

# Chemical Equilibrium

CHM 101

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# Chemical Equilibrium

- When some types of chemical reactions occur in the gas or solution phases, these reaction attain “chemical equilibrium”, *i.e.*, the reaction does not go to completion, but the reaction vessel will contain both reactant species and product species mixed together.

# Chemical Equilibrium

- This occurs when the concentrations of the reactants stop decreasing, and the concentrations of the products stop increasing.  $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$



(I will use  $\leftrightarrow$  to indicate an equilibrium process in my lecture notes)

# Chemical Equilibrium

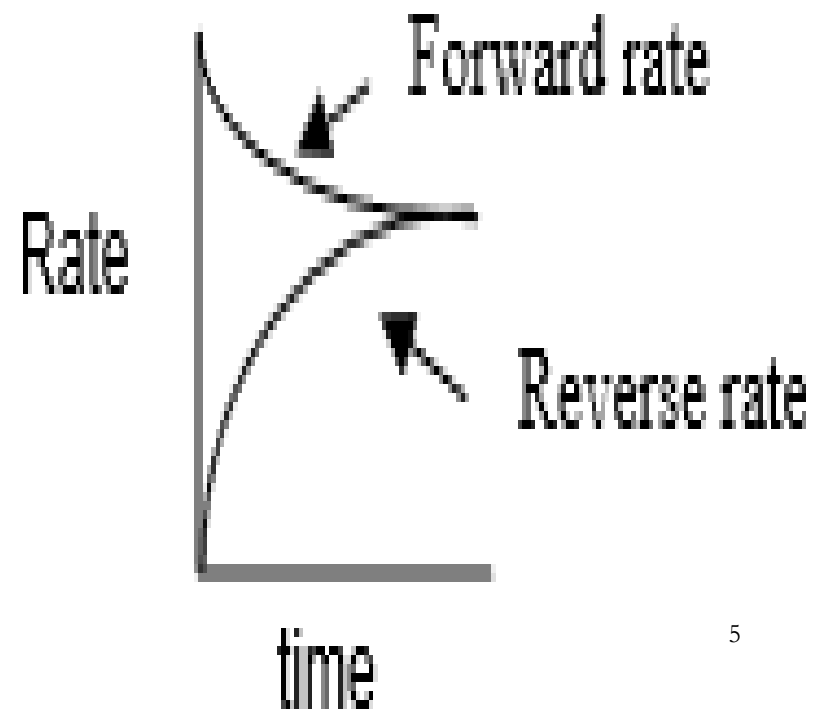
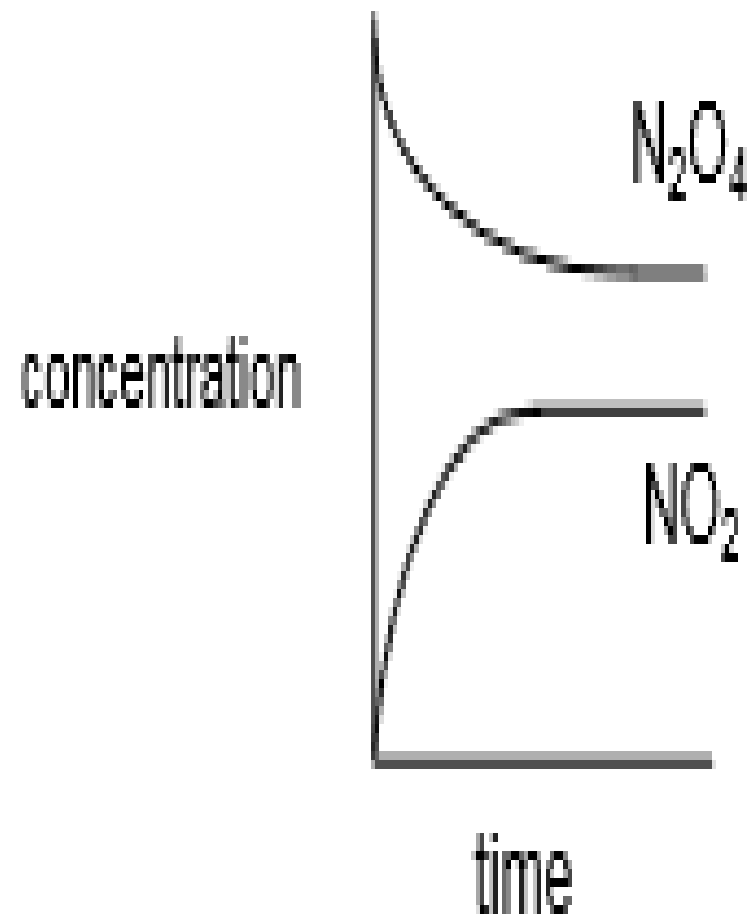
Many chemical reactions do not go to completion but instead attain a state of chemical equilibrium.

Chemical equilibrium: A state in which the rates of the forward and reverse reactions are equal and the concentrations of the reactants and products remain constant.

