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

$$2\ SO_2(g) + O_2(g) \rightarrow 2\ SO_3(g)$$

$\Delta H^\circ = \Sigma \Delta H^\circ_{f,prod} - \Sigma \Delta H^\circ_{f,react}$

$\Delta S^\circ = \Sigma S^\circ_{prod} - \Sigma S^\circ_{react}$

Compound	ΔH_f°	S°
$SO_3(g)$	–396 kJ/mol	257 J/K•mol
$SO_2(g)$	–297 kJ/mol	248 J/K•mol
$O_2(g)$	? 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.

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

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

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

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

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).

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

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

Since both ΔH° and ΔS° are negative, this reaction will be spontaneous (at $\Delta G^\circ < 0$) at $\boxed{T < 1060\ K}$ 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 $O_2(g)$ and $SO_3(g)$ are 0.50 atm and 2.0 atm respectively, calculate the equilibrium partial pressure of $SO_2(g)$. Note: T = 25°C

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

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

$K = \frac{P_{SO_3}^2}{P_{SO_2}^2 \times P_{O_2}}$, $7.76 \times 10^{24} = \frac{(2.0\ atm)^2}{P_{SO_2}^2 (0.50\ atm)}$

Solving: $\boxed{P_{SO_2} = 1.0 \times 10^{-12}\ atm}$

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

(Temp < 25°C) } K > 1, so $\Delta G^\circ < 0$

} K < 1, so $\Delta G^\circ > 0$

a) Is this reaction spontaneous at standard concentrations and T = 25°C? Explain.

From data, as temperatures decrease from 303°C, K increases to larger numbers greater than 1. So at T = 25°C, K > 1 which means $\Delta G^\circ < 0$. So yes, this reaction is spontaneous at T = 25°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. **(8 pts)** The gas arsine, AsH_3, decomposes by the following reaction:

Let $[AsH_3]_o$ = initial concentration of AsH_3

$$2\ AsH_3(g) \rightleftharpoons 2\ As(s) + 3\ H_2(g)$$

In an experiment, pure $AsH_3(g)$ of unknown initial concentration is placed into a 2.0 L container. After equilibrium is reached, 6.0 moles of $H_2(g)$ is produced and 12.0 moles of $AsH_3(g)$ is also present. Determine the initial concentration of $AsH_3(g)$ before the reaction took place ($[AsH_3]_{initial}$ = ?). Also, calculate the value of the equilibrium constant for this reaction (K = ?). Show all work for credit.

	$2\ AsH_3(g)$	$\rightleftharpoons$	$2\ As(s)$	$+\ 3\ H_2(g)$
I	$[AsH_3]_o$	—	0	
C	$-2x$	—	$+3x$	
E	$[AsH_3]_o - 2x$	—	$3x$	

From the problem:

$$[H_2]_e = \frac{6.0\ mol}{2.0\ L} = 3.0\ M = 3x, \quad x = 1.0\ M$$

$$[AsH_3]_e = [AsH_3]_o - 2x = \frac{12.0\ mol}{2.0\ L}$$

$$[AsH_3]_o - 2x = 6.0\ M, \quad [AsH_3]_o - 2(1.0) = 6.0\ M$$

$$[AsH_3]_o = 6.0 + 2.0 = \boxed{8.0\ M}$$

$$K = \frac{[H_2]^3}{[AsH_3]^2} = \frac{(3.0)^3}{(6.0)^2} = \frac{27}{36} = \boxed{0.75 = K}$$

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^-.

$$X^- + H_2O \rightleftharpoons HX + OH^- \qquad K_b = \frac{[HX][OH^-]}{[X^-]}$$

	X^-		HX	OH^-
I	0.10	—	0	0
C	−x		+x	+x
E	0.10−x		x	x

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

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

$$K_b = \frac{x^2}{0.10-x} = \frac{(3.16\times10^{-3})^2}{0.10-3.16\times10^{-3}}, \quad K_b = 1.0\times10^{-4}$$

So K_a for $HX = \frac{K_w}{K_b} = \frac{1.0\times10^{-14}}{1.0\times10^{-4}} = 1.0\times10^{-10}$

$$HX \rightleftharpoons H^+ + X^- \qquad K_a = 1.0\times10^{-10}$$

	HX		H^+	X^-
I	1.0		0	0
C	−x		+x	+x
E	1.0−x		x	x

$$K_a = 1.0\times10^{-10} = \frac{x^2}{1.0-x} \approx \frac{x^2}{1.0}$$

Solving: $x = 1.0\times10^{-5}\ M$

Assumptions good.

$[H^+] = 1.0\times10^{-5}\ M$, $pH = -\log(1.0\times10^{-5}) = 5.00$

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.

$$pH = pK_a + \log\frac{[base]}{[acid]} = -\log\frac{1\times10^{-14}}{1.1\times10^{-8}} + \log\left(\frac{1.60}{0.80}\right)$$

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

b) In order for this 1.0 L of solution to have pH = 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 pH = pK_a?

In order for $pH = pK_a$, we need a solution where $[HONH_2] = [HONH_3^+]$. We need to lower the $HONH_2$ concentration and raise the $HONH_3^+$ concentration. This will occur if we add strong acid (HCl):

$$H^+ + HONH_2 \rightarrow HONH_3^+$$

$$\text{mol } HONH_2 = 1.0\text{ L} \times 1.60\ \tfrac{\text{mol}}{\text{L}} = 1.60\text{ mol}$$

$$\text{mol } HONH_3^+ = 1.0\text{ L} \times 0.80\ \tfrac{\text{mol}}{\text{L}} = 0.80\text{ mol}$$

Let x = mol H^+ added from HCl

	H^+	+ $HONH_2$	→ $HONH_3^+$
I	x	1.60 mol	0.80 mol
C	−x	−x	+x
A	0	1.60 − x	0.80 + x

We want $[HONH_2] = [HONH_3^+]$ which will occur when mol $HONH_2$ = mol $HONH_3^+$.

$1.60 - x = 0.80 + x$, $2x = 0.80$, $x = 0.40$ mol H^+

0.40 mol HCl must be added to achieve $pH = pK_a$.

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

Volume $Ca(OH)_2$ to reach equiv. pt = 1000 mL

$0.10\ M\ Ca(OH)_2 = 0.20\ M\ OH^-$

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:

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

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

	H^+	+ OH^-	→ H_2O
I	20	30	—
C	−20	−20	—
A	0	10 mmol	—

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

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

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