Chem 103 Test 1 key version B

Part I
1) A
2) B (HNO₃ + CH₃COONa), (NaOH + CH₃COOH), (CH₃COONa + CH₃COOH)
3) A
4) B
5) B
6) C (pH transition at equivalence much higher than pKₐ = 4.7, phenolphthalein starts to change color at equiv. pt of HAc).
7) D
8) E
9) C [OH⁻]=1.2M; [H⁺]=10⁻¹⁴.₀₀/1.₂=8.₃x10⁻¹⁵
10) D \( \frac{\pi_{Mg(NO_3)_2}}{\pi_{NaNO_3}} = \frac{i'MRT}{i'MRT} = \frac{i}{i} = \frac{3}{2} = 1.5 \)
11) C
12) C
13) C (the more polar the O-H bond, the easier to lose the H⁺; related to electronegativity).
14) A (comes from HSO₄⁻ < > H⁺ + SO₄²⁻)
15) D
16) D

Part II
17) There are several ways to solve this problem. One way is outlined below:
Let the unknown concentration = x. Then [H₃O⁺] for the HCl(aq) solution = x, and hence [H₂O⁺] for the acetic acid solution is x/100. But for the acetic acid solution, we also have HA + H₂O<= => H₃O⁺ + A⁻

\[
K_a = \frac{[H_3O^+][A^-]}{[HA]} = \sqrt{x.K_a}
\]

Therefore: \( [H_3O^+]_{HA} = x = [H_3O^+]_{HA} \times \frac{100}{x/100} = \sqrt{x.K_a} \)

=> \( x^2/10^4 = x.K_a \)

=> \( x = 10^4.K_a = 0.19 \text{ M} \)

Check: For HCl(aq), [H₃O⁺] = 0.19 M => pH = 0.72.
For hydrazoic acid, [H₃O⁺] = \( \sqrt{0.19 \times 1.9 \times 10^{-5}} = 1.9 \times 10^{-3} \) => pH = 2.72.

18) \( \pi = iMRT \); use i=1,

\[
M = \frac{\pi}{RT} = \frac{0.0203 \text{ atm}}{(0.0821 \text{ atm}L/molK)(300K)} = 8.24 \times 10^{-4}
\]

\[
M = \frac{3.6\text{ g}}{0.200L} \Rightarrow MW = \frac{3.6\text{ g}}{0.200L} = 8.24 \times 10^{-4} = 2.18 \times 10^4 \text{ g/mol}
\]

19) Lewis acid = electron pair acceptor. Cu²⁺ 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. H₂O and NH₃ 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 Cu²⁺ ion.

In this demonstration, the Cu²⁺ ion acts as the Lewis acid. Adding water to the solid CuSO₄ dissociates the Cu²⁺ ion from SO₄²⁻. Water molecules donate their lone e pairs to the Cu²⁺ to form a bluish complex of Cu²⁺ and water, [Cu(H₂O)₅]²⁺. Adding NH₃ to the solution immediately displaces the water from the Cu²⁺ and the resulting Cu(NH₃)₄²⁺ complex is the dark blue complex formed. It shows that NH₃ is a stronger Lewis base than water.
20) HA & H+ + A- 
\[0.35M-x\] \[x \times x\] 
\[\frac{x^2}{0.35-x} = 4.2 \times 10^{-7}\] 
x = \sqrt{(0.35)(4.2 \times 10^{-7})} = 3.83 \times 10^{-4}\] 
& pH = -\log x = 3.42

21) D A E B C
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

Part III

22) Solution:
a) get emp formula: assume 100 g
mol C = 54.55 g (1 mol C/12.0 g) = 4.55 mol C
mol H = 9.09 g (1 mol H/1.0 g) = 9.09 mol H
mol O = (100-54.55-9.09) (1 mol/16.0 g) = 2.27 mol O

tentative emp formula: divide subscripts by 2.27:
C_2H_4O and Emp wt = 2(12.0) + 4(1.0) + 16.0 = 44.0 g/mol

b) by FP depression: \[\Delta T = \frac{K_f m}{C} = \frac{K_f (g/MW)}{(0.060 \text{ kg})}\]
\[= (6.55-(-1.45)) = (20.)(2.11 / \text{MW}) / 0.060 \text{ kg} \Rightarrow 8.00 = 703 / \text{MW}\]
\[\Rightarrow \text{MW} = 703/8.00 = 87.9 \text{ g/mol}\]

according to the % composition, the EW = 44.0 so the actual MW is double 44.0 or 88.0 g/mol. the molecular formula is C_4H_8O_2.

