

42. (a) } all lead to neutral solutions, so pH 7
 (b) }
 (c) }
 (d) }

(e) 0.1 L of 0.1 M HNO_3 diluted to 1 L

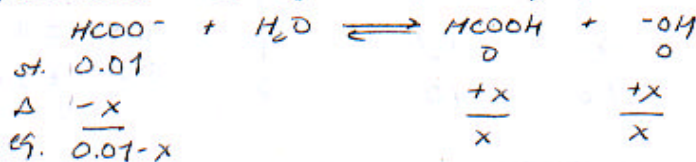
$$[\text{H}_3\text{O}^+] = \frac{0.01 \text{ mole}}{1 \text{ L}} = 1 \times 10^{-2} \text{ M}$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log 10^{-2} = 2$$

44. (a) all solutions contain

$$0.1 \text{ L} \times 0.1 \frac{\text{moles HCOOH}}{\text{L}} = 0.01 \text{ moles HCOOH}$$

(a) all HCOOH is consumed by OH^- to give 0.01 moles HCOO^- .
 pH is determined by the following equilibrium:



$$K_b = \frac{K_w}{K_a} = \frac{1 \times 10^{-14}}{1.8 \times 10^{-4}} = \frac{[\text{HCOOH}][\text{OH}^-]}{[\text{HCOO}^-]} = \frac{x^2}{0.01-x} \approx \frac{x^2}{0.01}$$

$$x = [\text{OH}^-] = 7.47 \times 10^{-7} \text{ M}$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log \frac{K_w}{7.47 \times 10^{-7}} = 7.87$$

(b) same as above: pH 7.87

(c) buffer! $[\text{HCOO}^-] = [\text{HCOOH}] = 0.0050 \text{ M}$

$$\text{pH} = \text{p}K_a + \log \frac{0.005}{0.005} = -\log(1.8 \times 10^{-4}) + \log 1 = 3.74$$

(d) same as above: 3.74

(e) pH due to strong acid: $0.1 \frac{\text{moles}}{\text{L}} \times 0.1 \text{ L} = 0.01 \text{ moles HCl}$

$$[\text{HCl}] = [\text{H}_3\text{O}^+] = 0.01 \quad \text{pH} = -\log [\text{H}_3\text{O}^+] = 2$$