

Chem 103 Test 1 key version A

<p>Part I</p> <p>1) B $\frac{540(100\%)}{540 + 580} = 48.2\%$</p> <p>2) D $\frac{\pi_{\text{Mg(NO}_3)_2}}{\pi_{\text{NaNO}_3}} = \frac{i' \text{MRT}}{i \text{MRT}} = \frac{i'}{i} = \frac{3}{2} = 1.5$</p> <p>3) A</p> <p>4) C</p> <p>5) E</p> <p>6) D</p> <p>7) C $[\text{OH}^-] = 1.2 \text{M}; [\text{H}^+] = 10^{-14.00}/1.2 = 8.3 \times 10^{-15}$</p> <p>8) C (the more polar the O-H bond, the easier to lose the H^+; related to electronegativity).</p>	<p>9) A (comes from $\text{HSO}_4^- \rightleftharpoons \text{H}^+ + \text{SO}_4^{2-}$)</p> <p>10) D</p> <p>11) B</p> <p>12) D</p> <p>13) C</p> <p>14) 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>15) A</p> <p>16) C (pH transition at equivalence much higher than $\text{pK}_a = 4.7$, phenolphthalein starts to change color at equiv. pt of HAC).</p>
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<p>Part II</p> <p>17) 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. 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. 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>	<p>18) There are several ways to solve this problem. One way is outlined below: 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 $\text{HA} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{A}^-$</p> $K_a = \frac{x}{[\text{H}_3\text{O}^+]^2/x} = \frac{x^2}{x^2 - x}$ $[\text{H}_3\text{O}^+]_{\text{HA}} = \sqrt{x \cdot K_a}$ <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 $\text{HCl}(\text{aq})$, $[\text{H}_3\text{O}^+] = 0.19 \text{ M}$ $\Rightarrow \text{pH} = 0.72$. 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>19) HA \rightleftharpoons H⁺ + A⁻</p> <p>0.35M-x x x</p> $\frac{x^2}{0.35 - x} = 4.2 \times 10^{-7}; x = \sqrt{(0.35)(4.2 \times 10^{-7})} = 3.83 \times 10^{-4} \Rightarrow \text{pH} = -\log x = 3.42$	<p>20) D A E B C</p> <p>pH of D ≈ 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>

$21) \Pi = iMRT; \text{ use } i=1,$ $M = \frac{\Pi}{RT} = \frac{(0.0203 \text{ atm})}{(0.0821 \text{ atmL/molK})(300 \text{ K})} = 8.24 \times 10^{-4} \text{ M}$ $M = \frac{3.6 \text{ g}}{0.200 \text{ L}} \Rightarrow MW = \frac{3.6 \text{ g}}{8.24 \times 10^{-4} \text{ mol}} = 2.18 \times 10^4 \text{ g/mol}$	
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Part III

<p>22) a) $\text{H}^+ + \text{C}_6\text{H}_5\text{CO}_2^- \rightarrow \text{HC}_6\text{H}_5\text{CO}_2$ or, $\text{H}^+ + \text{A}^- \rightleftharpoons \text{HA}$</p> <p>b) $M_A V_A = M_{\text{H}^+} V_e \Rightarrow V_e = M_A V_A / M_{\text{H}^+} = (0.120 \text{ M})(27.5) / (0.150 \text{ M}) = 22.0 \text{ mL}$</p> <p>c) $V_{\text{H}^+} = 0.$</p> $\text{A}^- + \text{H}_2\text{O} \rightleftharpoons \text{HA} + \text{OH}^- \quad K_a = \frac{[\text{HA}][\text{OH}^-]}{[\text{A}^-]} \approx \frac{x^2}{.120} = 6.3 \times 10^{-5} \Rightarrow x = 2.75 \times 10^{-3} = [\text{OH}^-]$ <p>0.120-x x x</p> <p>pOH = 2.56 \Rightarrow pH = 11.44</p> <p>d) $V_{\text{H}^+} = 20.0 \text{ mL}$; Buffer region;</p> $\text{pH} = \text{pK}_a + \log \frac{V_e - V_{\text{H}^+}}{V_{\text{H}^+}} = 4.20 + \log(2/20) = 3.20$ <p>e) $V_{\text{H}^+} = V_e$: pure weak acid; $[\text{HA}] = \frac{M_{\text{H}^+} V_{\text{H}^+}}{V_{\text{HA}}^0 + V_{\text{H}^+}} = \frac{(0.120)(27.5 \text{ mL})}{(27.5 \text{ mL} + 22.0 \text{ mL})} = 0.067$</p> $\text{HA} \rightleftharpoons \text{H}^+ + \text{A}^-$ <p>0.067-x x x</p> $\frac{x^2}{.067 - x} \approx \frac{x^2}{.067} \approx 6.3 \times 10^{-5} \Rightarrow x = 2.05 \times 10^{-3}; \text{ pH} = 2.69$ <p>f) $V_{\text{H}^+} = 30.0 \text{ mL}$ (excess strong acid):</p> $[\text{H}^+] = \frac{M_{\text{H}^+} (V_{\text{H}^+} - V_e)}{V_{\text{HA}} - V_{\text{H}^+}} = \frac{(0.150)(30.0 - 22.0)}{27.5 + 30.0} = 0.0209 \text{ M}; \text{ pH} = 1.68$

