

$$K = \frac{[\text{Prod}]^{\text{coeff}} * [\text{Prod}]^{\text{coeff}}}{[\text{React}]^{\text{coeff}}}$$

$$K = \frac{\text{small}}{\text{Big}} = < 1$$

favors reactants
 small #

3 moles → 1 mole
 Rate \xrightarrow{f}
 \xleftarrow{r} Rate

favored product formation

$$K = \frac{(P)}{(R)} > 1$$

Feb 7-7:47 AM

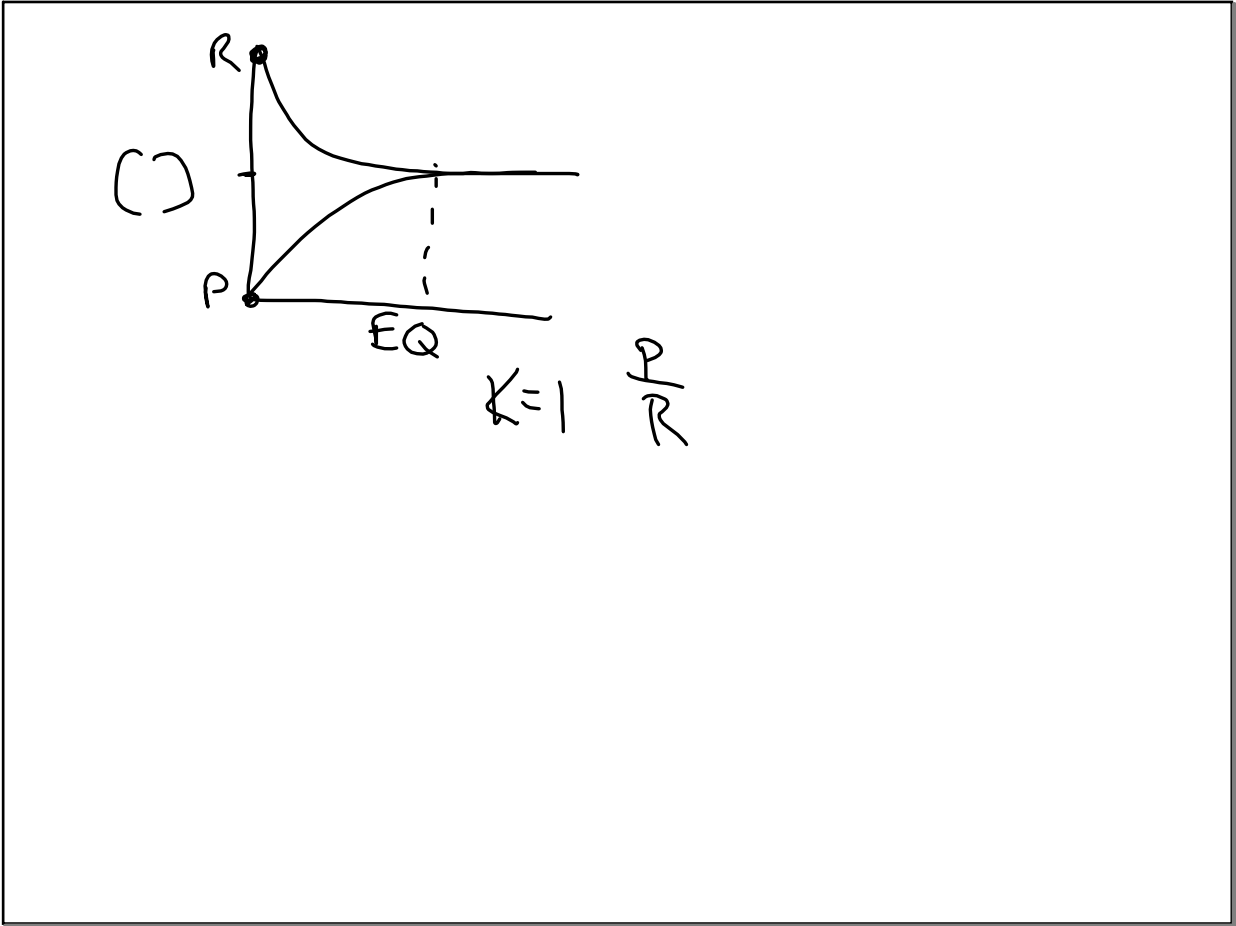
$$K = \frac{P}{R} = \frac{\text{small}}{\text{Big}} = < 1$$

favors react.

