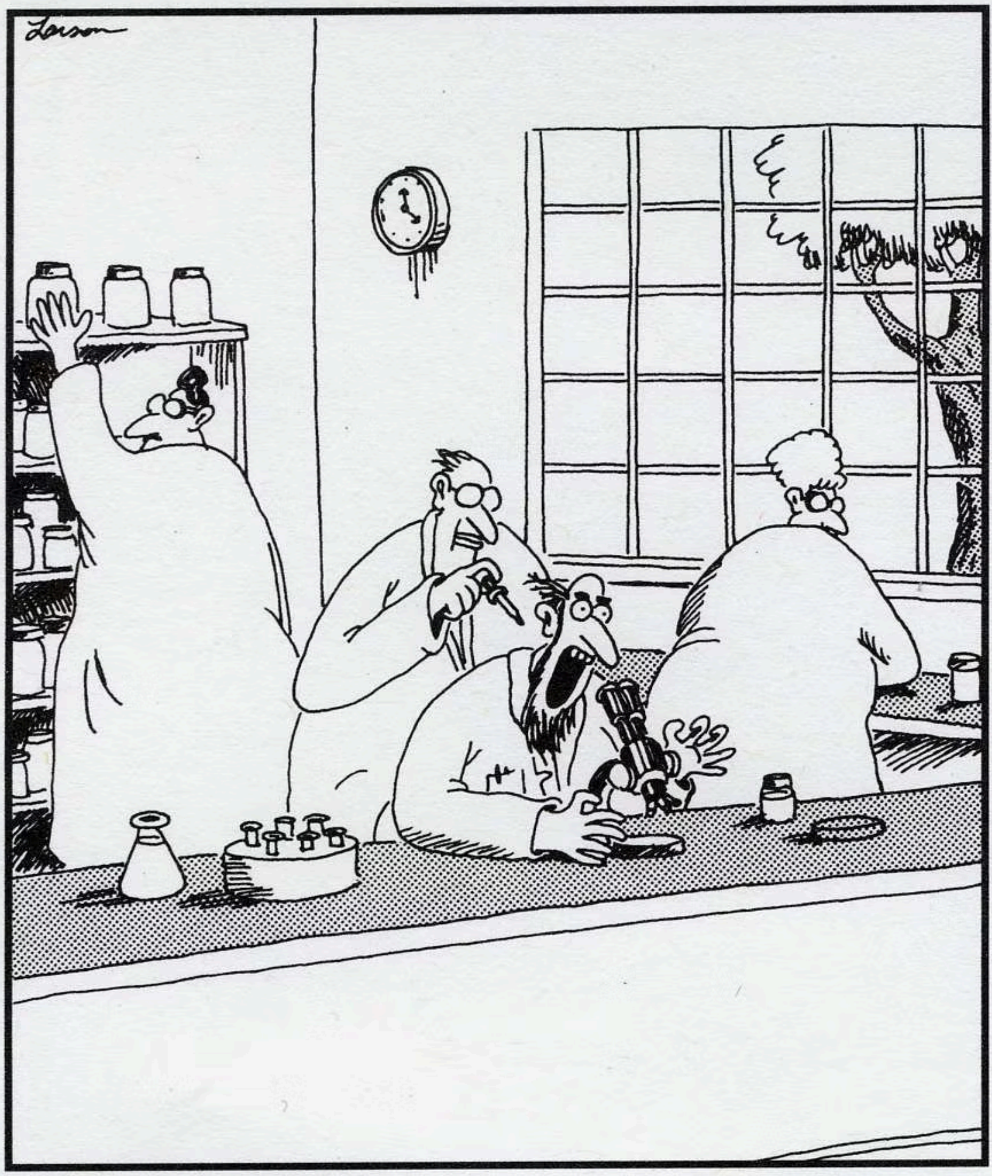
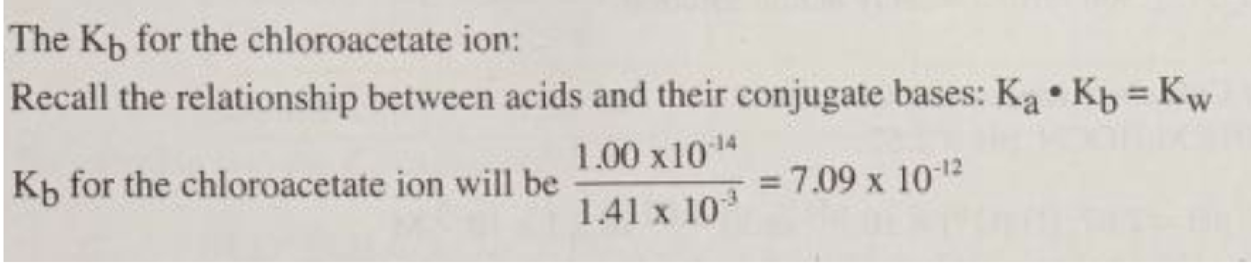
Exam 2 Review R and R Worksheet
7/31/24



Professor Glickman, the lab practical joker, deftly places a single drop of hydrochloric acid on the back of Professor Bingham's neck.

Acids and Bases 🍋

1) Chloroacetic acid ($ClCH_2CO_2H$) has $K_a = 1.41 \times 10^{-3}$. What is the value of K_b for the chloroacetate ion ($ClCH_2CO_2^-$)?



The K_b for the chloroacetate ion:

Recall the relationship between acids and their conjugate bases: $K_a \bullet K_b = K_w$

K_b for the chloroacetate ion will be $\dfrac{1.00 \text{ x} 10^{-14}}{1.41 \text{ x } 10^{-3}} = 7.09 \text{ x } 10^{-12}$

2) For each of the following reactions, predict whether the equilibrium lies predominantly to the left or to the right. Explain your predictions briefly.