⇒ Equilibrium is a dynamic process the conversions of reactants to products and products to reactants are still going on, although there is no net change in the number of reactant and product molecules. For the reaction:  $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$

# Chemical Equilibrium

For the reaction:  $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$





# Chemical Equilibrium

- At any given time in a container of  $\text{NO}_2$ , some fraction of the gas will be in the form of  $\text{NO}_2$ , and some fraction will be in the form of  $\text{N}_2\text{O}_4$ .
- $\text{NO}_2$  is a brown gas while  $\text{N}_2\text{O}_4$  is colorless



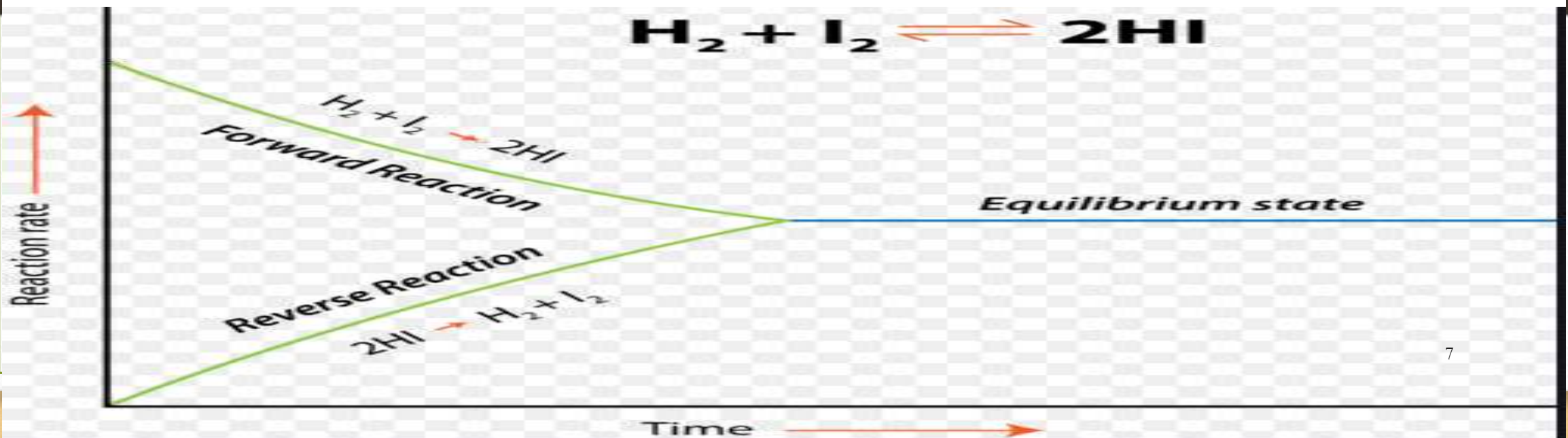
# Chemical Equilibrium

Hydrogen and iodine gases react to form hydrogen iodide according to the following reaction:



Forward reaction:  $\text{H}_2 (g) + \text{I}_2 (g) \rightarrow 2\text{HI} (g)$

Reverse reaction:  $2\text{HI} (g) \rightarrow \text{H}_2 (g) + \text{I}_2 (g)$



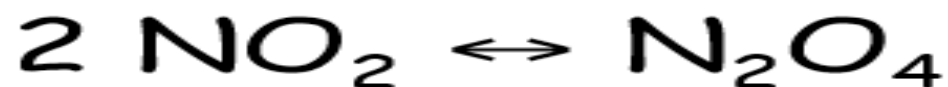
# Chemical Equilibrium

- Initially, only the forward reaction occurs because no HI is present.
- As soon as some HI has formed, it begins to decompose back into  $H_2$  and  $I_2$ .
- Gradually, the rate of the forward reaction decreases while the rate of the reverse reaction increases. Eventually the rate of combination of  $H_2$  and  $I_2$  to produce HI becomes equal to the rate of decomposition of HI into  $H_2$  and  $I_2$ .
- When the rates of the forward and reverse reactions have become equal to one another, the reaction has achieved a state of balance.
- **Chemical equilibrium** is the state of a system in which the rate of the forward reaction is equal to the rate of the



