Practice Final Exam

1. Calculate the amount of ***heat, in kJ***, needed to increase the temperature of 263.8 g of frozen phenol, C_6H_6OH, at -25.0°C (melting point is 40.5°C) to liquid at 65°C. s_l=1.43 J g^{-1} °C^{-1}, ΔH_{fus}=11.51 kJ mol^{-1},s_s=0.306 cal g^{-1} °C^{-1}.

Warm phenol to freezing point

$$q = ms\Delta T = (263.8\text{ g})(0.306\text{ cal g}^{-1}\ {}^\circ\text{C}^{-1})\left(\frac{4.184\text{ J}}{1\text{ cal}}\right)(40.5\ {}^\circ\text{C} - (-25.0\ {}^\circ\text{C}))\left(\frac{1\text{ kJ}}{10^3\text{ J}}\right) = 22.\underline{1}2224\text{ kJ}$$

Melt the phenol

$$q = n\Delta H_{fus} = (263.8\text{ g})\left(\frac{1\text{ mol}}{94.1112\text{ g}}\right)\left(\frac{11.51\text{ kJ}}{\text{mol}}\right) = 32.2\underline{6}328\text{ kJ}$$

Warm the liquid to the final temperature

$$q = ms\Delta T = (263.8\text{ g})(1.43\text{ J g}^{-1}\ {}^\circ\text{C}^{-1})(65\ {}^\circ\text{C} - 40.5\ {}^\circ\text{C}) = 9.\underline{2}42233\text{ kJ}$$

Total heat

$$q_{total} = 22.\underline{1}2224\text{ kJ} + 32.2\underline{6}328\text{ kJ} + 9.\underline{2}42233\text{ kJ} = 63.6\text{ kJ}$$

2. Quantum Chemistry:

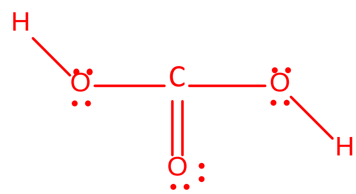
a. Give the ***electron configuration*** and ***orbital diagram*** of molybdenum using the Noble gas shorthand.

Mo: [Kr]$5s^1$ $4d^5$

Mo: [Kr] ↑ ↑ ↑ ↑ ↑ ↑

5s 4d

b. Draw the ***electron dot structure*** of H_2CO_3 (hint: the hydrogens are attached to the oxygens). Give the electron ***group*** and ***molecular geometries***. Give the ***hybridization*** on the central atom. Is the ion ***polar***?



EGG = MG = trigonal planar

Hybridization = sp^2

Non-polar

3. Hydrogen is prepared from natural gas (mainly methane, CH_4) by partial oxidation.

$$2\ CH_{4(g)} + O_{2(g)} \rightarrow 2\ CO_{(g)} + 4\ H_{2(g)}$$

Calculate the enthalpy change $\Delta H°$ for this reaction, using standard enthalpies of formation.

$$\Delta H = \sum_{\text{products}} n\Delta H_f^\circ - \sum_{\text{reactants}} n\Delta H_f^\circ$$

$$= \left[\left(\frac{2\text{ mol CO}}{\text{mol rxn}}\right)\left(\frac{-110.525\text{ kJ}}{\text{mol CO}}\right) + \left(\frac{4\text{ mol H}_2}{\text{mol rxn}}\right)\left(\frac{0.00\text{ kJ}}{\text{mol H}_2}\right)\right] - \left[\left(\frac{2\text{ mol CH}_4}{\text{mol rxn}}\right)\left(\frac{-74.81\text{ kJ}}{\text{mol CH}_4}\right) + \left(\frac{1\text{ mol O}_2}{\text{mol rxn}}\right)\left(\frac{0.00\text{ kJ}}{\text{mol O}_2}\right)\right]$$

$$= -71.43\text{ kJ}$$

4. A compound of carbon, hydrogen, and oxygen was burned in oxygen, and 1.000 g of the compound produced 1.434 g CO_2 and 0.783 g H_2O. In another experiment, 0.1107 g of the compound was dissolved in 25.0 g of water. This solution had a freezing point of –0.0894°C. What is the molecular formula of the compound? K_f = 1.86°C m^{-1}