(a) $CH_3CO_2H(aq) + Br^+(aq) \Leftrightarrow CH_3CO_2^-(aq) + HBr(aq)$

The equilibrium will favor the weaker acid which can be determined by comparing pk_a of the acids on each side of the reaction. On the left side, the acid is acetic acid, which has a pk_a of 4.7. On the right side, the acid is hydrobromic acid of -8.7. Therefore, the equilibrium will **favor the left side.**

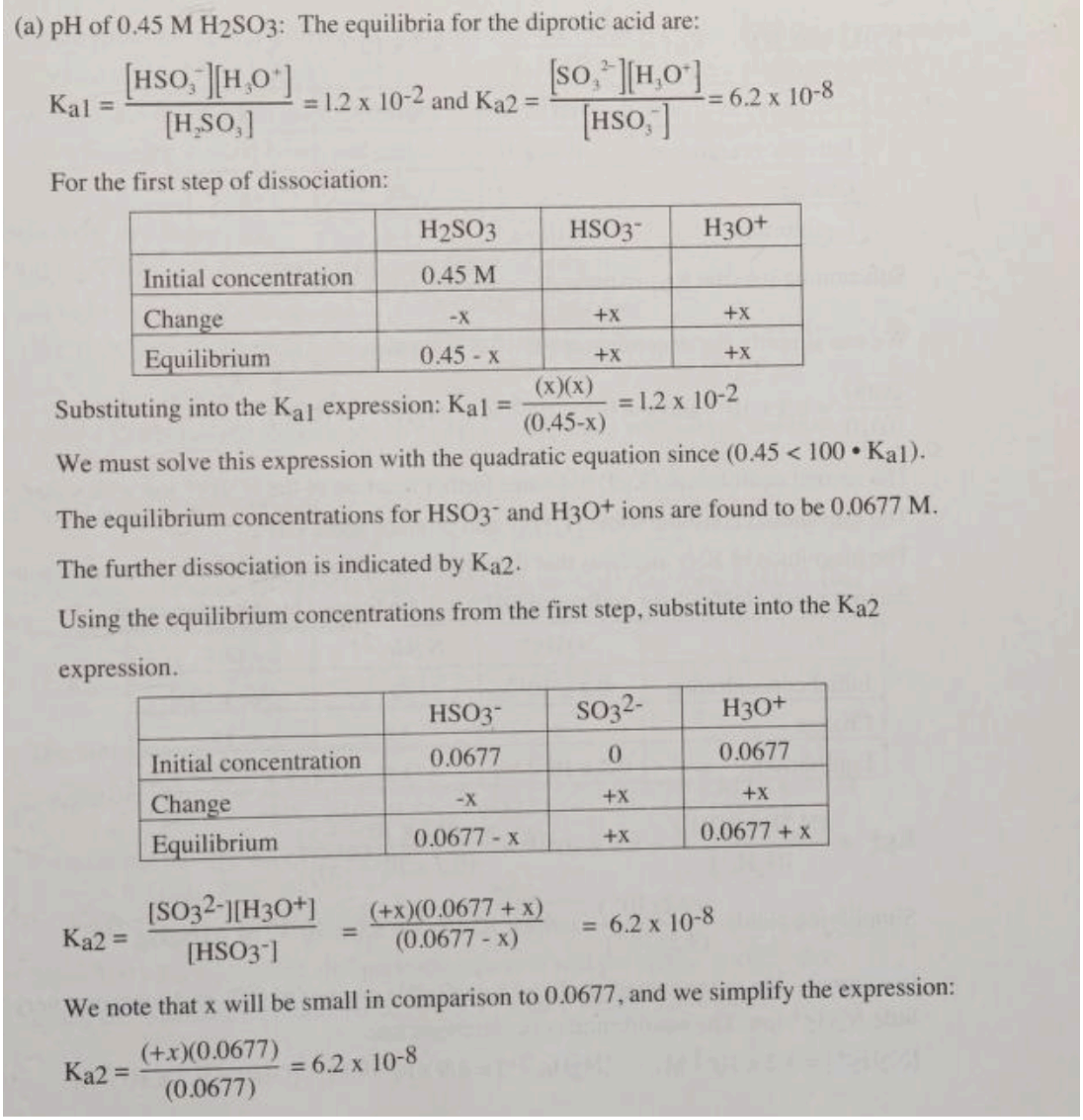
(b) $H_3PO_4(aq) + F^+(aq) \Leftrightarrow H_2PO_4^-(aq) + HF(aq)$

On the left side, the acid is phosphoric acid, a polyprotic acid. The most acidic proton on phosphoric acid has a pk_a of 2.1. On the right side, the acid is hydrofluoric acid, which has a pk_a of 3.2. Therefore, the equilibrium will **favor the right side.**

3) Sulfurous acid, H_2SO_3, is a weak acid capable of providing two H^+ ions.

(a) What is the pH of a 0.45 M solution of H_2SO_3?

(b) What is the equilibrium concentration of the sulfite ion, SO_3^{2-}, in the 0.45 M solution of H_2SO_3?



(a) pH of 0.45 M H_2SO_3: The equilibria for the diprotic acid are:

$$K_{a1} = \frac{[HSO_3^-][H_3O^+]}{[H_2SO_3]} = 1.2 \times 10^{-2} \text{ and } K_{a2} = \frac{[SO_3^{2-}][H_3O^+]}{[HSO_3^-]} = 6.2 \times 10^{-8}$$

For the first step of dissociation:

	H_2SO_3	HSO_3^-	H_3O^+
Initial concentration	0.45 M		
Change	-x	+x	+x
Equilibrium	0.45 - x	+x	+x

Substituting into the K_{a1} expression: $K_{a1} = \frac{(x)(x)}{(0.45\text{-}x)} = 1.2 \times 10^{-2}$

We must solve this expression with the quadratic equation since ($0.45 < 100 \bullet K_{a1}$).

The equilibrium concentrations for HSO_3^- and H_3O^+ ions are found to be 0.0677 M.

The further dissociation is indicated by K_{a2}.

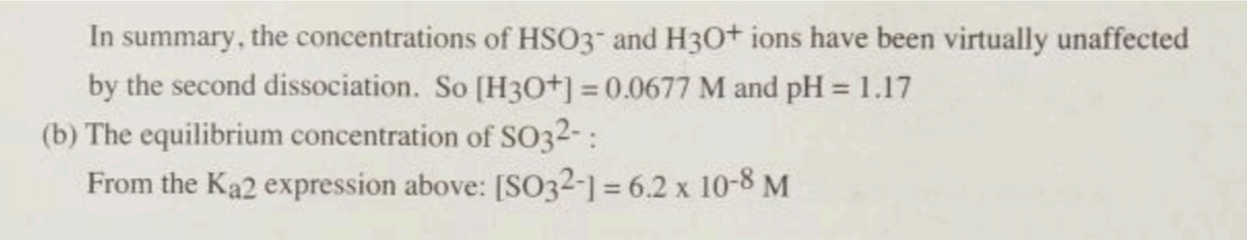
Using the equilibrium concentrations from the first step, substitute into the K_{a2} expression.

