

Chem 103 Test 1 key version B

<p>Part I</p> <p>1) A</p> <p>2) B ($\text{HNO}_3 + \text{CH}_3\text{COONa}$), ($\text{NaOH} + \text{CH}_3\text{COOH}$), ($\text{CH}_3\text{COONa} + \text{CH}_3\text{COOH}$)</p> <p>3) A</p> <p>4) B</p> <p>5) B $\frac{540(100\%)}{540 + 580} = 48.2\%$</p> <p>6) C (pH transition at equivalence much higher than $\text{pK}_a = 4.7$, phenolphthalein starts to change color at equiv. pt of HAc).</p> <p>7) D</p>	<p>8) E</p> <p>9) C $[\text{OH}^-]=1.2\text{M}$; $[\text{H}^+]=10^{-14.00}/1.2=8.3\times 10^{-15}$</p> <p>10) D $\frac{\pi_{\text{Mg}(\text{NO}_3)_2}}{\pi_{\text{NaNO}_3}} = \frac{i'\text{MRT}}{i\text{MRT}} = \frac{i'}{i} = \frac{3}{2} = 1.5$</p> <p>11) C</p> <p>12) C</p> <p>13) C (the more polar the O-H bond, the easier to lose the H^+; related to electronegativity).</p> <p>14) A (comes from $\text{HSO}_4^- \rightleftharpoons \text{H}^+ + \text{SO}_4^{2-}$)</p> <p>15) D</p> <p>16) D</p>
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<p>Part II</p> <p>17) There are several ways to solve this problem. One way is outlined below:</p> <p>Let the unknown concentration = x. Then $[\text{H}_3\text{O}^+]$ for the $\text{HCl}(\text{aq})$ solution = x, and hence $[\text{H}_3\text{O}^+]$ for the acetic acid solution is x/100. But for the acetic acid solution, we also have</p> $\text{HA} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{A}^-$ $x \qquad \qquad \qquad [\text{H}_3\text{O}^+] \quad [\text{H}_3\text{O}^+]$ $K_a = [\text{H}_3\text{O}^+]^2/x$ $[\text{H}_3\text{O}^+]_{\text{HA}} = \sqrt{(x \cdot K_a)}$ <p>Therefore: $[\text{H}_3\text{O}^+]_{\text{HCl}} = x = [\text{H}_3\text{O}^+]_{\text{HA}} \cdot 100$</p> $(x/100) = \sqrt{(x \cdot K_a)}$ $\Rightarrow x^2/10^4 = x \cdot K_a$ $\Rightarrow x = 10^4 \cdot K_a = 0.19 \text{ M}$ <p>Check: For $\text{HCl}(\text{aq})$, $[\text{H}_3\text{O}^+] = 0.19 \text{ M}$ $\Rightarrow \text{pH} = 0.72$.</p> <p>For hydrazoic acid, $[\text{H}_3\text{O}^+] = \sqrt{(0.19 \times 1.9 \times 10^{-5})} = 1.9 \times 10^{-3} \Rightarrow \text{pH} = 2.72$.</p>	<p>18) $\Pi = i\text{MRT}$; use $i=1$,</p> $M = \frac{\Pi}{RT} = \frac{(0.0203\text{atm})}{(0.0821\text{atmL/molK})(300\text{K})} = 8.24 \times 10^{-4} \text{ M}$ $M = \frac{3.6\text{g}}{0.200\text{L}} \Rightarrow \text{MW} = \frac{3.6\text{g}}{0.200\text{L}} = 2.18 \times 10^4 \text{ g/mol}$ <p>19) Lewis acid = electron pair acceptor. Cu^{2+} ion is an example. Most metal cations have empty shells that can be filled by electron pair donors like ligands.</p> <p>Lewis base = electron pair donor. H_2O and NH_3 are examples. Like other Lewis bases, there are available lone electron pairs that can be donated to form a coordinate covalent bond with a Lewis acid, like Cu^{2+} ion.</p> <p>In this demonstration, the Cu^{2+} ion acts as the Lewis acid. Adding water to the solid CuSO_4 dissociates the Cu^{2+} ion from SO_4^{2-}. Water molecules donate their lone e pairs to the Cu^{2+} to form a bluish complex of Cu^{2+} and water, $[\text{Cu}(\text{H}_2\text{O})_5]^{2+}$. Adding NH_3 to the solution immediately displaces the water from the Cu^{2+} and the resulting $\text{Cu}(\text{NH}_3)_4^{2+}$ complex is the dark blue complex formed. It shows that NH_3 is a stronger Lewis base than water.</p>
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<p>20) HA \rightleftharpoons H⁺ + A⁻</p> <p>0.35M-x x x</p> $\frac{x^2}{0.35-x} = 4.2 \times 10^{-7}; x = \sqrt{(0.35)(4.2 \times 10^{-7})} =$ <p>x = 3.83x10⁻⁴ => pH=-logx=3.42</p>	<p>21) D A E B C</p> <p>pH of D \approx 7 (extremely dilute); pH of A = -log(4.2 \times 10⁻⁶)=5.38, pH of E = 3.5, pH of B = 2.65; pH of C = -log(0.25)=0.60</p>
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Part III

