17.7a) Since both acids have one "jump" in pH, they are both monoprotic. Given that each acid has the same volume, the more concentrated acid will require more base to neutralize it. As such, the "red" acid is more concentrated.

b) To determine the Ka values of each acid it is helpful to remember that when half the acid has been converted into conjugate base the concentration of acid and conjugate base will be equal. At these conditions, pH = pKa. Relative to the titration curve, these conditions exist at a titrant volume half way to equivalence. For the blue acid, the half titration point is about pH = 7. As such, the pKa is 7 and the K_a is about 1×10^{-7}. For the red acid, the half titration point is about pH = 4. As such, the pKa is 4 and the K_a is about 1×10^{-4}. This would indicate that the red acid has a larger K_a. A good general principle is stronger acids have "lower" curves before equivalence.

17.45a) $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

	CH_3COOH	CH_3COO^-	H^+
I	0.150M	0 M	0 M
C	-x	+x	+x
E	0.150 - x	x	x

$$K_a = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]} \qquad 1.8 \times 10^{-5} = \frac{x^2}{0.150 - x} \approx \frac{x^2}{0.150}$$

$x = 0.00164M \qquad pH = -\log[H_3O^+] = -\log[0.00164M] = 2.78$

17.45c) $35.0\text{mL} \times \frac{1\text{ L}}{1000\text{ mL}} \times \frac{0.150\text{ mol } CH_3COOH}{1\text{ L } CH_3COOH} = 0.00525\text{ mol } CH_3COOH$

$34.5\text{mL} \times \frac{1\text{ L}}{1000\text{ mL}} \times \frac{0.150\text{ mol NaOH}}{1\text{ L NaOH}} = 0.00518\text{ mol NaOH}$

	CH_3COOH	$\rightleftharpoons$	CH_3COO^-	+	H^+
I	0.00525 mol		0 mol		0 mol
R	-0.00518		+0.00518		0 mol
	0.00007mol/0.0695L		0.00518/0.0695L		0 mol
A	0.001007M		0.0745M		0M
C	-x		+x		+x
E	0.001007 - x		0.0745 + x		x

$$K_a = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]} \quad 1.8 \times 10^{-5} = \frac{(0.0745+x)x}{0.001007 - x} \approx \frac{0.0745x}{0.001007}$$

$x = 2.4 \times 10^{-7}$ $\quad pH = -\log[H^+] \quad pH = -\log[2.4 \times 10^{-7}] \quad pH = 6.62$

17.45d)

$$\frac{35.0\,\text{mL}}{} \left| \frac{1\text{ L}}{1000\text{ mL}} \right| \frac{0.150\text{ mol } CH_3COOH}{1\text{ L } CH_3COOH} = 0.00525\text{ mol } CH_3COOH$$

$$\frac{35.0\,\text{mL}}{} \left| \frac{1\text{ L}}{1000\text{ mL}} \right| \frac{0.150\text{ mol NaOH}}{1\text{ L NaOH}} = 0.00525\text{ mol NaOH}$$

$$M = \frac{\text{moles}}{\text{liters}} = \frac{0.00525\text{ mol } CH_3COO^-}{0.0700\text{ L}} = 0.075\, M\ CH_3COO^-$$

	CH_3COO^-	+	H_2O	$\rightleftharpoons$	CH_3COOH	+	OH^-
I	0.075M		———		0 M		0 M
C	-x		———		+x		+x
E	0.075 - x		———		x		x

$K_b = K_w \div K_a$ $K_b = 1.0\text{x}10^{-14} \div 1.8\text{x}10^{-5}$ $K_b = 5.6\text{ x}10^{-10}$

$$K_b = \frac{[CH_3COOH][OH^-]}{[CH_3COO^-]} \quad 5.6 \times 10^{-10} = \frac{x^2}{0.075 - x} \approx \frac{x^2}{0.075}$$

$x = 6.48 \times 10^{-6}$

$pOH = -\log[OH^-]$ $pOH = -\log[6.48 \times 10^{-6}]$ $pOH = 5.19$

$pH + pOH = 14$ $pH = 14 - pOH$ $pH = 14 - 5.19 = 8.81$

17.45e)

$$\frac{35.0\text{mL}}{} \left| \frac{1\text{ L}}{1000\text{ mL}} \right| \frac{0.150\text{ mol } CH_3COOH}{1\text{ L } CH_3COOH} = 0.00525\text{ mol } CH_3COOH$$

$$\frac{35.5\text{mL}}{} \left| \frac{1\text{ L}}{1000\text{ mL}} \right| \frac{0.150\text{ mol NaOH}}{1\text{ L NaOH}} = 0.00533\text{ mol NaOH}$$

$$\overline{0.00008\text{ mol NaOH}}$$

$$M = \frac{\text{moles}}{\text{liters}} = \frac{0.00008\text{ mol NaOH}}{0.0705\text{ L}} = 0.00113\ M\text{ NaOH}$$

$$\text{pOH} = -\log[OH^-] = -\log[0.00113\ M] = 2.95$$

$$\text{pH} = 14 - \text{pOH} = 14 - 2.95 = 11.05$$

17.46a)

	NH_3	+	H_2O	$\rightleftharpoons$	NH_4^+	+	OH^-
I	0.050M		———		0 M		0 M
C	-x		———		+x		+x
E	0.050 - x		———		x		x

$$K_b = \frac{[NH_4^+][OH^-]}{NH_3} \qquad 1.8 \times 10^{-5} = \frac{x^2}{0.050 - x} \approx \frac{x^2}{0.050}$$