	HSO_3^-	SO_3^{2-}	H_3O^+
Initial concentration	0.0677	0	0.0677
Change	-x	+x	+x
Equilibrium	0.0677 - x	+x	0.0677 + x

$$K_{a2} = \frac{[SO_3^{2-}][H_3O^+]}{[HSO_3^-]} = \frac{(+x)(0.0677 + x)}{(0.0677 - x)} = 6.2 \times 10^{-8}$$

We note that x will be small in comparison to 0.0677, and we simplify the expression:

$$K_{a2} = \frac{(+x)(0.0677)}{(0.0677)} = 6.2 \times 10^{-8}$$



In summary, the concentrations of HSO_3^- and H_3O^+ ions have been virtually unaffected by the second dissociation. So $[H_3O^+] = 0.0677$ M and pH = 1.17

(b) The equilibrium concentration of SO_3^{2-} :

From the K_{a2} expression above: $[SO_3^{2-}] = 6.2 \times 10^{-8}$ M

4) What is the pH of 1.0 L of 0.20 M acetic acid to which 16.4 g of sodium acetate, $NaCH_3CO_2$, is added?

16.4 g of Sodium Acetate * (1 mol Sodium Acetate / 82.0343 g) = 0.2 moles Sodium acetate

0.2 M acetic acid * (1 L) = 0.2 moles acetic acid

pH = pk_a + log(conjugate base/acid)

pH = 4.7 + log(0.2 / 0.2)

pH = 4.7

5) What is the pH of the buffer solution that contains 2.2 g of NH_4Cl in 250 mL of 0.12 M NH_3? Is the final pH lower or higher than the pH of the 0.12 M ammonia solution?

2.2 g NH_4Cl * (1 mol / 53.491 g) = 0.0411 moles NH_4Cl

0.12 M NH_3 * (.250 L) = 0.03 moles NH_3

pH = pk_a + log(conjugate base/acid)

pH = 9.25 + log(0.03 / 0.0411)

pH = 9.11

$NH_3(aq) + H_2O(l) \Leftrightarrow OH^-(aq) + NH_4^+(aq)$

	NH_3	H_2O	OH^-	NH_4^+
I (M)	0.12		0	0
C (M)	-x		+x	+x
E (M)	0.12 - x		x	x

$1.8 * 10^{-5} = x^2 / 0.12 - x$

Assume x to be small relative to 0.12 so this expression simplifies to:

$1.8 * 10^{-5} = x^2 / 0.12$

$x = 0.0015$

$[OH^-] = 0.0015$ M

pH = 14 - pOH

pH = 14 - (-log(0.0015))

pH = 11.18

The final pH is lower than the pH of the 0.12 M ammonia solution.

6) What must the ratio of acetic acid to acetate ion be to have a buffer with a pH value of 4.50?

pH = pk_a + log(conjugate base/acid)

pH - pk_a = log(conjugate base/acid)

$10^{4.50 - 4.74}$ = conjugate base / acid

0.575 = conjugate base / acid

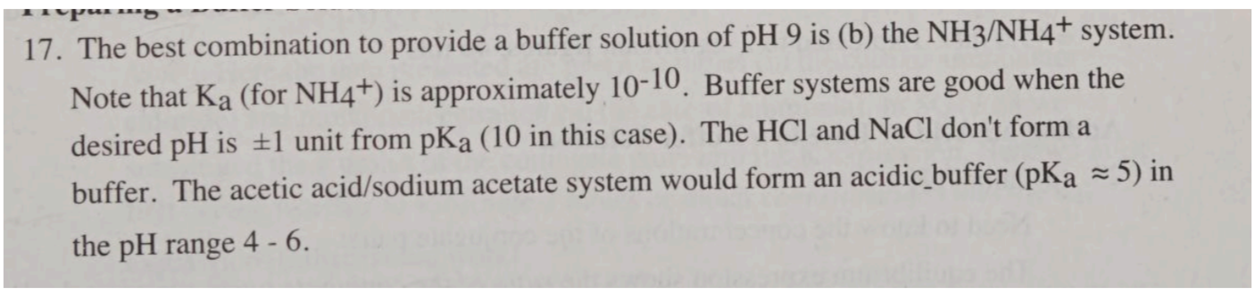
Acetic acid / Acetate Ion = 1.76

7) Which of the following combinations would be the best to buffer the pH of a solution at approximately 9?

(a) HCl and NaCl

(b) NH_3 and NH_4Cl

(c) CH_3CO_2H and $NaCH_3CO_2$



17. The best combination to provide a buffer solution of pH 9 is (b) the NH_3/NH_4^+ system. Note that K_a (for NH_4^+) is approximately 10^{-10}. Buffer systems are good when the desired pH is ±1 unit from pK_a (10 in this case). The HCl and NaCl don't form a buffer. The acetic acid/sodium acetate system would form an acidic_buffer ($pK_a \approx 5$) in the pH range 4 - 6.

8) You titrate 25.0 mL of 0.10 M NH_3 with 0.10 M HCl.

(a) What is the pH of the NH_3 solution before the titration begins?

0.025 L * 0.10 M NH_3 = 0.0025 mol NH_3

$NH_3 + H_2O \Leftrightarrow NH_4^+ + OH^-$

	NH_3	NH_4^+	OH^-
I	0.1 M	0	0
C	-x	+x	+x
E	0.1 M	x	x

$k_b = 1.8*10^{-5} = x^2/(0.1\ M)$
$x = [OH^-] = 0.001342\ M$
pOH = 2.87
pH = 14 - 2.87 = 11.13

(b) What is the pH at the equivalence point?

$NH_3 + H_3O^+ \rightarrow NH_4^+ + H_2O$

	NH_3	H_3O^+	NH_4^+	H_2O
I	0.0025 mol	0.0025 mol	0	—
C	-0.0025 mol	-0.0025 mol	+0.0025 mol	—
E	0 mol	0 mol	0.0025 mol	—

$NH_4^+ + H_2O \Leftrightarrow NH_3 + H_3O^+$

	NH_4^+	NH_3	H_3O^+
I	0.0025 mol	0	0
C	-x	+x	+x
E	0.0025 mol	x	x

$K_a = 5.8 * 10^{-10} = x^2/(0.0025 \text{ mol} /0.05 \text{ L})$
$x = [H_3O^+] = 5.39 * 10^{-6}$ M
pH = 5.28

(c) What is the pH at the halfway point of the titration?

