Chemistry 192
Practice 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. The equilibrium degree of dissociation of the gas-phase reaction

$$\mathrm{NOBr}_{(g)} \rightleftharpoons \mathrm{NO}_{(g)} + \frac{1}{2}\mathrm{Br}_{2(g)}$$

at 298. K and a total pressure of 1.00 bar is found to be $\alpha = 0.24$. A gas-phase mixture of NOBr, NO and Br_2 are placed in a container of fixed volume such that [NOBr]=2.53 M, [NO]=0.753 M and [Br_2]=1.57 M. Under these initial conditions, do the necessary calculations and predict whether the reaction will proceed spontaneously to the right or to the left.
Answer:

	n_{NOBr}	n_{NO}	n_{Br_2}
initial	n_0	0	0
change	$-\alpha n_0$	αn_0	$\alpha n_0/2$
final	$(1-\alpha)n_0$	αn_0	$\alpha n_0/2$

$$n_{tot} = n_0(1+\alpha/2)$$

$$K_P = \frac{\left(\frac{\alpha}{1+\alpha/2}P\right)\left(\frac{\alpha/2}{1+\alpha/2}P\right)^{1/2}}{\frac{1-\alpha}{1+\alpha/2}P} = 0.10$$

$$K_C = K_P(RT)^{-\Delta n_{gas}} = (0.10)[(0.08314)(298)]^{-1/2} = 0.020$$

$$Q_C = \frac{[\mathrm{NO}][\mathrm{Br}_2]^{1/2}}{[\mathrm{NOBr}]} = \frac{(0.753)(1.57)^{1/2}}{2.53} = 0.37$$

$$Q_C > K_C \qquad \text{to left}$$

Name:

2. The acid ionization constant of chlorous acid according to the reaction

$$HClO_{2(aq)} + H_2O_{(\ell)} \rightleftharpoons ClO^-_{2(aq)} + H_3O^+_{(aq)}$$

is $K_a = 1.1 \times 10^{-2}$. Calculate the pH of an aqueous 0.0230 M $KClO_2$ solution given potassium chlorite is completely soluble in water. Approximations work in the solution to this problem.
Answer:

$$ClO^-_{2(aq)} + H_2O_{(\ell)} \rightleftharpoons HClO_{2(aq)} + OH^-_{(aq)}$$

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{1.1 \times 10^{-2}} = 9.1 \times 10^{-13}$$

	$[ClO_2^-]$	$[HClO_2]$	$[OH^-]$
initial	0.0230 M	0 M	0 M
change	$-y$ M	y M	y M
equilibrium	$0.0230 - y$ M	y M	y M

$$\frac{y^2}{0.0230 - y} \approx \frac{y^2}{0.0230} = 9.1 \times 10^{-13}$$

$$y = [OH^-] = 1.4 \times 10^{-7}$$

$$pOH = -\log_{10}(1.4 \times 10^{-7}) = 6.84 \qquad pH = 14.00 - pOH = 7.16$$

Name:

3. The volume of a buffer solution that is made from 0.00100 moles of benzoic acid (C_6H_5COOH) and 0.00200 moles of benzoate anion ($C_6H_5COO^-$) is 0.100 L. Calculate the pH of a solution formed by mixing the buffer with 0.0500 L of 0.0100 M HCl. You are given the acid ionization constant of benzoic acid is 6.3×10^{-5}, and you can assume the standard approximations work for this system.

Answer:

Method 1

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

$$pK_a = -\log_{10}(6.3 \times 10^{-5}) = 4.20$$

$$pH = pK_a + \log_{10} \frac{[C_6H_5COO^-]}{[C_6H_5COOH]}$$

$$n_{H_3O^+} = (0.0100 \text{ mol L}^{-1})(0.0500 \text{ L}) = 5.00 \times 10^{-4} \text{ mol}$$

$$pH = 4.20 + \log_{10} \frac{\frac{0.00200 - 5.00 \times 10^{-4}}{0.150}}{\frac{0.00100 + 5.00 \times 10^{-4}}{0.150}} = 4.20$$

Method 2, ICE Table

	$n_{C_6H_5COOH}$	$n_{C_6H_5COO^-}$	$n_{H_3O^+}$
initial	0.00100	0.00200	0
change	5.00×10^{-4}	-5.00×10^{-4}	y
equilibrium	0.0015	0.0015	y

$$\frac{y(0.0015/.150)}{0.0015/.150} = 6.3 \times 10^{-5}$$

$$y = [H_3O^+] = 6.3 \times 10^{-5} \text{ M}$$

$$pH = -\log_{10}(6.3 \times 10^{-5}) = 4.20$$

Name: