



Chemistry 1302B Practice Final Exam

- Solutions -

Introduction

Dear Student,

Thank you for opening the Chemistry 1302B practice final exam solutions for Chemistry 1302B. This resource has been created by the Education Team at WebStraw. The Education Team consists of students that have previously taken and/or students that are currently taking Chemistry 1302B.

Purpose

This resource focuses on key concepts that are important for students to understand to succeed within this course. This resource was created by students for other students. Our goal is to help students (1) further develop their understanding of course content and (2) achieve greater academic success. (3) Our resource is also open access meaning there are no financial or legal barriers to students who wish to access and use our resource.

Instructions

Upon completion of the Practice Questions, compare your thought-process and answers to the solutions provided.

Disclaimer

This resource is supplementary to your course content and is not meant to (1) replace any of the resources provided to you by your instructor nor is it meant to (2) be used as a tool to learn the course material from scratch. We assume that students who use this resource will have a basic understanding of the course content. As such, many of the solutions expect you to have the basic knowledge of understanding (e.g, knowing the formula for equilibrium). This resource does not contain everything you need to know for your evaluations. Please refer to the course material provided by your instructors if there are any discrepancies between our resource and your course content.

We wish you the best of luck on your exams!
The WebStraw Team

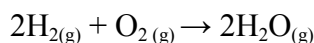
Note to Instructors: If this resource has been created for your course and you would like to collaborate with us, please email us at team@webstraw.ca

1. Ideal Gas Laws

Question Type: “Partial Pressure of Gas in a Beaker”

Thought Process:

To begin this question, balancing the chemical reaction is crucial. Without this, molar/volume/pressure calculations cannot be made as they are dependent on the correct ratios of reagents in the mixture. This helps you calculate the limiting reagent.



Once the chemical reaction has been balanced the limiting reagent can be identified. For context this question stated that equal volumes of hydrogen and oxygen gas were mixed together, but as you can see, we need 2 moles of hydrogen gas for every mole of oxygen gas to create water. This means that given equal amounts of hydrogen and oxygen, hydrogen will be used up before we run out of oxygen → it is the **limiting reagent**.

Solution:

Given that 30g of water were formed, we know that there will have been **1.66 moles** of water synthesized (since Moles = Mass / Molar mass)

Since hydrogen is the limiting reagent, and the question states that the reaction was pushed to completion, we can assume that there is no hydrogen gas left in the vessel, and that only O₂ remains. Hence, the partial pressure of hydrogen is effectively zero.

Since 1.66 moles of water were generated, then 1.66 moles of hydrogen (and by extension 1.66 moles of oxygen) were pumped into the vessel. However, since the mole ratio for O₂ and water is 1:2, only 0.83 moles of oxygen is required to produce 1.66 moles of water, so **0.83 moles** of oxygen remains.

We can then plug in the 0.83 moles of oxygen into **PV = nRT** to find the final partial pressure of oxygen within the vessel.

Solution: 1953 Pa of oxygen only

1. Ideal Gas Laws

Question Type: “Relationship Between the Flow of Two Gases”

Thought process:

The fact that a small hole is made in the vessel's wall should remind you of the idea of **effusion**, and this is what we will consider to answer this question. In order to find the moles of CH_4 that remain, we can find the ratio for the rate of effusion between the two gases.

$$r_{\text{eff}}\text{He}/r_{\text{eff}}\text{CH}_4 = \sqrt{(\text{MM}_{\text{CH}_4}/\text{MM}_{\text{He}})}$$
$$r_{\text{eff}}\text{He}/r_{\text{eff}}\text{CH}_4 = \sqrt{(16.042 \text{ gmol}^{-1}/4.003 \text{ gmol}^{-1})}$$
$$r_{\text{eff}}\text{He}/r_{\text{eff}}\text{CH}_4 = 2.00$$

Since the rate of effusion is 2 times greater for He than for methane gas, we can expect the amount effused for CH_4 to be $\frac{1}{2}$ the amount of the helium effused. This means that for the 0.4 mol (1.0 mol - 0.6 mol) of He which effused, half the amount or 0.2 mol CH_4 effused within the same time. Subtracting this from the starting amount of 1.0mol, we are left with **0.8 mol of CH_4** .

