Spring 2023 CHM2046 Exam 2 Review

The material covered in this review is from Chapters 16-21, and 23

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Chapter 16: Kinetics

1. Ammonia is generated on an industrial scale using the Haber-Bosch process. The reaction is shown below:

$$N_{2\,(g)} + 3H_{2\,(g)} \rightarrow 2NH_{3\,(g)}$$

Find the rate law, individual, and overall reaction orders and the average value of k for the reaction.

Experiment	Initial rate (mol/L*s)	Initial $[N_2]$ (mol/L)	Initial $[H_2]$ (mol/L)
1	1.9×10^{-12}	0.0113	0.0011
2	1.7×10^{-11}	0.0220	0.0033
3	9.3×10^{-12}	0.0550	0.0011
4	4.9×10^{-11}	0.0220	0.0056

Rate Law: Rate=$k[N_2]^m[H_2]^n$

N2 Order: 1st order

H2 Order: 2nd order

Overall Reaction Order: 3rd order

Average Value of k: 1.39*10^{-6} L/mol*s

$\frac{Rate\,3}{Rate\,1} = \frac{k[N_2]_3^m[H_2]_3^n}{k[N_2]_1^m[H_2]_1^n} \rightarrow \frac{9.3\times10^{-12}}{1.9\times10^{-12}} = \left(\frac{[N_2]_3}{[N_2]_1}\right)^m = \left(\frac{0.0550\,M}{0.0113\,M}\right)^m$

$\hookrightarrow 4.9 = 4.9^m \rightarrow m = 1$ 1st order N_2

$1+2=3$

$\frac{Rate\,4}{Rate\,2} = \frac{k[N_2]_4^m[H_2]_4^n}{k[N_2]_2^m[H_2]_2^n} \rightarrow \frac{4.9\times10^{-11}}{1.7\times10^{-11}} = \left(\frac{0.0056}{0.0033}\right)^n$

$\hookrightarrow 2.9 = (1.7)^n \rightarrow n = 2$ 2nd order H_2

$Rate = k[N_2]^1[H_2]^2$

$k = \frac{Rate}{[N_2][H_2]^2} = \frac{1.9\times10^{-12}}{(0.0113M)(0.0011)^2} = 1.39\times10^{-6}\,\frac{L}{mol\,s}$

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$\ln \frac{[X]_0}{[X]_t} = kt$

2. H_2O_2 decomposes into H_2 and O_2 in a first order reaction. If the initial concentration is 4.38 M, the final concentration is 2.91 M, and the decomposition takes place over 10 minutes, what is k? Using the calculated k, how long will it take to decompose 25% of the initial amount?

a. k=0.035/min; 7 minutes
b. k=0.041/min; 7 minutes
c. k=0.035/min; 10 minutes
d. k=0.041/min; 10 minutes
e. k=0.059/min; 7 minutes
f. k=0.059/min; 10 minutes

$\ln \left(\frac{4.38}{2.91}\right) = k(10\,\text{min}) \rightarrow k = 0.041/\text{min}$

25% decomposed = 1 − .25 = 0.75 → 75% remains

$\ln \frac{[H_2O_2]_0}{[H_2O_2]_t} = kt$

$\ln \frac{(1\,M)}{(0.75\,M)} = (0.041/\text{min})t \rightarrow t = 7\,\text{min}$

3. Which of the following statements are true regarding exothermic reactions?

I. Heat is absorbed Endothermic
II. Heat is released Exothermic
III. Heat is a reactant Endo
IV. Heat is a product EXO
V. Heat and enthalpy will be on the same side of the equation All
VI. Heat and enthalpy will be on opposite sides of the equation Wrong
VII. The energy of the reactants is higher than the products EXO
VIII. The energy of the reactants is lower than the products Endo

a. I, III, V, VII
b. II, IV, V, VII
c. I, II, III, IV
d. V, VI, VII, VIII
e. I, IV, V, VIII
f. II, III, V, VII

4. Which of the following statements are true regarding endothermic reactions?

I. Heat is absorbed
II. Heat is released
III. Heat is a reactant
IV. Heat is a product
V. Heat and enthalpy will be on the same side of the equation
VI. Heat and enthalpy will be on opposite sides of the equation
VII. The energy of the reactants is higher than the products
VIII. The energy of the reactants is lower than the products

a. I, III, V, VIII
b. II, IV, VI, VII
c. I, II, III, IV
d. V, VI, VII, VIII
e. I, IV, V, VIII
f. II, III, V, VII

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5. Which of the following statements are true regarding catalysts?

