MC, Chem1B, Fall11, Lec Exam3 $62-2=60$ 48 8 $\frac{116}{110}$ Name KEY [redacted]

Read questions **carefully** to **understand what is being asked**, before answering. No outside paper is allowed. Use the reverse side of your answer paper as scratch. Use the important equation table and periodic table provided. (Total points = 62 + 39 + 9 = 110).

Show your calculation first with set up equation. Then use the raw data with units in the equation in the equation and then complete the calculation.

1) The following information is available for the reaction at 25°C: 1) ____________

	$CaCO_3$ (s) ---->	CaO (s) +	CO_2 (g)
$\Delta G_f°$ (kJ/mol)	-1129.16	-603.42	-394.36
$\Delta H_f°$ (kJ/mol)	-1207.6	-635.09	-393.51
$S_f°$ (J/K.mol)	91.7	38.2	213.74

(a) Calculate the Gibbs free energy change of the reaction (3pts.). $\Delta G^0 = \Delta H^0 - T\Delta S^0$

$\Delta H_f^0 = \sum H_f^0(\text{product}) - \sum H_f^0(\text{Reactant}) = \sum(-393.51 + (-635.09)) - \sum(-1207.6) = (-1028 + 1207.6)\frac{kJ}{mol} = 179.0\ kJ/mol$

$\Delta_f S^0 = \sum S_f^0(\text{product}) - \sum S_f^0(\text{reactant}) = \sum(213.74 + 38.2) - \sum(91.7) = 251.94 - 91.7 = 160.24\ J/K.mol$

$\Delta G^0 = \Delta H^0 - T\Delta S^0 = 179000\frac{J}{mol} - 298K \cdot 160.24\frac{J}{K.mol} = 131248\frac{J}{mol} = \underline{\underline{131.248\ kJ/mol}}$

(b) Calculate the temperature in °C when the reaction will be favorable (5 pts.).

$\Delta G^0 = \Delta H - T\Delta S$ the reaction is favorable when $\Delta G°$ is negative so we set $\Delta H = T\Delta S$

$\Delta H = T\Delta S \Rightarrow T = \frac{\Delta H}{\Delta S} = \frac{179000\ J/mol}{160.24\ J/K.mol} = \underline{\underline{1117.07\ K}}$

$K = °C + 273$ $°C = K - 273$

$= (1117.07 - 273)°C = \underline{844.07°C}$ ✓

– the reaction will be favorable when the temperature is $> \underline{\underline{844.07°C}}$

2) The standard entropy of formation of $SiCl_4$(g), is -134.1 J/K·mol. 2) ____________

$Si(s) + 2\ Cl_2\ (g) \rightarrow SiCl_4\ (g)$

Calculate the standard molar entropy of $SiCl_4$(g) given: $S°[Si(s)] = 18.8$ J/K·mol-rxn and $S°[Cl_2(g)] = 223.1$ J/K·mol-rxn. (6 pts.)

$\Delta S_f^0 = \sum S^0(\text{product}) - \sum S^0(\text{reactant})$

$-134.1\frac{J}{K.mol} = \sum S^0(SiCl_4) - \sum S^0(Si(s) + 2S^0(Cl_2))$

$= S^0(SiCl_4) - \sum(18.8 + 2(223.1))\ J/K.mol)$

$-134.1 = S^0(SiCl_4) - 465.00$

$S^0(SiCl_4) = -134.1 + 465.00$

$\underline{\underline{S^0(SiCl_4) = 330.90\ J/K.mol}}$ ✓

3) Given: 3) ______________

$C (s) + O_2 (g) \rightarrow CO_2 (g)$ $\Delta G^0 = -394.4$ kJ

$2 CO (g) + O_2 (g) \rightarrow \cancel{2}CO_2 (g)$ $\Delta G^0 = -514.4$ kJ

Calculate the standard free energy of formation of CO (8 pts.):

$C (s) + 1/2\ O_2 (g) \rightarrow CO (g)$

$$C(s) + O_2(g) \longrightarrow CO_2(g) \qquad \Delta G^\circ = -394.4\,kJ$$

$$\tfrac{1}{2}\Big(2CO_2(g) \longrightarrow 2CO(g) + O_2(g) \qquad = +514.4\,kJ\Big)$$

$$C(s) + O_2(g) \longrightarrow \cancel{CO_2(g)} \qquad \Delta G^\circ = -394.4\,kJ$$

$$\cancel{CO_2(g)} \longrightarrow CO(g) + \tfrac{1}{2}O_2(g) \qquad = 257.2\,kJ/mol$$

$$\boxed{C(s) + \tfrac{1}{2}O_2(g) \longrightarrow CO(g) \qquad = -137.2\,kJ} \checkmark$$

