

# Practice Test C.1 (Ch 5, 19, 20.5–6) Thermo & Electro

Name \_\_\_\_\_ Per \_\_\_\_\_

- *This is practice - Do NOT cheat yourself of finding out what you are capable of doing. Be sure you follow the testing conditions outlined below.*
- *DO NOT USE A CALCULATOR. You may use ONLY your blue periodic table.*
- *Try only for 1.2 minutes per question (50 questions = 60 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

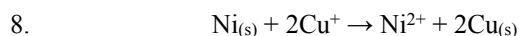
- (A) ionization energy
- (B) free energy
- (C) activation energy
- (D) lattice energy
- (E) kinetic energy

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.

Questions 4 – 7 refer to the following

- (A) 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 substance is taken in when a substance melts.
7. The energy released when cooling 1 g of a substance 1°C.



If the equilibrium constant for the reaction above is greater than 1, which of the following correctly describes the standard voltage,  $E^\circ$ , and the standard free energy change,  $\Delta G^\circ$ , for this reaction?

- a.  $E^\circ$  and  $\Delta G^\circ$  are both positive
- b.  $E^\circ$  and  $\Delta G^\circ$  are both negative
- c.  $E^\circ$  and  $\Delta G^\circ$  are both zero
- d.  $E^\circ$  is positive and  $\Delta G^\circ$  is negative
- e.  $E^\circ$  is negative and  $\Delta G^\circ$  is positive



An electrolytic 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,  $E^\circ$ , of 1.10 volts. Which of the following could correctly account for this observation?

