Write the expression for the equilibrium constant for the following reactions:

\[ \text{N}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{NO}(g) \]

\[ K_c = \frac{[\text{NO}]^2}{[\text{N}_2][\text{O}_2]} \quad K_p = \frac{P_{\text{NO}}^2}{P_{\text{N}_2}P_{\text{O}_2}} \]

\[ 2\text{NaHCO}_3(s) = \text{Na}_2\text{CO}_3(s) + \text{CO}_2(g) + \text{H}_2\text{O}(g) \]

\[ K_c = \frac{[\text{CO}_2][\text{H}_2\text{O}]}{[\text{NaHCO}_3]} \quad K_p = \frac{(P_{\text{CO}_2})(P_{\text{H}_2\text{O}})}{P_{\text{NaHCO}_3}} \]

\[ \text{HCOOH}(aq) = \text{H}^+(aq) + \text{HCOO}^-(aq) \]

\[ K_c = \frac{[\text{H}^+][\text{HCOO}^-]}{[\text{HCOOH}]} \]

Pure phosgene gas (COCl\(_2\)), 3.00x10\(^{-2}\) mol, was placed in a 1.50L container. It was heated to 800K and the pressure of CO was found to be 0.497 atm. Calculate the equilibrium constant \(K_p\) for the reaction:

\[ \text{CO}(g) + \text{Cl}_2(g) \rightleftharpoons \text{COCl}_2(g) \]

\[ P = \frac{nRT}{V} \]

\[ P = \frac{(0.0300)(0.0821)(800)}{1.50 \text{ L}} \]

\[ P = 1.31 \text{ atm} \]

\[ K_p = \frac{[\text{COCl}_2]}{[\text{CO}][\text{Cl}_2]} = \frac{(0.816 \text{ atm})}{(0.497)(0.497)} = 3.31 \]
Find $K_p$ for the decomposition of phosphorus pentachloride into phosphorus trichloride and chlorine gas, given that the equilibrium partial pressures are 0.875 atm, 0.463 atm, and 1.98 atm respectively at 250°C. (review naming if needed). Then find its $K_c$.

$$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$$

$$K_p = \frac{[PCl_3][Cl_2]}{[PCl_5]} = \frac{(0.463)(1.98)}{0.875} = 1.0477 = 1.05$$

$$K_p = K_c(RT)^\Delta n \Rightarrow K_c = \frac{K_p}{(RT)^\Delta n} = \frac{1.0477}{(0.821 * (250 + 273))^\Delta n}$$

$$= 0.024$$

When is $K_c$ equal to $K_p$?

$$K_p = K_c(RT)^\Delta n$$

When $(RT)^\Delta n = 1$

so $\Delta n = 0$

So when moles gas reactant = moles gas product
Given each of the following equilibrium constants, find the unknown equilibrium constant.

Example 1

1. \[ H_2C_2O_4^{\text{aq}} \rightleftharpoons H^+(\text{aq}) + HC_2O_4^{-}\text{(aq)} \quad K'_{c} = 6.5 \times 10^{-2} \]
2. \[ HC_2O_4^{-}(\text{aq}) \rightleftharpoons H^+(\text{aq}) + C_2O_4^{2-}\text{(aq)} \quad K''_{c} = 6.1 \times 10^{-5} \]
3. \[ H_2C_2O_4(\text{aq}) \rightleftharpoons 2H^+(\text{aq}) + C_2O_4^{2-}(\text{aq}) \quad K_{c} = ? \]

1 & 2 add to equal equation 3, so multiply the \( K_c \) of each.

\[ K_c = (6.5 \times 10^{-2})(6.1 \times 10^{-5}) = 3.97 \times 10^{-6} \]

Example 2

\[ C(s) + CO_2(g) \rightleftharpoons 2CO(g) \quad K'_{p} = 1.3 \times 10^{14} \]

\[ \times 2 \quad \rightarrow \quad 2CO(g) + Cl_2(g) \rightleftharpoons COCl_2(g) \quad K''_{p} = (6.0 \times 10^{-3})^2 \] < squares

\[ C(s) + CO_2(g) + 2Cl_2(g) \rightleftharpoons 2COCl_2(g) \quad K_{p} = ? \]

\[ K_p = K'_{p}(K''_{p})^2 \]
\[ = (1.3\times 10^{14})(6.0\times 10^{-3})^2 \]
\[ = 4.7 \times 10^9 \]
Calculate the number of moles of H₂ that are present at equilibrium if a mixture of 0.300 mol of CO and 0.300 mol of H₂O is heated to 700°C (Kc=0.534) in a 10.0 L container.

