

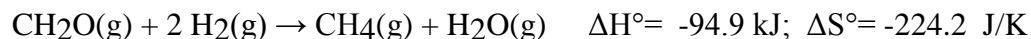
NAME \_\_\_\_\_

S2014/CHEM1451/Dooley/Exam 4

Multiple Choice: (3Pts each)

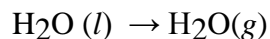
- \_\_\_\_\_ 1. Which of the following processes have a  $\Delta S > 0$ ?
- $\text{CH}_3\text{OH}(\text{l}) \rightarrow \text{CH}_3\text{OH}(\text{s})$
  - $\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightarrow 2 \text{NH}_3(\text{g})$
  - $\text{CH}_4(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightarrow \text{CO}(\text{g}) + 3 \text{H}_2(\text{g})$
  - $\text{Na}_2\text{CO}_3(\text{s}) + \text{H}_2\text{O}(\text{g}) + \text{CO}_2(\text{g}) \rightarrow 2 \text{NaHCO}_3(\text{s})$
  - $2 \text{NH}_3(\text{g}) + \text{CO}_2(\text{g}) \rightarrow \text{NH}_2\text{CONH}_2(\text{aq}) + \text{H}_2\text{O}(\text{l})$
- \_\_\_\_\_ 2. Consider the following reaction at constant P. Use the information here to determine the value of  $\Delta S_{\text{surr}}$  at 298 K.
- $$\text{N}_2(\text{g}) + 2 \text{O}_2(\text{g}) \rightarrow 2 \text{NO}_2(\text{g}) \Delta H = +66.4 \text{ kJ}$$
- $\Delta S_{\text{surr}} = +223 \text{ J/K}$
  - $\Delta S_{\text{surr}} = -223 \text{ J/K}$
  - $\Delta S_{\text{surr}} = -66.4 \text{ J/K}$
  - $\Delta S_{\text{surr}} = +66.4 \text{ kJ/K}$
  - $\Delta S_{\text{surr}} = -66.4 \text{ J/K}$
- \_\_\_\_\_ 3. Consider a reaction that has a positive  $\Delta H$  and a positive  $\Delta S$ . Which of the following statements is TRUE?
- This reaction will be spontaneous only at high temperatures.
  - This reaction will be spontaneous at all temperatures.
  - This reaction will be nonspontaneous at all temperatures.
  - This reaction will be nonspontaneous only at high temperatures.
  - It is not possible to determine without more information.
- \_\_\_\_\_ 4. Consider a reaction that has a negative  $\Delta H$  and a positive  $\Delta S$ . Which of the following statements is TRUE?
- This reaction will be spontaneous only at high temperatures.
  - This reaction will be spontaneous at all temperatures.
  - This reaction will be nonspontaneous at all temperatures.
  - This reaction will be nonspontaneous only at high temperatures.
  - It is not possible to determine without more information.

\_\_\_\_\_ 5. Estimate  $\Delta G^\circ_{\text{rxn}}$  for the following reaction at 449.0 K.



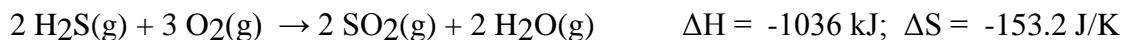
- a) +5.8 kJ
- b) +12.9 kJ
- c) -101 kJ
- d) +2.4 kJ
- e) -4.2 kJ

\_\_\_\_\_ 6. For the following example, which signs would you expect for enthalpy and entropy?



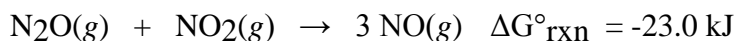
- a) a negative  $\Delta H$  and a negative  $\Delta S$
- b) a positive  $\Delta H$  and a negative  $\Delta S$
- c) a negative  $\Delta H$  and a positive  $\Delta S$
- d) a positive  $\Delta H$  and a positive  $\Delta S$
- e) It is not possible to determine without more information.

\_\_\_\_\_ 7. Above what temperature does the following reaction become nonspontaneous?



- a)  $6.762 \times 10^3 \text{ K}$
- b) 158.7 K
- c) 298 K
- d) This reaction is nonspontaneous at all temperatures.
- e) This reaction is spontaneous at all temperatures.

\_\_\_\_\_ 8. Given the following equation,

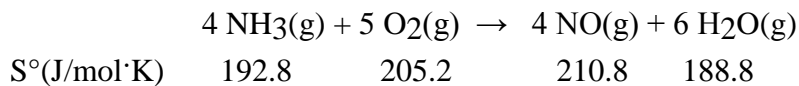


Calculate  $\Delta G^\circ_{\text{rxn}}$  for the following reaction.



