KWANTLEN
POLYTECHNIC
UNIVERSITY

CHEMISTRY 1210 – SPRING 2017

EXAM 2

March 23rd 2017

Name: ______________________________

Student #: ______________________________

_____ **/ 37**

Time allowed: 1h50

Only approved calculators are permitted
Cell phones and other electronics must be turned off
"What is 'HIJKLMNO'? – H_2O"
Good Luck – Bonne chance – Suerte

$T_k = T_c + 273.15$
$K_w = 1.0 \times 10^{-14}$ (at 25°C)   Assume all acid-base questions are at 25°C unless specified

1) ***(2 pts)*** Which of the above salts would have the highest molar solubility?

a) AgCl ksp = $1.8x10^{-10}$ $= x^2$ $x = 1.34 \times 10^{-8}$

b) AgBr ksp= $3.4x10^{-18}$ $= x^2$ $x = 1.84 \times 10^{-9}$

c) Ag_2S ksp = $1.9x10^{-12}$ $= 4x^3$ $x = 7.80 \times 10^{-5}$

d) Ag_3PO_4 ksp = $2.1x10^{-16}$ $= 27x^4$ $x = 5.28 \times 10^{-5}$

2) ***(2 pts)*** Calculate the molar solubility of the compound AgCl (K_{sp} =1.8×10^{-10}) in a solution containing 0.010 M $CaCl_2$. $\longrightarrow$ 0.020 M Cl^-

$\hookrightarrow Ca^{2+} + 2Cl^-$

$1.8 \times 10^{-10} = (x)(0.020\ M)$

Assume x ≪ 0.020 M

$x = \frac{1.8 \times 10^{-10}}{0.020\ M} = \boxed{9.0 \times 10^{-9}\ M}$

3) ***(2 pts)*** Will a precipitate of magnesium fluoride form when 300. mL of 1.1×10^{-3} M $MgCl_2$ are added to 500. mL of 1.2×10^{-3} M NaF? (K_{sp} (MgF_2) = 6.9×10^{-9}) (Assume NaCl is highly soluble and does not precipitate)

$MgF_2 \rightleftharpoons Mg^{2+} + 2F^-$ $K_{sp} = [Mg^{2+}][F^-]^2$

$[Mg^{2+}] = 1.1 \times 10^{-3} M \times \frac{300\ mL}{800\ mL} = 4.125 \times 10^{-4} M$

$[F^-] = 1.2 \times 10^{-3} M \times \frac{500\ mL}{800\ mL} = 7.5 \times 10^{-4} M$

$Q_{sp} = (4.125 \times 10^{-4})(7.5 \times 10^{-4})^2$

$= 2.32 \times 10^{-10} < K_{sp}$

$\hookrightarrow$ NO PPT

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4) ***(2 pts)*** The solubility of tin(I) iodide (SnI) is 4.77 g/L. What is K_{sp} for this compound

$\rightarrow 245.6\ g/mol$

$$\frac{4.77g}{L} \times \frac{1\ mol}{245.6g} = 0.0194\ \frac{mol}{L}$$

$$SnI \rightleftharpoons Sn^{2+} + I^{-}$$
$$0.0194M \rightarrow 0.0194 \quad 0.0194$$

$$K_{sp} = [Sn^{2+}][I^{-}] = (0.0194M)^2 = \boxed{3.77\times10^{-4}}$$

5) ***(3 pts)*** A solution contains 0.10 M Ba^{2+} and 0.10 M Sr^{2+}. A solution of CrO_4^{2-} is slowly added to increase the concentration of CrO_4^{2-} in solution. Considering that ksp $BaCrO_4 = 1.2x10^{-10}$ and ksp of $SrCrO_4 = 3.5x10^{-5}$.

a) Which of ion (Ba^{2+} or Sr^{2+}) will precipitate first

Ba^{2+}

b) What <u>range</u> of concentrations will precipitate one cation but not the other?

