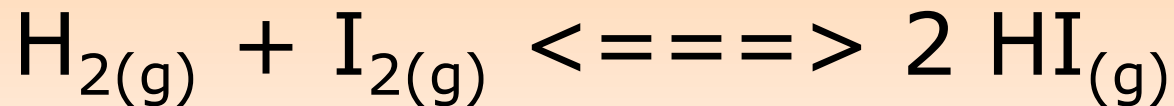


EQUILIBRIUM LAW - K_{eq}

EQUILIBRIUM LAW - K_{eq}

A simple mathematical relationship defines each reaction at chemical equilibrium.



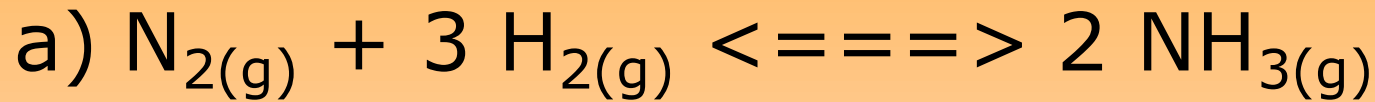
$$K_{eq} = \frac{[\text{HI}_{(g)}]^2}{[\text{H}_{2(g)}][\text{I}_{2(g)}]}$$

EQUILIBRIUM LAW - K_{eq}

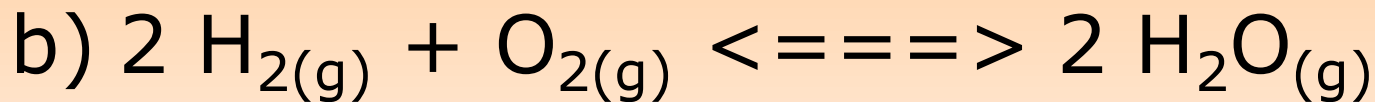
- can always be developed from the balanced chemical equation
- K_{eq} will always have a specific value at specific environmental conditions
 - if the conditions change, the K_{eq} will also change
- units for K_{eq} will never be used

EQUILIBRIUM LAW - K_{eq}

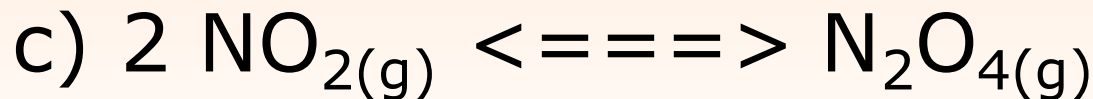
Example #1



$$K_{eq} = \frac{[\text{NH}_{3(g)}]^2}{[\text{N}_{2(g)}][\text{H}_{2(g)}]^3}$$



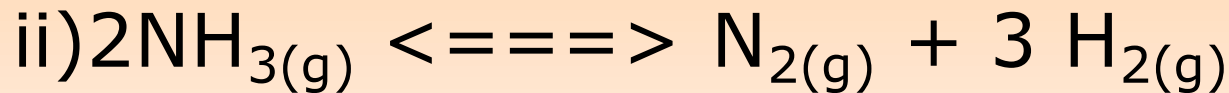
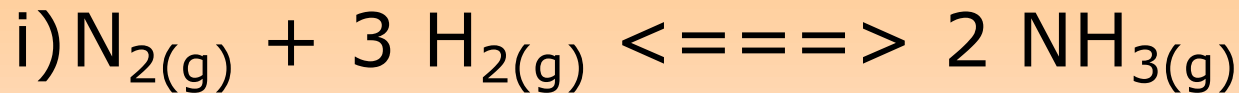
$$K_{eq} = \frac{[\text{H}_2\text{O}_{(g)}]^2}{[\text{H}_{2(g)}]^2 [\text{O}_{2(g)}]}$$



$$K_{eq} = \frac{[\text{N}_2\text{O}_{4(g)}]}{[\text{NO}_{2(g)}]^2}$$

EQUILIBRIUM LAW - K_{eq}

How are the equilibrium laws of the following equations related?



$$\text{i) } K_{eq} = \frac{[\text{NH}_{3(g)}]^2}{[\text{N}_{2(g)}][\text{H}_{2(g)}]^3}$$

$$\text{ii) } K_{eq} = \frac{[\text{N}_{2(g)}][\text{H}_{2(g)}]^3}{[\text{NH}_{3(g)}]^2}$$

