

discard

ANSWER KEY

CHEMISTRY F1404
PROFESSOR J. MORROW

FIRST EXAM

2/16/00

PRINT NAME, LAST: _____

FIRST: _____

I.D.# : _____

EACH QUESTION IS WORTH 10 POINTS

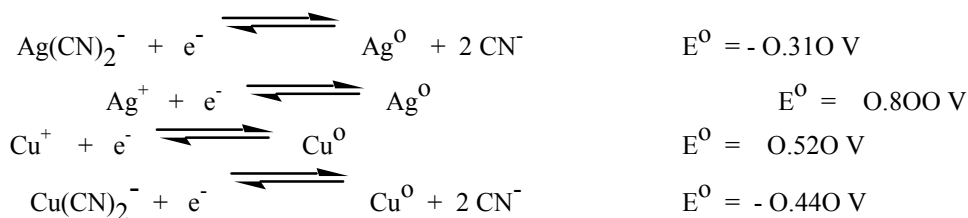
- | | | |
|----------|-----------|-----------|
| 1. _____ | 7. _____ | 13. _____ |
| 2. _____ | 8. _____ | 14. _____ |
| 3. _____ | 9. _____ | 15. _____ |
| 4. _____ | 10. _____ | 16. _____ |
| 5. _____ | 11. _____ | 17. _____ |
| 6. _____ | 12. _____ | 18. _____ |

COLUMN TOTALS:

EXAM TOTAL (140 pts) _____

OUT OF 100

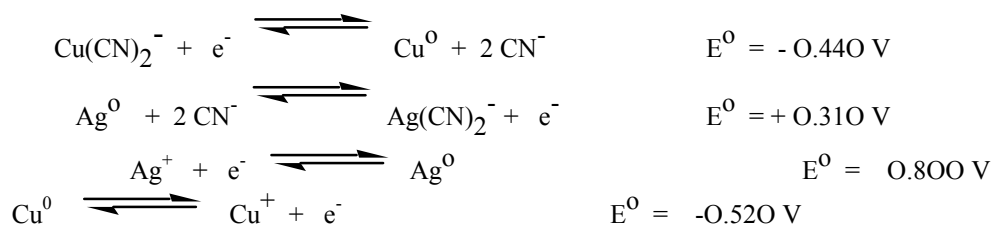
1) GIVEN THE FOLLOWING STANDARD REDUCTION POTENTIALS (25°C)



Calculate ΔG^0_{rxn} (in kJ) for the following reaction; SHOW WORK



- a) -14.5 b) 12.5 c) 14.5 d) -12.5 e) 25.0 f) -25.0



$$E^0_{\text{cell}} = +0.150 \text{ volts}$$

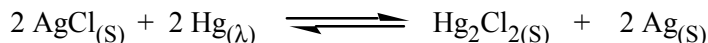
$$\Delta G^0_{\text{rxn}} = -nFE^0 = -(1)(96500)(0.150) = -14475 \text{ J} = -14.5 \text{ kJ}$$

GIVE 5 PTS IF ANSWER GIVEN IS d ANSWER: $\Delta G^0_{\text{rxn}} = \underline{\text{a}}$

2) For the reaction, $2 \text{Hg}_{(\text{l})} + \text{Cl}_{2(\text{g})} \longrightarrow \text{Hg}_2\text{Cl}_{2(\text{s})}$, $E^0 = +1.058 \text{ volts}$.

For the reaction, $2 \text{Ag}_{(\text{s})} + \text{Cl}_{2(\text{g})} \longrightarrow 2 \text{AgCl}_{(\text{s})}$, $E^0 = +1.178 \text{ volts}$.

