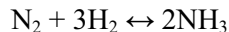


II Increasing the Pressure on the System



Increasing the pressure favours reaction with fewer molecules

Decreasing the pressure favours reaction with more molecules

If there is the same number of molecules on both sides of the equations then no change occurs; pressure has no effect.

III Effect of Temperature

Increase in temperature will favour the equilibrium in the endothermic reaction (so absorbing heat).

Decrease in temperature will shift the position of equilibrium in the exothermic direction (so releasing heat and increasing the temperature).



Note: in industrial plants, often moderate temperatures are used in reactions where lower temperatures favour the forward reaction. Since the rate of the reaction in these is generally much lower, thus to compensate the lower rate, a compromise moderate temperature is used to facilitate a faster rate but a lower yield of products.

IV Changes in Volume ($P \propto \frac{1}{V}$)

An increase in volume (i.e. the system wants to decrease the volume), i.e. you have effectively decreased the pressure, hence you must increase the pressure i.e. must move to side with more molecules, (thus more pressure and less volume).

NOTE: Volume changes have No Effect on a reaction with the same number of molecules on each side.

Also a change in volume changes the concentration, and hence the rate of the reaction.

V Adding a Noble Gas

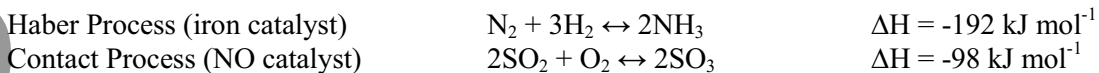
No effect

VI Catalyst

Has no effect on both the forward and reverse reaction.

It changes the time taken for the system to attain equilibrium, but does not alter the proportions of reactants and products in the equilibrium mixture.

VII Industrial Application of Le Châtelier's Principle



Equilibrium position is used to describe qualitative whether the reactants or products are more plentiful in the system. Equilibrium position lies to the LEFT \leftarrow i.e. the reactants predominate or to the RIGHT \rightarrow products predominate.

Note : Haber process is an example of heterogeneous catalysis, recall: mechanism for catalysis.

Contact process is an example of homogeneous catalysis.

The Equilibrium Law

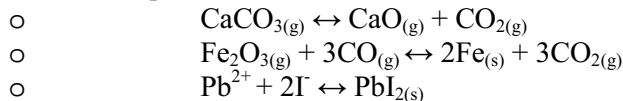
- Relates the concentrations of reactants and products
- States that for any system at equilibrium, there is a numerical relationship between the concentrations of products raised to the power of their stoichiometric numbers, and the concentrations of the reactants, raised to the power of their stoichiometric numbers. This is called the Equilibrium constant K_c .

$K_c = \frac{[products]}{[reactants]}$ - For heterogeneous equilibria, pure solid or pure liquid – their concentrations remain constant.

example :

$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

- For example:



Reversing an equation (such as the one above) can be done as follows: $K_c = \frac{1}{K_c}$

Multiplying an equation:

$$K_c = (K_c)^n$$

Dividing an equation:

$$K_c = \sqrt[n]{K_c}$$

I Significance of the Numerical Value of the Equilibrium Constant

Indicates whether there are more reactants or products in the systems at equilibrium

If K_c is:

Large \rightarrow products predominate

Small \rightarrow reactants predominate

Between 10^{-2} and 10^2 \rightarrow the reaction will not happen

II Units of K_c

- Equilibrium constants for different systems can have different units.

- For example:

$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3} \Rightarrow K_c = \frac{(mol\ dm^{-3})^2}{(mol\ dm^{-3})(mol\ dm^{-3})^3} \Rightarrow K_c = dm^6\ mol^{-2}$$

III Effect of Temperature on the Numerical Value of the Equilibrium Constant

- If the forward reaction is endothermic, an increase in temperature will increase K_c .
- If the forward reaction is exothermic, an increase in temperature will decrease K_c .

IV Calculations Involving K_c

- 1) Calculation of K_c from equilibrium concentrations given.
- 2) Calculation of K_c from initial concentration and concentration of products at 1L container
- 3) Calculation of K_c from initial concentration and concentration of products with different volume of container.
- 4) Calculation of equilibrium concentration, given K_c and given initial concentration.
Ex. $H_2 + I_2 \leftrightarrow 2HI$ $K_c = 25$
0.10 M 0.10 M
- 5) Calculations of equilibrium concentration given K_c is very, very small and the initial concentration, (make assumption ...)
- 6) Reaction Quotient vs. reaction at equilibrium? If not, in which direction?

Relationship of Gibbs Free Energy to K_c

$$\Delta G = -RT \ln K_c \quad \text{large } K_c \text{ means reaction is spontaneous}$$