Practice Test C (Ch 5, 19, 20.5−6) Thermo & more Electro

Name______________________Per___

- ***This is practice − Do NOT cheat yourself of finding out what you are capable of doing under pressure. Be sure you follow the testing conditions outlined below.***
- ***DO NOT USE A CALCULATOR. You may use ONLY your periodic table and formula sheets.***
- ***Try only for 1.5 minutes per question (51 questions = 77 min total). Time yourself. It is important that you practice working for speed. Then when time is up, continue working, (with a calculator if you must), but use a different color writing utensil so you will have a sense of how many you were able to complete under the time constraints and without the calculator.***

Questions 1 − 3 refer to the following

(A) ionization energy
(B) free energy
(C) activation energy
(D) lattice energy
(E) kinetic energy

1. The average energy will be the same for any ideal gas at a given temperature.

2. The energy released when gaseous ions combine to form an ionic solid.

3. The energy required to convert a gaseous atom, in the ground state, to a gaseous cation.

The next 4 questions refer to the following choices:

(A) specific heat capacity
(B) free energy
(C) ionization energy
(D) heat of vaporization
(E) heat of fusion

4. This always has a negative value for a reaction that occurs spontaneously

5. This energy is given off when a substance condenses.

6. This energy is taken in when a substance melts.

7. The energy released when cooling 1 g of a substance 1ºC.

8. $Ni_{(s)} + 2Cu^+ \rightarrow Ni^{2+} + 2Cu_{(s)}$

If the equilibrium constant for the reaction above is greater than 1, which of the following correctly describes the standard voltage, E^o, and the standard free energy change, ΔG^o, for this reaction?

a. E^o and ΔG^o are both positive
b. E^o and ΔG^o are both negative
c. E^o and ΔG^o are both zero
d. E^o is positive and ΔG^o is negative
e. E^o is negative and ΔG^o is positive

9. $Zn + Cu^{2+} \rightarrow Cu + Zn^{2+}$

A Voltaic cell based on the reaction represented above was constructed from zinc and copper half-cells. The observed voltage was found to be 1.20 volts instead of the standard cell potential, E^o, of 1.10 volts. Which of the following could correctly account for this observation?

a. The zinc electrode contained more mass than the copper electrode.
b. The Zn^{2+} electrolyte was 0.5 M $Zn(NO_3)_2$, whereas the Cu^{2+} electrolyte was 1.0 M $Cu(NO_3)_2$.
c. The Zn^{2+} solution was colorless, whereas the Cu^{2+} solution was blue.
d. The solutions in the half-cells began at different temperatures.
e. The salt bridge contained NaBr as the electrolyte.

10. Which of the following must be true for a reaction that proceeds in reverse from initial standard state conditions?

a. ΔG^o is positive and K_{eq} is greater than 1
b. ΔG^o is positive and K_{eq} is less than 1
c. ΔG^o is negative and K_{eq} is greater than 1
d. ΔG^o is negative and K_{eq} is less than 1
e. ΔG^o is equal to 0 and K_{eq} is equal to 1

11. A sample of gallium metal is sealed inside a well insulated, rigid container, and the temperature inside the container is at the melting point of gallium metal. What can be said about the energy and the entropy of the system? Assume the insulation prevents any energy change with the surroundings.

a. The total energy increases, and the total entropy will increase.
b. The total energy is constant, and the total entropy change is constant.
c. The total energy is constant, and the total entropy change will decrease.
d. The total energy is constant, and the total entropy change will increase.
e. The total energy decreases, and the total entropy change will decrease.

12. When ammonium nitrate dissolves in water the temperature drops. Which of the following conclusions is true?
    a. The hydration energies are very high.
    b. Ammonium nitrate produces an ideal solution in water.
    c. The heat of solution for ammonium nitrate is exothermic.
    d. Ammonium nitrate has a low lattice energy.
    e. Ammonium nitrate is more soluble in hot water.

13. Choose the reaction expected to have the greatest increase in entropy.
    a. $H_2O_{(g)} \rightarrow H_2O_{(L)}$
    b. $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$
    c. $Ca_{(s)} + H_{2(g)} \rightarrow CaH_{2(s)}$
    d. $2NH_{3(g)} \rightarrow 3H_{2(g)} + N_{2(g)}$
    e. $2KClO_{3(s)} \rightarrow 2KCl_{(s)} + 3O_{2(g)}$

