

# Chemical Equilibrium

WORKOUT Example CHM 101

# Example 1: Thermal Decomposition of $\text{NH}_4\text{SH}_{(s)}$

Consider the thermal decomposition of  $\text{NH}_4\text{SH}_{(s)}$ : This also is related to  $K_{sp}$

but since  $\text{NH}_4\text{SH}$  is a solid, we get:



$$K_c = \frac{[\text{NH}_3][\text{H}_2\text{S}]}{[\text{NH}_4\text{SH}]}$$

As for  $K_p$ , it is the same as  $K_c$ , but instead of brackets  $[ ]$ ,  $K_p$  uses parentheses  $( )$ :

$$K_c = \frac{[\text{NH}_3][\text{H}_2\text{S}]}{[1]}$$

$$K_c = [\text{NH}_3][\text{H}_2\text{S}]$$

$$K_p = \frac{(\text{NH}_3)(\text{H}_2\text{S})}{(\text{NH}_4\text{SH})}$$

$$K_p = \frac{(\text{NH}_3)(\text{H}_2\text{S})}{(1)}$$

$$K_p = (\text{NH}_3)(\text{H}_2\text{S})$$

Example 2: Hydrogen and Iodine. Consider the double replacement reaction of hydrogen and iodine gas:



$$K_c = \frac{[HI]^2}{[H_2][I_2]}$$

$$K_p = \frac{(HI)^2}{(H_2)(I_2)}$$

Example 3 Given: NOBr = 0.46 M, NO = 0.1 M, Br<sub>2</sub> = 0.3M

To set up  $K_c$ , it is  $\frac{\text{Products}}{\text{Reactants}}$



$$K_c = \frac{[HI]^2}{[H_2][I_2]}$$

$$K_p = \frac{(HI)^2}{(H_2)(I_2)}$$

Answer:  $K_c = 0.0142$  M

## Example 5

$\text{N}_2\text{O}_4$  (l) is an important component of rocket fuel, At  $25\text{ }^\circ\text{C}$   $\text{N}_2\text{O}_4$  is a colorless gas that partially dissociates into  $\text{NO}_2$ . The color of an equilibrium mixture of these 2 gasses depends on their relative proportions, which are dependent on temperature. Equilibrium is established in the reaction  $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$  at  $25\text{ }^\circ\text{C}$ .

Given: 3.00 L container, 7.64 g  $\text{N}_2\text{O}_4$  1.56 g  $\text{NO}_2$

What is the  $K_c$  for this reaction?

**Step 1: Convert grams to moles**

$$\text{mol } \text{N}_2\text{O}_4 = 7.64 \text{ g} * \frac{1 \text{ mol } \text{N}_2\text{O}_4}{92.01 \text{ g}} = 8.303 * 10^{-2} \text{ mol}$$

$$\text{mol } \text{NO}_2 = 1.56 \text{ g} * \frac{1 \text{ mol } \text{NO}_2}{46.01 \text{ g}} = 3.391 * 10^{-2} \text{ mol}$$

**Step 2: Convert moles to Molarity (moles/L)**

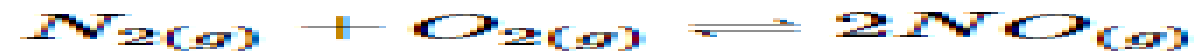
**Step 3: Write the Equilibrium constant for  $K_c$ :**

$$[\text{N}_2\text{O}_4] \text{ M} = \frac{8.303 * 10^{-2} \text{ mol } \text{N}_2\text{O}_4}{3.00 \text{ L}} = 0.0277 \text{ M}$$

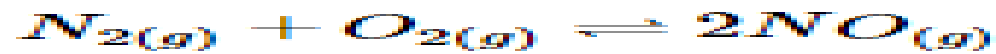
$$[\text{NO}_2] \text{ M} = \frac{3.391 * 10^{-2} \text{ mol } \text{NO}_2}{3.00 \text{ L}} = 0.0113 \text{ M}$$

$$K_c = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]} = \frac{[0.0113]^2}{[0.0277]} = 4.61 \times 10^{-3}$$

Example 6 Given:  $N_2 = 0.79$  moles,  $O_2 = 0.21$  moles, Temp = 2500 K, When equilibrium is established the mole percent of Nitrogen Oxide (NO) at 1.8%. Calculate  $K_p$  for the reaction



1st step: Create an [ICE table](#):



ICE	$N_2$	$O_2$	NO
Initial	0.79 mol	0.21 mol	0 mol
Change	-x	-x	+2x
Equilibrium	(0.79 - x)	(0.21 - x)	(2x + 0)

