## Topic 3.3. Chemical Equilibrium

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## Outline

- Introduction
- Main part
- What is equilibrium?
- Expressions for equilibrium constants, $\boldsymbol{K}_{\mathbf{c}}$;
- Calculating $K_{c}$ using equilibrium concentrations;

Calculating equilibrium concentrations using initial concentration and $\boldsymbol{K}_{\mathbf{c}}$ value;

- Relationship between $K_{\mathrm{c}}$ and $K_{\mathrm{p}}$;
- Factors that affect equilibrium;
- Le Chatelier's Principle
- Conclusion
- Literature




## Consider the following reactions:

$\mathrm{CaCO}_{3}(s)+\mathrm{CO}_{2}(a q)+\mathrm{H}_{2} \mathrm{O}(l) \rightarrow \mathrm{Ca}^{2+}(a q)+2 \mathrm{HCO}_{3}^{-}(a q)$
..(1)
and
$\mathrm{Ca}^{2+}(a q)+2 \mathrm{HCO}_{3}^{-}(a q) \rightarrow \mathrm{CaCO}_{3}(s)+\mathrm{CO}_{2}(a q)+\mathrm{H}_{2} \mathrm{O}(l)$
..(2)

Reaction (2) is the reverse of reaction (1).
At equilibrium the two opposing reactions occur at the same rate.

Concentrations of chemical species do not change once equilibrium is established.

Consider the following equilibrium system:

$$
\begin{aligned}
w \mathrm{~A}+x \mathrm{~B} & \rightleftarrows \mathrm{C}+z \mathrm{D} \\
K_{\mathrm{c}} & =\frac{[\mathrm{C}]^{y}[\mathrm{D}]^{2}}{[\mathrm{~A}]^{v}[\mathrm{~B}]^{x}}
\end{aligned}
$$

The numerical value of $K_{\mathrm{c}}$ is calculated using the concentrations of reactants and products that exist at equilibrium.

## Expressions for Equilibrium

Examples:

$$
\begin{aligned}
& \mathrm{N}_{2}(g)+3 \mathrm{H}_{2}(g) \nleftarrow 2 \mathrm{NH}_{3}(g) ; \quad K_{\mathrm{c}}=\frac{\left[\mathrm{NH}_{3}\right]^{2}}{\left[\mathrm{~N}_{2}\right]\left[\mathrm{H}_{2}\right]^{3}} \\
& \mathrm{PCl}_{5}(g) \nleftarrow \mathrm{PCl}_{3}(g)+\mathrm{Cl}_{2}(g) ; \quad K_{\mathrm{c}}=\frac{\left[\mathrm{PCl}_{3}\right]\left[\mathrm{Cl}_{2}\right]}{\left[\mathrm{PCl}_{5}\right]} \\
& \mathrm{CH}_{4}(\mathrm{~g})+\mathrm{H}_{2}(\mathrm{~g}) \nrightarrow \mathrm{CO}(\mathrm{~g})+3 \mathrm{H}_{2}(\mathrm{~g}) ; \\
& K_{\mathrm{c}}=\frac{[\mathrm{CO}]\left[\mathrm{H}_{2}\right]^{3}}{\left[\mathrm{CH}_{4}\right]\left[\mathrm{H}_{2} \mathrm{O}\right]}
\end{aligned}
$$

## 3Cabulating Equilibrium Constant

## Example-1:

1 mole of $\mathrm{H}_{2}$ gas and 1 mole of $\mathrm{I}_{2}$ vapor are introduced into a 5.00-liter sealed flask. The mixture is heated to a certain temperature and the following reaction occurs until equilibrium is established.

$$
\mathrm{H}_{2}(\mathrm{~g})+\mathrm{I}_{2}(\mathrm{~g}) \rightleftarrows 2 \mathrm{HI}(\mathrm{~g})
$$

At equilibrium, the mixture is found to contain 0,316 mole of HI . (a) What are the concentrations of $\mathrm{H}_{2}, \mathrm{I}_{2}$ and HI at equilibrium? (b) Calculate the equilibrium constant $K_{c}$.

$$
\mathrm{H}_{2}(\mathrm{~g})+\mathrm{I}_{2}(\mathrm{~g}) \underset{ }{\rightleftarrows} \mathrm{HI}(\mathrm{~g})
$$

Initial [ ], M:
0.200
0.200
0.000

Change in [ ], M: $-0.158 \quad-0.158+0.316$
Equilibrium [ ], M
0.042
0.042
0.316

$$
K_{\mathrm{c}}=\frac{[\mathrm{HI}]^{2}}{\left[\mathrm{H}_{2}\right]\left[\mathrm{I}_{2} \mathrm{~F}\right.} \quad \frac{(0.316)^{2}}{(0.0 \overline{\overline{4}} 2)^{5 / 2}}
$$