14. $$2\ Fe^{2+} + Cu^{2+} \rightarrow 2\ Fe^{3+} + Cu_{(s)}$$

    If the equilibrium constant for the reaction above is 2.5×10^{-4}, which of the following correctly describes the standard voltage, E^o, and the standard free energy change, ΔG^o, for this reaction?
    a. E^o and ΔG^o are both zero
    b. E^o and ΔG^o are both positive
    c. E^o and ΔG^o are both negative
    d. E^o is positive and ΔG^o is negative
    e. E^o is negative and ΔG^o is positive

15. For which of the following processes would ΔS have a positive value?
    I. $MgCO_{3(s)} \rightarrow MgO_{(s)} + CO_{2(g)}$
    II. $Ba^{2+} + SO_4^{2-} \rightarrow BaSO_{4(s)}$
    III. $Cl_{2(g)} + C_3H_{6(g)} \rightarrow C_3H_6Cl_{2(g)}$
    a. I only
    b. I and II only
    c. I and III only
    d. II and III only
    e. I, II, and III

16. Under standard conditions calcium metal reacts readily with chlorine gas. What conclusions may be drawn from this fact?
    a. $K_{eq} < 1$, $\Delta G^o > 0$
    b. $K_{eq} > 1$ and $\Delta G^o = 0$
    c. $K_{eq} < 1$ and $\Delta G^o < 0$
    d. $K_{eq} > 1$ and $\Delta G^o < 0$
    e. $K_{eq} > 1$ and $\Delta G^o > 0$

17. $$SF_{4(g)} + F_{2(g)} \rightarrow SF_{6(g)}$$

    The reaction above is thermodynamically spontaneous at 298, but becomes non spontaneous at higher temperatures. Which of the following is true at 298 K?
    a. ΔG, ΔH, and ΔS are all positive
    b. ΔG, ΔH, and ΔS are all negative
    c. ΔG and ΔH are negative, but ΔS is positive
    d. ΔG and ΔS are negative, but ΔH is positive
    e. ΔG and ΔH are positive, but ΔS is negative

18. $$Zn + Pb^{2+} \rightarrow Zn^{2+} + Pb$$

    A Galvanic cell based on the reaction represented above was constructed from zinc and lead half-cells. The observed voltage was found to be 0.58 volts instead of the standard cell potential, E°, of 0.63 volts. Which of the following could correctly account for this observation?
    a. The lead electrode contained more mass than the zinc electrode.
    b. The Zn^{2+} electrolyte was 0.5 M $ZnSO_4$, whereas the Pb^{2+} electrolyte was 1.0 M $Pb(NO_3)_2$
    c. The Pb^{2+} solution was less concentrated than the Zn^{2+} solution.
    d. The solutions in the half-cells had different volumes
    e. The salt bridge contained KNO_3 as the electrolyte.

19. $$H_2O_{(L)} \rightarrow H_2O_{(s)}$$

    When water freezes at its normal freezing point, 0°C and 1 atmosphere, which of the following is true (V = volume)
    a. $\Delta H < 0$, $\Delta S > 0$, and $\Delta V > 0$
    b. $\Delta H < 0$, $\Delta S < 0$, and $\Delta V > 0$
    c. $\Delta H > 0$, $\Delta S < 0$, and $\Delta V < 0$
    d. $\Delta H > 0$, $\Delta S > 0$, and $\Delta V > 0$
    e. $\Delta H > 0$, $\Delta S > 0$, and $\Delta V < 0$

20. Calcium iodide, CaI_2, has a ΔH^o_{soln} of −104 kJ per mole. Which of the following is probably true when calcium iodide is dissolved in water?
    a. The energy to break the crystal structure must be less than the energy of solvating the ions.
    b. As the solution dissolves, the temperature of the solution decreases.
    c. The resulting solution has a high vapor pressure
    d. Heating the solution causes more salt to dissolve.
    e. The ΔG for the dissolving process is > 0

21. Which of the following reactions has the largest positive value of Δ*S*?

a. $2\ H_2S_{(g)} + SO_{2(g)} \rightarrow 3\ S_{(s)} + 2\ H_2O_{(g)}$

b. $2\ SO_{3(g)} \rightarrow 2\ SO_{2(g)} + O_{2(g)}$

c. $Mg_{(s)} + Cl_{2(g)} \rightarrow MgCl_{2(s)}$

d. $Fe_2O_{3(s)} + 3\ H_{2(g)} \rightarrow 2\ Fe_{(s)} + 3\ H_2O_{(g)}$

e. $H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow H_2O_{(L)}$

22. Consider the reactions below to answer the question.

$$2C_6H_{6(L)} + 15O_{2(g)} \rightarrow 12CO_{2(g)} + 6H_2O_{(L)}\ \ \Delta G^o = -6400\text{ kJ}$$

$$C_{(s)} + O_{2(g)} \rightarrow CO_{2(s)}\ \ \Delta G^o = -400\text{ kJ}$$

$$H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow H_2O_{(L)}\ \ \Delta G^o = -250\text{ kJ}$$

What is the standard free energy change for the reaction below, as calculated from the data above?

$$6\ C_{(s)} + 3\ H_{2(g)} \rightarrow C_6H_{6(L)}$$

a. −250 kJ

b. −100.0 kJ

c. −50.0 kJ

d. 50.0 kJ

e. 100.0 kJ

23. $N_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow N_2O_{(g)}$

A reaction is only spontaneous above 550ºC. Which of the following is true at 400ºC?

