

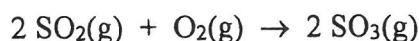
CHEMISTRY 104 – Summer 2025
Hour Exam I Answers

Multiple Choice (3 points each)

- | | | | |
|-----|---|-----|---|
| 1. | A | 14. | B |
| 2. | A | 15. | C |
| 3. | D | 16. | A |
| 4. | E | 17. | D |
| 5. | B | 18. | C |
| 6. | B | 19. | E |
| 7. | E | 20. | C |
| 8. | E | 21. | D |
| 9. | C | 22. | D |
| 10. | C | 23. | A |
| 11. | B | 24. | D |
| 12. | E | 25. | B |
| 13. | C | 26. | E |
| | | 27. | A |

WRITTEN OUT PROBLEMS – Show all work for partial credit.

28. Consider the following reaction and data:
(12 pts.)



$$\Delta H^\circ = \sum \Delta H_{f, \text{prod}}^\circ - \sum \Delta H_{f, \text{react}}^\circ$$

$$\Delta S^\circ = \sum S_{\text{prod}}^\circ - \sum S_{\text{react}}^\circ$$

Compound	ΔH_f°	S°
$\text{SO}_3(\text{g})$	-396 kJ/mol	257 J/K•mol
$\text{SO}_2(\text{g})$	-297 kJ/mol	248 J/K•mol
$\text{O}_2(\text{g})$	0	205 J/K•mol

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

- a) Calculate the value of ΔG° for the reaction at 25°C.

$$\Delta H^\circ = [2(-396)] - [2(-297) + 1(0)] = -198 \text{ kJ}$$

$$\Delta S^\circ = [2(257)] - [2(248) + 1(205)] = -187 \text{ J/K}$$

$$\Delta G^\circ = -198 \text{ kJ} - (298 \text{ K})(-0.187 \text{ kJ/K})$$

$$\Delta G^\circ = -198 \text{ kJ} + 55.7 \text{ kJ} = \boxed{-142 \text{ kJ}}$$

- b) Assuming ΔH° and ΔS° are temperature independent, calculate the temperatures at which this reaction is spontaneous (assuming all gases are at 1 atm).

At equilibrium, $\Delta G^\circ = 0$ (at standard concentrations).

$$\Delta G^\circ = 0 = \Delta H^\circ - T\Delta S^\circ, \quad \Delta H^\circ = T\Delta S^\circ, \quad T = \frac{\Delta H^\circ}{\Delta S^\circ}$$

$$T = \frac{-198 \text{ kJ}}{-0.187 \text{ kJ/K}} = 1060; \text{ this is the temperature where } \Delta G^\circ = 0.$$

Since both ΔH° and ΔS° are negative, this reaction will be spontaneous at (temp) $T < 1060 \text{ K}$ where the favorable ΔH° term will dominate.

- c) At standard conditions, this reaction is allowed to react to reach equilibrium. If, at equilibrium, the partial pressures of $\text{O}_2(\text{g})$ and $\text{SO}_3(\text{g})$ are 0.50 atm and 2.0 atm respectively, calculate the equilibrium partial pressure of $\text{SO}_2(\text{g})$. Note: $T = 25^\circ\text{C}$

$$\Delta G^\circ = -RT \ln K, \quad \ln K = \frac{-\Delta G^\circ}{RT} = \frac{-(-142,000 \text{ J})}{8.3145 \text{ J/(K}\cdot\text{mol)}(298 \text{ K})} = 57.31$$

$$K = e^{57.31} = 7.76 \times 10^{24}$$

$$K = \frac{P_{\text{SO}_3}^2}{P_{\text{SO}_2}^2 \times P_{\text{O}_2}}, \quad 7.76 \times 10^{24} = \frac{(2.0 \text{ atm})^2}{P_{\text{SO}_2}^2 (0.50 \text{ atm})}$$

$$\text{Solving: } P_{\text{SO}_2} = 1.0 \times 10^{-12} \text{ atm}$$

$$\Delta G^\circ = -RT \ln K$$

29. Consider the following equilibrium constant vs. temperature data for some reaction:
(7 pts)

K	Temp
2.54×10^4	109°C
5.04×10^2	225°C
6.33×10^1	303°C
2.25×10^{-1}	412°C
3.03×10^{-3}	539°C

$\left. \begin{array}{l} < 25^\circ\text{C} \\ \} K > 1, \text{ so } \Delta G^\circ < 0 \end{array} \right\}$
 $\left. \begin{array}{l} \\ \} K < 1, \text{ so } \Delta G^\circ > 0 \end{array} \right\}$

- a) Is this reaction spontaneous at standard concentrations and $T = 25^\circ\text{C}$? Explain.

