

# Chemistry 2720 Fall 2001 Test 2 Solutions

1. The wavelength of the neutrons must be

$$\lambda = \frac{2d \sin \theta}{n} = \frac{2(1.8 \text{ \AA}) \sin(10^\circ)}{1} = 0.625 \text{ \AA}.$$

The momentum of the neutrons is therefore

$$p = \frac{h}{\lambda} = \frac{6.6260688 \times 10^{-34} \text{ J/Hz}}{0.625 \times 10^{-10} \text{ m}} = 1.06 \times 10^{-23} \text{ kg m/s}.$$

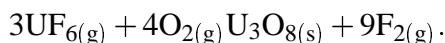
The speed is

$$v = \frac{p}{m} = \frac{1.06 \times 10^{-23} \text{ kg m/s}}{1.67492716 \times 10^{-27} \text{ kg}} = 6328 \text{ m/s}.$$

The kinetic energy is therefore

$$K = \frac{1}{2}mv^2 = \frac{1}{2}(1.67492716 \times 10^{-27} \text{ kg})(6328 \text{ m/s})^2 = 3.35 \times 10^{-20} \text{ J}.$$

2. (a) The reaction is



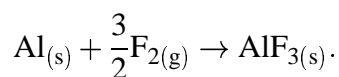
For this reaction,

$$\begin{aligned} \Delta \bar{G}^\circ &= \Delta \bar{G}_{f(\text{U}_3\text{O}_8)}^\circ - 3\Delta \bar{G}_{f(\text{UF}_6)}^\circ \\ &= -3369.5 - 3(-2029.1) \text{ kJ/mol} = 2717.8 \text{ kJ/mol}. \end{aligned}$$

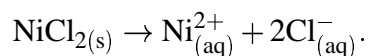
$$\begin{aligned} \Delta \bar{G} &= \Delta \bar{G}^\circ + RT \ln \left( \frac{(a_{\text{U}_3\text{O}_8})(a_{\text{F}_2})^9}{(a_{\text{UF}_6})^3(a_{\text{O}_2})^4} \right) \\ &= 2717.8 \text{ kJ/mol} + (8.314472 \times 10^{-3} \text{ kJ K}^{-1} \text{ mol}^{-1})(298.15 \text{ K}) \ln \left( \frac{(1)(10^{-4})^9}{(1)^3(0.2)^4} \right) \\ &= 2528.3 \text{ kJ/mol}. \end{aligned}$$

Since  $\Delta \bar{G} > 0$ , the process is nonspontaneous.

- (b) Good candidates would either produce the reactant oxygen or remove the product fluorine. Since fluorine is a strong oxidizing agent, its reaction with a good reducing agent might be sufficiently spontaneous to make the desired conversion go. Since oxygen is necessarily also present, we have to choose a reactant which isn't oxidized too fast in air. One possibility would be to try adding metal shavings to the reaction chamber. For instance, we could try to add aluminium shavings. Aluminium is passivated by a thin oxide layer in air, but it might still react with fluorine according to



3. The reaction is



For this reaction,

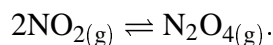
$$\begin{aligned}\Delta\bar{G}^{\circ} &= \Delta\bar{G}_{f(\text{Ni}^{2+})}^{\circ} + 2\Delta\bar{G}_{f(\text{Cl}^{-})}^{\circ} - \Delta\bar{G}_{f(\text{NiCl}_2)}^{\circ} \\ &= -46 + 2(-131.218) - (-259.03) \text{ kJ/mol} = -49 \text{ kJ/mol.} \\ \therefore K &= e^{-\Delta\bar{G}^{\circ}/(RT)} \\ &= \exp\left(\frac{49000 \text{ J/mol}}{(8.314472 \text{ J K}^{-1} \text{ mol}^{-1})(298.15 \text{ K})}\right) \\ &= 4.52 \times 10^8.\end{aligned}$$

Suppose that the solubility is  $s$ . Then the concentration of  $\text{Ni}^{2+}$  is  $s$  and the concentration of  $\text{Cl}^{-}$  ions is  $2s$ . Accordingly,

$$\begin{aligned}K &= s(2s)^2 = 4s^3. \\ \therefore s^3 &= (4.52 \times 10^8)/4 = 1.13 \times 10^8. \\ \therefore s &= 483 \text{ mol/L}\end{aligned}$$

(since the standard concentration is 1 mol/L). This is an enormous value, clearly much higher than the solubility could possibly be. (The mole density of water is only about 55 mol/L.) This means that nickel chloride is extremely soluble in water.

4. The reaction is



Note that the results which follow depend on exactly how you wrote down the reaction. The equilibrium constant can be calculated from the pressures by

$$K = \frac{a_{\text{N}_2\text{O}_4}}{(a_{\text{NO}_2})^2} = \frac{P_{\text{N}_2\text{O}_4}/P^{\circ}}{(P_{\text{NO}_2}/P^{\circ})^2}.$$

Thus

$$K_{298} = \frac{23 \text{ Torr}/750.062 \text{ Torr}}{(46 \text{ Torr}/750.062 \text{ Torr})^2} = 8.15$$

and

$$K_{305} = \frac{30 \text{ Torr}/750.062 \text{ Torr}}{(68 \text{ Torr}/750.062 \text{ Torr})^2} = 4.87.$$

