

CHM 1311 - E

First Midterm

Oct 14 – 2015

(Prof. S. Gambarotta)

Your Name: _____

Student #: _____

1. The solution key will be posted today on the web. Solutions will be worked out in the next DGD.
2. You must respond to all exercises.
3. Periodic Table and important values are at the end.
4. A few scratch sheets are at the very end. Report in the booklet the minimum amount of calculations to show your reasoning.

1. (1 point)

Calculate the number of oxygen atoms in 38.0 g of oxygen gas.

Plan: Divide the mass given by the molar mass of O_2 to find amount (mol). Since 1 mole of oxygen molecules contains 2 moles of oxygen atoms, multiply the amount (mol) by 2 to obtain amount (mol) of atoms and then multiply by Avogadro's number to obtain the number of atoms.

Solution:

$$\text{Amount (mol) of } O_2 = (38.0 \text{ g } O_2) \left(\frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} \right) = 1.1875 \text{ mol } O_2$$

$$\text{Amount (mol) of O atoms} = (1.1875 \text{ mol } O_2) \left(\frac{2 \text{ mol O atoms}}{1 \text{ mol } O_2} \right) = 2.375 \text{ mol O atoms}$$

$$\text{Number of O atoms} = (2.375 \text{ mol O atoms}) \left(\frac{6.022 \times 10^{23} \text{ O atoms}}{1 \text{ mol O atoms}} \right) = 1.430225 \times 10^{24} = 1.43 \times 10^{24} \text{ O atoms}$$

2. (5 points)

Calculate the number of atoms of Fe in one molecule of hemoglobin, knowing that iron makes up 0.33% of the mass of hemoglobin and whose molar mass is 6.8×10^4 g/mol

Plan: If the molecular formula for hemoglobin (Hb) were known, the number of Fe^{2+} ions in a molecule of hemoglobin could be calculated. It is possible to calculate the mass of iron from the percentage of iron and the molar mass of the compound. Assuming you have 1 mole of hemoglobin, take 0.33% of its molar mass as the mass of Fe in that 1 mole. Divide the mass of Fe by its molar mass to find amount (mol) of Fe in 1 mole of hemoglobin which is also the number of ions in 1 molecule.

Solution:

$$\text{Mass of Fe} = \left(\frac{0.33\% \text{ Fe}}{100\% \text{ Hb}} \right) \left(\frac{6.8 \times 10^4 \text{ g}}{\text{mol}} \right) = 224.4 \text{ g Fe}$$

$$\text{Amount (mol) of Fe} = (224.4 \text{ g Fe}) \left(\frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} \right) = 4.0179 \text{ mol} = 4.0 \text{ mol } Fe^{2+}/\text{mol Hb}$$

Thus, there are **4 Fe^{2+} /molecule Hb**.

3. (4 points)

A sample of 0.1595 g of menthol (Molar mass = 156.3) was analyzed in a combustion analyzer and yielded 0.449 g CO_2 and 0.184 g of H_2O . What is menthol's molecular formula?

Plan: In combustion analysis, finding the amount (mol) of carbon and hydrogen is relatively simple because all of the carbon present in the sample is found in the carbon of CO_2 , and all of the hydrogen present in the sample is found in the hydrogen of H_2O . Convert the mass of CO_2 to moles and use the ratio between CO_2 and C to find the amount (mol) and mass of C present. Do the same to find the amount (mol) and mass of H from H_2O . The amount (mol) of oxygen are more difficult to find, because additional O_2 was added to cause the combustion reaction. Subtracting the masses of C and H from the mass of the sample gives the mass of O. Convert the mass of O to moles of O. Take the moles of C, H, and O and divide by the smallest value to convert to whole numbers to get the empirical formula. Determine the empirical formula mass and compare it to the molar mass given in the problem to see how the empirical and molecular formulas are related. Finally, determine the molecular formula.

