

Lecture Presentation

Chapter 14 Chemical Equilibrium

14.1, 14.2, 14.3, 14.4,
14.7, 14.9, 14.10, 14.11,
14.12, 14.14, 14.15, 14.19,
14.23, 14.24, 14.25, 14.27,
14.31, 14.34, 14.50, 14.52,
14.54, 14.56, 14.58, 14.60.
14.68, 14.74, 14.88, 14.98,
14.100, 14.110

John E. McMurry
Robert C. Fay

The Equilibrium State

(kinetic chapter no reversible rxn this chapter all reversible rxn)

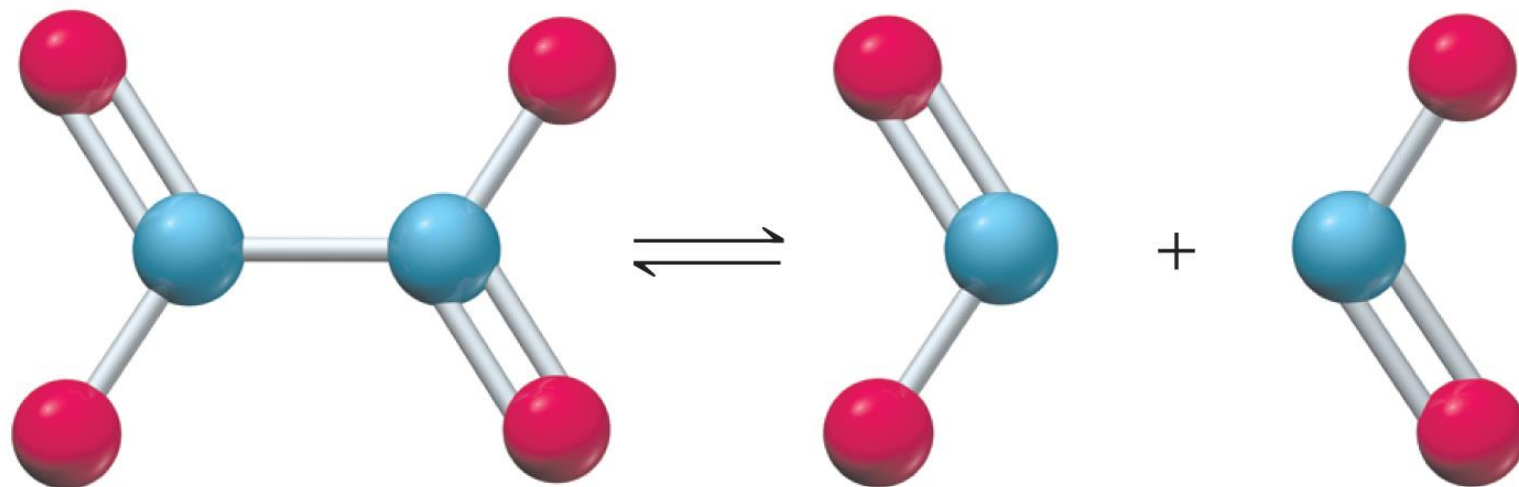
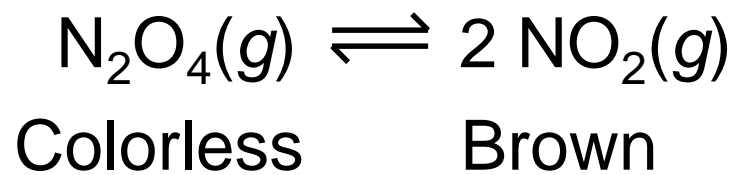
Chemical Equilibrium: The state reached when the concentrations of reactants and products remain constant over time (dynamic equilibrium \rightleftharpoons)
(dynamic equilibrium: rxn to product = rxn to reactant)

This Chapter: How far forward does the reaction go ?
NOT how fast but the ratio of products and reactants



Liquid bromine in
dynamic equilibrium
with gaseous
bromine.

The Equilibrium State

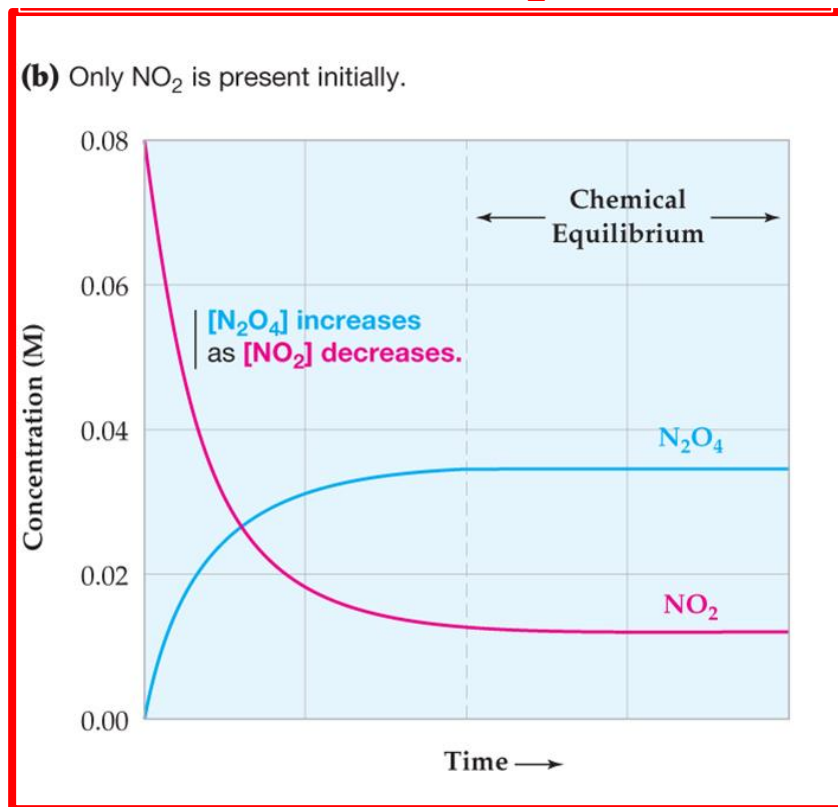
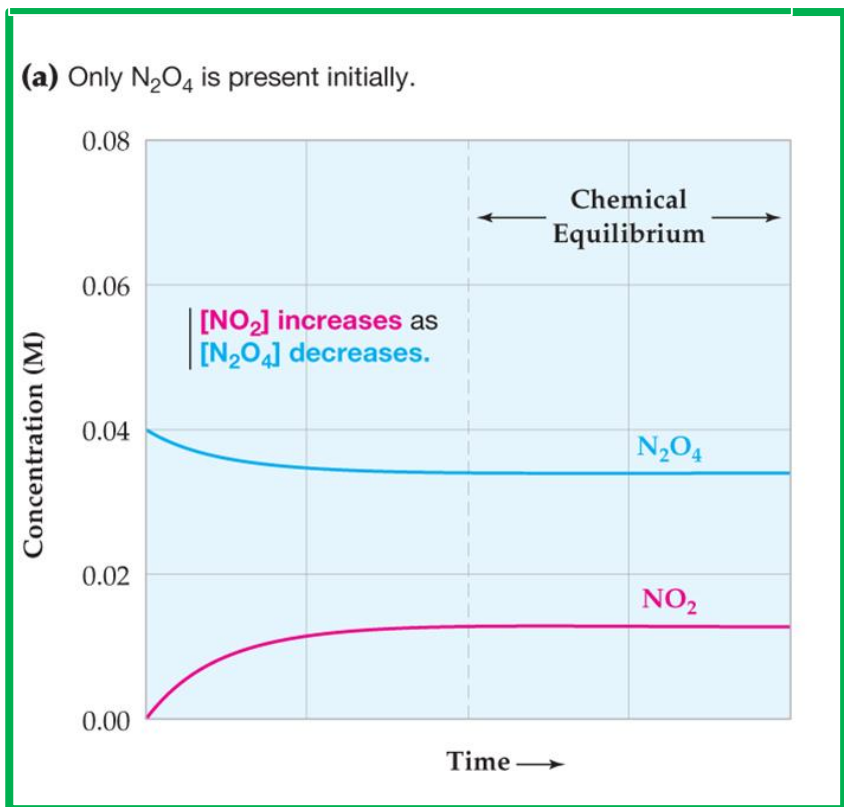