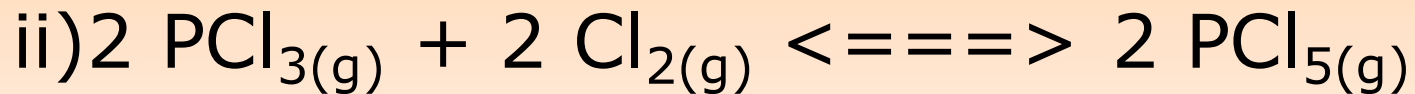
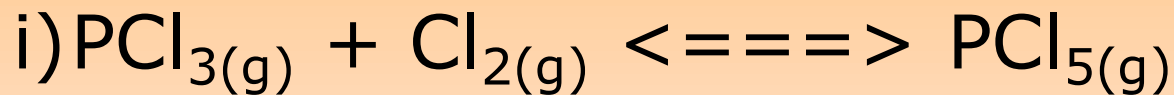
EQUILIBRIUM LAW - K_{eq}

If the direction of an equation is reversed, the equilibrium constant is the reciprocal of the original equilibrium constant.

$$K_{eq \text{ forward}} = 1 / K_{eq \text{ reverse}}$$

EQUILIBRIUM LAW - K_{eq}

How are the equilibrium laws of the following equations related?



$$\text{i) } K_{eq} = \frac{[\text{PCl}_{5(g)}]}{[\text{PCl}_{3(g)}][\text{Cl}_{2(g)}]}$$

$$\text{ii) } K_{eq} = \frac{[\text{PCl}_{5(g)}]^2}{[\text{PCl}_{3(g)}]^2[\text{Cl}_{2(g)}]^2}$$

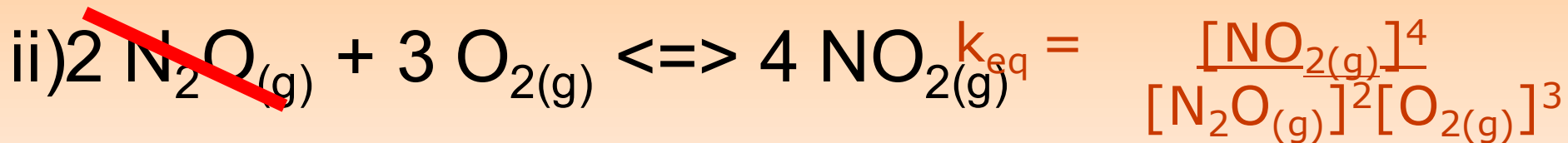
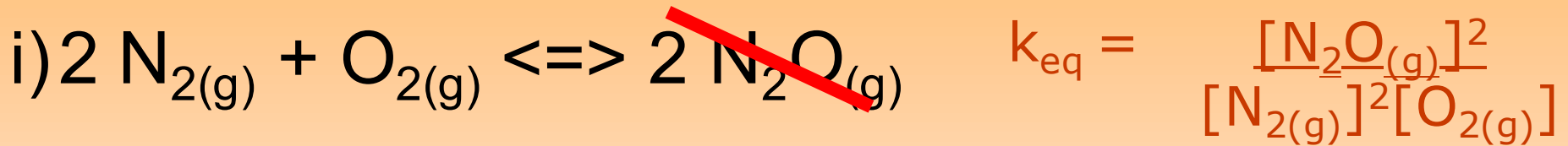
EQUILIBRIUM LAW - K_{eq}

When the coefficients of a balanced equation are multiplied by a factor, the equilibrium constant is raised to the exponent of the same factor.

$$K_{eq \text{ rxn ii)}} = [K_{eq \text{ rxn i)}}]^x$$

EQUILIBRIUM LAW - K_{eq}

What is the equilibrium law of the sum of the following reactions?



$$k_{eq} = \frac{[\text{NO}_{2(g)}]^4}{[\text{N}_{2(g)}]^2[\text{O}_{2(g)}]^4}$$

$$K_{eq \text{ final}} = \frac{[\text{N}_2\text{O}_{(g)}]^2}{[\text{N}_{2(g)}]^2[\text{O}_{2(g)}]} \times \frac{[\text{NO}_{2(g)}]^4}{[\text{N}_2\text{O}_{(g)}]^2[\text{O}_{2(g)}]^3}$$

EQUILIBRIUM LAW - K_{eq}

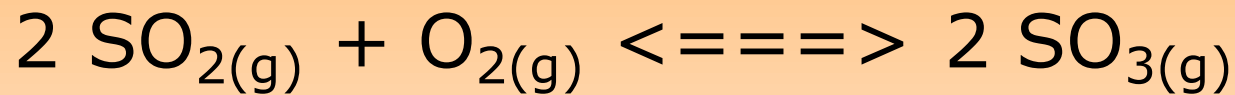
When chemical equilibria are added together, the equilibrium constants are multiplied together.

$$K_{eq \text{ final rxn}} = K_{eq \text{ rxn 1}} \times K_{eq \text{ rxn 2}}$$

EQUILIBRIUM LAW - K_{eq}

Example #2

At 25°C, $K_{eq} = 7.0 \times 10^{25}$ for:



What is the value of K_{eq} for:



$$\begin{aligned} K_{eq} &= 7.0 \times 10^{25} \text{ inversed, to the power of } 0.5 \\ &= 1.195 \times 10^{-13} \\ &= 1.2 \times 10^{-13} \end{aligned}$$

EQUILIBRIUM LAW - K_{eq}

MAGNITUDE OF K_{eq}

The value of K_{eq} (large or small) can provide a hint to the ratio of reactants to products at equilibrium.