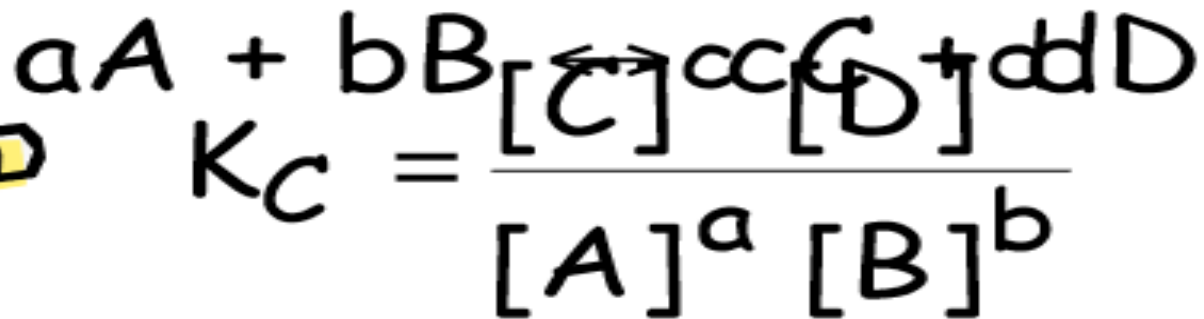
# Chemical Equilibrium

- At any given time in a container of  $\text{NO}_2$ , some fraction of the gas will be in the form of  $\text{NO}_2$ , and some fraction will be in the form of  $\text{N}_2\text{O}_4$ .
- Chemical equilibrium is a dynamic process—an individual molecule will repeatedly move from the  $\text{NO}_2$  form to the  $\text{N}_2\text{O}_4$  form, the overall concentrations of  $\text{NO}_2$  and  $\text{N}_2\text{O}_4$  do not change at a given temperature



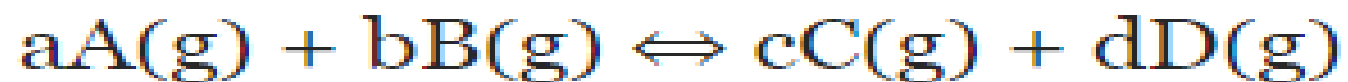
# Equilibrium Constants

- The "equilibrium constant",  $K_{eq}$ , for a chemical reaction indicates whether the reactants or the products will be favored in an equilibrium process
- The equilibrium constant in terms of concentrations is defined as:



# Equilibrium Constant

Law of mass action also forms the basis which states that the rate of a chemical reaction is directly proportional to the product of the concentrations of the reactants raised to their respective stoichiometric coefficients. Therefore, given the reaction -



- By using the law of mass action here,
  - The forward reaction rate would be  $k_+ [A]^a[B]^b$
  - The backward reaction rate would be  $k_- [C]^c[D]^d$

# Equilibrium Constant

where,  $[A]$ ,  $[B]$ ,  $[C]$  and  $[D]$  being the active masses and  $k_+$  and  $k_-$  are rate constants of forward and backward reactions, also the  $a$ ,  $b$ ,  $c$ ,  $d$  are the stoichiometric coefficients related to  $A$ ,  $B$ ,  $C$  and  $D$  respectively. However, at the equilibrium - the forward and the backward rates are equal, stating -

Rate of forward reaction = Rate of backward reaction

$$K_f [A]^a [B]^b = K_b [C]^c [D]^d$$



# Equilibrium Constant

or,

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

or,

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



# Equilibrium Constant

where,

$$K_c = \frac{K_f}{K_b}$$

$K_c$  is the equilibrium constant expressed in terms of the molar concentrations. The equation  $K_c = [C]^c \cdot [D]^d / [A]^a \cdot [B]^b$

**Law of mass action** - The value of the equilibrium constant expression,  $K_c$ , is constant for a given reaction at equilibrium and at a constant temperature.  
 $\Rightarrow$  The equilibrium concentrations of reactants and products may vary, but the value for  $K_c$  remains the same.

## **The Equilibrium Constant**

For a reaction:  $aA + bB \rightleftharpoons cC + dD$

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

# Equilibrium Constant

## The Equilibrium Constant

For a reaction:  $aA + bB \rightleftharpoons cC + dD$

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

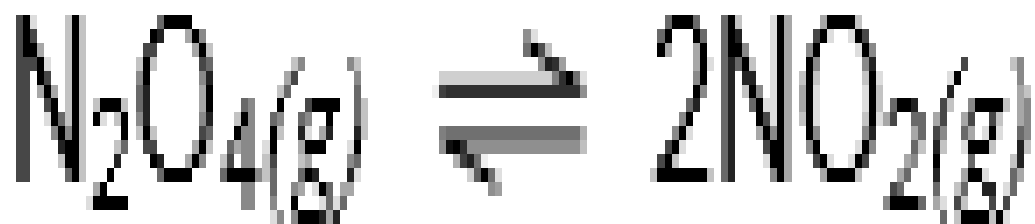
The **equilibrium constant**,  $K_c$ , is the ratio of the equilibrium concentrations of products over the equilibrium concentrations of reactants each raised to the power of their stoichiometric coefficients.

Example. Write the equilibrium constant,  $K_c$ , for  $N_2O_4(g) \rightleftharpoons 2NO_2(g)$

# Equilibrium Constant

- For the  $2 \text{NO}_2 \leftrightarrow \text{N}_2\text{O}_4$  reaction, the equilibrium constant is given as:

$$K_C = \frac{[\text{N}_2\text{O}_4]}{[\text{NO}_2]^2}$$



$$K_c = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]}$$

# Reaction Quotient

The reaction quotient,  $Q$ , is used when questioning if we are at equilibrium. The calculation for  $Q$  is **exactly** the same as for  $K$  but we can only use  $K$  when we know we are at equilibrium. Comparing  $Q$  and  $K$  allows the direction of the reaction to be predicted.