## Calculating Equilibrium Constâtit

## Example-2:

0.500 mole of HI is introduced into a 1.00 liter sealed flask and heated to a certain temperature. Under this condition HI decomposes to produce $\mathrm{H}_{2}$ and $\mathrm{I}_{2}$ until an equilibrium is established. An analysis of the equilibrium mixture shows that 0.105 mole of HI has decomposed. Calculate the equilibrium concentrations of $\mathrm{H}_{2}, \mathrm{I}_{2}$ and HI , and the equilibrium constant $K_{\mathrm{c}}$ for the following reaction:

$$
\mathrm{H}_{2}(\mathrm{~g})+\mathrm{I}_{2}(\mathrm{~g}) \rightleftarrows 2 \mathrm{HI}(\mathrm{~g}),
$$

## earleulating Equilibrium Constant

The reaction: $\mathrm{H}_{2}(\mathrm{~g})+\mathrm{I}_{2}(\mathrm{~g}) \nLeftarrow 2 \mathrm{HI}(\mathrm{g})$, proceeds from right to left.

$$
\mathrm{H}_{2}(\mathrm{~g})+\mathrm{I}_{2}(\mathrm{~g}) \rightleftarrows 2 \mathrm{HI}(\mathrm{~g})
$$

| Initial [ ], $M:$ | 0.000 | 0.000 | 0.500 |
| :--- | :--- | :--- | :---: |
| Change in [ ], $M:+0.0525$ | +0.0525 | -0.105 |  |
| Equil'm [ ], $M$ | 0.0525 | 0.0525 | 0.395 |

$$
K_{\mathrm{c}}=\frac{(0.395)^{2}}{(0.0525)^{2}}=56.6
$$

## Expression and value of Equilibrium Constant for a Reaction

* The expression for $K$ depends on the equation; The value of $K$ applies to that equation; it does not depend on how the reaction occurs;
Concentrations used to calculate the value of $K$ are those measured at equilibrium.

The expression of equilibrium constant depends on how the equilibrium equation is written. For example, for the following equilibrium:
$\mathrm{H}_{2}(g)+\mathrm{I}_{2}(g) \rightleftarrows 2 \mathrm{HI}(g) ; \quad K_{\mathrm{c}}=\frac{[\mathrm{HI}]^{2}}{\left[\mathrm{H}_{2}\right]\left[\mathrm{I}_{2}\right]}$
For the reverse reaction:
$2 \mathrm{HI}(g) \rightleftarrows \mathrm{H}_{2}(g)+\mathrm{I}_{2}(g) ; \quad K_{\mathrm{c}}^{\prime}=\frac{\left[\mathrm{H}_{2}\right]\left[\mathrm{I}_{2}\right]}{[\mathrm{HI}]^{2}}=1 / K_{\mathrm{c}}$
And for the reaction: $\mathrm{HI}(g) \rightleftarrows 1 / 2 \mathrm{H}_{2}(g)+1 / 2 \mathrm{I}_{2}(g)$;

$$
K_{\mathrm{c}}{ }^{\prime \prime}=\sqrt{\frac{\left[\mathrm{H}_{2}\right]\left[\mathrm{I}_{2}\right]}{[\mathrm{HI}]^{2}}}=\sqrt{K_{\mathrm{c}}{ }^{\prime}}=1 / \sqrt{K_{\mathrm{c}}}
$$

* Consider the following reaction involving gases:

$$
\begin{gathered}
2 \mathrm{SO}_{2}(g)+\mathrm{O}_{2}(g) \rightleftarrows 2 \mathrm{SO}_{3}(g) \\
K_{\mathrm{p}}=\frac{\left(\mathrm{P}_{\mathrm{SO} 3}\right)^{2}}{\left(\mathrm{P}_{\mathrm{SO} 2}\right)^{2}\left(\mathrm{P}_{\mathrm{O} 2}\right)}
\end{gathered}
$$

## The Relationship between $K_{c}$ and $K_{p}$

* Consider the reaction: $2 \mathrm{SO}_{2}(\mathrm{~g})+\mathrm{O}_{2}(\mathrm{~g}) \nLeftarrow 2 \mathrm{SO}_{3}(\mathrm{~g})$

$$
K_{\mathrm{c}}=\frac{[\mathrm{SO}]^{2}}{\left[\mathrm{SO}_{2}\right]^{2}\left[\mathrm{O}_{2}\right]} K_{\mathrm{p}}=\quad \frac{\left(\mathrm{P}_{\mathrm{SO} 3}\right)^{2}}{\left(\mathrm{P}_{\mathrm{SO} 2}\right)^{2}\left(\mathrm{P}_{\mathrm{O} 2}\right)}
$$

