

CHEMISTRY 123 – FALL 2010

Midterm #2



Test Booklet A - For Question 1

Your name: _____

Your Student ID number: _____

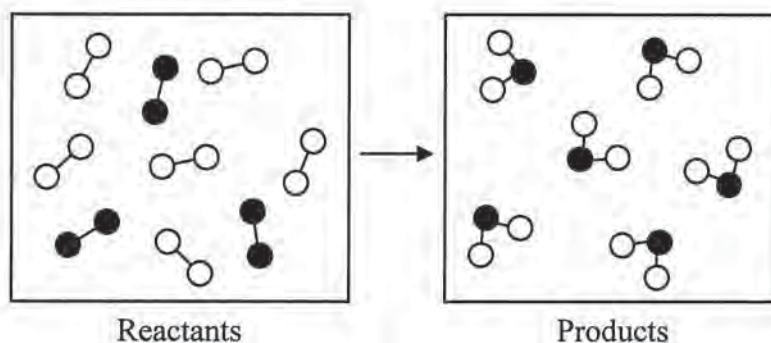
Your TA: _____

When you are finished with the exam, hand in the scantron sheet to your TA. Keep the exam booklet and check your answers with the key posted on Carmen. For homework, rework the questions you answered incorrectly, as well as the ones you answered correctly, but guessed the correct answer.

There are multiple versions of the exam. You are taking Version A.
Unless otherwise stated all reactions are assumed to occur at 25 °C, and 1 atm.

1. [0 Points] What test booklet do you have? This is on the first page.
A) A
B) B

For problems 1 and 2 consider the reaction depicted below (all molecules shown are in the gas phase, assume all bond enthalpies to be the same):



2. [7 points] Predict the sign of ΔH_{rxn} and ΔS_{rxn} for the reaction shown above.
- ΔS_{rxn} is positive, ΔH_{rxn} is positive
 - ΔS_{rxn} is positive, ΔH_{rxn} is negative
 - ΔS_{rxn} is negative, ΔH_{rxn} is positive
 - ΔS_{rxn} is negative, ΔH_{rxn} is negative
 - ΔS_{rxn} is negative, ΔH_{rxn} is zero
3. [7 points] What general statement can you make about the relationship between spontaneity and temperature for the reaction shown above?
- The reaction is spontaneous at all temperatures
 - The reaction is not spontaneous at any temperature
 - The reaction becomes more spontaneous as the temperature decreases
 - The reaction becomes more spontaneous as the temperature increases
 - There is not enough information given to answer this question
4. [7 points] For which of the following reactions would you expect ΔS_{rxn} to be negative?
- $\text{SO}_2(\text{g}) + 2\text{H}_2(\text{g}) \rightarrow \text{S}(\text{s}) + 2\text{H}_2\text{O}(\text{g})$
 - $\text{NH}_4\text{Cl}(\text{s}) \rightarrow \text{NH}_3(\text{g}) + \text{HCl}(\text{g})$
 - $\text{CoCl}_2(\text{s}) \rightarrow \text{Co}^{2+}(\text{aq}) + 2\text{Cl}^{-}(\text{aq})$
 - $2\text{H}_2\text{O}(\text{g}) \rightarrow 2\text{H}_2(\text{g}) + \text{O}_2(\text{g})$
 - $\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{O}(\text{g})$

5. [7 points] Which of the following reactions should occur faster?
- a. $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l})$ $\Delta G^\circ = -474.26 \text{ kJ/mol}$
- b. $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$ $\Delta G^\circ = -817.86 \text{ kJ/mol}$
- c. $2\text{Mg}(\text{s}) + \text{O}_2(\text{g}) \rightarrow 2\text{MgO}(\text{s})$ $\Delta G^\circ = -1139.2 \text{ kJ/mol}$
- d. $2\text{NaBr} \rightarrow 2\text{Na}(\text{s}) + \text{Br}_2(\text{l})$ $\Delta G^\circ = +698.6 \text{ kJ/mol}$
- e. There is not enough information given to answer this question

For questions 6-7 refer to the following table;

Chemical	ΔH_f° (kJ mol ⁻¹)	ΔG_f° (kJ mol ⁻¹)	S° (J mol ⁻¹ K ⁻¹)
TiCl ₄ (l)	-804.2	-728.2	221.9
TiCl ₄ (g)	-763.2	-726.9	354.9
H ₂ S(g)	-20.17	-33.01	205.6
Li ₂ S(s)	-441.8	-435.9	70.1
LiCl(s)	-408.3	-384.0	59.30
TiS ₂ (s)	-415.0	-402.6	52.9
Cl ₂ (g)	0	0	222.96
HCl(g)	-92.3	-95.27	186.69

6. [7 points] One of the current projects in Prof. Goldberger's lab involves the synthesis of TiS₂ as a thermoelectric material, a material that converts heat directly into electricity.

What is the ΔG° of the following reaction?

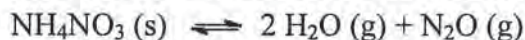


- a. 378 kJ/mol
- b. 423 kJ/mol
- c. -339 kJ/mol
- d. -360 kJ/mol
- e. -72 kJ/mol
7. [7 points] If Prof. Goldberger attempted to synthesize TiS₂, using H₂S as the sulfur source, at what temperature would he need to carry out the following reaction, if we assume entropy and enthalpy do not change as a function of temperature?



- a. 50 °C
- b. 150 °C
- c. 250 °C
- d. 350 °C
- e. None of the above

8. [7 points] The commercial production of laughing gas (N_2O) involves the careful decomposition of ammonium nitrate according to the following reaction.