$Q = K$  equilibrium

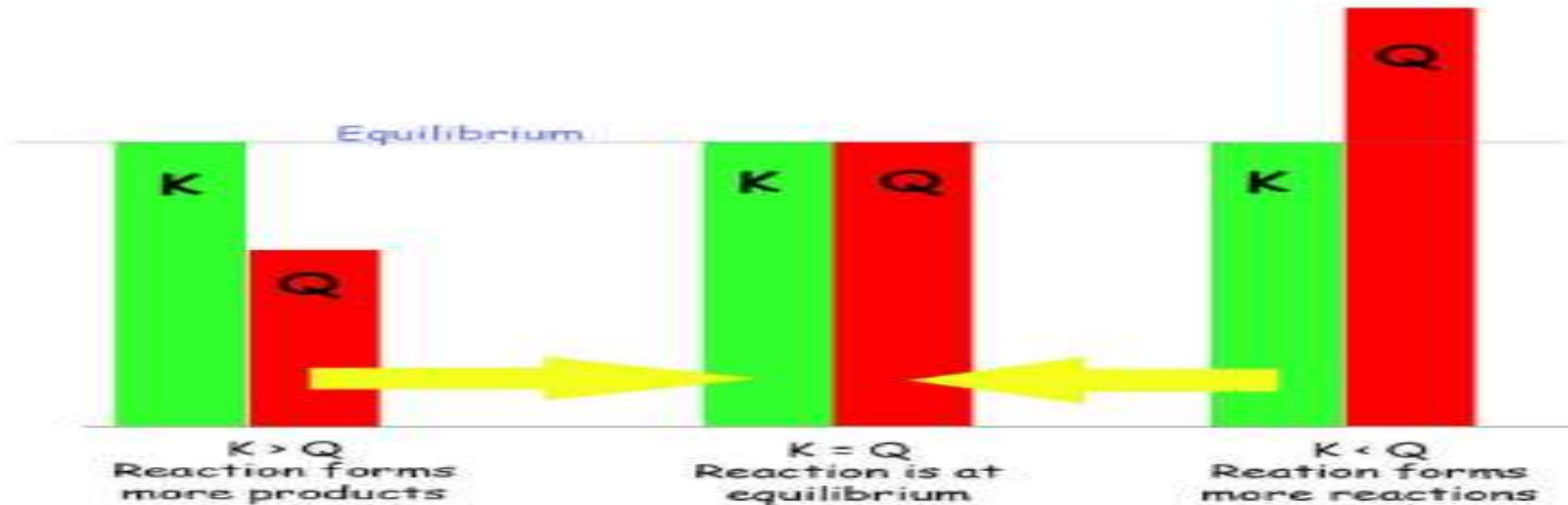
$Q < K$  reaction proceeds to the right to form more products and decrease amount of reactants so value of  $Q$  will increase.

$Q > K$  reaction proceeds to the left to form more reactants and decrease amount of products so value of  $Q$  will decrease.



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

The process of finding the Reaction Quotient ( $Q_c$ ) is the same as finding  $K_c$  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.



# Equilibrium Constant

The **equilibrium constant**,  $K_c$ , is the ratio of the equilibrium concentrations of products over the equilibrium concentrations of reactants each raised to the power of their stoichiometric coefficients.

Example. Write the equilibrium constant,  $K_c$ , for  $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$

## ***Other Characteristics of $K_c$***

- 1) Equilibrium can be approached from either direction.
- 2)  $K_c$  does not depend on the initial concentrations of reactants and products.
- 3)  $K_c$  does depend on temperature.

## ***Magnitude of $K_c$***

- ⇒ If the  $K_c$  value is large ( $K_c \gg 1$ ), the equilibrium lies to the right and the reaction mixture contains mostly products.
- ⇒ If the  $K_c$  value is small ( $K_c \ll 1$ ), the equilibrium lies to the left and the reaction mixture contains mostly reactants.
- ⇒ If the  $K_c$  value is close to 1 ( $0.10 < K_c < 10$ ), the mixture contains appreciable amounts of both reactants and products.

# WHAT DOES THE EQUILIBRIUM CONSTANT TELL US?

If a problem asks you to find which way the reaction will shift in order to achieve equilibrium, and  $K$  is given, you would have to calculate for  $Q$  and compare the two numbers.

When comparing  $K$  and  $Q$ :

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- $K < Q$  : Since there are more products than reactants, the reaction will produce more reactants to reach equilibrium, the reaction favors the reactants.
- $K > Q$  : Since there are more reactants than products, the reaction will produce more products to reach equilibrium, the reaction favors the products.
- $K = Q$  : There is no change in the products nor reactants, so equilibrium is achieved.

# WHAT DOES THE EQUILIBRIUM CONSTANT TELL US?

## Predicting the Direction of Reaction

- The reaction quotient,  $Q$ , is the resulting value when we substitute reactant and product concentrations into the equilibrium expression.
  1. If  $Q > K$ , the reaction will go to the left.
    - The ratio of products over reactants is too large & the reaction will move toward equilibrium by forming more reactants.
  2. If  $Q < K$ , the reaction will go to the right.
    - The ratio of products over reactants is too small & the reaction will move toward equilibrium by forming more products.
  3. If  $Q = K$ , the reaction mixture is already at equilibrium, so no shift occurs.