Assuming ideal behavior,
where $\mathrm{PV}=\mathrm{nRT}$ and $\mathrm{P}=(\mathrm{n} / \mathrm{V}) \mathrm{RT}=[\mathrm{M}] \mathrm{RT}$ and $\mathrm{P}_{\mathrm{SO} 3}=\left[\mathrm{SO}_{3}\right] \mathrm{RT} ; \mathrm{P}_{\mathrm{SO} 2}=\left[\mathrm{SO}_{2}\right] \mathrm{RT} ; \quad \mathrm{P}_{\mathrm{O} 2}=\left[\mathrm{O}_{2}\right] \mathrm{RT}$

$$
K_{\mathrm{p}}=\frac{\left[\mathrm{SO}_{3}\right]^{2}(\mathrm{RT})^{2}}{\left[\mathrm{SO}_{2}\right]^{2}(\mathrm{RT})^{2}\left[\mathrm{O}_{2}\right](\mathrm{RT})}=\frac{\left[\mathrm{SO}_{3}\right]^{2}}{\left[\mathrm{SO}_{2}\right]^{2}\left[\mathrm{O}_{2}\right]}(\mathrm{RT})^{-1}=K_{\mathrm{c}}(\mathrm{RT})^{-1}
$$

For reaction: $\mathrm{PCl}_{5}(\mathrm{~g}) \rightarrow \mathrm{PCl}_{3}(\mathrm{~g})+\mathrm{Cl}_{2}(g)$;

$$
\begin{aligned}
K_{\mathrm{p}} & =\frac{\left(\mathrm{P}_{\mathrm{PCl} 3}\right)\left(\mathrm{P}_{\mathrm{Cl} 2}\right)}{\left(\mathrm{P}_{\mathrm{PCl5}}\right)}=\frac{\left[\mathrm{PCl}_{3}\right](\mathrm{RT}) \times\left[\mathrm{Cl}_{2}\right](\mathrm{RT})}{\left[\mathrm{PCl}_{5}\right](\mathrm{RT})} \\
& =\frac{\left[\mathrm{PCl}_{3}\right]\left[\mathrm{Cl}_{2}\right]}{\left[\mathrm{PCl}_{5}\right]}(\mathrm{RT})^{1}=K_{C}(\mathrm{RT})^{1}
\end{aligned}
$$

In general, for reactions involving gases such that,

- $a \mathrm{~A}+b \mathrm{~B} \rightleftarrows c \mathrm{C}+d \mathrm{D}$ where $\mathrm{A}, \mathrm{B}, \mathrm{C}$, and D are all gases, and $a, b, c$, and $d$ are their respective coefficients,
- $K_{\mathrm{p}}=K_{\mathrm{c}}(\mathrm{RT})^{\Delta \mathrm{n}}$
and $\Delta \mathrm{n}=(c+d)-(a+b)$
(In heterogeneous systems, only the coefficients of the gaseous species are counted.)

For other reactions:

- 1. $2 \mathrm{NO}_{2}(\mathrm{~g}) \rightleftarrows \mathrm{N}_{2} \mathrm{O}_{4}(\mathrm{~g}) ; \quad K_{\mathrm{p}}=K_{\mathrm{c}}(\mathrm{RT})^{-1}$
- 2. $\mathrm{H}_{2}(g)+\mathrm{I}_{2}(g) \rightleftarrows 2 \mathrm{HI}(g) ; \quad K_{\mathrm{p}}=K_{\mathrm{c}}$
- 3. $\mathrm{N}_{2}(g)+3 \mathrm{H}_{2}(g) \rightleftarrows 2 \mathrm{NH}_{3}(g) ; \quad K_{\mathrm{p}}=K_{\mathrm{c}}(\mathrm{RT})^{-2}$


## Homogeneous equilibria:

$$
\begin{aligned}
& \mathrm{CH}_{4}(\mathrm{~g})+\mathrm{H}_{2} \mathrm{O}(\mathrm{~g}) \rightleftarrows \mathrm{CO}(\mathrm{~g})+3 \mathrm{H}_{2}(\mathrm{~g}) ; \\
& \mathrm{CO}(\mathrm{~g})+\mathrm{H}_{2} \mathrm{O}(\mathrm{~g}) \rightleftarrows \mathrm{CO}_{2}(\mathrm{~g})+\mathrm{H}_{2}(\mathrm{~g}) ;
\end{aligned}
$$

Heterogeneous equilibria:
$\mathrm{CaCO}_{3}(\mathrm{~s}) \rightleftarrows \mathrm{CaO}(\mathrm{s})+\mathrm{CO}_{2}(\mathrm{~g}) ;$
$\mathrm{HF}(a q)+\mathrm{H}_{2} \mathrm{O}(l) \rightleftarrows \mathrm{H}_{3} \mathrm{O}^{+}(a q)+\mathrm{F}^{-}(a q) ;$
$\mathrm{PbCl}_{2}(s) \not \rightleftarrows \mathrm{Pb}^{2+}(a q)+2 \mathrm{Cl}^{-}(a q) ;$