Molar mass:

$$\Delta T_f = -K_f m$$

$$m = \frac{-\Delta T_f}{K_f} = -\frac{(-0.0894\ ^\circ\text{C})}{1.86\ ^\circ\text{C m}^{-1}} = 0.048\underline{0}6\ \text{m}$$

$$n = 25.0\ \text{g} \times \frac{1\ \text{kg}}{10^3\ \text{g}} \times \frac{0.048\underline{0}6\ \text{mol}}{\text{kg}} = 0.0012015\ \text{mol}$$

$$\text{molar mass} = \frac{0.1107\ \text{g}}{0.00120\ \text{mol}} = 92.25\ \text{g mol}^{-1}$$

Empirical formula:

$$C: 1.434\ \text{g}\ CO_2 \times \frac{1\ \text{mol}\ CO_2}{44.0095\ \text{g}\ CO_2} \times \frac{1\ \text{mol C}}{1\ \text{mol}\ CO_2} = 0.0325\underline{8}387\ \text{mol C}$$

$$\times \frac{12.0107\ \text{g C}}{1\ \text{mol C}} = 0.3913\ \text{g C}$$

$$H: 0.783\ \text{g}\ H_2O \times \frac{1\ \text{mol}\ H_2O}{18.0153\ \text{g}\ H_2O} \times \frac{2\ \text{mol H}}{1\ \text{mol}\ H_2O} = 0.086\underline{9}2611\ \text{mol H}$$

$$\times \frac{1.00794\ \text{g H}}{1\ \text{mol H}} = 0.0876\ \text{g H}$$

$$\text{O}: 1.000\ \text{g} - 0.3913\ \text{g} - 0.0876\ \text{g} = 0.52\underline{1}1\ \text{g O} \times \frac{1\ \text{mol O}}{15.9994\ \text{g O}} = 0.032\underline{5}6895\ \text{mol O}$$

$$\frac{0.0325\underline{8}387\ \text{mol}}{0.032\underline{5}6895\ \text{mol}} = 1\ \text{C} \times 3 = 3$$

$$\frac{0.086\underline{9}2611\ \text{mol}}{0.032\underline{5}6895\ \text{mol}} = 2.66\ \text{H} \times 3 = 8$$

$$\frac{0.032\underline{5}6895\ \text{mol}}{0.032\underline{5}6895\ \text{mol}} = 1\ \text{O} \times 3 = 3$$

$$\text{Empirical mass} = 92.0938\ \text{g mol}^{-1}$$

$$n = \frac{92.25}{92.0938} = 1$$

$$\text{molecular formula} = C_3H_8O_3$$

5. Polonium is the only metal that has a simple cubic unit cell structure. The density of polonium is 9.31 g cm^{-3}. Calculate the ***edge length*** of the unit cell in pm. Calculate the ***atomic radius*** in pm.

$$\frac{cm^3}{9.31\ \text{g}} \times \frac{209\ \text{g}}{\text{mol}} \times \frac{\text{mol}}{6.022 \times 10^{23}\ \text{at}} \times \frac{1\ \text{at}}{\text{cell}} \times \left(\frac{10^{-2}\ \text{m}}{\text{cm}}\right)^3 \times \left(\frac{\text{pm}}{10^{-12}\ \text{m}}\right)^3 = 3.7\underline{2}7827 \times 10^7\ \text{pm}^3/\text{cell}$$

$$\text{edge length} = \sqrt[3]{3.7\underline{2}7827 \times 10^7\ \text{pm}^3} = 33\underline{4}.05\ \text{pm} = 334\ \text{pm}$$

$$e = 2r \Rightarrow r = \frac{e}{2} = \frac{334.05\ \text{pm}}{2} = 16\underline{7}.03\ \text{pm} = 167\ \text{pm}$$