4) A voltaic cell using the reaction below, has a standard cell potential of +1.19 V: 4) ______________

$Tl^{3+} (aq) + 2\ Cr^{2+} (aq) \longrightarrow Tl^{+} (aq) + 2\ Cr^{3+} (aq)$

-1

(a) Write the cathode cell reaction (2pt): $\overset{}{Ti}^{3+}(aq) + 2e^- \longrightarrow Ti^{+}(aq)$

-1

(b) Write the anode cell reaction (2pt): $\overset{Tl}{}\ \cancel{2}Cr^{2+}(aq) \longrightarrow \cancel{2}Cr^{3+}(aq) + \cancel{2}e^-$ $\;\overset{Tl}{}$

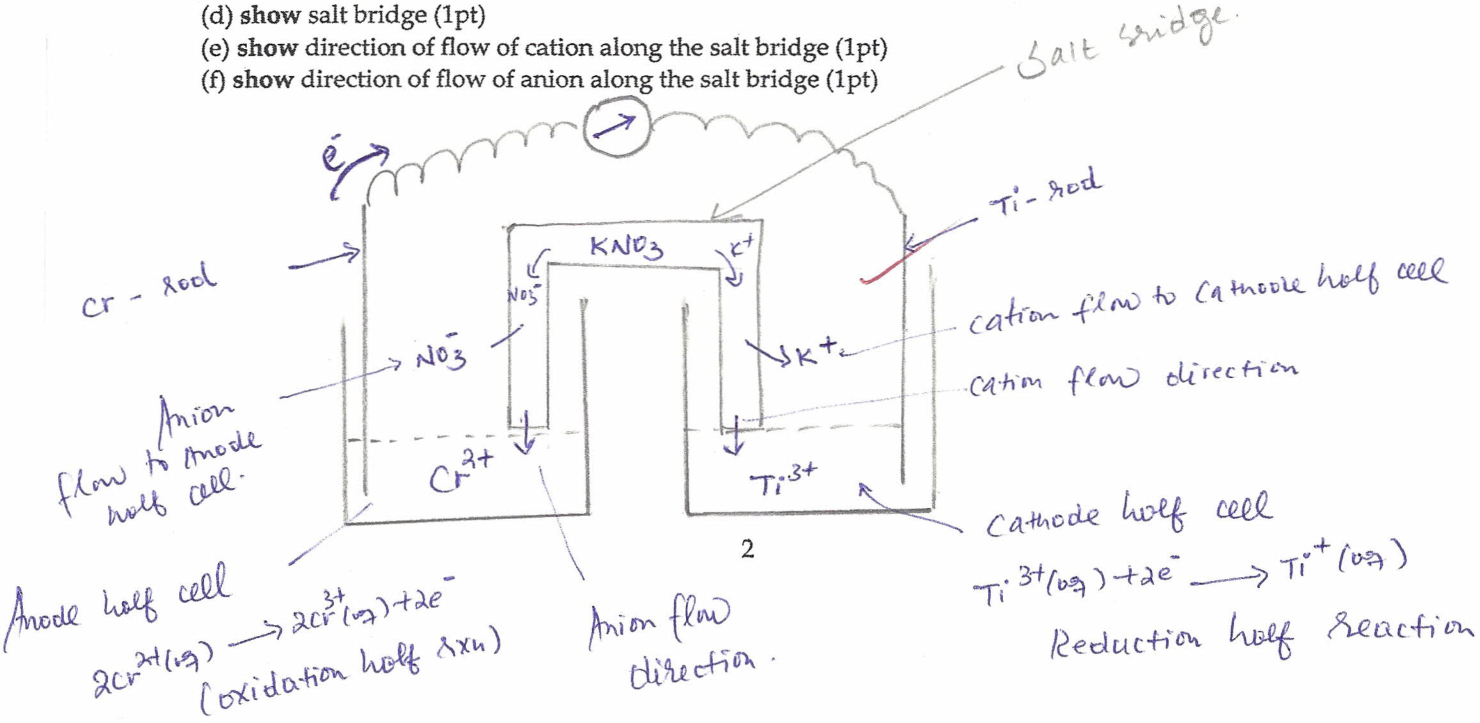
(c) If $Cr^{3+} (aq) + e^- \rightarrow Cr^{2+} (aq)$ has $E_{red}^0 = -0.43$ V; then calculate the E_{red}^0 for reduction of Tl^{3+} (aq) to Tl^{+} (aq) (4pt):

$$E^\circ_{cell} = E^\circ_{red}(\text{Cathode}) - E^\circ_{red}(\text{Anode}) = +1.19V = E^\circ_{red}(\text{cathode}) - (-0.43V)$$

$$E^\circ_{red}(\text{cathode}) = +1.19V - 0.43V = \underline{\underline{+0.76V}}$$

$$\therefore E^\circ_{red}(Ti^{3+} \rightarrow Ti^{+}) = \underline{\underline{+0.76V}}$$

