NAME______________________________

S2014/CHEM1451/Dooley/Exam 4

Multiple Choice: (3Pts each)

______1. Which of the following processes have a $\Delta S > 0$?

a) $CH_3OH(l) \rightarrow CH_3OH(s)$
b) $N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$
c) $CH_4(g) + H_2O\ (g) \rightarrow CO(g) + 3\ H_2(g)$
d) $Na_2CO_3(s) + H_2O(g) + CO_2(g) \rightarrow 2\ NaHCO_3(s)$
e) $2\ NH_3(g) + CO_2(g) \rightarrow NH_2CONH_2(aq) + H_2O(l)$

______2. Consider the following reaction at constant P. Use the information here to determine the value of ΔS_{surr} at 298 K.

$N_2(g) + 2\ O_2(g) \rightarrow 2\ NO_2(g)\ \Delta H = +66.4\ kJ$

a) $\Delta S_{surr} = +223$ J/K
b) $\Delta S_{surr} = -223$ J/K
c) $\Delta S_{surr} = -66.4$ J/K
d) $\Delta S_{surr} = +66.4$ kJ/K
e) $\Delta S_{surr} = -66.4$ J/K

______3. Consider a reaction that has a positive ΔH and a positive ΔS. Which of the following statements is TRUE?

a) This reaction will be spontaneous only at high temperatures.
b) This reaction will be spontaneous at all temperatures.
c) This reaction will be nonspontaneous at all temperatures.
d) This reaction will be nonspontaneous only at high temperatures.
e) It is not possible to determine without more information.

______4. Consider a reaction that has a negative ΔH and a positive ΔS. Which of the following statements is TRUE?

a) This reaction will be spontaneous only at high temperatures.
b) This reaction will be spontaneous at all temperatures.
c) This reaction will be nonspontaneous at all temperatures.
d) This reaction will be nonspontaneous only at high temperatures.
e) It is not possible to determine without more information.

______5. Estimate $\Delta G°_{rxn}$ for the following reaction at 449.0 K.

$CH_2O(g) + 2\ H_2(g) \rightarrow CH_4(g) + H_2O(g)$ $\Delta H° = -94.9$ kJ; $\Delta S° = -224.2$ J/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?

$H_2O\ (l) \rightarrow H_2O(g)$

a) a negative ΔH and a negative ΔS
b) a positive ΔH and a negative ΔS
c) a negative ΔH and a positive ΔS
d) a positive ΔH and a positive ΔS
e) It is not possible to determine without more information.

______7. Above what temperature does the following reaction become nonspontaneous?

$2\ H_2S(g) + 3\ O_2(g) \rightarrow 2\ SO_2(g) + 2\ H_2O(g)$ $\Delta H = -1036$ kJ; $\Delta S = -153.2$ J/K

a) 6.762×10^3 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,

$N_2O(g)\ +\ NO_2(g) \rightarrow 3\ NO(g)$ $\Delta G°_{rxn} = -23.0$ kJ

Calculate $\Delta G°_{rxn}$ for the following reaction.

$9\ NO(g) \rightarrow 3N_2O(g)\ +\ 3NO_2(g)$

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

______9. Calculate ΔS°_{rxn} for the following reaction. The S° for each species is shown below the reaction.

$$4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g)$$

S°(J/mol·K)	192.8	205.2	210.8	188.8

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 ΔH_f° of NH_3?
a) $N(g) + 3H(g) \rightarrow NH_3(g)$
b) $N_2(g) + 3H(g) \rightarrow NH_3(g)$
c) $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
d) $1/2\ N_2(g) + 3/2\ H_2(g) \rightarrow NH_3(g)$
e) $NH_3(g) \rightarrow 1/2\ N_2(g) + 3/2\ H_2(g)$

______11. What element is being oxidized in the following redox reaction?

$MnO4^-$ (aq) + H2C2O4(aq) → Mn2+(aq) + CO2(g)

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

______12. What is the reduction half-reaction for the following overall galvanic cell reaction?

Co2+(*aq*) + 2 Ag(*s*) → Co(*s*) + 2 Ag+(*aq*)

a) Ag(*s*) + e- → Ag+(*aq*)
b) Ag+(*aq*) + e- → Ag(*s*)
c) Co2+(*aq*) + 2 e- → Co(*s*)
d) Co2+(*aq*) + e- → Co(*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.

