**Partial Pressures and the Equilibrium Constant (Kp)**

**Because gas pressures are easily measured, equilibrium equations for gas – phase reactions are often written using partial pressures (at a constant volume and temperature, the total pressure of a mixture of gases can be found by adding up the pressures of each gas) rather than concentrations.**

**For Example:**

**The equilibrium equation for the decomposition of N2O4 is:**

**N2O4 (g) 2NO2 (g)**

**Kp = (PNO2)2 / (PN2O4)**

**PN2O4 = partial pressure of N2O4**

**PNO2 = partial pressure of NO2**

**Kp = Equilibrium constant using partial pressures.**

**As for Kc, values of Kp are dimensionless because the partial pressures in the equilibrium equation are actually ratios of partial pressures in atmospheres (atm).**

**The values for Kp can be calculated by using pressures instead of Molarity (concentration) and the same equation for Kc.**

**Example:**

**Methane (CH4) reacts with hydrogen sulfide to yield H2 and Carbon Disulfide (in cell phones):**

**CH4 (g) + 2H2S (g) CS2 (g) + 4 H2 (g)**

**What is the Kp value at 1000 K if the partial pressures are:**

**CH4 = .2 atm, CS2 = .52 atm, H2S = .25 atm and H2 = .1 atm**

**Kp = (.52)(.10)4 / (.2)(.25)2 = 4.2 x 10-3**