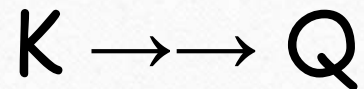
# How the Gas Equilibrium Constants Relate to Reaction Quotient (Q)

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.  $K > Q :$



The reaction will favor the products because products are on the right of the equation.  $K = Q : \text{NO}$

CHANGE



## Work Out Example

- Example. For the reaction,  $B \rightleftharpoons 2A$ ,  $K_c = 2$ . Suppose 3.0 moles of A and 3.0 moles of B are introduced into a 2.00 L flask.

(a) In which direction will the reaction proceed to attain equilibrium?

(b) Will the concentration of B increase, decrease or remain the same as the system moves towards equilibrium?

Solution:

## WRITING EQUILIBRIUM CONSTANT EXPRESSIONS

- Calculating Equilibrium Constants,  $K_c$
- $K_c$  values are listed without units  $\Rightarrow$  don't include units when calculating  $K_c$ .
- If equilibrium concentrations are known, simply substitute the concentrations into the equilibrium constant expression:
- Example.
- For the reaction,  $\text{CO} + 3\text{H}_2 \leftrightarrow \text{CH}_4 + \text{H}_2\text{O}$ ,
- Calculate  $K_c$  from the following equilibrium concentrations:

# Homogeneous & Heterogeneous Equilibria

- Homogeneous equilibria: reactants and products exist in a single phase.
- For the gas phase reaction:



- The equilibrium constant with the concentrations of reactants and products expressed in terms of molarity,  $K_c$ , is:

$$K_c = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]}$$



# Homogeneous & Heterogeneous Equilibria



$$K_c = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]}$$

$$K_c = \frac{[\text{N}_2\text{O}_4]}{[\text{NO}_2]^2}$$

**Homogeneous equilibria:** reactants and products exist in a single phase.

For the gas phase reaction:  $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$

The equilibrium constant with the concentrations of reactants and products expressed in terms of molarity,  $K_c$ , is:

$$K_c = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]}$$

# Heterogeneous Equilibria and Solvents in Homogeneous Equilibria

- Heterogeneous equilibria: reactants and products are present in more than one phase.
- pure solids and liquids: concentrations of pure solids and liquids are fixed by their density and molar mass (both constants) and do not vary with the amount.

$$[ ] = M = \frac{\text{Density}}{\text{Molar Mass}}$$

$$M = \frac{\text{mol}}{L} = \frac{g}{ml} \times \frac{10^3 \text{ ml}}{1L} \times \frac{\text{mol}}{g}$$

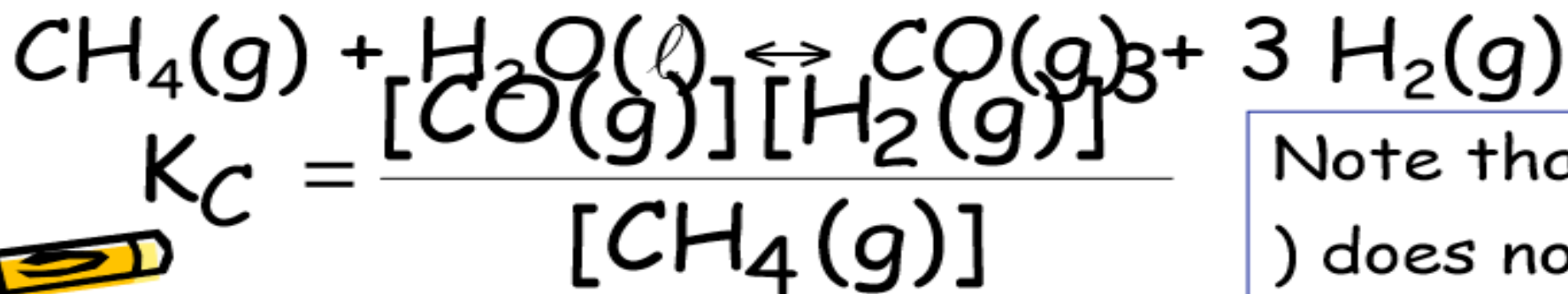
Thus, the concentrations of solids and liquids are incorporated in the  $K_c$  value; they are not part of the variable  $K_c$  expression:



# Homogeneous & Hetroenou Equilibria

- If the reaction involves a pure solid or pure liquid, these species do not appear in the equilibrium constant expression:

Example:



Note that  $\text{H}_2\text{O}(\text{l})$  does not appear in the denominator.