	Δ***G***	Δ***H***	Δ***S***
(A)	+	+	+
(B)	−	−	−
(C)	−	−	+
(D)	−	+	−
(E)	+	+	−

24. A voltaic cell contains one half−cell with a zinc electrode in a $Zn^{2+}_{(aq)}$ solution and a copper electrode in a $Cu^{2+}_{(aq)}$ solution. The reaction is shown below. At standard conditions, $E^o = 1.10$ V. Which condition below would cause the cell potential to be greater than 1.10 V?

$$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$$

a. 1.0 M $Zn^{2+}_{(aq)}$, 1.0 M $Cu^{2+}_{(aq)}$

b. 5.0 M $Zn^{2+}_{(aq)}$, 5.0 M $Cu^{2+}_{(aq)}$

c. 5.0 M $Zn^{2+}_{(aq)}$, 1.0 M $Cu^{2+}_{(aq)}$

d. 0.5 M $Zn^{2+}_{(aq)}$, 0.5 M $Cu^{2+}_{(aq)}$

e. 0.1 M $Zn^{2+}_{(aq)}$, 1.0 M $Cu^{2+}_{(aq)}$

25. Which of the following combinations is true when wax melts? (V = volume)

a. $\Delta H > 0, \Delta S > 0$, and $\Delta V > 0$

b. $\Delta H > 0, \Delta S > 0$, and $\Delta V < 0$

c. $\Delta H > 0, \Delta S < 0$, and $\Delta V > 0$

d. $\Delta H < 0, \Delta S < 0$, and $\Delta V < 0$

e. $\Delta H < 0, \Delta S > 0$, and $\Delta V > 0$

26. $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$

From the table below, determine the enthalpy change for the above reaction

bond	**average bond energy (kJ/mol^{-1})**
H−H	436
O = O	499
H−O	464

a. 464 kJ

b. 485 kJ

c. −485 kJ

d. 443 kJ

e. −443 kJ

27. $CO_{(g)} + 2H_{2(g)} \rightarrow CH_3OH_{(g)}$ ΔH = −91 kJ

Determine the ΔH for the reaction above if $CH_3OH_{(L)}$ were formed in the above reaction instead of $CH_3OH_{(g)}$. The Δ*H* of vaporization for CH_3OH is 37 kJ mol^{-1}.

a. −128 kJ

b. −58 kJ

c. +128 kJ

d. +54 kJ

e. −37 kJ

28. $Co^{2+} + 2e- \rightarrow Co$ $E^o = -0.28$ V

$Cd^{2+} + 2e- \rightarrow Cd$ $E^o = -0.40$ V

Given the above standard potential, estimate the approximate value of the equilibrium constant for the following reaction:

$$Cd + Co^{2+} \rightarrow Cd^{2+} + Co$$

Given: $E_{cell} = E^0_{cell} - \frac{0.06}{n}\log Q$

a. 10^{-4}

b. 10^{-2}

c. 10^{4}

d. 10^{16}

e. 10^{2}

For the following three questions, select from the following list of choices

I. $\Delta H = +$ and $\Delta S = +$
II. $\Delta H = +$ and $\Delta S = -$
III. $\Delta H = -$ and $\Delta S = -$
IV. $\Delta H = -$ and $\Delta S = +$

29. Which of the four processes above are spontaneous at all temperatures?
    a. I only
    b. II only
    c. III only
    d. IV only
    e. I, III, and IV only

30. For which of the four processes listed above, can spontaneity can change with varying temperatures
    a. I only
    b. II only
    c. III only
    d. IV only
    e. I and III only

31. Which of the four processes above is improbable at low temperature, but becomes more probable at high temperature
    a. I only
    b. II only
    c. III only
    d. IV only
    e. impossible to identify without more information

