 # 1. (10 points) Short Answer Questions * indicates partial points possible *

a) The conjugate acid and conjugate base of NH_3 are ____NH_4^+____ and ____ NH_2^-____, respectively. (2)*

b) Which of the following compound is categorized as soluble: *(circle one)* (1)*

Ag_2CO_3 **$BaSO_4$** [✓ **Na_2SO_4**] **$Mg(OH)_2$**

c) Using the following information: (1)*

$$P_4(s) + 6Cl_2(g) \rightleftharpoons 4PCl_3(g) \qquad K_c = 2.00 * 10^{19}$$

$K_1 = 1 / K_c = 5.00 \times 10^{-20}$

$$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g) \qquad K_c = 1.13 * 10^{-2}$$

$K_2 = (K_c)^4 = 1.63 \times 10^{-8}$

Determine K_c for: $4\,PCl_5(g) \rightleftharpoons P_4(s) + 10\,Cl_2(g)$

$K_c =$ $K_1 \times K_2 = 8.15 \times 10^{-28}$

d) Show the full ionic equation and the net ionic equation associated with the reaction of Na_2SO_4(*aq*) with $Pb(NO_3)_2$(*aq*).

Ionic equation: (1)*

$2\ Na^+(aq) + SO_4^{2-}(aq) + Pb^{2+}(aq) + 2\ NO_3^-(aq) \rightarrow PbSO_4(s) + 2\ Na^+(aq) + 2\ NO_3^-(aq)$

Net Ionic equation: (1)* (Partial points only if aqueous phases are missing)

$SO_4^{2-}(aq) + Pb^{2+}(aq) \rightarrow PbSO_4(s)$

e) What is the value of K_b associated with the reaction below (1) : $K_b = K_w/K_a = 1.6 \times 10^{-5}$

$$CN^- + H_2O \rightleftharpoons HCN + OH^-$$

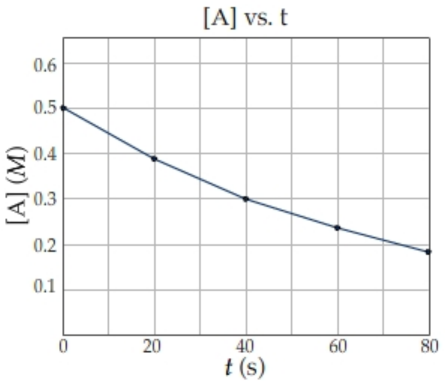
f) What is the half-life of a first order reaction with a rate constant of $1.34*10^{-3}$: (1)* (0.5 no units)

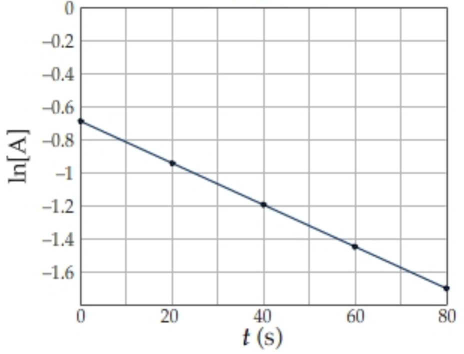
ln(2)/k = 517 s

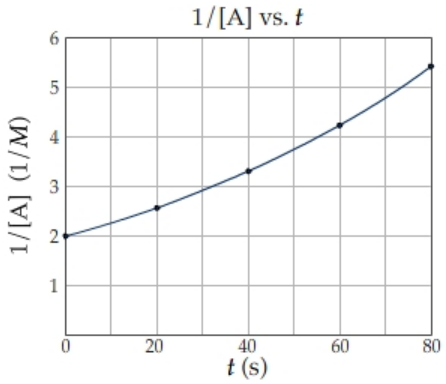
g) According to collision theory, what are the two requirements for a bimolecular reaction to occur? (2)

1. Enough kinetic energy to overcome the E_a
2. Correct orientation (geometry)

 # 2. (10 points) In a process described by the overall reaction $2A \longrightarrow B$, the following graphs can be obtained from experimental data, all at 295 K:







a) What is the order for the reaction? (*circle one answer*) **(1)**

Zero order **[✓ First order]** **Second order** **Third order**

b) What is the value of the rate constant for the reaction? (3) (1 calc / 1 answer / 1 units)

k = -slope = $-\left(\frac{\Delta \ln[A]}{\Delta t}\right) = -\left(\frac{\ln[A]_f - \ln[A]_i}{t_f - t_f}\right) = -\left(\frac{-1.7-(-0.7)}{80s - 0s}\right) = 1.25\text{x}10^{-2}$