$$BaCrO_4 \rightleftharpoons Ba^{2+} + CrO_4^{2-} \qquad K_{sp} = [Ba^{2+}][CrO_4^{2-}]$$

$$[CrO_4^{2-}] = \frac{K_{sp}}{[Ba^{2+}]} = \frac{1.2\times10^{-10}}{0.10} = 1.2\times10^{-9}$$

$$SrCrO_4 \rightleftharpoons Sr^{2+} + CrO_4^{2-} \qquad K_{sp} = [Sr^{2+}][CrO_4^{2-}]$$

$$[CrO_4^{2-}] = \frac{K_{sp}}{[Sr^{2+}]} = \frac{3.5\times10^{-5}}{0.10M} = 3.5\times10^{-4}$$

$$\boxed{1.2\times10^{-9}M < [CrO_4^{2-}] < 3.5\times10^{-4}M}$$

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6) ***(3 pts)*** Given the following <u>*balanced*</u> redox reaction:

$$2OH^{1-} + 2Re + 7ClO_3^{1-} \rightarrow 7ClO_2^{1-} + 2ReO_4^{1-} + H_2O$$

a) What is the oxidizing agent? ClO_3^-

b) What element(s) is/are undergoing oxidation? Re

c) How many electrons are involved in the reduction process? $14e^-$ ($2e^-$ for 1/2 pt)

ox

$\cancel{8H_2O} + 2Re \rightarrow 2ReO_4^- + \cancel{8H^+}\ 16H^+ \cancel{8H_2O}\ 8H_2O + 14e^-$

$16OH$ $16H_2O$

RED

$7H_2O + 7ClO_3^- + 14e^- \rightarrow 7ClO_2^- + \cancel{H_2O} + 14OH$

7) ***(2 pts)*** Balance the following oxidation reduction reaction occurring in basic solution.

$$Cr(OH)_4^{1-} + H_2O_2 \rightarrow CrO_4^{2-} + H_2O$$

$(4OH^- + Cr(OH)_4^- \rightarrow CrO_4^{2-} + 4H_2O + 3e^-) \times 2$

$2H_2O + H_2O_2 \rightarrow \cancel{2H_2O} + 2OH$

$(2e^- + H_2O_2 \rightarrow 2OH^-) \times 3$

$\cancel{8OH^-} + 2Cr(OH)_4^- \rightarrow 2CrO_4^{2-} + 8H_2O + \cancel{6e^-}$

$\cancel{6e^-} + 3H_2O_2 \rightarrow \cancel{6OH^-}$

$$\boxed{2OH^- + 2Cr(OH)_4^- + 3H_2O_2 \rightarrow 2CrO_4^{2-} + 8H_2O}$$

ATOMS ☑

CHARGES ☑

8) **(7 pts)** Use the following reduction potentials to answer this question:

$IO_4^{1-} + 2H^+ + 2e^- \leftrightarrow IO_3^{1-} + H_2O$ $E^o = 1.589$ V

$IO_3^{1-} + 5H^+ + 4e^- \leftrightarrow HOI + 2H_2O$ $E^o = 1.154$ V

$Cu^{2+} + 2e^- \leftrightarrow Cu$ $E^o = 0.339$ V REDUCTION

$IO_3^{1-} + 3H_2O + 6e^- \leftrightarrow I^{1-} + 6OH^{1-}$ $E^o = 0.269$ V OXIDATION

$Zn^{2+} + 2e^- \leftrightarrow Zn$ $E^o = -0.762$ V

a) The strongest reducing agent is: (1 pt) (gets oxidized) Zn (or s for Zn^{2+})

b) Write the shorthand cell notation for the *spontaneous* electrochemical cell formed by coupling a Cu/Cu^{2+} half cell with an IO_3^{1-} half cell in basic solution.

$Pt(s) \mid IO_3^-(aq), I^-(aq), OH^-(aq) \parallel Cu^{2+}(aq) / Cu(s)$

c) Consider an electrochemical cell prepared by connecting a 1.0 L Copper half-cell containing a copper electrode and 0.10 M Cu^{2+} and a 1.0 L Zinc half-cell containing a zinc electrode in a 0.50 M Zn^{2+} solution

i) Determine the initial cell potential of this electrochemical cell

$Cu^{2+} + 2e^- \rightarrow Cu$ 0.339 V

$Zn^{2+} + 2e^- \rightarrow Zn$ −0.762 V

$Cu^{2+} + Zn \rightarrow Cu + Zn^{2+}$ $E = 1.101$ V

$$Q = \frac{[Zn^{2+}]}{[Cu^{2+}]}$$

$$E = E^o - \frac{0.05916}{2} \log Q = 1.101\text{ V} - \frac{0.05916}{2} \log\left(\frac{0.50}{0.10}\right) = \boxed{1.08\text{ V}}$$

ii) Determine which electrode would be heavier and by what mass (what is the mass increase) after the cell operates for 121 seconds at 5.0 A.

