

Part 1. Short Answer [2 marks each]

1. What are three ways we have of predicting whether a reaction will be spontaneous or not?

Spontaneous means $\Delta G < 0$
 $\Delta S_{univ} > 0$ $K > 1$

-1 for any missed

2. What is the standard state for ions in aqueous solution?

1 Molar concentration (2)

3. Define enthalpy of formation.

Δ_{fH} = ΔH of the reaction forming a compound from its constituent elements in their standard states (1)

4. For each pair listed below, select which has the greater entropy.

$\text{H}_2\text{O}(l)$ vs $\text{H}_2\text{O}(g)$

$\text{O}_2(g)$ at 298 K vs $\text{O}_2(g)$ at 400 K

butane vs cyclo-butane

$\text{NaCl}(s)$ vs $\text{NaCl}(aq)$

1.5 each

5. Define a Bronsted base.

A Bronsted Base is a proton acceptor

6. For the following reactions, identify the acid and its conjugate base.



↑
acid
0

↑
conj. base
1

Part 2.

1: The reaction $I_2(g) \rightleftharpoons 2I(g)$ was studied by sealing solid iodine, I_2 , in a glass bulb and measuring the total equilibrium pressure in the bulb at various temperatures (the I_2 completely sublimes). The following data was obtained:

Volume of the bulb = 249.8 ml; original mass of $I_2 = 0.497$ g; $MW(I) = 127$ g/mol

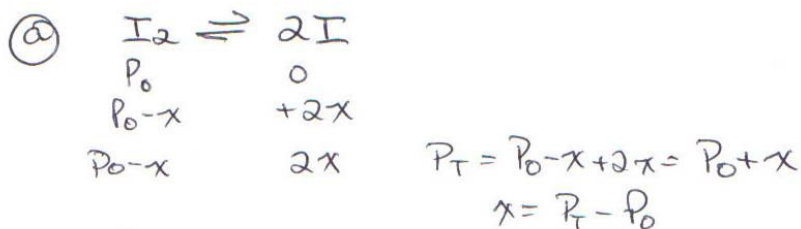
$T(^{\circ}C)$	Total Equilibrium Pressure (atm)
800	0.734
1000	0.984

(a) show that the equilibrium pressure of $I_2(g)$ is $2P_0 - P_T$ and that of $I(g)$ is $2(P_T - P_0)$, where P_0 is the original pressure of $I_2(g)$ (at the particular temperature in question, when no decomposition occurs) and P_T is the total pressure in the bulb at equilibrium. Assume ideal gas behaviour throughout. [3]

(b) Calculate K_p for each temperature. [2]

(c) Calculate ΔH° for the reaction in the temperature range $800^{\circ}C$ to $1000^{\circ}C$ (assume ΔH° is independent of temperature). [3]

(d) Calculate the free energy and entropy change for the reaction at $800^{\circ}C$. [4]



Equl. $(I_2) P_0 - (P_T - P_0) = 2P_0 - P_T$
 $(I) 2(P_T - P_0)$

(b) $K = \frac{[2(P_T - P_0)]^2}{2P_0 - P_T}$

$K_{800} = 1.20 \times 10^{-2}$

$K_{1000} = 1.64 \times 10^{-1}$

$P_0 = \frac{nRT}{V}$ at each T

$n = \frac{0.497g}{2(127g/mol)} = 1.96 \times 10^{-3}$ moles

$P_0(800^{\circ}C) = \frac{1.96 \times 10^{-3} \text{ mol} (R)(1073K)}{0.2498L}$

= 0.69 atm.

$P_0(1000^{\circ}C) = 0.82 \text{ atm.}$

(c) $\ln \frac{K_2}{K_1} = -\frac{\Delta H^{\circ}}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$

$\Delta H^{\circ} = 148 \text{ kJ/mol.}$

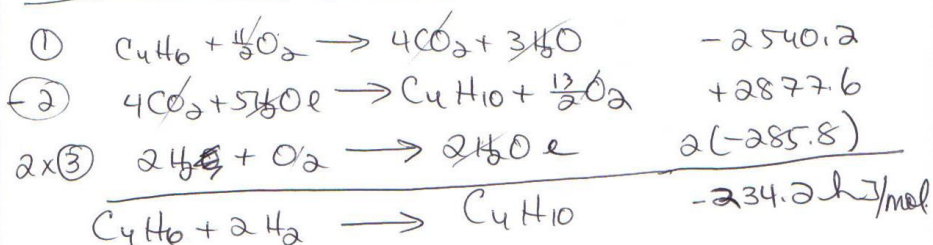
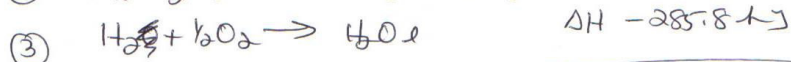
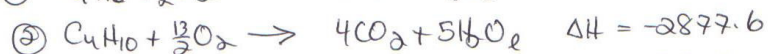
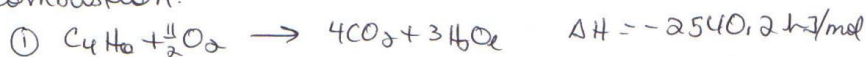
(d) at 1073K. $\Delta G^{\circ} = -RT \ln K$
 = 39.5 kJ/mol

