

Equilibrium Constant (A quantitative expression of Equilibrium)

- Equilibrium constant is a mathematical expression that describes the relationship between reactants and products - The equilibrium constant is only applicable at equilibrium.
- The equilibrium constant, K_{eq} , is defined as a relationship between the concentration of the products divided by the concentration of the reactants.
 - If we know the balanced equation, we can properly write out the equilibrium equation.
 - The coefficients that occurs in a balanced equation becomes the exponent to which a given concentration is raised to.
- For the general equation $aA + bB \rightleftharpoons cC + dD$
 - the $K_{eq} = \frac{[C]^c [D]^d}{[A]^a [B]^b}$
 - Example:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad K_{eq} = \frac{[NH_3]^2}{[N_2][H_2]^3}$$
- K_{eq} depends upon the [] of the reactants and products, meaning that we consider ONLY gaseous and aqueous substances.
 - The concentration of pure liquids and solids remains constant.
 - Example:

$$PBr_5(s) \rightleftharpoons PBr_3(g) + Br_2(g) \quad K_{eq} = \frac{[PBr_3][Br_2]}{[PBr_5]}$$

* But since PBr_5 is a solid, its concentration is constant – it is not expressed in the equation.

So this equation becomes:

$$K_{eq} = [PBr_3][Br_2]$$

* NOTE: K_{eq} is only applicable at a given temperature.
- To calculate K_{eq} you must:
 - Write the balanced equation if it is not provided.
 - Write the equilibrium constant expression.
 - Write the molar [], or calculate the molar [] of all the known species.
 - Substitute the molar [] into the K_{eq} expression and solve.

Numerical value of K_{eq}

- The K_{eq} has a specific value for each reaction and must be determined experimentally.
- If the K_{eq} is close to 1, the reactants and products are about equally favoured.
- If the K_{eq} is large (several positive powers of 10) the products are favoured.
- If the K_{eq} is small (several negative powers of 10) the reactants are favoured.

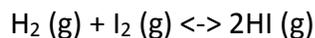
Effect of Temperature of K_{eq}

- Remember that all chemical reactions either release or absorb energy.
- ΔH is not part of the K_{eq} expression, but ΔH can be used to predict how the K_{eq} will react to temperature changes.
- Example: $H_2(g) + I_2(g) \rightleftharpoons 2HI(g) + 9.4 \text{ kJ}$
 - By decreasing the temperature, the reaction will shift towards the exothermic reaction, meaning that it will shift to the right. (Increases the products.)

- This means that Keq will increase, because the formula for Keq is [products] / [reactants]. If the [products] increases, so does the Keq.)
 - The reverse is also true. If we increase the temp, the reaction will shift towards the left (towards the endothermic reaction).
 - This means that Keq will decrease.
- For exothermic reactions, increased temp causes a shift to the left, which increases reactants and decreases Keq (and vice versa).
- For endothermic reactions, increased temp causes a shift to the right, which increases the products and increases Keq (and vice versa).

Example involving Keq

- A mixture of H₂ and I₂ react at 448 °C. When equilibrium is established, the [] of the participants is found to be: [H₂]=0.46 M, [I₂] = 0.39 M, and [HI] = 3.0 M. calculate the Keq at 448 °C from these data.



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

$$\text{Keq} = \frac{[3.0]^2}{[0.46] [0.39]}$$

$$\text{Keq} = 50.2$$