1. Write an expression for the equilibrium constant of each chemical equation.

   (a) $SbCl_5(g) \rightleftharpoons SbCl_3(g) + Cl_2(g)$ $K = \frac{[SbCl_3][Cl_2]}{[SbCl5]}$

   (b) $2\,BrNO(g) \rightleftharpoons 2\,NO(g) + Br_2(g)$ $K = \frac{[NO]^2[Br_2]}{[BrNO]^2}$

2. Consider the reaction: $NH_4H_2(s) \rightleftharpoons NH_3(g) + H_2S(g)$

   An equilibrium mixture of this reaction at a certain temperature has $[NH_3]$ = 0.278 M and $[H_2S]$ = 0.355 M. What is the value of K_c at this temperature?

   $$K_c = \frac{[NH_3][H_2S]}{1} = [NH_3][H_2S] = (0.278\ M)(0.355\ M) = 0.0987$$

3. Silver sulfate dissolves in water according to the reaction:
   $Ag_2SO_4(s) \rightleftharpoons 2\,Ag^+(aq) + SO_4^{\ 2-}(aq)$ $K_c = 1.1 \times 10^{-5}$ at 298 K.
   A 1.5 L solution contains 6.55 g of dissolved silver sulfate. If additional solid silver sulfate is added to the solution, will it dissolve? (Hint: Calculate Q and compare to K)

   Need concentration instead of mass: $(6.55\,g\ Ag_2SO_4)\left(\frac{1\ mol\ Ag_2SO_4}{311.80\ g\ Ag_2SO_4}\right) = 0.02101\,mol\ Ag_2SO_4$. Once dissolved, this is $(0.02101\,mol\ Ag_2SO_4)\left(\frac{2\ mol\ Ag^+}{1\ mol\ Ag_2SO_4}\right) = 0.04201\,mol\ Ag^+$, and 0.02101 moles of $SO_4^{\ 2-}$.

   Those moles are in 1.5 L of solution, so:
   $[Ag^+] = \frac{0.0420\ mol\ Ag^+}{1.5\ L\ soln.} = 0.0280\ M$
   $[SO_4^{2-}] = \frac{0.02101\ mol\ SO_4^{2-}}{1.5\ L\ soln} = 0.0140\ M$

   $Q = [Ag^+]^2[SO_4^{2-}] = (0.0280)^2(0.0140) = 1.1 \times 10^{-5}$.

   In this case, $Q = K$, so the system is already at equilibrium. The equilibrium will not shift to the right. If additional Ag_2SO_4 solid is added, it will not go into solution.

4. For the reaction: $HC_2H_3O_2(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + C_2H_3O_2^-(aq)$, $K_c = 1.8 \times 10^{-5}$. If a solution initially consisted of 1.0000 M $HC_2H_3O_2$, calculate the equilibrium concentrations of the reactants and products (ignore the water).

	$HC_2H_3O_2$	+	H_2O	$\rightleftharpoons$	H_3O^+	+	$C_2H_3O_2^-$
I	1.0000 M		–		0		0
C	-x		–		+x		+x
E	1.0000 - x		–		x		x

$K_c = 1.8 \times 10^{-5} = \frac{x^2}{1.0000-x}$.

Assume $x << 1.0000$, so $1.8 \times 10^{-5} = \frac{x^2}{1.0000}$.

Then $x^2 = 1.8 \times 10^{-5} \Rightarrow x = 0.00424\ M$.

Was our assumption valid? $\frac{0.00424}{1} \times 100\% = 0.424\%$. This is less than 5%, so our assumption was ok.

At equilibrium: $[HC_2H_3O_2]$ = 0.9958 M, $[H_3O^+]$ = 0.0042 M, and $[C_2H_3O_2^-]$ = 0.0042 M.

5. For the reaction $N_2O_4(g) \rightleftharpoons 2\,NO_2(g)$ $K_c = 0.513$ at 500 K.
If a reaction vessel initially contains an N_2O_4 concentration of 0.0500 M at 500 K, what are the equilibrium concentrations of N_2O_4 and NO_2 at 500 K?