$\Delta S^{\circ} = \frac{\Delta H^{\circ} - \Delta G^{\circ}}{T}$
 = 101 J/Kmol

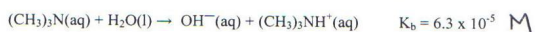
2: The standard enthalpies of combustion of 1,2-butadiene ($C_4H_6(g)$), butane ($C_4H_{10}(g)$) and $H_2(g)$ are -2540.2 kJ/mol , -2877.6 kJ/mol and -285.8 kJ/mol , respectively (in each case $CO_2(g)$ and $H_2O(l)$ are formed). Use this data to calculate the enthalpy of hydrogenation of butadiene, for which the unbalanced equation is: [12 marks]



combustion.



3. What must be the molarity of an aqueous solution of trimethylamine ($(CH_3)_3N$) if it has a $pH = 11.12$? (Hint, you will need to do your ICE table backwards!). [8 marks]



$$pH = 11.12 \quad pOH = 14 - 11.12 = 2.88$$

$$[OH^-] = 0.0013 \text{ M}$$

$$K_b = \frac{[OH^-][CH_3)_3NH^+]}{[(CH_3)_3N]} = 6.3 \times 10^{-5} \text{ M}$$

	$(CH_3)_3N$	OH^-	$(CH_3)_3NH^+$
I	x	0	0
C	x - 0.0013	0.0013	0.0013
E	0.027 M	0.0013	0.0013

at equil. $[OH^-] = [(CH_3)_3NH^+]$

$$K_b = \frac{(0.0013 \text{ M})^2}{[(CH_3)_3N]_{eq}}$$

$$[(CH_3)_3N]_{eq} = 0.027 \text{ M}$$

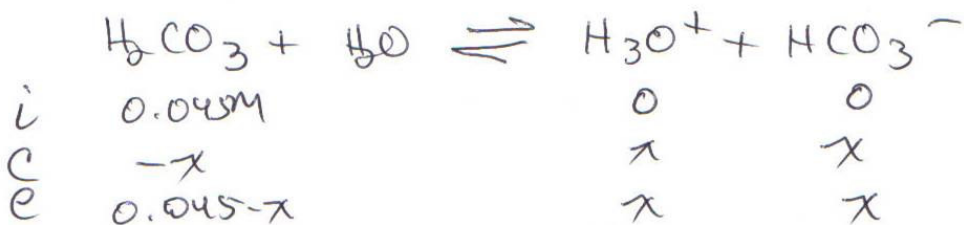
$$\therefore x = 0.027 + 0.0013 = \boxed{0.028 \text{ M}}$$

4. For 0.045 M H_2CO_3 , a weak diprotic acid, calculate (a) $[\text{H}_3\text{O}^+]$, (b) $[\text{HCO}_3^-]$ and (c) $[\text{CO}_3^{2-}]$.

[14 marks]

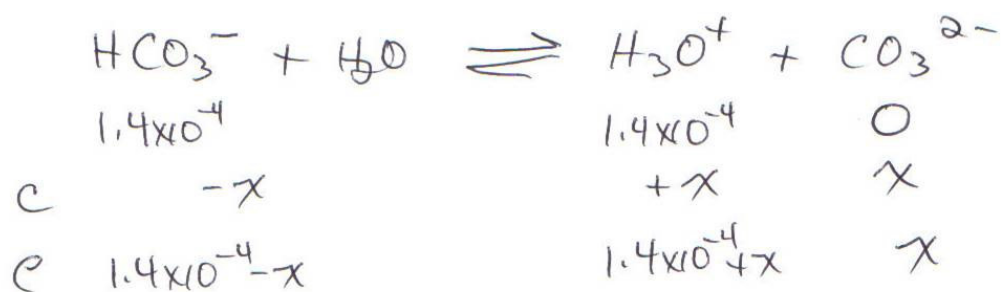
$$K_{a1} = 4.4 \times 10^{-7}, K_{a2} = 4.7 \times 10^{-11}$$

Start with 1st dissociation



$$K = \frac{[\text{H}_3\text{O}^+][\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]} = \frac{x^2}{0.045-x} = 4.4 \times 10^{-7} \text{ M}$$

solving. either by assumption or quadratic eq -
 $x = 1.4 \times 10^{-4} \text{ M}$



$$K = \frac{(1.4 \times 10^{-4} + x)(x)}{1.4 \times 10^{-4} - x} = 4.7 \times 10^{-11}$$

solving $x = 4.7 \times 10^{-11}$ ← you may also get "0"

so, $[\text{H}_3\text{O}^+] = 1.4 \times 10^{-4} \text{ M}$, $[\text{HCO}_3^-] = 1.4 \times 10^{-4} \text{ M}$

$$[\text{CO}_3^{2-}] = 4.7 \times 10^{-11} \text{ M}$$

↑ also "0"