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

23) a) OH^- + H_2SeO_3 \rightarrow H_2O + HSeO_3^-:
\[V_e = \frac{M_{H_2A}V_{H_2A}}{M_{OH}} = \frac{(0.200 M)(20.0 \text{ mL})}{(0.100 M)} = 40.0 \text{ mL}\]

b) OH^- + HSeO_3^- \rightarrow H_2O + SeO_3^{2-}:
Note: \[n_{H_2SeO_3} = n_{HSeO_3} + n_{HSeO_4} = (0.150+0.200)(20.0 \text{ mL}) = 7 \text{ mmol}\]
\[V_e = 7 \text{ mmol NaOH} \times 1 \text{ L} / 0.100 \text{ mol} = 70.0 \text{ mL}\]

c) i) \[V_{NaOH} = 0; \text{ we have buffer pH} = pK_1 + \log \frac{[HSeO_3^-]}{[H_2SeO_3]} = 2.57 + \log \frac{0.150}{0.200} = 2.45\]

ii) pH at \(\frac{1}{2} V_e = 20.0 \text{ mL} = pK_1 = 2.57\)

iii) pH at \(V_e = (\frac{1}{2})(pK_1 + pK_2) = (\frac{1}{2})(2.57 + 6.60) = 4.59\)

iv) pH at \(V_e + \frac{1}{2} V_{2e} = 40 + 35 = 75 \text{ mLs: pH} = pK_2\)

d) for graph, ask in class.
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solution:
FW of MgCl₂ = 24.305 + 2(35.453) = 95.211 g/mol
a) #M = 20.0 g (1 mol/95.211) / 0.0900 L = 2.33 M

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

c) # %m/V = 20.0 g/90.0 mL x 100% = 22.2 %
d) M1V1 = M2V2 => M2 = M1V1/V2 = (2.33)(5.0)/(500) = 0.0233 M
π = iMRT = (3)(0.0233 mol/L)(0.0821 atm L/mol K)(298 K) = 1.71 atm

25) a) H⁺ + C₆H₅CO₂⁻ → HC₆H₅CO₂⁻ or, H⁺ + A⁻ ⇌ HA
b) MₐVₐ = MₜₐV⁺ => V⁺ = MₐVₐ/Mₜₐ = (0.120 M)(27.5 mL)/(0.150 M) = 22.0 mL

c) Vₜ⁺=0.

A⁻ + H₂O ⇌ HA + OH⁻ Ka = [HA][OH⁻]/[A⁻] ≈ x² / 0.120 x = 6.3x10⁻⁵ => x = 2.75x10⁻⁴ [OH⁻]

pOH = 2.56 => pH = 11.44

d) Vₜ⁺=20.0 mL; Buffer region;

pH = pKₐ + log [Vₜ⁺/Vₜ⁺] [H⁺] = 4.20 + log(2/20) = 3.20

e) Vₜ⁺ = Vₜ⁻: pure weak acid; [HA] = Mₖₜ⁺Vₜ⁺ / (Vₐ [Vₜ⁻] + Vₜ⁺) = (0.120)(27.5 mL) / (27.5 mL + 22.0 mL) = 0.067

HA ⇌ H⁺ + A⁻

0.067 - x  x  x

x² / 0.067 - x ≈ x² / 6.3x10⁻⁵ => x = 2.05 x 10⁻³; pH = 2.69

f) Vₜ⁺ = 30.0 mL (excess strong acid):

[H⁺] = Mₖₜ⁺(Vₜ⁺ - Vₜ⁻) / Vₐ[Vₜ⁻] = (0.150)(30.0 - 22.0) / 27.5 + 30.0 = 0.0209 M; pH = 1.68