

5 Point Bonus for all Quizzes Submitted by 4:00 PM Friday, December 2

Complete the following individually. You may use your textbook and notes, but may not receive assistance from your classmates or anyone other than Dr. Lamp. *This signed statement must accompany the completed assignment.* By signing below, you certify that you completed the problems in accordance with these rules. No credit will be given to unsigned papers. Staple any additional sheets prior to turning the assignment in.

Signature _____ Date _____

Complete the following problems on separate paper and **staple the pages** to this sheet **write your final answers on this page.** You must show your work to receive full credit. Show your answers to the correct number of significant figures with the correct units.

1. A 0.431 g sample of HCl(g) was placed in a 625 mL reaction vessel at a 862K and allowed to dissociate to H₂ and Cl₂. When equilibrium is reached between HCl(g), H₂(g) and Cl₂(g), 0.0414 g Cl₂ is present. What is the K_c for the reaction at this temperature?

Answer: K_c = 0.00300₃

First we need a balanced reaction: $2\text{HCl} \rightleftharpoons \text{H}_2 + \text{Cl}_2$. To determine the equilibrium constant, we need values for equilibrium concentrations for each of the species. We are given information related to the initial amount of HCl and the equilibrium amount of Cl₂. One approach is to treat this with the ICE table. To do so, we need to convert the masses into molarities.

$$0.431 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.461 \text{ g}} \times \frac{1}{0.625 \text{ L}} = 0.0189_1 \text{ M HCl}$$

$$0.0414 \text{ g Cl}_2 \times \frac{1 \text{ mol Cl}_2}{70.9054 \text{ g}} \times \frac{1}{0.625 \text{ L}} = 9.34_2 \times 10^{-4} \text{ M Cl}_2$$

	2HCl	\rightleftharpoons	H ₂	+	Cl ₂
I	0.0189 ₁		0		0
C	-2(9.34 ₂ × 10 ⁻⁴)		9.34 ₂ × 10 ⁻⁴		9.34 ₂ × 10 ⁻⁴
E	0.0189 ₁ -2(9.34 ₂ × 10 ⁻⁴) = 0.0170 ₄ M		9.34 ₂ × 10 ⁻⁴ M		9.34 ₂ × 10 ⁻⁴ M

Therefore, K_c is:

$$K_c = \frac{[\text{H}_2][\text{Cl}_2]}{[\text{HCl}]^2} = \frac{(9.34_2 \times 10^{-4} \text{ M})(9.34_2 \times 10^{-4} \text{ M})}{(0.0170_4 \text{ M})^2} = 0.00300_3$$

2. Carbonyl bromide, $\text{COBr}_2(\text{g})$, decomposes to $\text{CO}(\text{g})$ and $\text{Br}_2(\text{g})$ with an equilibrium constant, K_c , of 0.190 at 73°C . If a 0.0150 mol sample of COBr_2 is heated in a 2.50 L flask until equilibrium is attained, what will be the concentrations of all COBr_2 , CO , and Br_2 at equilibrium?

Answer: $[\text{COBr}_2] = 0.000178 \text{ M}$

$[\text{CO}] = 0.00582 \text{ M}$

$[\text{Br}_2] = 0.00582 \text{ M}$

The ICE table approach works well here. First we need to get things in terms of concentration:

$$\frac{0.0150 \text{ mol Br}_2}{2.5\text{L}} = 0.00600 \text{ M COBr}_2$$

	COBr_2	\rightleftharpoons	CO	+	Br_2
I	0.00600		0		0
C	-x		+x		+x
E	$0.00600 - x$		x		x

Inserting into K_c gives:

$$K_c = \frac{[\text{CO}][\text{Br}_2]}{[\text{COBr}_2]} = \frac{(x)(x)}{0.00600 - x}$$

Now some algebra:

$$(0.00600 - x)K_c = x^2$$

$$0 = x^2 + K_c x - 0.00600K_c = (0.190)^2 + 0.190x - 0.00114$$

From the quadratic formula, we find $x = 0.00582_2$ or -0.196

Since x represents the equilibrium concentration of CO and Br_2 , a negative value makes no chemical sense, therefore, the value $x = 0.00582_2$ is the reasonable result.

Now to find the equilibrium concentrations of all species:

$$[\text{CO}] = [\text{Br}_2] = x = 0.00582 \text{ M}$$

$$[\text{COBr}_2] = 0.00600 - x = 0.000178 \text{ M}$$

3. In the reaction $\text{CO(g)} + \text{H}_2\text{O(g)} \rightleftharpoons \text{CO}_2\text{(g)} + \text{H}_2\text{(g)}$, $K_c = 31.4$ at 588 K. If 10.1 g of each reactant and product are brought together in a 1.00 L reaction vessel at 588 K, how many grams of H_2O will be present at equilibrium?