$x = 9.5 \times 10^{-4}$

$pOH = -\log[OH^-]$ $\quad$ $pOH = -\log[9.5 \times 10^{-4}]$ $\quad$ $pOH = 3.02$

$pH + pOH = 14$ $\quad$ $pH = 14 - pOH$ $\quad$ $pH = 14 - 3.02 = 10.98$

17.46c)

$$\frac{30.0\text{mL}}{} \left| \frac{1\text{ L}}{1000\text{ mL}} \right| \frac{0.050\text{ mol }NH_3}{1\text{ L }NH_3} = 0.00150\text{ mol }NH_3$$

$$\frac{59.0\text{mL}}{} \left| \frac{1\text{ L}}{1000\text{ mL}} \right| \frac{0.025\text{ mol HCl}}{1\text{ L HCl}} = 0.00148\text{ mol HCl}$$

	NH_3	+	H_2O	$\rightleftharpoons$	NH_4^+	+	OH^-
I	0.00150 mol		———		0 mol		0 mol
R	-0.00148		———		+ 0.00148		0 mol
	0.00002mol/0.089L		———		0.00148mol/0.089L		0 mol
A	$2.25 \times 10^{-4} M$		———		$1.67 \times 10^{-2} M$		$0\ M$
C	-x		———		+x		+x
E	2.24×10^{-4} - x		———		1.67×10^{-2} + x		+x

$$K_b = \frac{[NH_4^+][OH^-]}{NH_3} \qquad 1.8 \times 10^{-5} = \frac{(1.67\times10^{-2} + x)x}{2.24 \times 10^{-4} - x} \approx \frac{(1.67\times10^{-2})x}{2.24 \times 10^{-4}}$$

$x = 2.4 \times 10^{-7}$

$pOH = -\log[OH^-]$ $\quad$ $pOH = -\log[2.4 \times 10^{-7}]$ $\quad$ $pOH = 6.62$

$pH + pOH = 14$ $\quad$ $pH = 14 - pOH$ $\quad$ $pH = 14 - 6.62 = 7.38$

17.46d) $\frac{30.0\text{mL}}{} \left| \frac{1\text{ L}}{1000\text{ mL}} \right| \frac{0.050\text{ mol } NH_3}{1\text{ L } NH_3} = 0.00150\text{ mol } NH_3$

$$\frac{60.0\text{mL}}{} \left| \frac{1\text{ L}}{1000\text{ mL}} \right| \frac{0.025\text{ mol HCl}}{1\text{ L HCl}} = 0.00150\text{ mol HCl}$$

$$M = \frac{\text{moles}}{\text{liters}} = \frac{0.00150\text{ mol } NH_4^+}{0.0900\text{ L}} = 0.0167\ M\ NH_4^+$$

	NH_4^+	$\rightleftharpoons$	NH_3	+	H^+
I	0.0167M		0 M		0 M
C	-x		+x		+x
E	0.0167 - x		x		x

$K_a = K_w \div K_b$ $K_a = 1.0\text{x}10^{-14} \div 1.8\text{x}10^{-5}$ $K_a = 5.6\text{ x}10^{-10}$

$$K_a = \frac{[NH_3][H^+]}{[NH_4^+]} \quad 5.6 \text{ x } 10^{-10} = \frac{x^2}{0.0167 - x} \approx \frac{x^2}{0.0167}$$

$x = 3.06 \text{ x } 10^{-6} M$ $pH = -\log[H_3O^+] = -\log[3.06 \text{ x } 10^{-6} M] = 5.51$

17.46e)

$$\frac{30.0\text{mL}}{} \left| \frac{1\text{ L}}{1000\text{ mL}} \right| \frac{0.050\text{ mol }NH_3}{1\text{ L }NH_3} = 0.00150\text{ mol }NH_3$$

$$\frac{61.0\text{mL}}{} \left| \frac{1\text{ L}}{1000\text{ mL}} \right| \frac{0.025\text{ mol HCl}}{1\text{ L HCl}} = 0.00153\text{ mol HCl}$$

$$= 0.00003\text{ mol HCl}$$

$$M = \frac{\text{moles}}{\text{liters}} = \frac{0.00003\text{ mol HCl}}{0.0910\text{ L}} = 0.000330\ M\ H^+$$

$$pH = -\log[H_3O^+] = -\log[0.000330\ M] = 3.48$$

17.47a) When NaOH is titrated with HBr, the resulting solution at equivalence contains NaBr, a neutral salt. As such, the pH at equivalence is 7.

17.47b) Because the concentrations of the acid and base are the same, equal volumes of each are needed to reach equivalence. The resulting doubling in volume will decrease the concentration of the resulting salt solutions by half.

	NH_3OH^+	$\rightleftharpoons$	NH_2OH	+	H^+
I	0.100M		0 M		0 M
C	-x		+x		+x
E	0.100 - x		x		x

$K_a = K_w \div K_b \quad K_a = 1.0x10^{-14} \div 1.1x10^{-8} \quad K_a = 9.1\ x10^{-7}$

$$K_a = \frac{[NH_2OH][H^+]}{[NH_3OH^+]} \qquad 9.1\ x10^{-7} = \frac{x^2}{0.100 - x} \approx \frac{x^2}{0.100}$$

$x = 3.02\ x\ 10^{-4} M \quad pH = -log[H_3O^+] = -log[3.02\ x\ 10^{-4} M] = 3.52$