Chap 14 Team Problems

<u>Problem #1</u>
Ammonium carbamate, $NH_4CO_2NH_2$, decomposes according to the following reaction:

$NH_4CO_2NH_{2(s)} \leftrightarrow 2NH_{3(g)} + CO_{2(g)}$

Some ammonium carbamate is put into an evacuated container at 40ºC and allowed to come to equilibrium, the pressure in the container is found to be 0.363atm. What is the equilibrium constant K_p?

ANSWER:

Using Dalton's Law of Partial Pressures:
$P_{total} = P_{NH3} + P_{CO2} = 0.363atm$

Based on the reaction, there is a 2:1 ration of $NH_3 : CO_2$

Therefore, $2x + x = 0.363atm$
Solving for x
$x = 0.121atm$

The equilibrium expression for this reaction is:
$K_p = [NH_3]^2 [CO_2]$

This means:
$[2x]^2 [x] = K_p$

Using the value of "x" from above: $[2(0.121atm)]^2 [0.121atm] = K_p$

Solving: $K_p = 7.09e\text{-}3$

<u>Problem #2</u>
The dissociation of molecular iodine into iodine atoms is represented by:

$I_{2(g)} \leftrightarrow 2I_{(g)}$

At 1000 K the equilibrium constant, $K_c = 3.80e\text{-}5$. Suppose you start with 0.0456 mol of I_2 in a 2.30L flask at 1000 K and allow it to come to equilibrium. What are the concentrations of the molecular iodine and of the iodine atoms?

ANSWER
In order to use an ICE chart, we must have the concentration of the initial I_2 : 0.0456mol / 2.30 L = 0.0198M

Now run an ICE chart:

	$I_{2(g)}$	$\leftrightarrow$	$2I_{(g)}$
Initial	0.0198M		0M
Change	- x		+2x
Equil.	0.0198 – x		2x

Plug into the equilibrium expression: $K_c = [I]^2 / [I_2] = [2x]^2 / [0.0198 - x] = 3.80e\text{-}5$
Assuming "x" is less than 5% of 0.0198, solve for "x": x = 4.34e-4 (4.34e-4/0.0198 = 2% Hooray!)

Use this to calculate the equilibrium concentrations:
$[I_2] = 0.0194M$, $[I] = 8.68e\text{-}4M$