Answer: 1.25 x 10^{-2} s^{-1}

c) Using the *appropriate integrated rate law*, determine the expected concentration of A after 40 seconds of reaction, if the initial concentration of A is changed to 0.8 M. (3)

$$\ln[A]_t = -kt + \ln[A]_o \qquad \ln[\text{A}]_{40s} = -1.25\text{x}10^{-2} \bullet 40s + \ln[0.8\text{M}]$$

$$\ln[\text{A}]_{40s} = -7.23\text{x}10^{-1}$$

$$[\text{A}]_{40s} = e^{\ln[A]} = e^{-7.23\text{x}10^{-1}} = 0.485\text{M}$$

Answer: 0.485M

d) Knowing that the rate constant (k) is 0.0357 s^{-1} for the same reaction at 325 K, calculate the activation energy of the reaction. (3)

$$\ln\left(\frac{k_2}{k_1}\right) = \left(-\frac{\text{Ea}}{\text{R}}\right)\left(\frac{1}{\text{T}_2} - \frac{1}{\text{T}_1}\right) \qquad \ln\left(\frac{3.57\text{x}10^{-2}}{1.25\text{x}10^{-2}}\right) = \left(-\frac{\text{Ea}}{8.31451\ \text{JK}^{-1}\text{mol}^{-1}}\right)\left(\frac{1}{325} - \frac{1}{295}\right)$$

$$\text{Ea} = \ln\left(\frac{3.57\text{x}10^{-2}}{1.25\text{x}10^{-2}}\right) \bullet \frac{1}{\left(\frac{1}{325\text{K}} - \frac{1}{295\text{K}}\right)} \bullet -8.31451\ \text{JK}^{-1}\text{mol}^{-1} = 27886\ \text{J/mol} = 27.9\ \text{kJ/mol}$$

Answer: 27.9 kJ/mol

 # 3. (10 points) The production of sulfuric acid involves as one of its step the oxidation of sulfur dioxide with excess oxygen to produce sulfur trioxide, according to the following equilibrium:

$$2\ SO_2(g) + O_2(g) \rightleftharpoons 2\ SO_3(g) \qquad \Delta H = -\ 197\ kJ/mol$$

a) In a reaction, 1.00 mol of SO_2, 0.600 mol of O_2 and 0.416 mol of SO_3 are introduced simultaneously in a 2.00 L vessel at 1000 K. Once equilibrium is reached, 0.210 mol of O_2 remain in the vessel. Determine the equilibrium amount of SO_2 and SO_3 (*in mol*), as well as the equilibrium constant K_c for the reaction. (6)

Initial concentrations: 0.500 M SO_2, 0.300 M O_2, 0.208 M SO_3

Final concentrations: 0.105 M O_2

	$2\ SO_2(g)$ +	**O_2 (g) $\rightleftharpoons$**	**$2\ SO_3(g)$**
I *(1 pt)*	0.500	0.300	0.208
C *(1 pt)*	**- 0.390**	**- 0.195**	**+ 0.390**
E *(1 pt)*	**0.110**	0.105	**0.598**

$$\boxed{K_c = \frac{[SO_3]^2}{[SO_2]^2 \bullet [O_2]} = \frac{0.598^2}{0.110^2 \bullet 0.105} = 281 \quad \textit{(2 pts –1 formula, 1 concentrations \& answer)}}$$

Equilibrium SO_2: 0.220 mol *(0.5)* Equilibrium SO_3: 1.20 mol *(0.5)* **K_c: 281**

b) Determine the value of K_p at 1000 K (work with pressure units in atmospheres: i.e. *atm*)**. (2)**

$$\boxed{\begin{aligned} &K_p = K_c \bullet (RT)^{\Delta n} = 281 \bullet (0.08206\ \ atm{\cdot}L{\cdot}K\text{-}1{\cdot}mol\text{-}1\ x\ 1000K)^{2-3} \\ &K_p = 3.42 \end{aligned}}$$

K_p: 3.42

c) From the equilibrated mixture obtained in a), would a drop in temperature of the reaction to 800 K shift the equilibrium towards the products, the reactants or not affect the equilibrium at all? Provide brief reasoning for your answer. **(2)**

A shift of the equilibrium towards the products is expected, as a reduced temperature would favor the reaction releasing energy to its surroundings. (Known from negative ΔH).