(d) **Draw** a diagram of the voltaic cell. Make sure to
(a) **label** the **cathode** (1pt)
(b) **label** the **anode** (1pt)
(c) **indicate the direction** of electron flow outside the cell (1pt)
(d) **show** salt bridge (1pt)
(e) **show** direction of flow of cation along the salt bridge (1pt)
(f) **show** direction of flow of anion along the salt bridge (1pt)



5) The standard cell potential (E°) of a voltaic cell constructed using the cell reaction below is 0.76 V: 5) ______________

$$Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$$

With P_{H_2} = 1.0 atm and $[Zn^{2+}]$ = 1.0 M, the cell potential is 0.66 V. **Show set up and all your work** to calculate the concentration of H^+ in the cathode compartment (10

Given $E^0_{cell} = 0.76V$

$P_{H_2} = 1.0\ atm$

$[Zn^{2+}] = 1.0M$

$E_{cell} = 0.66V$

$Zn(s) \longrightarrow Zn^{2+} + 2e^-$

$2H^+(aq) + 2e^- \longrightarrow H_2(g)$

$\boxed{n = 2}$

$$E_{cell} = E^0_{cell} - \frac{0.0259}{n} \ln Q$$

$$Zn(s) + 2H^+(aq) \longrightarrow Zn^{2+}(aq) + H_2(g)$$

$$Q = \frac{[Zn^{2+}]P_{H_2}}{[H^+]^2}$$

$$E_{cell} = E^0_{cell} - \frac{0.0259}{n} \ln Q$$

$$0.66V = 0.76V - \frac{0.0259}{2} \ln \frac{[Zn^{2+}]P_{H_2}}{[H^+]^2}$$

$$0.66V = 0.76V - 0.01295 \ln \frac{1}{[H^+]^2}$$

$$0.66V - 0.76V = -0.01295 \ln \frac{1}{[H^+]^2}$$

$$-0.10V = -0.01290 \ln \frac{1}{[H^+]^2}$$

$$\frac{-0.100}{-0.012950} = \ln \frac{1}{[H^+]^2}$$

$$7.72 = \ln \frac{1}{[H^+]^2}$$

$$7.72 = \cancel{\ln 1} - \ln [H^+]^2 \ ; \quad -7.72 = \ln [H^+]^2$$

$$[H^+]^2 = e^{-7.72}$$

$$[H^+]^2 = 0.000444$$

$$\underline{\underline{[H^+] = 0.021M}}$$

✓

6) For the cell reaction : 6) ______________

$$3\,Ni^{2+}(aq) + 2\,Cr(OH)_3(s) + 10\,OH^-(aq) \longrightarrow 3\,Ni(s) + 2\,CrO_4^{2-}(aq) + 8\,H_2O(l);$$

Given: $CrO_4^{2-}(aq) + 4\,H_2O(l) + 3e^- \longrightarrow Cr(OH)_3(s) + 5\,OH^-(aq)$ $E^0_{red} = -0.13$ V; 2

$Ni^{2+}(aq) + 2e^- \longrightarrow Ni(s)$ $E^0_{red} = -0.28$ V; 3

$\boxed{n = 6}$

(a) Calculate the E^0 at T = 298 K (4 pts.)

$$E^0_{cell} = E^0_{red}(\text{Cathode}) - E^0_{red}(\text{Anode})$$

$$-0.28V - (-0.13V) = \underline{\underline{-0.15V}}$$

✓

(b) Calculate the standard free energy change ΔG^0 at T = 298 K for the above reaction (6 pts.);

$n = 6$, $E^0_{cell} = -0.15$] from the above equation.