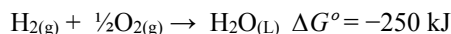
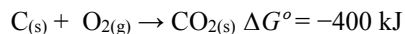
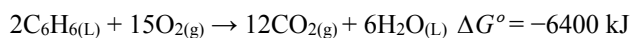
- a. The zinc electrode contained more mass than the copper electrode.
  - b. The  $\text{Zn}^{2+}$  electrolyte was 0.5 M  $\text{Zn}(\text{NO}_3)_2$ , whereas the  $\text{Cu}^{2+}$  electrolyte was 1.0 M  $\text{Cu}(\text{NO}_3)_2$ .
  - c. The  $\text{Zn}^{2+}$  solution was colorless, whereas the  $\text{Cu}^{2+}$  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 spontaneously from initial standard state conditions?
    - a.  $\Delta G^\circ$  is positive and  $K_{eq}$  is greater than 1
    - b.  $\Delta G^\circ$  is positive and  $K_{eq}$  is less than 1
    - c.  $\Delta G^\circ$  is negative and  $K_{eq}$  is greater than 1
    - d.  $\Delta G^\circ$  is negative and  $K_{eq}$  is less than 1
    - e.  $\Delta G^\circ$  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?
- The hydration energies are very high.
  - Ammonium nitrate produces an ideal solution in water.
  - The heat of solution for ammonium nitrate is exothermic.
  - Ammonium nitrate has a low lattice energy.
  - Ammonium nitrate is more soluble in hot water.
13. Choose the reaction expected to have the greatest increase in entropy.
- $\text{H}_2\text{O}_{(g)} \rightarrow \text{H}_2\text{O}_{(l)}$
  - $\text{C}_{(s)} + \text{O}_{2(g)} \rightarrow \text{CO}_{2(g)}$
  - $\text{Ca}_{(s)} + \text{H}_{2(g)} \rightarrow \text{CaH}_{2(s)}$
  - $2\text{NH}_{3(g)} \rightarrow 3\text{H}_{2(g)} + \text{N}_{2(g)}$
  - $2\text{KClO}_{3(s)} \rightarrow 2\text{KCl}_{(s)} + 3\text{O}_{2(g)}$
14.  $2\text{Fe}^{2+} + \text{Cu}^{2+} \rightarrow 2\text{Fe}^{3+} + \text{Cu}_{(s)}$
- If the equilibrium constant for the reaction above is  $2.5 \times 10^{-4}$ , which of the following correctly describes the standard voltage,  $E^\circ$ , and the standard free energy change,  $\Delta G^\circ$ , for this reaction?
- $E^\circ$  and  $\Delta G^\circ$  are both zero
  - $E^\circ$  and  $\Delta G^\circ$  are both positive
  - $E^\circ$  and  $\Delta G^\circ$  are both negative
  - $E^\circ$  is positive and  $\Delta G^\circ$  is negative
  - $E^\circ$  is negative and  $\Delta G^\circ$  is positive
15. For which of the following processes would  $\Delta S$  have a positive value?
- $\text{MgCO}_{3(s)} \rightarrow \text{MgO}_{(s)} + \text{CO}_{2(g)}$
  - $\text{Ba}^{2+} + \text{SO}_4^{2-} \rightarrow \text{BaSO}_{4(s)}$
  - $\text{Cl}_{2(g)} + \text{C}_3\text{H}_{6(g)} \rightarrow \text{C}_3\text{H}_6\text{Cl}_{2(g)}$
- I only
  - I and II only
  - I and III only
  - II and III only
  - I, II, and III
16. Under standard conditions calcium metal reacts readily with chlorine gas. What conclusions may be drawn from this fact?
- $K_{eq} < 1$ ,  $\Delta G^\circ > 0$
  - $K_{eq} > 1$  and  $\Delta G^\circ = 0$
  - $K_{eq} < 1$  and  $\Delta G^\circ < 0$
  - $K_{eq} > 1$  and  $\Delta G^\circ < 0$
  - $K_{eq} > 1$  and  $\Delta G^\circ > 0$
17.  $\text{SF}_{4(g)} + \text{F}_{2(g)} \rightarrow \text{SF}_{6(g)}$
- The reaction above is thermodynamically spontaneous at 298, but becomes non spontaneous at higher temperatures. Which of the following is true at 298 K?
- $\Delta G$ ,  $\Delta H$ , and  $\Delta S$  are all positive
  - $\Delta G$ ,  $\Delta H$ , and  $\Delta S$  are all negative
  - $\Delta G$  and  $\Delta H$  are negative, but  $\Delta S$  is positive
  - $\Delta G$  and  $\Delta S$  are negative, but  $\Delta H$  is positive
  - $\Delta G$  and  $\Delta H$  are positive, but  $\Delta S$  is negative
18.  $\text{Zn} + \text{Pb}^{2+} \rightarrow \text{Zn}^{2+} + \text{Pb}$
- An electrolytic 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,  $E^\circ$ , of 0.63 volts. Which of the following could correctly account for this observation?
- The lead electrode contained more mass than the zinc electrode.
  - The  $\text{Zn}^{2+}$  electrolyte was  $\text{ZnSO}_4$ , whereas the  $\text{Pb}^{2+}$  electrolyte was 1.0 M  $\text{Pb}(\text{NO}_3)_2$
  - The  $\text{Pb}^{2+}$  solution was less concentrated than the  $\text{Zn}^{2+}$ .
  - The solutions in the half-cells had different volumes
  - The salt bridge contained  $\text{KNO}_3$  as the electrolyte.
19.  $\text{H}_2\text{O}_{(l)} \rightarrow \text{H}_2\text{O}_{(s)}$
- When water freezes at its normal freezing point,  $0^\circ\text{C}$  and 1 atmosphere, which of the following ( $V = \text{volume}$ )
- $\Delta H < 0$ ,  $\Delta S > 0$ , and  $\Delta V > 0$
  - $\Delta H < 0$ ,  $\Delta S < 0$ , and  $\Delta V > 0$
  - $\Delta H > 0$ ,  $\Delta S < 0$ , and  $\Delta V < 0$
  - $\Delta H > 0$ ,  $\Delta S > 0$ , and  $\Delta V > 0$
  - $\Delta H > 0$ ,  $\Delta S > 0$ , and  $\Delta V < 0$
20. Calcium iodide,  $\text{CaI}_2$ , has a  $\Delta H^\circ_{\text{soln}}$  of  $-104$  kJ per mole. Which of the following is probably true when calcium iodide is dissolved in water?
- The energy to break the crystal structure must be less than the energy of solvating the ions.
  - As the solution dissolves, the temperature of the solution decreases.
  - The resulting solution has a high vapor pressure
  - Heating the solution causes more salt to dissolve.
  - The  $\Delta G$  for the dissolving process is  $> 0$

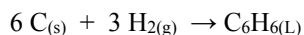
21. Which of the following reactions has the largest positive value of  $\Delta S$ ?