The Equilibrium State



Experiment 1: forward rxn
start with only $\text{N}_2\text{O}_4(\text{g})$

Experiment 2: backward rxn
Start with only $\text{NO}_2(\text{g})$

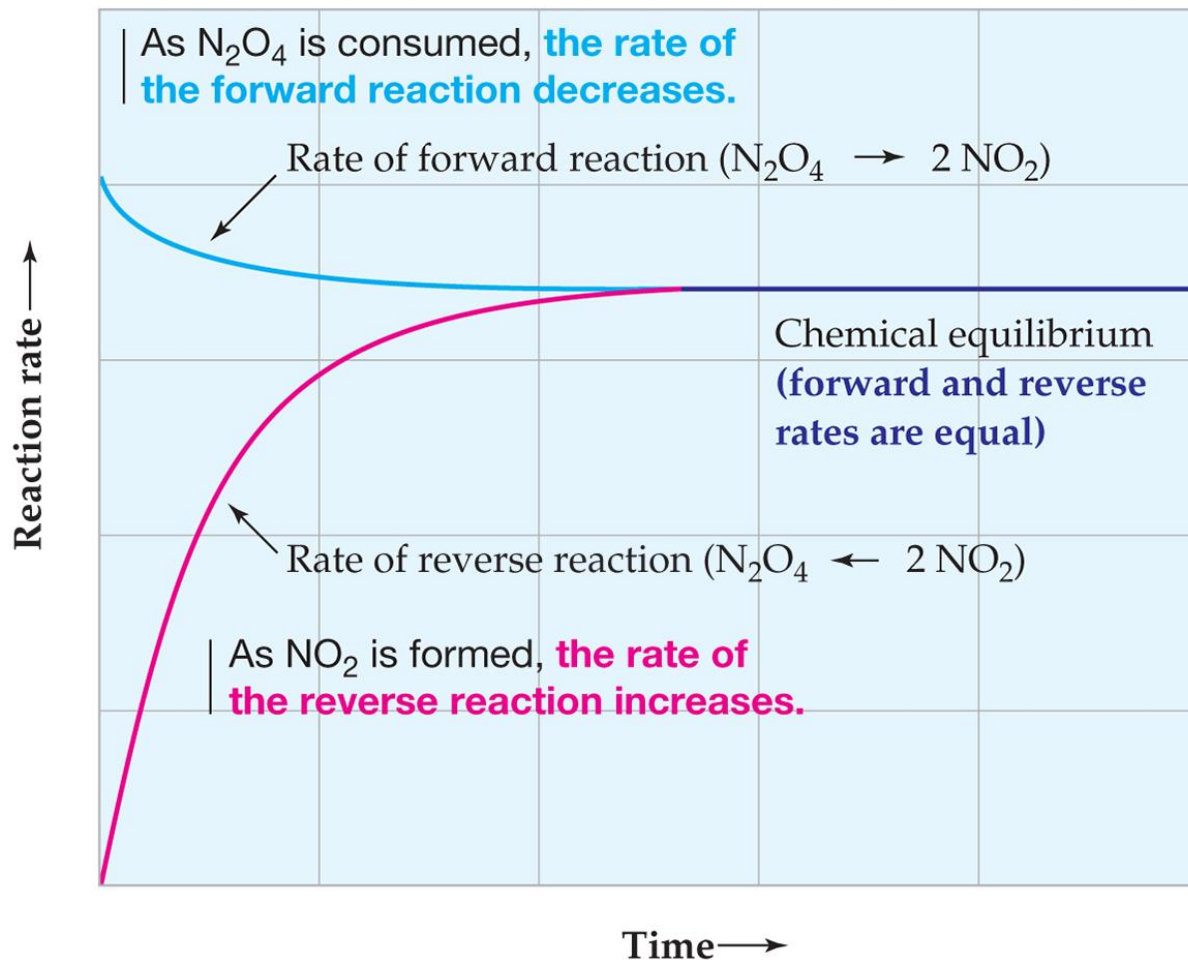


Both experiments end with exactly the same concentration of the N_2O_4 and NO_2

The Equilibrium State

rate forward rxn = rate backward rxn

figure for rxn
Starting from N_2O_4



When the two rates become equal, an equilibrium state is attained and there are no further changes in concentrations.

The Equilibrium Constant K_c



TABLE 14.1 Concentration Data at 25°C for the Reaction
 $\text{N}_2\text{O}_4(g) \rightleftharpoons 2 \text{NO}_2(g)$

Experiment	Initial Concentrations (M)		Equilibrium Concentrations (M)		Equilibrium Constant Expression
	$[\text{N}_2\text{O}_4]$	$[\text{NO}_2]$	$[\text{N}_2\text{O}_4]$	$[\text{NO}_2]$	$[\text{NO}_2]^2/[\text{N}_2\text{O}_4]$
1	0.0400	0.0000	0.0337	0.0125	4.64×10^{-3}
2	0.0000	0.0800	0.0337	0.0125	4.64×10^{-3}
3	0.0600	0.0000	0.0522	0.0156	4.66×10^{-3}
4	0.0000	0.0600	0.0246	0.0107	4.65×10^{-3}
5	0.0200	0.0600	0.0429	0.0141	4.63×10^{-3}

Elementary Reaction Mechanism Step

$$\text{rate} = k_1 [\text{N}_2\text{O}_4]$$

$$\text{rate (back)} = k_2 [\text{NO}_2]^2$$

Why?
Equilibrium
state

The Equilibrium Constant K_c

K_c is the same at the same temperature no matter how much of each reactant or product you have

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

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

Equilibrium constant \swarrow

Products \leftarrow

Reactants \leftarrow

Equilibrium constant expression

For the following reaction: $N_2O_4(g) \rightleftharpoons 2 NO_2(g)$

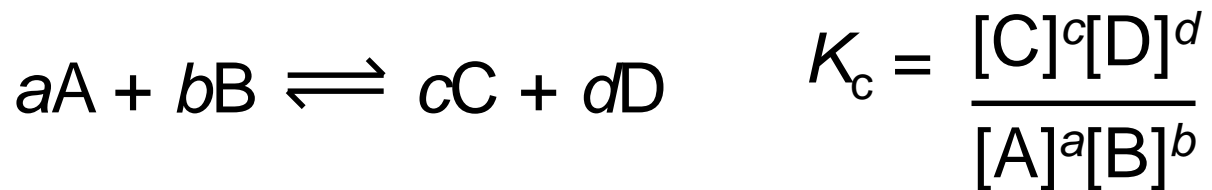
$$K_c = \frac{[NO_2]^2}{[N_2O_4]} = 4.64 \times 10^{-3} \text{ (at } 25 \text{ }^\circ\text{C)}$$