22)

Solution:

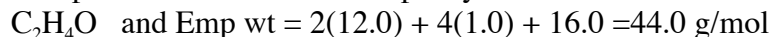
a) get emp formula: assume 100g

mol C = 54.55gC (1molC/12.0g) = 4.55 mol C

mol H = 9.09gH(1mol H/1.0g) = 9.09 mol H

mol O = (100-54.55-9.09)(1mol/16.0g)=2.27molO

tentative emp formula: divide subscripts by 2.27:

by FP depression: $\Delta T_f = K_f m = K_f(g/MW)/(0.060 \text{ kg})$

$$\Rightarrow (6.55 - (-1.45)) = (20.) (2.11 / MW) / 0.060 \text{ kg} \Rightarrow 8.00 = 703 / MW$$

$$\Rightarrow MW = 703 / 8.00 = 87.9 \text{ g/mol}$$

according to the % composition, the EW = 44.0 so the actual MW is double 44.0 or 88.0g/mol. the molecular formula is C₄H₈O₂.

b) The bp of solution is: $\Delta T_b = k_b m = (2.79)((2.11/88.0) / 0.060 \text{ kg}) = 2.79(0.400) = 1.11^\circ\text{C}$
 $T_b = 80.74 + 1.11 = 81.85^\circ\text{C}$

23) a) $\text{OH}^- + \text{H}_2\text{SeO}_3 \rightarrow \text{H}_2\text{O} + \text{HSeO}_3^-$: $V_e = \frac{M_{\text{H}_2\text{A}} V_{\text{H}_2\text{A}}}{M_{\text{OH}}} = \frac{(.200\text{M})(20.0\text{mL})}{(0.100\text{M})} = 40.0\text{mL}$

b) $\text{OH}^- + \text{HSeO}_3^- \rightarrow \text{H}_2\text{O} + \text{SeO}_3^{2-}$: Note: $n_{\text{HSeO}_3^-} = n_{\text{H}_2\text{SeO}_4} + n_{\text{HSeO}_4} = (.150 + .200)(20.0\text{mL}) = 7\text{mmol}$. $V_e = 7\text{mmol NaOH} \times (1 \text{ L} / 0.100\text{mol}) = 70.0 \text{ mL}$

c) i) $V_{\text{NaOH}} = 0$; we have buffer $\text{pH} = \text{p}K_1 + \log \frac{[\text{HSeO}_3^-]}{[\text{H}_2\text{SeO}_3]} = 2.57 + \log \frac{.150}{.200} = 2.45$

ii) pH at $\frac{1}{2} V_e = 20.0 \text{ mL} = \text{p}K_1 = 2.57$

iii) pH at $V_e = (1/2) (\text{p}K_1 + \text{p}K_2) = (1/2)(2.57 + 6.60) = 4.59$

iv) pH at $V_e + \frac{1}{2} V_{2e} = 40 + 35 = 75 \text{ mLs}$: $\text{pH} = \text{p}K_2$

d) for graph, ask in class.