Examples:

$$
\begin{aligned}
& \mathrm{CaCO}_{3}(s) \notin \mathrm{CaO}(s)+\mathrm{CO}_{2}(g) ; \\
& K_{\mathrm{c}}=\left[\mathrm{CO}_{2}\right] \quad K_{\mathrm{p}}=\mathrm{P}_{\mathrm{CO} 2} ; \quad K_{\mathrm{p}}=K_{\mathrm{c}}(\mathrm{RT})
\end{aligned}
$$

$\mathrm{HF}(a q)+\mathrm{H}_{2} \mathrm{O}_{(l)} \rightleftarrows \mathrm{H}_{3} \mathrm{O}^{+}(a q)+\mathrm{F}^{-}(a q) ;$

$$
K_{\mathrm{a}}=\frac{\left[\mathrm{H}_{3} \mathrm{O}^{+}\right]\left[\mathrm{F}^{-}\right]}{[\mathrm{HF}]}
$$

## Solubility Eqilibrium

$\mathrm{PbCl}_{2}(s) \rightleftarrows \mathrm{Pb}^{2+}(a q)+2 \mathrm{Cl}^{-}(a q) ;$

$$
K_{s p}=\left[\mathrm{Pb}^{2+}\right]\left[\mathrm{Cl}^{-}\right]^{2}
$$

( $K_{s p}$ is called solubility product)

## Counining Equations and Equilibrium

when two or more equations are added to yield a net equation, the equilibrium constant for the net equation, $K_{n e t}$, is equal to the product of equilibrium constants of individual equations.

- For example,

$$
\operatorname{Eqn}(1): \mathrm{A}+\mathrm{B} \nLeftarrow \mathrm{C}+\mathrm{D} ; \quad K_{1}=\frac{[\mathrm{C}][\mathrm{D}]}{[\mathrm{A}][\mathrm{D}]}
$$

$$
\operatorname{Eqn}(2): \mathrm{C}+\mathrm{E} \nLeftarrow \mathrm{~B}+\mathrm{F} ; \quad K_{2}=\frac{[\mathrm{B}][\mathrm{F}]}{[\mathrm{C}][\mathrm{E}]}
$$

## Constants

* Net equation: $\mathrm{A}+\mathrm{E} \rightleftarrows \mathrm{D}+\mathrm{F}$;

$$
K_{\text {net }}^{=} K \frac{[\mathrm{D} \nmid[\mathrm{~F}]}{[\mathrm{A}][\mathrm{E}]}
$$

If $\operatorname{Eqn}(1)+\operatorname{Eqn}(2)=$ Net equation, then $K_{1} \times K_{2}=K_{\text {net }}$

## 触 Equilibrium Exercise \#1

A flask is charged with 2.00 atm of nitrogen dioxide and 1.00 atm of dinitrogen tetroxide at $25^{\circ} \mathrm{C}$ and allowed to reach equilibrium. When equilibrium is established, the partial pressure of $\mathrm{NO}_{2}$ has decreased by 1.24 atm . (a) What are the partial pressures of $\mathrm{NO}_{2}$ and $\mathrm{N}_{2} \mathrm{O}_{4}$ at equilibrium? (b) Calculate $K_{\mathrm{p}}$ and $K_{\mathrm{c}}$ for following reaction at $25^{\circ} \mathrm{C}$.

$$
2 \mathrm{NO}_{2}(\mathrm{~g}) \rightleftarrows \mathrm{N}_{2} \mathrm{O}_{4}(\mathrm{~g})
$$

(Answer: $K_{\mathrm{p}}=2.80 ; K_{\mathrm{c}}=68.6$ )

## Equilibrium Exercise \#2a

Methanol is produced according to the following equation:

$$
\mathrm{CO}(\mathrm{~g})+2 \mathrm{H}_{2}(\mathrm{~g}) \rightleftarrows \mathrm{CH}_{3} \mathrm{OH}(\mathrm{~g})
$$

In an experiment, 1.000 mol each of CO and $\mathrm{H}_{2}$ were allowed to react in a sealed $10.0-\mathrm{L}$ reaction vessel at 500 K . When the equilibrium was established, the mixture was found to contain 0.0892 mole of $\mathrm{CH}_{3} \mathrm{OH}$. What are the equilibrium concentrations of $\mathrm{CO}, \mathrm{H}_{2}$ and $\mathrm{CH}_{3} \mathrm{OH}$ ? Calculate the equilibrium constants $K_{\mathrm{c}}^{2}$ and $K_{\mathrm{p}}$ for this reaction at 500 K ? ( $\mathrm{R}=0.0821 \mathrm{~L} . \mathrm{atm} / \mathrm{Mol} . \mathrm{K}$ )
(Answer: $[\mathrm{CO}]=0.0911 \mathrm{M} ;\left[\mathrm{H}_{2}\right]=0.0822 \mathrm{M} ;\left[\mathrm{CH}_{3} \mathrm{OH}\right]=0.00892 \mathrm{M}$; (b) $K_{\mathrm{c}}=14.5 ; K_{\mathrm{p}}=8.60 \times 10^{-3}$ )