End 2/26W A section

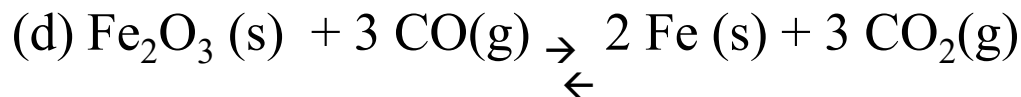
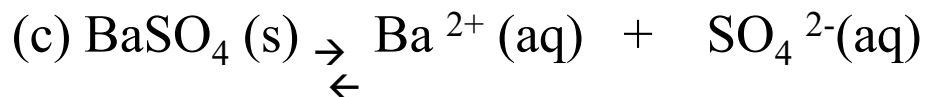
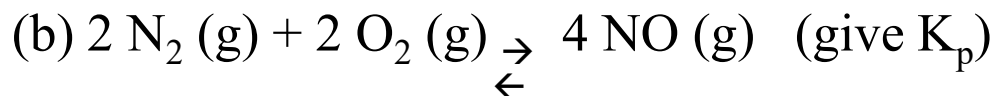
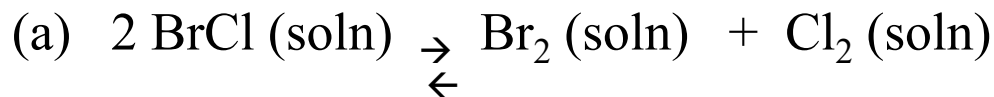
HW 14.1: The Equilibrium Constant K_c

Write the equilibrium constant expression for the

following reactions. (K_p = equilibrium constants with amounts given by pressure of each component)(if gas is in solution use K_c , if gas is in gas state use K_p)



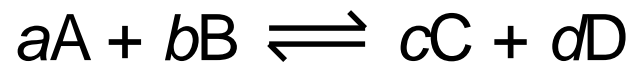
(s) or (l)
leave out of
K expression



HW 14.1: The Equilibrium Constant K_c

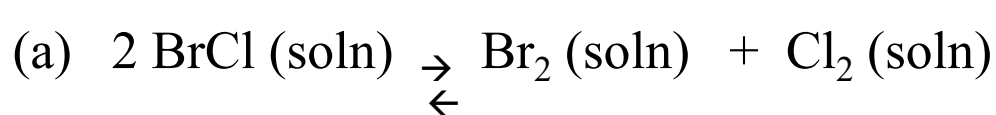
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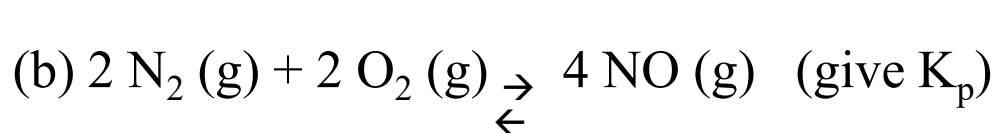


$$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

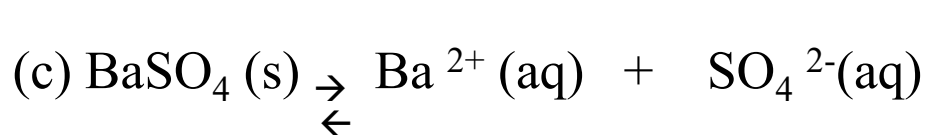
(s) or (l)
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K expression



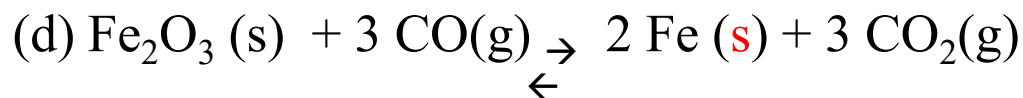
$$K_c = \frac{[\text{Br}_2][\text{Cl}_2]}{[\text{BrCl}]^2}$$



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



$$K_c = [\text{Ba}^{2+}][\text{SO}_4^{2-}]$$



2/28 Friday
C section

$$K_p = \frac{(P_{\text{CO}_2})^3}{(P_{\text{CO}})^3}$$

The Equilibrium Constant K_c

TABLE 14.1 Concentration Data at 25°C for the Reaction
 $\text{N}_2\text{O}_4(g) \rightleftharpoons 2 \text{NO}_2(g)$

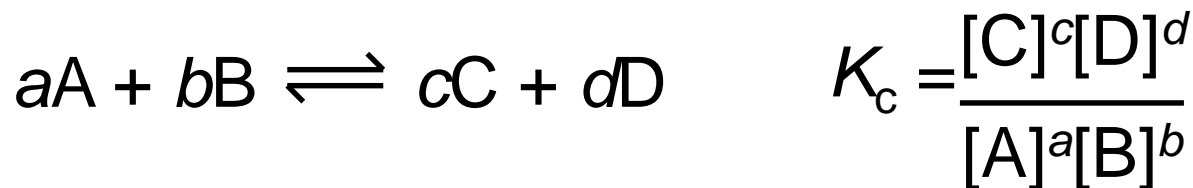
Experiment	Initial Concentrations (M)		Equilibrium Concentrations (M)		Equilibrium Constant Expression
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$$K_c = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]}$$

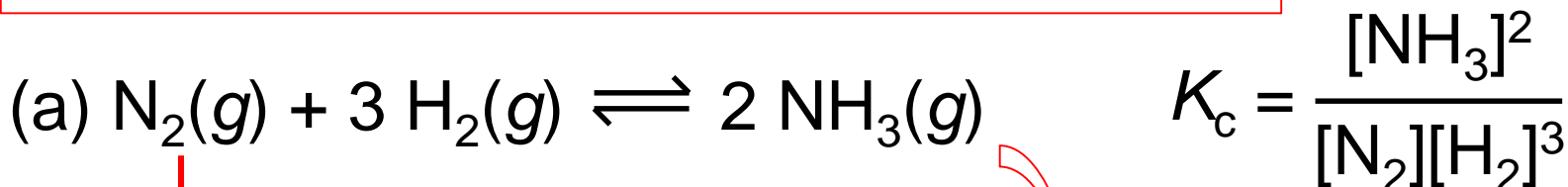
	Experiment 1	Experiment 5
	$\frac{(0.0125)^2}{0.0337} = 4.64 \times 10^{-3}$	$\frac{(0.0141)^2}{0.0429} = 4.63 \times 10^{-3}$

The Equilibrium Constant K_c

The equilibrium constant and the equilibrium constant expression are for the chemical equation *as written*.