Solution:

$$\text{Amount (mol) of C} = (0.449 \text{ g CO}_2) \left(\frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \right) \left(\frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \right) = 0.010202 \text{ mol C}$$

$$\text{Mass (g) of C} = (0.010202 \text{ mol C}) \left(\frac{12.01 \text{ g C}}{1 \text{ mol C}} \right) = 0.122526 \text{ g C}$$

$$\text{Amount (mol) of H} = (0.184 \text{ g H}_2\text{O}) \left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) \left(\frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \right) = 0.020422 \text{ mol H}$$

$$\text{Mass (g) of H} = (0.020422 \text{ mol H}) \left(\frac{1.008 \text{ g H}}{1 \text{ mol H}} \right) = 0.020585 \text{ g H}$$

$$\begin{aligned} \text{Mass (g) of O} &= \text{Sample mass} - (\text{mass of C} + \text{mass of H}) \\ &= 0.1595 \text{ g} - (0.122526 \text{ g C} + 0.020585 \text{ g H}) = 0.016389 \text{ g O} \end{aligned}$$

$$\text{Amount (mol) of O} = (0.016389 \text{ g O}) \left(\frac{1 \text{ mol O}}{16.00 \text{ g O}} \right) = 0.0010243 \text{ mol O}$$

$$\text{Preliminary formula} = \text{C}_{0.010202}\text{H}_{0.020422}\text{O}_{0.0010243}$$

Converting to integer subscripts (dividing all by the smallest subscript):

$$\begin{array}{ccc} \text{C}_{0.010202} & \text{H}_{0.020422} & \text{O}_{0.0010243} \\ \hline \frac{0.010202}{0.0010243} & \frac{0.020422}{0.0010243} & \frac{0.0010243}{0.0010243} \end{array} \rightarrow \text{C}_{10}\text{H}_{20}\text{O}_1$$

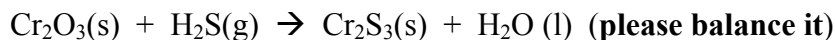
$$\text{Empirical formula} = \text{C}_{10}\text{H}_{20}\text{O}$$

$$\text{Empirical formula mass} = 10(12.01 \text{ g/mol C}) + 20(1.008 \text{ g/mol H}) + 1(16.00 \text{ g/mol O}) = 156.26 \text{ g/mol}$$

The empirical formula mass is the same as the given molar mass so the empirical and molecular formulas are the same. The molecular formula is $\text{C}_{10}\text{H}_{20}\text{O}$.

4. (1 point)

Chromium oxide reacts with H_2S according to the following equation:



To produce 421 g of $\text{Cr}_2\text{S}_3(\text{s})$ how many (a) moles and (b) grams of $\text{Cr}_2\text{O}_3(\text{s})$ are required?

Plan: Convert mass of Cr_2S_3 to amount (mol) by dividing by its molar mass. Use the mole ratio between Cr_2S_3 and Cr_2O_3 from the balanced chemical equation to determine the amount (mol) of Cr_2O_3 required. Multiply the amount (mol) of Cr_2O_3 by its molar mass to obtain the mass in grams.

Solution:

$$\text{a) Amount (mol) of } \text{Cr}_2\text{S}_3 = (421 \text{ g } \text{Cr}_2\text{S}_3) \left(\frac{1 \text{ mol } \text{Cr}_2\text{S}_3}{200.21 \text{ g } \text{Cr}_2\text{S}_3} \right) = 2.102792 \text{ mol } \text{Cr}_2\text{S}_3$$

$$\text{Amount (mol) of } \text{Cr}_2\text{O}_3 = (2.102792 \text{ mol } \text{Cr}_2\text{S}_3) \left(\frac{1 \text{ mol } \text{Cr}_2\text{O}_3}{1 \text{ mol } \text{Cr}_2\text{S}_3} \right) = 2.102792 \text{ mol} = \mathbf{2.10 \text{ mol } \text{Cr}_2\text{O}_3}$$

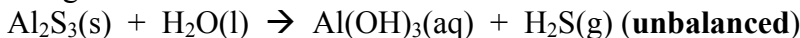
$$\text{b) Mass (g) of } \text{Cr}_2\text{O}_3 = (2.102792 \text{ mol } \text{Cr}_2\text{O}_3) \left(\frac{152.00 \text{ g } \text{Cr}_2\text{O}_3}{1 \text{ mol } \text{Cr}_2\text{O}_3} \right) = 319.624 \text{ g} = \mathbf{3.20 \times 10^2 \text{ g } \text{Cr}_2\text{O}_3}$$