- $2 \text{H}_2\text{S}_{(g)} + \text{SO}_{2(g)} \rightarrow 3 \text{S}_{(s)} + 2 \text{H}_2\text{O}_{(g)}$
- $2 \text{SO}_{3(g)} \rightarrow 2 \text{SO}_{2(g)} + \text{O}_{2(g)}$
- $\text{Mg}_{(s)} + \text{Cl}_{2(g)} \rightarrow \text{MgCl}_{2(s)}$
- $\text{Fe}_2\text{O}_{3(s)} + 3 \text{H}_{2(g)} \rightarrow 2 \text{Fe}_{(s)} + 3 \text{H}_2\text{O}_{(g)}$
- $\text{H}_{2(g)} + \frac{1}{2}\text{O}_{2(g)} \rightarrow \text{H}_2\text{O}_{(l)}$

22. Consider the reactions below to answer the question.



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



- 250 kJ
- 100.0 kJ
- 50.0 kJ
- 50.0 kJ
- 100.0 kJ

23.  $\text{N}_{2(g)} + \text{O}_{2(g)} \rightarrow \text{N}_2\text{O}_{(g)}$

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

	$\Delta G$	$\Delta H$	$\Delta S$
(A)	+	+	+
(B)	-	-	-
(C)	-	-	+
(D)	-	+	-
(E)	+	+	-

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

- 1.0 M  $\text{Zn}^{2+}_{(aq)}$ , 1.0 M  $\text{Cu}^{2+}_{(aq)}$
- 5.0 M  $\text{Zn}^{2+}_{(aq)}$ , 5.0 M  $\text{Cu}^{2+}_{(aq)}$
- 5.0 M  $\text{Zn}^{2+}_{(aq)}$ , 1.0 M  $\text{Cu}^{2+}_{(aq)}$
- 0.5 M  $\text{Zn}^{2+}_{(aq)}$ , 0.5 M  $\text{Cu}^{2+}_{(aq)}$
- 0.1 M  $\text{Zn}^{2+}_{(aq)}$ , 1.0 M  $\text{Cu}^{2+}_{(aq)}$

25. Which of the following combinations is true when wax melts? ( $V = \text{volume}$ )

- $\Delta H > 0$ ,  $\Delta S > 0$ , and  $\Delta V > 0$
- $\Delta H > 0$ ,  $\Delta S > 0$ , and  $\Delta V < 0$
- $\Delta H > 0$ ,  $\Delta S < 0$ , and  $\Delta V > 0$
- $\Delta H < 0$ ,  $\Delta S < 0$ , and  $\Delta V < 0$
- $\Delta H < 0$ ,  $\Delta S > 0$ , and  $\Delta V > 0$

26.  $2\text{H}_{2(g)} + \text{O}_{2(g)} \rightarrow 2\text{H}_2\text{O}_{(g)}$

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

bond	average bond energy (kJ/mol <sup>-1</sup> )
H-H	436
O=O	499
H-O	464

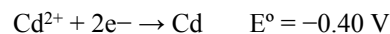
- 464 kJ
- 485 kJ
- 485 kJ
- 443 kJ
- 443 kJ

27.  $\text{CO}_{(g)} + 2\text{H}_{2(g)} \rightarrow \text{CH}_3\text{OH}_{(g)} \quad \Delta H = -91 \text{ kJ}$

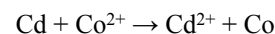
Determine the  $\Delta H$  for the reaction above if  $\text{CH}_3\text{OH}_{(l)}$  were formed in the above reaction instead of  $\text{CH}_3\text{OH}_{(g)}$ . The  $\Delta H$  of vaporization for  $\text{CH}_3\text{OH}$  is  $37 \text{ kJ mol}^{-1}$ .

- 128 kJ
- 58 kJ
- +128 kJ
- +54 kJ
- 37 kJ

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



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



- $10^{-4}$
- $10^{-2}$
- $10^4$
- $10^{16}$
- $10^2$

For questions 29 – 30 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. Which of the four processes 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.  $\text{N}_{(g)} + \text{O}_{2(g)} \rightarrow \text{NO}_{2(g)}$
- b.  $\frac{1}{2}\text{N}_{2(g)} + \text{O}_{2(g)} \rightarrow \text{NO}_{2(g)}$
- c.  $\frac{1}{2}\text{N}_2\text{O}_{4(g)} \rightarrow 2 \text{NO}_{2(g)}$
- d.  $\text{N}_{2(g)} + 2\text{O}_{2(g)} \rightarrow \text{N}_2\text{O}_{4(g)}$
- e.  $\text{N}_{2(g)} + 2\text{O}_{2(g)} \rightarrow 2\text{NO}_{2(g)}$

33. At a certain temperature  $\text{C}_{(s)} + \text{O}_{2(g)} \rightarrow \text{CO}_{2(g)}$  has a  $\Delta G$  of  $-339.4 \text{ kJ mol}^{-1}$ . Which statements from the following list are true at this temperature

- I. the system is exothermic
  - II. gaseous carbon dioxide spontaneous forms
  - III. gaseous carbon dioxide is unstable
- a. I only
  - b. II only
  - c. I and II only
  - d. II and II only
  - e. I, II, and III

For questions 34 – 35 refer to the following exothermic reactions involving gases.