Answer: mass $\text{H}_2\text{O} = 6.83_1$ g H_2O

The ICE table approach works well here. First we need to get things in terms of concentration:

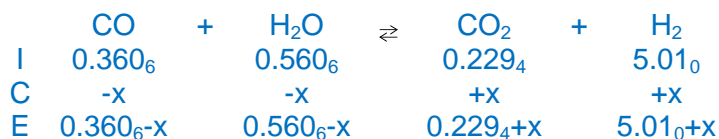
$$10.1 \text{ g CO} \times \frac{1 \text{ mol CO}}{28.010 \text{ g}} \times \frac{1}{1 \text{ L}} = 0.360_6 \text{ M CO}$$

Repeating the process for the other reactants and products we find the following initial concentrations: $[\text{H}_2\text{O}] = 0.560_6 \text{ M}$, $[\text{CO}_2] = 0.229_4 \text{ M}$, $[\text{H}_2] = 5.01_0 \text{ M}$

Now we need to figure out what direction the reaction will go on the way to equilibrium:

$$Q = \frac{[\text{CO}_2][\text{H}_2]}{[\text{CO}][\text{H}_2\text{O}]} = \frac{(0.229_4)(5.01_0)}{(0.360_6)(0.560_6)} = 5.69$$

Since Q is less than K_c , we need to make more products and the reaction will proceed to the right.



Inserting into K_c gives:

$$K_c = \frac{[\text{CO}_2][\text{H}_2]}{[\text{CO}][\text{H}_2\text{O}]} = \frac{(0.229_4+x)(5.01_0+x)}{(0.360_6-x)(0.560_6-x)} = \frac{1.14_9 + 5.23_9x + x^2}{0.202_1 - 0.921_2x + x^2}$$

Now some algebra:

$$\begin{aligned} (0.202_1 - 0.921_2x + x^2)K_c &= 1.14_9 + 5.23_9x + x^2 \\ 6.34_8 - 28.9_2x + 31.4x^2 &= 1.14_9 + 5.23_9x + x^2 \\ 30.4x^2 - 34.1_6x + 5.19_8 &= 0 \end{aligned}$$

From the quadratic formula, we find $x = 0.942_2$ or 0.181_4

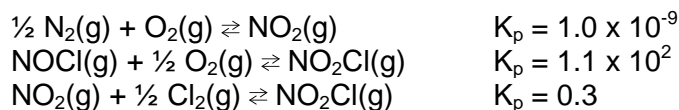
If we use the value $x = 0.942_2 \text{ M}$, we will arrive at negative concentrations for CO and H_2O , which makes no chemical sense. Therefore, $x = 0.181_4 \text{ M}$ is the reasonable result.

Thus, the equilibrium concentration of H_2O is $0.560_6 - 0.181_4 = 0.379_2 \text{ M}$

And the mass of H_2O is:

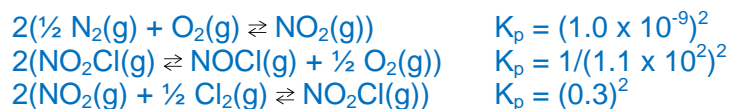
$$\frac{0.379_2 \text{ mol H}_2\text{O}}{1 \text{ L}} \times 1.00 \text{ L} \times \frac{18.015 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 6.83_1 \text{ g H}_2\text{O}$$

4. Determine K_c for the reaction: $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightleftharpoons 2\text{NOCl}(\text{g})$ from the following data at 298K:

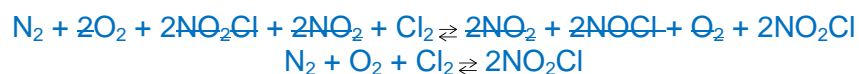


Answer: $K_c = 1.8 \times 10^{-22}$

We need to rearrange reactions to make our target reaction:



So, the sum of the reactions is:



$$K_p = \frac{(1.0 \times 10^{-9})^2 (0.3)^2}{(1.1 \times 10^2)^2} = 7.4 \times 10^{-24}$$

Now we need to convert from K_p to K_c :

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

In this reaction we have 2 moles of gaseous products and 1 mole of gaseous reactions, so $\Delta n = -1$
 $K_c = (7.4 \times 10^{-24}) / (0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1} \times 298 \text{ K})^{-1} = 1.8 \times 10^{-22}$