$V = J/C \Rightarrow C = J/V$

$$\Delta G^0 = -nF\Delta E^0_{cell}$$

$$= -(6 \times 96485 \times (-0.15)) = 86836.5$$

$$e^- \times \frac{J \cdot V}{V \cdot mol \cdot e}$$

$$= \underline{\underline{86.837\ kJ/mol}}$$

✓

Unit for $\Delta G^0 = \frac{kJ}{mol}$

(c) Calculate the equilibrium constant, K at T = 298 K for the above reaction (6 pts.);

$$K = e^{\frac{nF\Delta E^{\circ}_{cell}}{RT}}$$

$= V = \frac{J}{C}$

$C = J/V$

$$= e^{\frac{6 \times 96485 \times (-0.15)}{8.3145 \times 298}}$$

$$\frac{(e \times \text{coulomb}/\text{mol}.e^{-})(V)}{J/\text{mol}.K.K}$$

$$\frac{(\cancel{e} \times \cancel{J}/\cancel{V}.\text{mol}.\cancel{e^{-}})(\cancel{V})}{\cancel{J}/\cancel{\text{mol}}.\cancel{K}.\cancel{K}} = \underline{\underline{\text{no unit for K}}}$$

$$= e^{-35.047}$$

$$\underline{\underline{K = 6.02 \times 10^{-16}}}$$ ✓

MULTIPLE CHOICE. Please answer on your scantron starting at bubble 7. Show your work to select the one response that best completes the statement or answers the question (3 pts each).

7) ________ electrons appear in the following half-reaction when it is balanced. 7) A

$S_4O_6^{2-} \rightarrow 2S_2O_3^{2-}$

A) 2 B) 3 C) 4 D) 1 E) 6

8) What is the oxidation number of nitrogen in the HNO_3? 8) A

A) +5 B) +7 C) -1 D) +1 E) +3

9) The half-reaction occurring at the anode in the balanced reaction shown below is ________. 9) C

$3MnO_4^- (aq) + 24H^+ (aq) + 5Fe (s) \rightarrow 3Mn^{2+} (aq) + 5Fe^{3+} (aq) + 12H_2O (l)$

A) $2MnO_4^- (aq) + 12H^+ (aq) + 6e^- \rightarrow 2Mn^{2+} (aq) + 3H_2O (l)$
B) $Fe^{2+} (aq) \rightarrow Fe^{3+} (aq) + e^-$
C) $Fe (s) \rightarrow Fe^{3+} (aq) + 3e^-$
D) $MnO_4^- (aq) + 8H^+ (aq) + 5e^- \rightarrow Mn^{2+} (aq) + 4H_2O (l)$
E) $Fe (s) \rightarrow Fe^{2+} (aq) + 2e^-$

10) In a voltaic cell lab experiment the E_{red} of an Al electrode was found to be -0.515 V relative to Cu, which acted as a refenece cell. The E_{cell} is 10) B

A) - 1.018 V B) + 0.515 V C) + 1.018 V D) - 0.515 V

Table 20.2

Half-reaction	E° (V)
$Cr^{3+} (aq) + 3e^- \rightarrow Cr (s)$	-0.74
$Fe^{2+} (aq) + 2e^- \rightarrow Fe (s)$	-0.440
$Fe^{3+} (aq) + e^- \rightarrow Fe^{2+} (s)$	+0.771
$Sn^{4+} (aq) + 2e^- \rightarrow Sn^{2+} (aq)$	+0.154

11) Which of the following reactions will occur spontaneously as written? 11) B

A) $Sn^{4+} (aq) + Fe^{2+} (s) \rightarrow Sn^{2+} (aq) + Fe (s)$
B) $2Cr (s) + 3Fe^{2+} (s) \rightarrow 3Fe (s) + 2Cr^{3+} (aq)$
C) $2Cr^{3+} (aq) + 3Sn^{2+} (aq) \rightarrow 3Sn^{4+} (aq) + 2Cr (s)$
D) $Sn^{2+} (aq) + Fe^{2+} (s) \rightarrow Sn^{4+} (aq) + Fe^{3+} (aq)$
E) $3Fe^{2+} (aq) + Cr^{3+} (aq) \rightarrow Cr (s) + 3Fe^{3+} (aq)$

$Fe^{3+} (aq) + e^- \rightarrow Fe^{2+} (s) = 0.771$

* $Sn^{4+} (aq) + 2e^- \rightarrow Sn^{2+} (aq) = 0.154$

* $Fe^{2+} (aq) + 2e^- \rightarrow Fe (s) = -0.440$

$Cr^{3+} (aq) + 3e^- \rightarrow Cr (s) = -0.74$

Ⓑ $-0.440 + 0.74 =$ spon.