Solution: D

1. Ideal Gas Laws

Question Type: “Partial Pressures at Different Temperatures”

Thought process:

From Dalton’s Law of Partial Pressures, we know that the total pressure of a mixture of gases is equal to the sum of the partial pressures of each gas involved. We can first find the partial pressure of nitrogen gas at 65 °C by using this property since we are given the total pressure and the vapour pressure. Liquid water adds no pressure to the system.

$$P_{N_2} = 110.0 \text{ kPa} - 25.0 \text{ kPa} = 85.0 \text{ kPa}$$

Now that we have the partial pressure for nitrogen gas, we can use the combined Charles’s and Boyle’s Law in order to compute for the pressure of nitrogen gas at 50 °C. The reason we don’t do this directly with the vapour pressure is because we are told in the question that there is still some liquid water left over. This means the reaction **did not go to completion** and not all of the water became a gas.

$$P_1/T_1 = P_2/T_2$$

$$P_1 = P_2 T_1 / T_2$$

$$= (85.0 \text{ kPa})(50 + 273.15 \text{ K}) / (65 + 273.15 \text{ K})$$

$$= 81.23 \text{ kPa}$$

Using Dalton’s Law of Partial pressures again, we find that the vapour pressure at 50°C is **12.3 kPa**.

Solution: B

2. Thermodynamics

Question Type: “ Energy Required to Change States of Matter”

Thought Process:

In order to calculate the amount of energy required to convert the ice into water, we must consider the energy input at the 3 different stages: bringing the ice to 0°C, converting the ice into water, and heating the water to 10°C. Note that since we're talking about a change in temperature, celsius is equivalent to kelvin. The calculations are as follows:

1) Bringing the ice to 0°C

$$Q = mc\Delta T$$

$$Q = (45\text{g})(2.05\text{J/g } ^\circ\text{C})(16^\circ\text{C})$$

$$Q = 1.4760\text{kJ}$$

2) Melting the ice into liquid water

$$Q = n\Delta H_{\text{fus}}$$

$$Q = (2.4978)(6.01\text{kJ/mol})$$

$$Q = 15.0117\text{kJ}$$

3) Heating the water to 10°C.

$$Q = mc\Delta T$$

$$Q = (45\text{g})(4.184\text{J/g } ^\circ\text{C})(10^\circ\text{C})$$

$$Q = 1.8828\text{kJ}$$

$$Q = mc\Delta T + n\Delta H + mc\Delta T$$

Solution: 18.37kJ

2. Thermodynamics

Question Type: “Entropy”

Thought Process:

When doing these types of questions there are three things we want to look for, 1) the **states** of the products and reactants (g, l, s, and sometimes aq) 2) Whether it is a **synthesis or decomposition** reaction, and 3) the **molar coefficients** of the reactants and products.

Answers A,B, and C all result in a **reduction** in the total moles of gas. Gases have the highest entropy, and so a loss in the total amount of gas results in a reduced entropy. Hence A, B, and C are wrong. Answer E results in no change in the moles of gas, but is a synthesis reaction; compared to answer D, which is a decomposition, answer E thus has a lower change in entropy. As a result, D is the correct answer.

Solution: D

3. Electrochemistry

Question Type: “Lithium Ion Battery”

Thought Process:

To find the # of **moles of electrons**, use the two equations $Q = It$ and $Q = n_e F$, where Q is the total charge. Faraday’s constant (F) is the total charge in coulombs for one mole of electrons.

Note that the stoichiometric coefficients of the electrons are the same as both those of LiC_6 and CoO_2 . As a result, the moles of LiC_6 and CoO_2 are **equal** to the moles of electrons. To save time, we can take the molar masses of LiC_6 and CoO_2 and sum them together (only because there are equal moles of each); the molar mass of LiC_6 is 79.00g/mol and the molar mass of CoO_2 is 90.93 g/mol. The **total molar mass** is 169.93g/mol.

We may then use this to find the total mass of reactants (LiC_6 and CoO_2). If you were to find the mass of each reactant separately and added them together at the end, you would get the same result. Feel free to use that if you don’t quite understand why it’s mathematically correct to sum the molar masses first!

