Chemistry 192
Exam 1
Spring, 2018
Solutions

R =8.3144 J mol^{-1} K^{-1}

R=.0821 L atm mol^{-1} K^{-1}

R=.08314 L bar mol^{-1} K^{-1}

$N_A = 6.022 \times 10^{23}$ molecules mol^{-1}

$T = t + 273.15$

$[H_3O^+][OH^-] = 1.0 \times 10^{-14}$

Name:

1. At 1330K, the pressure equilibrium constant for the gas-phase reaction

$$GeW_2O_{7(g)} \rightleftharpoons GeO_{(g)} + W_2O_{6(g)}$$

is $K_P = 2.6 \times 10^{-3}$. If 110. g of gas-phase GeW_2O_7 are placed in a 7.00 L flask of fixed volume at 1330K, calculate the total pressure in the flask when equilibrium is reached. Approximations do <u>not</u> work for this problem. [Hint: Calculate and use K_C.] (33 Points)

Answer:

$$n = \frac{110.\text{ g}}{(72.6 + 2[183.8] + 7[16.0])\text{ g mol}^{-1}} = 0.199\text{ mol}$$

	$n_{GeW_2O_7}$	n_{GeO}	$n_{W_2O_6}$
initial	0.199	0	0
change	-x	x	x
equilibrium	(0.199 -x)	x	x

$$n_{tot} = 0.199 + x$$

$$K_C = K_P(RT)^{-\Delta n_{gas}} = (2.6 \times 10^{-3})[(0.083144)(1330)]^{-1} = 2.4 \times 10^{-5}$$

$$2.4 \times 10^{-5} = \frac{[\text{GeO}][\text{W}_2\text{O}_6]}{[\text{GeW}_2\text{O}_7]} = \frac{(x/7.00)^2}{(0.199 - x)/7.00}$$

$$x^2 + 1.6 \times 10^{-4}x - 3.3 \times 10^{-5} = 0$$

$$x = \frac{-1.6 \times 10^{-4} \pm [(1.6 \times 10^{-4})^2 + 4(3.3 \times 10^{-5})]^{1/2}}{2} = 5.7 \times 10^{-3}\text{ mol (ignoring the negative solution)}$$

$$P = \frac{nRT}{V} = \frac{(0.199 + x)RT}{V} = \frac{(0.205\text{ mol})(0.081344\text{ L bar mol}^{-1}\text{K}^{-1})(1330.\text{ K})}{7.00\text{ L}} = 3.23\text{ bar}$$

Name:

2. Calculate pK_b for the reaction of propylamine with water

$$CH_3(CH_2)_2NH_{2(aq)} + H_2O_{(\ell)} \rightleftharpoons CH_3(CH_2)_2NH^+_{3(aq)}OH^-_{(aq)}$$

given the pH of an aqueous 1.0 M propylammonium [$CH_3(CH_2)_2NH_3^+$] solution is 5.25. Approximations work for this problem. (33 Points)

Answer:

$$CH_3(CH_2)_2NH^+_{3(aq)} + H_2O_{(\ell)} \rightleftharpoons CH_3(CH_2)_2NH_{2(aq)} + H_3O^+_{(aq)}$$

$$[H_3O^+] = 10^{-5.25} = 5.6 \times 10^{-6}$$

	$[CH_3(CH_2)_2NH_3^+]$	$[CH_3(CH_2)_2NH_2]$	$[H_3O^+]$
initial	1.0	0	0
change	-5.6×10^{-6}	5.6×10^{-6}	5.6×10^{-6}
equilibrium	1.0	5.6×10^{-6}	5.6×10^{-6}

$$K_a = \frac{(5.6 \times 10^{-6})^2}{1.0} = 3.1 \times 10^{-11}$$

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{3.1 \times 10^{-11}} = 3.2 \times 10^{-4} \qquad pK_b = -\log_{10}(K_b) = 3.49$$

Name:

3. A buffer is prepared by mixing 0.35 moles of benzoic acid (C_6H_5COOH) and 0.45 moles of the benzoate anion ($C_6H_5COO^-$) with water to make a solution of total volume 0.35 L. The buffer is then combined with 0.10 L of 0.10 M potassium hydroxide (KOH, a strong base). Given $K_a = 6.3 \times 10^{-5}$ for benzoic acid, calculate the final pH of the mixture of buffer and potassium hydroxide. Approximations work for this problem. (34 Points)

Answer:

$$C_6H_5COOH_{(aq)} + H_2O_{(\ell)} \rightleftharpoons C_6H_5COO^-_{(aq)} + H_3O^+_{(aq)}$$

$$\text{p}K_a = -\log_{10}(6.3 \times 10^{-5}) = 4.20$$

$$n_{OH^-} = (0.10 \text{ mol L}^{-1})(0.10 \text{ L}) = 1.0 \times 10^{-2} \text{ mol}$$

