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

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

$\boldsymbol{pH = 1.7}$ *i.e.* in between 1 and 2.

62. (c) Basic

Word-friendly explanation (fast):
Molar mass of $Na_2CO_3 \cdot H_2O \approx 124\ g\ mol^{-1}$.
Moles added = 0.62 / 124 = 0.005 mol (this gives 0.005 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 M.
Moles of $(NH_4)_2SO_4$ = 0.05 × 0.100 L = 0.005 mol → 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 ⇒ exactly enough NH_4^+ to neutralize CO_3^{2-}.
Net reaction converts $CO_3^{2-} + 2\ NH_4^+ \rightarrow H_2CO_3 + 2\ NH_3$.
H_2CO_3 partly decomposes to $CO_2 + H_2O$, while NH_3 (a weak base) remains in solution.
Amount of NH_3 produced ≈ 0.01 mol in ≈ 0.1 L → $[NH_3] \approx 0.1$ 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 → $[OH^-] = 10^{-3}$ M.
pOH = 3 → pH = 14 − 3 = 11.
HCl (10^{-3} M): strong acid → $[H^+] = 10^{-3}$ M.
pH = 3.
NaCl (10^{-3} M): neutral salt (from strong acid + strong base) → solution is essentially neutral (pH ≈ 7) at this dilution (neglecting very small ionic effects).

65. (c) NaOH $\rightleftharpoons Na^{+} + OH^{-}$

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

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

66. (a) +1

Explanation (Word-friendly):
Let the acid be:
H2A (dibasic acid)
It gives 2 H+ ions when fully ionized.
Concentration of acid = 0.05 M
So, total H+ ions produced = 2 × 0.05 = 0.10 M
Now,
pH = –log (H+)
pH = –log (0.10) = 1

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

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

68. (d) A weak acid and its salt with a strong base maintain pH 4 – 5

69. (a) $NaOH \rightleftharpoons Na^{+} + OH^{-} = [OH^{-}] = 10^{-8} M.$

70. (c) $[OH^{-}] = 0.0001N, pOH = 4, pH + pOH = 14$

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

71. (d) 0.001 M KOH solution

$[OH^{-}] = 0.001M = 1 \times 10^{-3} M$
$\boldsymbol{[H^{+}] \times [OH^{-}] = 1 \times 10^{-14}}$

$[H^+] = \frac{1\times10^{-14}}{[OH^-]}$

$[H^+] = \frac{1\times10^{-14}}{1\times10^{-3}} = 1\times10^{-14}\times10^{+3}$

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

$pH = 11$

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.* $CH_3COOH + CH_3COONa$

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

$K_a = 1.8\times10^{-5}$; $pH = -\log K_a + \log\frac{[Salt]}{[Acid]}$

$= -\log1.8\times10^{-5} + \log\frac{0.1}{0.1} = -\log1.8\times10^{-5}$

$pH = 4.7$.

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

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

$pOH = 7$

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

78. (c) 0.01 $M\,Ba(OH)_2 = 0.02N Ba(OH)_2$

$N_1V_1 = N_2V_2$

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

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

$\mathbf{pOH} = \mathbf{2}$ or $pH = 12$

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

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