I. Catalysts cause products to form slower

II. Catalysts cause products to form faster

III. Catalysts increase activation energy

IV. Catalysts are not reformed

V. Catalysts lower activation energy

VI. Catalysis are reformed

VII. Catalysts affect reaction rate; it increases

VIII. Catalysts affect reaction rate; it decreases

IX. Catalysts don't affect reaction rate

a. I, III, VI, IX
b. II, III, VI, IX
c. II, V, VI, VII
d. I, IV, VI, VIII
e. II, IV, VI, VII

$K_p = K_c(RT)^{\Delta n_{gas}}$

Chapter 17: Equilibrium

1. Given the following chemical reaction, calculate the K_p given that the K_c is 0.28 at 900°C.

$$CS_{2\,(g)} + 4H_{2\,(g)} \leftrightarrow CH_{4\,(g)} + 2H_2S_{\,(g)}$$

a. 7.5×10^{-5}
b. 8.1×10^{-2}
c. 3.6×10^{-3}
d. 3.0×10^{-5}
e. 2.9×10^{-4}

$K_c = 0.28$
$T = 900°C + 273.15K$
$R = 0.0821 \frac{L\,atm}{mol\,K}$
$\Delta n_{gas} = 3 - 5 = -2$

$K_p = (0.28)\left((1173.15)(0.0821)\right)^{-2} = 3.0 \times 10^{-5}$

2. Which of the following statements regarding Q and K are true?

I. If K>Q, then the reaction proceeds to the right

II. If K=Q, then the reaction is at equilibrium

III. If the reaction proceeds to the right, it will create more products

IV. If the reaction proceeds to the left, it will create more products

V. If K>Q, then the reaction proceeds to the left

VI. If K<Q, then the reaction proceeds to the left

VII. If K=Q, then the reaction proceeds to the right

VIII. If K<Q, then the reaction is at equilibrium

a. II, III, V, VIII
b. I, II, III, VI
c. IV, V, VI, VII
d. VI, VII, VIII
e. I, III, VI, VIII

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3. Fill in the table summarizing the effects of Le Chatelier's Principle.

Change	Effect on Equilibrium (Left or Right)	Effect on the value of K (Equilibrium Constant)
Increase [reactant]	Right (products)	None
Increase [product]	Left (reactants)	None
Decrease [reactant]	Left (reactants)	None
Decrease [product]	Right (products)	None
Increase pressure	Towards side with fewer moles of gas	None
Increase volume	Towards side with more moles of gas	None
Decrease pressure	Toward side with more moles of gas	None
Decrease volume	Towards side with fewer moles of gas	None
Increase pressure (inert gas)	No change in volume, no change; concentrations unchanged	None
Increase temperature	Towards absorption of heat (Endothermic shift right) (Exothermic shift left)	Endothermic, increases Exothermic, decreases
Decrease temperature	Towards release of heat (Endothermic shift left) (Exothermic shift right)	Endothermic, decreases Exothermic, increases
Add catalyst	None; forward and reverse rates increase equally	None

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Chapter 18: Acid-Base Equilibria

1. Which of the following statements regarding acids, bases, and Kas is true?

I. The stronger the acid, the larger the Ka, the larger the pKa
II. The stronger the acid, the larger the Ka, the smaller the pKa
III. The weaker the acid, the lower the concentration of H3O+, the larger the pKa
IV. The larger the pKa, the smaller the Ka
V. A strong acid is a weak base
VI. Kw, Ka, and Kb are related to each other in the equation Kw=Ka*Kb
VII. The equilibrium of an acid base reaction goes from the stronger acid to the weaker acid
VIII. The equilibrium of an acid base reaction goes from the weaker acid to the stronger acid
IX. If the reaction proceeds to the right, Kc>1.

a. I, VIII
b. All but I, VIII
c. II, III, VII, VIII
d. IV, V, VII, VIII
e. I, III, IV, VII

2.Which of the following statements regarding pH is true?