32. The standard enthalpy of formation for nitrogen dioxide is the enthalpy change of which reaction
    a. $N_{(g)} + O_{2(g)} \rightarrow NO_{2(g)}$
    b. $\frac{1}{2}N_{2(g)} + O_{2(g)} \rightarrow NO_{2(g)}$
    c. $\frac{1}{2}N_2O_{4(g)} \rightarrow 2\ NO_{2(g)}$
    d. $N_{2(g)} + 2O_{2(g)} \rightarrow N_2O_{4(g)}$
    e. $N_{2(g)} + 2O_{2(g)} \rightarrow 2NO_{2(g)}$

33. At a certain temperature $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$ has a ΔG of −339.4 kJ mol^{-1}. Which statements from the following list are true at this temperature
    I. the reaction is exothermic
    II. gaseous carbon dioxide spontaneously forms
    III. gaseous carbon dioxide is unstable
    a. I only
    b. II only
    c. I and II only
    d. II and III only
    e. I, II, and III

The next two questions refer to the following exothermic reactions involving gases.

I. $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
II. $\frac{1}{2}C + \frac{1}{2}O_2 \rightarrow \frac{1}{2}CO_2$
III. $CH_4 + O_2 \rightarrow C + 2H_2O$
IV. $C + O_2 \rightarrow CO_2$

34. According to the data given above, which reaction liberates the most energy
    a. I only
    b. II only
    c. III only
    d. IV only
    e. This cannot be determined without more data.

35. According to the data given above, which reaction liberates the least energy
    a. I only
    b. II only
    c. III only
    d. IV only
    e. This cannot be determined without more data.

36. Given: $Mg^{2+} + 2e^- \rightarrow Mg$ $E^o = -2.37$ V
    $Fe^{3+} + 1e^- \rightarrow Fe^{2+}$ $E^o = -0.77$ V

    When the reaction $Mg + Fe^{3+} \rightarrow Fe^{2+} + Mg^{2+}$ comes to equilibrium, the E_{cell} value becomes
    a. 3.91 V
    b. 3.14 V
    c. −1.60 V
    d. 1.60 V
    e. 0.00 V

37. Solid sulfur is burned to produce sulfur dioxide gas.

$$S + O_2 \rightarrow SO_2 \qquad \Delta H = -296 \text{ kJ mol}^{-1}$$

The heat released by burning 1.00 kilograms of solid sulfur is captured by 250. L of water. Assuming all of the energy produced is used to heat the water that starts at 25ºC, what will be the final temperature be closest to?

$c_{water} = 4.18 \text{ J g}^{-1} \text{ ºC}^{-1}$

a. 10,000ºC

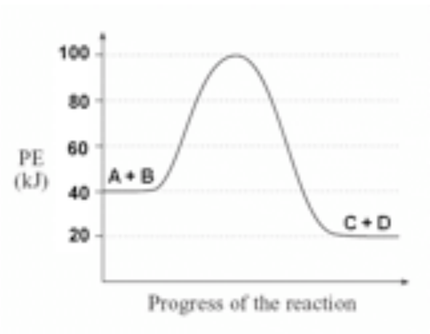
b. 1000ºC

c. 300ºC

d. 10ºC

e. 35ºC

38. Consider the potential energy diagram below



Which of the following thermodynamic quantities can be determined from the diagram

I. ΔG

II. ΔH

III. ΔS

a. I only

b. II only

c. III only

d. I and II only

e. I, II, and III

39. $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$ $\Delta H^o = X$

$H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow H_2O_{(L)}$ $\Delta H^o = Y$

$2C_{(s)} + H_{2(g)} \rightarrow C_2H_{2(g)}$ $\Delta H^o = Z$

Based on the information given above, what is the ΔHº for the following reaction?

$$C_2H_{2(g)} + {}^5/_2O_{2(g)} \rightarrow 2CO_{2(g)} + H_2O_{(L)}$$

a. $X + Y - Z$

b. $2X + Y - Z$

c. $2X + {}^3/_2Y - Z$

d. $-X - Y + Z$

e. $-2X - Y + Z$

40. When pure sodium is placed in an atmosphere of chlorine gas, the following spontaneous reaction occurs.

$$2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$$

Which of the following statements are true about the reaction?