6. Ammonia will burn in the presence of a platinum catalyst to produce nitric oxide, NO.

$$4\ NH_{3\,(g)} + 5\ O_{2\,(g)} \rightarrow 4\ NO_{(g)} + 6\ H_2O_{(g)}$$

What is the heat of reaction at constant pressure? Use the following thermochemical equations:

$$2(N_{2\,(g)} + O_{2\,(g)} \rightarrow 2\ NO_{(g)};\ \Delta H = 180.6\ kJ)$$
$$-2(N_{2\,(g)} + 3\ H_{2\,(g)} \rightarrow 2\ NH_{3\,(g)};\ \Delta H = -91.8\ kJ)$$
$$3(2\ H_{2\,(g)} + O_{2\,(g)} \rightarrow 2\ H_2O_{(g)};\ \Delta H = -483.7\ kJ)$$

$$\Delta H = (2)(+180.6\text{ kJ}) + (-2)(-91.8\text{ kJ}) + (3)(-438.7\text{ kJ}) = -771.3\text{ kJ}$$

7. A metal, M, was converted to the chloride, MCl_2. Then a solution of the chloride was treated with silver nitrate to give silver chloride crystals, which were filtered from the solution.

$$MCl_{2\,(aq)} + 2\ AgNO_{3\,(aq)} \rightarrow M(NO_3)_{2\,(aq)} + 2\ AgCl_{(s)}$$

If 2.434 g of the metal gave 7.964 g of silver chloride, what is the atomic weight of the metal? What is the metal?

$$\frac{2.434\text{ g M}}{7.964\text{ g AgCl}} \times \frac{143.321\text{ g AgCl}}{1\text{ mol AgCl}} \times \frac{2\text{ mol AgCl}}{1\text{ mol MCl}_2} \times \frac{1\text{ mol MCl}_2}{1\text{ mol M}} = 87.61\text{ g mol}^{-1}$$

The metal is strontium.

8. Lithium bromide and lithium chromate react in an oxidation-reduction reaction in nitric acid. Two of the products of the reaction are the molecular bromine and the chromium(II) ion. Write the ***balanced complete chemical equation***. If 24.55 mL of 0.05214 M lithium bromide solution reacts with 54.33 mL of 0.1244 M lithium chromate solution, what ***mass of bromine*** can be obtained? If only 0.0996 g of bromine are produced, what is the ***percent yield*** of the reaction?

$$LiBr + Li_2CrO_4 \rightarrow Br_2 + Cr^{2+}$$
$$2(2\ LiBr \rightarrow Br_2 + 2\ Li^+ + 2\ e^-)$$
$$4\ e^- + 8\ H^+ + Li_2CrO_4 \rightarrow Cr^{2+} + 2\ Li^+ + 4\ H_2O$$

$$4\ e^- + 8\ H^+ + 4\ LiBr + Li_2CrO_4 \rightarrow 2\ Br_2 + Cr^{2+} + 6\ Li^+ + 4\ H_2O + 4\ e^-$$
$$8\ H^+ + 4\ LiBr + Li_2CrO_4 \rightarrow 2\ Br_2 + Cr^{2+} + 6\ Li^+ + 4\ H_2O$$
$$+8\ NO_3^- \qquad\qquad\qquad\qquad + 8\ NO_3^-$$

$$8\ HNO_3 + 4\ LiBr + Li_2CrO_4 \rightarrow 2\ Br_2 + Cr(NO_3)_2 + 6\ LiNO_3 + 4\ H_2O$$

$$24.55\text{ mL LiBr} \times \frac{0.05214\text{ mol LiBr}}{1000\text{ mL LiBr}} \times \frac{2\text{ mol Br}_2}{4\text{ mol LiBr}} \times \frac{159.808\text{ g Br}_2}{1\text{ mol Br}_2} = 0.1022\text{ g Br}_2$$

$$54.33\text{ mL Na}_2\text{Cr}_2\text{O}_7 \times \frac{0.1244\text{ mol Li}_2\text{CrO}_4}{1000\text{ mL Li}_2\text{CrO}_4} \times \frac{2\text{ mol Br}_2}{1\text{ mol Li}_2\text{CrO}_4} \times \frac{159.808\text{ g Br}_2}{1\text{ mol Br}_2} = 2.160\text{ g Br}_2$$

0.1022 g of bromine can be produced.

$$\%\text{yield} = \frac{\textit{actual}}{\textit{theoretical}} \times 100 = \frac{0.0996}{0.1022} \times 100 = 94.5\%$$