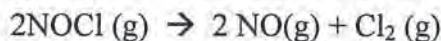
Given the following table of thermodynamic data at 298 K:

Substance	ΔH_f° (kJ mol ⁻¹)	S° (J mol ⁻¹ K ⁻¹)
$\text{NH}_4\text{NO}_3 (\text{s})$	-365.6	151
$\text{H}_2\text{O} (\text{g})$	-241.82	188.83
$\text{N}_2\text{O} (\text{g})$	81.6	220.0

Assuming entropy and enthalpy do not change as a function of temperature, the value of the equilibrium constant K_{eq} for the reaction at 170 °C is _____.

- a. $K_{\text{eq}} = 2.3 \times 10^2$
 b. $K_{\text{eq}} = 4.4 \times 10^{27}$
 c. $K_{\text{eq}} = 1.1$
 d. $K_{\text{eq}} = 2.3 \times 10^{-28}$
 e. $K_{\text{eq}} = 5.2 \times 10^{29}$
9. [7 points] Calculate ΔG for the following reaction, given the pressures listed

Substance	ΔG_f° (kJ mol ⁻¹)
$\text{NOCl} (\text{g})$	66.3
$\text{NO} (\text{g})$	86.71
$\text{Cl}_2 (\text{g})$	0



$P_{\text{NOCl}} = 0.30 \text{ atm}$, $P_{\text{NO}} = 2.4 \text{ atm}$, $P_{\text{Cl}_2} = 0.50 \text{ atm}$

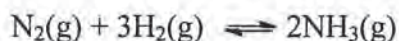
- a. 17 kJ/mol
 b. 32 kJ/mol
 c. 37 kJ/mol
 d. 41 kJ/mol
 e. 49 kJ/mol

10. [7 points] Calculate the K_{sp} of $FeCl_3$ at room temperature, given the following thermodynamic data;

Substance	ΔG_f° (kJ mol ⁻¹)
$FeCl_3$ (s)	-334
Fe^{3+} (aq)	-10.54
Cl^- (aq)	-131.2

- a. 2.1×10^{-34}
 b. 1.9×10^{12}
 c. 0.925
 d. 1.03
 e. 2.0×10^{-11}

11. [7 points] Consider the Haber Process.

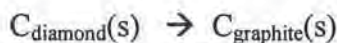


Substance	ΔG_f° (kJ mol ⁻¹)
NH_3 (g)	-16.7

If you prepared a vessel at room temperature containing $P_{H_2} = 0.05$ atm, $P_{N_2} = 0.03$ atm, and $P_{NH_3} = 1$ atm, in which direction would the equilibrium shift?

- a. The equilibrium would shift toward reactants.
 b. The equilibrium would shift towards products.
 c. The system is already under equilibrium
 d. It is impossible to determine with the information given.

12. [7 points] The old adage states "Diamonds are Forever." To determine if this is true, First, calculate the equilibrium constant for the interconversion of Carbon diamond into Carbon graphite; and Second, determine if this conversion is spontaneous at room temperature? The interconversion reaction is;



Compound	ΔH_f° (kJ mol ⁻¹)	S° (J mol ⁻¹ K ⁻¹)
C (s) diamond	1.88	2.43
C (s) graphite	0	5.69

- a. 0.317, Nonspontaneous
 b. 0.317, Spontaneous
 c. 2.85, Spontaneous
 d. 3.16, Spontaneous
 e. -974, Nonspontaneous

13. [7 points] What is the sign of the ΔH° , ΔS° , and ΔG° for the melting of ice at 0 °C?

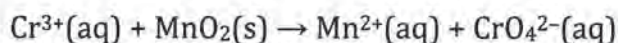
Answer	ΔH°	ΔS°	ΔG°
a.	+	+	0
b.	+	-	+
c.	-	+	0
d.	-	-	-
e.	0	+	-

14. [7 points] Which statement is true for the following electrochemical reaction?



- (a) Cu^{2+} is oxidized
- (b) The $\text{Ag}^+(\text{aq})|\text{Ag}(\text{s})$ half reaction occurs at the anode
- (c) Ag^+ acts as the reducing agent
- (d) Ag^+ is reduced
- (e) Both (c) and (d) are true

15. [7 points] Balance the following redox reaction in basic solution.



- (a) ~~$2 \text{H}_2\text{O}(\text{l}) + 2 \text{Cr}^{3+}(\text{aq}) + 3 \text{MnO}_2(\text{s}) \rightarrow 3 \text{Mn}^{2+}(\text{aq}) + 2 \text{CrO}_4^{2-}(\text{aq}) + 4 \text{H}^+(\text{aq})$~~
- (b) $2 \text{Cr}^{3+}(\text{aq}) + 4 \text{OH}^-(\text{aq}) + 3 \text{MnO}_2(\text{s}) \rightarrow 2 \text{CrO}_4^{2-}(\text{aq}) + 2 \text{H}_2\text{O}(\text{l}) + 3 \text{Mn}^{2+}(\text{aq})$
- (c) ~~$2 \text{OH}^-(\text{aq}) + \text{Cr}^{3+}(\text{aq}) + \text{MnO}_2(\text{s}) \rightarrow \text{Mn}^{2+}(\text{aq}) + \text{CrO}_4^{2-}(\text{aq}) + 2 \text{H}^+(\text{aq})$~~
- (d) $\text{Cr}^{3+}(\text{aq}) + 2 \text{O}_2(\text{g}) + \text{MnO}_2(\text{s}) \rightarrow \text{CrO}_4^{2-}(\text{aq}) + \text{Mn}^{2+}(\text{aq})$
- (e) $2 \text{OH}^-(\text{aq}) + \text{Cr}^{3+}(\text{aq}) + \text{MnO}_2(\text{s}) \rightarrow \text{Mn}^{2+}(\text{aq}) + \text{CrO}_4^{2-}(\text{aq}) + 4 \text{H}_2\text{O}(\text{l})$

16. [7 points] Which of the following species is the best oxidizing agent?