EQUILIBRIUM LAW - K_{eq}

MAGNITUDE OF K_{eq}

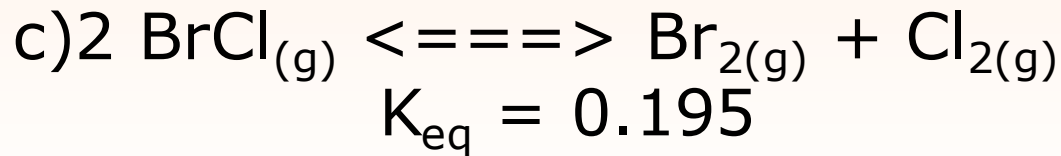
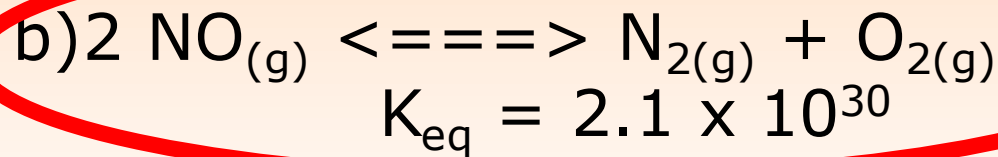
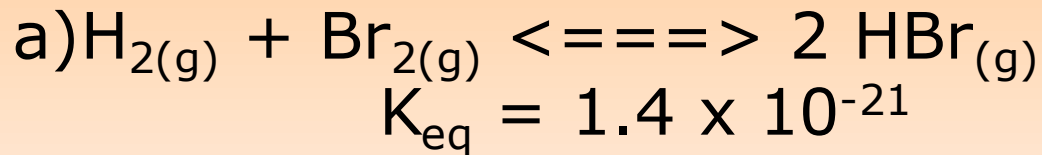
1. K_{eq} is very large ($K_{eq} > 1$)
 - $[products] > [reactants]$
2. $K_{eq} \approx 1$
 - $[products] \approx [reactants]$
3. K_{eq} is very small ($K_{eq} < 1$)
 - $[products] < [reactants]$

EQUILIBRIUM LAW - K_{eq}

MAGNITUDE OF K_{eq}

Example #3

Which of the following reactions will tend to proceed farthest toward completion?



EQUILIBRIUM LAW - K_{eq}

**EQUILIBRIUM INVOLVING
PURE SOLIDS AND
LIQUIDS**

EQUILIBRIUM LAW - K_{eq}

EQUILIBRIUM – SOLIDS AND LIQUIDS

What is the concentration of a solid or liquid? (i.e. H_2O)

Does the concentration of these pure compounds change?

ex. 1 mol $NaHCO_3$ occupies 38.9 cm^3

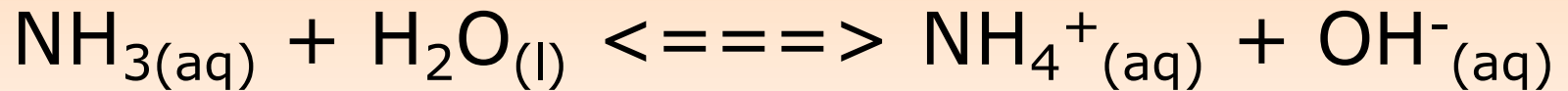
2 mol $NaHCO_3$ occupies 77.8 cm^3

Molar concentration remains the same. Solids and liquids are unaffected by concentration

EQUILIBRIUM LAW - K_{eq}

EQUILIBRIUM – SOLIDS AND LIQUIDS

In the equilibrium law, solids and liquids do not need to be included as it becomes part of the equilibrium constant.