 # 4. (10 points) Fractional separation involves precipitating one ion voluntarily while keeping the other in solution. In an attempt to separate chloride and iodide ions, an aqueous solution of 2.00 M $AgNO_3$ is slowly added to a solution of 0.0100 M in Cl^- and 0.250 M in I^-.

a) Which anion, between Cl^- and I^-, precipitates first? (*You need to show your work*) **(5)**

Calculate the $[Ag^+]$ at which both ions would start precipitate from K_{sp} values:

For chlorine: (2)

$K_{sp} = [Ag^+][Cl^-]$ hence: $[Ag^+] = K_{sp} / [Cl^-] = 1.8 \times 10^{-10} / 0.0100$ M $= 1.8 \times 10^{-8}$ M

For Iodine: (2)

$K_{sp} = [Ag^+][I^-]$ hence: $[Ag^+] = K_{sp} / [I^-] = 1.5 \times 10^{-16} / 0.250$ M $= 6.0 \times 10^{-16}$ M

Once the concentration of Ag^+ reaches 6.0×10^{-16} M (the smaller of the two conc.), AgI will start to precipitate (1)

Answer: I^- precipitates (as AgI) first.

b) When the second anion (Cl^- or I^-) starts to precipitate, what is the remaining concentration of the first anion in solution. (4)

When AgCl starts to precipitate, the concentration of Ag+ will be 1.8×10^{-8} M (found in part a – *1 pt*). At this point, concentration of I^- can be found from Ksp.

$K_{sp} = [Ag^+][I^-]$

hence: $[I^-] = K_{sp} / [Ag^+] = 1.5 \times 10^{-16} / 1.8 \times 10^{-8}$ M $= 8.3 \times 10^{-9}$ M (*1 pt formula, 1 pt answer*)

Answer: 8.3×10^{-9} M (***units 1 pt***)

c) Assume that the separation of ions is effective if > 99 % of the first ions have precipitated while the other type is still in solution. **Would you consider this technique appropriate for the separation of Cl^- from I^-?** (*Circle one answer*) **(1 – has to match answer in b)**

[✓ YES] **NO**

 # 5. (10 points) A student prepares a 0.240 M solution of an unknown monoprotic acid (HA). At equilibrium, he measures a pH for the solution of 2.56. **(6)**

a) What is the acid dissociation constant (K_a) for the acid?

At equilibrium: pH = - log [H+] = 2.56

[H+] = 10^{-pH} = $10^{-2.56}$ = $2.75 \text{ x } 10^{-3}$ (1 pt formula, 1 pt answer)

1 pt overall	**HA $\rightleftharpoons$**	**H+ +**	**A-**
I	0.240	0	0
C ***(1 pt stoichiometry)***	**$-2.75 \text{ x } 10^{-3}$**	**$+2.75 \text{ x } 10^{-3}$**	**$+2.75 \text{ x } 10^{-3}$**
E	**0.237**	**$2.75 \text{ x } 10^{-3}$**	**$2.75 \text{ x } 10^{-3}$**

$$\boxed{K_a = \frac{[H^+][A^-]}{[HA]} = \frac{[2.75 \text{ x } 10^{-3}]^2}{0.237} = 3.19 \text{ x } 10^{-5}}$$ **(1 pt formula, 1 pt answer)**

Answer: **$3.19 \text{ x } 10^{-5}$**

b) A lab technician is trying to determine the concentration of his solution of cinnamic acid ($C_9H_7O_2H$ – MM 148 g/mol), a *monoprotic acid* commonly used in the manufacture of flavors and perfumes. **A 10.00 mL sample of the cinnamic acid solution is titrated to completion with 32.55 mL of 0.1200M NaOH. What is the % mass of Cinnamic acid in the original solution (assume a density of 1.020 g/ml for the solution)? (4)**

(2 pts for the dimensional analysis / finding the grams of acid)

(1 pt for the % mass formula, 1 pt for final answer)

$$\text{g of acid} = 32.55 \text{ mL NaOH} \bullet \frac{1\text{L}}{1000\text{ml}} \bullet \frac{0.1200\text{mol NaOH}}{\text{L NaOH}} \bullet \frac{1 \text{ mol acid}}{1 \text{ mol NaOH}} \bullet \frac{148\text{g acid}}{1 \text{ mol acid}}$$

$$\text{g of acid} = 0.578\text{g}$$

$$\% \text{ mass acid} = \frac{0.578\text{g acid}}{10.00\text{mL}} \bullet \frac{1\text{mL}}{1.02\text{g}} * 100\% = 5.67 \%$$

Answer: **5.67 %**