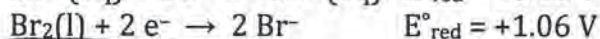
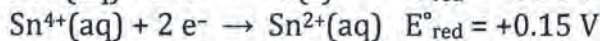
- (a) $\text{Al}(\text{s})$
- (b) $\text{Al}^{3+}(\text{aq})$
- (c) $\text{Fe}^{3+}(\text{aq})$
- (d) $\text{Fe}^{2+}(\text{aq})$
- (e) $\text{H}_2(\text{g})$

17. [7 points] An electrochemical process occurs in a voltaic cell. What is the cell EMF for this reaction?



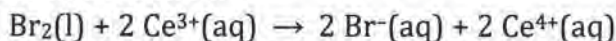
- (a) +1.76 V
- (b) +1.24 V
- (c) +2.22 V
- (d) -1.24 V
- (e) +3.78 V

18. [7 points] Which of the following statements is correct if each substance is in its standard state? Assume potentials are given in water at 25 °C.

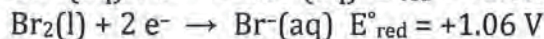
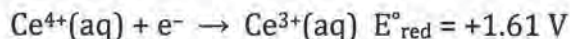


- (a) $\text{Sn}^{4+}(\text{aq})$ will oxidize $\text{Ni}(\text{s})$ to $\text{Ni}^{2+}(\text{aq})$
- (b) $\text{Br}_2(\text{l})$ will reduce $\text{Sn}^{4+}(\text{aq})$ to $\text{Sn}^{2+}(\text{aq})$
- (c) $\text{Sn}^{4+}(\text{aq})$ will oxidize $\text{Br}^{-}(\text{aq})$ to $\text{Br}_2(\text{l})$
- (d) $\text{Ni}^{2+}(\text{aq})$ will reduce $\text{Sn}^{4+}(\text{aq})$ to $\text{Sn}^{2+}(\text{aq})$
- (e) $\text{Sn}^{4+}(\text{aq})$ will reduce $\text{Ni}^{2+}(\text{aq})$ to $\text{Ni}(\text{s})$

19. [7 points] Consider this reaction:

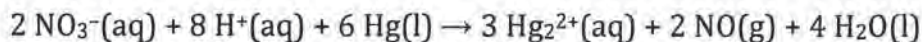


Calculate the cell potential, E , when $[\text{Ce}^{3+}] = 0.75 \text{ M}$, $[\text{Ce}^{4+}] = 0.015 \text{ M}$, and $[\text{Br}^{-}] = 0.040$.



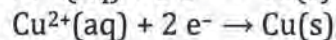
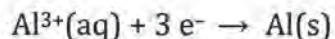
- (a) -0.18 V
- (b) -0.37 V
- (c) -0.73 V
- (d) $+2.67 \text{ V}$
- (e) -0.55 V

20. [7 points] Which of the following set of conditions given below would produce the most positive cell potential for the following reaction:



- (a) $[\text{NO}_3^{-}] = 1.0 \text{ M}$; $[\text{Hg}_2^{2+}] = 1.0 \text{ M}$; $P_{\text{NO}(\text{g})} = 1.0 \text{ atm}$; $\text{pH} = 7.0$
- (b) $[\text{NO}_3^{-}] = 2.0 \text{ M}$; $[\text{Hg}_2^{2+}] = 0.50 \text{ M}$; $P_{\text{NO}(\text{g})} = 0.75 \text{ atm}$; $\text{pH} = 10.0$
- (c) $[\text{NO}_3^{-}] = 0.5$; $[\text{Hg}_2^{2+}] = 2.0 \text{ M}$; $P_{\text{NO}(\text{g})} = 2.0 \text{ atm}$; $\text{pH} = 10.0$
- (d) $[\text{NO}_3^{-}] = 2.0$; $[\text{Hg}_2^{2+}] = 0.50$; $P_{\text{NO}(\text{g})} = 0.75 \text{ atm}$; $\text{pH} = 5.0$
- (e) $[\text{NO}_3^{-}] = 0.5$; $[\text{Hg}_2^{2+}] = 2.0 \text{ M}$; $P_{\text{NO}(\text{g})} = 2.0 \text{ atm}$; $\text{pH} = 2.0$

21. [7 points] Determine the value for ΔG° for the following reactions carried out in a voltaic cell:



- (a) -1158 kJ/mol
(b) -193 kJ/mol
(c) -579 kJ/mol
(d) -386 kJ/mol
(e) +579 kJ/mol
22. [7 points] Determine the value of the equilibrium constant, K , at 25°C for the following reaction:



- (a) 1×10^{42}
(b) 1×10^{-22}
(c) 1×10^{-11}
(d) 1×10^{22}
(e) 1×10^{11}
23. [7 points] What mass of aluminum could be plated on an electrode from the electrolysis of a $\text{Al}(\text{NO}_3)_3$ solution with a current of 2.50 A for 30.0 minutes?
- (a) 0.140 grams
(b) 0.00699 grams
(c) 0.419 grams
(d) 1.40 grams
(e) 0.0210 grams

24. [7 points] Manganese is a transition metal with varying oxidation states. Using a manganese salt, 2.89×10^5 Coulombs plate out 41.2 g of manganese. What form of manganese ion is in the solution of this salt?

- (a) Mn^+
(b) Mn^{2+}
(c) Mn^{3+}
(d) Mn^{4+}
(e) Mn^{7+}

25. [7 points] Which one or more of the following would serve to protect an iron pipe from corrosion by serving as the anode in a cathodic protection scheme?

(a) Zn(s)

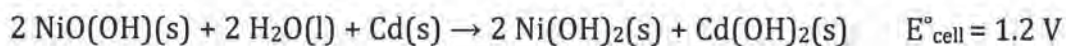
(b) Cu(s)

(c) Ag(s)

(d) Al(s)

(e) Both (a) and (d)

26. [7 points] Consider the cell reaction of a nickel-cadmium battery.



What is the $\Delta G^\circ_{\text{rxn}}$ of the nickel-cadmium battery reaction?