- I.  $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$
- II.  $\frac{1}{2}\text{C} + \frac{1}{2}\text{O}_2 \rightarrow \frac{1}{2}\text{CO}_2$
- III.  $\text{CH}_4 + \text{O}_2 \rightarrow \text{C} + 2\text{H}_2\text{O}$
- IV.  $\text{C} + \text{O}_2 \rightarrow \text{CO}_2$

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:  $\text{Mg}^{2+} + 2\text{e}^- \rightarrow \text{Mg} \quad E^\circ = -2.37 \text{ V}$   
 $\text{Fe}^{3+} + 1\text{e}^- \rightarrow \text{Fe}^{2+} \quad E^\circ = -0.77 \text{ V}$

When the reaction  $\text{Mg} + \text{Fe}^{3+} \rightarrow \text{Fe}^{2+} + \text{Mg}^{2+}$  comes to equilibrium, the  $E_{\text{cell}}$  value becomes

- a. 3.91 V
- b. 3.14 V
- c.  $-1.60 \text{ V}$
- d. 1.60 V
- e. 0.00 V

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

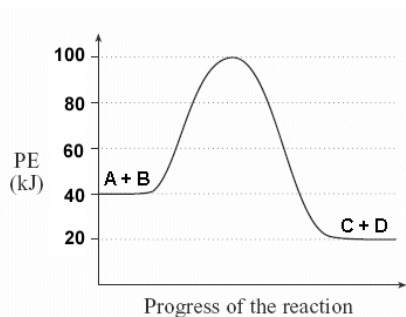


The heat released by burning 1.00 kilograms of solid sulfur is captured by 250. L of water. Assuming all of the energy produced is used to heat the water that starts at 25°C, what will be the final temperature be closest to?

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

- 10,000°C
- 1000°C
- 300°C
- 10°C
- 35°C

38. Consider the potential energy diagram below

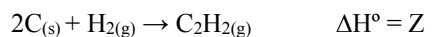
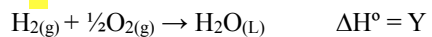


Which of the following thermodynamic quantities can be determined from the diagram

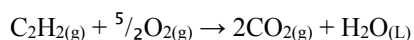
- $\Delta G$
- $\Delta H$
- $\Delta S$

- I only
- II only
- III only
- I and II only
- I, II, and III

39.  $\text{C}_{(\text{s})} + \text{O}_{2(\text{g})} \rightarrow \text{CO}_{2(\text{g})} \quad \Delta H^\circ = X$

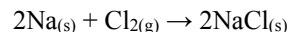


Based on the information given above, what is the  $\Delta H^\circ$  for the following reaction?



- $X + Y - Z$
- $2X + Y - Z$
- $2X + \frac{3}{2}Y - Z$
- $-X - Y + Z$
- $-2X - Y + Z$

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



Which of the following statements are true about the reaction?

- $\Delta S > 0$
- $\Delta H < 0$
- $\Delta G > 0$

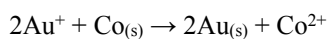
- I only
- II only
- I and II only
- II and III only
- I, II, and III only

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

- $G^\circ$  and  $E^\circ$  are positive and  $K$  is greater than 1
- $G^\circ$  is negative,  $E^\circ$  is positive and  $K$  is greater than 1
- $G^\circ$  is positive,  $E^\circ$  is negative and  $K$  is less than 1
- $G^\circ$  and  $E^\circ$  are negative and  $K$  is greater than 1
- $G^\circ$  and  $E^\circ$  are negative and  $K$  is less than 1

Questions # – # refer to the following electrochemical cell.

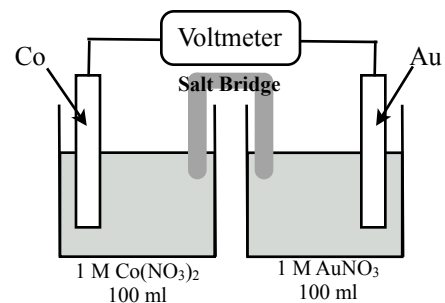
The spontaneous reaction that occurs when the cell below operates at 25°C is:



- (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 occurs for each of the following circumstances?

