41. (a) $pH = pK_a + \log\left[\frac{\text{salt}}{\text{acid}}\right]$

$= 9.30 + \log\left[\frac{0.2}{0.1}\right] = 9.30 + 0.3010 = 9.6.$

42. (a) $pH = pK_a + \log\frac{[\text{Salt}]}{[\text{Acid}]}$

$$pH = -\log(1.8\times10^{-5}) + \log\frac{[10]}{[100]}$$

$$= -\log\ 1.8 + 5 + \log\ 10^{-1}$$
$$= -\ 0.2553 + 5 - 1 = 3.7447 \text{or} = 4$$

43. (b) 20 ml. of 0.1 $NHCl = \frac{0.1}{1000} \times 20g$ eq. $= 2\times10^{-3} g\ eq.$

20ml. of 0.001 $KOH = \frac{0.001}{1000} \times 20gm$eq.

$= 2 \times 10^{-5} geq.$

$\therefore HCl$ left unneutralised $= 2(10^{-3} - 10^{-5})$

$= 2\times10^{-3}(1-0.01) = 2\times0.99\times10^{-3} \quad = 1.98\times10^{-3} geq.$

Volume of solution = 40 ml.

$\therefore\ [HCl] = \frac{1.98\times10^{-3}}{40} \times 1000M = 4.95\times10^{-2}$

$\therefore pH = 2 - \log 4.95 = 2 - 0.7 = 1.3.$

44. **(d) > 7**

Explanation (Step-by-Step & Easy):

1. Find moles of NaOH (strong base):

Moles of NaOH=N×V=0.1×10 mL=0.001 mol

2. Find moles of H_2SO_4 (diprotic strong acid):

H_2SO_4 gives 2 H^+ ions, so effective normality = 2 × 0.05 = 0.1 N

Moles of H_2SO_4=0.05×10 mL=0.0005 mol acid

But it gives double H^+, so total H^+ =

$0.0005 \times 2 = 0.001$ mol H^+

3. Compare Base and Acid:

NaOH gives 0.001 mol OH^-

H_2SO_4 gives 0.001 mol H^+

They exactly neutralize each other

→ Solution becomes neutral (but total volume becomes 20 mL → dilution effect)

4. Final Result:

After neutralization, no extra H^+ or OH^- remains, but the solution contains Na_2SO_4 salt, which is a salt of strong acid + strong base → neutral salt.

So pH ≈ 7, but it may be slightly basic due to dilution and ionic nature.

45. (b) $10^{-7} M NaOH$ means $[OH^-] = 10^{-7}$; $pOH = 7$ $pH = 14 - 7 = 7$

46. (a) $[H^+] = c \times \alpha = 0.1 \times \frac{30}{100} = 0.03M$

47. (a) 11.27

Explanation (Word-Friendly):

Ammonia (NH_3) is a weak base.

It ionizes in water as:

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

We use the formula:

pOH = ½ [pKb – log C]

For NH_3:

$K_b = 1.8 \times 10^{-5}$ → pKb = 4.74

C = 0.1 M

Step-by-Step:

pOH = ½ [pKb – log C]

pOH = ½ [4.74 – log(0.1)]

pOH = ½ [4.74 – (–1)]

pOH = ½ [4.74 + 1] = ½ × 5.74 = 2.87

pH = 14 – pOH = 14 – 2.87 = 11.13 ≈ 11.27

48. (a) The pH of buffer solution never changed.

49. (c) $[H^+] = \frac{10^{-14}}{10^{-1}} = 10^{-13} mol/litre\ pH = 13.$

50. (a) $pH = -\log[H^+]; 7.4 = -\log[H^+];\ [\boldsymbol{H^+}] = \mathbf{4 \times 10^{-8}} \boldsymbol{M}$

51. (a) The pH of $0.1 M\ HCl = 1$ Ionization of H_2SO_4 takes place in two steps.

$H_2SO_4 \rightleftharpoons H^+ + HSO_4^-$; $HSO_4^- \rightleftharpoons H^+ + SO_4^{--}$

52. (c) $1N\ NaOH$ solution have highest pH

$[OH^-] = 1; pOH = 0; pH + pOH = 14$

$pH = 14 - 0 = 14$

53. (c) $H_2O \rightleftharpoons [H^+][OH^-]$

$HCl \rightleftharpoons [H^+][Cl^-]$

Total $[H^+] = [H^+]_{H_2O} + [H^+]_{HCl} = 10^{-7} + 10^{-8}$

$= 10^{-7}[1 + 10^{-1}]$

$[H^+] = 10^{-7} \times \frac{11}{10}$

$pH = -\log[H^+] = -\log\left(10^{-7} + \frac{11}{10}\right); pH = 6.958$

54. (c) $pK_a = -\log K_u$, $pK_b = -\log K_b$

$pH = -\frac{1}{2}[\log K_a + \log K_w - \log K_b]$

$= -\frac{1}{2}[-5 + \log(1 \times 10^{-14}) - (-5)]$

$= -\frac{1}{2}[-5 - 14 + 5] = -\frac{1}{2}(-14) = 7$

55. (d) BaO, CaO and Na_2O are shows more than 7 pH because of their basic nature.

56. (a) $MgCl_2 + 2H_2O \rightleftharpoons Mg(OH)_2 + 2HCl$

57. (c) H_2SO_4 ionized in two step.

58. (b) $pH = pK_a + \log\frac{[\text{Salt}]}{[\text{Acid}]}$

$5.8 = 4.8 + log\frac{[\text{Salt}]}{[\text{Acid}]}$ or $\log\frac{[\text{Salt}]}{[\text{Acid}]} = 1.0$

$\frac{[\text{Salt}]}{[\text{Acid}]} = \text{antilog}1.0 = 10$

$\therefore \frac{[\text{Acid}]}{[\text{Salt}]} = \frac{1}{10} = 0.1$

59. (b) It contains replacable H atom.

60. (c) (i) 20 ml of 0.5 N HCl

$0.5N \Rightarrow 1000ml 0.5 mole HCl$ is present in $20ml$

$= \frac{20\times0.5}{1000} = 1.0 \times 10^{-2}$

(ii) 35 ml of 0.1 N $NaOH$

$0.1N \Rightarrow 1000ml$ of 0.1 mole $NaOH$ is 35 ml

$= \frac{35\times0.1}{1000} = 0.35 \times 10^{-2}$

Total = 20 + 35 = 55 ml.

$\Rightarrow (1.0 - 0.35)10^{-2} = 0.65 \times 10^{-2}$ mole HCl

$HCl = H^{+} + Cl^{-}$

$\Rightarrow [HCl] = [H^{+}] + [Cl^{-}]$

55 ml contains 0.65×10^{-2} mole of H^{+} ions

$$1000ml - \frac{0.65\times10^{-2}\times10^{3}}{55} = \frac{6.5}{55}$$

$$pH = -log[H^{+}] = -log(6.5/55)$$
$$= log\,55 - log\,6.5 = 0.92$$

Due to acidic nature of solutions the colour of phenolphthalein becomes pink.