(a) -231 kJ/mol

(b) -116 kJ/mol

(c) -347 kJ/mol

(d) -463 kJ/mol

(e) +116 kJ/mol

1) A

2) **Step 1**
→ determine sign of ΔS_{rxn}

9 molecules of gas \rightarrow 6 molecules of gas

• moves towards more order
 \therefore entropy decreases and ΔS_{rxn} is negative

Step 2
→ determine sign of ΔH_{rxn}

$$\begin{aligned}\Delta H_{rxn} &= (\text{bonds broken}) - (\text{bonds formed}) \\ &= 9 - 12 \\ &= -3\end{aligned}$$

ΔH_{rxn} is negative

3) Use the equation, $\Delta G = \Delta H - T\Delta S$.

• A rxn is spontaneous when ΔG is negative

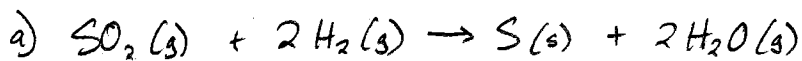
Using answers from previous question:

ΔH	ΔS	ΔG
-	-	+ @ High T
-	-	- @ Low T

\therefore rxn becomes more spontaneous as T decreases.

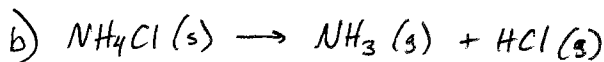
4) • ΔS_{rxn} is negative when there is a decrease in entropy (disorder or randomness)

Entropy: solid < liquid < gas



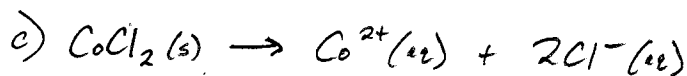
3 mol gas \rightarrow 2 mol gas

• entropy decreases, so ΔS_{rxn} is negative



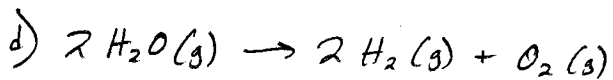
1 mol solid \rightarrow 2 mol gas

• entropy increases



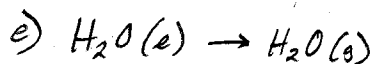
1 mol solid \rightarrow 3 mol aqueous ions

• entropy increases



2 mol gas \rightarrow 3 mol gas

• entropy increases



1 mol liquid \rightarrow 1 mol gas

• entropy increases

5) • Gibbs free energy, ΔG , only tells us whether or not a rxn is spontaneous. It does not tell us how fast a rxn takes place.

\therefore there is not enough information given to answer the question.

6) • Use given table

$$\Delta G_{\text{rxn}}^{\circ} = \sum \Delta G_{\text{f}}^{\circ} \text{ products} - \sum \Delta G_{\text{f}}^{\circ} \text{ reactants}$$

$$= \Delta G_{\text{f}}^{\circ} [\text{TiS}_2(\text{s})] + 4\Delta G_{\text{f}}^{\circ} [\text{LiCl}(\text{s})] - \Delta G_{\text{f}}^{\circ} [\text{TiCl}_4(\text{l})] - 2\Delta G_{\text{f}}^{\circ} [\text{Li}_2\text{S}(\text{s})]$$

$$= -402.6 + 4(-384.0) - (-728.2) - 2(-435.9)$$

$$= \boxed{-339 \text{ kJ/mol}}$$

Step 1
→ find ΔS_{rxn}°

$$\begin{aligned}\Delta S_{rxn}^{\circ} &= \sum S^{\circ}_{\text{products}} - \sum S^{\circ}_{\text{reactants}} \\ &= S^{\circ}[\text{TiS}_2(\text{s})] + 4S^{\circ}[\text{HCl}(\text{g})] - S^{\circ}[\text{TiCl}_4(\text{g})] - 2S^{\circ}[\text{H}_2\text{S}(\text{g})] \\ &= 52.9 + 4(186.69) - (354.9) - 2(205.6) \\ &= 33.56 \text{ J/mol}\cdot\text{K} \\ &= 33.56 \times 10^{-3} \text{ kJ/mol}\cdot\text{K}\end{aligned}$$

Step 2
→ find ΔH_{rxn}°

$$\begin{aligned}\Delta H_{rxn}^{\circ} &= \sum \Delta H_f^{\circ} \text{ products} - \sum \Delta H_f^{\circ} \text{ reactants} \\ &= \Delta H_f^{\circ}[\text{TiS}_2(\text{s})] + 4\Delta H_f^{\circ}[\text{HCl}(\text{g})] - \Delta H_f^{\circ}[\text{TiCl}_4(\text{g})] - 2\Delta H_f^{\circ}[\text{H}_2\text{S}(\text{g})] \\ &= (-415.0) + 4(-92.3) - (-763.2) - 2(-20.17) \\ &= 19.34 \text{ kJ/mol}\end{aligned}$$

Step 3
→ find T where rxn is spontaneous ($\Delta G < 0$)

$$\Delta G_{rxn}^{\circ} = \Delta H_{rxn}^{\circ} - T\Delta S_{rxn}^{\circ}$$

spont. rxn

$$\rightarrow \Delta G_{rxn}^{\circ} < 0$$

$$\Delta H_{rxn}^{\circ} - T\Delta S_{rxn}^{\circ} < 0$$

$$19.34 - T(33.56 \times 10^{-3}) < 0$$

$$T > 576 \text{ K}$$

$$T > 303^{\circ}\text{C}$$

∴ the answer is d) 350°C

8) **Step 1**
↳ find $\Delta S^\circ_{\text{rxn}}$

$$\begin{aligned}\Delta S^\circ_{\text{rxn}} &= \sum S^\circ_{\text{products}} - \sum S^\circ_{\text{reactants}} \\ &= 2S^\circ[\text{H}_2\text{O}(g)] + S^\circ[\text{N}_2\text{O}(g)] - S^\circ[\text{NH}_4\text{NO}_3(s)] \\ &= 2(188.83) + (220.0) - (151) \\ &= 447 \text{ J/mol}\cdot\text{K} \\ &= 0.447 \text{ kJ/mol}\cdot\text{K}\end{aligned}$$