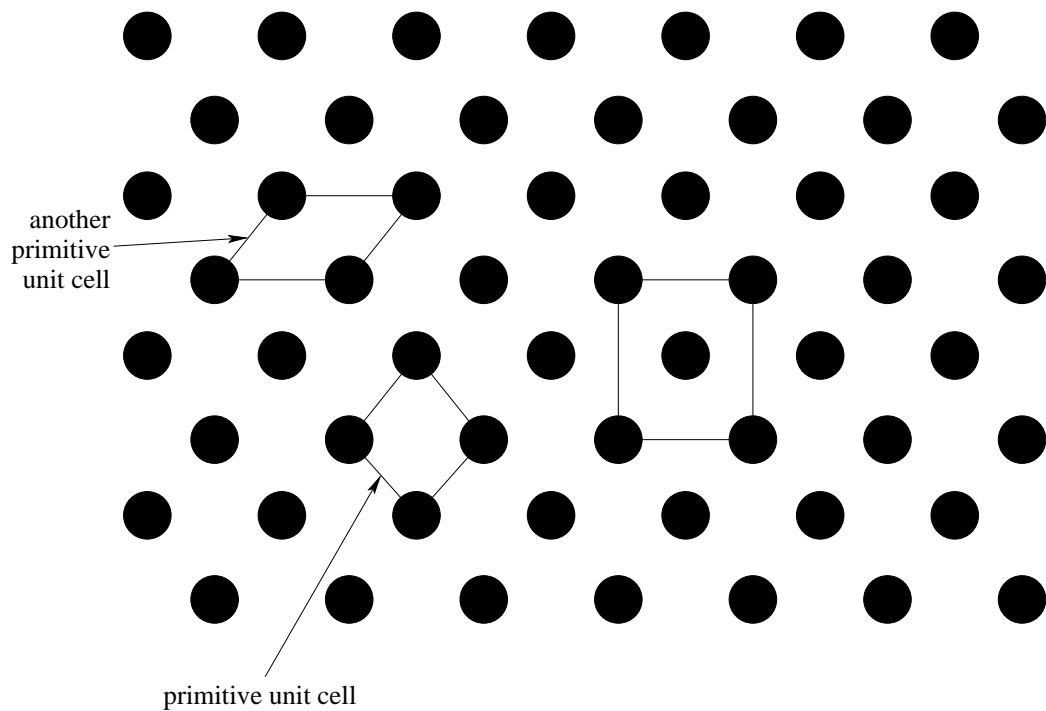
The standard free energy of reaction can be calculated directly from  $K_{298}$ :

$$\begin{aligned}\Delta\bar{G}^{\circ} &= -RT \ln K \\ &= -(8.314472 \text{ J K}^{-1} \text{ mol}^{-1})(298 \text{ K}) \ln(8.15) \\ &= -5.20 \text{ kJ/mol.}\end{aligned}$$

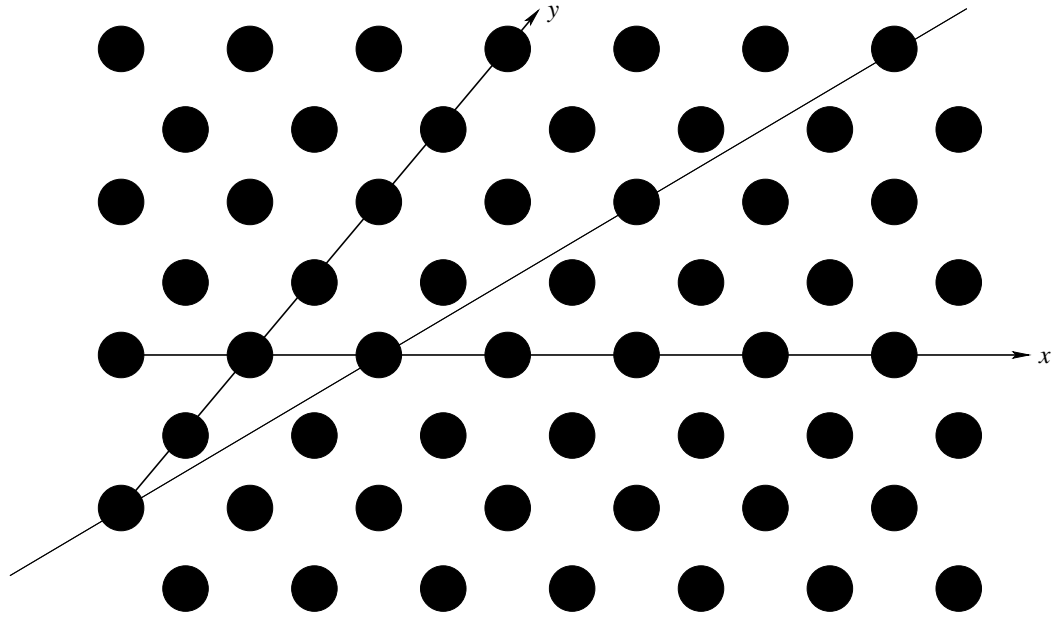
The standard enthalpy of reaction can be found from the variation of the equilibrium constant with temperature:

$$\begin{aligned} \Delta \bar{H}^\circ &= \frac{R \ln(K_1/K_2)}{\frac{1}{T_2} - \frac{1}{T_1}} \\ &= \frac{(8.314472 \text{ J K}^{-1} \text{ mol}^{-1}) \ln(8.15/4.87)}{\frac{1}{305 \text{ K}} - \frac{1}{298 \text{ K}}} \\ &= -55.59 \text{ kJ/mol.} \end{aligned}$$

5. (a)



(b) We start by drawing some  $x$  and  $y$  axes which follow the major crystallographic axes:



The  $x$  intercept is 1 (in units of the unit cell length) and the  $y$  intercept is  $-2$ . The reciprocals of these numbers are 1 and  $-\frac{1}{2}$ . We multiply both of these values by 2 to get integers. The Miller index of this plane is therefore  $(2\bar{1})$ .