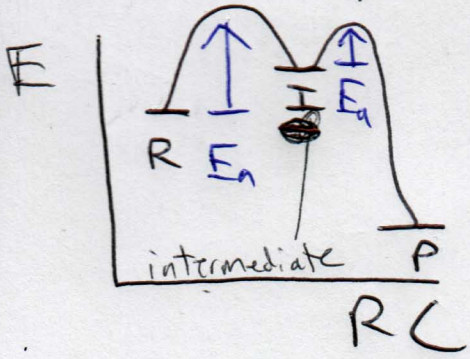
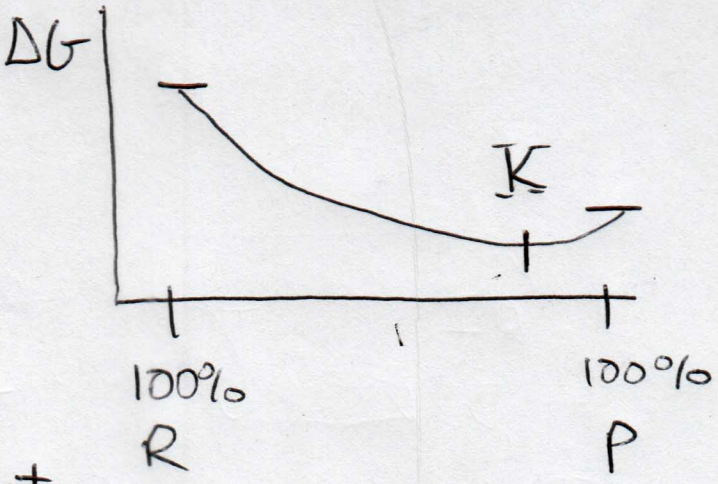
‡ ‡ transition state





K = equilibrium constant

$$= \frac{\text{products}}{\text{reactants}}$$

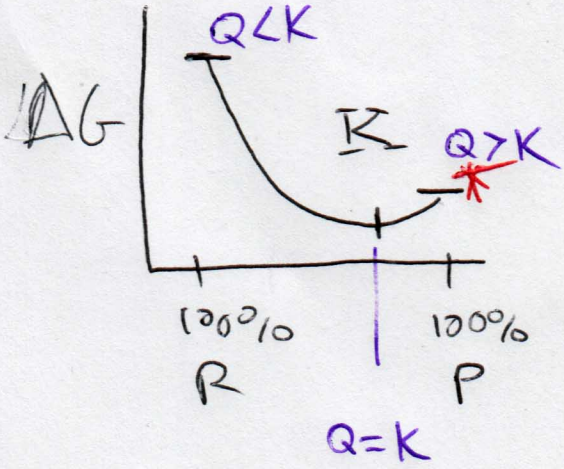
$$= \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

$aA + bB \rightleftharpoons cC + dD$

Only substances in the same phase are included in an equilibrium constant.

$\Delta G = -RT \ln K$

$\Delta G = \Delta H - T\Delta S$



$NaOH_{(aq)} + CH_3COOH_{(aq)} \rightarrow CH_3COONa_{(aq)} + H_2O_{(l)}$

Q → reaction quotient.
→ has the same mathematical form as K, but non-equilibrium concentrations can be used.

Imagine that a strong base and a weak acid react to 100% extent. This reaction would cause the weak acid to dissociate beyond its equilibrium point. This means the system is not at its lowest possible energy point ($\Delta E \neq 0$), there are too many products ($Q > K$), and the rate of reverse rxn is greater than the rate of the forward rxn.

This means that the salt formed in neutralization will react with water to reform hydroxide since this will lower Q, which will raise the rate of the forward rxn, bringing the system back into equilibrium.

Since hydroxide was formed, there will be an imbalance of $[H^+] + [OH^-]$ in solution. So, although equal moles of acid and base might have been reacted (neutralized), the solution will not be neutral ($[H^+] \neq [OH^-]$).

$\Delta G = \Delta H - T\Delta S$

A favorable entropy change ($\Delta S > 0$) can compensate for an unfavorable enthalpy change ($\Delta H > 0$), resulting in an overall release of free energy ($\Delta G < 0$).

$\Delta G < 0$ exergonic — $K > 1$ spontaneous

$\Delta G > 0$ endergonic — $K < 1$ non-spontaneous

$\Delta G = -RT \ln K$

↳ This means the rxn will have a large K, which means it is favorable for it to occur.

K_{sp} → solubility product

$$AgCl_{(s)} \rightleftharpoons Ag^+_{(aq)} + Cl^-_{(aq)}$$

Because this solid does not have a calculatable concentration in solution, it is not included in the equilibrium expression.

$$K_{sp} = [Ag^+][Cl^-]$$

Precipitation occurs when $Q_{sp} > K_{sp}$.