

## Chemical equilibrium

- P46. Determine the direction in which the  $2 \text{NH}_3 \rightleftharpoons \text{N}_2 + 3 \text{H}_2$  reaction goes in the system where the partial pressure of ammonia is 10.0 bar, the partial pressures of nitrogen and hydrogen are 20.0 bar. The equilibrium constant for the reaction is  $K_p = 4370$ . *[The reaction takes place in the direction of the upper arrow.]*
- P47. Calculate the equilibrium constant for the reaction  $3 \text{O}_2 \rightleftharpoons 2 \text{O}_3$  at 25 °C temperature and 1.00 bar pressure, using the following thermodynamic data: standard formation entropy of oxygen is  $205.1 \text{ J mol}^{-1} \text{ K}^{-1}$ ; standard heat of formation for ozone is  $142.7 \text{ kJ mol}^{-1}$ , standard formation entropy of ozone is  $238.9 \text{ J mol}^{-1} \text{ K}^{-1}$ . *[ $K = 6.52 \times 10^{-58}$ ]*
- P48. Into a 1.00-liter vessel, from which the air was first removed, 1.10 g of solid NOBr was added at  $-55 \text{ °C}$  and then heated to  $0 \text{ °C}$ . The substance is converted to gas and decomposes according to the following reaction equation:  $2 \text{NOBr} \rightleftharpoons 2 \text{NO} + \text{Br}_2$ . The equilibrium pressure measured is  $3.00 \times 10^4 \text{ Pa}$ . Calculate  $K_x$  and  $K_p$  at  $0 \text{ °C}$ . The molar mass of nitrosyl bromide is  $109.91 \text{ g mol}^{-1}$ . Which of the two equilibrium constants does not change if we increase the volume of the system to 2.00 liters? *[ $K_p = 0.2296$ ,  $K_x = 0.7655$ ;  $K_p$  does not change with changing volume.]*
- P49. The value of  $K_p$  for the  $2 \text{NOBr} \rightleftharpoons 2 \text{NO} + \text{Br}_2$  reaction is 0.397 at 25 °C and 0.230 at 0 °C. Calculate the average reaction heat in the 0-25 °C temperature range. *[ $\Delta_r H = 14.78 \text{ kJ/mol}$ ]*
- P50. Calculate the ionic strength of an aqueous solution containing (a) sodium sulfate, (b) potassium chloride, (c) sodium phosphate, or, (d) calcium chloride with 0.0500 mol/kg molality. *[0.1500 mol/kg (a); 0.0500 mol/kg (b); 0.3000 mol/kg (c); 0.1500 mol/kg (d)]*
- P51. Calculate the ionic strength and the mean activity coefficient of a solution containing (a) sodium chloride or (b) copper sulfate with 0.0010 mol/kg molality at 25 °C. It is assumed that the Debye–Hückel limiting law can be used in both solutions. *[ $I = 0.0010 \text{ mol/kg}$ ,  $\gamma_{\pm} = 0.964$  (a);  $I = 0.0040 \text{ mol/kg}$ ,  $\gamma_{\pm} = 0.743$  (b)]*
- P52. For a calcium chloride solution, give the formula for the mean activity coefficient using the activity coefficients of the individual ions. *[ $\gamma_{\pm} = (\gamma_+ \times \gamma_-^2)^{1/3}$ ]*
- P53. The mean activity coefficient in a 0.500 mol/kg lanthanum(III) chloride solution at 25 °C is 0.303. Give the % error of the value calculated with the Debye–Hückel limiting law. *[−99.2%]*
- P54. Calculate the value of  $\Delta_f G^0$  for the biological standard state at 37.0 °C, if  $\Delta_f G^{\ominus} = -21.8 \text{ kJ mol}^{-1}$  for the following reaction:  
$$\text{NADH}(aq) + \text{H}^+(aq) \rightarrow \text{NAD}^+(aq) + \text{H}_2(g) \quad [19.75 \text{ kJ/mol}]$$
- P55. Calculate the solubility constant and the solubility for mercury(II) chloride at 25.0 °C using the following standard formation Gibbs energies:  $\Delta_f G^{\ominus}(\text{Hg}^{2+}, aq) = 164.40 \text{ kJ mol}^{-1}$ ,  $\Delta_f G^{\ominus}(\text{Cl}^-, aq) = -131.23 \text{ kJ mol}^{-1}$ ,  $\Delta_f G^{\ominus}(\text{HgCl}_2, s) = -178.6 \text{ kJ mol}^{-1}$ . *[ $K = 7.75 \times 10^{-15}$ ;  $S = 1.26 \times 10^{-5} \text{ mol/dm}^3$ ]*
- P56. The solubility product of barium sulfate in water at 25.0 °C is  $9.16 \cdot 10^{-11}$ . Calculate its solubility in  
a. pure water, *[ $9.85 \times 10^{-6} \text{ mol/dm}^3$ ]*  
b. 0.0100 mol  $\text{dm}^{-3}$  sodium chloride solution, *[ $1.53 \times 10^{-5} \text{ mol/dm}^3$ ]*  
c. 0.0100 mol  $\text{dm}^{-3}$  ammonium sulfate solution. *[ $4.65 \times 10^{-8} \text{ mol/dm}^3$ ]*