161. (c) pH = pKa

Explanation:

A buffer **resists pH change best** when the concentrations of the weak acid and its conjugate base are equal.

At this point: **pH = pKa** (from Henderson-Hasselbalch equation).

162. (d) $pH = -\log K_b + \log \frac{[\text{Salt}]}{[\text{Acid}]}$

$pH = -\log[1.8 \times 10^{-5}] + \log \frac{[\text{Salt}]}{1.0}$

$9 = 4.7 + log \frac{[\text{Salt}]}{1.0}$; $log \frac{[\text{Salt}]}{1.0} = 4.7 - 9 = -4.3$

$\frac{[\text{Salt}]}{1.0} = \text{Antilog} \frac{1}{4.3}$; $[Salt] = 1.8$

163. (b) $pH = -\,log\, K_b + log \frac{[\text{salt}]}{[\text{acid}]}$

$5 = -\,log\, 1\, 0^{-4} + log \frac{[\text{salt}]}{[\text{acid}]}$

$log \frac{[\text{salt}]}{[\text{acid}]} = 1$

$\frac{[\text{salt}]}{[\text{acid}]} = \text{antilog}1 = 10:1$

164. (a) $1\ M\ KOH$ show highest pH value because it is a strong base.

165. (d) $NH_4OH$is a weak acid and $NH_4Cl$is a strong base salt.

166. (a) $pH = 13.6$

$\boldsymbol{pOH = 14 - 13.6 = 0.4}$

$[\boldsymbol{OH^-}] = \textbf{Antilog}(\boldsymbol{-0.4}) = \boldsymbol{0.3979}$. So the value of $[OH^-]$between$0.1 M and 1M$

167. (d) Aspirin is a weak acid. Due to common ion effect it is unionised in acid medium but completely ionised in alkaline medium.

168. (b) $[H^+][OH^-] = 10^{-14}$;$(10^{-7})(10^{-7}) = 10^{-14}$

169. (c) $HCl = 10^o M$has $pH = 0$. The value of *pH* decreases as concentration further increases.

170. (a) Because pure water has a 7 *pH*.

171. (c) When concentration of $[H^+]$ increased then the value of *pH* is decreases.

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

172. (c) The concentration of $[H^+] = 10^{-2}$mole/litre

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

173. (d) Due to common ion effect.

174. (b) In water solution.

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

concentration of $OH^-$is increased so that solution become more basic and the *pH* is increased.

175. (a) Na_2CO_3 is basic in nature. So its *pH* is greater than 7.

176. (c) It is not a mixture of weak acid or base and their strong salt.

177. (a) $[H^+] = \text{Antilog}(-4.58)$;

$[H^+] = 2.63 \times 10^{-5} moles/litre$

178. (c) $10^{-2} MNaOH$will give $[OH^-] = 10^{-2}$

$\therefore pOH = 2$, Also $pH + pOH = 14$

$\therefore pH = 12$.

179. (a) $pH = pK_a + \log\frac{[\text{Salt}]}{[\text{Acid}]} = -log\, 2 \times 10^{-5} + log\frac{10\times1}{50\times2} = 4$.

180. (b) 0.001 $M\, NaOH$means $[OH^-] = 10^{-3}$; $pOH = 3$

$pH + pOH = 14; pH = 14 - 3$

$pH = 11$; $[H^+] = 10^{-11} mole\text{-}litre^{-1}$