Combining all steps gives:

$$\begin{aligned} \text{Mass (g) of } \text{Cr}_2\text{O}_3 &= (421 \text{ g } \text{Cr}_2\text{S}_3) \left(\frac{1 \text{ mol } \text{Cr}_2\text{S}_3}{200.21 \text{ g } \text{Cr}_2\text{S}_3} \right) \left(\frac{1 \text{ mol } \text{Cr}_2\text{O}_3}{1 \text{ mol } \text{Cr}_2\text{S}_3} \right) \left(\frac{152.00 \text{ g } \text{Cr}_2\text{O}_3}{1 \text{ mol } \text{Cr}_2\text{O}_3} \right) \\ &= 319.624 \text{ g} = \mathbf{3.20 \times 10^2 \text{ g } \text{Cr}_2\text{O}_3} \end{aligned}$$

5. (3 points)

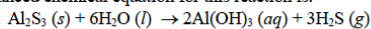
Calculate the amount of H_2S that can form when 158 g of aluminum sulfide reacts with 131 g of water.



Plan: First, balance the chemical equation. To determine which reactant is limiting, calculate the amount of H_2S formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the amount of H_2S formed and the amount of the excess reactant that reacts. The difference between the amount of excess reactant that reacts and the initial amount of reactant supplied gives the amount of excess reactant remaining.

Solution:

The balanced chemical equation for this reaction is:



Finding the amount (mol) of H_2S from the amount (mol) of Al_2S_3 (if H_2O is limiting):

$$\text{Amount (mol) of } \text{Al}_2\text{S}_3 = (158 \text{ g } \text{Al}_2\text{S}_3) \left(\frac{1 \text{ mol } \text{Al}_2\text{S}_3}{150.17 \text{ g } \text{Al}_2\text{S}_3} \right) = 1.05214 \text{ mol } \text{Al}_2\text{S}_3$$

$$\text{Amount (mol) of } \text{H}_2\text{S} \text{ from } \text{Al}_2\text{S}_3 = (1.05214 \text{ mol } \text{Al}_2\text{S}_3) \left(\frac{3 \text{ mol } \text{H}_2\text{S}}{1 \text{ mol } \text{Al}_2\text{S}_3} \right) = 3.15642 \text{ mol} = 3.16 \text{ mol } \text{H}_2\text{S}$$

Finding the amount (mol) of H_2S from the amount (mol) of H_2O (if Al_2S_3 is limiting):

$$\text{Amount (mol) of } \text{H}_2\text{O} = (131 \text{ g } \text{H}_2\text{O}) \left(\frac{1 \text{ mol } \text{H}_2\text{O}}{18.02 \text{ g } \text{H}_2\text{O}} \right) = 7.26970 \text{ mol } \text{H}_2\text{O}$$

$$\text{Amount (mol) of } \text{H}_2\text{S} \text{ from } \text{H}_2\text{O} = (7.26970 \text{ mol } \text{H}_2\text{O}) \left(\frac{3 \text{ mol } \text{H}_2\text{S}}{6 \text{ mol } \text{H}_2\text{O}} \right) = 3.63485 \text{ mol} = 3.63 \text{ mol } \text{H}_2\text{S}$$

Al_2S_3 is the limiting reagent and **3.16 mol of H_2S** will form.

$$\text{Mass (g) of } \text{H}_2\text{S} = (3.15642 \text{ mol } \text{H}_2\text{S}) \left(\frac{34.09 \text{ g } \text{H}_2\text{S}}{1 \text{ mol } \text{H}_2\text{S}} \right) = 107.602 \text{ g} = \mathbf{108 \text{ g } \text{H}_2\text{S}}$$

