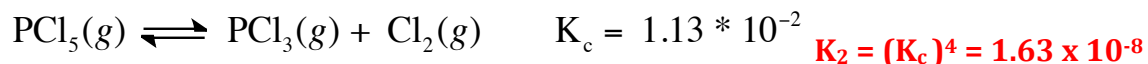


1. (10 points) Short Answer Questions * indicates partial points possible *

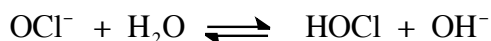
a) Which of the following compound(s) is(are) categorized as soluble: (1)*



b) Using the following information: (1)*

Determine K_c for: $4 \text{PCl}_5(g) \rightleftharpoons \text{P}_4(s) + 10 \text{Cl}_2(g)$

$$K_c = K_1 * K_2 = 8.15 * 10^{10}$$

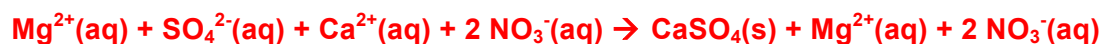
c) What is the value of K_b associated with the reaction below: (1) K_b = K_w/K_a = 3.4 x 10⁻⁷d) The conjugate acid and conjugate base of HCO₃⁻ are ___ H₂CO₃___ and ___ CO₃²⁻___, respectively. (2)*

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

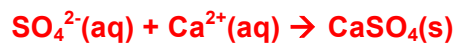
1. Enough kinetic energy to overcome the E_a
2. Correct orientation (geometry)

f) Show the full ionic equation and the net ionic equation associated with the reaction of MgSO₄(aq) with Ca(NO₃)₂(aq).

Ionic equation: (1)*

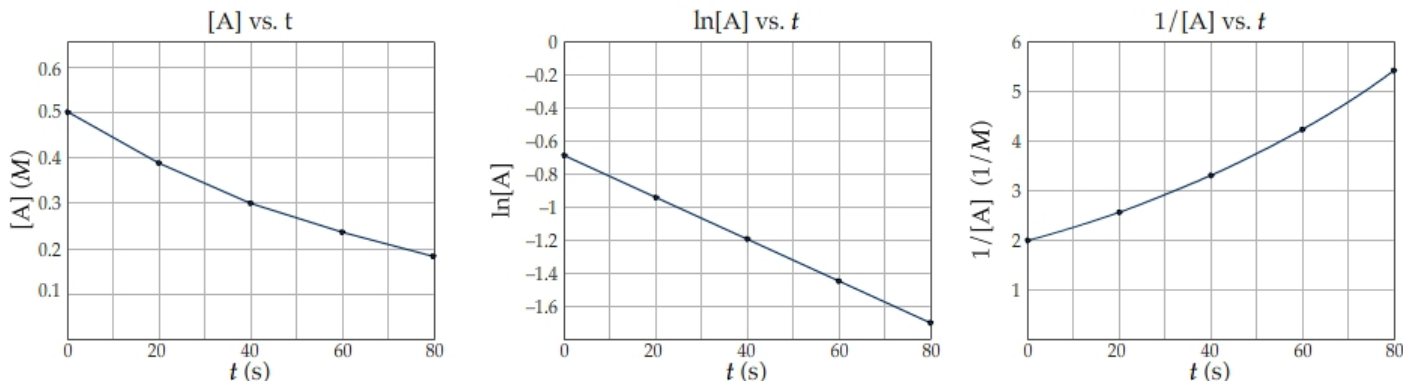


Net ionic equation: (1)* (Partial points only if aqueous phases are missing)

g) What is the half-life of a first order reaction with a rate constant of 1.52*10⁻⁴: (1)* (0.5 no units)

$$\ln(2)/k = 4.56 * 10^3 \text{ s}$$

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)

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

Answer: $1.25 \times 10^{-2} \text{ s}^{-1}$

c) Using the appropriate integrated rate law, determine the expected concentration of A at 52 seconds. (3)

$$\begin{aligned} \ln[A]_t &= -kt + \ln[A]_0 & \ln[A]_{52\text{s}} &= -1.25 \times 10^{-2} \cdot 52\text{s} + \ln[0.5\text{M}] \\ & & \ln[A]_{52\text{s}} &= -1.34 \\ [A]_{52\text{s}} &= e^{\ln[A]} = e^{-1.34} = 0.261\text{M} \end{aligned}$$

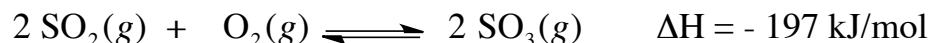
Answer: **0.261M**

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)

$$\begin{aligned} \ln\left(\frac{k_2}{k_1}\right) &= \left(-\frac{E_a}{R}\right)\left(\frac{1}{T_2} - \frac{1}{T_1}\right) & \ln\left(\frac{3.57 \times 10^{-2}}{1.25 \times 10^{-2}}\right) &= \left(-\frac{E_a}{8.31451 \text{ JK}^{-1}\text{mol}^{-1}}\right)\left(\frac{1}{325} - \frac{1}{295}\right) \\ E_a &= \ln\left(\frac{3.57 \times 10^{-2}}{1.25 \times 10^{-2}}\right) \cdot \frac{1}{\left(\frac{1}{325\text{K}} - \frac{1}{295\text{K}}\right)} \cdot -8.31451 \text{ JK}^{-1}\text{mol}^{-1} = 27886 \text{ J/mol} = 27.9 \text{ kJ/mol} \end{aligned}$$