$pH = pK_a$
$pH = -\log(5.8*10^{-10})$
pH = 9.25

Solubility

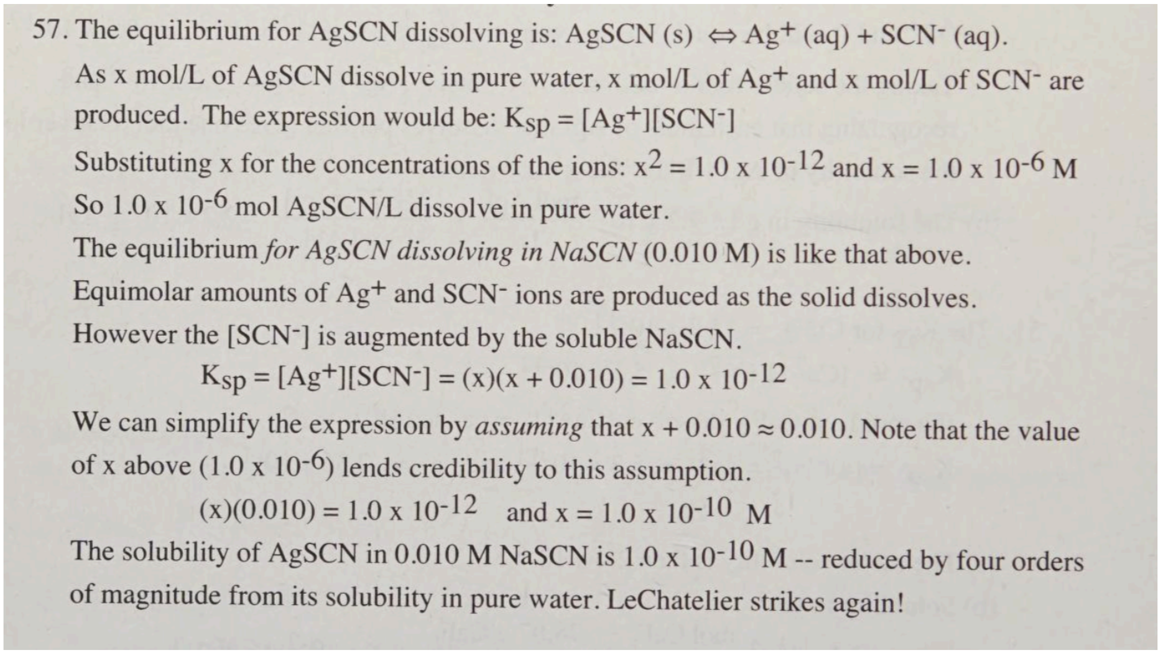
9) When 1.55 g of solid thallium(I) bromide is added to 1.00 L of water, the salt dissolves to a small extent.

$$TlBr(s) \Leftrightarrow Tl^+(aq) + Br^-(aq)$$

The thallium(I) and bromide ions in equilibrium with TlBr each have a concentration of 1.9×10^{-3} M. What is the value of K_{sp} for TlBr?

$K_{sp} = [Tl^+][Br^-] = (1.9 * 10^{-3})^2 = 3.6 * 10^{-6}$

10) Calculate the molar solubility of silver thiocyanate, AgSCN, in pure water and in water containing 0.010 M NaSCN.



57. The equilibrium for AgSCN dissolving is: AgSCN (s) $\Leftrightarrow$ Ag^+ (aq) + SCN^- (aq).

As x mol/L of AgSCN dissolve in pure water, x mol/L of Ag^+ and x mol/L of SCN^- are produced. The expression would be: $K_{sp} = [Ag^+][SCN^-]$

Substituting x for the concentrations of the ions: $x^2 = 1.0 \times 10^{-12}$ and $x = 1.0 \times 10^{-6}$ M

So 1.0×10^{-6} mol AgSCN/L dissolve in pure water.

The equilibrium *for AgSCN dissolving in NaSCN* (0.010 M) is like that above. Equimolar amounts of Ag^+ and SCN^- ions are produced as the solid dissolves. However the $[SCN^-]$ is augmented by the soluble NaSCN.

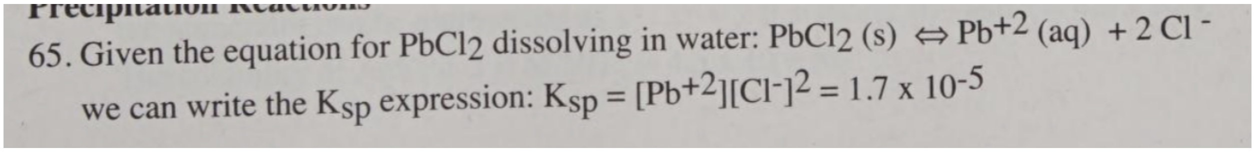
$$K_{sp} = [Ag^+][SCN^-] = (x)(x + 0.010) = 1.0 \times 10^{-12}$$

We can simplify the expression by *assuming* that $x + 0.010 \approx 0.010$. Note that the value of x above (1.0×10^{-6}) lends credibility to this assumption.

$$(x)(0.010) = 1.0 \times 10^{-12} \quad \text{and } x = 1.0 \times 10^{-10} \text{ M}$$

The solubility of AgSCN in 0.010 M NaSCN is 1.0×10^{-10} M -- reduced by four orders of magnitude from its solubility in pure water. LeChatelier strikes again!

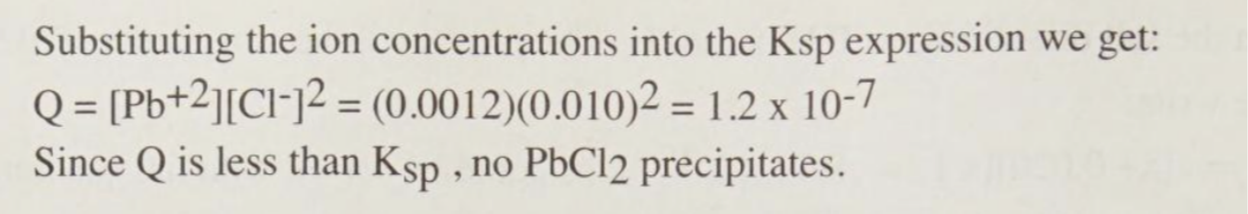
11) You have a solution that has a lead(II) ion concentration of 0.0012 M. If enough soluble chloride-containing salt is added so that the Cl^- concentration is 0.010 M, will $PbCl_2$ precipitate?