Method 1: Henderson-Hasselbalch

$$\text{pH} = \text{p}K_a + \log_{10}\frac{[C_6H_5COO^-]}{[C_6H_5COOH]} = 4.20 + \log_{10}\frac{(0.45 + 1.0 \times 10^{-2})/V}{(0.35 - 1.0 \times 10^{-2})/V} = 4.33$$

Method 2: ICE Table

	$n_{C_6H_5COOH}$	$n_{C_6H_5COO^-}$	$n_{H_3O^+}$
initial	0.35	0.45	0
change	-1.0×10^{-2}	1.0×10^{-2}	y
equilibrium	0.34	0.46	y

$$6.3 \times 10^{-5} = \frac{y(0.46)}{0.34} \quad y = [H_3O^+] = 4.7 \times 10^{-5} \text{ M} \quad \text{pH} = 4.33$$

Name:

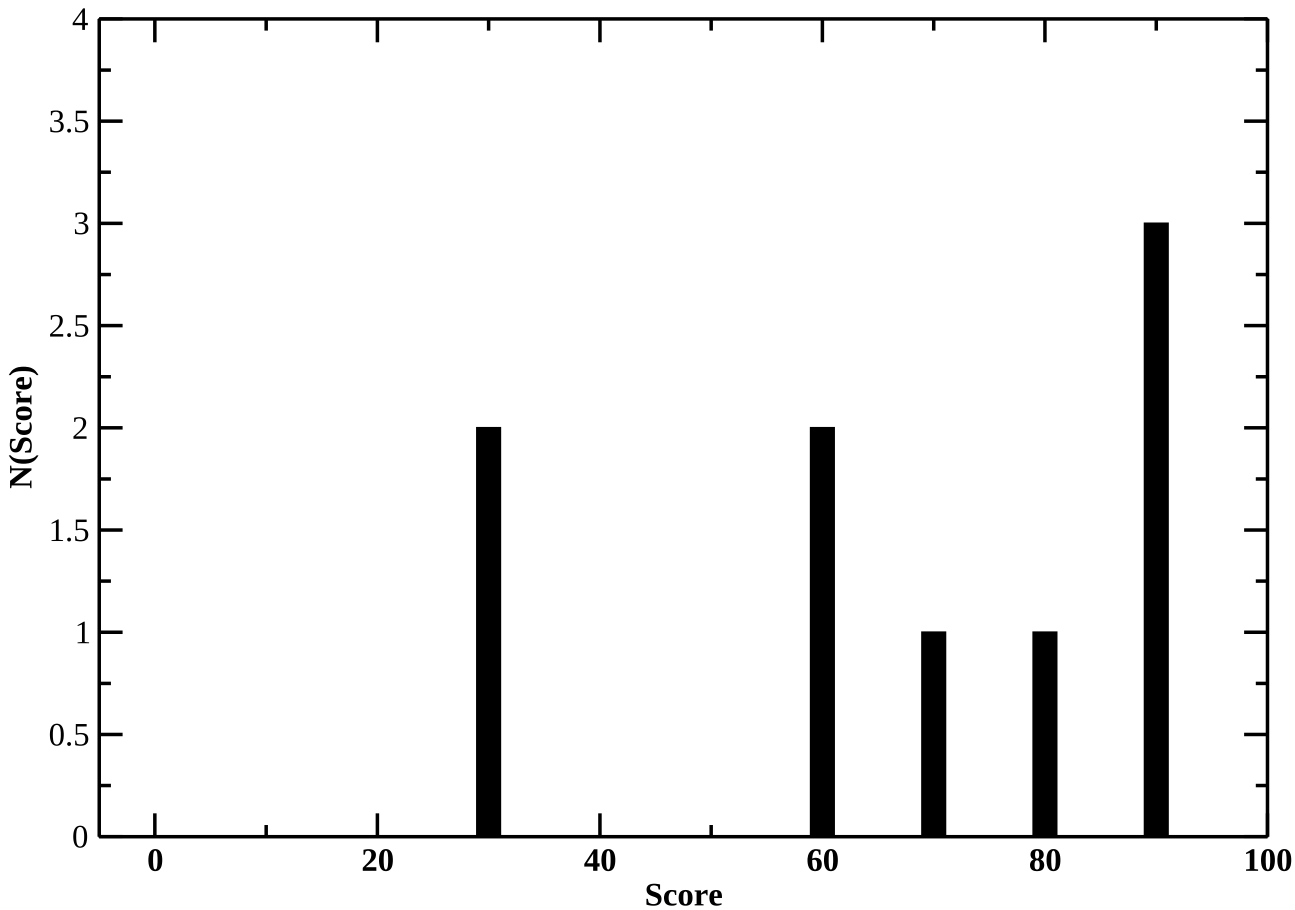

Figure 1: **High = 100, Median = 72, Mean = 72**