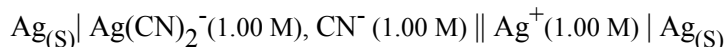
Calculate the equilibrium constant for the following reaction;



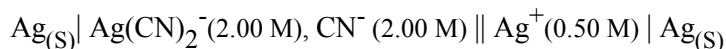
$$\begin{array}{l}
 E^0_{\text{cell}} = -0.120 \text{ volts} \\
 K = 10^{nE/0.0592} = 10^{(2)(-0.120)/(0.0592)} \\
 K = 8.83 \times 10^{-5}
 \end{array}$$

ANSWER: $K = 8.83 \times 10^{-5}$

3) For the equilibrium reaction, $\text{Ag}^+ + 2 \text{CN}^- \rightleftharpoons \text{Ag(CN)}_2^-$, the cell notation when the reactants and products are in standard state is;



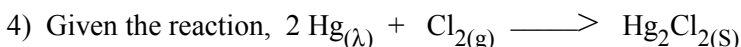
If the cell with the new indicated concentrations is now set up,



then,

- The reaction will shift to the right
- The reaction will shift to the left
- The cell voltage will become more positive
- The cell voltage will become less positive
- The cell voltage will be unchanged
- The equilibrium constant will increase in value
- The equilibrium constant will decrease in value

YOUR ANSWER(S) IS/ARE: e



Assume that the standard E° cell voltage is related to the Celsius temperature by the following equation and all substances are at standard state. In doing this, you can assume any temperature(s) you desire.

At 298 K (25°C), $E^\circ_{\text{cell}} = +1.058$ volts.

SHOW WORK

$$E^\circ_{\text{cell}} = 1.058 - 0.00017 (t - 25^\circ\text{C})$$

The ΔH° (in kJ) for the reaction (as written) at 298 K is;

- a) -321 b) +107 c) +214 d) -214 e) -107 f) +321

$$\Delta H^\circ = -nF \left\{ E^\circ - T \left(\frac{E^\circ_2 - E^\circ_1}{\Delta T} \right) \right\} = - (2)(96500) \{ 1.058 - (298)(-0.00017) \} \quad \Delta H^\circ = -214 \text{ kJ}$$

GIVE 5 PTS IF ANSWER GIVEN IS e ANSWER IS: d

5) What is the highest pH at which 0.050 M Mn^{2+} remains entirely in a solution that is saturated with H_2S , at a concentration of, $[\text{H}_2\text{S}] = 0.10 \text{ M}$.

Given: For the reaction, $\text{MnS}_{(s)} \rightleftharpoons \text{Mn}^{2+} + \text{S}^{2-}$, $K_{\text{sp}} = 3.6 \times 10^{-15}$ OR

for the reaction, $\text{MnS}_{(s)} + \text{H}_2\text{O} \rightleftharpoons \text{Mn}^{2+} + \text{HS}^- + \text{OH}^-$, $K = 3.0 \times 10^{-14}$

For H_2S : $K_{a1} = 9.2 \times 10^{-8}$ and $K_{a2} = 1.2 \times 10^{-15}$ SHOW WORK

- a) 2.9 b) 3.9 c) 4.4 d) 4.9 e) 5.4

$$K = 3.0 \times 10^{-14} = [\text{Mn}^{2+}][\text{HS}^-][\text{OH}^-] = [\text{Mn}^{2+}] \left(\frac{[\text{H}_2\text{S}]K_{a1}}{[\text{H}^+]} \right) \left(\frac{K_w}{[\text{H}^+]} \right)$$

$$[\text{H}^+]^2 = \frac{[\text{Mn}^{2+}][\text{H}_2\text{S}]K_{a1}K_w}{3.0 \times 10^{-14}} = 1.53 \times 10^{-10} \quad \text{AND} \quad [\text{H}^+] = 1.23 \times 10^{-5}$$

$$\text{pH} = 4.9$$

ANSWER IS: d

6) Find among the pH values on the right, the value which most closely

matches the pH of the solutions described in the first column. SHOW ALL WORK FOR PART B ON THIS PAGE!