Solution:

$$Q = It$$

$$Q = (0.200 \text{ A}) (36000\text{s}) = 7200\text{C}$$

$$n_e = Q/F$$

$$n_e = (7200\text{C}) / (96485\text{C/mol}) = 0.074623 \text{ mol}$$

$$m = n_e \times \text{MM}$$

$$m = 0.074623 \text{ mol} \times 169.93 \text{ g/mol}$$

$$\mathbf{m = 12.7 \text{ grams}}$$

3. Electrochemistry

Question Type: “Cell Diagrams and Cell Potential I”

Thought Process:

We must identify both the reduction and oxidation half-reactions and look up their standard potentials. Make sure to take the value with the **opposite sign** for the oxidation half-reactions. Note that in the cell abbreviations the anode is on the left and the cathode is on the right.

$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{red}} + E^{\circ}_{\text{ox}}$$

Solution: C

Oxidation reaction: $\text{Sn}^{2+} \rightarrow \text{Sn} + 2\text{e}^{-}$ ($E_{\text{ox}} = -0.13 \text{ V}$)

Reduction reaction: $2\text{Fe}^{3+} + 2\text{e}^{-} \rightarrow 2\text{Fe}^{2+}$ ($E_{\text{red}} = 0.77 \text{ V}$)

$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{red}} + E^{\circ}_{\text{ox}}$$

$$E^{\circ}_{\text{cell}} = \mathbf{0.64 \text{ V}}$$

D is incorrect because the question indicates that the reaction is spontaneous in the **forward direction**.

3. Electrochemistry

Question Type: “Cell Diagrams and Cell Potential II”

Thought Process:

The **KEY** to solving this problem is understanding the relationship between cell potential and spontaneity. Spontaneity (ΔG) and cell potential (E°_{cell}) are related by the formula $\Delta G = -nFE^\circ_{\text{cell}}$

A system at equilibrium has a ΔG value of 0, so E°_{cell} must also be 0 in order for this to be true. To solve this problem, simply calculate for the cell potentials using standard cell potentials of redox reactions and the Nernst equation if the reaction does not take place under standard conditions.

For option a) and b), E°_{cell} can be found using the standard redox potentials by plugging them into the formula $E^\circ_{\text{cell}} = E^\circ_{\text{red}} + E^\circ_{\text{ox}}$ where both do not produce a cell potential of 0V.

Although c) appears to be at equilibrium since it's the same reaction at the anode and cathode, Ag^+ at the cathode has a concentration of **0.5M** so $E^\circ_{\text{red}} \neq E^\circ_{\text{ox}}$

Solution: d) $\text{Cu(s)} \mid \text{Cu}^{2+} \text{(aq, 0.175M)} \parallel \text{Ag}^+ \text{(aq, 7.05 x 10}^{-9}\text{M)} \mid \text{Ag(s)}$

$$E_{\text{cell}} = E^\circ_{\text{cell}} - (RT/nF)\ln Q$$

$$E_{\text{cell}} = (0.8 - 0.34) - (0.0257/2) \ln [0.175 / (7.05 \times 10^{-9})^2]$$

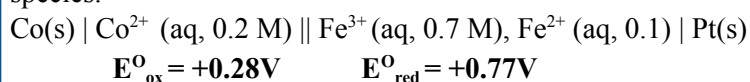
$$E_{\text{cell}} = 0\text{V}$$

3. Electrochemistry

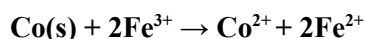
Question Type: “Cell Potential for an Electrochemical Cell”

Thought Process:

For a question such as this one, begin by noting down the oxidation and reduction potential for each species.



Next, write down the balanced equation for this electrochemical cell. Remember that this is a redox reaction, and that electrons have to be balanced in each half reaction such that they cancel out in the total reaction. The one for said example is as follows:



The balanced equation contains the coefficients that are required to solve the problem.

The equation to determine the non-standard cell potential is $E_{\text{cell}} = E_{\text{cell}}^{\circ} - (RT/nF) \ln Q$. We can substitute the appropriate values into our equation and solve accordingly.