I. Acidic solutions have a higher concentration of OH^-

II. Basic solutions have a higher concentration of OH^-

III. A neutral solution has an equal concentration of H_3O^+ and OH^-

IV. $Kw=\frac{[H_3O^+]}{[OH^-]}$

V. $Kw=[H_3O^+]*[OH^-]$

VI. pH+pOH=14

VII. pH-pOH=14

a. I, II, IV, VII
b. II, III, V, VI
c. III, IV, V, VII
d. II, IV, VI

3. If an unknown weak acid is 0.798% dissociated in a 2.15M solution. What is the Ka of the acid, the pKa, and the identity of the acid?

a. $2.46*10^{-3}$, 10.5, Formic acid
b. $5.12*10^{-5}$, 3.14, Lactic acid
c. $1.38*10^{-4}$, 3.86, Lactic acid
d. $9.17*10^{-4}$, 4.68, Formic acid

$0.798\% = \frac{x}{2.15M} \times 100\% \rightarrow x = 0.01715M$

$K_a = \frac{(x)(x)}{2.15-x} = \frac{x^2}{2.15-x} \rightarrow \frac{(0.01715)^2}{2.15-0.01715} = 1.38\times10^{-4}$

$pK_a = -\log K_a = -\log(1.38\times10^{-4}) = 3.86$

Lactic Acid

$HA + H_2O \rightleftharpoons A^- + H_3O^+$

	HA	H_2O	A^-	H_3O^+
I	2.15		0	0
C	−x		+x	+x
E	2.15−x		x	x

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4. What are the equilibrium values of carbonic acid and the pH of a 1.34M solution?

$[H_2CO_3] = 1.34M$

$[HCO_3^-] = 0.00077M$

$[CO_3^{2-}] = 4.7*10^{-11}M$

pH = 3.11

$\frac{[H_2CO_3]_{init}}{K_{a_1}} = \frac{1.34\,M}{4.5\times10^{-7}} = 2.97\times10^6 >>> 400 \therefore 1.34 - x \approx 1.34\,M$

$H_2CO_3 + H_2O \rightleftharpoons HCO_3^- + H_3O^+$

I: 1.34 | 0 | 0

C: −x | +x | +x

E: 1.34 − x | x | x

$X = [H_3O^+] = [HCO_3^-] = 0.00077M$

$K_{a_1} = \frac{[H_3O^+][HCO_3^-]}{[H_2CO_3]} = \frac{x^2}{1.34} \rightarrow x = 0.00077\,M$

$pH = -\log[H_3O^+] = -\log[0.00077] = 3.11$

$HCO_3^- + H_2O \rightleftharpoons H_3O^+ + CO_3^{2-}$

I: 0.00077 | 0.00077 | 0

C: −x | +x | +x

E: 0.00077 − x | 0.00077 | x

$K_{a_2} = \frac{[CO_3^{2-}][H_3O^+]}{[HCO_3^-]} \rightarrow [CO_3^{2-}] = \frac{K_{a_2}[HCO_3^-]}{[H_3O^+]} = \frac{(4.7\times10^{-11})(0.00077)}{(0.00077)} =$

$[CO_3^{2-}] = 4.7\times10^{-11}$

5. Which salts yield neutral solutions?

a. NH_4Cl

b. $CaCl_2$

c. $LiNO_3$

d. $Fe(NO_3)_3$

e. B and C

f. A and D

g. None of the above

↳ anions of strong hydrohalic acids: Cl^-, Br^-, I^-

Strong oxoacid ions: NO_3^-, ClO_4^-

Group 1A ions

Group 2A ions $Ca^{2+}, Sr^{2+}, Ba^{2+}$

6. Which of the following act as Lewis acids?

a. Ba^{2+}

b. NH_3 → Lewis Base $:NH_3$ ↳ electron pair acceptor

c. $AlCl_3$

d. H_2O → Lewis Base $H\ddot{O}H$

e. A and C

f. B and D

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Chapter 19: Ionic Equilibria in Aqueous Systems