I. $\Delta S > 0$

II. $\Delta H < 0$

III. $\Delta G > 0$

a. I only

b. II only

c. I and II only

d. II and III only

e. I, II, and III only

41. Which of the following are true about the oxidation reduction reaction that takes place in a galvanic cell at standard conditions.

a. G^o and E^o are positive and K is greater than 1

b. G^o is negative, E^o is positive and K is greater than 1

c. G^o is positive, E^o is negative and K is less than 1

d. G^o and E^o are negative and K is greater than 1

e. G^o and E^o are negative and K is less than 1

42. A voltaic cell contains one half−cell with a zinc electrode in a $Zn^{2+}{}_{(aq)}$ solution and a copper electrode in a $Cu^{2+}{}_{(aq)}$ solution. The reaction is shown below. At standard conditions, $E^o = 1.10$ V. If both solutions were reduced to 0.05 M, how would the voltage be affected?

$$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$$

a. the voltage would increase

b. the voltage would decrease

c. the voltage would remain the same

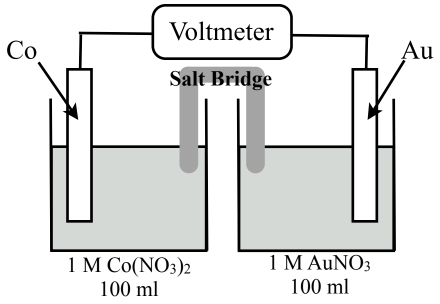
d. without more information, any voltage changes can not be determined

The next 9 questions refer to the following electrochemical cell.

The spontaneous reaction that occurs when the cell below operates at 25ºC is:

$$2Au^{+} + Co_{(s)} \rightarrow 2Au_{(s)} + Co^{2+}$$

(A) Voltage increases
(B) Voltage decreases
(C) Voltage becomes zero and remains at zero.
(D) No change in voltage occurs.
(E) Direction of voltage change cannot be predicted without additional information.



Which of the above choices occurs for each of the following circumstances?

43. The salt bridge is removed.

44. The salt bridge is replaced with a platinum wire.

45. Current is allowed to flow for 20 minutes.

46. Half of the $AuNO_3$ solution is removed and is replaced with the same volume of distilled water.

47. 50.0 g of solid $AuNO_3$ is added to the right beaker. (Assume no volume increase.)

48. The gold electrode is replaced with a platinum electrode

49. The Co electrode is made twice as large.

50. 20 ml of 1 M Na_2CrO_4 is added to the both beakers.

51. 100 ml of distilled water is added to both compartments.

$$4H_{2(g)} + CS_{2(g)} \leftrightarrows CH_{4(g)} + 2H_2S_{(g)}$$

52. A mixture of 2.50 mol $H_{2(g)}$, 1.50 mol $CS_{2(g)}$, 1.5 mol $CH_{4(g)}$ and 2.00 mol $H_2S_{(g)}$ is placed in a 5.0 L rigid reaction vessel, and the system reaches equilibrium according to the equation above. When the equilibrium is achieved, the concentration of $CH_{4(g)}$ has become 0.25 mol L^{-1}.

    Which correctly describes the values for ΔG, the free energy change, and Q, the reaction quotient, when the original mixture was prepared.

    a. $\Delta G = 0, Q = K_{eq}$
    b. $\Delta G > 0, Q < K_{eq}$
    c. $\Delta G > 0, Q > K_{eq}$
    d. $\Delta G < 0, Q > K_{eq}$
    e. $\Delta G < 0, Q < K_{eq}$

1. e Kinetic energy is proportional to temperature. However at any given temperature, the kinetic energy is an *average* value; some molecules will have more kinetic energy, some will have less.

2. d Lattice energy refers to the quantity of energy that must be put in to break apart the + − lattice of an ionic compound into gaseous ions. The lattice energy will be *released* when the opposite process occurs − formation of a lattice from gaseous ions.

3. a Although we haven't studied it yet, you may recall from first year, that ionization energy means just what it sounds like, it is the energy needed to ionize and atom to make ions; in particular to remove an electron and produce a cation.

4. b Free energy is ΔG, which is always negative for spontaneous processes

5. d Heat of vaporization, ΔH_{vap} is given off during condensation, exothermic, and is put in during vaporization, endothermic.

6. e Heat of fusion, ΔH_{fus} is put in during melting, endothermic, and is released during freezing, exothermic.

7. a the definition of *specific* heat capacity, c, the energy required t change 1 g of a substance 1ºC. Remember that heat capacity, C, is the energy required to change some substance (whatever its mass is) 1ºC.

8. d If the equilibrium constant is greater than 1, the reaction has proceeded to products, it is spontaneous, which corresponds to a positive E^o and a negative ΔG^o

9. b Using the Nernst Equation $E = E^o - \frac{0.0592}{2}\log Q$ If the voltage is greater than the standard cell, the Q must be < 1, so that log Q will be negative, which would increase the voltage. Q will be < 1 when the concentration of the products is decreased since $Q = \frac{[Zn^{2+}]}{[Cu^{2+}]}$

10. b Opposite to question #8, if a reaction is not spontaneous, ΔG^o must be positive and K_{eq} will be less than 1.