HINT: In PART B the following reaction occurs, $\text{H}_2\text{CO}_3 + \text{CO}_3^{2-} \longrightarrow 2 \text{HCO}_3^-$

| <u>SOLUTION</u> | <u>pH</u> | |
|---|-----------|-------|
| A. 0.460 M CH_3COOH (HOAc) | 5.04 | |
| that is also 0.0460 M in HNO_3 | 4.44 | |
| | 1.34 | |
| B. 0.020 M H_2CO_3 that is also | 11.12 | 0.250 |
| M in Na_2CO_3 | 10.72 | |
| | 5.13 | |
| | 2.70 | 7.41 |
| | 2.00 | |

(4 pts) Solution A pH = 1.34

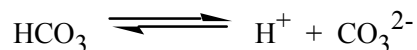
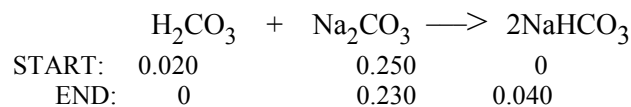
(6 pts) Solution B pH = 11.12

K_a values: H_2CO_3 ($K_{a1} = 4.3 \times 10^{-7}$), HCO_3^- ($K_{a2} = 4.4 \times 10^{-11}$), HOAc ($K_a = 1.8 \times 10^{-5}$)

Solution A:

FROM STRONG ACID ONLY: $[\text{H}^+] = 0.046 \text{ M}$
pH = 1.34

Solution B:



$$K_{a2} = 4.4 \times 10^{-11} = \frac{[\text{H}^+][\text{CO}_3^{2-}]}{[\text{HCO}_3^-]} = \frac{[\text{H}^+][0.230]}{[0.040]}$$

$$[\text{H}^+] = \frac{(4.4 \times 10^{-11})(0.040)}{0.230} = 0.765 \times 10^{-12}$$

pH = 11.12

7) If 0.77 g of $\text{Cu}(\text{IO}_3)_2$ will dissolve in 1.0 L of 0.10 M CuSO_4 solution, what is the solubility product, K_{sp} , of $\text{Cu}(\text{IO}_3)_2$? Molar mass = 413.3 g/mol

SHOW WORK

- a) 1.4×10^{-5} b) 1.4×10^{-6} c) 2.6×10^{-7} d) 3.5×10^{-7} e) 2.6×10^{-8}

$$K_{\text{sp}} = [\text{Cu}^{2+}][\text{IO}_3^-]^2 = (0.10)\left(\frac{(2)(0.77)}{413.3}\right)^2$$

$$K_{\text{sp}} = 1.4 \times 10^{-6}$$

IF ANSWER GIVEN IS d GIVE 5 PTS. ANSWER IS: b

THE FOLLOWING INFORMATION IS FOR PROBLEMS 8, 9, AND 10.

40.00 mL of 0.100 M methylamine, CH_3NH_2 , is titrated with 0.150 M HCl.

Calculate the pH; ($K_b = 4.0 \times 10^{-5}$ for $\text{CH}_3\text{NH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{NH}_3^+ + \text{OH}^-$)

8) before any HCl is added.

$$K_b = 4.0 \times 10^{-5} = \frac{[\text{CH}_3\text{NH}_3^+][\text{OH}^-]}{[\text{CH}_3\text{NH}_2]} = \frac{x^2}{0.10}$$

$$x = [\text{OH}^-] = 2.0 \times 10^{-3} \quad \text{pOH} = 2.7 \quad \text{AND} \quad \text{pH} = 11.3$$

$$\text{pH} = \underline{11.3}$$

9) halfway to the equivalence point.

$$\text{pH} = \text{pK} = -\log(4.0 \times 10^{-5})$$

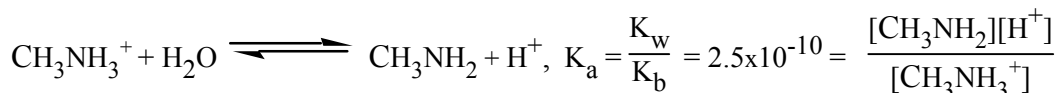
$$\text{pH} = 4.4$$

$$\text{pH} = \underline{4.4}$$

10) at the equivalence point. (This is now an hydrolysis problem)