Step 2
↳ find $\Delta H^\circ_{\text{rxn}}$

$$\begin{aligned}\Delta H^\circ_{\text{rxn}} &= \sum \Delta H^\circ_f \text{ products} - \sum \Delta H^\circ_f \text{ reactants} \\ &= 2\Delta H^\circ_f[\text{H}_2\text{O}(g)] + \Delta H^\circ_f[\text{N}_2\text{O}(g)] - \Delta H^\circ_f[\text{NH}_4\text{NO}_3(s)] \\ &= 2(-241.82) + (81.6) - (-365.6) \\ &= -36.4 \text{ kJ/mol}\end{aligned}$$

Step 3
↳ Find $\Delta G^\circ_{\text{rxn}}$ using $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$

$$\begin{aligned}\Delta G^\circ_{\text{rxn}} &= \Delta H^\circ_{\text{rxn}} - T\Delta S^\circ_{\text{rxn}} \\ &= (-36.4) - (443)(0.447) \\ &= -234.4 \text{ kJ/mol}\end{aligned}$$

Step 4
↳ Solve for K_{eq}

$$\begin{aligned}\Delta G^\circ &= -RT \ln(K_{\text{eq}}) \\ -234.4 &= -(8.314 \times 10^{-3})(443) \ln(K_{\text{eq}})\end{aligned}$$

$$\rightarrow \boxed{K_{\text{eq}} = 4.4 \times 10^{27}}$$

9) Under Non-Standard Conditions

$$\Delta G_{\text{rxn}} = \Delta G_{\text{rxn}}^{\circ} + RT \ln(Q)$$

• The standard pressure for gas is 1 atm

Step 1

→ Find $\Delta G_{\text{rxn}}^{\circ}$

$$\begin{aligned}\Delta G_{\text{rxn}}^{\circ} &= \sum \Delta G_f^{\circ} \text{ products} - \sum \Delta G_f^{\circ} \text{ reactants} \\ &= 2\Delta G_f^{\circ}[\text{NO}(g)] + \Delta G_f^{\circ}[\text{Cl}_2(g)] - 2\Delta G_f^{\circ}[\text{NOCl}(g)] \\ &= 2(86.71) + (0) - 2(66.3) \\ &= 40.82\end{aligned}$$

Step 2

→ Find Q of the rxn

$$Q = \frac{(P_{\text{NO}})^2 (P_{\text{Cl}_2})}{(P_{\text{NOCl}})^2} = \frac{(2.4)^2 (0.5)}{(0.3)^2} = 32$$

Step 3

→ Solve for ΔG_{rxn}

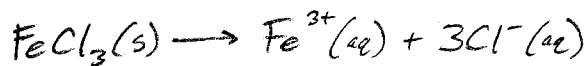
$$\begin{aligned}\Delta G_{\text{rxn}} &= \Delta G_{\text{rxn}}^{\circ} + RT \ln(Q) \\ &= 40.82 + (8.314 \times 10^{-3})(298) \ln(32) \\ &= \boxed{49 \text{ kJ/mol}}\end{aligned}$$

10) Under Standard Conditions

$$\Delta G = \Delta G^\circ = -RT \ln(K)$$

Step 1

↳ Find $\Delta G^\circ_{\text{rxn}}$



$$\Delta G^\circ_{\text{rxn}} = \sum \Delta G^\circ_f \text{ products} - \sum \Delta G^\circ_f \text{ reactants}$$

$$= \Delta G^\circ_f [\text{Fe}^{3+}(aq)] + 3\Delta G^\circ_f [\text{Cl}^-(aq)] - \Delta G^\circ_f [\text{FeCl}_3(s)]$$

$$= (-10.54) + 3(-131.2) - (-334)$$

$$= -170.1 \text{ kJ/mol}$$

Step 2

↳ Solve for K_{sp}

$$\Delta G^\circ_{\text{rxn}} = -RT \ln(K_{\text{sp}})$$

$$-170.1 = -(8.314 \times 10^{-3})(298) \ln(K_{\text{sp}})$$

$$\rightarrow \boxed{K_{\text{sp}} = 1.9 \times 10^{12}}$$

11) Question 11 was discarded

12)

Step 1

→ Find $\Delta S_{\text{rxn}}^{\circ}$ + $\Delta H_{\text{rxn}}^{\circ}$

• room temperature = $25^{\circ}\text{C} = 298\text{ K}$

$$\begin{aligned}\Delta S_{\text{rxn}}^{\circ} &= \sum S^{\circ}_{\text{products}} - \sum S^{\circ}_{\text{reactants}} \\ &= S^{\circ}[\text{C}_{\text{graphite}}(\text{s})] - S^{\circ}[\text{C}_{\text{diamond}}(\text{s})] \\ &= 5.69 - 2.43 \\ &= 3.26 \text{ J/mol}\cdot\text{K} \\ &= 3.26 \text{ kJ/mol}\cdot\text{K}\end{aligned}$$

$$\begin{aligned}\Delta H_{\text{rxn}}^{\circ} &= \sum \Delta H_f^{\circ} \text{ products} - \sum \Delta H_f^{\circ} \text{ reactants} \\ &= \Delta H_f^{\circ}[\text{C}_{\text{graphite}}(\text{s})] - \Delta H_f^{\circ}[\text{C}_{\text{diamond}}(\text{s})] \\ &= 0 - 1.88 \\ &= -1.88 \text{ kJ/mol}\end{aligned}$$