## Equilibrium Exercise \#2a

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(Answer: $[\mathrm{CO}]=0.0911 \mathrm{M} ;\left[\mathrm{H}_{2}\right]=0.0822 \mathrm{M} ;\left[\mathrm{CH}_{3} \mathrm{OH}\right]=0.00892 \mathrm{M}$; (b) $K_{\mathrm{c}}=14.5 ; K_{\mathrm{p}}=8.60 \times 10^{-3}$ )

For any system or reaction:

1. Knowing the equilibrium constant, we can predict whether or not a reaction mixture is at equilibrium, and we can predict the direction of net reaction.

- $Q_{\mathrm{c}}=K_{\mathrm{c}} \square$ equilibrium (no net reaction)
- $Q_{\mathrm{c}}<K_{\mathrm{c}} \square$ a net forward reaction;
- $Q_{\mathrm{c}}>K_{\mathrm{c}} \square$ a net reverse reaction

2. The value of $K$ tells us whether a reaction favors the products or the reactants.

For a reaction of known $K_{\mathrm{c}}$ value, the direction of net reaction can be predicted by calculating the reaction quotient, $Q_{c}$.
$Q_{\mathrm{c}}$ is called the reaction quotient, where for a reaction such as:

$$
a \mathrm{~A}+b \mathrm{~B} \nrightarrow c \mathrm{C}+d \mathrm{D} ; \quad Q_{\mathrm{c}}=\frac{[\mathrm{C}]^{\mathrm{c}}[\mathrm{D}]^{\mathrm{d}}}{[\mathrm{~A}]^{\mathrm{a}}[\mathrm{~B}]^{\mathrm{b}}}
$$

$Q_{\mathrm{c}}$ has the same expression as $K_{\mathrm{c}}$, but
$Q_{\mathrm{c}}$ is calculated using concentrations that are not necessarily at equilibrium.

## What does the reaction quotient tell us?

If $Q_{\mathrm{c}}=K_{\mathrm{c}}, \square$ the reaction is at equilibrium;
If $Q_{\mathrm{c}}<K_{\mathrm{c}}, \square$ the reaction is not at equilibrium and there's a net forward reaction;
If $Q_{\mathrm{c}}>K_{\mathrm{c}}$, the reaction is not at equilibrium and there's a net reaction in the opposite direction.

Equation: $\quad \mathrm{H}_{2}(\mathrm{~g})+\mathrm{I}_{2}(\mathrm{~g}) \not \rightleftarrows 2 \mathrm{HI}(\mathrm{g})$,

| Initial [ ], M | 0.1000 | 0.1000 | 0.0000 |
| :--- | :--- | :--- | :--- |
| Change [ ], M | $-x$ | $-x$ | $+2 x$ |

Equilibrium [ ], $M(0.1000-x)(0.1000-x) \quad 2 x$

$$
K_{\mathrm{c}}=\frac{[\mathrm{HI}]^{2}}{\left[\mathrm{H}_{2}\right]\left[\mathrm{I}_{2}\right]}=\frac{(2 x)^{2}}{(0.100-x)^{2}}=55.6
$$

$$
\begin{aligned}
& \frac{2 x}{(0.100-x}=\sqrt{55.6}=7.46 \\
& 2 x=0.746-7.46 x ; \Rightarrow 9.46 x=0.746 \\
& x=0.0789 ; \\
& {\left[\mathrm{H}_{2}\right]=\left[\mathrm{I}_{2}\right]=0.0211 \mathrm{M} ;[\mathrm{HI}]=0.158 \mathrm{M}}
\end{aligned}
$$

## For the reaction:

$2 \mathrm{NO}_{2}(\mathrm{~g}) \nless \mathrm{N}_{2} \mathrm{O}_{4}(\mathrm{~g}) ; \quad K_{\mathrm{p}}=1.27$ at 353 K.
If the initial pressure of $\mathrm{NO}_{2}$ was 3.92 atm , and initially there was no $\mathrm{N}_{2} \mathrm{O}_{4}$, what are the partial pressures of the gases at equilibrium at 353 K ? What is the total gas pressure at equilibrium?
(Answer: $\mathrm{P}_{\mathrm{NO} 2}=1.06 \mathrm{~atm} ; \mathrm{P}_{\mathrm{N} 2 \mathrm{O} 4}=1.43 \mathrm{~atm} ; \mathrm{P}_{\text {total }}=2.49 \mathrm{~atm}$ )

## Le Châtelier's Principle

The Le Châtelier's principle states that: when factors that influence an equilibrium are altered, the equilibrium will shift to a new position that tends to minimize those changes.