Precipitation Reactions

65. Given the equation for $PbCl_2$ dissolving in water: $PbCl_2$ (s) $\Leftrightarrow$ Pb^{+2} (aq) + 2 Cl^-

we can write the K_{sp} expression: $K_{sp} = [Pb^{+2}][Cl^-]^2 = 1.7 \times 10^{-5}$

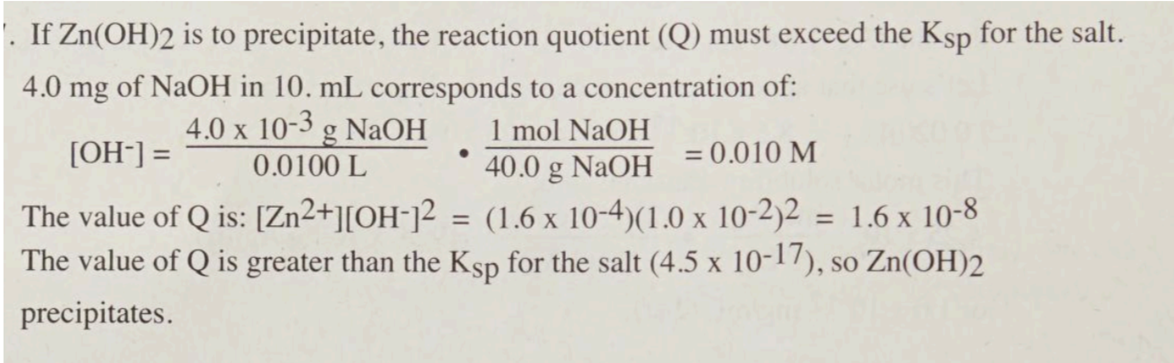


Substituting the ion concentrations into the Ksp expression we get:

$Q = [Pb^{+2}][Cl^-]^2 = (0.0012)(0.010)^2 = 1.2 \times 10^{-7}$

Since Q is less than K_{sp} , no $PbCl_2$ precipitates.

12) If the concentration of Zn^{2+} in 10.0 mL of water is 1.63×10^{-4} M, will zinc hydroxide, $Zn(OH)_2$, precipitate when 4.0 mg of NaOH is added?



. If $Zn(OH)_2$ is to precipitate, the reaction quotient (Q) must exceed the K_{sp} for the salt.

4.0 mg of NaOH in 10. mL corresponds to a concentration of:

$$[OH^-] = \frac{4.0 \times 10^{-3} \text{ g NaOH}}{0.0100 \text{ L}} \bullet \frac{1 \text{ mol NaOH}}{40.0 \text{ g NaOH}} = 0.010 \text{ M}$$

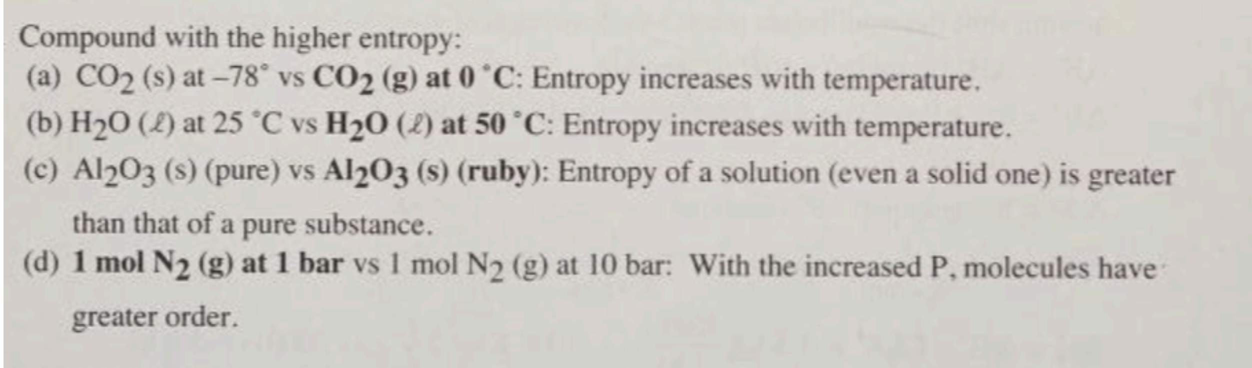
The value of Q is: $[Zn^{2+}][OH^-]^2 = (1.6 \times 10^{-4})(1.0 \times 10^{-2})^2 = 1.6 \times 10^{-8}$

The value of Q is greater than the K_{sp} for the salt (4.5×10^{-17}), so $Zn(OH)_2$ precipitates.

Thermodynamics

13) Which substance in each pair has the higher entropy?

(a) Dry ice (solid CO_2) at −78 °C or CO_2(g) at 0 °C.
(b) Liquid water at 25 °C or liquid water at 50 °C.
(d) One mole of N_2(g) at 1 bar pressure or one mole of N_2(g) at 10 bar pressure (both at 298 K).



Compound with the higher entropy:
(a) CO_2 (s) at –78° vs **CO_2 (g) at 0 °C**: Entropy increases with temperature.
(b) H_2O (ℓ) at 25 °C vs **H_2O (ℓ) at 50 °C**: Entropy increases with temperature.
(c) Al_2O_3 (s) (pure) vs **Al_2O_3 (s) (ruby)**: Entropy of a solution (even a solid one) is greater than that of a pure substance.
(d) **1 mol N_2 (g) at 1 bar** vs 1 mol N_2 (g) at 10 bar: With the increased P, molecules have greater order.

14) Calculate the standard entropy change for the following reactions at 25°C.
(a) 2 Al(s) + 3 Cl_2(g) → 2 $AlCl_3$(s)

$\Delta_r S^\circ$ = (2 mol)(109.29 J/K*mol) - [(3 mol)(223.08 J/K*mol) + (2 mol)(28.3 J/K*mol)]

$\Delta_r S^\circ$ = -507.26 J/K

(b) 2 CH_3OH(l) + 3 O_2(g) → 2 CO_2(g) + 4 H_2O(g)

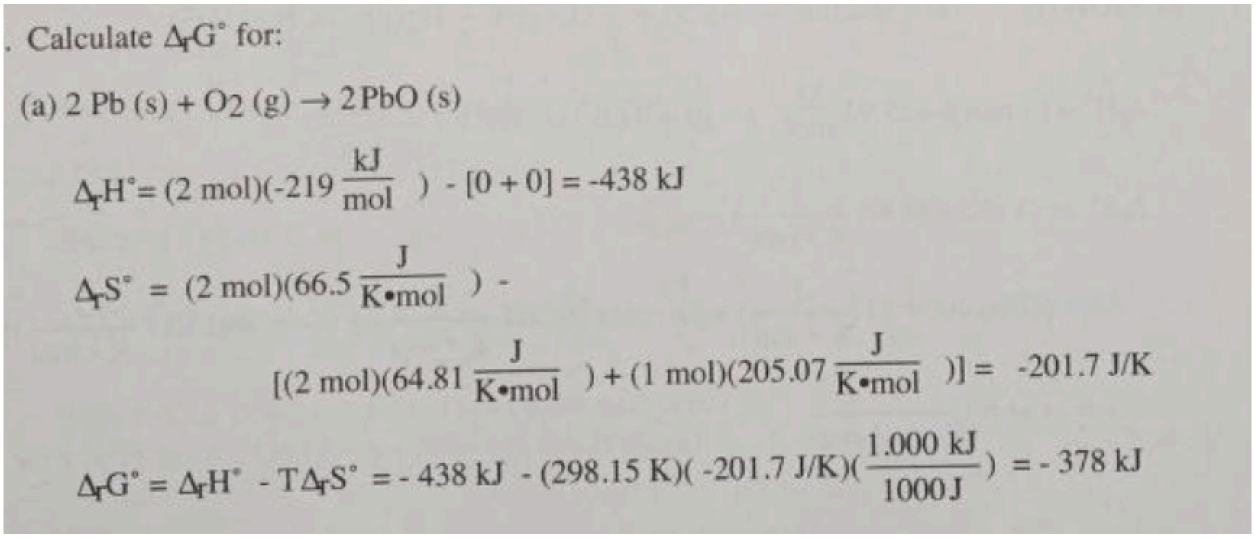
$\Delta_r S^\circ$ = [(4 mol *188.84 J/K*mol)+(2 mol * 213.74 J/K*mol)] - [(2 mol * 127.19 J/K*mol) + (3 mol * 205.07 J/K*mol)]