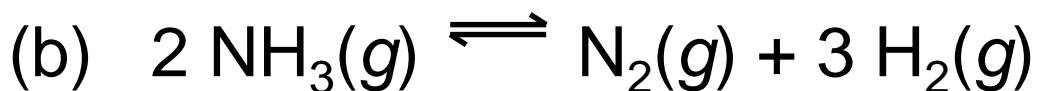


End 2/28F A
section



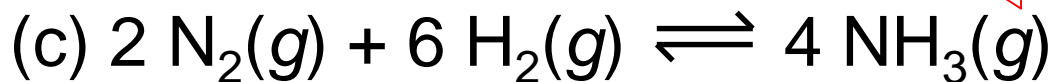
reverse rxn

double rxn



reversed rxn (a) - turn K upside down

$$K_c' = \frac{[\text{N}_2][\text{H}_2]^3}{[\text{NH}_3]^2} = \frac{1}{K_c}$$

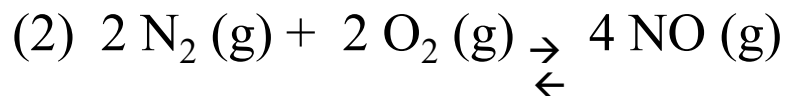
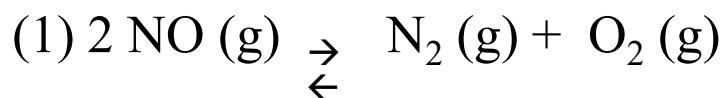
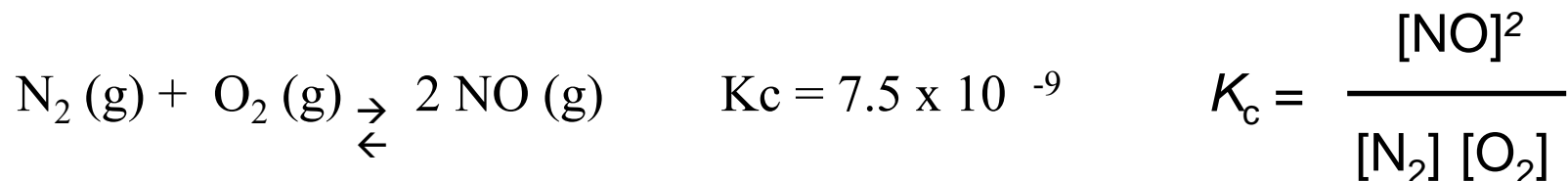


doubled rxn (a) - square all terms

$$K_c' = \frac{[\text{NH}_3]^4}{[\text{N}_2]^2[\text{H}_2]^6} = K_c^2$$

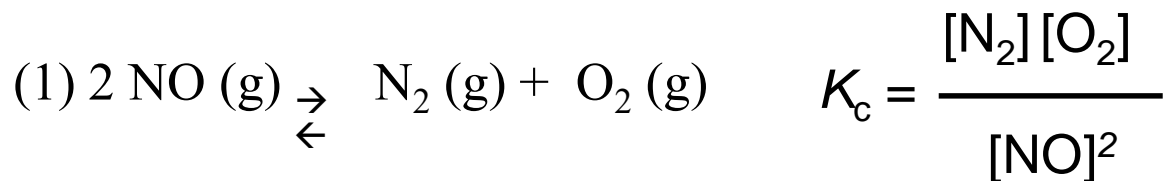
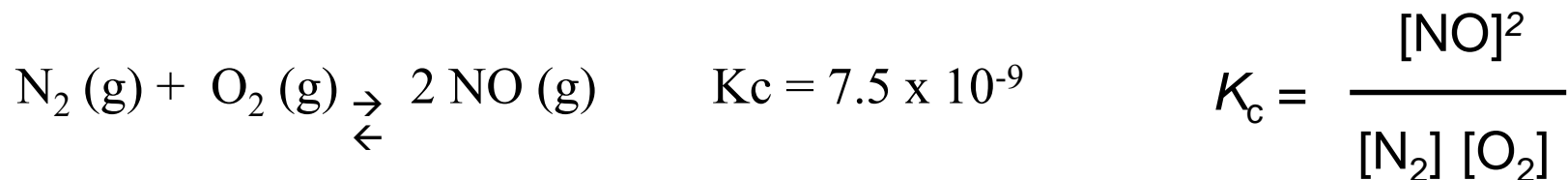
HW 14.2: The Equilibrium Constant K_c

Give the (a) equilibrium constant expression and the (b) numerical value of the equilibrium constant expression for the following reactions given the reaction. (assume dissolved in solvent)

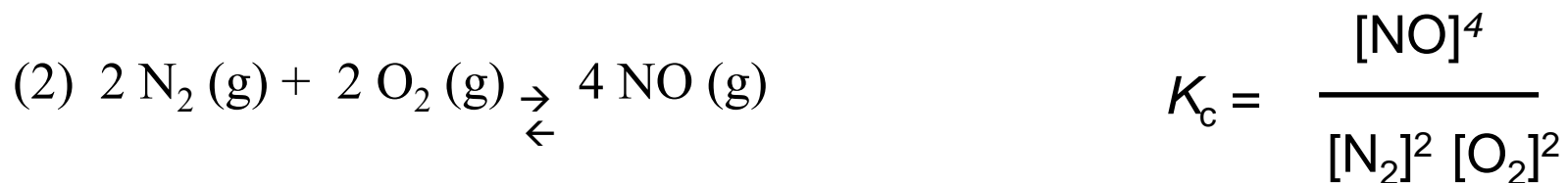


HW 14.2: The Equilibrium Constant K_c

Give the (a) equilibrium constant expression and the (b) numerical value of the equilibrium constant expression for the following reactions given the reaction.

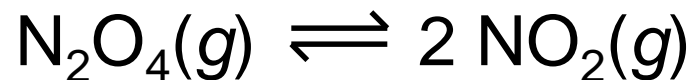


$$K_c = 1 / 7.5 \times 10^{-9} = 1.33 \times 10^8$$



$$K_c = (7.5 \times 10^{-9})^2 = 5.6 \times 10^{-17}$$

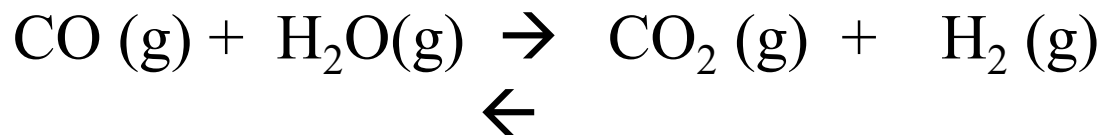
The Equilibrium Constant K_c (K_p)



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

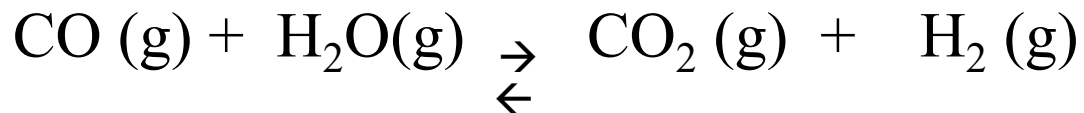
P is the partial pressure of that component.

HW 14.3: The Equilibrium Constant K_c (K_p)



What is the value of K_p at 700 K if the partial pressure in an equilibrium mixture at 700 K are 1.31 atm of CO, 10.0 atm of H₂O, 6.12 atm of CO₂ and 20.3 atm of H₂?

HW 14.3: The Equilibrium Constant K_c (K_p)



What is the value of K_p at 700 K if the partial pressure in an equilibrium mixture at 700 K are 1.31 atm of CO, 10.0 atm of H₂O, 6.12 atm of CO₂ and 20.3 atm of H₂ ?

$$K_p = \frac{P_{\text{CO}_2} P_{\text{H}_2}}{P_{\text{CO}} P_{\text{H}_2\text{O}}} = \frac{(6.12 \text{ atm})(20.3 \text{ atm})}{(1.31 \text{ atm})(10.0 \text{ atm})} = 9.48$$

The Equilibrium Constant K_c

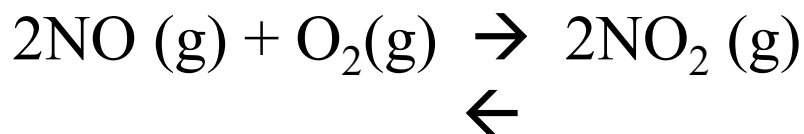
$$K_p = K_c(RT)^{\Delta n} \quad R \text{ is the gas constant, } 0.08206 \frac{\text{L atm}}{\text{K mol}}.$$

T is the absolute temperature (kelvin).

$\Delta n = \#$ moles of gaseous products
minus $\#$ moles of gaseous reactants.