\[
H_2(g) + CO_2(g) \rightleftharpoons H_2O(g) + CO(g)
\]

\[
\begin{array}{ccc}
I & 0 & 0 \\
C & 0 & 0 \\
E & +x & +x \\
\end{array}
\]

\[K_c = 0.534 = \frac{(0.0300-x)(0.0300-x)}{(x)(x)} \Rightarrow \sqrt{0.534} = \frac{(0.0300-x)}{x}
\]

\[0.7307 = \frac{(0.0300-x)}{x} \Rightarrow x = 0.0733M = [H_2]
\]

\[\frac{0.7307}{L} \times 10.0L = 7.307 \text{ mol L}^{-1} \Rightarrow 0.173 \text{ mol L}^{-1} \text{ H}_2
\]

At a certain temperature, the equilibrium constant, Kc, for this reaction is 53.3. At this temperature, 0.300 mol of H₂ and 0.300 mol of I₂ were placed in a 1.00L container to react. What concentration of HI is present at equilibrium?

\[
H_2(g) + I_2(g) \rightleftharpoons 2HI(g)
\]

\[
\begin{array}{ccc}
I & 0.300m & 0.300m \\
C & -x & -x \\
E & 0.300-x & 0.300-x \\
\end{array}
\]

\[k_c \times 533 = \frac{(2x)^2}{L(0.300-x)} \Rightarrow 7.301 = \frac{2x}{0.300-x}
\]

\[2.19 - 7.301x = 2x
\]

\[2.19 = 9.301x
\]

\[x = 0.235 \text{ M}
\]

\[[HI] = 2x = 2(0.235) = 0.472 \text{ M}
\]
For the decomposition of phosphorous pentachloride to phosphorous trichloride and chlorine at 400 K the Kc is 1.1x10⁻². Given that 1.0g of phosphorous pentachloride is added to a 250mL reaction flask, find the final concentrations of each species and the percent decomposition.

I 0.0192  0  0  1) Calculate initial [PCl₅]
C  -x  +x  +x  \[ 1.0g \left( \frac{1 \text{ mol}}{208.239g} \right) = 0.019208 \text{ mol} \]
E 0.0192 -x  x  x  \[ \frac{0.0192}{0.250L} \]

2) Make ice chart

3) Fill into \[ K_c = 1.1 \times 10^{-2} = \frac{x^2}{0.0192-x} \] quadratic with no quick tricks

4) \[ x^2 + 1.1 \times 10^{-2}x - 2.112 \times 10^{-4} = 0 \]
\[ \frac{-b \pm \sqrt{b^2-4ac}}{2a} \Rightarrow x = 0.0101 \text{ M} = [PCl₃] = [Cl₂] \]
\[ [PCl₅] = 0.0192 - 0.010 = 0.0092 \text{ M} \]

Recent decamp \[ \frac{0.0101 \times 100}{0.0192} = 53\% \]

Consider the reaction below. A reaction mixture at 780 °C initially contains [CO]=0.500 M and [H₂]=1.00M. At equilibrium, the CO concentration is 0.15 M. What is the value of the equilibrium constant?

\[ CO(g) + 2H₂(g) ⇌ CH₂OH(g) \]

I 0.500  1.00  0
C -x  0.35  -2x = 0.70
E 0.15  0.30  0.35

\[ K_c = \frac{0.35}{0.30^2 \times 0.15} = 25.9 = 26 \]
Week 1 Review Examples:

16.42 For the reaction $2A(g) \rightleftharpoons B(g) + 2C(g)$, a reaction vessel initially contains only $A$ at a pressure of $P_A = 0.296$ atm. At equilibrium, $P_A = 0.0724$ atm.