	N_2O_4	$\rightleftharpoons$	$2\,NO_2$
I	0.0500 M		0
C	-x		+2x
E	0.0500 - x		2x

$K_c = \frac{[NO_2]^2}{[N_2O_4]} = \frac{(2x)^2}{(0.0500-x)} = \frac{4x^2}{0.0500-x} = 0.513$

Try assuming $x << 0.0500$: $\frac{4x^2}{0.0500} = 0.513$

$x = 0.0801\ M$. In this case, $\frac{0.0801}{0.0500} \times 100 = 160\%$, so this was a terrible assumption (it's greater than 0.05!).

Use the quadratic equation or successive approximations to get $x = 0.0385\ M$.

Then at equilibrium: $[\mathrm{N_2O_4}] = 0.0500 - x = 0.0115\ M$ and $[\mathrm{NO_2}] = 2x = 0.0770\ M$.

6. Consider the reaction: $CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$, $K_c = 102$ at 500 K.
   If a reaction mixture initially contains 0.110 M CO and 0.110 M H_2O, what will the equilibrium concentration of each of the reactants and products be?

	CO	+	H_2O	$\rightleftharpoons$	CO_2	+	H_2
I	0.110 M		0.110		0		0
C	-x		-x		+x		+x
E	0.110 - x		0.110 - x		x		x

$$K_c = 102 = \frac{x^2}{(0.110-x)^2} = \left(\frac{x}{0.110-x}\right)^2$$

$$\sqrt{102} = 10.1 = \frac{x}{0.110-x}$$

$$1.11 - 10.1x = x$$

$$1.11 = 11.1x$$

$$x = 0.100\ M$$

At equilibrium: [CO] = [H_2O] = 0.010 M and [CO_2] = [H_2] = 0.100 M

You can check your answer by solving for K_c: $K_c = \frac{[CO_2][H_2]}{[CO_2][H_2]} = \frac{(0.100)(0.100)}{(0.010)(0.010)} = 100$, which matches the K_c given in the problem (within rounding).

7. Consider this reaction at equilibrium: $2\,BrNO(g) \rightleftharpoons 2\,NO(g) + Br_2(g)$.
   Predict whether the reaction will shift left, shift right, or remain unchanged after each disturbance.

   (a) NO is added to the reaction mixture (assume the pressure is constant). Shift left

   (b) BrNO is added to the reaction mixture (assume the pressure is constant). Shift right

   (c) Nothing is added, but the volume is decreased (pressure is increased). Shift left

8. This reaction is exothermic (ΔH_{rxn} is negative):
   $C_6H_{12}O_6(s) + O_2(g) \rightleftharpoons 6\,CO_2(g) + 6\,H_2O(g)$.
   Predict the effect (shift right, shift left, or no effect) of increasing and decreasing the reaction temperature. How does the value of the equilibrium constant depend on temperature?

   Because it is exothermic, we could think of heat as being a product. Increasing temperature would shift the equilibrium to the left (decrease K). Decreasing temperature would shift to the right (increase K).

9. Coal can be used to generate hydrogen gas (a potential fuel) by the endothermic reaction:
   $C(s) + H_2O(g) \rightleftharpoons CO(g) + H_2(g)$
   If this reaction mixture is at equilibrium, predict whether each disturbance will result in the formation of additional hydrogen gas, the formation of less hydrogen gas, or have no effect on the quantity of hydrogen gas.

   (a) adding more C to the reaction mixture – no change (solids do not appear in K expression) This assumes that C is already present and you would be adding more.

   (b) adding more H_2O to the reaction mixture – Shift right

   (c) raising the temperature of the reaction mixture – Endothermic, so shift right

   (d) increasing the volume of the reaction mixture – lower P, shift right

   (e) adding a catalyst to the reaction mixture – no change in equilibrium, it would just get to equilibrium faster

   (f) adding an inert gas to the reaction mixture – no change