AP Chemistry 2009 – Thanksgiving Day problem Set! GOBBLES!!!!!!

$$2\ Fe_{(s)} + \frac{3}{2}\ O_{2(g)} \rightarrow Fe_2O_{3(s)} \quad \Delta H_f^\circ = -824\ kJ\ mol^{-1}$$

1. Iron reacts with oxygen to produce iron(III) oxide as represented above. A 75.0 g sample of $Fe_{(s)}$ is mixed with 11.5 L of $O_{2(g)}$ at 2.66 atm and 298 K.

(a) Calculate the number of moles of each of the following before the reaction occurs.
   (i) $Fe_{(s)}$
   (ii) $O_{2(g)}$

(b) Identify the limiting reactant when the mixture is heated to produce Fe_2O_3. Support your answer with calculations.

(c) Calculate the number of moles of Fe_2O_3 produced when the reaction proceeds to completion.

(d) The standard free energy of formation, ΔG_f° of Fe_2O_3 is –740. kJ mol^{-1} at 298 K.
   (i) Calculate the standard entropy of formation ΔS_f° of Fe_2O_3 at 298 K. Include units with your answer.
   (ii) Which is more responsible for the spontaneity of the formation reaction at 298K, the standard enthalpy or the standard entropy?

The reaction represented below also produces iron(III) oxide. The value of ΔH° for the reaction is –280 kJ per mol.

$$2\ FeO_{(s)} + \frac{1}{2}\ O_{2(g)} \rightarrow Fe_2O_{3(s)}$$

(e) Calculate the standard enthalpy of formation, ΔH_f° of $FeO_{(s)}$.

$$BCl_{3(g)} + NH_{3(g)} \rightarrow Cl_3BNH_{3(s)}$$

2. The reaction represented above is a reversible reaction.

(a) Predict the sign of the entropy change, ΔS, as the reaction proceeds to the right. Explain your prediction.

(b) If the reaction spontaneously proceeds to the right, predict the sign of the enthalpy change, ΔH. Explain your prediction.

(c) The direction in which the reaction spontaneously proceeds changes as the temperature is increased above a specific temperature. Explain.

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3. Write full balanced net ionic equations for the following reaction:

(a) Zinc reacts with a solution of copper(II) nitrate

(b) Hypochlorous acid reacts with a solution potassium hydroxide

(c) Calcium carbonate decomposes upon heating

(d) Hydrogen sulfide gas is bubbled through a solution of silver nitrate

(e) A solution of potassium dichromate is added to an acidified solution of iron(II) chloride (hint: dichromate is a strong oxidizer and it reduces to Chromium (III)

4. The mass percent of MnO_2 in a sample of a mineral is determined by reacting it with a measured excess of As_2O_3 in acid solution, and then titrating the remaining As_2O_3 with standard $KMnO_4$. A 0.225 g sample of the mineral is ground and boiled with 75.0 mL of 0.0125 M As_2O_3 solution containing 10 mL of concentrated sulfuric acid. After the reaction is complete, the solution is cooled, diluted with water, and titrated with 2.28 $x10^{-3}$ M $KMnO_4$, requiring 16.34 mL to reach the endpoint.
Note: 5 mol of As_2O_3 react with 4 mol of MnO_4^-.

(a) Write a balanced equation for the reaction of As_2O_3 with MnO_2 in acid solution. The products are Mn^{2+} and AsO_4^{3-}.

(b) Calculate the number of moles of
i. As_2O_3 added initially.
ii. MnO_4^- used to titrate the excess As_2O_3.
iii. MnO_2 in the sample. Ans: 1.78 $x10^{-3}$ mol MnO_2

(c) Determine the mass percent of MnO_2 in the sample. Ans: 68.9% MnO_2 in sample

(d) Describe how the endpoint is detected in the $KMnO_4$ titration.