From data, as temperatures decrease from 303°C , K increases to larger numbers greater than 1. So at $T = 25^\circ\text{C}$, $K > 1$ which means $\Delta G^\circ < 0$. So yes, this reaction is spontaneous at $T = 25^\circ\text{C}$ (assuming standard concentrations).

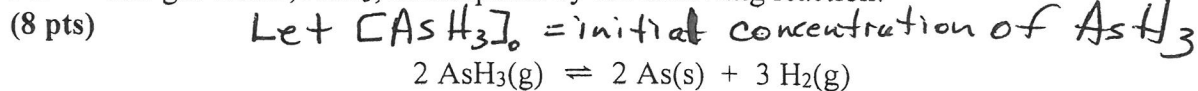
- b) Predict the signs of ΔH° and ΔS° for the reaction. Explain.

As temperature increases, note that the data shows that K goes from a number greater than 1 to a value less than 1. This indicates that as temperature increases, ΔG° goes from a negative value to a positive value.

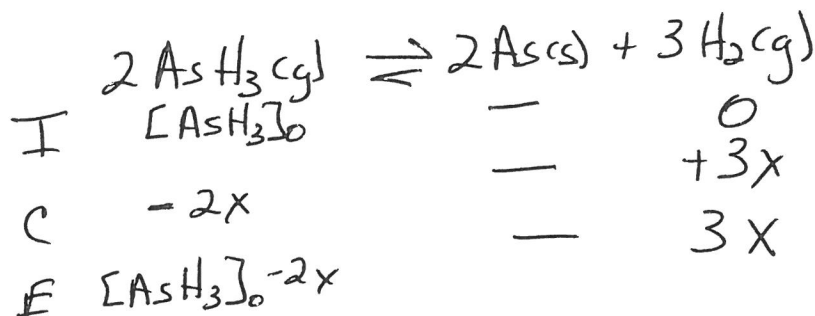
$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

This can only be explained if both ΔH° and ΔS° are negative values. When this is the case, ΔG° is negative only below some temperature (as we have here).

30. The gas arsine, AsH_3 , decomposes by the following reaction:



In an experiment, pure $\text{AsH}_3(\text{g})$ of unknown initial concentration is placed into a 2.0 L container. After equilibrium is reached, 6.0 moles of $\text{H}_2(\text{g})$ is produced and 12.0 moles of $\text{AsH}_3(\text{g})$ is also present. Determine the initial concentration of $\text{AsH}_3(\text{g})$ before the reaction took place ($[\text{AsH}_3]_{\text{initial}} = ?$). Also, calculate the value of the equilibrium constant for this reaction ($K = ?$). Show all work for credit.



From the problem:

$$[\text{H}_2]_e = \frac{6.0 \text{ mol}}{2.0 \text{ L}} = 3.0 \text{ M} = 3x, \quad x = 1.0 \text{ M}$$

$$[\text{AsH}_3]_e = [\text{AsH}_3]_0 - 2x = \frac{12.0 \text{ mol}}{2.0 \text{ L}}$$

$$[\text{AsH}_3]_0 - 2x = 6.0 \text{ M}, \quad [\text{AsH}_3]_0 - 2(1.0) = 6.0 \text{ M}$$

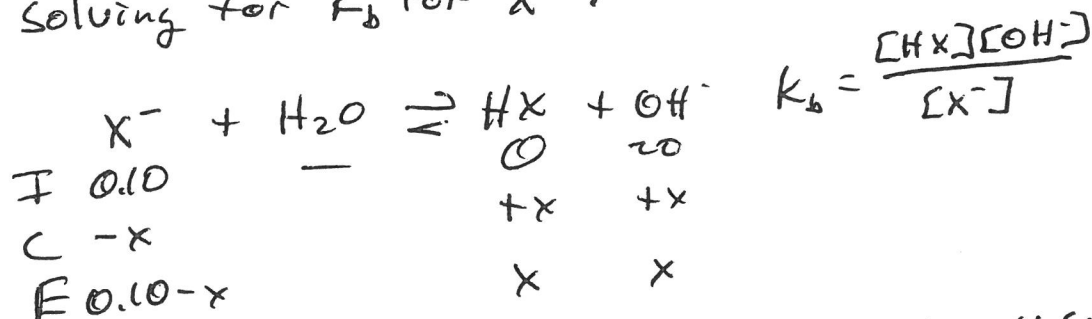
$$[\text{AsH}_3]_0 = 6.0 + 2.0 = \boxed{8.0 \text{ M}}$$