The Equilibrium Constant K_c

$$K_p = K_c(RT)^{\Delta n} \quad \mathbf{R}$$
 is the gas constant, $0.08206 \frac{\text{L atm}}{\text{K mol}}$.
$$\Delta n = (c+d) - (a+b)$$



If $K_c = 6.9 \times 10^5$ at 227°C , what is the value of K_p ?

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

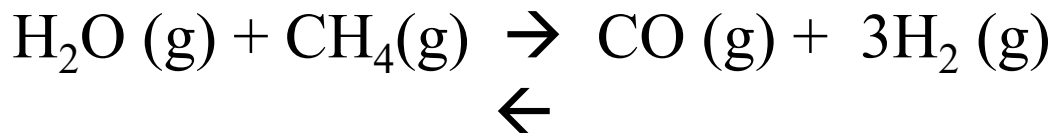
$$\Delta n = (2) - (2+1) = -1$$

$$K_p = (6.9 \times 10^5) [(0.08206 \text{ L atm/mol K})(227+ 273.15)]^{-1}$$

$$K_p = 16811.9 = 1.7 \times 10^4$$

HW 14.4: The Equilibrium Constant K_c

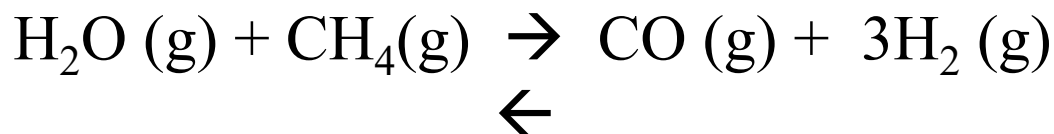
$$K_p = K_c(RT)^{\Delta n} \quad \mathbf{R}$$
 is the gas constant, $0.08206 \frac{\text{L atm}}{\text{K mol}}$.
$$\Delta n = (c+d) - (a+b)$$



If $K_c = 3.8 \times 10^{-3}$ at 1000 K, what is the value of K_p ?

HW 14.4: The Equilibrium Constant K_c

$$K_p = K_c(RT)^{\Delta n} \quad \mathbf{R}$$
 is the gas constant, $0.08206 \frac{\text{L atm}}{\text{K mol}}$.
$$\Delta n = (c+d) - (a+b)$$



End 3/2 Monday
A section

If $K_c = 3.8 \times 10^{-3}$ at 1000 K, what is the value of K_p ?

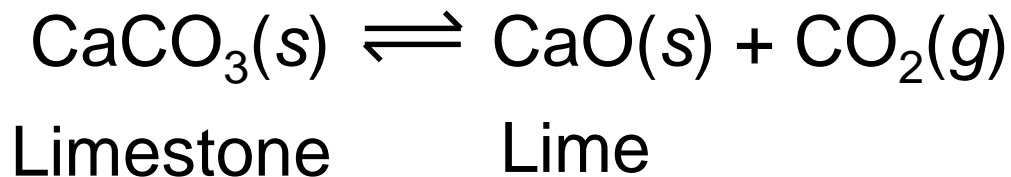
$$K_p = K_c (RT)^{\Delta n}$$

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

$$K_p = (3.8 \times 10^{-3}) [(0.08206 \text{ L atm/mol K})(1000 \text{ K})]^2$$

$$K_p = 25.588 = 26$$

Heterogeneous Equilibria



$$K_c = \frac{[\text{CaO}][\text{CO}_2]}{[\text{CaCO}_3]} = \frac{(1)[\text{CO}_2]}{(1)} = [\text{CO}_2]$$

Pure solids and pure liquids are not included in equilibrium expression.

$$K_c = [\text{CO}_2]$$

Molarity

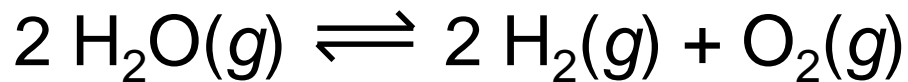
If dissolved.

$$K_p = P_{\text{CO}_2}$$

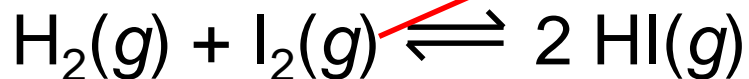
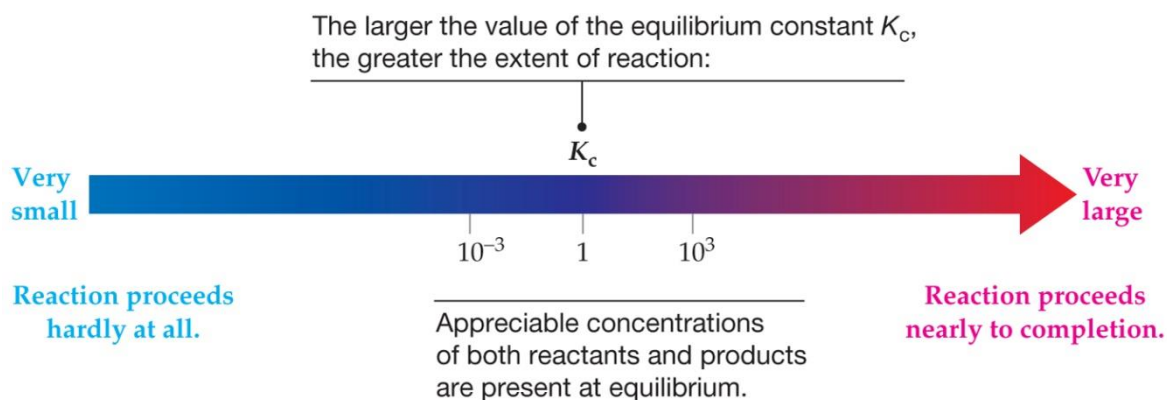
Pressure

If just gas – no solvent.

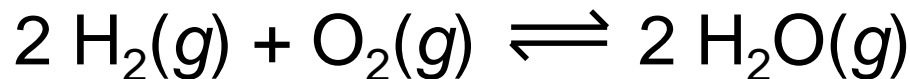
Using the Equilibrium Constant



(at 500 K) $K_c = 4.2 \times 10^{-48}$

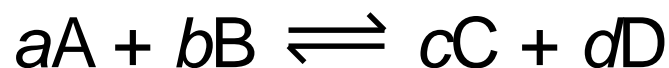


(at 700 K) $K_c = 57.0$



(at 500 K) $K_c = 2.4 \times 10^{47}$

Using the Equilibrium Constant



Reaction quotient: $Q_c = \frac{[C]_t^c [D]_t^d}{[A]_t^a [B]_t^b}$

The **reaction quotient**, Q_c , is defined in the same way as the equilibrium constant, K_c , except that the concentrations in Q_c are **not necessarily equilibrium values**. (value of mixed reaction before equilibrium)

Using the Equilibrium Constant

- If $Q_c < K_c$ net reaction goes from left to right (*reactants to products, \rightarrow*).
- If $Q_c > K_c$ net reaction goes from right to left (*products to reactants, \leftarrow*).
- If $Q_c = K_c$ no net reaction occurs.
(*at equilibrium*)

End 3/3 Tuesday
C section

HW 14.5: Using the Equilibrium Constant

- If $Q_c < K_c$ net reaction goes from left to right (*reactants to products*, \rightarrow).
- If $Q_c > K_c$ net reaction goes from right to left (*products to reactants*, \leftarrow).
- If $Q_c = K_c$ no net reaction occurs. (*at equilibrium*)

If the equilibrium constant $K_c = 2.78 \times 10^{-10}$ and you have a reaction $A + 2 B \rightarrow 3 C + D$ and $[A] = 1.78 \times 10^{-4}$, $[B] = 1.1 \times 10^{-5}$ and $[C] = 1.23 \times 10^{-6}$, $[D] = 7.8 \times 10^{-8}$ does the reaction go forward to produce more products or backwards to produce more reactant or stay in equilibrium ?