11. b When a system is isolated, and no work can be done on or by the system (since the container is rigid), the energy will be constant. At the melting point (which for gallium, by the way is just above room temperature), the system is at equilibrium and some of the gallium will melt and an equal amount will freeze, thus the entropy change of these two processes cancel each other out so there is no net change.

12. e According to LeChatelier, endothermic processes are "helped" by higher temperatures. Raising the temperature, induces the system to respond by shifting to the right (more soluble) to "use up" some the the heat added.For dissolving processes there are three energy steps with together result in the overall ΔH of the reaction.
    - The energy to break the crystal structure, this ΔH is always endothermic
    - The energy of separating the water molecules to make room for the ions, this process is alway endothermic
    - The energy of solvation aka hydration; the forming of "bonds" (interactions, really) when the water molecules surround the ions, this process is always exothermic
    - The net result of these three processes determines if the $\Delta H_{dissolve}$ (aka $\Delta H_{solution}$, ΔH_{soln}) is negative (exothermic) or positive (endothermic).

13. e Remember that we are looking for the biggest *change* in entropy. You should be looking for greatest increase in the number moles of gaseous products. Choice (e) increases from no gas molecules to 3 gas molecules, whereas choice (d) forms 4 moles of gas, it comes from 2 moles of gas and no solid, making it less of a change than in (e).

14. e Similar to questions 8 & 10, if K is < 1, then the standard ΔG must be positive and E must be negative, both of which indicate the forward reaction not spontaneous, thus the reverse reaction is spontaneous.

15. a when gas formed from solids occur as in (I), entropy is certainly increased, ΔS +. Precipitation as in (II) would be an unfavorable entropy factor, and 1 mol of gas from 2 mol of gas as in (III) would also be unfavorable entropy.

16. d Since the reaction happens, it is spontaneous and ΔG^o must be negative, and a negative free energy means a larger K_{eq}.

17. b Consider Gibbs equation; $\Delta G = \Delta H - T\,\Delta S$. If the reaction is spontaneous at 298 ΔG must be negative. If the reaction can shift to become non-spontaneous at high temp, increasing temp, making $|T\Delta S|$ larger in magnitude than $|\Delta H|$, eventually the $-(T\Delta S)$ will become more positive than the negative ΔH, and causing the reaction to become non-spontaneous.

18. c This question is very similar to #9. Realize that voltage will change when the concentration of the ions change in the electrochemical cell. The size of the metal electrodes does not affect voltage, the fact that the ions are put into solution as nitrates or sulfates, it only matters that they are soluble salts. The volume of the solutions in the compartments does not affect voltage, and of course the salt bridge should have potassium and chloride ions. Thus the decreased concentration of lead ions compared to the zinc ions will reduce the voltage.

19. b Remember that freezing is an exothermic process, and certainly the entropy is decreasing, and since ice is less dense, water expands as it freezes, the volume increases, thus $\Delta V > 0$

20. a This question is very similar to #12 We are told in the problem that the dissolving of calcium iodide is exothermic (negative ΔH), thus we know that the single exothermic piece of the three parts of the dissolving puzzle must be greater in magnitude, though opposite in sign that even the other two part combined. Heating any exothermic reaction will cause a shift back towards reactants, which is the solid, which is to say it may cause dissolved ions to precipitate out of solution not help the salt to dissolve.

21. b Three of the five reactions form solids, less entropy, one forms a liquid and only b produces 3 moles of gas from 2 moles of gas which is a higher entropy situation. I can see how you might have been tempted to choose "d", however if you look closely, 3 gas molecules turn to 3 gas molecules and 1 solid molecule turn to 2 solid molecules, that is likely to have a smaller entropy change, since no improvement in the number of molecules

22. d This is a Hess' Law problem. To "build" the reaction:

Rx 1: reverse and multiply by ½	$\Delta H = +3200$ kJ
Rx 2: multiply by 6	$\Delta G^o = -2400$ kJ
Rx 3: multiply by 3	$\Delta G^o = -750$ kJ

For a net result of +50 kJ

23. a Since we know that the reaction is non-spontaneous below 550ºC, we know that ΔG is positive, and since high temperatures will make $T\Delta S$ larger, we know that ΔS must be a positive quantity. Thus in the equation $\Delta G = \Delta H - T\Delta S$ when $T\Delta S$ gets larger, the subtraction of such term will overshadow the ΔH quantity. We know that ΔH must be positive which is what makes the reaction non-spontaneous at lower temperatures, when ΔH would be larger than $T\Delta S$ promoting non-spontaneity.