Combining all steps gives:

$$\text{Grams } \text{H}_2\text{S} = (158 \text{ g } \text{Al}_2\text{S}_3) \left(\frac{1 \text{ mol } \text{Al}_2\text{S}_3}{150.17 \text{ g } \text{Al}_2\text{S}_3} \right) \left(\frac{3 \text{ mol } \text{H}_2\text{S}}{1 \text{ mol } \text{Al}_2\text{S}_3} \right) \left(\frac{34.09 \text{ g } \text{H}_2\text{S}}{1 \text{ mol } \text{H}_2\text{S}} \right) = 107.602 \text{ g} = \mathbf{108 \text{ g } \text{H}_2\text{S}}$$

The remaining mass of the excess reagent can be calculated from the amount of H_2O combining with the limiting reagent.

$$\begin{aligned} \text{Amount (mol) of } \text{H}_2\text{O} \text{ required to react with 158 g of } \text{Al}_2\text{S}_3 &= (1.05214 \text{ mol } \text{Al}_2\text{S}_3) \left(\frac{6 \text{ mol } \text{H}_2\text{O}}{1 \text{ mol } \text{Al}_2\text{S}_3} \right) \\ &= 6.31284 \text{ mol } \text{H}_2\text{O} \end{aligned}$$

$$\text{Mass (g) of } \text{H}_2\text{O} \text{ required to react with 158 g of } \text{Al}_2\text{S}_3 = (6.31284 \text{ mol } \text{H}_2\text{O}) \left(\frac{18.02 \text{ g } \text{H}_2\text{O}}{1 \text{ mol } \text{H}_2\text{O}} \right) = 113.757 \text{ g } \text{H}_2\text{O}$$

$$\text{Mass (g) of remaining } \text{H}_2\text{O} = 131 \text{ g } \text{H}_2\text{O} - 113.757 \text{ g } \text{H}_2\text{O} = 17.243 \text{ g} = \mathbf{17 \text{ g } \text{H}_2\text{O}}$$

6. (3 points)

Calculate the mass (g) of calcium nitrate in a milliliter of a solution prepared by diluting 64.0 mL of 0.745 mmol/L calcium nitrate to a final volume of 0.100 L

$$d) c_1 = 0.745 \text{ mol/L} \quad V_1 = (64.0 \text{ mL}) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) = 0.0640 \text{ L} \quad c_2 = ? \quad V_2 = 0.100 \text{ L}$$

$$c_1 V_1 = c_2 V_2$$

$$c_2 = \frac{c_1 \times V_1}{V_2} = \frac{(0.745 \text{ mol/L})(0.0640 \text{ L})}{(0.100 \text{ L})} = 0.4768 \text{ mol/L}$$

The concentration in concentration (mol/L) must now be converted to grams per millilitre.

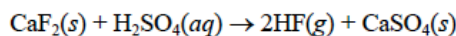
$$\left(\frac{0.4768 \text{ mol Ca(NO}_3)_2}{\text{L}} \right) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) \left(\frac{164.10 \text{ g Ca(NO}_3)_2}{1 \text{ mol Ca(NO}_3)_2} \right) = 0.07824 \text{ g/mL} = 0.0782 \text{ g Ca(NO}_3)_2/\text{mL}$$

7. (3 points)

When 15.0 g of CaF_2 reacts with excess of H_2SO_4 , HF gas is collected at 97.9 kPa and 25.5°C. Solid CaSO_4 is the other product of the reaction. What temperature is required to store the gas in a 8.63 L container at 1.17 bar?

Plan: Write a balanced equation. Convert mass of CaF_2 to amount in moles and use the molar ratio from the balanced equation to find the amount in moles of gas produced. Use the ideal gas law to find the temperature required to store that amount of HF gas at the given conditions of temperature and pressure.

Solution:



$$\text{amount in moles HF gas} = (15.0 \text{ g CaF}_2) \left(\frac{1 \text{ mol CaF}_2}{78.08 \text{ g CaF}_2} \right) \left(\frac{2 \text{ mol HF}}{1 \text{ mol CaF}_2} \right) = 0.3842213 \text{ mol HF}$$

Finding the temperature:

$$V = 8.63 \text{ L}$$

$$T = \text{unknown}$$

$$p = 1.17 \text{ bar}$$

$$n = 0.3842213 \text{ mol}$$

$$pV = nRT$$

Solving for T:

$$T = \frac{pV}{nR} = \frac{(1.17 \text{ bar})(8.63 \text{ L})}{(0.3842213 \text{ mol HF}) \left(0.08314 \frac{\text{L}\cdot\text{bar}}{\text{mol}\cdot\text{K}} \right)} = 316.0860 \text{ K}$$

The gas must be heated to **316 K**.

