Part A: Understanding K and Q Conceptual

True or False and briefly Discuss your answer below.

_____ *Q* measures the relative amounts of products and reactants present during a reaction at a particular point in time.

_____ Q < K the reaction will favor the formation of products

_____ Q > K the reaction will favor the formation of products

_____ Q = K means the reaction is at equilibrium

_____ Q can have many values unlike an equilibrium constant for a chemical reaction.

Why?____

Part B: The reaction quotient, Q

1. Consider the following reaction at equilibrium:

$$2\ KClO_3\ (s) \rightleftharpoons 2\ KCl\ (s) + 3\ O_2\ (g) \quad K_c = 10.1$$

a) Write the form of the reaction quotient (Q_c) for this reaction:

b) Determine the value of the Q_c if 1.0 mol of O_2, 2.0 mol of $KClO_3$, and 1.0 mol of KCl are mixed in a 1.0-L container at 472°C.

c) Based on your value of Q_c, which direction does the reaction have to shift to reach equilibrium?

d) Predict which direction the reaction will shift when O_2 is removed and KCl is added.

Part C: Finding equil. concentrations given *K* and all but one equil. concentration

2. Consider the reaction: $2\ COF_2(g) \rightleftharpoons CO_2(g) + CF_4(g)$ $K_c = 2.00$

   a) Write the equilibrium constant expression for this reaction.

   b) If at equilibrium the concentration of CF_4 = 0.118 M and the concentration of CO_2 is 1.10 M, what must be the concentration of COF_2 at equilibrium. Notice from the K_c expression above, the only variable we don't have is $[COF_2]$.

   c) This reaction is an endothermic reaction. Your lab partner takes your flask, which is at equilibrium, and places it in an oven for 20min, then takes it out and puts it on the lab bench. You come back and determine the concentrations and find $[COF_2]$ = 0.0322, $[CO_2]$ = 2.20M and $[CF_4]$ = 0.222M. Based on this information, has the flask returned to equilibrium? What has to happen?

Part D: Finding equil. concentrations when we know *K* and the initial concentrations

3. Consider the reaction: $2\ H_2S(g) \rightleftharpoons 2\ H_2(g) + S_2(g)$ $K_c = 1.67 \times 10^{-7}$

a) Write the form of K_c for this reaction:

If we start with an initial $[H_2S] = 0.010$ M determine the equilibrium concentrations. The method we used in question 2 won't work since we don't know ANY equilibrium concentrations.

b) Instead, start with the ICE table below and fill in the initial concentrations.

	$2\ H_2S$ $\rightleftharpoons$	$2\ H_2(g)$ +	$S_2(g)$
Initial			
Change			
Equil.			

c) Next use variables (x) to fill in the "change" and "equilibrium" rows in the table.

d) Plug the equilibrium values and the value of K_c into the expression you found in question 3a.

e) We can solve this problem without the quadratic equation or without using the method of successive approximations by realizing that because K_c is so small, it is likely that the "x" in the denominator is also small. Drop the "x" in the denominator and solve for the "x" in the numerator.

f) Now that you know the value of "x", go back to the ICE table and solve for each equilibrium concentration.

g) Check your answer by plugging your equilibrium concentrations into the equation you wrote in question 3a and verify that it gives the known value of K_c.

OK! When do you need to use the quadratic formula? If $100 \times K_c <$ $[A]_{init}$ then you can ignore x in denominator. If not, you have to use the quadratric equation.