$pH = 4.7.$$

76. (a) NH_4Cl and NH_4OH is a buffer solution (weak base and salt of strong acid).

77. (a) $pH + pOH = 14$; $pH = 14 - pOH$ $\therefore [OH^-] = 10^{-7}$

$$pOH = 7$$

$$\therefore pH = 14 - 7 = 7.$$



78. (c) $0.01\text{ M Ba(OH)}_2 = 0.02\text{ N Ba(OH)}_2$

$$N_1 V_1 = N_2 V_2$$

$$[0.02\text{ N}] \times [50\text{ ml}] = N_2 \times 100\text{ ml}$$

$$N_2 = \frac{0.02 \times 50}{100} = 10^{-2}\text{ N}; [\text{OH}^-] = 10^{-2}\text{ N}$$

$$p\text{OH} = 2 \text{ or } pH = 12$$

79. (b) $pH = -\log [H^+]$.

80. (a) Na_2CO_3 is a mixture of weak acid and strong base, so it is a base.