1. What is the pH of a buffer of 0.83M $(CH_3)_2NH_2Cl$ and 1.2M $(CH_3)_2NH$ before and after adding 125mL of 0.75M HCl to 1 L of the buffer. (Info: pKb of $(CH_3)_2NH$=3.23).

a. 9.776 -> 10.93
b. 2.726 -> 7.901
c. 10.93 -> 10.85
d. 7.901 ->2.726

Before

$K_b\ (CH_3)_2NH = 10^{-pK_b} = 10^{-3.23} = 5.9\times10^{-4}$

$K_a = \frac{K_w}{K_b} = \frac{10^{-14}}{5.9\times10^{-4}} = 1.7\times10^{-11} = K_a$ of $(CH_3)_2NH_2^+$

$[(CH_3)_2NH_2^+] = 0.83M$

$[(CH_3)_2NH] = 1.2M$

$[H_3O^+] = K_a\frac{[Acid]}{[Base]} = (1.7\times10^{-11})\left(\frac{0.83}{1.2}\right) = 1.176\times10^{-11}$

$pH = -\log[H_3O^+] = -\log[1.176\times10^{-11}] = 10.93$

After adding 125mL of 0.75M HCl to 1L buffer

$[(CH_3)_2NH_2^+]_{mix} = \frac{M_aV_a}{M_{dilute}} = \frac{(0.83M)(1L)}{(1L+0.125mL)} = 0.737777M$

$[(CH_3)_2NH] = \frac{(1.2M)(1L)}{(1.125L)} = 1.0666M$

$[H_3O^+]_{added} = \frac{(.75M)(0.125L)}{(1.125L)} = 0.0833M$

$(CH_3)_2NH + H_3O^+ \rightleftharpoons (CH_3)_2NH_2^+ + H_2O$

	$(CH_3)_2NH$	H_3O^+	$(CH_3)_2NH_2^+$	H_2O
I	1.0666	0.0833	0.737777	
C	−0.0833	−0.0833	+0.0833	
E	0.9833	0	0.8211	

$[H_3O^+] = K_a\frac{Acid}{Base} = (1.7\times10^{-11})\left(\frac{0.8211}{0.9833}\right) = 1.42\times10^{-11}$

$pH = -\log[H_3O^+] = 10.85$

2. Given a pH of 10.73, what is the ratio of a buffer of $[NH_3]/[NH_4^+]$, and using that what masss of NH_4Cl would need to be added to 2438 mL of 0.56M NH_3 to make the buffer? (Kb of NH_3=1.76*10^{-5})

a. 1.48; 97.2 g
b. 1.48; 49.3 g
c. 7.51; 97.2 g
d. 7.51; 49.3 g
e. 8.61; 52.3 g

$K_a\ NH_4^+ = \frac{K_w}{K_b NH_3} = \frac{10^{-14}}{1.76\times10^{-5}} = 5.68\times10^{-10}$

$pK_a = -\log K_a = 9.25$

$pH = pK_a + \log\frac{[NH_3]}{[NH_4^+]}$

$10.73 = 9.25 + \log\frac{[NH_3]}{[NH_4^+]}$ → ratio = 1.48

Mass NH_4Cl: $(2.438L)\left(\frac{0.56mol}{1L}\right)\left(\frac{1mol\ NH_4Cl}{1.48mol\ NH_3}\right)\left(\frac{53.49g\ NH_4Cl}{1mol\ NH_4Cl}\right) = 49.3g$

3. Magnesium phosphate is an anticaking agent for silicone-containing cleaning agents and salt. Its K_{sp} is 1.04*10^{-24}. If $[Mg^{2+}]=[PO_4^{3-}]=3.6*10^{-10}$ M, will magnesium phosphate precipitate?

a. Yes, Qsp>Ksp
b. No, Qsp>Ksp
c. No, Qsp=Ksp
d. Yes, Qsp<Ksp
e. No, Qsp<Ksp

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4. Does the addition of HNO_3 affect the solubility of calcium fluoride?