SHOW WORK AT EQUIVALENCE POINT, VOLUME OF ADDED HCl is 26.7 mL

TOTAL VOLUME IS $40.0 + 26.7 = 66.7$ mL AND $[\text{CH}_3\text{NH}_3^+] = \frac{0.0040 \text{ mol}}{0.0667 \text{ L}} = 0.0600 \text{ M}$



$$2.5 \times 10^{-10} = \frac{x^2}{0.0600} \quad x = [\text{H}^+] = 3.87 \times 10^{-6} \quad \text{AND} \quad \text{pH} = 5.41$$

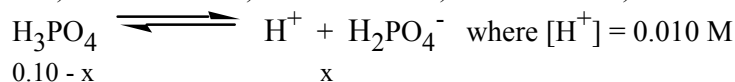
GIVE 5 PTS FOR pH = 5.30 pH = 5.41

11) How many milliliters (mL) of a 0.0500 M NaOH (a strong base) solution should be added to 1.00 L of 0.100 M H_3PO_4 solution, to produce

a buffer of pH = 2.00 ? For H_3PO_4 , $K_{a1} = 6.67 \times 10^{-3}$ (SHOW WORK)

GIVEN: $\text{H}_3\text{PO}_4 + \text{OH}^- \text{ (from NaOH)} \longrightarrow \text{H}_2\text{PO}_4^- + \text{H}_2\text{O}$

a) 100 b) 200 c) 400 d) 600 e) 800 f) 900



$$K_{a1} = 6.67 \times 10^{-3} = \frac{[\text{H}^+][\text{H}_2\text{PO}_4^-]}{[\text{H}_3\text{PO}_4]} = \frac{(0.0100)[\text{H}_2\text{PO}_4^-]}{[\text{H}_3\text{PO}_4]}$$

$$\frac{[\text{H}_2\text{PO}_4^-]}{[\text{H}_3\text{PO}_4]} = 0.667 = \frac{x}{0.100 - x} \quad \therefore x = 0.0400$$

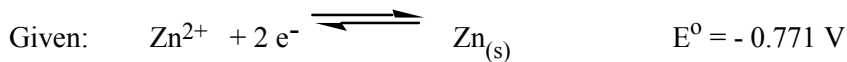
$$M_{\text{NaOH}} V_{\text{NaOH}} = 0.0400 \quad \therefore V_{\text{NaOH}} = 0.800 \text{ L}$$

ANSWER IS: e

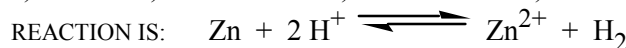
12) Given the following cell; $\text{Zn}_{(\text{s})} | \text{Zn}^{2+} (1 \text{ M}) || \text{H}^+ (? \text{ M}) | \text{H}_{2(\text{g})} (1 \text{ atm})$.

The measured cell voltage is $E_{\text{cell}} = 0.653 \text{ V}$. Calculate the pH of the solution in the cathode

compartment. HINT: First calculate E°_{cell} . MOLARITIES AND PRESSURE ARE GIVEN IN THE CELL NOTATION ABOVE.



a) 1.0 b) 1.5 c) 2.0 d) 3.0 e) 3.5 f) 4.0



$$E = E^{\circ} - \frac{0.0592}{2} \log \frac{[\text{Zn}^{2+}]}{[\text{H}^+]^2} \quad 0.653 = 0.771 - \frac{0.0592}{2} \log \frac{1}{[\text{H}^+]^2}$$

$$[\text{H}^+] = 10^{-2} \text{ M} \quad \text{AND} \quad \text{pH} = 2.0$$

ANSWER IS: c

13) Complete the following table. (For $\text{Ca}_3(\text{PO}_4)_2$, $K_{\text{sp}} = 1.3 \times 10^{-32}$)

| $[\text{Ca}^{2+}]$ | $[\text{PO}_4^{3-}]$ |
|---|--|
| <u>1.09×10^{-8}</u> | $1.0 \times 10^{-4} \text{ M}$ |
| $6.0 \times 10^{-3} \text{ M}$ | <u>2.45×10^{-13}</u> |