$\Delta_r S^\circ$ = 313.25 J/K

15) Using values of $\Delta_f H^\circ$ and S°, calculate $\Delta_r G^\circ$ for each of the following reactions at 25°C.

(a) 2 Pb(s) + O_2(g) → 2 PbO(s)
(b) NH_3(g) + HNO_3(aq) → NH_4NO_3(aq)

Which of these reactions is (are) predicted to be product-favored at equilibrium? Are the reactions enthalpy or entropy-driven?



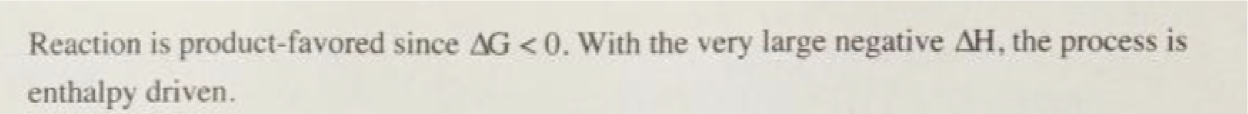
. Calculate $\Delta_r G^\circ$ for:

(a) $2\,\text{Pb (s)} + \text{O}_2\text{ (g)} \rightarrow 2\,\text{PbO (s)}$

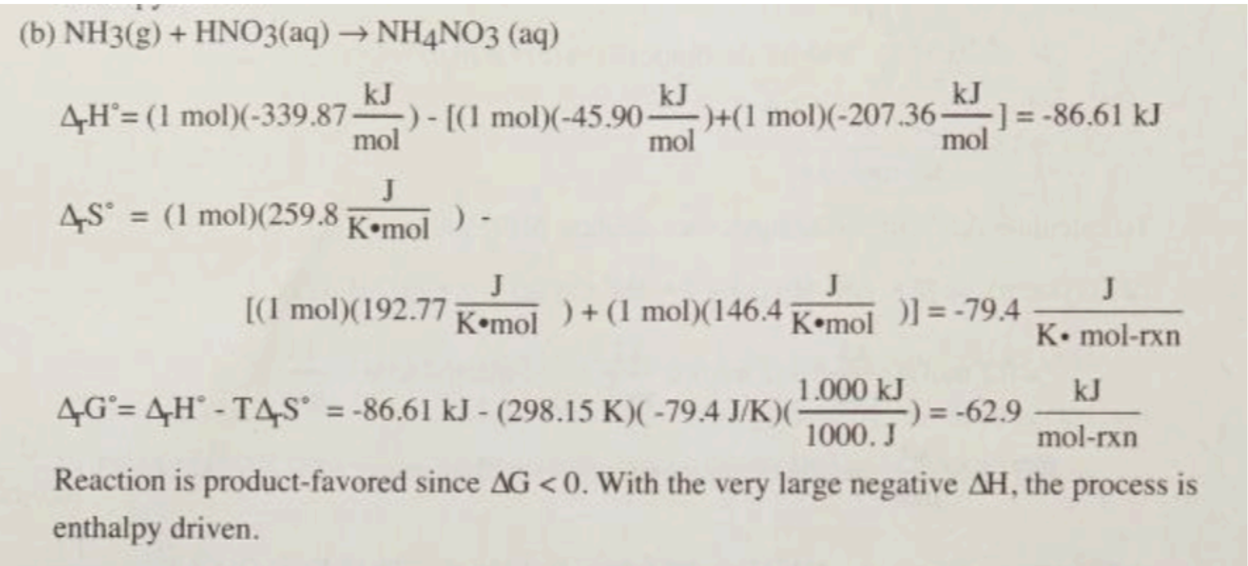
$$\Delta_r H^\circ = (2\text{ mol})(-219\,\frac{\text{kJ}}{\text{mol}}) - [0 + 0] = -438\text{ kJ}$$

$$\Delta_r S^\circ = (2\text{ mol})(66.5\,\frac{\text{J}}{\text{K}\bullet\text{mol}}) - [(2\text{ mol})(64.81\,\frac{\text{J}}{\text{K}\bullet\text{mol}}) + (1\text{ mol})(205.07\,\frac{\text{J}}{\text{K}\bullet\text{mol}})] = -201.7\text{ J/K}$$

$$\Delta_r G^\circ = \Delta_r H^\circ - T\Delta_r S^\circ = -438\text{ kJ} - (298.15\text{ K})(-201.7\text{ J/K})(\frac{1.000\text{ kJ}}{1000\text{ J}}) = -378\text{ kJ}$$



Reaction is product-favored since $\Delta G < 0$. With the very large negative ΔH, the process is enthalpy driven.