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<p>23) a) $\text{OH}^- + \text{H}_2\text{SeO}_3 \rightarrow \text{H}_2\text{O} + \text{HSeO}_3^-$; $V_e = \frac{M_{\text{H}_2\text{A}} V_{\text{H}_2\text{A}}}{M_{\text{OH}}} = \frac{(.200 \text{ M})(20.0 \text{ mL})}{(0.100 \text{ M})} = 40.0 \text{ mL}$</p> <p>b) $\text{OH}^- + \text{HSeO}_3^- \rightarrow \text{H}_2\text{O} + \text{SeO}_3^{2-}$: Note: $n_{\text{HSeO}_3} = n_{\text{H}_2\text{SeO}_4} + n_{\text{HSeO}_4} = (.150 + .200)(20.0 \text{ mL}) = 7 \text{ mmol}$. $V_e = 7 \text{ mmol NaOH} \times (1 \text{ L} / 0.100 \text{ mol}) = 70.0 \text{ mL}$</p> <p>c) i) $V_{\text{NaOH}} = 0$; we have buffer $\text{pH} = \text{pK}_1 + \log \frac{[\text{HSeO}_3^-]}{[\text{H}_2\text{SeO}_3]} = 2.57 + \log \frac{.150}{.200} = 2.45$</p> <p>ii) pH at $\frac{1}{2} V_e = 20.0 \text{ mL} = \text{pK}_1 = 2.57$</p> <p>iii) pH at $V_e = (1/2) (\text{pK}_1 + \text{pK}_2) = (1/2)(2.57 + 6.60) = 4.59$</p> <p>iv) pH at $V_e + \frac{1}{2} V_{2e} = 40 + 35 = 75 \text{ mLs}$: $\text{pH} = \text{pK}_2$ d) for graph, ask in class.</p>
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Solution:

a) get emp formula: assume 100g

$$\text{mol C} = 54.55\text{gC} (1\text{molC}/12.0\text{g}) = 4.55 \text{ mol C}$$

$$\text{mol H} = 9.09\text{gH}(1\text{mol H}/1.0\text{g}) = 9.09 \text{ mol H}$$

$$\text{mol O} = (100-54.55-9.09)(1\text{mol}/16.0\text{g})=2.27\text{molO}$$

tentative emp formula: divide subscripts by 2.27:

$$\text{C}_2\text{H}_4\text{O} \quad \text{and Emp wt} = 2(12.0) + 4(1.0) + 16.0 = 44.0 \text{ g/mol}$$

by FP depression: $\Delta T_f = K_f m = K_f(\text{g}/\text{MW})/(\text{.060 kg})$

$$\Rightarrow (6.55 - (-1.45)) = (20.) (2.11 / \text{MW}) / .060\text{kg} \Rightarrow 8.00 = 703 / \text{MW}$$

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

according to the % composition, the EW = 44.0 so the actual MW is double 44.0 or 88.0g/mol. the molecular formula is $\text{C}_4\text{H}_8\text{O}_2$.

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

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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 = $20.0/95.211 = 0.210 \text{ mol MgCl}_2$
 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: $m = 0.210\text{mol}/.0772\text{kg} = 2.72 \text{ molal}$

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

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

$$\pi = iMRT = (3)(.0233\text{mol/L})(.0821\text{atmL/molK})(298\text{K}) = 1.71 \text{ atm}$$