HW 14.5: Using the Equilibrium Constant

- If $Q_c < K_c$ net reaction goes from left to right (*reactants to products, \rightarrow*).
- If $Q_c > K_c$ net reaction goes from right to left (*products to reactants, \leftarrow*).
- If $Q_c = K_c$ no net reaction occurs. (*at equilibrium*)

If the equilibrium constant $K_c = 2.78 \times 10^{-10}$ and you have a reaction $A + 2 B \rightarrow 3 C + D$ and $[A] = 1.78 \times 10^{-4}$, $[B] = 1.1 \times 10^{-5}$ and $[C] = 1.23 \times 10^{-6}$, $[D] = 7.8 \times 10^{-8}$ does the reaction go forward to produce more products or backwards to produce more reactant or stay in equilibrium ?

$$K_c = \frac{[C]^3[D]}{[A][B]^2} \quad Q = \frac{[1.23 \times 10^{-6}]^3[7.8 \times 10^{-8}]}{[1.78 \times 10^{-4}][1.1 \times 10^{-5}]^2}$$

$$Q = 6.74 \times 10^{-12} \quad \& \quad 6.74 \times 10^{-12} < 2.78 \times 10^{-10}$$

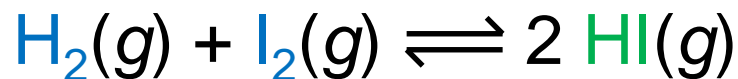
Reaction goes forward until equilibrium is reached.

End 3/2/20
Monday
section A

Using the Equilibrium Constant – ICE method [Initial, Change, Equilibrium] (given K & initial concentrations, calculate equilibrium concentrations)

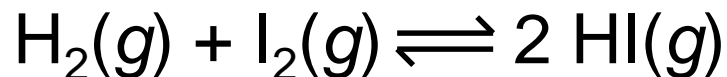
At 700 K, 0.500 mol of HI is added to a 2.00 L container and allowed to come to equilibrium. Calculate the equilibrium concentrations of H₂, I₂, and HI. K_c is 57.0 at 700 K.

Initial: Zero
of reactants



Using the Equilibrium Constant

Set up a table: Initial, Change, Equilibrium



0	0	0.250
+x	+x	-2x
x	x	0.250 - 2x

Initial

Change

Equilibrium

0.500 mol in 2
L is 0.250 M

Substitute values into the equilibrium expression:

$$K_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} \quad 57.0 = \frac{(0.250 - 2x)^2}{x^2}$$

Using the Equilibrium Constant

Solve for “x”:

$$\sqrt{57.0} = \sqrt{\frac{(0.250 - 2x)^2}{x^2}} \quad x = 0.0262$$

$$7.55 = \frac{0.250 - 2x}{x} \quad 7.55x = 0.250 - 2x$$

$$9.55x = 0.250 \quad x = 0.250 / 9.55 = 0.0262$$

Determine the equilibrium concentrations:

$$H_2 = x = 0.0262 \text{ M}$$

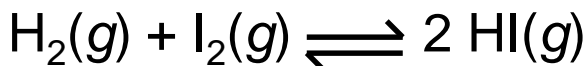
$$I_2 = x = 0.0262 \text{ M}$$

$$HI = 0.250 - 2x = 0.250 - 2(0.0262) = 0.198 \text{ M}$$

HW 14.6: Using the Equilibrium Constant

For $\text{H}_2(g) + \text{I}_2(g) \rightleftharpoons 2\text{HI}(g)$, for initial concentrations of H_2 (0.00623 M), I_2 (0.00414 M) and HI (0.0224 M). Calculate the concentrations at equilibrium given K_c (54.3) assuming forward rxn.

- Set up the ICE table. (where x = change of H_2)
- show the expression for K_c
- If your solution to K_c gives $x = 0.00156$, calculate the concentration of HI .



Initial

Change

Equilibrium

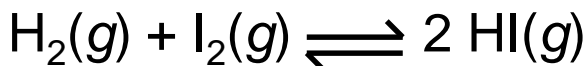
$K_c =$ _____

Solution $x = 0.00156$ $[\text{HI}] =$ _____

HW 14.6: Using the Equilibrium Constant

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- Set up the ICE table. (where x = change of H_2)
- show the expression for K_c
- If your solution to K_c gives $x = 0.00156$, calculate the concentration of HI .



0.00623	0.00414	0.0224	Initial
-x	-x	+2x	Change
0.00623-x	0.00414-x	0.0224 + 2x	Equilibrium

$$K_c = \frac{[0.0224+2x]^2}{[0.00623-x][0.00414-x]} = 54.3$$

Solution $x = 0.00156$ $[\text{HI}] = 0.0224 + 2(0.00156) = 0.0255 \text{ M}$

Using the Equilibrium Constant – showing the math not responsible for on exam – too long for quiz/exam

Step 2: The equilibrium constant is

$$54.3 = \frac{(0.0224 + 2x)^2}{(0.00623 - x)(0.00414 - x)}$$

$$54.3 = \frac{5.02 \times 10^{-4} + 0.0448x + 0.0448x + 4x^2}{(2.58 \times 10^{-5} - 0.00623x - 0.00414x + x^2)}$$

$$54.3(2.58 \times 10^{-5} - 0.0104x + x^2) = 5.02 \times 10^{-4} + 0.0896x + 4x^2$$

$$1.40 \times 10^{-3} - 0.56472x + 54.3x^2 = 5.02 \times 10^{-4} + 0.0896x + 4x^2$$

$$50.3x^2 - 0.654x + 8.98 \times 10^{-4} = 0$$

Using the Equilibrium Constant – showing the math not responsible for on exam – too long for quiz/exam

For equation: $50.3x^2 - 0.654x + 8.98 \times 10^{-4} = 0$

This is a quadratic equation of the form $ax^2 + bx + c = 0$. The solution for a quadratic equation (see Appendix 4) is

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

Here we have $a = 50.3$, $b = -0.654$, and $c = 8.98 \times 10^{-4}$

$$x = \frac{0.654 \pm \sqrt{(-0.654)^2 - 4(50.3)(8.98 \times 10^{-4})}}{2 \times 50.3}$$

$$x = 0.0114 M \quad \text{or} \quad x = 0.00156 M$$

The first solution is physically impossible because the amounts of H_2 and I_2 reacted would be more than those originally present. The second solution gives the correct answer. Note that in solving quadratic equations of this type, one answer is always physically impossible, so choosing a value for x is easy.

