141. (b) $[H^+] = 1.00 \times 10^{-6}\ mole/litre$

$pH = -\log\ [1.00 \times 10^{-6}]$; $pH = 6$.

142. (a) $[H^+]$ is in moles per litre.

143. (d) As the solution is acidic, $pH < 7$. This is because $[H^+]$ from H_2O $(10^{-7}M)$ cannot be neglected in comparison to 10^{-8}.

145. (b) pH of $0.001\ M\ HCl = 10^{-3}M[H^+]$, $pH = 3$.

146. (d) Because it can furnish H^+ ions in solutions.

147. (c) Because it is a strong acid.

$H^+ = 10^{-1}$

$\boldsymbol{pH = -\log\ [H^+] = -\log\ [10^{-1}]; pH = 1}$.

148. (b) Buffer solution is a combination of weak acid and conjugate base. $NaCl$ is a salt and $NaOH$ is the base.

149. (a) $[H^+] = \sqrt{Kc} = \sqrt{10^{-5} \times 0.1} = 10^{-3}, pH = 3$.

150. (d) In $\frac{N}{10} NaOH$ have $[OH^-] = 10^{-1}M$ means $pOH = 1$ and then $pH + pOH = 14$

$pH = 14 - pOH = 13$.

151. (b) Borate ions are hydrolyzed to develop alkaline nature in solution.

152. (d) Less the pH, more acidic is the solution.

153. (b) Acidic

Explanation (Simple & Word-Friendly):

pH = 7 → Neutral (e.g., pure water)

pH > 7 → Basic (e.g., NaOH solution)

pH < 7 → Acidic (e.g., vinegar, lemon juice)

Since pH = 6, it is less than 7, so the solution is slightly acidic.

154. (b) The equal conc. of salt and acid.

155. (c) $pH = -\log \quad K_a + \log \frac{[KCN]}{[HCN]}$

$pH = -\log[5 \times 10^{-10}] + \log\left(\frac{0.15}{1.5}\right) = 8.302$

156. (d) $NH_4OH + NH_4Cl$

Explanation (Simple & Word-Friendly):

Type of Buffer	Example	pH
Acidic buffer	Weak acid + its salt	pH < 7
Basic buffer	Weak base + its salt	pH > 7

(a) $CH_3COOH + CH_3COONa$ → weak acid + salt → acidic buffer

(b) HCOOH + HCOOK → weak acid + salt → acidic buffer

(c) CH_3COONH_4 → not a buffer

(d) $NH_4OH + NH_4Cl$ → weak base + its salt → basic buffer → pH > 7

157. (c) $pH = pK_a + \log\frac{[\text{Salt}]}{[\text{Acid}]}$ equimolar means

$\frac{[\text{Salt}]}{[\text{Acid}]} = 1;\ pH = 4.74 + 0 = 4.74$

158. (a) Because of $NaCl$ is a salt of strong acid and strong base. So that it is neutral.

159. (c) When strong acid and strong base are react neutral salt are formed. So that $NaCl$ is a neutral salt.

160. (b) Addition of 25 mL of 0.02 M HCl

Explanation:

pH decreases when the **concentration of H^+ increases**.

Adding **0.02 M HCl** increases the H^+ concentration → lowers the pH.

Adding 0.005 M HCl **dilutes the solution slightly** → pH increases slightly.

Mg reacts with HCl, but this does not decrease pH significantly in a buffered way.