(b) $\text{NH}_3\text{(g)} + \text{HNO}_3\text{(aq)} \rightarrow \text{NH}_4\text{NO}_3\text{ (aq)}$

$$\Delta_r H^\circ = (1\text{ mol})(-339.87\,\frac{\text{kJ}}{\text{mol}}) - [(1\text{ mol})(-45.90\,\frac{\text{kJ}}{\text{mol}}) + (1\text{ mol})(-207.36\,\frac{\text{kJ}}{\text{mol}})] = -86.61\text{ kJ}$$

$$\Delta_r S^\circ = (1\text{ mol})(259.8\,\frac{\text{J}}{\text{K}\bullet\text{mol}}) - [(1\text{ mol})(192.77\,\frac{\text{J}}{\text{K}\bullet\text{mol}}) + (1\text{ mol})(146.4\,\frac{\text{J}}{\text{K}\bullet\text{mol}})] = -79.4\,\frac{\text{J}}{\text{K}\bullet\text{ mol-rxn}}$$

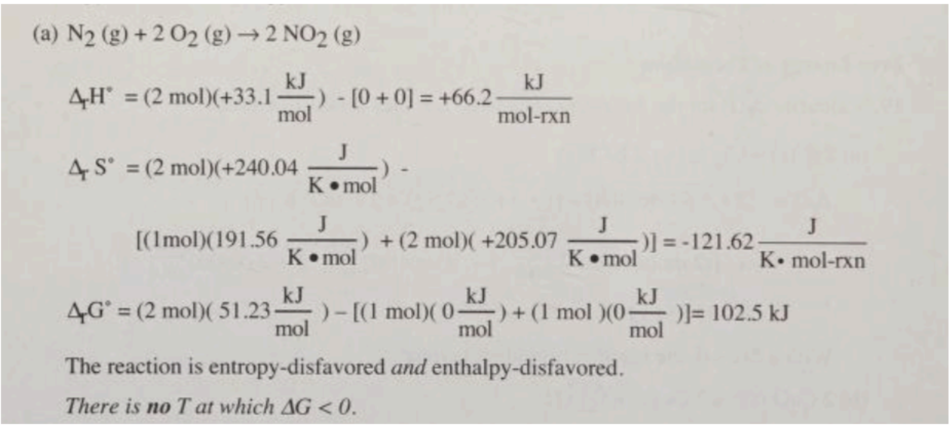
$$\Delta_r G^\circ = \Delta_r H^\circ - T\Delta_r S^\circ = -86.61\text{ kJ} - (298.15\text{ K})(-79.4\text{ J/K})(\frac{1.000\text{ kJ}}{1000.\text{ J}}) = -62.9\,\frac{\text{kJ}}{\text{mol-rxn}}$$

Reaction is product-favored since $\Delta G < 0$. With the very large negative ΔH, the process is enthalpy driven.

16) Determine whether the reactions listed below are entropy-favored or disfavored under standard conditions. Predict how an increase in temperature will affect the value of $\Delta_r G^\circ$.

(a) $\text{N}_2\text{(g)} + 2\,\text{O}_2\text{(g)} \rightarrow 2\,\text{NO}_2\text{(g)}$

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



(a) $N_2\ (g) + 2\ O_2\ (g) \rightarrow 2\ NO_2\ (g)$

$\Delta_r H^\circ = (2\text{ mol})(+33.1 \frac{\text{kJ}}{\text{mol}}) - [0 + 0] = +66.2 \frac{\text{kJ}}{\text{mol-rxn}}$

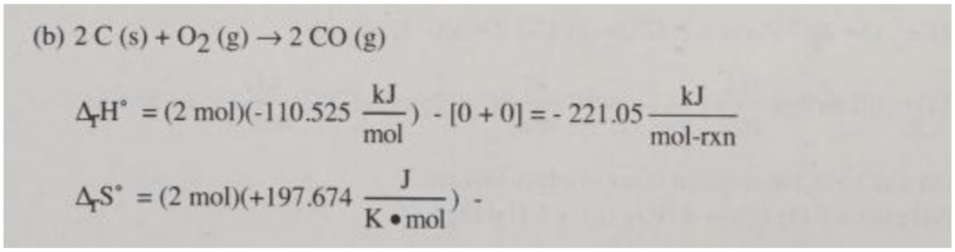
$\Delta_r S^\circ = (2\text{ mol})(+240.04 \frac{\text{J}}{\text{K} \bullet \text{mol}}) -$

$[(1\text{mol})(191.56 \frac{\text{J}}{\text{K} \bullet \text{mol}}) + (2\text{ mol})(+205.07 \frac{\text{J}}{\text{K} \bullet \text{mol}})] = -121.62 \frac{\text{J}}{\text{K} \bullet \text{mol-rxn}}$

$\Delta_r G^\circ = (2\text{ mol})(51.23 \frac{\text{kJ}}{\text{mol}}) - [(1\text{ mol})(0 \frac{\text{kJ}}{\text{mol}}) + (1\text{ mol})(0 \frac{\text{kJ}}{\text{mol}})] = 102.5\text{ kJ}$

The reaction is entropy-disfavored *and* enthalpy-disfavored.

*There is **no** T at which ΔG < 0.*



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

$\Delta_r H^\circ = (2\text{ mol})(-110.525 \frac{\text{kJ}}{\text{mol}}) - [0 + 0] = -221.05 \frac{\text{kJ}}{\text{mol-rxn}}$

$\Delta_r S^\circ = (2\text{ mol})(+197.674 \frac{\text{J}}{\text{K} \bullet \text{mol}}) -$

$\Delta_r S^\circ$ = [2 mol* 197.674 J/(K*mol)] - [1 mol *205.07 J/(K*mol) + 2 mol * 5.6 J/(K*mol)] = 179.078 J/K

$\Delta_r G^\circ = \Delta_r H^\circ - T\Delta_r S^\circ$

$\Delta_r G^\circ = -221.05\text{ kJ} - 298\text{ K}\ (179.078 \times 10^{-3}\text{ kJ/K})$

$\Delta_r G^\circ = -274.45$ kJ/mol

17) Calculate $\Delta_r G^\circ$ at 25°C for the formation of 1.00 mol of $C_2H_6(g)$ from $C_2H_4(g)$ and $H_2(g)$. Use this value to calculate K_p for the equilibrium.

$C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g)$

Comment on the sign of $\Delta_r G^\circ$ and the magnitude of K_p.

$\Delta_r G^\circ = \Sigma\Delta_f G^\circ_{\text{products}} - \Sigma\Delta_f G^\circ_{\text{reactants}}$

= -31.89 - (68.35 + 0) = -100.24 kJ/(mol*rxn)

$\Delta G°$ = -RT*ln(K)

-100.24 kJ/(mol*rxn) = -8.314 J/(mol*K) * 298 K * ln(K)

K = 3.7 * 10^{17} … Large K value means product favored.

Electrochemistry

18) Balance the following redox equation in acidic conditions.