$0.771 - 0.154$

12) The standard cell potential ($E°_{cell}$) for the voltaic cell based on the reaction below is ______ V. 12) B

$Sn^{2+} (aq) + 2Fe^{3+} (aq) \rightarrow 2Fe^{2+} (aq) + Sn^{4+} (aq)$

A) +1.21 B) +0.617 C) +0.46 D) +1.39 E) -0.46

13) The reduction half reaction occurring in the standard hydrogen electrode is ______. 13) D

A) H_2 (g, 1 atm) → $2H^+$ (aq, 1M) + $2e^-$
B) $2H^+ (aq) + 2OH^- \rightarrow H_2O (l)$
C) $2H^+$ (aq, 1M) + Cl_2 (aq) → 2HCl (aq)
D) $2H^+$ (aq, 1M) + $2e^-$ → H_2 (g, 1 atm)
E) $O_2 (g) + 4H^+ (aq) + 4e^- \rightarrow 2H_2O (l)$

14) The standard cell potential ($E°_{cell}$) for the reaction below is +1.10 V. The cell potential for this reaction is ______ V when the concentration of $[Cu^{2+}] = 1.0 \times 10^{-5}$ M and $[Zn^{2+}] = 1.0$ M. 14) A

$Zn (s) + Cu^{2+} (aq) \rightarrow Cu (s) + Zn^{2+} (aq)$

A) 0.95 B) 0.80 C) 1.25 D) 1.10 E) 1.40

15) The thermodynamic quantity that expresses the degree of disorder in a system is ______. 15) B

A) bond energy
B) entropy
C) internal energy
D) enthalpy
E) heat flow

16) The normal boiling point of water is 100.0°C and its molar enthalpy of vaporization is 40.67 kJ/mol. What is the change in entropy in the system in J/K when 39.3 grams of steam at 1 atm condenses to a liquid at the normal boiling point? 16) E

Vap $\xrightarrow{con}$ liqi

Exothermic

A) 373 B) 88.8 C) -40.7 D) -88.8 E) -238

17) ΔS is positive for the reaction ______. 17) C

A) $2NO_2 (g) \rightarrow N_2O_4 (g)$
B) $2Hg (l) + O_2 (g) \rightarrow 2HgO (s)$
C) $BaF_2 (s) \rightarrow Ba^{2+} (aq) + 2F^- (aq)$
D) $CO_2 (g) \rightarrow CO_2 (s)$
E) $2H_2 (g) + O_2 (g) \rightarrow 2H_2O (g)$

18) Of the following, the entropy of ______ is the largest. 18) E

A) HCl (s) B) HCl (g) C) HCl (l) D) HBr (g) E) HI (g)

19) The standard Gibbs free energy of formation of ________ is zero. 19) D

(a) H_2O (l)
(b) Na (s)
(c) H_2 (g)

A) (a) only
B) (b) only
C) (c) only
D) (b) and (c)
E) (a), (b), and (c)

TRUE/FALSE. Select A in the scantron if the statement is TRUE and B if the statement is FALSE (3 pts).

20) The standard reduction potential, E^{o}_{red}, is proportional to the stoichiometric coefficient. T or [✓ F]

21) The entropy of a pure crystalline substance at 0°C is zero. 0K T or [✓ F]

22) The more negative ΔG° is for a given reaction, the larger the value of the corresponding equilibrium constant, K. [✓ T] or F

Show your calculation first with set up equation. Then use the raw data with units in the equation in the equation and then complete the calculation.

23) **Extra Point Question:** 23) ____________

Given	$Cl_2\ (g) + 2\ e^- \longrightarrow 2\ Cl^-\ (aq)$	$E^0 = +\ 1.36\ V$
	$I_2\ (s) + 2\ e^- \longrightarrow 2\ I^-\ (aq)$	$E^0 = +\ 0.54\ V$

Use above information for the redox reaction $Cl_2\ (g) + 2\ I^-\ (aq) \rightarrow 2\ Cl^-\ (aq) + I_2\ (s)$ and write the

(i) cathode half reaction (2 pts.)

$$\rightarrow Cl_2(g) + 2e^- \longrightarrow 2Cl^-(aq)$$ ✓

(ii) anode half reaction (2 pts.)

$$2I^-(aq) \longrightarrow I_2(s) + 2e^-$$ ✓

(iii) Show all your quantitave work to decide if the above redox reaction is spontaneous or not. (4 pts):

$$E^0_{cell} = E^0_{red}\ (\text{Cathode}) - E^0_{red}\ (\text{Anode})$$

$$= +1.36V - 0.54$$

$$E^0_{cell} = \underline{0.82V}$$ ✓

Since the standard cell potential is positive, then the redox reaction is <u>spontaneous.</u>