Factors that influence equilibrium:
Concentration, temperature, and partial pressure (for gaseous)

Consider the reaction: $\mathrm{N}_{2}(\mathrm{~g})+3 \mathrm{H}_{2}(\mathrm{~g}) \nLeftarrow 2 \mathrm{NH}_{3}(\mathrm{~g}) ;$

$$
K_{\mathrm{c}}=\frac{[\mathrm{NH}]^{2}}{\left[\mathrm{~N}_{2}\right]\left[\mathrm{H}_{2}\right]^{3}}
$$

* If $\left[\mathrm{N}_{2}\right]$ and/or $\left[\mathrm{H}_{2}\right]$ is increased, $Q_{\mathrm{c}}<K_{\mathrm{c}}$
- $\square$ a net forward reaction will occur to reach new equilibrium position.
* If $\left[\mathrm{NH}_{3}\right]$ is increased, $Q_{\mathrm{c}}>K_{\mathrm{c}}$, and a net reverse reaction will occur to come to new equilibrium position. and shift left when nressure decreases

Consider the reaction:
$2 \mathrm{SO}_{2}(\mathrm{~g})+\mathrm{O}_{2}(\mathrm{~g}) \nless 2 \mathrm{SO}_{3}(\mathrm{~g})$,

1. The total moles of gas decreases as reaction proceeds in the forward direction.
2. If pressure is increased by decreasing the volume (compression), a forward reaction occurs to reduce the stress.
3. Reactions that result in fewer moles of gas favor high pressure conditions.

Consider the reaction: $\mathrm{PCl}_{5}(\mathrm{~g}) \nleftarrow \mathrm{PCl}_{3}(\mathrm{~g})+\mathrm{Cl}_{2}(\mathrm{~g})$;

1. Forward reaction results in more gas molecules.
2. Pressure increases as reaction proceeds towards equilibrium.
3. If mixture is compressed, pressure increases, and reverse reaction occurs to reduce pressure;
4. If volume expands and pressure drops, forward reaction occurs to compensate.
5. This type of reactions favors low pressure condition

Consider the following reactions:

$$
\begin{aligned}
& \text { 1. } \mathrm{CO}(\mathrm{~g})+\mathrm{H}_{2} \mathrm{O}(\mathrm{~g}) \not \longrightarrow \mathrm{CO}_{2}(\mathrm{~g})+\mathrm{H}_{2}(\mathrm{~g}) ; \\
& \text { 2. } \mathrm{H}_{2}(\mathrm{~g})+\mathrm{Cl}_{2}(\mathrm{~g}) \rightleftarrows 2 \mathrm{HCl}_{(\mathrm{g}) ;}
\end{aligned}
$$

2. Reactions have same number of gas molecules in reactants and products.
3. Reducing or increasing the volume will cause equal effect on both sides - no net reaction will occur.
4. Equilibrium is not affected by change in pressure.

## 317) Effect Temperature on Equilibrium

- Consider the following exothermic reaction:

$$
\mathrm{N}_{2}(\mathrm{~g})+3 \mathrm{H}_{2}(\mathrm{~g}) \nLeftarrow 2 \mathrm{NH}_{3}(\mathrm{~g}) ; \quad \Delta H^{0}=-92 \mathrm{~kJ},
$$

The forward reaction produces heat $=>$ heat is a product. When heat is added to increase temperature, reverse reaction will take place to absorb the heat;
If heat is removed to reduce temperature, a net forward reaction will occur to produce heat.

* Exothermic reactions favor low temperature conditions.


## 解 Equilibrium Exercise \#8

Determine whether the following reactions favor high or low pressures?

1. $2 \mathrm{SO}_{2}(\mathrm{~g})+\mathrm{O}_{2}(\mathrm{~g}) \nleftarrow 2 \mathrm{SO}_{3}(\mathrm{~g})$;
2. $\mathrm{PCl}_{5}(\mathrm{~g}) \not \rightleftarrows \mathrm{PCl}_{3}(\mathrm{~g})+\mathrm{Cl}_{2}(\mathrm{~g})$;
3. $\mathrm{CO}(\mathrm{g})+2 \mathrm{H}_{2}(\mathrm{~g}) \not \rightleftarrows \mathrm{CH}_{3} \mathrm{OH}(\mathrm{g})$;
4. $\mathrm{N}_{2} \mathrm{O}_{4}(\mathrm{~g}) \nLeftarrow 2 \mathrm{NO}_{2}(\mathrm{~g})$;
5. $\mathrm{H}_{2}(\mathrm{~g})+\mathrm{F}_{2}(\mathrm{~g}) \rightleftarrows 2 \mathrm{HF}(\mathrm{g}) ;$

## Equilibrium Exercise \#9

Determine whether the following reactions favors high or low temperature?