24. e According to the Nernst equation, $E = E^o - \frac{0.0592}{2}\log\frac{[Zn^{2+}]}{[Cu^{2+}]}$ the only way to have a value greater than 1.10 V is if $\frac{0.0592}{2}\log\frac{[Zn^{2+}]}{[Cu^{2+}]}$ produces a negative value. This can only happen when the $[Zn^{2+}]$ is less than the $[Cu^{2+}]$. If $[Zn^{2+}] < [Cu^{2+}]$ then Q will be smaller than 1 and the log of any number smaller than 1 will be a negative value, and of course subtracting a negative will add to the E^o

25. a If the wax is melting, bonds are being broken, endothermic, thus $\Delta H = +$, and since liquid wax will have more entropy, $\Delta S = +$, and lastly, liquid wax is less dense than solid wax you know this because solid wax sinks in liquid wax indicating larger volume for the liquid, $\Delta V = +$.

26. c Remember that bond forming (products) is exothermic and bond breaking (reactants) is endothermic, $(+BDE_{reactants}) + (-BDE_{products})$. Thus: $[(2\times436) + 499] + [-(4\times464)] = -485$ kJ

27. a Use Hess' Law

$$\begin{array}{ll}
CO_{(g)} + 2H_{2(g)} \rightarrow CH_3OH_{(g)} & \Delta H = -91 \text{ kJ} \\
\underline{CH_3OH_{(g)} \rightarrow CH_3OH_{(L)} \qquad\quad \Delta H = -37 \text{ kJ}} & (g \rightarrow L, \text{ condensation is } -\text{Heat of vaporization}) \\
CO_{(g)} + 2H_{2(g)} \rightarrow CH_3OH_{(L)} & \Delta H = -128 \text{ kJ}
\end{array}$$

28. c Use the equation that relates E° with K $E^\circ = \dfrac{0.0592 \log K}{n}$ at first this calculation seems challenging without a calculator, however, putting in the values and making a few estimations: $\log K = \dfrac{E^\circ n}{0.0592}$ After substitution: $\log K = \dfrac{(0.12)(2)}{0.06}$ this reduces to $\log K = 4$ and the log of $10^4 = 4$, thus the K = ~10^4

29. d A process will be spontaneous at all temperatures when both driving forces, enthalpy and entropy are both favorable, thus ΔH is − and ΔS is +

30. e Spontaneity can change with varying temperature if one of the two driving forces is unfavorable and the other is favorable, thus when ΔH and ΔS have the same sign, as with choices I and III.

31. a When temperatures increase, the $|T\Delta S|$ term becomes greater, and when ΔS is positive, the $-T\Delta S$ term can overshadow the + ΔH term, flipping the ΔG from + to −.

32. b The definition of standard enthalpy of formation is the energy involved in forming *one* mole of a compound from its elements in their standard (or free) state at 298 K and 1 atm. This of course only applies to (b).

33. c A system is only at equilibrium when $\Delta G = 0$. The negative ΔG tell us that the reaction will occur spontaneously. Remember that spontaneously does not mean "without ignition energy" because indeed this burning of coal would require an ignition energy, perhaps the high temperature of the system is enough to provide that, but whether or not it does, when ΔG is negative, we can still say the product will form spontaneously. Lastly, a negative ΔG tells us that the products are more stable than the reactants. We know the reaction is exothermic because we can see this is a combustion reaction, and all combustion reactions are exothermic.

34. a *This is a tricky question. Use Hess' Law to help determine some information about the size of the enthalpy of reactions.* Since II is exactly half of IV, II must be less endothermic than IV. Also, if you add III to IV, you get I, thus I must be greater than III and IV, and must be the most exothermic of these four processes (since each of the parts must be less than the whole).

35. e From this data, and the discussion in #34, we know that II is less than IV, but there is no way to determine how the energy of II compared to III.

36. e When a voltaic cell reaches equilibrium, E^o always = 0

37. e On first inspection, this calculation looks not to be such easy math. Take a closer look. We should use the equation $(\Delta H)\left(\dfrac{mass_{sulfur}}{MolarMass}\right) = (c)(mass_{water})(\Delta T)$ solve for ΔT $\dfrac{(mass_{sulfur})(\Delta H)}{(c)(mass_{water})(MolarMass)} = (\Delta T)$ now substituting the values $\dfrac{(1000g)(296{,}000J)}{\left(\dfrac{4.18J}{g^\circ C}\right)(250{,}000g)(32)} = (\Delta T)$ First, it is worth eliminating the three zeros off the two factors 296,000 and 250,000. Then consider that 250 into 1000 is ~4, divided by 4.18 is ~1. Next consider that 296 is approximately 300 divided by 32 which is approximately 30, and thus results in a temp change of about 10°. This of course would be a 10° increase, which would give a final temp change of 35°C.

