Problem Set 3 (For January 27)

Schedule before Exam 1: There will be a practice exam posted on the web site (use the Exam Stuff link to find it) on Monday, January 24, and we will use part or all of x-hour on Thursday, January 27, to review this Problem Set along with the practice exam. Solutions to this problem set will be posted on Wednesday, January 26. **The exam will cover those sections of the text through 6.8 that were emphasized in lecture and on the web Course Calendar with the exceptions of 4.7, 4.11, and 4.12.**

Recommended Text Problems: 5.27, 5.28, 5.31, 5.35, 5.54, 5.56, 6.21, 6.23, 6.31, 6.34, 6.63

1 . A common laboratory vacuum system can attain pressures on the order of 8.0×10^{-7} torr at room temperature (20 °C). How many molecules are in a 1.0 L volume under these conditions?

2 . If C(*s*) is reacted with ("burned in") pure $O_2(g)$, the product is $CO_2(g)$. Suppose 0.631 g of C is burned in a 425 mL vessel at 20.0 °C containing pure O_2 at a pressure of 3.75 atm. (The carbon solid has a volume of 0.32 mL and can be neglected.) After burning, the temperature rises 2.0 °C. What is the final pressure of the vessel, and what is the composition of the gas in it?

3 . In the apparatus below, the left and right compartments are each 1.00 L in volume, and throughout the experiment, the apparatus is maintained at 300 K. 0.052 mol of Cl_2 gas is in the left hand side, and the right side contains 0.258 g solid S in a vacuum. The partition between the two sides is removed, and the Cl_2 reacts with the solid sulfur to form S_2Cl_2 gas. What is the final state of this system? What gas or gases are present? What is the final pressure? What mass of S (if any) remains? (Assume solid S has negligible volume.)

0.052 mol $Cl_2(g)$ 1.00 L	0.258 g S(*s*) 1.00 L

4 . Write the equilibrium constant expression, K_P for the following reactions.

(a) $CO_2(g) + H_2(g) \rightleftarrows CO(g) + H_2O(g)$

(b) $CO_2(g) + H_2(g) \rightleftarrows CO(g) + H_2O(l)$

(c) $F_2(g) + Br_2(l) \rightleftarrows 2\ BrF(g)$

(d) $5\ O_2(g) + 4\ NH_3(g) \rightleftarrows 4\ NO(g) + 6\ H_2O(g)$

5 . An experiment starts with 1.00 mol N_2O_4 at low temperature. This is placed in a container, and the temperature is raise to 423 K. At this temperature, the N_2O_4 is a gas that is significantly dissociated into $NO_2(g)$ according to

$$N_2O_4(g) \rightleftarrows 2\ NO_2(g) \qquad K_P = 115$$

Write the equilibrium constant expression for this reaction first in terms of partial pressures and then in terms of gas mole fractions and the total pressure. If the container is maintained at a total pressure of 1.00 atm as this equilibrium is attained, what are the final, equilibrium amounts and partial pressures of each gas? What would happen (qualitatively—no need for a new calculation here) to this equilibrium if the volume of the container was increased, keeping the temperature constant? (See also Problem 6.37, and its solution, but not until you attempt this one on your own!)