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - (RT/nF) \ln Q$$

$$E_{\text{cell}} = (0.28+0.77) - (0.0257/2) \ln [(0.2)(0.1)^2/(0.7)^2]$$

$$E_{\text{cell}} = 1.12\text{V}$$

4. Chemical Kinetics

Question Type: “Catalyst”

Thought Process:

The **KEY** to this question is recognizing that the **activation energy always stays the same** regardless of the temperature.

First, calculate the activation energy using the rate enhancement factor for the reaction 1 (7×10^4)

$$\ln(k_2/k_1) = \Delta E_a/RT$$

$$\ln(7 \times 10^4) = \Delta E_a/(8.314)(322\text{K})$$

$$\Delta E_a = 29866.49\text{J}$$

Substitute the value for ΔE_a into the same equation, but instead of using the rate enhancement factor for reaction 1, use the value given for reaction 2 (10^3). Rearrange the equation to solve for T.

$$\ln(k_2/k_1) = \Delta E_a/RT$$

$$T = E_a/R[\ln(10^3)]$$

$$T = 29866.49/(8.314 \times \ln(10^3))$$

$$\mathbf{T = 520\text{ K}}$$

4. Chemical Kinetics

Question Type: “Rate Laws”

Thought Process:

First, compute for x by dividing rate 3 by rate 1. We do this because $[IV]^y$ is the same for both rates and will therefore **cancel out**. The order of the reactants will almost always be whole numbers or clean fractions (ie. $\frac{1}{2}$). It is reasonable to round them to such numbers for this course.

Determine y using the same process, comparing either reactions 1 or 3 to reaction 2. Note that $[MF]^x$ does NOT cancel out, and you will have to include the MF values in your rate ratio equations. It does not matter which two reactions you choose to compare.

The second last step is using the rate and concentrations from any one run to determine the value of k using the **rate formula**. Since you have already found the values of x and y , you only need to isolate for k .

Once you have found all the unknowns, plug in $[MF] = 0.22 \text{ mol L}^{-1}$ and $[IV] = 0.11 \text{ mol L}^{-1}$ as the respective concentrations within the rate formula and determine the rate.

Solution:

$$x = 0.5$$

$$y = 1$$

$$k = 1.93 \times 10^{-3}$$

$$\text{rate} = 9.95 \times 10^{-5}$$

4. Chemical Kinetics

Question Type: “Half Life”

Thought Process:

The **KEY** to this question is determining the **number of half lives using $t/t_{1/2}$** and then using the half life formula:

$$N_t = N_0(0.5)^{t/t_{1/2}}$$

Since you are given the amount in grams at the start, you can isolate for $N(t)$ and determine how much gas will be present in 44 days.

Solution:

$$t/t_{1/2} = 44\text{days}/4.8\text{days}$$

$$t/t_{1/2} = 9.167$$

$$N(t) = 4000 \text{ mg} \times (0.5^{9.167})$$

$$\mathbf{N(t) = 6.96 \text{ mg}}$$

4. Chemical Kinetics

Question Type: “Plotted Reaction Kinetics”

Thought Process:

A straight line for a plot of $\ln[\text{Ni}_2\text{O}_3]$ would indicate that this reaction follows **first order kinetics**.

To solve for half-life, 2 methods are possible:

1) Integrated rate law

We can first rearrange the equation such that it resembles the equation of a line “ $y = mx + b$ ”

$$\ln[A]_t - \ln[A]_0 = -kt \rightarrow \ln[A]_t = -kt + \ln[A]_0$$

In order to solve for t (half-life), we simply plug in the given values for the initial concentration of NiO ($[A]_0$) and the slope ($-k$). Since we know what the initial concentration of Ni_2O_3 is, we can assume that the half-life concentration would be half of the initial concentration.

$$\ln[A]_t = -kt + \ln[A]_0$$

$$\ln[0.10 \text{ M}]_t = (-0.40 \text{ hour}^{-1})t + \ln[0.20 \text{ M}]_0$$

$$t = 1.7328 \text{ hour}^{-1} \times 3600 \text{ s/hr}$$

$$t = \mathbf{6238.32 \text{ s}}$$

2) Shortcut

Knowing this reaction follows first order kinetics, we can also use the equation $\ln 2/k$ to solve for half-life.

