Exam 3 Study Guide

Here are the central topics for the final exam, along with some practice problems.

Chapter 9 (Sections 9.5 and 9.6): Reaction Enthalpy

Can you: (these topics were also on the Exam 2 study guide) interpret a reaction enthalpy change (endothermic versus exothermic)? combine reactions and their enthalpy changes using Hess's Law? understand and use enthalpies of formation (and formation reactions)?

For practice:
The measured standard molar enthalpy of reaction for burning the compound called *ketene*, CH_2CO, is –981.1.kJ mol^{-1} at 298 K:

$$CH_2CO(g) + 2\ O_2(g) \rightleftarrows 2\ CO_2(g) + H_2O(g) \qquad \Delta H_r^o = -981.1\ kJ\ mol^{-1}$$

Using data in Appendix 4 as well, calculate the enthalpy of reaction for

$$2\ CH_4(g) + 2\ O_2(g) \rightleftarrows CH_2CO(g) + 3\ H_2O(g)$$

Then use these two reactions to calculate the enthalpy of reaction for

$$CH_4(g) + 2\ O_2(g) \rightleftarrows CO_2(g) + 2\ H_2O(g)$$

Chapter 10: Spontaneity, Entropy, and Free Energy

Can you: describe entropy in terms of configurational (positional) probabilities? relate entropy changes in gases to temperature changes? relate absolute entropies to a compound's physical state, molar mass, and molecular complexity? relate enthalpy and entropy changes to free energy changes? use free energy changes to predict the spontaneous direction of a physical or chemical change? calculate reaction entropy changes from standard absolute entropies? use free energy of formation data? relate free energy changes to a reaction's equilibrium constant or reaction quotient? calculate an equilibrium constant at various temperatures?

For practice:
Predict in advance the sign of the reaction entropy change for the following reaction, then use Appendix 4 data to calculate its ΔS_r^o at 298 K:

$$N_2(g) + 2\ O_2(g) \rightleftarrows 2\ NO_2(g)$$

The molar constant pressure heat capacity of $O_2(g)$ is 29.36 J K^{-1} mol^{-1}. What is the entropy of $O_2(g)$ at 320 K? What is the equilibrium constant for this reaction at 298 K? at 320 K?

Chapter 11: Electrochemistry (and those sections in Chapter 4 on oxidation-reduction reactions)

Can you: assign oxidation numbers? balance oxidation-reduction reactions? describe an oxidation-reduction reaction in terms of its half-reactions? describe an electrochemical cell in terms of its constituent parts: its electrodes, its electrolytes, its interconnection (such as a salt bridge), its anode and cathode, its polarity, its oxidation and reduction half-reactions, and its net chemical reaction? use a table of standard reduction potentials to: calculate cell voltages, predict spontaneous reaction directions, and calculate new half-reaction reduction potentials? apply the Nernst equation

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to an electrochemical cell at arbitrary conditions? relate cell voltages to free energy changes and equilibrium constants? explain how electrolysis works?

For practice:
An electrochemical cell is constructed with a Zn electrode in a Zn^{2+} solution connected via a salt bridge to an acidic solution of permanganate ion, MnO_4^- and manganese(II) ion, Mn^{2+} into which a Pt wire is placed for electrical contact. What net reaction does this cell represent? Which electrode is positive? What is the half reaction at the cathode? What is the standard reduction potential for this cell? If the pH = 2.00 in the acidic solution, $[Zn^{2+}] = 0.015$ M, $[Mn^{2+}] = 0.0010$ M, and $[MnO_4^-] = 0.12$ M, what is the cell voltage? How many electrons are transferred in the net reaction? What is the standard free energy change of the net reaction? How does the cell voltage change if the pH is reduced to 1.00? If Zn and Zn^{2+} are replaced by Cu and Cu^{2+}, will the cell voltage increase or decrease?