2nd Step: Find the Mole Percent (%) of products

$$X_{NO} = 0.018$$

$$X_{total} = (0.79 - x) + (0.21 - x) + (2x) = 1$$

$$X_{NO} = \frac{2x}{X_{total}}$$

$$0.018 = \frac{2x}{1}$$

$$x = 0.009$$

$$K_p = \frac{p(NO)^2}{p(N_2)p(O_2)}$$

$$PV = nRT$$

$$P = \frac{\frac{[n(NO)(RT)]^2}{V^2}}{\frac{n(N_2)(RT)}{V} \frac{n(O_2)(RT)}{V}}$$

[Volume Cancels out]

$$P = \frac{n(NO)^2}{n(N_2)n(O_2)}$$

4th Step: Plug in values

$$K_p = \frac{p(\text{NO})^2}{p(\text{N}_2)p(\text{O}_2)}$$

$$PV = nRT$$

$$P = \frac{\frac{[n(\text{NO})(RT)]^2}{V^2}}{\frac{n(\text{N}_2)(RT)}{V} \frac{n(\text{O}_2)(RT)}{V}}$$

[Volume Cancels out]

$$P = \frac{n(\text{NO})^2}{n(\text{N}_2)n(\text{O}_2)}$$

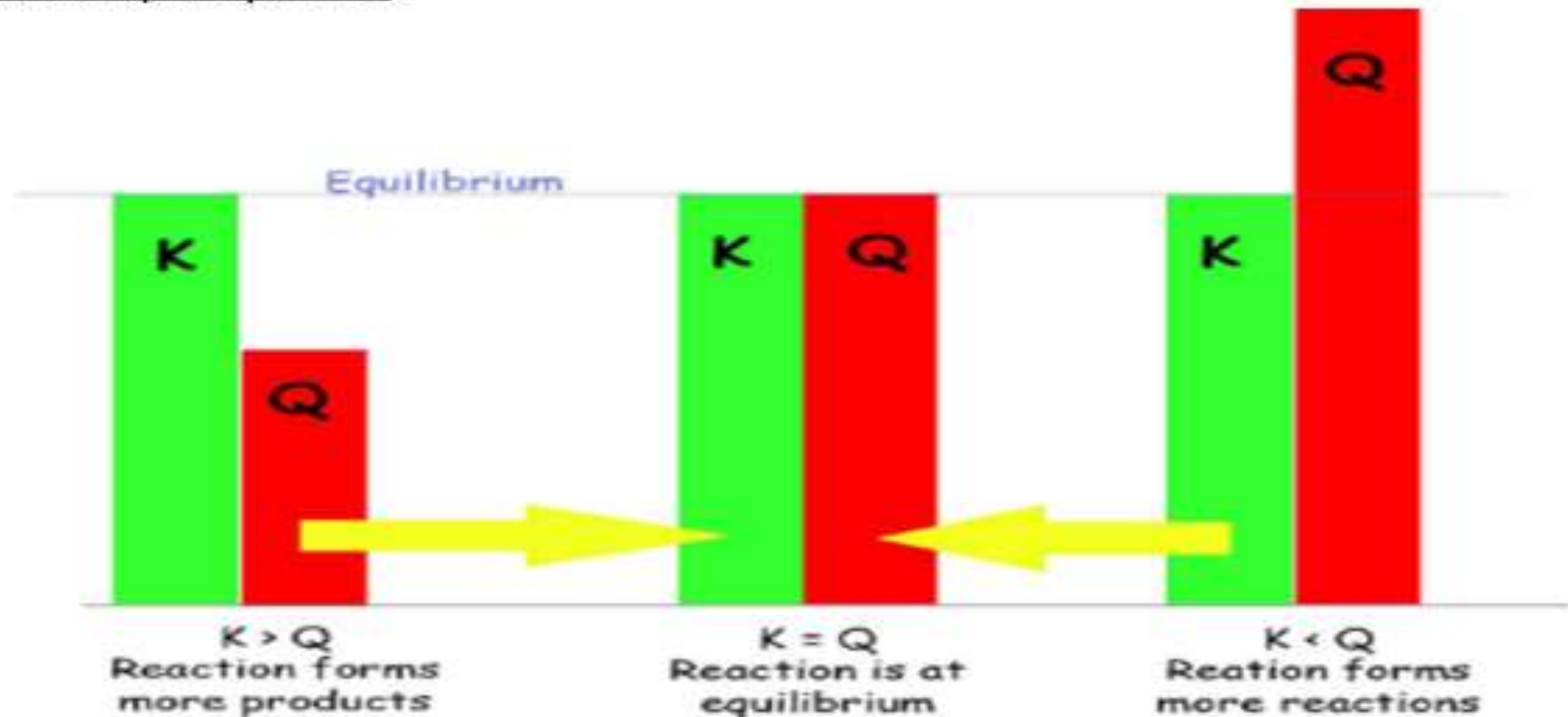
$$x = 0.009$$

$$K_p = 2.1 \times 10^{-3}$$

$$K_p = \frac{(2x)^2}{(0.79 - x)(0.21 - x)}$$

# How the Gas Equilibrium Constants Relate to Reaction Quotient (Q)

The process of finding the Reaction Quotient ( $Q$ ) is the same as finding  $K$ , and  $K_p$ , where the products of the reaction is divided by the reactants of the reaction (Products/Reactants) at any time not necessarily at equilibrium.





# How the Gas Equilibrium Constants Relate to Reaction Quotient (Q)

A trick to remember to which what the reaction will favor is:  
Put:

$K \_ Q$  (in alphabetical order! - or it will not work)



The reaction will favor the reactants because reactants are on the left of the equation.



The reaction will favor the products because products are on the right of the equation.



**Heterogeneous Equilibria:** Reactants/Products in more than one phase.  
For example:



The relationship between the two equilibrium constants are:

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

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

- $\Delta n$  = (Total moles of gas on the products side) - (Total moles of gas on the reactants side). Hence  $\Delta n = (d + c) - (a + b)$  [The lower case numbers are the exponents]
- R is the gas constant found in the

$$\text{ideal gas law } (0.0821 \frac{\text{Liter Atm}}{\text{Mole Kelvin}})$$