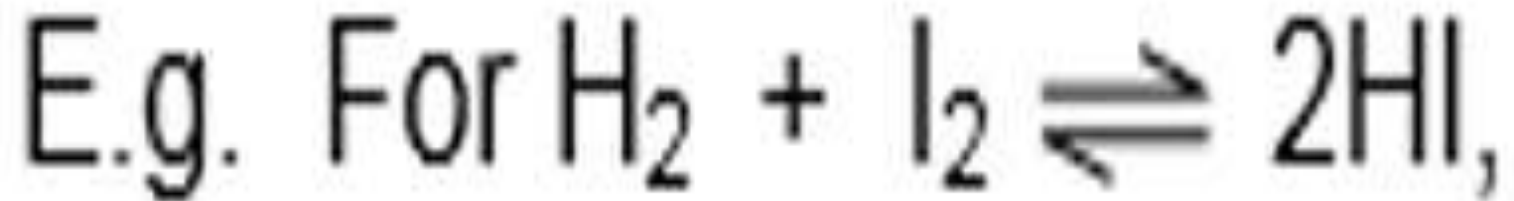
# Heterogeneous equilibria

Omit concentration terms for solids and liquids from  $K_c$  and  $K_p$  expressions; only include terms for gases (g) and aqueous substances (aq).

Example. Write the  $K_c$  expression for  $\text{CaCO}_3(s) \rightleftharpoons \text{CaO}(s) + \text{CO}_2(g)$



Example. For  $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$ , does the equilibrium shift left or right if the pressure is increased?





# Gas Phase Expressions can also be expressed by $K_p$

⇒ The  $K_p$  expression is written using equilibrium partial pressures of reactants & products. For the reaction given above, the  $K_p$  expression is:

$$K_p = \frac{P_{\text{NO}_2}^2}{P_{\text{N}_2\text{O}_4}}$$

Worked out Examples: For:  $\text{H}_2 + \text{I}_2 \leftrightarrow 2\text{HI}$ , does the equilibria shift left or right if we: a) add  $\text{H}_2$ ? b) remove  $\text{I}_2$ ?

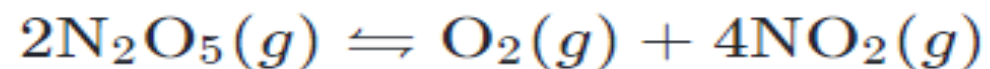
a)

b)



# finding $K_p$ , from partial pressures

- Example 1: finding  $K_p$ , from partial pressures
- Let's try finding  $K_p$ , for the following gas-phase reaction:



We know the partial pressure for each component at equilibrium for some temperature  $T$ , Given that  $P_{\text{N}_2\text{O}_5} = 2.00 \text{ atm}$ ,  $P_{\text{O}_2} = 0.296 \text{ atm}$ ,  $P_{\text{NO}_2} = 1.70 \text{ atm}$ . At temperature  $T$ , what is  $K_p$  for the reaction?

**Solution:** At temperature  $T$ , what is  $K_p$ , for this reaction?

First we can write the  $K_p$  expression for our balanced equation:

We can now solve for  $K_p$ , by plugging in the equilibrium partial pressures in the equilibrium expression:

$$K_p = \frac{(P_{\text{O}_2})(P_{\text{NO}_2})^4}{(P_{\text{N}_2\text{O}_5})^2}$$

$$K_p = \frac{(0.296)(1.70)^4}{(2.00)^2} = 0.618$$

## K<sub>p</sub> is related to K<sub>c</sub>

- Since pressure and molarity are related by the Ideal Gas Law, the following equation relates K<sub>p</sub> and K<sub>c</sub>:

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

where  $R = 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{K} \cdot \text{mol}}$ ; T = temperature in Kelvin

$\Delta n$  = moles of gaseous products – moles of gaseous reactants

⇒ Note that  $K_c = K_p$  when the number of gas molecules are the same on both sides.

# K<sub>c</sub> and K<sub>p</sub>

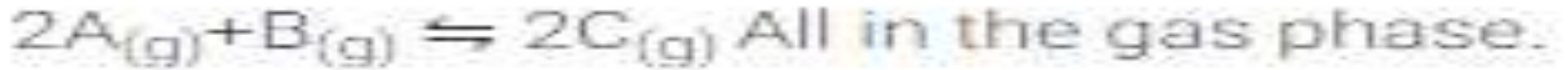


$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

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



Consider an example



The  $K_p$  is given by-

$$K_p = \frac{P_C^2}{P_A^2 P_B} \text{ --- (1)}$$

## Ideal Gas Equation

Each of these ideal gas molecules behaves similarly.

So for each of them,

$$PV = nRT$$

On rearranging we get-

$$P = \frac{n}{V} RT$$



Substituting these in equation (1)

$$P = \frac{n}{V} RT$$

$$\Rightarrow K_p = \frac{[C]^2 (RT)^2}{[A]^2 (RT)^2 [B] (RT)} \Rightarrow K_p = \frac{[C]^2}{[A]^2 [B]} \times \frac{(RT)^2}{(RT)^2 (RT)}$$

On canceling like terms and substituting  $K_c = \frac{[C]^2}{[A]^2 [B]}$  we get-

$$\Rightarrow K_p = \frac{K_c}{RT}$$

Or

$$K_p = K_c(RT)^{-1}$$

In general,

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

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

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

- Where,  $\Delta n$  represents the change in the number of moles of gas molecules. [That is  $\Delta n = \text{product} - \text{reactant}$  in moles only for gas molecules]
- When the change in the number of moles of gas molecules is zero, that is  $\Delta n = 0$

$$\Rightarrow K_p = K_c$$

In general, for any chemical reactions of gas molecules relation between  $K_p$  And  $K_c$  is-

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

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

## Equilibrium Constant and Pressure

- How does the expression for the equilibrium constant change if pressure is used as the variable instead of concentration?
- Using the Ideal Gas Law:

$$P_A = \frac{nRT}{V} = [A]RT$$

$$\therefore [A] = \frac{P_A}{RT}$$





## Example 2: finding $K_p$ , from $K_c$ .

- Now let's look at a different reversible reaction:



If  $K_c$  for this reaction is  $4.5 \times 10^4$  at 400 K, what is the equilibrium constant,  $K_p$ , at the same temperature?

