

Equilibrium Problems #2

When 3.556g of ammonia {NH₃(g)} and 9.924g of oxygen {O₂(g)} are sealed together in a 5.000L vessel, they reach equilibrium with nitrogen dioxide {NO₂(g)} and water {H₂O(g)}. If the equilibrium concentration of water is found to be 2.473x10⁻³ M:

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

	4 NH ₃ (g) +	7 O ₂ (g) ⇌	4 NO ₂ (g) +	6 H ₂ O(g)
Initial	(3.556g/28.014 ^g /mol) /5.00L = 2.539x10 ⁻² M	(9.924g/31.998 ^g /mol) /5.00L = 6.203x10 ⁻² M	0 M	0 M
Change	- 4x	- 7x	+ 4x	+ 6x
Equilibrium	(2.539x10 ⁻² - 4x) M	(6.203x10 ⁻² - 7x) M	4x M	6x M

Since [H₂O]_{eq} = 2.473x10⁻³ M, x = 4.122x10⁻⁴. Plugging in to concentrations:

Equilibrium	2.374x10 ⁻² M	5.914x10 ⁻² M	1.649x10 ⁻³ M	2.473x10 ⁻³ M
-------------	--------------------------	--------------------------	--------------------------	--------------------------

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

Plugging the above concentrations into the equilibrium constant expression:

$$K_c = \frac{(1.648 \times 10^{-3})^4 (2.473 \times 10^{-3})^6}{(2.374 \times 10^{-2})^4 (5.914 \times 10^{-2})^7}$$

$$K_c = \frac{(7.376 \times 10^{-12})(2.287 \times 10^{-16})}{(3.177 \times 10^{-7})(2.530 \times 10^{-9})}$$

$$K_c = 2.099 \times 10^{-12}$$

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 $7.442 \times 10^{-2} \text{ M}$.

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

	$\text{CuCl}_4^{2-}(\text{aq}) \rightleftharpoons$	$\text{Cu}^{2+}(\text{aq}) +$	$4 \text{Cl}^-(\text{aq})$
Initial	$(6.854\text{g}/205.358 \text{ g/mol})$ $/0.2500\text{L} =$ 0.1335 M		
Change	$- x$	$+ x$	$+ 4x$
Equilibrium	$(0.1335 - x) \text{ M}$	$x \text{ M}$	$4x \text{ M}$

Since $[\text{Cl}^-]_{\text{eq}} = 7.442 \times 10^{-2} \text{ M}$, $x = 1.861 \times 10^{-2} \text{ M}$. 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?

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

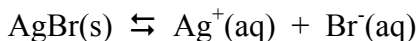
c. Is the reaction product-favored or reactant-favored?

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 $\text{NaCl}(\text{aq})$ instead of pure water, how would the equilibrium concentrations of all species change?

Since Cl^- is a *product* in this reaction, using $\text{NaCl}(\text{aq})$ would increase the $[\text{Cl}^-(\text{aq})]$ 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 5.0×10^{-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?



$$K_{sp} = [\text{Ag}^+]_{eq}[\text{Br}^-]_{eq}$$

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

$$K_{sp} = x^2$$

$$x = 7.071 \times 10^{-7} \text{ M}$$

So in 5.00L of solution, there are 3.536×10^{-6} mols of $\text{Ag}^+(\text{aq})$. This means that 3.536×10^{-6} mols of AgBr(s) must have dissolved, so 6.639×10^{-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?

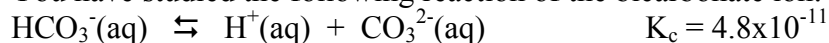
This one probably requires a table:

	$\text{AgBr(s)} \rightleftharpoons$	$\text{Ag}^+(\text{aq}) +$	$\text{Br}^-(\text{aq})$
Initial		0.100 M	0 M
Change		+ x	+ x
Equilibrium		(0.100 + x) M	x M

$$K_{sp} = [\text{Ag}^+]_{eq}[\text{Br}^-]_{eq} = (0.100 + x)(x) = 5.0 \times 10^{-13}$$

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, $x = 5.0 \times 10^{-12}$ M. 4.694×10^{-9} g of AgBr dissolves.

You have studied the following reaction of the bicarbonate ion:



If you prepare a solution that has an initial bicarbonate concentration $\{[\text{HCO}_3^-(\text{aq})]_o\}$ of 1.388 M, what are the equilibrium concentrations of all species?

Set up a table and you get to:

$$K_c = \{[\text{H}^+]_{eq}[\text{CO}_3^{2-}]_{eq}\} / \{[\text{HCO}_3^-]_{eq}\} = (x \cdot x) / (1.388 - x)$$

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

$$4.8 \times 10^{-11} = x^2 / 1.388$$

$$x = 8.162 \times 10^{-6}$$

$$[\text{H}^+]_{eq} = 8.162 \times 10^{-6} \text{ M}$$

$$[\text{CO}_3^{2-}]_{eq} = 8.162 \times 10^{-6} \text{ M}$$

$$[\text{HCO}_3^-]_{eq} = 1.388 \text{ M}$$