$$K = \frac{[\text{H}_2]^3}{[\text{AsH}_3]^2} = \frac{(3.0)^3}{(6.0)^2} = \frac{27}{36} = \boxed{0.75 = K}$$

31. A 0.10 M solution of the salt NaX has a pH of 11.50. Calculate the pH of a 1.0 M solution of HX.

(6 pts)

X^- must be a weak base, so HX is a weak acid. In order to solve for the pH of HX solution, we need to determine K_a for HX. we will do this by solving for K_b for X^- .

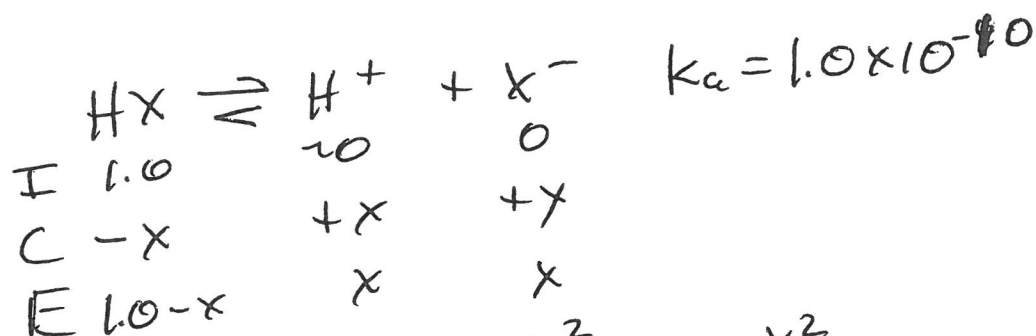


From problem, pH = 11.50, so pOH = 14 - 11.50 = 2.50.

$$[OH^-] = x = 10^{-2.50} = 3.16 \times 10^{-3} M$$

$$K_b = \frac{x^2}{0.10-x} = \frac{(3.16 \times 10^{-3})^2}{0.10-3.16 \times 10^{-3}}, \quad K_b = 1.0 \times 10^{-4}$$

$$\text{So } K_a \text{ for HX} = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-4}} = 1.0 \times 10^{-10}$$



$$K_a = 1.0 \times 10^{-10} = \frac{x^2}{1.0-x} \approx \frac{x^2}{1.0}$$

Solving: $x = 1.0 \times 10^{-5} M$

Assumptions good.

$$[H^+] = 1.0 \times 10^{-5} M, \quad pH = -\log(1.0 \times 10^{-5}) = \boxed{5.00}$$

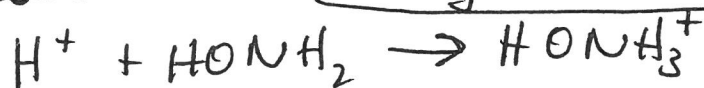
32. Consider 1.0 L of a solution composed of 1.60 M HONH_2 ($K_b = 1.1 \times 10^{-8}$) and 0.80 M HONH_3NO_3 . This is a buffer solution containing a weak base (HONH_2) and its conjugate acid (HONH_3^+).
(8 pts) a) Calculate the pH of this solution.

$$\text{pH} = \text{pK}_a + \log \frac{[\text{base}]}{[\text{acid}]} = -\log \frac{1 \times 10^{-14}}{1.1 \times 10^{-8}} + \log \left(\frac{1.60}{0.80} \right)$$

$$\text{pH} = 6.04 + \log 2 = 6.04 + 0.30 = \boxed{6.34}$$

- b) In order for this 1.0 L of solution to have $\text{pH} = \text{pK}_a$, would you add HCl or KOH? What quantity (moles) of HCl or KOH would you add to the solution in order to get $\text{pH} = \text{pK}_a$?