Step 2

→ Find $\Delta G_{\text{rxn}}^{\circ}$ and determine if rxn is spontaneous

$$\begin{aligned}\Delta G^{\circ} &= \Delta H^{\circ} - T\Delta S^{\circ} \\ &= -1.88 - (298)(3.26 \times 10^{-3}) \\ &= -2.85 \text{ kJ/mol}\end{aligned}$$

$$\Delta G^{\circ} < 0, \therefore \boxed{\text{rxn is spontaneous}}$$

Step 3

→ find K using $\Delta G^{\circ} = -RT \ln K$

$$\begin{aligned}\Delta G^{\circ} &= -RT \ln(K) \\ -2.85 &= -(8.314 \times 10^{-3})(298) \ln(K)\end{aligned}$$

$$\rightarrow \boxed{K = 3.16}$$

13)

0°C = melting/freezing pt. of water



• rxn is in equilibrium at 0°C , $\therefore \Delta G^{\circ} = 0$

Ice Melting

1 mol solid \rightarrow 1 mol liquid

• entrop increases, $\therefore \Delta S^{\circ}$ is positive

• In order for ice to melt, it has to absorb energy. \therefore the rxn is endothermic and ΔH° is positive

14)



• Ag^+ is reduced

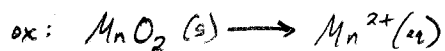
• Cu is oxidized

Helpful Saying: LEO the lion says GER

Lose Electrons = Oxidized

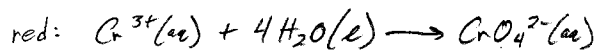
Gain Electrons = Reduced

(5) **Step 1** → write reduction and oxidation reactions separately

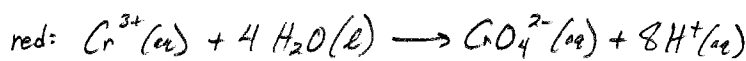


Step 2 → balance each expression separately

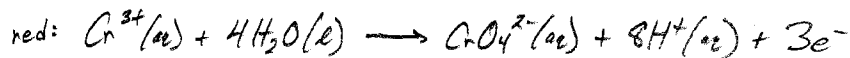
1) • Balance # of Oxygen atoms by adding H_2O



2) • Balance # of Hydrogen Atoms by adding H^+ ions

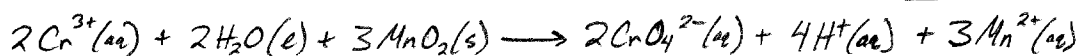
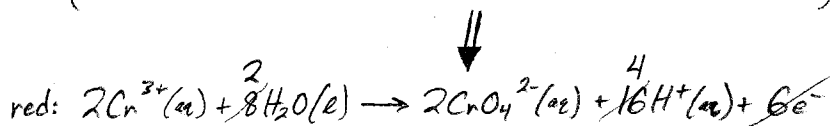


3) • Balance charge by adding e^-



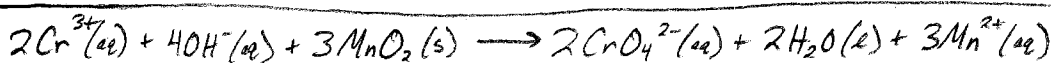
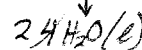
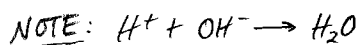
Step 3

→ Balance # of e^- in each rxn and combine them into one expression



Step 4

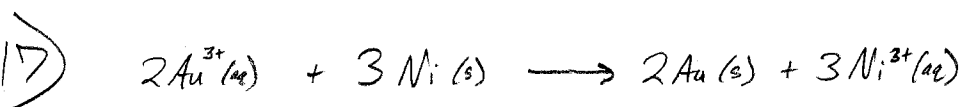
→ neutralize H^+ w/ OH^- ions because we are told the rxn is in a basic solution



(6) Best oxidizing agent = agent that most spontaneously reduces

• To reduce spontaneously, the $E_{red}^{\circ} = (+)$

<u>Reduction of Agent</u>	<u>E_{red}°</u>
$Al(s) + 1e^{-} \rightarrow Al^{-}(aq)$	• Not listed \rightarrow Very nonspontaneous
$Al^{3+}(aq) + 3e^{-} \rightarrow Al(s)$	-1.66
$Fe^{3+}(aq) + 3e^{-} \rightarrow Fe(s)$	-0.04 \leftarrow most positive E_{red}° , $\therefore Fe^{3+}$ is the best oxidizing agent
$Fe^{2+}(aq) + 2e^{-} \rightarrow Fe(s)$	-0.45
$H_2(s) + 2e^{-} \rightarrow 2H^{-}$	• Not listed \rightarrow Very nonspontaneous



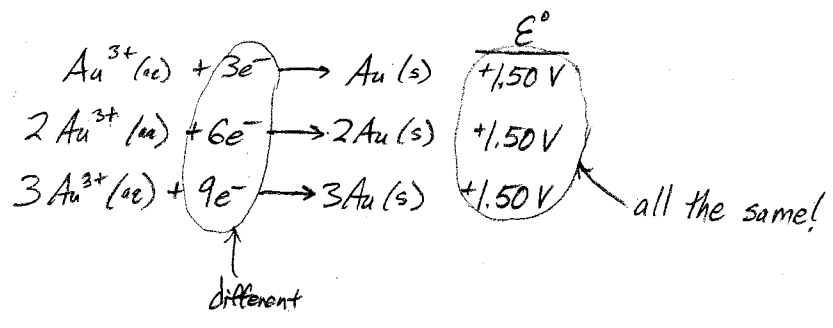
red: $Au^{3+}(aq) + 3e^{-} \rightarrow Au(s) \quad \frac{E^{\circ}}{+1.50V}$