a. Increases solubility
b. Decreases solubility
c. No effect on solubility

Strong acids increase solubility

5. What is the pH at the equivalence point of 912 mL of 10.67 M HBrO with 15.02 M NaOH?

a. 12.84
b. 13.74
c. 2.29
d. 11.71
e. 6.91

Initial $HBrO = (0.912L)\left(\frac{10.67\ mol\ HBrO}{1L}\right) = 9.73104\ mol\ HBrO$

Equivalence Point → Same mols of NaOH → 9.73104 mol NaOH

Volume (L) NaOH added

$(9.73104\ mol)\left(\frac{1L}{15.02\ mol\ NaOH}\right) = 0.647787L$

$[BrO^-] = \frac{9.73104\ mol}{0.912L + 0.647787L} = 6.2384M$

$K_b\ BrO^- = \frac{K_w}{K_a\ HBrO} = \frac{1\times10^{-14}}{2.3\times10^{-9}} = 4.3\times10^{-6}$

$BrO^- + H_2O \rightleftharpoons HBrO + OH^-$

	6.2384	0	0
	−x	+x	+x
	6.2384−x	x	x

$K_b = \frac{[HBrO][OH^-]}{[BrO^-]} = 4.3\times10^{-6} = \frac{x^2}{6.2384}$

$x = [OH^-] = 0.005171M$

$pOH = -\log[OH^-] = -\log[0.005171M]$

$pOH = 2.2865$

$pH = 14 - pOH = 14 - 2.2865 = 11.71$

$pH = 11.71$

Chapter 20: Thermodynamics

1. Fill in the following table on the spontaneity of ΔG, ΔH, ΔS

ΔG=ΔH-TΔS					
ΔG	**ΔH**	**-TΔS**	**ΔS**	**Spontaneity**	**Exothermic or Endothermic**
-	-	-	+	Spontaneous at all temperatures (T)	Exothermic
+	+	+	-	Nonspontaneous at all T	Endothermic
+ or -	+	-	+	Spontaneous at high T (ΔG<0), Nonspontaneous at low T (ΔG>0)	Endothermic
+ or -	-	+	-	Spontaneous at low T (ΔG<0), Nonspontaneous at high T (ΔG>0)	Exothermic

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2. Find K at 298 K of the following reaction:

$$2HCl_{(g)} + Br_{2\,(l)} \leftrightarrow 2HBr_{(g)} + Cl_{2\,(g)}$$

a. $8.41*10^{-51}$
b. $5.89*10^{-28}$
c. $1.64*10^{-11}$
d. $9.37*10^{-15}$

$\Delta G°_{rxn} = [(2mol\ HBr)(-53.5\ kJ/mol) + (1mol\ Cl_2)(0\ kJ/mol)] - [(2mol\ HCl)(-131.17\ kJ/mol) + (1mol\ Br_2)(0\ kJ)]$

$\Delta G°_{rxn} = 155.34\ kJ$

$\Delta G = -RT\ln K$

$K = e^{\frac{\Delta G}{-RT}} = e^{\frac{(155.34\ kJ)(\frac{1000\ J}{1\ kJ})}{-(8.314\ J/mol\,K)(298\ K)}} = 5.99 \times 10^{-28}$

3. FeO(s) oxidizes into Fe_2O_3 (s). If there is 1 mole of Fe_2O_3, does this reaction occur spontaneously at 298K? Given the values of ΔH_{rxn}, ΔS_{rxn}, and ΔS_{univ}.

a. It is spontaneous; ΔH_{rxn}=-560.7 kJ, ΔS_{rxn}=-273.44 J/K and ΔS_{univ}=+1608 J/K
b. It is not spontaneous; ΔH_{rxn}=-560.7 kJ, ΔS_{rxn}=-68.44 J/K and ΔS_{univ}=+1813 J/K
c. It is spontaneous; ΔH_{rxn}=+560.7 kJ, ΔS_{rxn}=+68.44 J/K and ΔS_{univ}=-1813 J/K
d. It is not spontaneous; ΔH_{rxn}=+560.7 kJ, ΔS_{rxn}=+68.44 J/K and ΔS_{univ}=-1813 J/K