- 42. The salt bridge is removed.
- 43. The salt bridge is replaced with a platinum wire.
- 44. Current is allowed to flow for 20 minutes.
- 45. Half of the  $\text{AuNO}_3$  solution is removed and is replaced with the same volume of distilled water.
- 46. 50.0 g of solid  $\text{Au}(\text{NO}_3)_2$  is added to the right beaker. (Assume no volume increase.)
- 47. The gold electrode is replaced with a platinum electrode
- 48. 20 ml of 1 M  $\text{Na}_2\text{CrO}_4$  is added to the both beakers.
- 49. The Co electrode is made twice as large.
- 50. 100 ml of distilled water is added to both compartments.



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 is the means just what it sounds like, energy needed to make ions; in particular to remove an electron and produce a cation.
4. b Free energy is  $\Delta G$ , which is always negative for spontaneous processes
5. d Heat of vaporization,  $\Delta H_{\text{vap}}$  is given off during condensation, exothermic, and is put in during vaporization, endothermic.
6. e Heat of fusion,  $\Delta H_{\text{fus}}$  is put in during melting, endothermic, and is released during freezing, exothermic.
7. a the definition of *specific* heat capacity,  $c$ , the energy required to change 1 g of a substance 1°C. Remember that heat capacity,  $C$ , is the energy required to change some substance (whatever its mass is) 1°C.
8. d If the equilibrium constant is greater than 1, the reaction has proceeded to products, it is spontaneous, which corresponds to a positive  $E^\circ$  and a negative  $\Delta G^\circ$
9. b Using the Nernst Equation  $E = E^\circ - \frac{0.0592}{2} \log Q$  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  $Q = \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$
10. c Similar to question #8, if a reaction proceeds spontaneously,  $\Delta G$  must be negative and  $K_{\text{eq}}$  will be greater 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) 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. Further, for dissolving processes there are three energy steps with together result in the overall  $\Delta H$  of the reaction.
  - The energy to break the crystal structure, this  $\Delta H$  is always endothermic
  - The energy of separating the water molecules to make room for the ions, this process is always endothermic
  - The energy of solvation; 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  $\Delta H_{\text{dissolve}}$  (aka  $\Delta H_{\text{solution}}$ ,  $\Delta H_{\text{soln}}$ ) 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  $\Delta G$  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,  $\Delta S +$ . 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  $\Delta G^\circ$  must be negative, and a negative free energy means a larger  $K_{\text{eq}}$ .
17. b Consider Gibbs equation;  $\Delta G = \Delta H - T \Delta S$ . If the reaction is spontaneous at 298  $\Delta G$  must be negative. If the reaction can shift to become nonspontaneous 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  $\Delta H$ , and causing the reaction to become nonspontaneous.
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  $\Delta H$ ), 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
22. d This is a Hess' Law problem. To "build" the reaction:  
 Rx 1: reverse and multiply by  $\frac{1}{2}$   $\Delta H = +3200$  kJ  
 Rx 2: multiply by 6  $\Delta G^\circ = -2400$  kJ  
 Rx 3: multiply by 3  $\Delta G^\circ = -750$  kJ  
 For a net result of +50 kJ
23. a Since we know that the reaction is nonspontaneous below 550°C, we know that  $\Delta G$  is positive, and since high temperatures will make  $T\Delta S$  larger, we know that  $\Delta S$  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  $\Delta H$  quantity. We know that  $\Delta H$  must be positive which is what makes the reaction nonspontaneous at lower temperatures, when  $|\Delta H|$  would be larger than  $T\Delta S$  promoting spontaneity.
24. e According to the Nernst equation,  $E = E^\circ - \frac{0.0592}{2} \log \frac{[Zn^{2+}]}{[Cu^{2+}]}$  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  $E^\circ$
25. a If the wax is melting, then it is spontaneous,  $\Delta G = -$ , 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  $BDE_{\text{reactants}} - BDE_{\text{products}}$  (since bond breaking is endothermic and bond forming is exothermic)  $[2(436) + 499] - 4(464)$
27. a Use Hess' Law  $\begin{array}{l} CO_{(g)} + 2H_{2(g)} \rightarrow CH_3OH_{(g)} \quad \Delta H = -91 \text{ kJ} \\ CH_3OH_{(g)} \rightarrow CH_3OH_{(L)} \quad \Delta H = -37 \text{ kJ} \\ \hline CO_{(g)} + 2H_{2(g)} \rightarrow CH_3OH_{(L)} \quad \Delta H = -128 \text{ kJ} \end{array}$  (g  $\rightarrow$  L, condensation is  $-\text{Heat of vaporization}$ )
28. c Use the equation that relates  $E^\circ$  with K  $E^\circ = \frac{0.0592 \log K}{n}$  at first this calculation seems challenging without a calculator, however, putting in the values and making a few estimations:  $\log K = \frac{E^\circ n}{0.0592}$  After substitution:  
 $\log K = \frac{(0.12)(2)}{0.06}$  this reduces to  $\log K = 4$  and the log of  $10^4 = 4$ , thus the  $K = \sim 10^4$
29. d A process will be spontaneous at all temperatures when both driving forces, enthalpy and entropy are both favorable, thus  $\Delta H$  is  $-$  and  $\Delta S$  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  $\Delta H$  and  $\Delta S$  have the same sign, as with choices I and III.
31. a When temperatures increase, the  $|T\Delta S|$  term becomes greater, and when  $\Delta S$  is positive, the  $-T\Delta S$  term can overshadow the  $+\Delta H$  term, flipping the  $\Delta G$  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. b A system is only at equilibrium when  $\Delta G = 0$ . The negative  $\Delta G$  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  $\Delta G$  is negative, we can still say the product will form spontaneously. Lastly, a negative  $\Delta G$  tells us that the products are more stable than the reactants.
34. a This is a tricky question. 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 the most exothermic of these four processes (since each of the parts must be less than the whole).
35. e From this data, there is no way to determine whether it's II or III that is less energetic than the other.
36. e When a voltaic cell reaches equilibrium,  $E^\circ$  always = 0