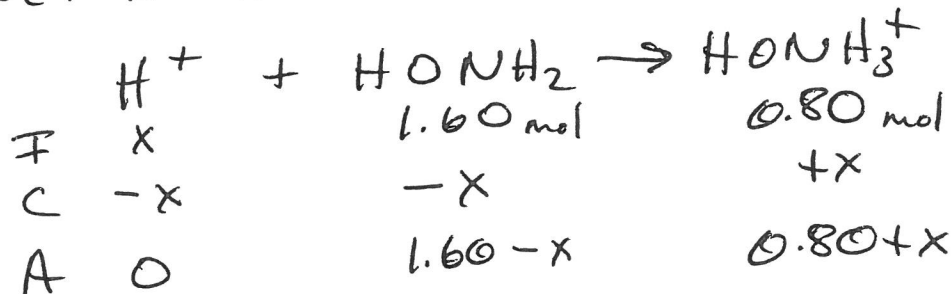
In order for $\text{pH} = \text{pK}_a$, we need a solution where $[\text{HONH}_2] = [\text{HONH}_3^+]$. We need to lower the HONH_2 concentration and raise the HONH_3^+ concentration. This will occur if ~~we~~ we add strong acid (HCl):



$$\text{mol HONH}_2 = 1.0 \text{ L} \times 1.60 \frac{\text{mol}}{\text{L}} = 1.60 \text{ mol}$$

$$\text{mol HONH}_3^+ = 1.0 \text{ L} \times 0.80 \frac{\text{mol}}{\text{L}} = 0.80 \text{ mol}$$

Let $x = \text{mol H}^+$ added from HCl



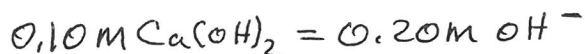
We want $[\text{HONH}_2] = [\text{HONH}_3^+]$ which will occur when $\text{mol HONH}_2 = \text{mol HONH}_3^+$.

$$1.60 - x = 0.80 + x, \quad 2x = 0.80, \quad x = 0.40 \text{ mol H}^+$$

0.40 mol HCl must be added to achieve $\text{pH} = \text{pK}_a$.

For a strong acid titrated by a strong base,
 $pH = 7.00$ at the equivalence point.

CHEMISTRY 104 Volume $Ca(OH)_2$ to reach equiv. pt = 100 mL Summer 2025
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33. Consider the titration of 200.0 mL of 0.10 M HNO_3 titrated by 0.10 M $Ca(OH)_2$. Sketch the titration curve for this titration. On your plot, label the axis and indicate the pH and volume of the equivalence point. Also indicate the initial pH before any $Ca(OH)_2$ has been added as well as the pH at 150.0 mL of $Ca(OH)_2$ added.

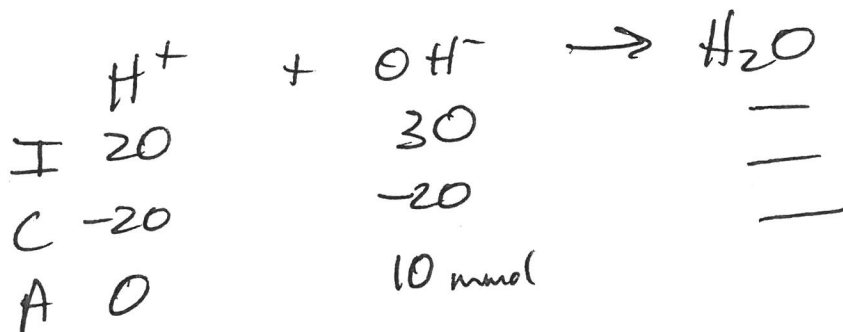
(9 pts)

Initial pH: 0.10 M HNO_3
 $[H^+] = 0.10 M$, $pH = -\log(0.10) = 1.00$

pH at 150.0 mL $Ca(OH)_2$ added:

$$200.0 mL (0.10 M) = 20 \text{ mmol } H^+ \text{ initially}$$

$$150.0 mL (0.20 \frac{\text{mol } OH^-}{L}) = 30 \text{ mmol } OH^- \text{ added}$$



$$[OH^-]_{\text{excess}} = \frac{10 \text{ mmol } OH^-}{200.0 \text{ mL} + 150.0 \text{ mL}} = 2.86 \times 10^{-2} M$$

$$pOH = -\log(2.86 \times 10^{-2}) = 1.54, pH = 14 - 1.54 = 12.46$$

In order to reach the equivalence point, we must add 20 mmol OH^- , which occurs at 100.0 mL $Ca(OH)_2$ added.