8. (1 point)

What is the concentration (mol/L) of a solution formed when 10.0L of ammonia gas at 33°C and 96.7kPa dissolves in enough water to give a final volume of 0.750L?

Plan: Use the ideal gas law to find the amount in moles of ammonia gas in 10.0 L at this pressure and temperature. Concentration (mol/L) is amount (mol) per litre. Use the amount in moles of ammonia and the final volume of solution (0.750 L) to get the concentration (mol/L).

Solution:

$$\begin{aligned} V &= 10.0 \text{ L} & T &= 33^\circ\text{C} + 273 = 306 \text{ K} \\ p &= 96.7 \text{ kPa} & n &= \text{unknown} \\ pV &= nRT \end{aligned}$$

Solving for n :

$$\text{amount in moles of ammonia} = n = \frac{pV}{RT} = \frac{(96.7 \text{ kPa})(10.0 \text{ L})}{\left(8.31446 \frac{\text{L}\cdot\text{kPa}}{\text{mol}\cdot\text{K}}\right)(306 \text{ K})} = 0.380076 \text{ mol}$$

$$\text{concentration (mol/L)} = c = \frac{\text{amount(mol) ammonia}}{V(\text{L}) \text{ of solution}} = \frac{0.380076 \text{ mol}}{0.750 \text{ L}} = 0.50677 \text{ mol/L} = \mathbf{0.507 \text{ mol/L}}$$

9. (5 points)

When gaseous F_2 and solid I_2 are heated to high temperature, I_2 sublimes and gaseous IF_5 forms. If 46.7 kPa of F_2 and 2.50 g of solid I_2 are placed into a 2.50 L container at 250K and the container is heated to 550K, what is the final total pressure (bar) and partial pressure of I_2 gas?

Plan: Since the amounts of two reactants are given, this is a limiting reactant problem. Write the balanced equation and use molar ratios to find the amount in moles of IF_7 produced by each reactant. The mass of I_2 is converted to amount in moles using its molar mass and the amount in moles of F_2 is found using the ideal gas law. The smaller amount in moles of product indicates the limiting reagent. Determine the amount in moles of excess reactant gas and the amount in moles of product gas and use the ideal gas law to solve for the total pressure.

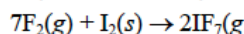
Solution:

amount in moles of F_2 :

$$\begin{aligned} V &= 2.50 \text{ L} & T &= 250. \text{ K} \\ p &= 46.7 \text{ kPa} & n &= \text{unknown} \\ pV &= nRT \end{aligned}$$

Solving for n :

$$n = \frac{pV}{RT} = \frac{(46.7 \text{ kPa})(2.50 \text{ L})}{\left(8.31446 \frac{\text{L}\cdot\text{kPa}}{\text{mol}\cdot\text{K}}\right)(250. \text{ K})} = 0.0561672 \text{ mol } \text{F}_2$$



$$\text{amount in moles } \text{IF}_7 \text{ from } \text{F}_2 = (0.0561672 \text{ mol } \text{F}_2) \left(\frac{2 \text{ mol } \text{IF}_7}{7 \text{ mol } \text{F}_2} \right) = 0.01604777 \text{ mol } \text{IF}_7 \text{ (unrounded)}$$

$$\text{amount in moles } \text{IF}_7 \text{ from } \text{I}_2 = (2.50 \text{ g } \text{I}_2) \left(\frac{1 \text{ mol } \text{I}_2}{253.8 \text{ g } \text{I}_2} \right) \left(\frac{2 \text{ mol } \text{IF}_7}{1 \text{ mol } \text{I}_2} \right) = 0.019700551 \text{ mol } \text{IF}_7 \text{ (unrounded)}$$