# Definition of $K_c$ and $K_p$

$K_c$  is an equilibrium constant in terms of molar **concentrations** and is usually defined as:

in the general reaction,  $aA + bB \rightleftharpoons cC + dD$

$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

If a large  $K_c$  is formed then there are more products formed. Inversely, a small  $K_c$  indicates that the reaction favors the reactants.  $K_p$  is an equilibrium constant in terms of **partial pressures**. and is usually defined as:

$$K_p = \frac{(C)^c (D)^d}{(A)^a (B)^b}$$

for the general reaction



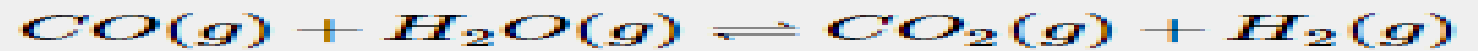
**Homogeneous Equilibria:** Reactants/Products all in a single phase. For example:



### Example 7: Relating K to Q

Given:  $K_c = 1.00$  at about 1100 K  
 $CO = 1.00$  mol  $H_2O = 1.00$  mol  $CO_2 = 2.00$  mol  $H_2 = 2.00$  mol.

Compared with their initial amounts, which of the substances will be present in a greater amount and which is in a lesser amount when equilibrium is established?



Step 1: Write out the expression for  $Q_c$

$$Q_c = \frac{[CO_2][H_2]}{[CO][H_2O]}$$

Step 2: Plug in the number of Molarity, since volume is not given, assume it is 1 Liters

$$Q_c = \frac{[2.00][2.00]}{[1.00][1.00]}$$

$$K_c = 1.00 \text{ (unitless)} \quad Q_c = 4.00 \text{ (unitless)}$$

Step 3: Compare  $K_c$  with  $Q_c$

$$K_c < Q_c$$

$$1.00 < 4.00$$

$$Q_c = 4.00$$

Therefore, the reaction will shift to the LEFT towards the reactants.

### Example 8

A mixture of hydrogen, iodine, and hydrogen iodide, each at 0.0020 M, was introduced into a container heated to 783 Kelvins. At this temperature  $K_c=46$ , Predict if more HI or less will be formed.

Step 1: Write out the reaction



Step 2: Write out the expression for  $Q_c$

$$Q_c = \frac{[HI]^2}{[I_2][H_2]}$$

Step 3: Plug in the Molarity given

$$\text{Molarity} = 0.0020 \text{ M}$$

$$K_c = 46 \text{ (unitless)} \quad Q_c = 1.00 \text{ (unitless)}$$

$$Q_c = \frac{[0.0020]^2}{[0.0020][0.0020]}$$

$$Q_c = 1.00$$

Step 4: Compare  $K_c$  with  $Q_c$

$$K_c > Q_c$$

$$46 > 1.00$$

The reaction will shift to the RIGHT towards the products. Therefore, more HI will be produced.