$\Delta n = \text{mol of product gas} - \text{mol of reactant gas}$

$$= 2 \text{ mol NH}_3 - (1 \text{ mol N}_2 + 3 \text{ mol H}_2)$$

$$= -2 \text{ mol gas}$$

will use  $R = 0.08314 \frac{\text{L} \cdot \text{bar}}{\text{K} \cdot \text{mol}}$ .

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

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

$$= (4.5 \times 10^4)(R \cdot 400)^{-2}$$

$$= (4.5 \times 10^4)(0.08314 \cdot 400)^{-2}$$

$$= 41$$

# WORKED OUT EXAMPLES

Example. Does  $K_c = K_p$  for (a)  $\text{H}_2(\text{g}) + \text{F}_2(\text{g}) \rightleftharpoons 2\text{HF}(\text{g})$ ? (b)  $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})$ ?

Example. For the reaction,  $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})$  (a) write the equilibrium constant expression,  $K_p$ . (b) What is the value for  $K_p$  if  $K_c = 2.8 \times 10^2$  at 1000 K?



# Le Chatelier's Principle

- Le Chatelier's principle states that if a **dynamic equilibrium** is disturbed by changing the conditions, the position of equilibrium shifts to counteract the change to reestablish an equilibrium.
- If a chemical reaction is at equilibrium and experiences a change in pressure, temperature, or concentration of products or reactants, the equilibrium shifts in the opposite direction to offset the change.
- This page covers changes to the position of equilibrium due to such changes and discusses briefly why catalysts have no effect on the equilibrium position.

# Recall factors that Le Chatelier's principle states will affect the equilibrium of a system

## Key Points

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Le Chatelier's principle can be used to predict the behavior of a system due to changes in pressure, temperature, or concentration.

- Le Chatelier's principle implies that the addition of heat to a reaction will favor the endothermic direction of a reaction as this reduces the amount of heat produced in the system.
- Increasing the concentration of reactants will drive the reaction to the right, while increasing the concentration of products will drive the reaction to the left.



# Changes in Volume and Pressure

- Because the pressure of gases is related directly to the concentration by  $P = n/V$ ,  
changing the pressure by increasing/decreasing the volume of a container will disturb an equilibrium system.
  - ⇒ If  $P$  increases ( $V$  decreases), the system shifts to the side with a smaller number of gas molecules (this effectively reestablishes equilibrium by decreasing the pressure).
  - ⇒ If  $P$  decreases ( $V$  increases), the system shifts to the side with a greater number of gas molecules.

# Changes in Temperature

- Heat can be considered a reactant in an endothermic rxn and a product in an exothermic rxn. Heat can be considered a reactant in an endothermic rxn and a product in an exothermic rxn.



Endothermic ( $\Delta H > 0$ )



Exothermic ( $\Delta H < 0$ )



# Changes in Temperature

- $K_c$  is larger when the reaction shifts right. This occurs if  $T$  is increased for an Endothermic Reaction or  $T$  is decreased for an Exothermic reaction.
- $K_c$  is smaller when the reaction shifts left. This occurs if  $T$  is decreased for an Endothermic Reaction or  $T$  is increased for an Exothermic reaction.

Example. If the temperature is decreased for the reaction:  $2\text{CO}_2 \rightleftharpoons 2\text{CO} + \text{O}_2$ ,  $\Delta H = 566 \text{ kJ}$ .