Cu(s) increases in mass as it is the reduction product

$$n = \frac{It}{F} = \frac{(5.0A)(121s)}{96485} = 6.27\times10^{-3} \text{ mole } e^- \times \frac{1 Cu}{2e^-} \times \frac{63.5}{1 mol}$$

$\boxed{Cu(s) \quad 0.20g}$

+ve

9) ***(1 pts)*** An experiment in a coffee cup calorimeter results in q_{rxn}= 2.345 kJ. Is the reaction *exothermic or endothermic and will the calorimeter temperature increase or decrease?*

Endo or Exo *Calorimeter T:* *increase or decrease*

10) **(2 pts)** High purity benzoic acid (MW: 122.122) has ΔH = -3227 kJ/mol for its combustion reaction. A 1.221 g sample burns in a calorimeter (heat capacity = 1365 J/°C) that also contains 1201 g of water (heat capacity = 4.184 J/g °C). What is the temperature change following complete combustion of the benzoic acid sample?

$$-3227\frac{kJ}{mol} \times 1.221 g \times \frac{1\ mol}{122.122} = -32.26\ kJ = -32\ 260\ J$$

$$q_{rxn} + q_{cal} + q_{H_2O} = 0$$

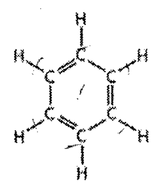
$$-32260 J + 1365\frac{J}{°C}\Delta T + (1201 g)(4.184\frac{J}{g\ °C})\Delta T = 0$$

$$32260 J = 6389.98\ \Delta T$$

$$\boxed{\Delta T = 5.05\ °C}$$

11) **(2 pts)** Determine the change in enthalpy (ΔH) resulting from the combustion of 1.00 mol of benzene(C_6H_6)? Consider the <u>unbalanced</u> reaction: $2C_6H_6 + {}_{15}O_2 \rightarrow {}_{12}CO_2 + {}_{6}H_2O$

H-O-H

Structure of benzene:  or 

Bond energies in kJ/mol: C-H: 415; C=C: 615; C-C: 345; O=O: 495; C=O: 750; O-H: 460

BOND BREAKING

$$12(C-H) + 6(C-C) + 6(C=C) + 15\ O=O$$

$$12(415) + 6(345) + 6(615) + 15(495) = 18165\frac{kJ}{mol}$$

BOND MAKING

$$24(C=O) + 12(O-H)$$

$$24(750) + 12(460) = 23520$$

$$\Delta H = 18165 - 23520 = -5355\frac{kJ}{mol\ rxn} \times \frac{1\ rxn}{2\ C_6H_6} = \boxed{-2678\ kJ}$$

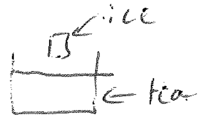
12) ***(3 pts)*** One way to quickly cool your tea after steeping the tea leaves in hot water is to add an ice cube. Determine the final temperature of a 251 mL tea that was initially at 85.0 °C when a 25 g ice cube at -10.0 °C is added to the tea.

(For the tea, assume heat capacity is the same as water c=4.184 J/g°C and density of water 1.000 g/mL)

(assume this takes place in a perfect insulator, c = 0 J/°C, no calorimeter!)