F_2 is limiting. All of the F_2 is consumed.

amount in mole I_2 remaining = original amount of moles of I_2 – amount of I_2 (mol) reacting with F_2

$$\text{amount in mole } \text{I}_2 \text{ remaining} = (2.50 \text{ g } \text{I}_2) \left(\frac{1 \text{ mol } \text{I}_2}{253.8 \text{ g } \text{I}_2} \right) - (0.0561672 \text{ mol } \text{F}_2) \left(\frac{1 \text{ mol } \text{I}_2}{7 \text{ mol } \text{F}_2} \right) = 1.82639 \times 10^{-3} \text{ mol } \text{I}_2$$

$$\begin{aligned} \text{Total amount in moles of gas} &= (0 \text{ mol } \text{F}_2) + (0.01604777 \text{ mol } \text{IF}_7) + (1.82639 \times 10^{-3} \text{ mol } \text{I}_2) \\ &= 0.01787416 \text{ mol gas} \end{aligned}$$

$$\begin{aligned} V &= 2.50 \text{ L} & T &= 550. \text{ K} \\ p &= \text{unknown} & n &= 0.01787416 \text{ mol} \\ pV &= nRT \end{aligned}$$

Solving for P :

$$p \text{ (kPa)} = \frac{nRT}{V} = \frac{(0.01787416 \text{ mol}) \left(8.31446 \frac{\text{L}\cdot\text{kPa}}{\text{mol}\cdot\text{K}} \right) (550. \text{ K})}{2.50 \text{ L}} = 32.6950 \text{ kPa}$$

$$p \text{ (bar)} = (32.6950 \text{ kPa}) \left(\frac{1 \text{ bar}}{100 \text{ kPa}} \right) = 0.32695 \text{ bar} = \mathbf{0.327 \text{ bar}}$$

$$p_{\text{iodine}} \text{ (bar)} = X_{\text{iodine}} P_{\text{total}} = [(1.82639 \times 10^{-3} \text{ mol I}_2) / (0.01787416 \text{ mol})] (0.32695 \text{ bar}) \\ = 0.033408 \text{ bar} = \mathbf{33.4 \times 10^{-3} \text{ bar}}$$

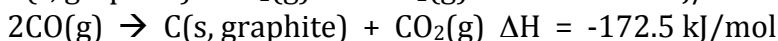
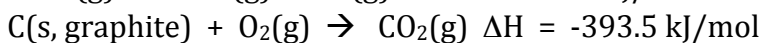
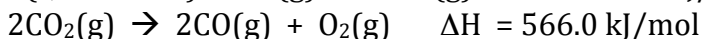
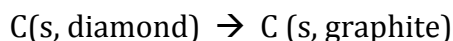
10. (3 points)

A 30.5 g sample of an alloy sample at 93.0 C is placed in 50.0 g of water at 22.0 C (heat capacity of water 4.184 J/g.K) of a calorimeter with heat capacity of 9.2 J/K (mass of the calorimeter is build in the value of the heat capacity). If the final temperature is 31.1 C what is the specific heat capacity of the alloy?

$$\begin{aligned} -q_{\text{lost}} &= q_{\text{gained}} = q_{\text{water}} + q_{\text{calorimeter}} \\ -(30.5 \text{ g alloy})(c_{\text{alloy}})(31.1 - 93.0)^\circ\text{C} &= (50.0 \text{ g H}_2\text{O})(4.184 \text{ J/g}\cdot\text{K})(31.1 - 22.0)^\circ\text{C} + (9.2 \text{ J/K})(31.1 - 22.0)^\circ\text{C} \\ -(30.5 \text{ g})(c_{\text{alloy}})(-61.9\text{K}) &= (50.0 \text{ g})(4.184 \text{ J/g}\cdot\text{K})(9.1\text{K}) + (9.2 \text{ J/K})(9.1\text{K}) \\ 1887.95 \text{ g K}(c_{\text{alloy}}) &= 1903.72 \text{ J} + 83.72 \text{ J} = 1987.44 \text{ J} \\ c_{\text{alloy}} &= 1987.44\text{J}/1887.95 \text{ g}\cdot\text{K} = 1.052697 \text{ J/g}\cdot\text{K} = \mathbf{1.1 \text{ J/g}\cdot\text{K}} \end{aligned}$$