ox: $Ni(s) \rightarrow Ni^{2+}(aq) + 2e^{-} \quad \frac{+0.26V}{}$

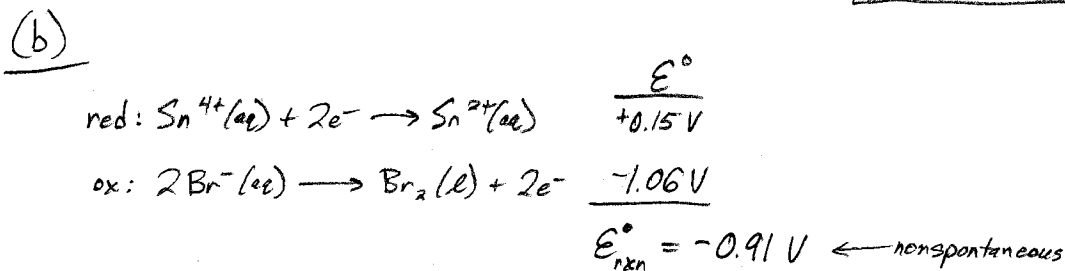
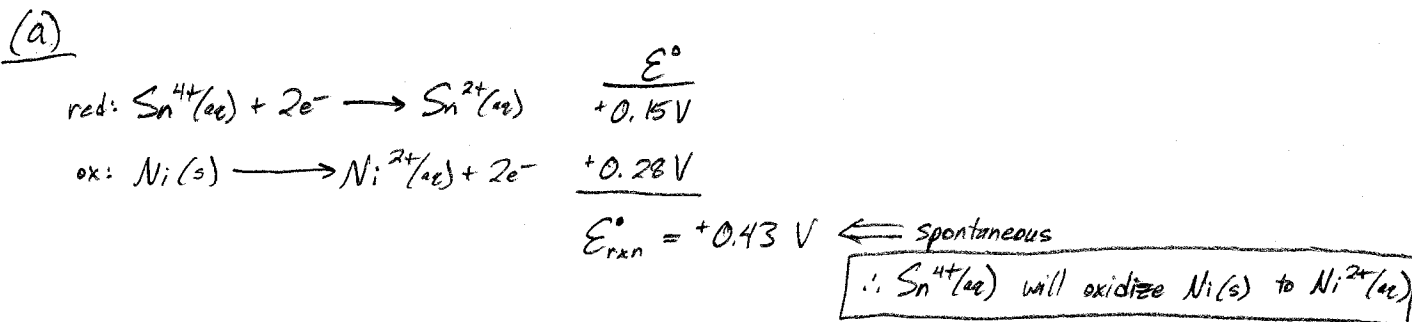
$E_{rxn}^{\circ} = +1.76V$

NOTE:

E° is an intensive property, so E° is always the same no matter how many moles of e^{-} react. For example:



- 18) • Lose Electrons = Oxidized
 • Gain Electrons = Reduced



(c)

- same reaction as (b)

(d)

- Ni^{2+} must oxidize (lose electrons) to reduce $\text{Sn}^{4+}(\text{aq})$
 \hookrightarrow the oxidation of Ni^{2+} is nonspontaneous

(e)

- Sn^{4+} must oxidize (lose electrons) to reduce $\text{Ni}^{2+}(\text{aq})$
 \hookrightarrow the oxidation of Sn^{4+} is nonspontaneous

19) At 25°C, under nonstandard conditions

$$E = E^\circ - \frac{0.0592}{n} \log(Q)$$

$n = \text{moles of } e^-$

Standard
Conditions: concentrations of aqueous ions = 1 M

Step 1

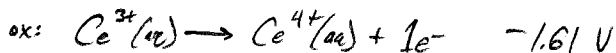
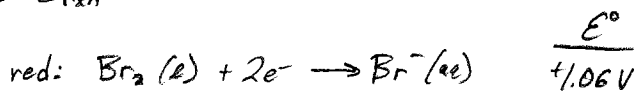
→ Find Q of the rxn



$$Q = \frac{[\text{Br}^-]^2 [\text{Ce}^{4+}]^2}{[\text{Ce}^{3+}]^2} = \frac{[0.04]^2 [0.015]^2}{[0.75]^2} = 6.4 \times 10^{-7}$$

Step 2

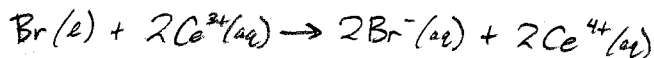
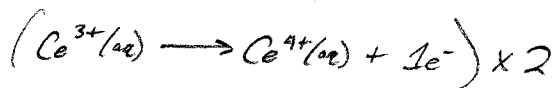
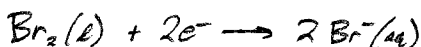
→ Find E°_{rxn}



$$E^\circ_{\text{rxn}} = -0.55 \text{ V}$$

Step 3

→ Find n



$$\rightarrow n = 2 \text{ mol } e^-$$

Step 4

→ solve for E

$$E = E^\circ - \frac{0.0592}{n} \log(Q)$$

$$= (-0.55) - \frac{0.0592}{2} \log(6.4 \times 10^{-7})$$

$$= \boxed{-0.37 \text{ V}}$$

$$20) \quad E = E^\circ - \frac{0.0592}{n} \log Q$$

Step 1

→ write Q expression



$$Q = \frac{[\text{H}_3^{2+}]^3 (P_{\text{NO}})^2}{[\text{NO}_3^-]^2 [\text{H}^+]^8}$$

Step 2

→ Find and compare Q values

- in order to make E most positive, we need $\log(Q)$ as negative as possible
- in order to make $\log(Q)$ as negative as possible, we want Q to be as small as possible.

$$a) \quad Q = \frac{(1)^3 (1)^2}{(1)^2 (1 \times 10^{-9})^8} = 1 \times 10^{56}$$

$$b) \quad Q = \frac{(1.5)^3 (0.75)^2}{(2)^2 (1 \times 10^{-10})^8} = 1.8 \times 10^{78}$$

$$c) \quad Q = \frac{(2)^3 (2)^2}{(5)^2 (1 \times 10^{-10})^8} = 5.1 \times 10^{82}$$

$$d) \quad Q = \frac{(1.5)^3 (0.75)^2}{(2)^2 (1 \times 10^{-9})^8} = 8.8 \times 10^{37}$$