1. Gaseous Hydrogen Iodide is placed in a closed container at 425 °C, Where it partially decomposes to Hydrogen and Iodine:

$$[\text{HI}] = 3.53 \times 10^{-3} \text{ M} \quad [\text{H}_2] = 4.79 \times 10^{-4} \text{ M} \quad [\text{I}_2] = 4.79 \times 10^{-4} \text{ M}$$

1. What is the value of  $K_c$  at this temperature?



3. Determine values of  $K_c$  from the  $K_p$  value given: (number 7 from p. 655 in the textbook)



$K_p$

1. Write the  $K_p$  for the reaction and state where the reaction is Homogeneous or Heterogeneous.

- 3. The two common chlorides of Phosphorus,  $\text{PCl}_3$  and  $\text{PCl}_5$ , both important in the production of other phosphorous compounds, coexist in equilibrium through: (number 17 from p. 655 in the textbook)



At 250 °C, an equilibrium mixture in a 2.50 L flask contains 0.105 g  $\text{PCl}_5$ , 0.220 g  $\text{PCl}_3$ , and 2.12 g

$K_c$



$$[\text{HI}] = 3.53 \times 10^{-3} \text{ M} \quad [\text{H}_2] = 4.79 \times 10^{-4} \text{ M} \quad [\text{I}_2] = 4.79 \times 10^{-4} \text{ M}$$

$$K_c = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2}$$

$$K_c = \frac{[4.79 \times 10^{-4} \text{ M}][4.79 \times 10^{-4} \text{ M}]}{[3.53 \times 10^{-3} \text{ M}]^2}$$

$$K_c = \frac{[2.29441 \times 10^{-7}] \text{ M}^2}{[1.24609 \times 10^{-5}] \text{ M}^2}$$

$$\underline{K_c = 1.841 \times 10^{-2}}$$



$$K_c = \frac{[\text{NO}]^2}{[\text{N}_2][\text{O}_2]}$$

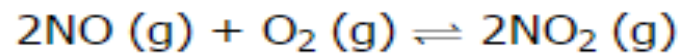
The reaction is a homogeneous reaction because the reactants/products all have the same phase.



$$K_c = \frac{[\text{H}_2\text{O}]}{[\text{H}_2]}, \text{ FeO and Fe are solids so they are not included in equilibrium constants.}$$

The reaction is a heterogeneous reaction because the reactants/products have different phases.

3. Converting to  $K_c$  from  $K_p$  (number 7 from p. 655 in the textbook)



$$K_p = 1.48 \times 10^4 \text{ at } 184^\circ\text{C}$$

We know that  $K_p = K_c (RT)^{\Delta n}$ , we are given  $K_p$  but not  $K_c$ , you can rearrange the equation to:  $K_c = \frac{K_p}{(RT)^{-\Delta n}}$

Which can also be written as:  $K_c = K_p (RT)^{-\Delta n}$

Now that we have your formula, we need to convert  $184^\circ\text{C}$  to Kelvin,  $K = 184 + 273 = 457\text{K}$



Which can also be written as:  $K_c = K_p (RT)^{-\Delta n}$

Now that we have your formula, we need to convert 184 °C to Kelvin,  $K = 184 + 273 = 457K$

$$K_c = \frac{[NO_2]^2}{[NO]^2[O_2]}$$

$-\Delta n = (\text{total number of moles of products}) - (\text{total number of moles in reactants})$

$$-\Delta n = (2) - (3) = -1$$

$$\Delta n = -(-1)$$

$$R = 0.08206 \frac{\text{Liter Atm}}{\text{Mole Kelvin}}$$

Now, plug in all the numbers we found:  $K_c = K_p (RT)^{-\Delta n}$

$$K_c = (1.48 * 10^4)[(0.08206)(457K)]^{-(-1)}$$

Now, plug in all the numbers we found:  $K_c = K_p (RT)^{-\Delta n}$

$$K_c = (1.48 * 10^4)[(0.08206)(457K)]^{-(-1)}$$

$$\underline{K_c = 5.5 * 10^5}$$

[it would be the same if you used this equation:  $K_c = K_p (RT)$ ]

$K_p$



$$K_p = \frac{[C_2H_6]^2 [O_2]}{[C_2H_4]^2 [H_2O]^2}$$

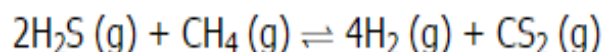
The reaction is a homogeneous reaction because the reactants/products all have the same phase.



$$K_p = \frac{[TiCl_4]}{[Cl_2]^2}, \text{ Ti is a solid so it is not included in equilibrium constants.}$$

The reaction is a heterogeneous reaction because the reactants/products have different phases.