Calculate the value of $K_p$. (Assume no changes in volume or temperature.)

\[
\begin{align*}
2A & \rightleftharpoons B + 2C \\
I & \ 0.296 \ 0 \ 0 \\
C & \ 0.2236 \ 0.1118 \ 0.2236 \\
E & \ 0.0724 \ 0.1118 \ 0.2236 \\
K_p &= \frac{(C)^2}{(B)(A)} = \frac{0.2236^2}{(0.1118)(0.0724)} = 6.48
\end{align*}
\]

16.45 Consider the following reaction:

$\text{H}_2(g) + \text{I}_2(g) \rightleftharpoons 2\text{HI}(g)$

A reaction mixture in a 3.67 L flask at a certain temperature initially contains 0.763 g $\text{H}_2$ and 96.9 g $\text{I}_2$. At equilibrium, the flask contains 90.4 g HI.

Calculate the equilibrium constant ($K_c$) for the reaction at this temperature.

\[
\begin{align*}
\text{H}_2 + \text{I}_2 & \rightleftharpoons 2\text{HI} \\
I & \ 0.103 \ 0.104 \ 0 \\
C & -0.0963 \ -0.0963 \ +0.193 \\
E & \ 0.0067 \ 0.0077 \ 0.193 \\
K_c &= \frac{0.193^2}{(0.0067)(0.0077)} = 722
\end{align*}
\]

Initially \[
\begin{align*}
0.763 \text{ g (1 mol)} & = 0.103 \text{ M H}_2 \\
3.67 \text{ L}
\end{align*}
\]

Equilibrium \[
\begin{align*}
96.9 \text{ g (1 mol)} & = 0.104 \text{ M I}_2 \\
3.67 \text{ L}
\end{align*}
\]

Equilibrium \[
\begin{align*}
90.4 \text{ g (1 mol)} & = 0.198 \text{ M HI} \\
3.67 \text{ L}
\end{align*}
\]
In the laboratory studying the extraction of iron metal from iron ore, the following reaction was carried out at 1270K in a reaction vessel of volume 10.0 L. At equilibrium the partial pressure of CO was 4.24 bar and that of CO$_2$ was 1.71 bar. The pressure of the CO$_2$ was reduced to 0.62 bar by reacting some of it with NaOH and the system was allowed to reach equilibrium again. What will be the partial pressure of each gas once equilibrium is re-established?

$$FeO(s) + CO(g) ⇌ Fe(s) + CO_2(g)$$

Find $K_p$ using first set of equilibrium pressures:

$$K_p = \frac{[CO_2]}{[CO]} = \frac{1.71}{4.24} = 0.4033$$

Now make ICE chart with new initial conditions:

$$FeO(s) + CO(g) ⇌ Fe(s) + CO_2(g)$$

- 4.24
- $x$
+ $x$

$$4.24 - x$$
$$3.42$$

$$K = 0.403 = \frac{0.62 + x}{4.24 - x}$$

$$1.709 - 0.403x = 0.62 + x$$

$$x = 0.776 M$$ Plug back in
Deriving the Van't hoff equation.
We'll do this on the document camera.

\[
\ln \frac{K_2}{K_1} = \frac{\Delta H}{R} \left( \frac{1}{T_1} - \frac{1}{T_2} \right)
\]

\[\Delta G_1 = -RT_1 \ln K_1 \quad \Delta G_2 = -RT_2 \ln K_2\]

Solve each block

\[\ln K_1 = \frac{-\Delta G_1}{RT_1} \quad \ln K_2 = \frac{-\Delta G_2}{RT_2}\]

Subtract each side from each other

\[\ln K_1 - \ln K_2 = \frac{-\Delta G_1}{RT_1} + \frac{\Delta G_2}{RT_2}\]

\[\ln K_1 - \ln K_2 = -\frac{1}{R} \left( \frac{\Delta H_1 - T_1 \Delta S_1}{T_1} - \frac{\Delta H_2 - T_2 \Delta S_2}{T_2} \right)\]

rearranging

\[\ln K_1 - \ln K_2 = -\frac{1}{R} \left( \frac{\Delta H_1}{T_1} - \frac{\Delta H_2}{T_2} - \Delta S_1 + \Delta S_2 \right)\]

\[\Delta H \text{ is temp dependent so can be factored}\]

\[\ln K_1 - \ln K_2 = -\frac{\Delta H}{R} \left( \frac{1}{T_1} - \frac{1}{T_2} \right)\]

\[\ln K_2 - \ln K_1 = \frac{\Delta H}{R} \left( \frac{1}{T_1} - \frac{1}{T_2} \right)\]

\[\ln \frac{K_2}{K_1} = \frac{\Delta H}{R} \left( \frac{1}{T_1} - \frac{1}{T_2} \right)\]