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:



a) In a reaction, 1.05 mol of SO_2 , 0.600 mol of O_2 and 0.852 mol of SO_3 are introduced simultaneously in a 2.00 L vessel at 1000 K. Once equilibrium is reached, 0.220 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: 0.525 M SO_2 , 0.300 M O_2 , 0.426 M SO_3

Final concentrations: 0.110 M O_2

	2 $\text{SO}_2(\text{g})$	+	$\text{O}_2(\text{g})$	\rightleftharpoons	2 $\text{SO}_3(\text{g})$
I (1 pt)	0.525		0.300		0.426
C (1 pt)	- 0.380		- 0.190		+ 0.380
E (1 pt)	0.145		0.110		0.806

$$K_c = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2 \cdot [\text{O}_2]} = \frac{0.806^2}{0.145^2 \cdot 0.110} = 281 \text{ (1 ps - formula, 1 pt concentrations \& answer)}$$

Equilibrium SO_2 : 0.290 mol (0.5pt) Equilibrium SO_3 : 1.61 mol (0.5pt) K_c : 281

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

$$K_p = K_c \cdot (RT)^{\Delta n} = 281 \cdot (0.08206 \text{ atm}\cdot\text{L}\cdot\text{K}^{-1}\cdot\text{mol}^{-1} \times 1000\text{K})^{2-3}$$

$$K_p = 3.42$$

K_p : 3.42

c) From the equilibrated mixture obtained in a), would an increase in temperature of the reaction to 1500 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 reactants is expected, as an increased temperature would favor the reaction absorbing energy from 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 0.0150 M in Cl^- and 0.280 M in I^- .

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

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

For chlorine: (2)

$$K_{\text{sp}} = [\text{Ag}^+][\text{Cl}^-] \text{ hence: } [\text{Ag}^+] = K_{\text{sp}} / [\text{Cl}^-] = 1.8 \times 10^{-10} / 0.0150 \text{ M} = 1.2 \times 10^{-8} \text{ M}$$

For Iodine: (2)

$$K_{\text{sp}} = [\text{Ag}^+][\text{I}^-] \text{ hence: } [\text{Ag}^+] = K_{\text{sp}} / [\text{I}^-] = 1.5 \times 10^{-16} / 0.280 \text{ M} = 5.4 \times 10^{-16} \text{ M}$$

Once the concentration of Ag^+ reaches $5.4 \times 10^{-16} \text{ M}$ (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 $1.2 \times 10^{-8} \text{ M}$ (found in part a – 1 pt). At this point, concentration of I^- can be found from K_{sp} .

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

$$\text{hence: } [\text{I}^-] = K_{\text{sp}} / [\text{Ag}^+] = 1.5 \times 10^{-16} / 1.2 \times 10^{-8} \text{ M} = 1.25 \times 10^{-8} \text{ M} \text{ (1 pt formula, 1 pt answer)}$$

Answer: $1.25 \times 10^{-8} \text{ M}$ (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 system 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 0.310 M solution of an unknown monoprotic acid (HA). At equilibrium, he measures a pH for the solution of 1.59. (6)

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

At equilibrium: $\text{pH} = -\log [\text{H}^+] = 1.59$

$[\text{H}^+] = 10^{-\text{pH}} = 10^{-1.59} = 2.57 \times 10^{-2}$ (1 pt formula, 1 pt answer)

<i>1 pt overall</i>	HA	\rightleftharpoons	H+	+	A-
I	0.310		0		0
C (1 pt stoichiometry)	-2.57×10^{-2}		$+2.57 \times 10^{-2}$		$+2.57 \times 10^{-2}$
E	0.284		2.57×10^{-2}		2.57×10^{-2}

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} = \frac{[2.57 \times 10^{-2}]^2}{0.284} = 2.33 \times 10^{-3} \quad (1 \text{ pt formula, 1 pt answer})$$

Answer: 2.33×10^{-3}

b) A lab technician is trying to determine the concentration of his solution of cinnamic acid ($\text{C}_9\text{H}_7\text{O}_2\text{H}$ – 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 27.45 mL of 0.1300M 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)

$$\begin{aligned} \text{g of acid} &= 27.45 \text{ mL NaOH} \cdot \frac{1\text{L}}{1000\text{ml}} \cdot \frac{0.1300\text{mol NaOH}}{\text{L NaOH}} \cdot \frac{1 \text{ mol acid}}{1 \text{ mol NaOH}} \cdot \frac{148\text{g acid}}{1 \text{ mol acid}} \\ \text{g of acid} &= 0.528\text{g} \\ \% \text{ mass acid} &= \frac{0.528\text{g acid}}{10.00\text{mL}} \cdot \frac{1\text{mL}}{1.02\text{g}} * 100\% = 5.18 \% \end{aligned}$$

Answer: 5.18 %