4. Chemical Kinetics

Question Type: “Activation Energy with a Catalyst”

Thought Process:

The **KEY** to this question is to remember that the temperature remains constant. So, the equation used is simple $\ln(\text{catalyzed rate}/\text{uncatalyzed rate}) = \Delta E_a/(RT)$

Input the given values to find the activation energy. Make sure to change the temperature value from °C to Kelvin in order to match the units of the gas constant R.

Solution:

$$\ln(\text{catalyzed rate}/\text{uncatalyzed}) = \Delta E_a/(RT)$$

$$\ln[1.0/(1.3 \times 10^{-3})] = \Delta E_a/(8.314 \times 298.15)$$

$$\Delta E_a = 16.47 \text{ kJ}$$

5. Equilibrium

Question Type: “Buffer Solutions with a Specific pH”

Thought Process:

Write out the reaction based on HF since NaF is simply added to the acid. Then, construct an ice table to illustrate the change in concentration for each reactant/product.

HF	\rightleftharpoons	H ⁺	+	F ⁻
0.95M		0		c
-x		+x		+x
0.95		x		c

In order to solve for the concentration of fluoride ions, we can write an expression for K_a and isolate for $[F^-]$.

So, $K_a = \frac{[H^+][F^-]}{[HF]}$. We know the K_a for this reaction and we are given the concentration of HF in the question. The last important piece of information that we have not used from the problem is the pH. Given the pH, we are able to calculate for $[H^+]$ using the formula $[H^+] = 10^{-pH}$.

Once you have found the concentration of fluoride ions, multiply this value by the volume in order to determine the moles of fluoride consumed. This value is also the number of moles of NaF which is what we want, but how did we come to this conclusion?

If we write out the decomposition reaction for NaF which is $NaF \rightarrow Na^+ + F^-$, we can see that there is a 1:1 ratio between NaF and F^- . Thus, we can conclude that the moles found for F^- equals the moles of NaF.

Finally, multiply the number of moles by its molar mass in order to get the mass of NaF required. The **KEY** to getting the correct solution is multiplying the moles by the **molar mass of NaF** rather than fluoride.

5. Equilibrium

Note: You could reach the same answer using the Henderson-Hasselbalch equation, $\text{pH} = \text{pK}_a + \log([\text{base}]/[\text{acid}])$, isolating for the concentration of the base (F^-) and solving in the same manner as described above. However, you would have to convert K_a to pK_a using the relationship $\text{pK}_a = -\log(K_a)$

Solution: 4.02g

5. Equilibrium

Question Type: “Making Buffer Solutions”

Thought Process:

Recall that for a mixture to be considered a buffer solution, an acid or a base and its conjugate species must be within a **1:10 ratio**. Knowing this, we simply need to calculate the final concentrations of the conjugate acid base pair following the neutralization reactions.

How to approach each option:

Reaction 1- NaOH is the strong base and will break down into its ions Na^+ and OH^- . Na^+ is a spectator ion (since it is the conjugate acid of a strong base) and will therefore not react with the conjugate acid, CH_3COOH . The OH^- which does react will receive an extra proton from CH_3COOH to produce water. From here, construct a reaction table to determine the final concentrations and the ratio between CH_3COOH and CH_3COO^- . The initial concentration of OH^- is the concentration of NaOH and the change in concentration will be all of the $[\text{OH}^-]$ used up.

CH_3COOH	+	OH^-	→	$\text{CH}_3\text{COO}^- + \text{H}_2\text{O}$
0.15 mol		0.15 mol		0
-0.15		-0.15		+0.15
0		0		0.15

Since the ratio is not within 1:10, this reaction will not produce a buffer solution. There is no conjugate acid, and so the final solution is basic. This is rather intuitive, since a **strong** base is being mixed with **equal amounts** of a **weak** acid.

Repeat for the other options. Below are the reactions for each.