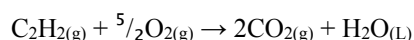
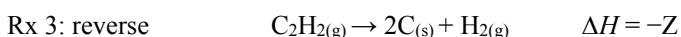
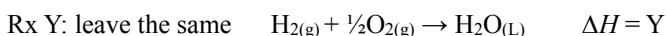
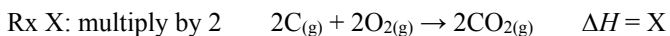
37. e On first inspection, this calculation looks not to be such easy math. Take a closer look. We should use the equation
- $$(\Delta H) \left( \frac{\text{mass}_{\text{sulfur}}}{\text{MolarMass}} \right) = (c)(\text{mass}_{\text{water}})(\Delta T) \quad \text{solve for } \Delta T \quad \frac{(\text{mass}_{\text{sulfur}})(\Delta H)}{(c)(\text{mass}_{\text{water}})(\text{MolarMass})} = (\Delta T) \quad \text{now substituting the}$$

values  $\frac{(1000\text{g})(296,000\text{J})}{\left(\frac{4.18\text{J}}{\text{g}^\circ\text{C}}\right)(250,000\text{g})(32)} = (\Delta T)$  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.

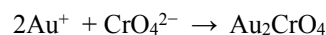
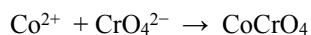
38. b Potential energy diagrams give us information about  $\Delta H$  only.

39. b To apply Hess' Law correctly,



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

40. c Remember  $\Delta G = \Delta H - T \Delta S$  and since solid is formed,  $\Delta S$  must be negative, unfavorable, in order for the reaction to be spontaneous, as it is, then the  $\Delta H$  factor must be favorable, negative.
41. b If the reaction is taking place as the problem says,  $E^\circ$  must be positive,  $G^\circ$  must be negative, and  $K$  would be greater than one.
42. 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.
43. 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.
44. b As a cell runs, the reaction proceeds in the forward direction (as written in the equation above) 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  $E_{\text{cell}}$ .
45. 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.
46. 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.
47. d Replacing the gold electrode will cause no change in the voltage because the gold electrode is an inert electrode.
48. 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,



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.

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 *This is a tough problem* – Consider the Nernst equation shown below:

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.0592}{2} \log \frac{[\text{Co}^{2+}]}{[\text{Au}^+]^2}$$

Any concentration changes that impact the ions in solution will impact will impact the “Q factor,” aka:  $\log[\text{Co}^{2+}]/[\text{Au}^+]^2$  Concentration changes that affect both ions equally have greater impact on the  $[\text{Au}^+]^2$  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  $E_{\text{cell}}^\circ$  reducing the voltage of  $E_{\text{cell}}$ .