38. b Potential energy diagrams give us information about ΔH, not ΔG or ΔS. Potential energy diagrams also tell us about activation energy of both the forward and reverse reactions.

39. b To apply Hess' Law correctly,

Rx X: multiply by 2 $2C_{(g)} + 2O_{2(g)} \rightarrow 2CO_{2(g)}$ $\Delta H = X$

Rx Y: leave the same $H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow H_2O_{(L)}$ $\Delta H = Y$

Rx 3: reverse $\underline{C_2H_{2(g)} \rightarrow 2C_{(s)} + H_{2(g)}}$ $\Delta H = -Z$

$$C_2H_{2(g)} + {}^{5}/{}_{2}O_{2(g)} \rightarrow 2CO_{2(g)} + H_2O_{(L)}$$

For a net result of $\Delta H = 2X + Y - Z$

40. b We are told the reaction is spontaneous, so we know III is false, and since a solid is formed, we know that ΔS must be negative, unfavorable thus I is false. Thus in order for the reaction to be spontaneous, then the ΔH factor must be favorable, negative, thus II is true.

41. b If the reaction is taking place as the problem says, E^o must be positive, G^o must be negative, and K would be greater than one.

42. c Since the stoichiometry is all ones, and the reaction quotient, Q does not change, it would still equal 1, the log of 1 is zero, and the voltage remains the same. Nernst Equation: $E_{cell} = E^o_{cell} - \frac{0.0592}{2}\log\frac{[\downarrow Zn^{2+}]}{[\downarrow Cu^{2+}]}$

43. c Removing the salt bridge will stop the reaction in its tracks since the charge build up will immediately put a stop to any flow of electrons.

44. c The salt bridge is necessary to maintain a neutral charge in the solution as the reaction proceeds and cobalt ions are formed in one compartment and gold ions are lost in the other compartment. The salt bridge must contain free ions (anions and cations) that can migrate into each compartment as necessary, a piece of metal cannot push ions into each compartment to maintain a balanced charge.

45. b As a cell runs, the reaction proceeds in the forward direction (as written in the equation shown) the concentration of the products increase and the concentration of the reactants decrease causing Q > 1, and log Q will be +, thus reducing the voltage of E_{cell}.

46. b This of course decreases the concentration of gold ions, in the denominator of Q, making Q > 1 and the log Q will be +, thus reducing the voltage.

47. a This of course increases the concentration of the gold ions, causing Q < 1, and the log Q will be −, resulting in an increase in voltage.

48. d Replacing the gold electrode will cause very little change in the voltage because the gold electrode is an inert electrode. What is meant by an inert electrode is that the gold atoms do not participate. The gold ions turn into gold atoms and in doing so, the newly formed gold atoms gets deposited on the gold electrode. But this gold electrode could be made of any conducting material as long as that material is less active than the cobalt.

49. d Changing the size of the electrode will never affect the voltage. It might change the reaction rate due to an increase of surface area, but it will not change the voltage. You can see that there is no place for solid in the Nernst equation.

50. b Sodium chromate is soluble, thus adding chromate ions will cause precipitates to form in both compartments. These precipitates will serve to reduce both the cobalt and gold ions, however, because of the stoichiometry of the precipitation reactions,

$$Co^{2+} + CrO_4^{2-} \rightarrow CoCrO_4 \qquad 2Au^+ + CrO_4^{2-} \rightarrow Au_2CrO_4$$

0.02 moles of chromate will precipitate twice as many moles of gold ions out of solution causing the gold ion concentration to decrease more than the cobalt, as well as the fact that the decrease is accentuated by the squaring effect. Thus after the precipitation removes some ions from solution, Q > 1, and log Q will be +, decreasing the voltage.

51. b *This is a tough problem* – Consider the Nernst equation shown below: $E_{cell} = E^o_{cell} - \frac{0.0592}{2}\log\frac{[\downarrow Co^{2+}]}{[\downarrow Au^{+}]\downarrow^{2}}$

Any concentration changes that impact the ions in solution will impact will impact the "Q factor," aka: $\log[Co^{2+}]/[Au^+]^2$ Concentration changes that affect both ions equally have greater impact on the $[Au^+]^2$ more, because of the square. You might think that diluting both solutions would act equally, however, lowering both concentrations equally from 1 to some concentration less than 1 will make the Q > 1, causing log Q to be + which will result in voltage subtracted from E^o_{cell} reducing the voltage of E_{cell}.

52. c Since the reaction shifts to the left, the forward reaction is not spontaneous, thus + ΔG, and the shift left indicates that Q must be larger than K