2) $\text{CH}_3\text{COO}^- + \text{H}^+ \rightarrow \text{CH}_3\text{COOH}$ (For this reaction, the strong acid is higher in concentration than the weak base, so all the base will be consumed. This is NOT a buffer)

3) $\text{NH}_4^+ + \text{OH}^- \rightarrow \text{NH}_3 + \text{H}_2\text{O}$ (buffer)

5. Equilibrium

4) $\text{CH}_3\text{COOH} + \text{OH}^- \rightarrow \text{CH}_3\text{COO}^- + \text{H}_2\text{O}$ (2 bases cannot produce the conjugate acid, therefore this cannot produce a buffer solution)

5) $\text{NH}_4^+ + \text{OH}^- \rightarrow \text{NH}_3 + \text{H}_2\text{O}$ (buffer)

Solution: 2

5. Equilibrium

Question Type: “Solubility Constant”

Thought Process:

Remember to convert units when approaching this question. Right now, the solubility is given in mg/L and the molar mass is given in g/L.

With any solubility question, you must first find the dissolved concentration in mol/L (we'll denote this as x , as represented in the ICE table below. Once we obtain this value, we simply set up the equation to calculate K_{sp} and plug in the concentration of the compound.

	X_2Z	\rightarrow	$2X$	+	Z
I	c		0		0
C	$-x$		$+2x$		$+x$
E	$c-x$		$2x$		x

Note that the value for c here is irrelevant, and is just used for the sake of completion in the ICE table.

Since solubility is defined as the amount a substance is able to dissolve in a solvent, the equation is written as follows:

$$x = \text{Solubility/MM} = 31.6 \times 10^{-3} / 285.6 = 1.1 \times 10^{-4}$$

$$K_{sp} = [2x]^2[x] = 4x^3$$

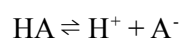
$$K_{sp} = 5.42 \times 10^{-12}$$

5. Equilibrium

Question Type: “Percent Ionization”

Thought Process:

First, develop an equilibrium equation for a weak acid and identify the products and reactants.



Considering we assume the duodenum is neutral in pH, we can assume that our acid will dissociate into **equal amounts of H⁺ and A⁻**. In order to solve for the % ionization, we must find the equilibrium concentration or in other words, the amount of acid dissolved in solution. Because the pH of the solution is given, we can identify the H⁺ concentration at equilibrium, hence find the concentration needed to calculate the % ionization. Once we have this value, the % ionization is simply the quotient of the ionized product to the unionized reactant multiplied by 100.

Solution:

$$x = [\text{H}^+] = 10^{-\text{pH}} = 10^{-3.41} = 3.89 \times 10^{-4}$$

$$\begin{aligned} \% \text{ Ionization} &= x/c, \text{ where } c \text{ is the initial concentration of the weak acid} \\ &= 3.89 \times 10^{-4} / 0.46 \\ &= \mathbf{0.85 \%} \end{aligned}$$

5. Equilibrium

Question Type: “Reaction Mechanisms”

Thought Process:

The **KEY** to this question is to look for the **rate determining step (RDS)** and then write the rate law in terms of its concentrations.

Since step 2 is slow, this will be the RDS and the rate law will be determined based on the reactants in this step.

To isolate for the intermediate in step 2, we must use the **unimolecular rate law** from step 1 in order to isolate for Cl. Recall that **intermediates cannot be expressed in the overall rate law**, since their concentrations are not known directly.

Keep in mind that **ONLY** for an elementary step, the coefficients of the reactants become the exponents in the rate law for that step.

Any fast steps following the RDS are ignored, therefore step 3 can be ignored.

Solution:

$$\text{Rate} = k[\text{CHCl}_3][\text{Cl}]$$

$$\text{From Step 1: } k(\text{forward})[\text{Cl}_2] = k(\text{reverse})[\text{Cl}]^2 \longrightarrow [\text{Cl}] = \sqrt{\frac{K_f[\text{Cl}_2]}{K_r}}$$

Input rearranged expression from Step 1 into Step 2

$$\text{Rate} = \sqrt{\frac{K_f[\text{Cl}_2]}{K_r}} \cdot K[\text{CHCl}_3]$$

$$\text{Rate} = K'[\text{CHCl}_3][\text{Cl}_2]^{1/2}$$

5. Equilibrium

Question Type: “Equilibrium Constant Using Pressure Values”

Thought Process:

Begin the question by writing out a **balanced equation**. For a question such as this one, one must determine the pressures for the reactants and products at the start and end of the reaction. Using the coefficients of the balanced equation, the ‘change in’ or the delta value can be determined. Using this, the end pressures can be deduced as shown below.