At equilibrium, the concentrations are

$$[\text{H}_2] = (0.00623 - 0.00156) \text{ M} = \mathbf{0.00467 \text{ M}}$$

$$[\text{I}_2] = (0.00414 - 0.00156) \text{ M} = \mathbf{0.00258 \text{ M}}$$

$$[\text{HI}] = (0.0224 + 2 \times 0.00156) \text{ M} = \mathbf{0.0255 \text{ M}}$$

Le Châtelier's Principle

Le Châtelier's Principle: If a stress is applied to a reaction mixture *at equilibrium*, net reaction occurs in the direction that relieves the stress. (possible changes)
(equilibrium shifts to undo stress applied to system)

- concentration of reactants or products

- pressure and volume

only effect for rxn with gases, otherwise not much effect of P & V for liquid & solids

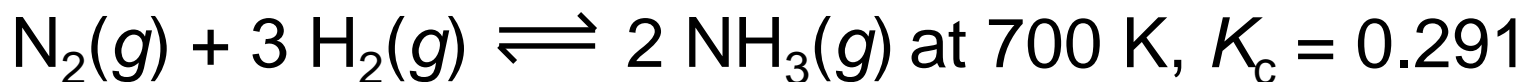
- temperature can be changed. (can predict only if given enthalpy) ($-\Delta H$ exo \rightarrow lower T)

Altering an Equilibrium Mixture: Changes in Concentration

In general, when an equilibrium is disturbed by the **addition or removal of any reactant or product**, Le Châtelier's principle predicts that

- the concentration stress of an **added reactant or product** is relieved by net reaction in the direction that **consumes the added substance**.
- the concentration stress of a **removed reactant or product** is relieved by net reaction in the direction that **replenishes the removed substance**.

Altering an Equilibrium Mixture: Changes in Concentration

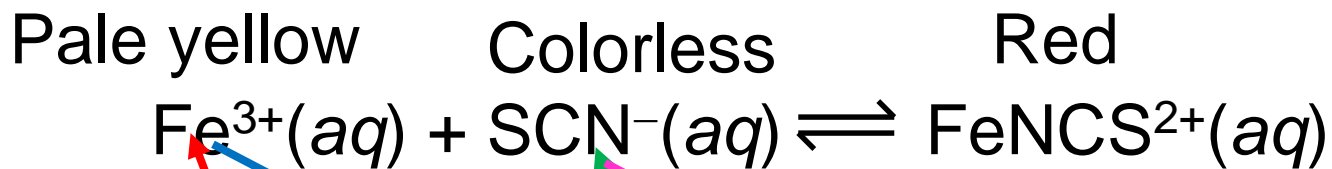


An equilibrium mixture of **0.50 M N₂**, **3.00 M H₂**, and **1.98 M NH₃** is disturbed by increasing the **N₂** concentration to **1.50 M**.

$$Q_c = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} = \frac{(1.98)^2}{(1.50)(3.00)^3} = 0.0968 < K_c$$

Since $Q_c < K_c$, more reactants will be *consumed* and the net reaction will be from *left to right*.

Altering an Equilibrium Mixture: Changes in Concentration



End 3/4/20
A section,
C section

Add subtract Add subtract

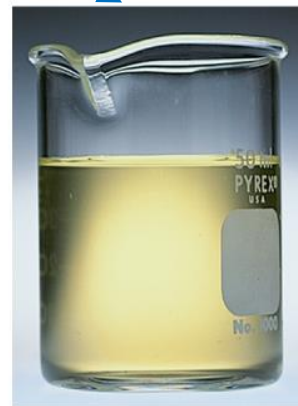
(a) Original solution:
 Fe^{3+} (pale yellow),
 SCN^{-} (colorless),
and FeNCS^{2+} (red).

(b) After adding FeCl_3
to **(a)**: $[\text{FeNCS}^{2+}]$
increases.

(c) After adding KSCN
to **(a)**: $[\text{FeNCS}^{2+}]$
increases.

(d) After adding
 $\text{H}_2\text{C}_2\text{O}_4$ to **(a)**:
 $[\text{FeNCS}^{2+}]$ decreases
as $[\text{Fe}(\text{C}_2\text{O}_4)_3^{3-}]$
increases.

(e) After adding HgCl_2
to **(a)**: $[\text{FeNCS}^{2+}]$
decreases as
 $[\text{Hg}(\text{SCN})_4^{2-}]$
increases.



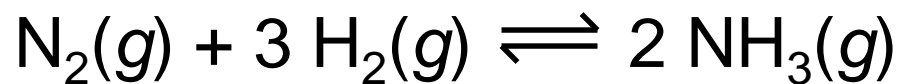
add reactant, rxn goes \rightarrow
add product, rxn goes \leftarrow

subtract reactant, rxn goes \leftarrow
subtract product, rxn goes \rightarrow

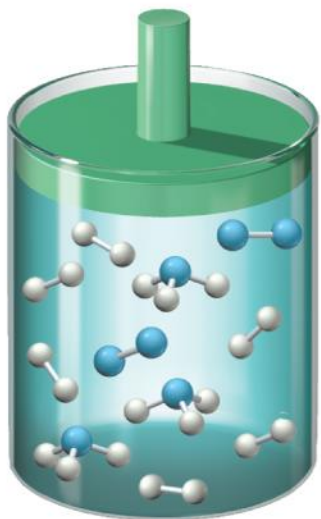
End Quiz 6

End Test 3

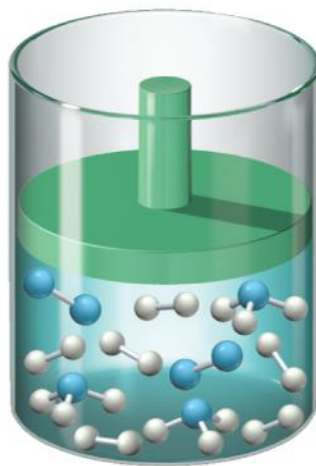
Altering an Equilibrium Mixture: Changes in Pressure & Volume



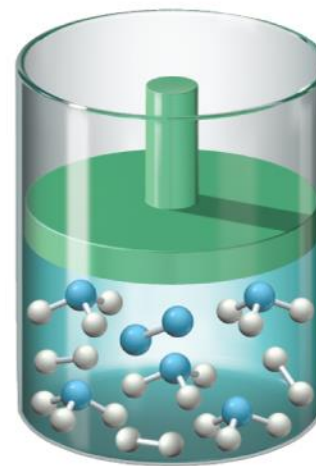
(a) A mixture of gaseous N_2 , H_2 , and NH_3 at equilibrium ($Q_c = K_c$).



(b) When the pressure is increased by decreasing the volume, the mixture is no longer at equilibrium ($Q_c < K_c$).



(c) Net reaction occurs from reactants to products, decreasing the total number of gaseous molecules until equilibrium is re-established ($Q_c = K_c$).



4 moles gas reactant \rightarrow 2 moles gas product (this example)

* increase P or decrease V rxn goes \rightarrow (toward side w less moles gas)

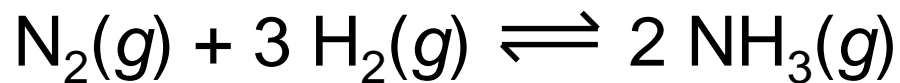
* decrease P or increase V rxn goes \leftarrow (toward side w more moles gas)

