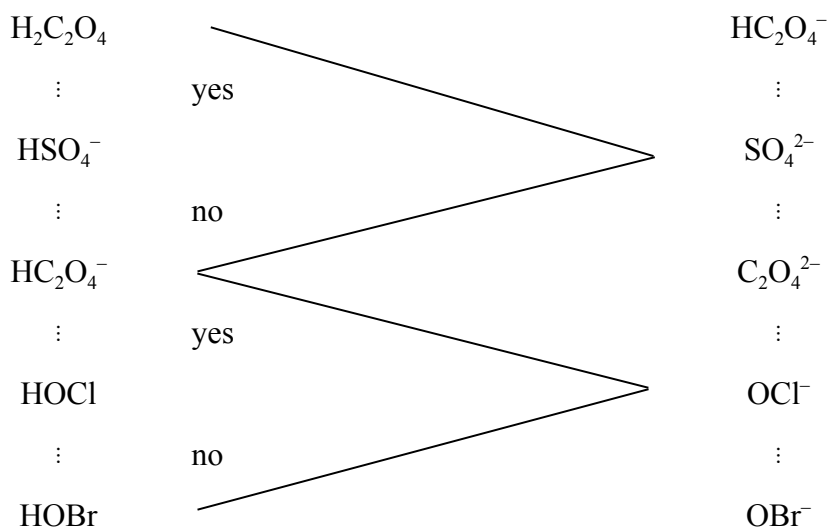


Chem 104
Test 3 Practice Problems
Solutions

1.

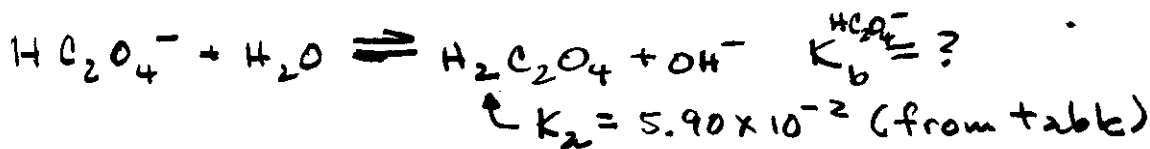
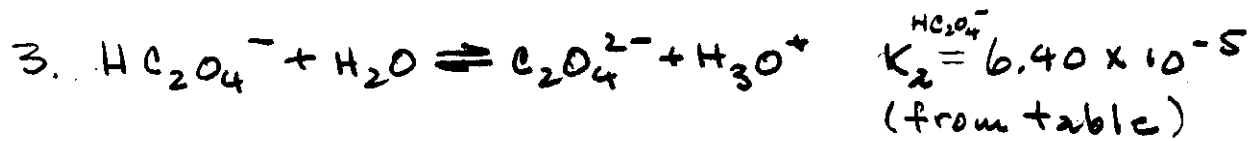
$[\text{H}_3\text{O}^+]$	$[\text{OH}^-]$	pH	pOH	acidic or basic?
$5.0 \times 10^{-8} \text{ M}$	$2.0 \times 10^{-7} \text{ M}$	7.30	6.70	basic

2. The stronger acid reacts with the stronger base to form weaker conjugates. From the table, reactions that lie more than 50% to the right are those between acids higher up on the left column with bases lower down on the right column. The relative placements of the conjugate pairs are shown below with the combinations that lie right (“yes”) and those that lie left (“no”).

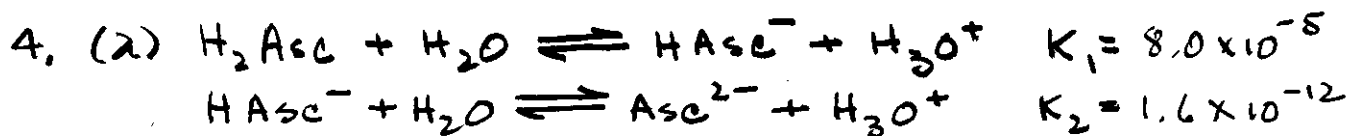
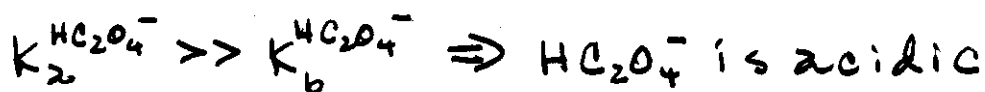


Conclusions:

$\text{H}_2\text{C}_2\text{O}_4(aq) + \text{SO}_4^{2-}(aq) \rightleftharpoons \text{HC}_2\text{O}_4^-(aq) + \text{HSO}_4^-(aq)$	right
$\text{HC}_2\text{O}_4^-(aq) + \text{SO}_4^{2-}(aq) \rightleftharpoons \text{C}_2\text{O}_4^{2-}(aq) + \text{HSO}_4^-(aq)$	left
$\text{HC}_2\text{O}_4^-(aq) + \text{OCl}^-(aq) \rightleftharpoons \text{C}_2\text{O}_4^{2-}(aq) + \text{HOCl}(aq)$	right
$\text{HOBr}(aq) + \text{OCl}^-(aq) \rightleftharpoons \text{OBr}^-(aq) + \text{HOCl}(aq)$	left



$$K_b^{\text{HC}_2\text{O}_4^-} = \frac{K_w}{K_2^{\text{H}_2\text{C}_2\text{O}_4}} = \frac{1.00 \times 10^{-14}}{5.90 \times 10^{-2}} = 1.69 \times 10^{-13}$$