14) You have a solution of sodium hydrogen sulfate, NaHSO_4 . It's initial concentration is C_0 . The pH = 2.0 for this solution at equilibrium. Calculate the starting concentration, C_0 .

(For the reaction, $\text{HSO}_4^- \rightleftharpoons \text{H}^+ + \text{SO}_4^{2-}$, $K_{a2} = 0.020$)

$$0.020 = \frac{[\text{H}^+][\text{SO}_4^{2-}]}{[\text{HSO}_4^-]} = \frac{(0.01)^2}{C_0 - 0.01} \quad C_0 = \underline{0.015 \text{ M}}$$

15) For Ag_2CO_3 , $K_{\text{sp}} = 6.2 \times 10^{-12}$. For AgCl , $K_{\text{sp}} = 2.8 \times 10^{-10}$.

Solid Ag_2CO_3 and solid AgCl are added to a beaker containing 1.00 M

$\text{Na}_2\text{CO}_3(\text{aq})$. Under these conditions the $[\text{CO}_3^{2-}] = 1.00 \text{ M}$. Calculate the $[\text{Cl}^-]$ in solution when equilibrium is established.

ASSUME THE CO_3^{2-} ION DOES NOT HYDROLYZE.

SHOW WORK

- a) 1.1×10^{-4} b) 1.26×10^{-8} c) 6.7 d) 0.15 e) 2.8×10^{-6}

$$\text{Ag}_2\text{CO}_3 \rightleftharpoons 2 \text{Ag}^+ + \text{CO}_3^{2-} \quad \therefore K_{\text{sp}} = [\text{Ag}^+]^2[\text{CO}_3^{2-}]$$

$$[\text{Ag}^+]^2 = \frac{6.2 \times 10^{-12}}{[\text{CO}_3^{2-}]} = \frac{6.2 \times 10^{-12}}{1.0} = 6.2 \times 10^{-12} \quad \therefore [\text{Ag}^+] = 2.49 \times 10^{-6}$$

$$\text{For AgCl, } K_{\text{sp}} = 2.8 \times 10^{-10} = [\text{Ag}^+][\text{Cl}^-] = (2.49 \times 10^{-6})[\text{Cl}^-]$$

$$\therefore [\text{Cl}^-] = \frac{2.8 \times 10^{-10}}{2.49 \times 10^{-6}} = 1.1 \times 10^{-4} \text{ M}$$

ANSWER IS: a

16) Given the following K_{sp} values: for $\text{M}(\text{OH})_4$, $K_{\text{sp}(\text{M})} = 4.0 \times 10^{-19}$;

for $\text{Z}(\text{OH})_2$, $K_{\text{sp}(\text{Z})} = 1.0 \times 10^{-14}$. One mole of each of the above solids is placed in a beaker containing 1 (one) liter of pure water. These solids go into equilibrium with their ions. Calculate the pH required for the metal ion concentrations to be equal, $[\text{M}^{2+}] = [\text{Z}^{3+}]$. (SHOW WORK)