Altering an Equilibrium Mixture: Changes in Pressure and Volume

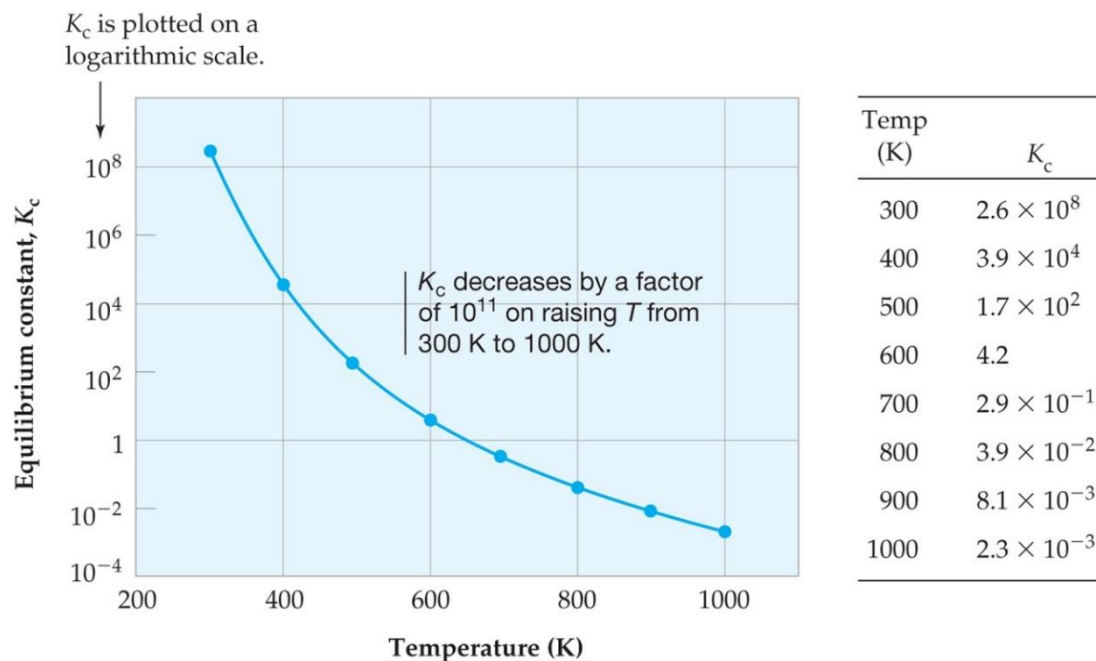
In general, when an equilibrium is disturbed by a **change in volume** that results in a corresponding change in pressure, Le Châtelier's principle predicts that

- an increase in P (reducing V) - rxn goes toward side with **less moles gas**
- a decrease in P (increasing V) - rxn goes toward side with **more moles gas**

Altering an Equilibrium Mixture: Changes in Temperature



$\Delta H^\circ = -2043 \text{ kJ}$ means
+ heat (heat is product)



As the temperature increases, the equilibrium shifts from products to reactants.

exothermic reaction ($-\Delta H^\circ$) rxn releases heat

higher T = add heat rxn goes \leftarrow (to more reactant)

Altering an Equilibrium Mixture: Changes in Temperature

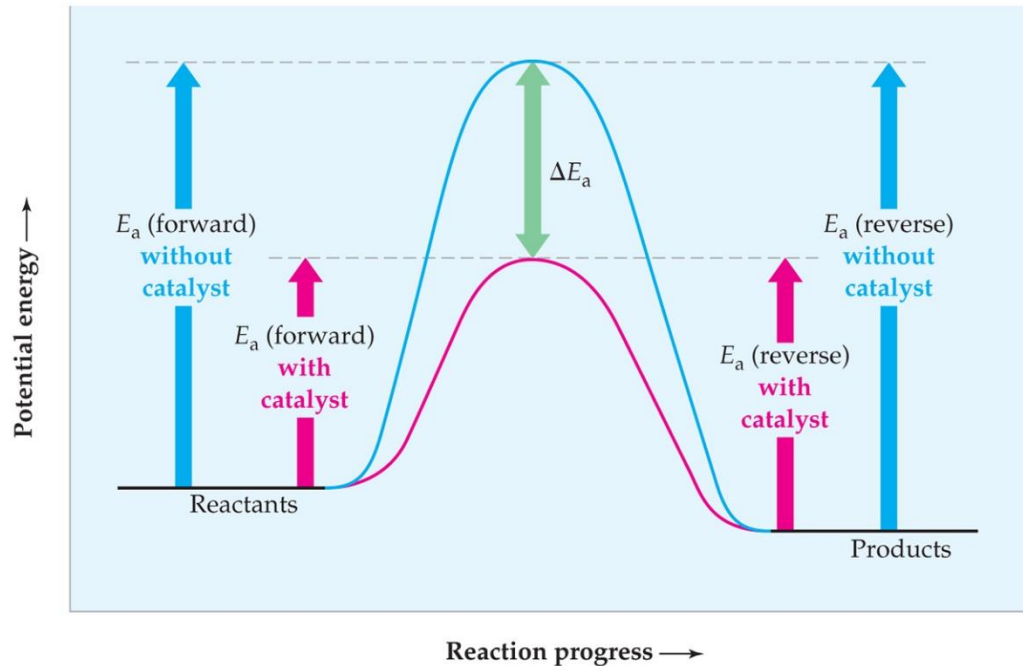
In general, when an equilibrium is disturbed by a change in **temperature**, Le Châtelier's principle predicts that

- the equilibrium constant ($K = \text{product} / \text{reactant}$) for an **exothermic reaction (negative ΔH°)** *decreases* (get more reactant) *as the temperature increases*.
- the equilibrium constant for an **endothermic reaction (positive ΔH°)** *increases* (get more product) *as the temperature increases*.

End 3/6 A, C
section

The Effect of a Catalyst on Equilibrium

The activation energy for the **catalyzed pathway** (red curve) is lower than that for the **uncatalyzed pathway** (blue curve) by an amount ΔE_a .



A catalyst lowers the activation energy barrier for the forward and reverse reactions by the same amount. The catalyst therefore accelerates the forward and reverse reactions by the same factor, and the composition of the equilibrium mixture is unchanged.

catalysis – no effect on equilibrium

activation energy for catalyzed path is lower than for uncatalyzed –

same activation energy for both \rightarrow (forward) and \leftarrow (backward) reactions

HW 14.7: Le Châtelier's Principle: If a stress is applied to a reaction mixture *at equilibrium*, net reaction occurs in the direction that relieves the stress. (equilibrium shifts to undo stress applied to system)

$2 A(g) + B(g) \rightarrow 3 C(g) + 2 D(g)$ $\Delta H = \text{negative}$ (circle one parenthesis under which direction will equilibrium shift each letter)

hint: ΔH negative is exothermic = (+ heat, heat is product)

(a) add A rxn goes [(\rightarrow) or (\leftarrow)]

(b) remove B rxn goes [(\rightarrow) or (\leftarrow)]

(c) add C rxn goes [(\rightarrow) or (\leftarrow)]

(d) remove C rxn goes [(\rightarrow) or (\leftarrow)]

(e) higher T rxn goes [(\rightarrow) or (\leftarrow)] end Test 3 (c) & below

(f) higher P rxn goes [(\rightarrow) or (\leftarrow)]
(assume all reactants & products in reaction are gases)

(g) higher V rxn goes [(\rightarrow) or (\leftarrow)]
(assume all reactant & products in reaction are gases)

HW 14.7: Le Châtelier's Principle: If a stress is applied to a reaction mixture *at equilibrium*, net reaction occurs in the direction that relieves the stress. (equilibrium shifts to undo stress applied to system)

$2 A (g) + B(g) \rightarrow 3 C(g) + 2 D(g)$ $\Delta H = \text{negative}$ (circle one parenthesis under which direction will equilibrium shift each letter)

hint: ΔH negative is exothermic = (+ heat, heat is product)

(a) add A rxn goes [(\rightarrow) or (\leftarrow)]

(b) remove B rxn goes [(\rightarrow) or (\leftarrow)]

(c) add C rxn goes [(\rightarrow) or (\leftarrow)]

(d) remove C rxn goes [(\rightarrow) or (\leftarrow)]

(e) higher T rxn goes [(\rightarrow) or (\leftarrow)] end Test 3 (e) & below

(higher T = add heat)

(f) higher P rxn goes [(\rightarrow) or (\leftarrow)]

(assume all reactants & products in reaction are gases) (moves to fewer moles gas)

(g) higher V rxn goes [(\rightarrow) or (\leftarrow)]

(assume all reactant & products in reaction are gases) (moves to fewer moles gas)