1. $2 \mathrm{SO}_{2}(\mathrm{~g})+\mathrm{O}_{2}(\mathrm{~g}) \not \rightleftarrows 2 \mathrm{SO}_{3}(\mathrm{~g}) ; \quad \Delta H^{\mathrm{o}}=-180 \mathrm{~kJ}$
2. $\mathrm{CO}(\mathrm{g})+\mathrm{H}_{2} \mathrm{O}(\mathrm{g}) \nless \mathrm{CO}_{2}(\mathrm{~g})+\mathrm{H}_{2}(\mathrm{~g}) ; \quad \Delta H^{\mathrm{o}}=-46 \mathrm{~kJ}$
3. $\quad \mathrm{CO}(\mathrm{g})+\mathrm{Cl}_{2}(\mathrm{~g}) \nVdash \mathrm{COCl}_{2}(\mathrm{~g}) ; \quad \Delta H^{\circ}=-108 \mathrm{~kJ}$
4. $\mathrm{N}_{2} \mathrm{O}_{4}(\mathrm{~g}) \nLeftarrow 2 \mathrm{NO}_{2}(\mathrm{~g}) ; \quad \Delta H^{\circ}=+57 \mathrm{~kJ}$
5. $\quad \mathrm{CO}(\mathrm{g})+2 \mathrm{H}_{2}(\mathrm{~g}) \nrightarrow \mathrm{CH}_{3} \mathrm{OH}(\mathrm{g}) ; \quad \Delta H^{0}=-270 \mathrm{~kJ}$

## Teminical Equilibria in Industrial Processes

Production of Sulfuric Acid, $\mathrm{H}_{2} \mathrm{SO}_{4}$;

1. $\mathrm{S}_{8}(\mathrm{~s})+8 \mathrm{O}_{2}(\mathrm{~g}) \rightarrow 8 \mathrm{SO}_{2}(\mathrm{~g})$
2. $\quad 2 \mathrm{SO}_{2}(\mathrm{~g})+\mathrm{O}_{2}(\mathrm{~g}) \rightleftarrows 2 \mathrm{SO}_{3}(\mathrm{~g}) ; \quad \Delta H=-198 \mathrm{~kJ}$
3. $\mathrm{SO}_{3}(\mathrm{~g})+\mathrm{H}_{2} \mathrm{SO}_{4}(l) \rightarrow \mathrm{H}_{2} \mathrm{~S}_{2} \mathrm{O}_{7}(l)$
4. $\mathrm{H}_{2} \mathrm{~S}_{2} \mathrm{O}_{7}^{(l)}+\mathrm{H}_{2} \mathrm{O}(t) \rightarrow 2 \mathrm{H}_{2} \mathrm{SO}_{4}^{(l)}$

The second reaction is exothermic and has high activation energy;
though thermodynamically favored the reaction is very slow at low temperature,.
At high temperature reaction goes faster, but the yield would be very low.
An optimum condition is achieved at moderate temperatures and using catalysts to speed up the reaction. Reaction also favors high pressure.

## Chemical Equilibria in Industrial Processes

The production of ammonia by the Haber-Bosch process:

$$
\mathrm{N}_{2}(\mathrm{~g})+3 \mathrm{H}_{2}(\mathrm{~g}) \rightleftarrows 2 \mathrm{NH}_{3}(\mathrm{~g}) ; \Delta H=-92 \mathrm{~kJ}
$$

* This reaction is exothermic and very slow at low temperature. Increasing the temperature will increase reaction rate, but will lower the yield.
An optimum condition is achieved at moderate temperature of 250 to $300^{\circ} \mathrm{C}$ with catalyst added to increase the reaction rate. Increasing the pressure will favor product formation.
* Reaction favors low temperature and high pressure conditions.


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$$

* This reaction is exothermic and very slow at low temperature. Increasing the temperature will increase reaction rate, but will lower the yield.
An optimum condition is achieved at moderate temperature of 250 to $300^{\circ} \mathrm{C}$ with catalyst added to increase the reaction rate. Increasing the pressure will favor product formation.
* Reaction favors low temperature and high pressure conditions.


## Chemical Equilibria in Industrial Processes

The production of ammonia by the Haber-Bosch process:

$$
\mathrm{N}_{2}(\mathrm{~g})+3 \mathrm{H}_{2}(\mathrm{~g}) \rightleftarrows 2 \mathrm{NH}_{3}(\mathrm{~g}) ; \Delta H=-92 \mathrm{~kJ}
$$

* This reaction is exothermic and very slow at low temperature. Increasing the temperature will increase reaction rate, but will lower the yield.
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* Reaction favors low temperature and high pressure conditions.


