15.1a) Given the energy profile, I predict that $k_f > k_r$ because the activation energy for the forward reaction is lower than the activation energy for the reverse reaction.

b) Based on this prediction, I predict the K_c value (which is the ratio of products to reactants) will be greater than 1.

15.2 Beginning the in the 4th box from the left, this reaction has reached equilibrium. This is evident in the constant amounts of both reactant and product, indicating that the forward and reverse processes must be occurring at equal rates.

15.3 In the mixture, there are 8 “moles” of product AX, 3 “moles” of recant A and 2 moles of reactant B. As such the Kc expression will be: [8]/([2][3]) or 1.3, making K greater than 1.

15.4 a) $A_2 + B \Rightarrow A_2B$ ($5A_2 + 5B \Rightarrow 4A_2B + A_2 + B$)

b,c) $K_c = \dfrac{[A_2B]}{[A_2][B]}$ $K_c = \dfrac{[.4]}{[.1][.1]} = 4$

d) $\Delta n = \text{moles}_{\text{final}} - \text{moles}_{\text{initial}} = .6 - 1.0 = -.4$

15.5 This friend is wrong. The equilibrium constant (K_c) only speaks of the make-up of the equilibrium mixture. It doesn't say anything about how quickly the reaction reaches equilibrium.

15.6 Equilibrium a has the fewest reactants particles and the most product particles, giving it the highest Kc. Equilibrium c has the most reactant particles and the fewest product particles, giving it the lowest Kc. Equilibrium b is then in the middle.

15.9 a&b) $K_c = 1.5$

For mixture i: $Q = \dfrac{[AB]^2}{[A_2][B_2]} = \dfrac{[6]^2}{[1][1]} = 36$

Since $Q > K_c$, this reaction will proceed toward reactants

For Mixture ii: $Q = \dfrac{[AB]^2}{[A_2][B_2]} = \dfrac{[3]^2}{[3][2]} = 1.5$

Since $Q = K_c$, this reaction is at equilibrium

For Mixture iii: $Q = \dfrac{[AB]^2}{[A_2][B_2]} = \dfrac{[2]^2}{[3][3]} = 0.44$

Since $Q < K_c$, this reaction will proceed toward products

15.10 a) $$K_c = \frac{[AB]^2}{[A_2][B]^2} = \frac{[2]^2}{[1][1]} = 4$$

b) A decrease in volume favors the side of the reaction with fewer moles of gas. As such, the decrease in volume favors the products. Note, however, that this decrease in volume doesn't change the value of K_c. Instead, the concentration of AB must increase to maintain K_c.

15.13a) $$K_c = \frac{K_f}{K_r} = \frac{4.7x10^{-3}\ s^{-1}}{5.8x10^{-1}\ s^{-1}} = 8.1\ x\ 10^{-3}$$

b) Since K_c is less than 1, the reaction favors reactants. As such, the partial pressure of A will be greater.

15.16 a) homogeneous

$$K_c = \frac{[O_2]^3}{[O_3]^2}$$

b)heterogeneous

$$K_c = \frac{1}{[Cl_2]^2}$$

e)heterogeneous

$$K_c = \frac{[Cl_2]^2}{[O_2][HCl]^4}$$

g)heterogeneous

$$K_c = \frac{[CO_2]^{16}}{[O_2]^{25}}$$

15.21 $K_p = K_c(RT)^{\Delta n} = 0.042(0.08206 \cdot 500K)^{-1} = 1.0 \text{ x}10^{-3}$

15.23a) Given that the K_c value is smaller than 1, the reaction favors the reactants NO and Br_2.

b) The Kc for the reverse reaction will be the inverse of the forward reaction, or $1/1.3 \text{ x } 10^{-2}$, which equals 77.

c) Because the coefficients in the balanced equation become exponents in the K_c expression, dividing all the coefficients in a balanced equation by 2 will cause the K_c for the new equation to be raised to the ½ power, which is the square root. As such, the Kc will be $77^{1/2}$ or 8.8.

15.25 a) Never mind

b) Doubling the coefficients in the balanced equation will cause the Kc value to increase by a power of 2. As such, the Kp for this reaction will be 1.85^2 or 3.42.

c) $K_p = K_c(RT)^{\Delta n}$

$3.42 = K_c(0.08206 \cdot 1000K)^{-1}$

$K_c = 281$

15.29 We exclude pure solids and liquids from equilibrium constant expressions because their concentrations or partial pressures are understood to be unchanging (or zero).