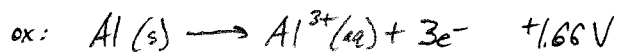
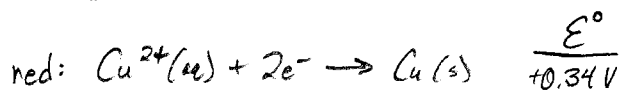
$$e) \quad Q = \frac{(2)^3 (2)^2}{(5)^2 (1 \times 10^{-2})^8} = 1.3 \times 10^{18}$$

Smallest Q , \therefore conditions in answer (e) will produce the most positive cell potential

$$21) \Delta G^\circ = -nFE^\circ$$

Step 1

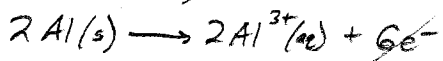
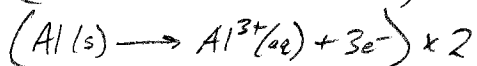
→ Find E_{rxn}°



$$E_{\text{rxn}}^\circ = +2.00\text{V}$$

Step 2

→ Find n



$$n = 6 \text{ mol } e^-$$

Step 3

→ Solve for ΔG°

$$\Delta G^\circ = -nFE^\circ$$

$$= -(6)(96500)(2)$$

$$= -1158000 \text{ J/mol}$$

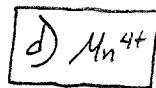
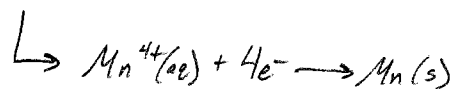
$$= \boxed{-1158 \text{ kJ/mol}}$$

$$23) \text{ Ampere (A)} = \frac{\text{Coulomb (C)}}{\text{second (s)}}$$

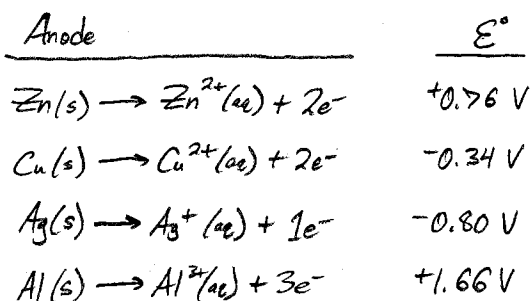
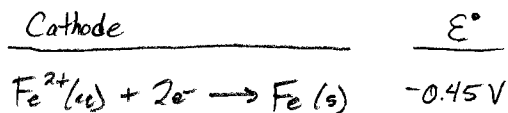


$$\text{g Al} = \frac{2.5\text{C}}{1\text{s}} \times \frac{30\text{min}}{1\text{min}} \times \frac{60\text{s}}{1\text{min}} \times \frac{1\text{mol e}^-}{96485\text{C}} \times \frac{1\text{mol Al}}{3\text{mol e}^-} \times \frac{26.982\text{g}}{1\text{mol Al}} = \boxed{0.419\text{g Al}}$$

$$24) \frac{x\text{ mol e}^-}{1\text{ mol Mn}} = \frac{2.89 \times 10^5\text{C}}{96485\text{C}} \times \frac{1\text{ mol e}^-}{41.2\text{g Mn}} \times \frac{54.938\text{g Mn}}{1\text{ mol Mn}} = \frac{4\text{ mol e}^-}{1\text{ mol Mn}}$$



- 25) • anode = oxidation
• cathode = reduction



Answers

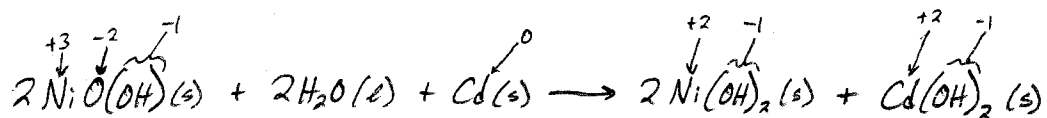
- (a) $\text{Fe}^{2+}(\text{aq}) + \text{Zn}(\text{s}) \rightarrow \text{Fe}(\text{s}) + \text{Zn}^{2+}(\text{aq})$ $E_{\text{cell}}^{\circ} +0.31\text{V} \Leftarrow \text{spontaneous}$
- (b) $\text{Fe}^{2+}(\text{aq}) + \text{Cu}(\text{s}) \rightarrow \text{Fe}(\text{s}) + \text{Cu}^{2+}(\text{aq})$ -0.79 V
- (c) $\text{Fe}^{2+}(\text{aq}) + 2\text{Ag}(\text{s}) \rightarrow \text{Fe}(\text{s}) + 2\text{Ag}^+(\text{aq})$ -1.25 V
- (d) $3\text{Fe}^{2+}(\text{aq}) + 2\text{Al}(\text{s}) \rightarrow 3\text{Fe}(\text{s}) + 2\text{Al}^{3+}(\text{aq})$ +1.21 V $\Leftarrow \text{spontaneous}$

$\boxed{\text{(e) Both a) and d)}}$

26)

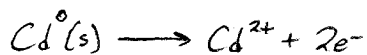
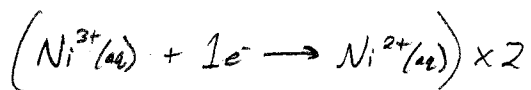
Step 1

→ Determine oxidation states



Step 2

→ Find n



$$\rightarrow n = 2 \text{ mol e}^-$$

Step 3

→ solve for ΔG°

$$\Delta G^\circ = -n F E_{\text{cell}}^\circ$$

$$= -(2) \left(\frac{96485}{1000} \right) (1.2)$$

$$= \boxed{-231.6 \text{ kJ/mol}}$$