(b) Assume first dissociation is the only important source of H_3O^+ in solution; i.e., $[\text{H}_3\text{O}^+] = [\text{HAsc}^-]$ (Assumption I). $C_{\text{H}_2\text{A}} \gg K_1$, so assume $[\text{H}_2\text{Asc}] = C_{\text{H}_2\text{Asc}}$ (Assumption II).

$$[\text{H}_3\text{O}^+] = \sqrt{(0.10)(8.0 \times 10^{-5})} = 2.83 \times 10^{-3} \text{ M} = [\text{HAsc}^-]$$

$$\text{pH} = 2.55$$

$$\% \text{ dissociation} = \frac{2.83 \times 10^{-3}}{0.10} \times 100\% = 2.8\%$$

(c) Since $[\text{H}_3\text{O}^+] = [\text{HAsc}^-]$, in K_2 we get

$$K_2 = \frac{[\text{H}_3\text{O}^+][\text{Asc}^{2-}]}{[\text{HAsc}^-]} = 1.6 \times 10^{-12} \text{ M} = [\text{Asc}^{2-}]$$

5. (a) Equivalence point:

$$V_t = \frac{M_a V_a}{M_t} = \frac{(0.120M)(25.0mL)}{0.100M} = 30.0mL$$

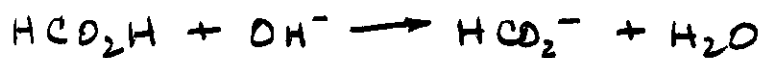
(b) mmol $HCO_2H = (0.120M)(25.0mL) = 3.00mmol$

(c) Initial pH

$$[H_3O^+] = \sqrt{(0.120)(1.9 \times 10^{-4})} = 4.775 \times 10^{-3} = 4.8 \times 10^{-3}M$$

$$pH = 2.32$$

(d) After 5.00mL \times 0.100M $NaOH = 0.500mmol OH^-$



Add 3.00 0.500

Get 2.50 ~0 0.50

$$K_a = 1.9 \times 10^{-4} = \frac{[H_3O^+](0.50)}{2.50}$$

$$[H_3O^+] = 9.5 \times 10^{-4}M$$

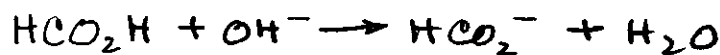
$$pH = 3.02$$

(e) After 15.0mL \times 0.100M $NaOH = 1.50mmol OH^-$
 \Rightarrow Half-titration point, where $pH = pK_a$.

$$pH = -\log(1.9 \times 10^{-4}) = 3.72$$

$$(f) \text{ Equiv. point} = 30.0 \text{ mL} \times 0.100 \text{ M NaOH} \\ = 3.00 \text{ mmol OH}^-$$

$$V = 25.0 \text{ mL} + 30.0 \text{ mL} = 55.0 \text{ mL}$$



$$\text{Add} \quad 3.00 \quad 3.00$$

$$\text{Get} \quad \sim 0 \quad \sim 0 \quad 3.00$$

$$C_{\text{HCO}_2^-} = 3.00 \text{ mmol} / 55.0 \text{ mL} = 0.054545... \text{ M}$$

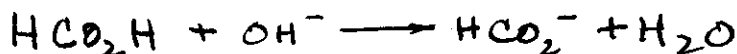
$$K_b^{\text{HCO}_2^-} = \frac{1.00 \times 10^{-14}}{1.9 \times 10^{-4}} = 5.26 \times 10^{-11}$$

$$[\text{OH}^-] = \sqrt{C_{\text{HCO}_2^-} K_b^{\text{HCO}_2^-}} = \sqrt{(0.0545)(5.26 \times 10^{-11})}$$

$$= \sqrt{2.87 \times 10^{-12}} = 1.7 \times 10^{-6}$$

$$\text{pOH} = 5.77 \Rightarrow \text{pH} = 8.23$$

$$(g) 5.00 \text{ mL excess NaOH} = 35.0 \text{ mL NaOH total} \\ 35.0 \text{ mL} \times 0.100 \text{ M NaOH} = 3.50 \text{ mmol OH}^-$$



$$\text{Add} \quad 3.00 \quad 3.50$$

$$\text{Get} \quad \sim 0 \quad 0.50 \quad 3.00$$

$$V = 25.0 \text{ mL} + 35.0 \text{ mL} = 60.0 \text{ mL}$$

$$[\text{OH}^-] = \frac{0.50 \text{ mmol}}{60.0 \text{ mL}} = 8.3 \times 10^{-3}$$

$$\text{pOH} = 2.08 \Rightarrow \text{pH} = 11.92$$