$FeO + O_2 \rightarrow Fe_2O_3$

$4FeO_{(s)} + O_{2(g)} \rightarrow 2Fe_2O_{3(s)}$

$\Delta H°_{rxn} = [(2mol\ Fe_2O_3)(-824.248\ kJ/mol)] - [(4mol\ FeO)(-271.96\ kJ/mol) + (1mol\ O_2)(0)] = -560.7\ kJ$

$\Delta S_{surr} = \frac{-\Delta H_{rxn}}{T} = \frac{-(-560.7\ kJ)(\frac{1000\ J}{1\ kJ})}{298\ K} = 1881.4\ J/K$

$\Delta S°_{rxn} = [(2mol\ Fe_2O_3)(87.28\ J/mol\,K)] - [(4mol\ FeO)(60.75\ J/mol\,K) + (1mol\ O_2)(205\ J/mol\,K)] = -273.44\ J/K$

$\Delta S_{univ} = \Delta S_{surr} + \Delta S°_{rxn} = 1881.4\ J/K + (-273.44\ J/K) = +1608\ J/K$

Because ΔS_{univ} is positive, the reaction is spontaneous

4. Calculate the $\Delta G°_{rxn}$ at 25°C of the following reaction:

$$4NH_{3\,(g)} + 5O_{2\,(g)} \leftrightarrow 4NO_{(g)} + 6H_2O_{(g)}$$

a. +961 kJ
b. +852 kJ
c. -961 kJ
d. -852 kJ
e. +134 kJ
f. -134 kJ

$\Delta G°_{rxn} = [(4mol\ NO)(86.60\ kJ/mol) + (6mol\ H_2O)(-228.6\ kJ/mol)] - [(4mol\ NH_3)(-16\ kJ/mol) + (5mol\ O_2)(0\ kJ/mol)]$

$= -961\ kJ$

5. Calculate the standard entropy of the combustion of butane. C_4H_{10}

a. -876 J/K
b. +876 J/K
c. -876 kJ
d. +876 kJ

$2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$

$\Delta S°_{rxn} = \sum n\Delta S°_f\ Products - \sum n\Delta S°_f\ Reactants$

$= [(8mol\ CO_2)(213.7\ J/mol\,K) + (10mol\ H_2O)(69.90\ J/mol\,K)] - [(2mol\ C_4H_{10})(310\ J/mol\,K) + (13mol\ O_2)(205\ J/mol\,K)]$

$= -876\ J/K$

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Chapter 21: Electrochemistry

1. Given the following reaction, what is the E° value for vanadium, given that E°_{cell} is 0.62 V and E° of nitrate is 0.96 V?

$$3V^{3+}_{(aq)} + NO_3^-{}_{(aq)} + H_2O_{(l)} \rightarrow 3VO^{2+}_{(aq)} + NO_{(g)} + 2H^+_{(aq)}$$

a. -0.34V
b. +0.67V
c. -0.97V
d. **+0.34V**
e. +0.97V

$V^{3+} + H_2O \rightarrow VO^{2+} + 2H^+ + e^-$ oxidation → anode

$NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$ reduction → cathode

or $V^{3+} \rightarrow VO^{2+}$ ($+3 \rightarrow +2$)

$3+ \rightarrow 4+ \; -1e^-$ Oxidation Anode

$NO_3^- \rightarrow NO$ (+5 → +2, $+3e^-$) Reduced Cathode

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

$0.62V = 0.96V - E^\circ_{anode}$

$E^\circ_{vanadium} = 0.96V - 0.62V = +0.34V$

2. Is the following reaction spontaneous? (Hint: Use the E°_{cell} value)

$$Cl_{2\,(g)} + Co^{2+}_{(aq)} \rightarrow Cl^-_{(aq)} + Co^{3+}_{(aq)}$$

a. It is spontaneous; E°_{cell} =+0.46V
b. It is spontaneous; E°_{cell} =+0.92V
c. **It is not spontaneous; E°_{cell} =-0.46V**
d. It is not spontaneous; E°_{cell} =-0.92V

$E^\circ_{cell} = 1.36V - 1.82V = -.46V$ Nonspontaneous

$Cl_2 \rightarrow 1.36V$

$Co^{2+} \rightarrow 1.82V$

0 → −1 reduction Cathode

2+ → 3+ oxidation Anode

3. Is the following reaction spontaneous? If it is, what would the nonspontaneous form look like and calculate E°_{cell} of the nonspontaneous reaction.