2. Find  $K_p$ , when  $K_c$  is given: (number 8 from p. 655 in the textbook)



$$K_c = 5.27 * 10^{-8} \text{ at } 973 \text{ K.}$$

$$K_p = K_c (RT)^{\Delta n}, \text{ since Temperature is already converted to Kelvin and } R = 0.08206 \frac{\text{Liter Atm}}{\text{Mole Kelvin}}$$

We need to find  $\Delta n$ :  $K_p = \frac{[H_2]^4 [CS_2]}{[H_2S]^2 [CH_4]}$

$\Delta n = (\text{total number of moles of products}) - (\text{total number of moles in reactants})$

$$\Delta n = (5) - (3) = 2$$

We can plug in our numbers:  $K_p = (5.27 * 10^{-8})[(0.08206)(973)]^2$

$$\underline{K_p = 3.6 * 10^{-4}}$$

(b)  $K_p$

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

We need to find:  $\Delta n$  and  $K$ ,  $R = 0.08206 \frac{\text{Latm}}{\text{molK}}$

$$K = 250\text{ }^{\circ}\text{C} + 273 = 523\text{ K}$$

$$\Delta n = (1) - (2) = -1$$

$$K_p = (26.32) [(0.08206)(523)]^{(-1)}$$

$$\underline{K_p = 0.6133}$$

$$\text{PCl}_5 = \frac{0.105\text{g}}{2.50\text{L}} * \frac{1\text{mol}}{137.3\text{g}} \text{ (Molar Mass)} = 2.0173 * 10^{-4}\text{ M}$$

$$\text{PCl}_3 = \frac{0.220\text{g}}{2.50\text{L}} * \frac{1\text{mol}}{137.3\text{g}} \text{ (Molar Mass)} = 6.4093 * 10^{-4}\text{ M}$$

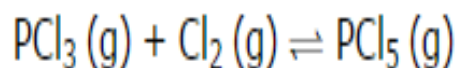
$$\text{Cl}_2 = \frac{2.12\text{g}}{2.50\text{L}} * \frac{1\text{mol}}{70.9\text{g}} \text{ (Molar Mass)} = 0.0119605$$

$$K_c = \frac{[\text{PCl}_5]}{[\text{Cl}_2][\text{PCl}_3]}$$

$$K_c = \frac{[2.0173 * 10^{-4}]}{[0.0119605][6.4093 * 10^{-4}]}$$

$$K_c = 26.32$$

3. The two common chlorides of Phosphorus,  $\text{PCl}_3$  and  $\text{PCl}_5$ , both important in the production of other phosphorous compounds, coexist in equilibrium through:  
(number 17 from p. 655 in the textbook)



At  $250\text{ }^{\circ}\text{C}$ , an equilibrium mixture in a  $2.50\text{ L}$  flask contains  $0.105\text{ g PCl}_5$ ,  $0.220\text{ g PCl}_3$ , and  $2.12\text{ g Cl}_2$ . What are the values of:

(a)  $K_c$

[We need to convert grams to Molarity (mol/L), so we multiply grams with the molar mass and divide by Liters.]

## Relating Gas Equilibrium Constants to Equilibrium (K)

The value of K depends on whether the solution being calculated for is using concentrations or partial pressures. The gas equilibrium constants relate to the equilibrium (K) because they are both derived from [the ideal gas law](#) ( $PV = nRT$ ).

$K_c$  is the concentration of the reaction, it is usually shown as:

$$\frac{c[C]c[D]}{c[A]c[B]}$$

$K_p$  is the amount of partial pressure in the reaction, usually shown as:

$$\frac{p(C)p(D)}{p(A)p(B)}$$

As we have seen above,  $K_p = K_c (RT)^{\Delta n}$ , we can derive this formula from the Ideal Gas Law.

We know that  $K_c$  is in terms Molarity  $\left(\frac{\text{Moles}}{\text{Liters}}\right)$ , and we can also arrange the Ideal Gas Law ( $PV = nRT$ ) as:  $\left(\frac{n}{L}\right) = \left(\frac{P}{RT}\right)$ .

We know that Partial Pressure is directly proportional to Concentration:

$$P = \left(\frac{n}{L}\right) * RT \text{ Pressure can be in units of: Pascal (Pa), Atmosphere (atm), or Torr.}$$

Therefore we can replace  $K_c$  with Molarity: the equation become,  $K_p = K_c (RT)^{\Delta n}$

$$(RT)^{\Delta n} = \frac{(RT)^c (RT)^d}{(RT)^a (RT)^b}$$



$$K = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

