Equilibrium Problems #2

When 3.556g of ammonia {$NH_3(g)$} and 9.924g of oxygen {$O_2(g)$} are sealed together in a 5.000L vessel, they reach equilibrium with nitrogen dioxide {$NO_2(g)$} and water {$H_2O(g)$}. If the equilibrium concentration of water is found to be 2.473×10^{-3} M:

a. What are the equilibrium concentrations of all products and reactants?

	$4\ NH_3(g)$ +	$7\ O_2(g)$ $\leftrightarrows$	$4\ NO_2(g)$ +	$6\ H_2O(g)$
Initial	$(3.556g/28.014\ ^{g}/_{mol})$ /5.00L = 2.539×10^{-2} M	$(9.924g/31.998\ ^{g}/_{mol})$ /5.00L = 6.203×10^{-2} M	0 M	0 M
Change	- 4x	- 7x	+ 4x	+ 6x
Equilibrium	$(2.539\times10^{-2} - 4x)$ M	$(6.203\times10^{-2} - 7x)$ M	4x M	6x M

Since $[H_2O]_{eq} = 2.473\times10^{-3}$ M, $x = 4.122\times10^{-4}$. Plugging in to concentrations:

Equilibrium	2.374×10^{-2} M	5.914×10^{-2} M	1.649×10^{-3} M	2.473×10^{-3} M

b. What is the value of the equilibrium constant for this reaction?

Plugging the above concentrations into the equilibrium constant expression:

$$K_c = \{(1.648\times10^{-3})^4(2.473\times10^{-3})^6\}/\{(2.374\times10^{-2})^4(5.914\times10^{-2})^7\}$$
$$K_c = \{(7.376\times10^{-12})(2.287\times10^{-16})\}/\{(3.177\times10^{-7})(2.530\times10^{-9})\}$$
$$K_c = 2.099\times10^{-12}$$

c. Is the reaction Product-favored or reactant-favored?

Since K_c is much less than 1, the reaction is reactant-favored.

When the tetrachlorocuprate ion ($CuCl_4^{2-}$) is dissolved in water, it reaches equilibrium with Cu^{2+} and Cl^- ions. If 6.854g of $CuCl_4^{2-}$ is dissolved in water to make 250.0 mL of solution, the equilibrium concentration of chloride ion is found to be 7.442×10^{-2}M.

a. What are the equilibrium concentrations of all products and reactants?

	$CuCl_4^{2-}(aq)$ ⇆	$Cu^{2+}(aq)$ +	4 $Cl^-(aq)$
Initial	(6.854g/205.358 $^g/_{mol}$) /0.2500L = 0.1335 M	0 M	0 M
Change	- x	+ x	+ 4x
Equilibrium	(0.1335 - x) M	x M	4x M

Since $[Cl^-]_{eq} = 7.442 \times 10^{-2}$ M, x = 1.861×10^{-2} M. Plugging in to concentrations:

Equilibrium	0.1149 M	0.01861 M	0.07442 M

b. What is the value of the equilibrium constant for this reaction?

$K_c = 4.968 \times 10^{-6}$

c. Is the reaction product-favored or reactant-favored?

Since K_c is much less than 1, the reaction is reactant-favored.

d. If the $CuCl_4^{2-}$ listed above were dissolved in a 0.0155 M solution of NaCl(aq) instead of pure water, how would the equilibrium concentrations of all species change?

Since Cl^- is a *product* in this reaction, using NaCl(aq) would increase the $[Cl^-(aq)]$ and force the reaction to shift toward *reactants* to return to equilibrium. Therefore, the concentration of tetrachlorocuprate would increase and the concentration of copper(II) ions would decrease.

Silver bromide has a K_{sp} of $5.0x10^{-13}$. If you were to prepare 5.00 L of a saturated solution of silver bromide, how many grams of silver bromide would be dissolved in this solution?

$$AgBr(s) \leftrightarrows Ag^+(aq) + Br^-(aq)$$
$$K_{sp} = [Ag^+]_{eq}[Br^-]_{eq}$$

Every AgBr that dissolves gives equal amounts of $Ag^+(aq)$ and $Br^-(aq)$ so the K_{sp} expression can be simplified to:

$$K_{sp} = x^2$$
$$x = 7.071x10^{-7}\ M$$

So in 5.00L of solution, there are $3.536x10^{-6}$ mols of $Ag^+(aq)$. This means that $3.536x10^{-6}$ mols of AgBr(s) must have dissolved, so $6.639x10^{-4}$ g of AgBr must have dissolved.

If you were to prepare the same solution as above, but instead of water you used 0.100 M potassium bromide solution as your solvent, how much silver bromide would dissolve?

This one probably requires a table:

	$AgBr(s) \leftrightarrows$	$Ag^+(aq)$ +	$Br^-(aq)$
Initial		0.100 M	0 M
Change		+ x	+ x
Equilibrium		(0.100 + x) M	x M

$K_{sp} = [Ag^+]_{eq}[Br^-]_{eq} = (0.100 + x)(x) = 5.0x10^{-13}$

Based upon the first part, it's probably safe to assume that "x" is small compared to 0.100, so this can be simplified as solved, $x = 5.0x10^{-12}$ M. $4.694x10^{-9}$ g of AgBr dissolves.

You have studied the following reaction of the bicarbonate ion:

$HCO_3^-(aq) \leftrightarrows H^+(aq) + CO_3^{2-}(aq)$ $\qquad K_c = 4.8x10^{-11}$

If you prepare a solution that has an initial bicarbonate concentration $\{[HCO_3^-(aq)]_o\}$ of 1.388 M, what are the equilibrium concentrations of all species?

Set up a table and you get to:

$K_c = \{[H^+]_{eq}[CO_3^{2-}]_{eq}\}/\{[HCO_3^-]_{eq}\} = (x \bullet x)/(1.388-x)$

Since K_c is very small, the value of x is probably insignificant compared to 1.388, so this could be simplified to:

$4.8x10^{-11} = x^2/1.388$

$x = 8.162x10^{-6}$

$[H^+]_{eq} = 8.162x10^{-6}$ M

$[CO_3^{2-}]_{eq} = 8.162x10^{-6}$ M

$[HCO_3^-]_{eq} = 1.388$ M