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solution:

$$\text{FW of MgCl}_2 = 24.305 + 2(35.453) = 95.211 \text{ g/mol}$$

$$\text{a) } \#M = 20.0\text{g} / (1\text{mol}/95.211) / 0.0900 \text{ L} = 2.33 \text{ M}$$

b) # m = mol solute/kg solvent. where mol solute = $20.0/95.211 = 0.210$ mol MgCl₂
and where kg solvent = $(90.0\text{mL})(1.08\text{g/mL}) - 20.0\text{g} = 77.2\text{g water} = 0.0772 \text{ kg}$
so that: $m = 0.210\text{mol}/0.0772\text{kg} = 2.72 \text{ molal}$

$$\text{c) } \# \%m/V = 20.0\text{g}/90.0\text{mL} \times 100\% = 22.2 \%$$

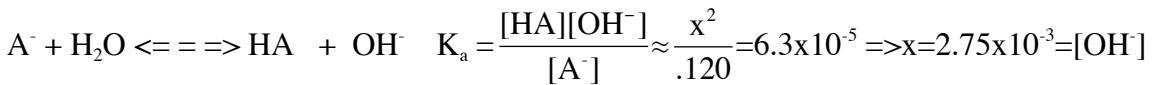
$$\text{d) } M_1V_1 = M_2V_2 \Rightarrow M_2 = M_1V_1/V_2 = (2.33)(5.0)/(500) = 0.0233 \text{ M}$$

$$\pi = iMRT = (3)(0.0233\text{mol/L})(0.0821\text{atmL/molK})(298\text{K}) = 1.71 \text{ atm}$$



$$\text{b) } M_A V_A = M_{\text{H}^+} V_e \Rightarrow V_e = M_A V_A / M_{\text{H}^+} = (0.120\text{M})(27.5) / (0.150\text{M}) = 22.0 \text{ mL}$$

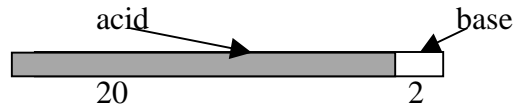
$$\text{c) } V_{\text{H}^+} = 0.$$



$$0.120 - x \quad \quad \quad x \quad \quad \quad x$$

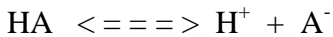
$$\text{pOH} = 2.56 \Rightarrow \text{pH} = 11.44$$

d) $V_{\text{H}^+} = 20.0\text{mL}$; Buffer region;



$$\text{pH} = \text{p}K_a + \log \frac{V_e - V_{\text{H}^+}}{V_{\text{H}^+}} = 4.20 + \log(2/20) = 3.20$$

$$\text{e) } V_{\text{H}^+} = V_e : \text{ pure weak acid; } [\text{HA}] = \frac{M_{\text{H}^+} V_{\text{H}^+}}{V_{\text{HA}} + V_{\text{H}^+}} = \frac{(0.120)(27.5\text{mL})}{(27.5\text{mL} + 22.0\text{mL})} = 0.067$$



$$0.067 - x \quad \quad \quad x \quad \quad \quad x$$

$$\frac{x^2}{.067 - x} \approx \frac{x^2}{.067} \approx 6.3 \times 10^{-5} \Rightarrow x = 2.05 \times 10^{-3}; \text{ pH} = 2.69$$

f) $V_{\text{H}^+} = 30.0 \text{ mL}$ (excess strong acid):

$$[\text{H}^+] = \frac{M_{\text{H}^+} (V_{\text{H}^+} - V_e)}{V_{\text{HA}} + V_{\text{H}^+}} = \frac{(0.150)(30.0 - 22.0)}{27.5 + 30.0} = 0.0209 \text{ M}; \text{ pH} = 1.68$$