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

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

$[H_3O^+] = \frac{0.01 \text{ mole}}{1 L} = 1 \times 10^{-2} M$

$pH = -\log [H_3O^+] = -\log 10^{-2} = 2$

44. all solutions contain

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

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

	$HCOO^-$ + H_2O	$\rightleftharpoons$	$HCOOH$	+	^-OH
st.	0.01		0		0
Δ	$-x$		$+x$		$+x$
eq.	$0.01-x$		x		x

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

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

$pH = -\log [H_3O^+] = -\log \frac{K_w}{7.47 \times 10^{-7}} = 7.87$

(b) same as above: pH 7.87

(c) buffer! $[HCOO^-] = [HCOOH] = 0.0050 M$

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

(d) same as above: 3.74

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

$[HCl] = [H_3O^+] = 0.01 \qquad pH = -\log [H_3O^+] = 2$