Balanced Reaction: $2\text{SO}_3(\text{g}) \rightarrow 2\text{SO}_2(\text{g}) + \text{O}_2(\text{g})$

I	0.75 atm	0	0
C	-2x	+2x	+x
E	0.75 - 2x	2x	x

Using the end pressure values, an equation for the total pressure can be made, as follows:

$$\begin{aligned}\text{Total Pressure} &= (0.75-2x) + 2x + x \\ 0.93 &= (0.75-2x) + 2x + x \\ 0.93 &= 0.75 + x \\ x &= 0.18\end{aligned}$$

Now that we know the value of x, the individual pressures of the reactants and products at equilibrium can be calculated as follows:

$$P_{\text{SO}_3} = 0.75 - 2(0.18) = 0.39$$

$$P_{\text{SO}_2} = 2(0.18) = 0.36$$

$$P_{\text{O}_2} = 0.18$$

Lastly, the equilibrium constant can be calculated as shown below.

$$K = \frac{\text{Products}}{\text{Reactants}}$$
$$K = \frac{(P_{\text{O}_2})(P_{\text{SO}_2})^2}{(P_{\text{SO}_3})^2}$$

5. Equilibrium

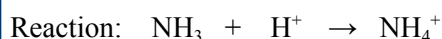
$$K = \frac{(0.18)(0.36)^2}{(0.39)^2}$$

$$K = 0.1534$$

Question Type: “Determining pH using Equilibrium Concentrations”

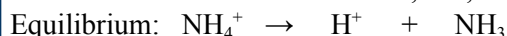
Thought Process:

Begin the question by writing out the chemical reaction and creating an ICE table with the given concentrations included. It is important to note that H_2SO_4 is a strong, diprotic acid and will react with the weak base which is NH_3 . Notice that the initial buffer solution contains conjugate species of acid and base.



0.5mol	0.1mol	0.5mol
-0.1	-0.1	+0.1
0.4mol	0	0.6mol

The total volume of the reaction is **2L** since there are 1L of each solution. Therefore, the concentrations for the above substances are 0.2M, 0M, and 0.3M respectively. ($C=n/V$)



0.30 M	0	0.20 M
-x	+ x	+x
0.30 - x	x	0.20+x

Create a K_a expression by using the equilibrium values acquired from the table above.

$$K_a = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = \frac{x(0.40 + x)}{(0.60 - x)}$$

Using the small x approximation,

5. Equilibrium

$$x = \frac{(0.6)(1 \times 10^{-14})}{(0.4)(1.8 \times 10^{-5})} = 8.33 \times 10^{-10}$$

After solving the following equation for x you are left with a value that equals the $[H^+]$ concentration.

$$x = 8.33 \times 10^{-10} = [H^+]$$

As the $[H^+]$ has been distinguished, the resultant pH can be determined.

$$pH = -\log[H^+]$$

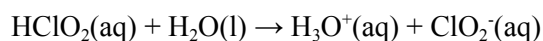
$$pH = 9.08$$

5. Equilibrium

Question Type: “Acid Dissociation Constant Using Concentration Values”

Thought Process:

Begin by writing the balanced equation for said chemical reaction. Next, note down the initial and equilibrium concentrations as shown below:



Initial	0.100	0	0
Equilibrium	0.100-x	x	x

Using the percent ionization and the initial concentration, the concentration of $[\text{ClO}_2^-]$ can be calculated.

$$\% \text{ ionization} = \frac{[\text{ClO}_2^-]}{[\text{HClO}_2] + [\text{ClO}_2^-]}$$

*It is important to note that the initial concentration is 0.100 M and this represents the denominator of the equation.

$$0.281 = \frac{[\text{ClO}_2^-]}{0.100}$$

$$[\text{ClO}_2^-] = 0.0281 = x$$

Based on the equilibrium concentrations, write down the K_a equation. Using the ‘x’ value calculated, we can determine the K_a value while yielding the $\text{p}K_a$ value.

$$K_a = \frac{x^2}{0.100 - x} = \frac{(0.0281)^2}{0.100 - 0.0281}$$

5. Equilibrium

$$K_a = 0.010982$$

$$\mathbf{pK_a = 1.959}$$