$Mg(s) \mid Mg^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$

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

______15. Determine the cell notation for the redox reaction given below.

$3\ Cl_2(g) + 2\ Fe(s) \rightarrow 6\ Cl^{-}(aq) + 2\ Fe^{3+}(aq)$

a) $Cl_2(g) \mid Cl^{-}(aq) \mid Pt \parallel Fe(s) \mid Fe^{3+}(aq)$
b) $Cl^{-}(aq) \mid Cl_2(g) \mid Pt \parallel Fe^{3+}(aq) \mid Fe(s)$
c) $Fe^{3+}(aq) \mid Fe(s) \parallel Cl^{-}(aq) \mid Cl_2(g) \mid Pt$
d) $Fe(s) \mid Cl_2(g) \parallel Fe^{3+}(aq) \mid Cl^{-}(aq) \mid Pt$
e) $Fe(s) \mid Fe^{3+}(aq) \parallel Cl_2(g) \mid Cl^{-}(aq) \mid Pt$

______16. Which of the following is the strongest reducing agent?

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

______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.)

$3\ Cl_2(g) + 2\ Fe(s) \rightarrow 6\ Cl^-(aq) + 2\ Fe^{3+}(aq)$

$Cl_2(g) + 2\ e^- \rightarrow 2\ Cl^-(aq)$ E° = +1.36 V

$Fe^{3+}(aq) + 3\ e^- \rightarrow Fe(s)$ E° = -0.04 V

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

______20. How many electrons are transferred in the following reaction? (The reaction is unbalanced.)

$I_2(s) + Fe(s) \rightarrow Fe^{3+}(aq) + I^-(aq)$

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

Short Answer/Problems: Show your work!

1. Consider the following reaction:

$$2\ Hg(g) + O_2(g) \rightarrow 2\ HgO(s) \qquad \Delta H° = -304.2\ kJ;\ \ \Delta S° = -414.2\ J/K$$

a) Calculate ΔG°_{RXN} for the reaction.

b) Calculate K_P for the reaction.

c) Calculate Q_P for the reaction if $P_{Hg} = 0.001$ atm, and $P_{O2} = 0.52$ atm

d) What is the value of ΔG_{RXN} under the conditions in part c?

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

$NO(g) + O(g) \rightarrow NO_2(g)$ $\Delta G°_{rxn} = ?$

$2\ O_3(g) \rightarrow 3\ O_2(g)$ $\Delta G°_{rxn} = +489.6$ kJ
$O_2(g) \rightarrow 2\ O(g)$ $\Delta G°_{rxn} = +463.4$ kJ
$NO(g) + O_3(g) \rightarrow NO_2(g) + O_2(g)$ $\Delta G°_{rxn} = -199.5$ kJ

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

$MnO_4^-(aq) + H_2C_2O_4(aq) \rightarrow Mn^{2+}(aq) + CO_2(g)$

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:

$$Pb(s) \rightarrow Pb^{2+}(aq) + 2e-$$

$$MnO_4^-(aq) + 4\ H^+(aq) + 3e- \rightarrow MnO_2(s) + 2\ H_2O\ (l)$$

a) Fill in the galvanic cell below. Indicate the direction of e- flow, the positions of each Reactant and Product (Pb(s), Pb^{2+}, MnO_4^-, H^+ and MnO_2). Because 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 KNO_3 solution, and show the directions the ions flow.

b) Calculate E°_{cell}.

c) If the concentrations of the aqueous ions are changed to the following, what is the new, non-standard E_{cell}? $[Pb^{2+}] = 0.15M$, $[MnO_4^-] = 1.50M$, and $[H^+] = 2.0\ M$