## Questions for self control-in

1.The expression of the equilibrium constant for following reaction :
A)

$$
\mathrm{wA}+\mathrm{xB} \nLeftarrow \mathrm{yC}+\mathrm{zD}
$$

B) $\mathrm{K}=\quad \frac{[\mathrm{A}]^{x}[\mathrm{~B}]^{x}}{[\mathrm{C}]^{x}[\mathrm{D}]^{2}}$
C) $\mathrm{K}=\quad \frac{[\mathrm{yCC}[\mathrm{zD}]}{[\mathrm{wA}][\mathrm{xB}]}$

## Questions for self control

2) 1.000 mole of $\mathrm{H}_{2}$ gas and 1.000 mole of $\mathrm{I}_{2}$ vapor are introduced into a 5.00 -liter sealed flask. The mixture is heated to a certain temperature and the following reaction occurs until equilibrium is established.

$$
\mathrm{H}_{2}(\mathrm{~g})+\mathrm{I}_{2}(\mathrm{~g}) \rightleftarrows 2 \mathrm{HI}(\mathrm{~g})
$$

At equilibrium, the mixture is found to contain 0,316 mole of HI. What are the concentration of $\mathrm{H}_{2}$, at equilibrium?
A) $0 \mathrm{~mol} / \mathrm{l}$
B) $0,042 \mathrm{~mol} / 1$
C) $1 \mathrm{~mol} / \mathrm{l}$
D) $5 \mathrm{~mol} / 1$
3.On the base Le-Shatelie principle predict the reaction direction of following reaction :

$$
2 \mathrm{SO}_{2}(\mathrm{~g})+\mathrm{O}_{2}(\mathrm{~g}) \rightleftarrows 2 \mathrm{SO}_{3}(\mathrm{~g}) ; \quad \Delta \mathrm{H}^{\mathrm{o}}=-180 \mathrm{~kJ}
$$

A) When pressure is increased
B) When pressure is decreased
C) When temperature is increased
D) When temperature is decreased
E) When concentration of SO3 is increased
F) When concentration of SO 3 is decreased
J) When concentration of O 2 is increased
H) When concentration of O 2 is decreased

## Literature

## 1.Basic literature :

1. Jenkins, Chemistry, ISBN 978-0-17-628930-0
2. Alberta Learning, Chemistry data booklet 2010, product №755115, ISBN 10645246
3.М.К.Оспанова, К.С.Аухадиева, Т.Г. Белоусова Химия: Учебник 1,2 часть для 10 класса естественно-математического направления общеобразовательных школ Алматы: Мектеп, 2019г.
4.М.К.Оспанова, К.С.Аухадиева, Т.Г. Белоусова Химия: Учебник 1,2 часть для 11 класса естественно-математического направления общеобразовательных школ Алматы: Мектеп, 2020 г.
3. М.Оспанова, К.Аухадиева, Т.Белоусова Химия. Дәрислик. 1, 2-кисим Алматы: Мектеп, 2019
4. М.Успанова, К.Аухадиева, Т. Белоусова Химия. Дарслик. 1, 2 - қисм Алматы: Мектеп, 2019
5. Т.Г.Белоусова, К.С. Аухадиева Химия: Методическое руководство 1,2 часть естественноматематического направления общеобразовательных школ Алматы: Мектеп, 2019 г.
6. Темирбулатова А., Сагимбекова Н., Алимжанова С.,Химия. Сборник задач и упражнений Алматы: Мектеп, 2019 г.

## 2.Additional literature :

1.Б.А.Мансуров «Химия» 10-11 кл., Атамура 2015 г
2.Б.Мансуров., Н.Торшина «Методика преподавания органической химии» Атамура 2015г.
3.А.Е.Темирбулатова, Н.Н.Нурахметов, Р.Н.Жумадилова, С.К.Алимжанова

Химия: Учебник для 11 класса естественно-математического направления общеобразовательной школы Алматы: Мектеп, 2015г. -344 стр.
4.Г.Джексембина «Методическое руководство» Алматы: Мектеп, 2015г
5.А.Темирболатова., А.Казымова., Ж.Сагымбекова «Книга для чтения» Мектеп 2015г.
6. Торгаева Э., Шуленбаева Ж. и др Химия.Электронный учебник.10класс. 2016 Национальный центр информатизации
7. Жакирова Н., Жандосова И. и др Химия.Электронный учебник.11класс. 2016 Национальный центр информатизации
8.Эектронные ресурсы с www.bilimland.kz

## Do you have any questions?