- a) -23.0 kJ
- b) 69.0 kJ
- c) -69.0 kJ
- d) -7.67 kJ
- e) 23.0 kJ

\_\_\_\_\_9. Calculate  $\Delta S^\circ_{\text{rxn}}$  for the following reaction. The  $S^\circ$  for each species is shown below the reaction.

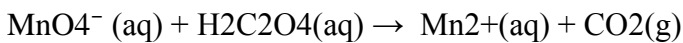


- a) +287.4 J/K
- b) -401.2 J/K
- c) +160.0 J/K
- d) -336.6 J/K
- e) +178.8 J/K

\_\_\_\_\_10. Which of the following is the reaction associated with the  $\Delta H_f^\circ$  of  $\text{NH}_3$ ?

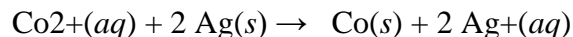
- a)  $\text{N}(\text{g}) + 3\text{H}(\text{g}) \rightarrow \text{NH}_3(\text{g})$
- b)  $\text{N}_2(\text{g}) + 3\text{H}(\text{g}) \rightarrow \text{NH}_3(\text{g})$
- c)  $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$
- d)  $1/2 \text{N}_2(\text{g}) + 3/2 \text{H}_2(\text{g}) \rightarrow \text{NH}_3(\text{g})$
- e)  $\text{NH}_3(\text{g}) \rightarrow 1/2 \text{N}_2(\text{g}) + 3/2 \text{H}_2(\text{g})$

\_\_\_\_\_11. What element is being oxidized in the following redox reaction?



- a) C
- b) O
- c) Mn
- d) H

\_\_\_\_\_12. What is the reduction half-reaction for the following overall galvanic cell reaction?

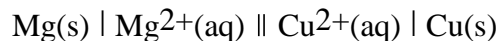


- a)  $\text{Ag}(\text{s}) + \text{e}^- \rightarrow \text{Ag}^+(\text{aq})$
- b)  $\text{Ag}^+(\text{aq}) + \text{e}^- \rightarrow \text{Ag}(\text{s})$
- c)  $\text{Co}^{2+}(\text{aq}) + 2 \text{e}^- \rightarrow \text{Co}(\text{s})$
- d)  $\text{Co}^{2+}(\text{aq}) + \text{e}^- \rightarrow \text{Co}(\text{s})$

\_\_\_\_\_13. Identify the location of oxidation in an electrochemical cell.

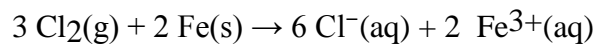
- a) the anode
- b) the cathode
- c) the electrode
- d) the salt bridge
- e) the socket

\_\_\_\_\_ 14. Determine the redox reaction represented by the following cell notation.



- a)  $\text{Cu(s)} + \text{Mg}^{2+}(\text{aq}) \rightarrow \text{Mg(s)} + \text{Cu}^{2+}(\text{aq})$
- b)  $\text{Mg(s)} + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Cu(s)} + \text{Mg}^{2+}(\text{aq})$
- c)  $2 \text{Mg(s)} + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Cu(s)} + 2 \text{Mg}^{2+}(\text{aq})$
- d)  $2 \text{Cu(s)} + \text{Mg}^{2+}(\text{aq}) \rightarrow \text{Mg(s)} + 2 \text{Cu}^{2+}(\text{aq})$
- e)  $3 \text{Mg(s)} + 2 \text{Cu}^{2+}(\text{aq}) \rightarrow 2 \text{Cu(s)} + 3 \text{Mg}^{2+}(\text{aq})$

\_\_\_\_\_ 15. Determine the cell notation for the redox reaction given below.



- a)  $\text{Cl}_2(\text{g}) \mid \text{Cl}^{-}(\text{aq}) \mid \text{Pt} \parallel \text{Fe(s)} \mid \text{Fe}^{3+}(\text{aq})$
- b)  $\text{Cl}^{-}(\text{aq}) \mid \text{Cl}_2(\text{g}) \mid \text{Pt} \parallel \text{Fe}^{3+}(\text{aq}) \mid \text{Fe(s)}$
- c)  $\text{Fe}^{3+}(\text{aq}) \mid \text{Fe(s)} \parallel \text{Cl}^{-}(\text{aq}) \mid \text{Cl}_2(\text{g}) \mid \text{Pt}$
- d)  $\text{Fe(s)} \mid \text{Cl}_2(\text{g}) \parallel \text{Fe}^{3+}(\text{aq}) \mid \text{Cl}^{-}(\text{aq}) \mid \text{Pt}$
- e)  $\text{Fe(s)} \mid \text{Fe}^{3+}(\text{aq}) \parallel \text{Cl}_2(\text{g}) \mid \text{Cl}^{-}(\text{aq}) \mid \text{Pt}$

