

101. (c) For NH_4OH .

$$[OH^-] = C \cdot \alpha ; C = \frac{1}{10} M, \alpha = 0.2$$

$$[OH^-] = \frac{1}{10} \times 0.2 = 2 \times 10^{-2} M$$

$$pOH = -\log [OH^-] = \log [2 \times 10^{-2}] ; pOH = 1.7$$

$$pH = 14 - pOH = 14 - 1.7 = 12.30.$$

102. (c) $pH = pK_a + \log \frac{[Salt]}{[Acid]}$. For small concentration of buffering agent and for maximum buffer capacity $\frac{[Salt]}{[Acid]} \approx 1$.

103. (a) $[H^+] =$ increased ten fold means pH of solution decreased by one.

$$pH = \log \frac{1}{[H^+]}$$

104. (a) Because the pH of buffer are not changed.

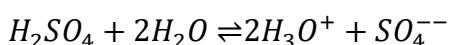
$$105. (c) pH = pK_a + \log \frac{[Salt]}{[Acid]}, 5.5 = 4.5 + \log \frac{[Salt]}{[0.1]}$$

$$\log \frac{[Salt]}{0.1} = 5.5 - 4.5 = 1$$

$$\frac{[Salt]}{0.1} = \text{antilog } 1 = 10 ; [Salt] = 1$$

106. (a) Moles of $H_2SO_4 = \frac{0.49}{98} = 5 \times 10^{-3}$ moles of H_2SO_4 present per litre of solution

$$(\text{molarity}) = \frac{.005}{1} = .005 M.$$



one mole of H_2SO_4 give 2 moles of H_3O^+ ions.

$$H_3O^+ = 2 \times (H_2SO_4) = 2 \times 0.005 = 0.01 M$$



$$[H^+] = 10^{-2} M; pH = 2$$

107. (c) CH_3COONH_4 is a simple buffer and called salt of weak acid.

108. (c) N.eq. for $HCl = \frac{0.4}{1000} \times 50 = 0.02$

N.eq. for $NaOH = \frac{0.2}{1000} \times 50 = 0.1$

Now $[OH^-]$ left = $0.1 - 0.02$

$$[OH^-] = .08 = 8 \times 10^{-2} M$$

$$pOH = -\log 8 \times 10^{-2} M; pOH = 1.0$$

109. (d) Buffer is mixture of weak base and its acid salt.

110. (b) $[NaOH] = 0.4/40 \text{ mole/l.} = 0.1M$

$$[OH^-] = 10^{-1} M, [H^+] = 10^{-13} M, pH = 13$$

111. (d) $pH + pOH = 14, pH = 4, H^+ = 10^{-4} \text{ mole/litre.}$

112. (d) Buffer solution have constant pH . When we add the water into this buffer solution. So no effect on it.

113. (b) $Ba(OH)_2 \rightleftharpoons Ba^{2+} + 2OH^-$

One molecule on dissociation furnishes $2OH^-$ ions.

$$\text{So, } [OH^-] = 2 \times 10^{-4} N$$

$$N = M \times 2; M = \frac{N}{2} = \frac{2 \times 10^{-4}}{2} = 10^{-4}$$



$$pOH = -\log[OH^-] = -\log(1 \times 10^{-4}) = -4$$

$$pH + pOH = 14; pH = 14 - 4 = 10.$$

114. (a) M.eq. of $0.10 \text{ M} HCl = \frac{0.10}{1000} \times 40 = 0.004 \text{ M}$

$$\text{M.eq. of } 0.45 \text{ M} NaOH = \frac{0.45 \times 10}{1000} = 0.0045 \text{ M}$$

$$\text{Now left } [OH^-] = 0.0045 - 0.004 = 5 \times 10^{-4} \text{ M}$$

Total volume = 50 ml.

$$[OH^-] = \frac{5 \times 10^{-4}}{50} \times 1000; [OH^-] = 1 \times 10^{-2}$$

$$pOH = 2; pH = 14 - pOH = 12.$$

115. (c) $0.001 \text{ M} HCl = 10^{-3} \text{ M} [H^+], pH = 3.$

116. (d) $[NaOH] = \frac{0.4}{40} = 0.01 \text{ M}; [OH^-] = 10^{-2} \text{ M}$

$$[H^+] = 10^{-12}, pH = -\log[H^+] = 12$$

117. (b) Those substance which give a proton is called Bronsted acid while

CH_3COO^- doesn't have proton so it is not a Bronsted acid.

118 (b) 2

Explanation (Word-friendly):

H_2SO_4 is a strong dibasic acid \rightarrow it gives 2 H^+ ions.

Concentration of $H_2SO_4 = 0.005 \text{ M}$

Total H^+ produced = $2 \times 0.005 = 0.01 \text{ M}$

Now,

$$pH = -\log(H^+)$$

$$pH = -\log(0.01) = 2$$



119. (d) Weak acid and conjugate base

Explanation (Word-friendly):

A buffer solution resists change in pH.

It is usually made from:

A weak acid (like CH₃COOH)

And its salt with a strong base, which provides the conjugate base (like CH₃COONa)

Example:



120. (c) $pH = -\log [H^+]$