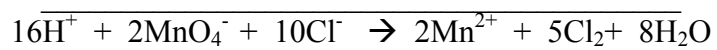
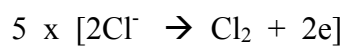
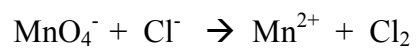
11. (3 points)

Using equations from the list below determine ΔH for this reaction”



12. (2 points)

Balance the following redox equation in acidic conditions (use either the semi-equation or the direct method. Whatever you feel more comfortable with):



Data For Water

Density = 1.00 g/mL (at 25°C)

 $s = 2.13 \text{ J g}^{-1} \text{ K}^{-1}$ (solid) $s = 4.184 \text{ J g}^{-1} \text{ K}^{-1}$ (liquid) $s = 2.01 \text{ J g}^{-1} \text{ K}^{-1}$ (gas) $\Delta H_{\text{fus}}^{\circ} = 6.02 \text{ kJ mol}^{-1}$ $\Delta H_{\text{vap}}^{\circ} = 40.7 \text{ kJ mol}^{-1}$ **Constants and Conversion Factors**

1 mmHg = 1 torr 760 mmHg = 1 atm 1 atm = 101.325 kPa 1 atm = 1.013125 bar

1 cm³ = 1 mL 1000 mL = 1 L 1000 L = 1 m³

Avogadro's Number	N	$6.022 \times 10^{23} \text{ mol}^{-1}$	
Boltzmann's constant	k	$1.30866 \times 10^{-23} \text{ J} \cdot \text{K}^{-1}$	
Faraday's constant	F	$96,485 \text{ C} \cdot \text{mol}^{-1}$	
Gas constant	R	$8.31451 \text{ J} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}$	
	R	$0.08206 \text{ atm} \cdot \text{L} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}$	
	R	$8.31451 \text{ m}^3 \cdot \text{Pa} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}$	
	R	$0.0831451 \text{ bar} \cdot \text{L} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}$	
Planck's constant	h	6.62608×10^{-34}	J·s
Speed of Light	c	2.99792458×10^8	m·s ⁻¹

Gas Laws

$$PV = nRT$$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$P_T = P_1 + P_2 + P_3 + \dots$$

$$d = \frac{m}{V} = \frac{P \cdot MM}{RT}$$

$$E_K = \frac{1}{2}mv^2$$

$$u_{rms} = \sqrt{\frac{3RT}{MM}}$$

$$\frac{\text{Rate A}}{\text{Rate B}} = \sqrt{\frac{MM_B}{MM_A}}$$

$$\left(P + \frac{n^2a}{V^2}\right)(V - nb) = nRT$$

Equilibrium

$$K_p = K_c(RT)^{\Delta n}$$

Acid/Base

$$pOH = -\log[OH^-]$$

$$pH = -\log[H^+]$$

$$pH + pOH = 14$$

$$K_a \times K_b = K_w$$

$$pH = pK_a + \log \frac{[A^-]}{[HA]}$$

$$pH = \frac{pK_{a1} + pK_{a2}}{2}$$

Thermochemistry

$$\Delta U = q + W$$

$$W_{\text{system}} = -P\Delta V = -\Delta nRT$$

$$\Delta H = \Delta U + P\Delta V$$

$$q_p = \Delta U + P\Delta V$$

$$q = ms\Delta T$$

$$\Delta H_{\text{rxn}}^\circ = \sum n\Delta H_f^\circ(\text{pds}) - \sum n\Delta H_f^\circ(\text{rxts})$$

The atom

$$E = hv$$

$$c = v\lambda$$

$$E = -B/n^2$$

Kinetics

$$[A]_t = [A]_o - kt$$

$$\ln[A]_t = \ln[A]_o - kt$$

$$1/[A]_t = 1/[A]_o + kt$$

$$k = Ae^{(-E_a/RT)}$$

$$\ln(k_2/k_1) = (-E_a/R)(1/T_2 - 1/T_1)$$