- a) 12.6 b) 9.6 c) 6.7 d) 2.2 e) 11.8

$$K_{\text{sp}(\text{M})} = 4.0 \times 10^{-19} = [\text{M}^{4+}][\text{OH}^-]^4 \quad K_{\text{sp}(\text{Z})} = 1.0 \times 10^{-14} = [\text{Z}^{2+}][\text{OH}^-]^2$$

$$\frac{K_{\text{sp}(\text{M})}}{K_{\text{sp}(\text{Z})}} = \frac{4.0 \times 10^{-19}}{1.0 \times 10^{-14}} = 4.0 \times 10^{-5} = [\text{OH}^-]^2$$

$$[\text{OH}^-] = (4.0 \times 10^{-5})^{1/2} = 6.32 \times 10^{-3} \quad \therefore \text{pOH} = 2.20 \quad \text{AND} \quad \text{pH} = 11.80$$

pH = e

17) Given the following K_{a} and K_{b} values: HF ($K_{\text{a}} = 10^{-4}$), HCN ($K_{\text{a}} = 10^{-9}$),

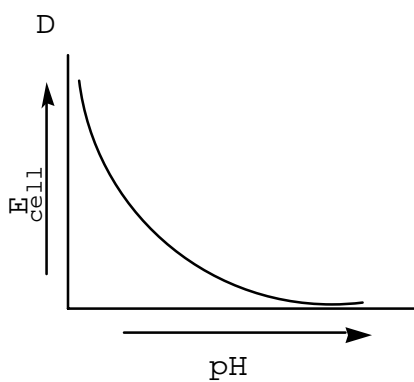
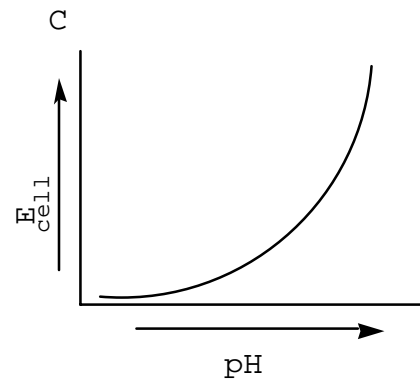
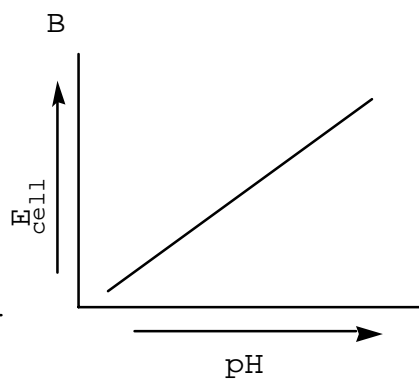
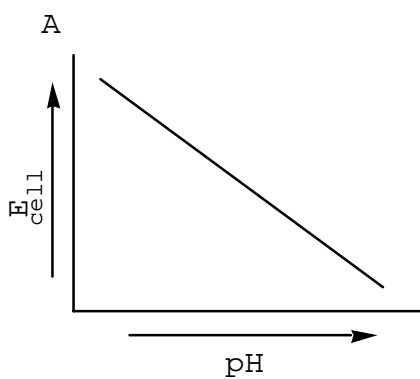
NH_3 ($K_{\text{b}} = 10^{-5}$), HOAc ($K_{\text{a}} = 10^{-5}$), and HCl ($K_{\text{a}} = \infty$)

List the following aqueous salts in order of increasing acidity (decreasing basicity); NaF , NH_4Cl , NaCN , NaCl , and NaOAc . THIS IS AN HYDROLYSIS PROBLEM.

LEAST ACIDIC $\text{NaCN} < \text{NaOAc} < \text{NaF} < \text{NaCl} < \text{NH}_4\text{Cl}$ MOST ACIDIC

18) Which graph correctly illustrates the dependence of the cell voltage, E_{cell} , on the pH for the reaction, $2 \text{Ag}^+_{(\text{aq})} + \text{H}_{2(\text{g})} \longrightarrow 2 \text{Ag}_{(\text{s})} + 2 \text{H}^+_{(\text{aq})}$.

Assume the silver ion remains constant at, $[\text{Ag}^+] = 1.0 \text{ M}$ and the H_2 pressure remains constant at 1 atm.



ANSWER IS: B