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

a. $Fe(s) + 2Fe^{3+}(aq) \rightarrow 3Fe^{2+}(aq)$; E°_{cell}=+1.21V
b. $3Fe^{2+}(aq) \rightarrow Fe(s) + 2Fe^{3+}(aq)$; E°_{cell}=+1.21V
c. $Fe(s) + 2Fe^{3+}(aq) \rightarrow 3Fe^{2+}(aq)$; E°_{cell} =-1.21 V
d. **$3Fe^{2+}(aq) \rightarrow Fe(s) + 2Fe^{3+}(aq)$; E°_{cell}=-1.21V**

$2(Fe^{3+}_{(aq)} + e^- \rightarrow Fe^{2+})$ 0.77V Cathode

$Fe(s) \rightarrow Fe^{2+} + 2e^-$ −0.44V Anode

$E^\circ_{cell} = 0.77V - (-0.44V) = +1.21V$ Spontaneous

Nonspontaneous: flip reaction & sign

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

$E^\circ_{cell,rev} = -0.44V - (0.77V) = -1.21V$

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4. What is the cell potential for the following concentration cell? (Hint: Use the Nernst Equation)

$$Zn\ (s)\ |\ Zn^{2+}\ (aq,\ 1.5\ M)\ ||\ Zn^{2+}\ (aq,\ 6\ M)\ |\ Zn\ (s)$$

a. +0.021V
b. +0.018V
c. -0.021V
d. +0.97V
e. -1.64V

Anode: $Zn(s) \rightarrow Zn^{2+}(aq, 1.5M) + 2e^-$ $n = 2$ $E^\circ_{anode} = -0.7618V$

Cathode: $Zn^{2+}(aq, 6M) + 2e^- \rightarrow Zn(s)$ $E^\circ_{cathode} = -0.7618V$

Overall: $Zn^{2+}(aq, 6M) \rightarrow Zn^{2+}(aq, 1.5M)$ $E^\circ_{cell} = 0V$

$$E_{cell} = E^\circ_{rxn} - \frac{0.0592V}{n} \log Q$$

$$E_{cell} = 0V - \frac{0.0592V}{2} \log \frac{[1.5M]}{[6M]} = +0.018V$$

5. Given the following reaction, the ΔG° = -143 kJ/mol of reaction. What is the K at 25°C, and E°_{cell}?

$Cd_{(s)} + Cu^{2+}_{(aq)} \rightarrow Cd^{2+}_{(aq)} + Cu_{(s)}$ (−2e⁻ oxidation; +2e⁻ reduction)

a. $K=1.17*10^{25}$; $E^\circ_{cell}=0.74V$
b. $K=1.25*10^{-25}$; $E^\circ_{cell}=-0.58V$
c. $K=5.51*10^{15}$; $E^\circ_{cell}=0.76V$
d. $K=8.25*10^{25}$; $E^\circ_{cell}=0.98V$
e. $K=1.81*10^{25}$; $E^\circ_{cell}=-1.67V$

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

$= 0.34V - (-0.40V)$

$= 0.74V$

$\Delta G = -RT \ln K$

$\frac{\Delta G}{-RT} = \frac{-RT \ln K}{-RT}$

$\ln K = \frac{-\Delta G}{RT}$

$e^{\ln K} = e^{\frac{-\Delta G}{RT}}$

$K = e^{\left(\frac{-\Delta G}{RT}\right)}$

$K = e^{\left(\frac{-(-143\ kJ/mol)\left(\frac{1000\ J}{1\ kJ}\right)}{(8.314\ J/mol\cdot K)(298K)}\right)} = 1.17 \times 10^{25}$