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

Step 1)

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

$H^+(aq) \rightarrow H_2(g)$

Step 2)

$H^+(aq) + H^+(aq) \rightarrow H_2(g)$

Step 3)

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

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

Step 4)

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

19) Balance the following redox equation in basic conditions.

$$CrO_4^{2-}(aq) + SO_3^{2-}(aq) \rightarrow Cr(OH)_3(s) + SO_4^{2-}(aq)$$

Step 1)

$CrO_4^{2-}(aq) \rightarrow Cr(OH)_3(s)$

$SO_3^{2-}(aq) \rightarrow SO_4^{2-}(aq)$

Step 2)

$CrO_4^{2-}(aq) \rightarrow Cr(OH)_3(s) + H_2O(l)$

$H_2O(l) + SO_3^{2-}(aq) \rightarrow SO_4^{2-}(aq)$

Step 3)

$5\ H^+ + CrO_4^{2-}(aq) \rightarrow Cr(OH)_3(s) + H_2O(l)$

$H_2O(l) + SO_3^{2-}(aq) \rightarrow SO_4^{2-}(aq) + 2\ H^+$

Step 4)

$5\ H^+ + CrO_4^{2-}(aq) + 3e^- \rightarrow Cr(OH)_3(s) + H_2O(l)$

$H_2O(l) + SO_3^{2-}(aq) \rightarrow SO_4^{2-}(aq) + 2\ H^+ + 2e^-$

Step 5)

$10\ H^+ + 2\ CrO_4^{2-}(aq) + 6e^- \rightarrow 2\ Cr(OH)_3(s) + 2\ H_2O(l)$

$3\ H_2O(l) + 3\ SO_3^{2-}(aq) \rightarrow 3\ SO_4^{2-}(aq) + 6\ H^+ + 6e^-$

Step 6)

$4\ H^+ + 2\ CrO_4^{2-}(aq) + H_2O(l) + 3\ SO_3^{2-}(aq) \rightarrow 2\ Cr(OH)_3(s) + 3\ SO_4^{2-}(aq)$

Step 7)

$5\ H_2O(l) + 2\ CrO_4^{2-}(aq) + 3\ SO_3^{2-}(aq) \rightarrow 2\ Cr(OH)_3(s) + 3\ SO_4^{2-}(aq) + 4OH^-(aq)$

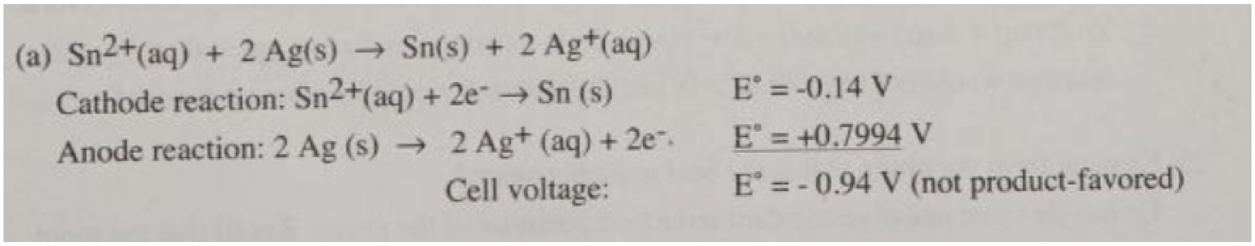
20) A voltaic cell is constructed using the reaction of chromium metal and iron(II) ions.

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

Complete the following sentences: Electrons in the external circuit flow from the Cr electrode to the Fe electrode. Negative ions move in the salt bridge from the $Fe|Fe^{2+}$ half-cell to the $Cr|Cr^{3+}$ half-cell. The half-reaction at the anode is $Cr(s) \rightarrow Cr^{3+}(aq) + 3e^-$, and that at the cathode is $Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$.

21) Balance each of the following unbalanced equations; then calculate the standard potential, *E*°cell, and decide whether each is product-favored at equilibrium as written. (All reactions are carried out in acid solution.)

(a) $Sn^{2+}(aq) + Ag(s) \rightarrow Sn(s) + Ag^+(aq)$



(a) $Sn^{2+}(aq) + 2\ Ag(s) \rightarrow Sn(s) + 2\ Ag^+(aq)$

Cathode reaction: $Sn^{2+}(aq) + 2e^- \rightarrow Sn\ (s)$ $E^\circ = -0.14\ V$

Anode reaction: $2\ Ag\ (s) \rightarrow 2\ Ag^+\ (aq) + 2e^-$. $\underline{E^\circ = +0.7994}\ V$

Cell voltage: $E^\circ = -\ 0.94\ V$ (not product-favored)

(b) $Al(s) + Sn^{4+}(aq) \rightarrow Sn^{2+}(aq) + Al^{3+}(aq)$

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

Cathode reaction: $Sn^{4+} + 2e^- \rightarrow Sn^{2+}$

Anode reaction: $Al \rightarrow Al^{3+} + 3e^-$

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode} = 0.15\ V - (-1.66\ V) = 1.81\ V$

Product favored at equilibrium

22) Calculate $\Delta_r G°$ and the equilibrium constant for the following reactions.

(a) $2\ Fe^{3+}(aq) + 2\ I^{-}(aq) \Leftrightarrow 2\ Fe^{2+}(aq) + I_2(aq)$

Cathode reaction: $2\ Fe^{3+} + 2\ e^{-} \rightarrow 2\ Fe^{2+}$

Anode reaction: $2\ I^{-} \rightarrow I_2 + 2e^{-}$

$E°_{cell} = E°_{cathode} - E°_{anode} = 0.771\ V - 0.621\ V = 0.15\ V$

$\Delta_r G° = -nFE°_{cell}$
$\Delta_r G° = -2 * 96{,}000\ C * 0.15\ V = -28.8\ kJ$

$E°_{cell} = (0.0257/n) * \ln K$
$0.15 = 0.0257/2 * \ln(K)$
$K = 1.2 * 10^{5}$

(b) $I_2(aq) + 2\ Br^{-}(aq) \Leftrightarrow 2\ I^{-}(aq) + Br_2(l)$

Cathode reaction: $I_2 + 2e^{-} \rightarrow 2\ I^{-}$

Anode reaction: $2\ Br^{-} \rightarrow Br_2 + 2e^{-}$

$E°_{cell} = E°_{cathode} - E°_{anode} = 0.621\ V - 1.08\ V = -0.46\ V$

$\Delta_r G° = -nFE°_{cell}$
$\Delta_r G° = -2 * 96{,}000\ C * -0.46\ J = 88.3\ kJ$

$E°_{cell} = (0.0257/n) * \ln K$
$-0.46 = (0.0257/2) * \ln K$
$K = 2.8 * 10^{-16}$

23) The standard potential, E°cell, for the reaction of Zn(s) and $Cl_2(g)$ is +2.12 V. What is the standard free energy change, $\Delta_r G°$, for the reaction?

$\Delta_r G° = -nFE°_{cell}$
$\Delta_r G° = -2 * 96000 * 2.12 = -407\ kJ$