Equilibrium Constants

$$React \rightleftharpoons Prod$$

$$K = \frac{[Prod]}{[React]}$$

EQ: RATE Forward = RATE Reverse

IGNORE Pure Solids + Liquids

(s) (l)

WHY?

Feb 26-7:36 AM

(aq) = solid dissolved in a liquid.

$NaCl(aq) \rightarrow Na^+(aq)$

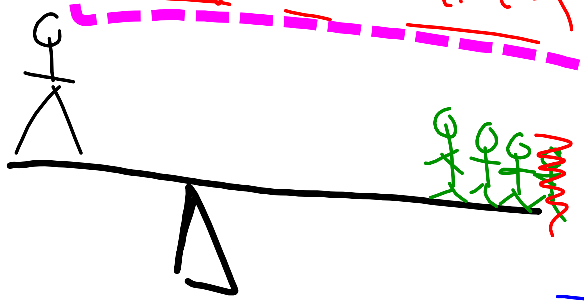
$\rightarrow Cl^-(aq)$

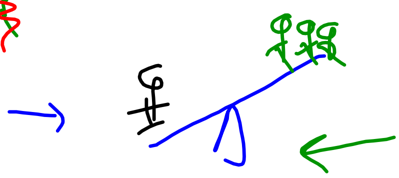
Can affect K

→ Change the concentration → Add/subtract

NaCl

H_2O





Feb 26-8:10 AM

✓ gas (g)

$O_2(g)$ → Add $O_2(g)$
↘ Remove $O_2(g)$

Δ Pressure
↓
Concentration of a gas

Feb 26-8:14 AM

DON'T INCLUDE in K (equilibrium) expression

Pure solids	Pure liquids
$NaCl(s)$ → Add $NaCl(s)$ ↘ Remove $NaCl(s)$	$H_2O(l)$
No Δ concentration	Add/remove $H_2O(l)$ No Δ [ ]

Feb 26-8:15 AM

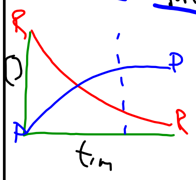
What does $\underline{K}$ (K_c or K_{eq})
equilibrium constant tell us about a reaction?

$$K = \frac{[\text{Prod}]}{[\text{React}]} \text{ AT EQ.}$$

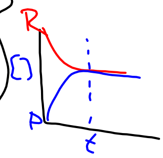
$K >> 1$ BIG #
$K = 1$
$K << 1$ Small #

$\frac{[P]}{[R]} = \frac{1000}{2}$

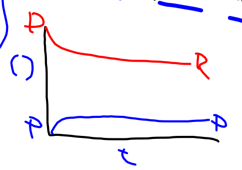
$K >> 1$
AT EQ
$[P] >> [R]$
rxn proceeds forward
make lots of product



$K = 1$
AT EQ
$[P] = [R]$



$K << 1$ $\frac{2}{1000}$
AT EQ
$[P] < [R]$
Very little product formed



Feb 26-8:22 AM

K_c
concentration (aq)
Molarity = $\frac{\text{moles solute}}{\ell \text{ solution}}$
[ ]

K_p
concentration in terms of pressure
GASES! (g)

$N_2O_{4(g)} \rightleftharpoons 2NO_{2(g)}$

$$K_p = \frac{P(NO_2)^2}{P(N_2O_4)^1}$$

Feb 26-8:32 AM

$$PV = nRT$$

$$P = \frac{n}{V}RT$$ (n: moles; V: l)

$$P = M\,RT$$

$$K_p = K_c(RT)^{\Delta n}$$

Δn = change in # moles = (Σ moles product) − (Σ moles react)

Feb 26-8:36 AM

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$$

$$K_p = \frac{(P_{NH_3})^2}{(P_{N_2})^1 (P_{H_2})^3}$$

$K_c = 9.6$ at 300°C

$K_p =$?

$$K_p = K_c(RT)^{\Delta n}$$

$2 - (3+1) = -2$

$$K_p = 9.6((0.08206)(573))^{-2}$$

$$K_p = 4.34 \times 10^{-3}$$

Feb 26-8:39 AM

$$\frac{K_p}{(RT)^{\Delta n}} = \frac{K_c (RT)^{\Delta n}}{(RT)^{\Delta n}}$$

$$K_c = \frac{K_p}{(RT)^{\Delta n}} = K_p (RT)^{-\Delta n}$$

Feb 26-8:51 AM

Hess's Law

K_c

① $HF_{(aq)} \rightleftharpoons H^+_{(aq)} + F^-_{(aq)}$ $K_1 = 6.8 \times 10^{-4}$

② $H_2C_2O_{4(aq)} \rightleftharpoons 2H^+_{(aq)} + C_2O_{4(aq)}^{-2}$ $K_2 = 3.8 \times 10^{-6}$

Find K for: $2HF_{(aq)} + C_2O_{4(aq)}^{-2} \rightleftharpoons 2F^-_{(aq)} + H_2C_2O_{4(aq)}$

×2 $2HF \rightleftharpoons 2H^+ + 2F^-$ $K_1 = (6.8 \times 10^{-4})^2$

Flip $2H^+ + C_2O_4^{-2} \rightleftharpoons H_2C_2O_4$ $K_2 = \left(\frac{1}{3.8 \times 10^{-6}}\right)$

$2HF + C_2O_4^{-2} \rightleftharpoons 2F^- + H_2C_2O_4$ $K_1 * K_2$

$K = 0.12$

Feb 26-8:52 AM

Difference between ΔH an K
Hess's Law

ΔH	EQN	K (one power of math higher)
Change Sign +/-	Flipped	$\frac{1}{K}$
Double (*2)	Double	K^2
Add The ΔH's	Find Final Eqn. Add Eqns.	Mult K's

Feb 26-8:59 AM

RICE Tables

R = reaction

I = [Initial] initial concentration at time = 0

C = [Change] change in concentration

E = [Equilibrium] Eq. conc.

Feb 26-9:08 AM

Find K

Molarity ↓

$$H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$$

	H_2	I_2	$2HI$
[I]	1×10^{-3}	2×10^{-3}	Ø
(Δ)	① -0.935×10^{-3}	① -0.935×10^{-3}	② $+1.87\times10^{-3}$ (MOLE RATIO)
⇒ [E]	0.065×10^{-3} = 6.5×10^{-5} M	1.065×10^{-3} M	1.87×10^{-3}

$$K = \frac{(P_{HI})^2}{(P_{H_2})(P_{I_2})} = \frac{(1.87\times10^{-3})^2}{(6.5\times10^{-5})(1.065\times10^{-3})} = 50.515$$

K >> 1

Favors product formation

Feb 26-9:10 AM

HW

15/ 23, 26, 32

Feb 26-9:27 AM

(15.31)

$2NO + 2H_2 \rightleftharpoons N_2 + 2H_2O$

I 0.1 mole 0.05 0.1 mole 1ℓ

E 0.062M

	(2)	(2)	(1)	(2)
I	0.1	0.05	0	0.1
Δ	− 0.938	−0.938	+½(0.938)	+ 0.938
E	0.062M			

Feb 26-9:29 AM