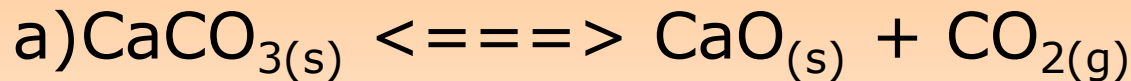
$$K_{eq} = \frac{[\text{NH}_4^+_{(aq)}][\text{OH}^-_{(aq)}]}{[\text{NH}_{3(aq)}]}$$

EQUILIBRIUM LAW - K_{eq}

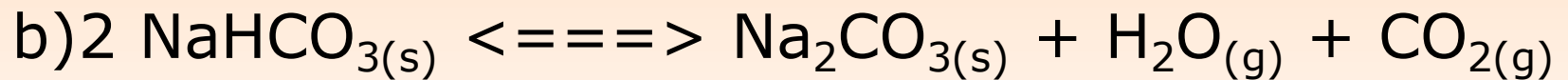
EQUILIBRIUM – SOLIDS AND LIQUIDS

Example #4

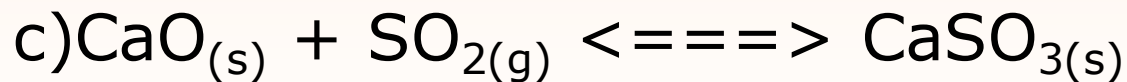
Write the equilibrium law for the following reactions:



$$K_{eq} = [\text{CO}_{2(g)}]$$



$$K_{eq} = [\text{H}_2\text{O}_{(g)}][\text{CO}_{2(g)}]$$

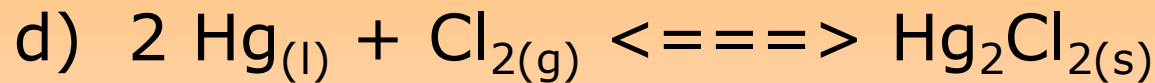


$$K_{eq} = \frac{1}{[\text{SO}_{2(g)}]}$$

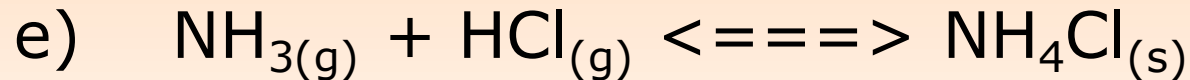
EQUILIBRIUM LAW - K_{eq}

EQUILIBRIUM – SOLIDS AND LIQUIDS

Example #4



$$K_{eq} = \frac{1}{[\text{Cl}_{2(g)}]}$$



$$K_{eq} = \frac{1}{[\text{NH}_{3(g)}][\text{HCl}_{(g)}]}$$

EQUILIBRIUM LAW - K_{eq}

EFFECT OF TEMPERATURE ON K_{eq}

EQUILIBRIUM LAW - K_{eq}

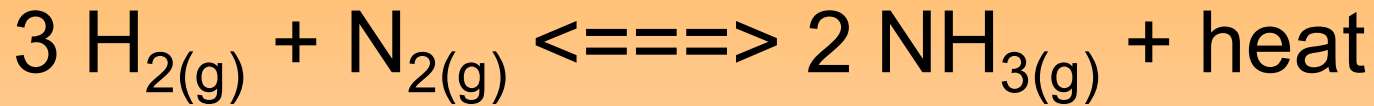
EQUILIBRIUM – TEMPERATURE

Will temperature change the value of K_{eq} ? Why or why not?

Reactions are endothermic or exothermic and therefore will be affected by the addition or removal of heat.

EQUILIBRIUM LAW - K_{eq}

EQUILIBRIUM – TEMPERATURE



a) What is the equilibrium law? $K_{eq} = \frac{[\text{NH}_{3(g)}]^2}{[\text{H}_{2(g)}]^3 [\text{N}_{2(g)}]}$

b) Which way does equilibrium shift when temperature increases? How will K_{eq} change?

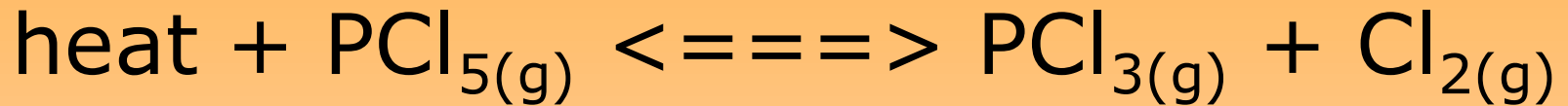
Shifts to the left. Since the denominator increases, K_{eq} decreases (is smaller in value)

c) When temperature decreases?

Shifts to the right. Since the numerator increases, K_{eq} increases (is larger in value)

EQUILIBRIUM LAW - K_{eq}

EQUILIBRIUM - TEMPERATURE



a) What is the equilibrium law? $K_{eq} = \frac{[\text{PCl}_{3(g)}][\text{Cl}_{2(g)}]}{[\text{PCl}_{5(g)}]}$

b) Which way does equilibrium shift when temperature increases? How will K_{eq} change?

Shifts to the right. Since the numerator increases, K_{eq} increases (is larger in value)

c) When temperature decreases?

Shifts to the left. Since the denominator increases, K_{eq} decreases (is smaller in value)

EQUILIBRIUM LAW - K_{eq}

Textbook homework:

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