(ΔH_{fus}(ice) = 6100 J/mol, c_{ice} = 2.04 J/g°C, melting temperature or ice is 0°C, boiling point of water is 100°C)

(Hint: this is a multi-step process)



$$\underset{\text{}}{q_{tea}} + \underset{\text{heat(s)}}{q_{ice}} + \underset{\text{melt}}{q_{ice}} + \underset{\text{heat(l)}}{q_{ice}} = 0$$

$$q_{tea} \rightarrow mc\Delta T = 251g\,(4.184)(T_f - 85°C) = 1050\,T_f - 89266\,J$$

$$q_{ice} \rightarrow mc\Delta T = 25g\,\left(2.04\,\frac{J}{g°C}\right)(0 - -10°C) = 510\,J$$

$$q_{ice} \rightarrow 6100\,\frac{J}{mol} \cdot 25g \times \frac{1\,mol}{18.0\,g} = 8472\,J$$

$$q_{ice} \rightarrow 25g\,(4.184)(T_f - 0) = 104.6\,T_f$$

$$1050\,T_f - 89266\,J + 510\,J + 8472\,J + 104.6\,T_f = 0$$

$$1154.6\,T_f = 80284$$

$$\boxed{T_f = 69.5\,°C}$$

13) ***(2 pts)*** A calorimeter is calibrated by adding a fixed amount of heat to it and measuring its temperature change. When 20.3 kJ of heat are added to a 25.5 °C calorimeter, its temperature rises to 30.4 °C. The same calorimeter is used to analyze a combustion reaction. Determine the change in temperature for the calorimeter when 10.5 g of diesel is burned inside the calorimeter. (ΔH_{comb}(diesel) = -44.8 kJ/g)

$$20.3\,kJ = C_{cal}\,\Delta T = C_{cal}(4.9°C) \qquad C_{cal} = 4.14\,\frac{kJ}{°C}$$

$$q_{cal} + q_{rxn} = 0$$

$$4.14\,\frac{kJ}{°C}\,\Delta T + \left(-44.8\,\frac{kJ}{g} \cdot 10.5g\right) = 0$$

$$\boxed{\Delta T = 113.6\,°C}$$

14) ***(2 pts)*** Acetylene, C_2H_2, is used as a fuel in welding because it produces a very hot flame.

$$2C_2H_2(g) + 5\,O_2(g) \rightarrow 4\,CO_2(g) + 2\,H_2O(l)$$ $\Delta H^o_{rxn} = \Sigma \Delta H^o_f(\text{prod}) - \Sigma \Delta H^o_f(\text{react})$

Determine the standard enthalpy change for the combustion reaction considering the following: ΔH^o_f = -227.0, -393.5, -285.8 kJ/mol for $C_2H_2(g)$, $CO_2(g)$ and $H_2O(l)$, respectively.

$$\Delta H^o_{rxn} = 4(-393.5) + 2(-285.8) - [2(-227)]$$

$$\boxed{\Delta H^o_{rxn} = -1691.6\ \text{KJ/mol}}$$

15) ***(2 pts)*** The chemistry of nitrogen oxides is very versatile. Determine the standard enthalpy for the following reaction:

$N_2O_3(g) + N_2O_5(s) \rightarrow 2N_2O_4(g)$ $\Delta H^o = ?$

Given the list of reactions below:

~~$NO(g) + NO_2(g) \rightarrow N_2O_3(g)$ $\Delta H^o = -39.8$ kJ/mol~~ × -1

~~$NO(g) + NO_2(g) + O_2(g) \rightarrow N_2O_5(g)$ $\Delta H^o = -112.5$ kJ/mol~~ × -1

~~$2NO_2(g) \rightarrow N_2O_4(g)$ $\Delta H^o = -57.2$ kJ/mol~~ × 2

~~$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$ $\Delta H^o = -114.2$ kJ/mol~~ × 1

~~$N_2O_5(s) \rightarrow N_2O_5(g)$ $\Delta H^o = 54.1$ kJ/mol~~ × 1

$4\cancel{NO_2} \rightarrow 2N_2O_4$ (-114.4)

$N_2O_5(s) \rightarrow \cancel{N_2O_5(g)}$ $(54.1\ \text{KJ/mol})$

$N_2O_3(g) \rightarrow \cancel{NO} + \cancel{NO_2}$ $(+39.8\ \text{KJ/mol})$

$\cancel{N_2O_5(g)} \rightarrow \cancel{NO} + \cancel{NO_2} + \cancel{O_2}$ $(+112.5\ \text{KJ/mol})$

$2\cancel{NO} + \cancel{O_2} \rightarrow 2\cancel{NO_2}$ (-114.2)

$N_2O_5(s) + N_2O_3(g) \rightarrow 2N_2O_4$ $\boxed{\Delta H_{rxn} = -22.2\ \text{KJ/mol}}$