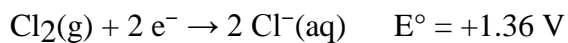
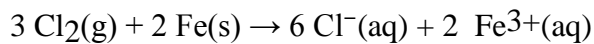
\_\_\_\_\_ 16. Which of the following is the strongest reducing agent?

- a)  $\text{Al(s)}$
- b)  $\text{Zn(s)}$
- c)  $\text{Mg(s)}$
- d)  $\text{Al}^{3+}(\text{aq})$
- e)  $\text{Mg}^{2+}(\text{aq})$

\_\_\_\_\_ 17. Which of the following metals will dissolve in nitric acid but not hydrochloric?

- a) Fe
- b) Pb
- c) Cu
- d) Sn
- e) Ni

\_\_\_\_\_18. Use the standard half-cell potentials listed below to calculate the standard cell potential for the following reaction occurring in an electrochemical cell at 25°C. (The equation is balanced.)



- a) +4.16 V
- b) -1.40 V
- c) -1.32 V
- d) +1.32 V
- e) +1.40 V

\_\_\_\_\_19. Determine which of the following pairs of reactants will result in a spontaneous reaction at 25°C.

- a)  $\text{Pb}^{2+}(\text{aq}) + \text{Cu}(\text{s})$
- b)  $\text{Ag}^+(\text{aq}) + \text{Br}^-(\text{aq})$
- c)  $\text{Li}^+(\text{aq}) + \text{Al}(\text{s})$
- d)  $\text{Fe}^{3+}(\text{aq}) + \text{Ni}(\text{s})$
- e) None of the above pairs will react.

\_\_\_\_\_20. How many electrons are transferred in the following reaction? (The reaction is unbalanced.)



- a) 1
- b) 2
- c) 6
- d) 3

Short Answer/Problems: Show your work!

1. Consider the following reaction:



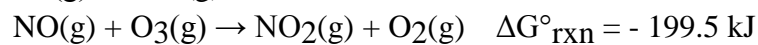
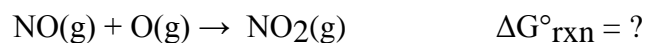
a) Calculate  $\Delta G_{\text{RXN}}^\circ$  for the reaction.

b) Calculate  $K_P$  for the reaction.

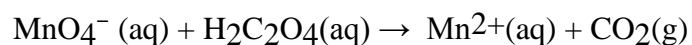
c) Calculate  $Q_P$  for the reaction if  $P_{\text{Hg}} = 0.001 \text{ atm}$ , and  $P_{\text{O}_2} = 0.52 \text{ atm}$

d) What is the value of  $\Delta G_{\text{RXN}}$  under the conditions in part c?

2. Use Hess's law to calculate  $\Delta G^\circ_{\text{rxn}}$  using the following information.



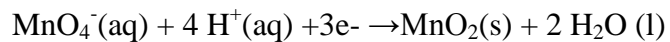
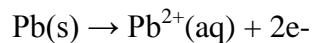
3. Balance the following redox reaction if it occurs in acidic solution.



Oxidation Half Reaction:	Reduction Half Reaction:

a) If this reaction occurred in basic conditions, how would the balancing process have changed?

4. An electrochemical cell is based on these two half-reactions:



a) Fill in the galvanic cell below. Indicate the direction of e- flow, the positions of each Reactant and Product (Pb(s),  $\text{Pb}^{2+}$ ,  $\text{MnO}_4^-$ ,  $\text{H}^+$  and  $\text{MnO}_2$ ). Because  $\text{MnO}_2$  does not conduct electricity, you will need to use a Pt electrode in one half-cell. Do not forget your salt bridge! Use a salt bridge containing  $\text{KNO}_3$  solution, and show the directions the ions flow.

b) Calculate  $E^\circ_{\text{cell}}$ .

c) If the concentrations of the aqueous ions are changed to the following, what is the new, non-standard  $E_{\text{cell}}$ ?  $[\text{Pb}^{2+}] = 0.15\text{M}$ ,  $[\text{MnO}_4^-] = 1.50\text{M}$ , and  $[\text{H}^+] = 2.0 \text{ M}$