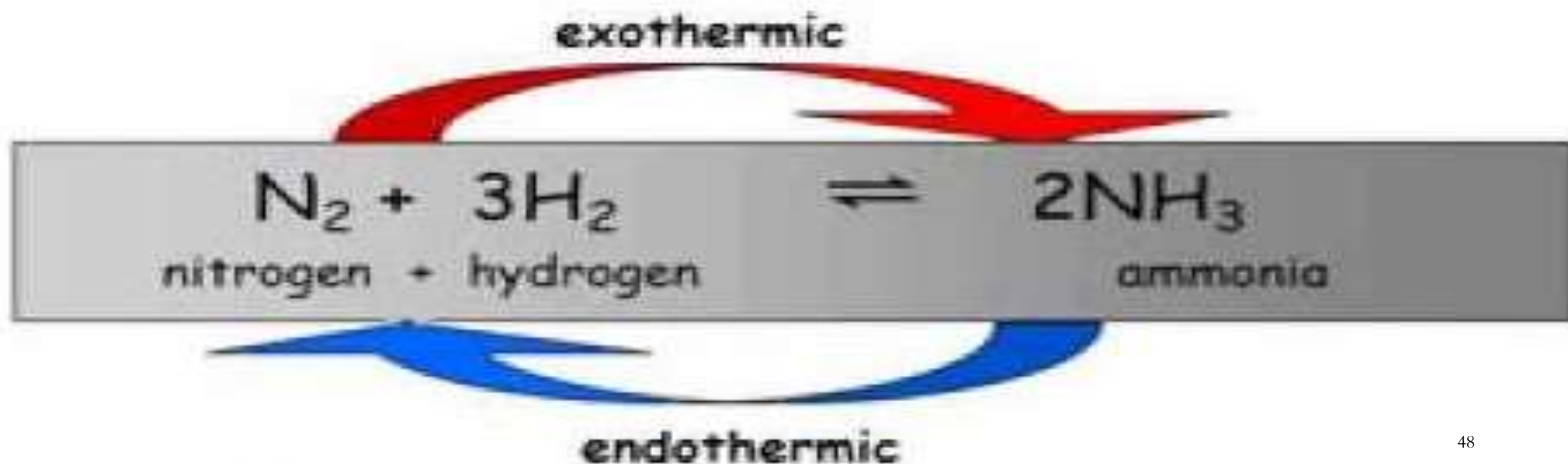
a) Will the equilibrium shift left or right? b) Does  $K_c$  become larger or smaller?



## Energy Changes in reversible reactions

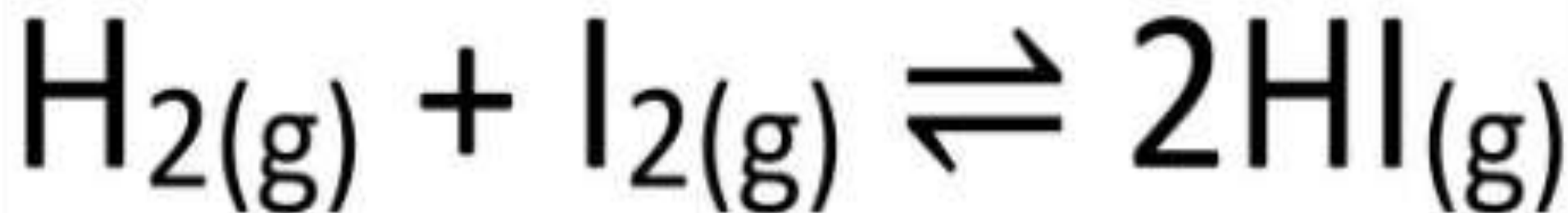
Reversible reactions are **exothermic** (give out heat) in one direction and **endothermic** (take in heat) in the other.

The **same amount** (Joules) of heat energy is given out in one direction and taken in in the other direction.





# Changes in concentration



Equilibrium shifts to left (reactants side)

Adding HI - equilibrium shifts to the left to 'use up' added HI ( $K_c$  doesn't change).

# Le Châtelier's Principle Summary

Variable	Type of Change	Response of System
concentration	increase	shifts to consume some of the added reactant or product
	decrease	shifts to replace some of the removed reactant or product
temperature	increase	shifts to consume some of the added thermal energy
	decrease	shifts to replace some of the removed thermal energy
volume	increase (decrease in pressure)	shifts toward the side with the larger total amount of gaseous entities
	decrease (increase in pressure)	shifts toward the side with the smaller total amount of gaseous entities
<b>Variables That Do Not Affect Chemical Equilibria</b>		
catalysts	—	no effect
inert gases	—	no effect



# Le Chatelier's Principle and Chemical Equilibrium

- When changing temperature is the applied stress, the exothermic or endothermic nature of the reaction must be considered.
- When heat is added to endothermic reactions they shift forward. When heat is removed from endothermic reactions they shift in reverse.
- When heat is added to an exothermic reactions they shift reverse. When heat is removed from exothermic reactions they shift forward.
- Endothermic reactions are identified with a positive  $\Delta H$  or heat written into the equation as a reactant. Exothermic reactions are identified with a negative  $\Delta H$  or heat written into the equation as a product.

# Effect of a Catalyst

- Choosing Optimum Conditions
- Le Chateliers principle can be used to select optimum conditions to form a substance. e.g. To form more  $\text{NH}_3$ , predict the optimum conditions for temperature and pressure.





# Example 1: Thermal Decomposition of $\text{NH}_4\text{SH}_{(s)} \rightleftharpoons \text{NH}_3(g) + \text{H}_2\text{S}(g)$

This also is related to  $K_{sp}$



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

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



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

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

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

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

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

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

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

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

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

## 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<sub>c</sub> it is  $\frac{\text{Products}}{\text{Reactants}}$

$$K_c = \frac{[\text{NO}]^2 [\text{Br}_2]}{[\text{NOBr}]^2}$$

$$K_c = \frac{[0.1]^2 [0.3]}{[0.46]^2}$$



Answer: K<sub>c</sub> = 0.0142 M