6. Fill the following table on the comparison of voltaic and electrolytic cells

				Electrode		
Cell Type	**ΔG**	**E_{cell}**	**Spontaneity**	**Name**	**Process**	**Sign**
Voltaic	<0	>0	Spontaneous	Anode	Oxidation	-
				Cathode	Reduction	+
Electrolytic	>0	<0	Nonspontaneous	Anode	Oxidation	+
				Cathode	Reduction	-

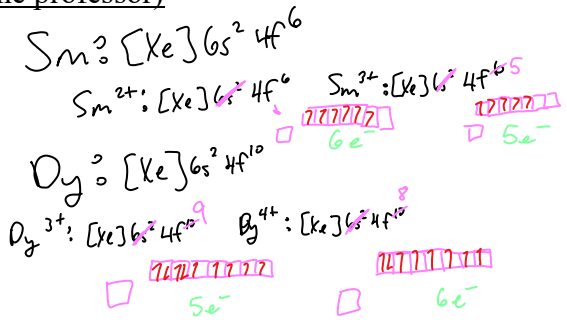
Different professors cover different material

Chapter 23: Transition Elements (This has only been covered by one professor)

1. How many unpaired electrons are Sm^{2+}, Sm^{3+}, Dy^{3+}, and Dy^{4+}?

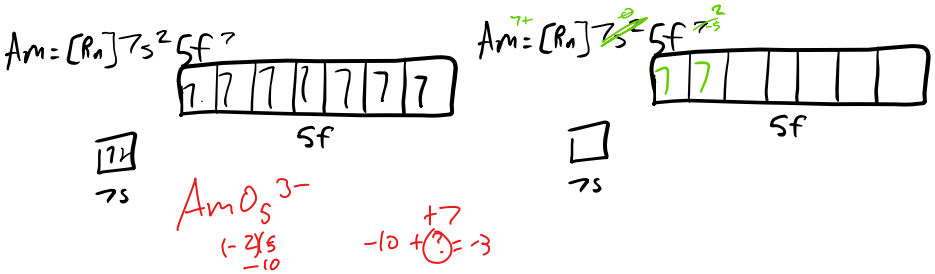
a. 6, 5, 5, 6
b. 6, 6, 6, 6
c. 5, 5, 5, 5
d. 5, 6, 6, 5
e. 5, 7, 4, 6
f. 7, 4, 4, 5

Sm^{2+} 6 unpaired
Sm^{3+} 5 unpaired
Dy^{3+} 5 unpaired
Dy^{4+} 6 unpaired



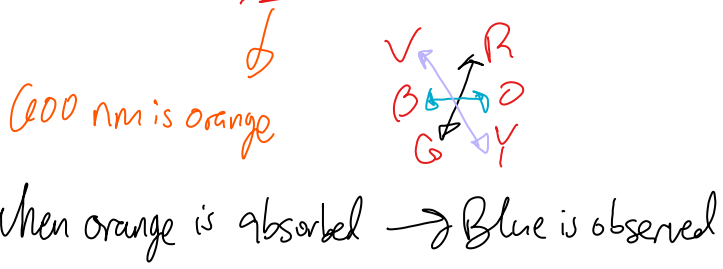
2. AmO_5^{3-} has a green color when in an aqueous solution. What is the oxidation state of Am in this molecule, how many unpaired electrons does it have, and what does its electron configuration look like?

a. 3+, 5 unpaired
b. 2+, 7 unpaired
c. 6+, 3 unpaired
d. 7+, 2 unpaired
e. 5+, 5 unpaired



3. If an absorbed color has a wavelength of 600 nm, what is the observed color?

a. Blue
b. Green
c. Orange
d. Violet
e. Red
f. Yellow



4. What is the coordination number and charge of the central metal ion in $[Co(en)_2Br_2]NO_3$?

a. 4, 2+
b. 4, 4+
c. 6, 3+
d. 5, 5+
e. 6, 6+

5. How many unpaired electrons would you expect for $[CoF_6]^{3-}$, and give its energy diagram. Is it high spin or low spin?

a. 6, low spin
b. 4, high spin
c. 6, high spin
d. 3, low spin