$$K = \frac{P}{R} = \frac{\text{Big}}{\text{small}} > 1$$

favors prod. formation

Feb 7-8:18 AM

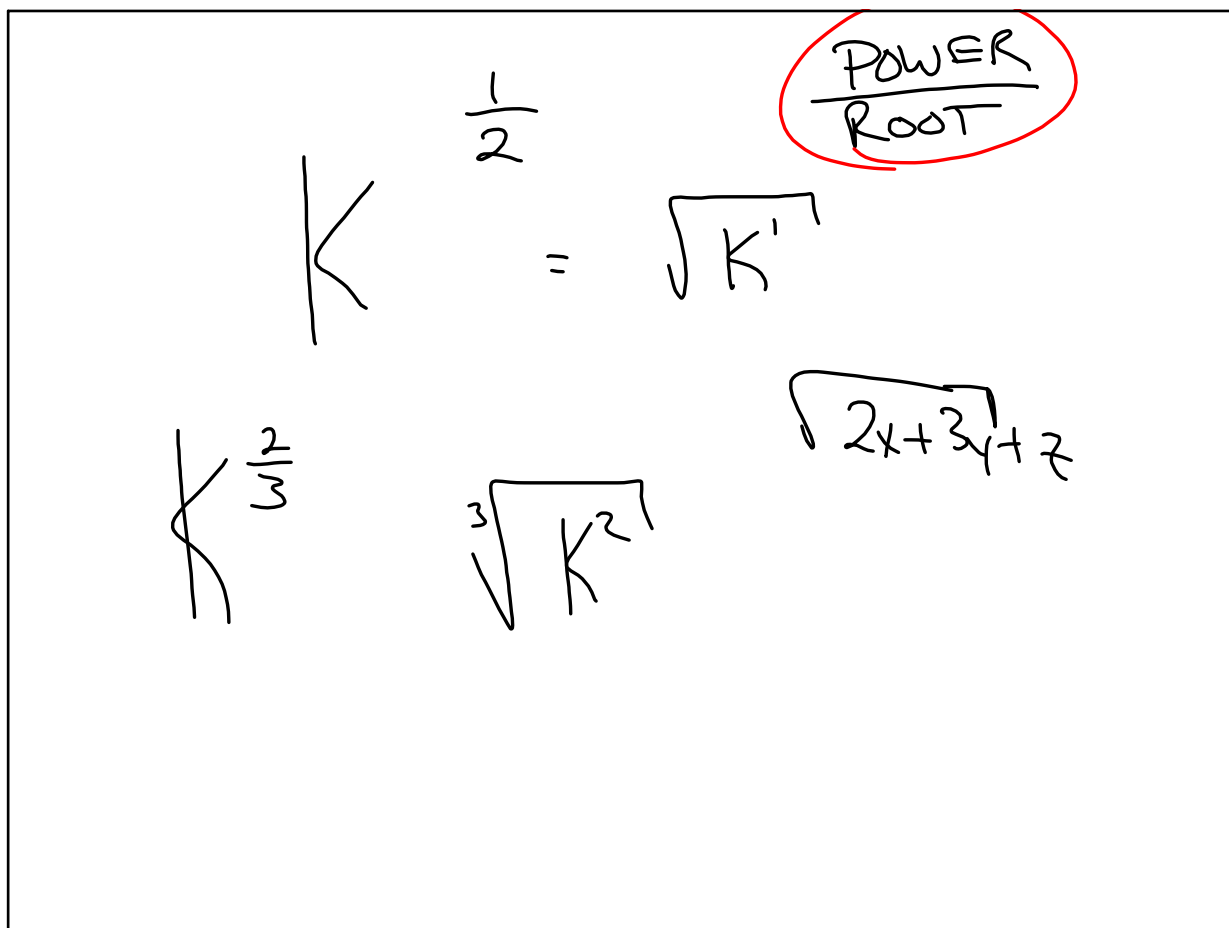


Feb 7-8:21 AM

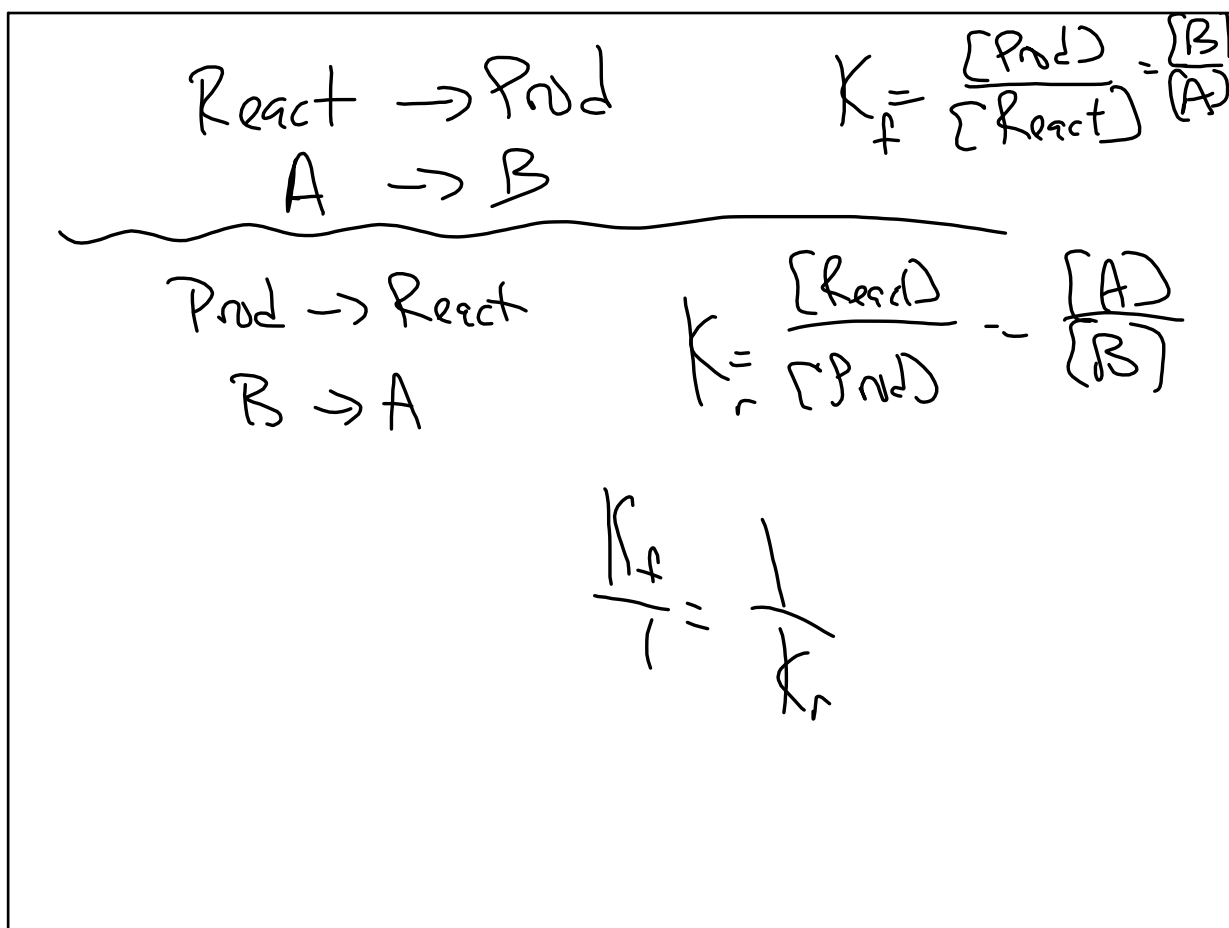
Diff between ΔH and K

<u>Hess's law</u>	ΔH <small>part of mole ratio</small>	K
Double the Eqn 1:3:2 2:6:4 <small>mole RATIO</small>	Double ΔH $2(\Delta H)$	Square K K^2
half Eqn	$\frac{1}{2}(\Delta H)$	$K^{1/2}$ or \sqrt{K}
Flip eqn A \rightarrow B B \rightarrow A	Negate ΔH $-\Delta H$	$\frac{1}{K}$ invert K^{-1}
Add Eqns ① A \rightarrow B ΔH ② B \rightarrow C ΔH <hr/> A \rightarrow C	Add ΔH $\Sigma \Delta H$ $\Delta H_1 + \Delta H_2$	mult K $K_1 \times K_2$

Feb 7-8:25 AM



Feb 7-8:30 AM



Feb 7-8:33 AM

$$\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$$

$K_c = 9.6$ at 300°C Find K_p

$\Delta n = 2 - 4 = -2$

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

$$= 9.6 \left[(0.08206)(573) \right]^{-2}$$

KELVIN!
↓

$K_p = 0.004$

Feb 7-8:41 AM

Change Δ \rightarrow $\begin{matrix} [R] \\ [P] \end{matrix}$ decreases / increases from $t=0 \Rightarrow t=t$ EQ

Tables

R	Balanced Reaction	Used to calc K_c or K_p
I	[Initial] or P_{initial}	
C	change in $[\]$ or P	
E	Equilibrium $[\]$ or P	

Start
difference
E_{AN}

Feb 7-8:46 AM

448°C

$$\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2 \text{HI}(\text{g})$$

$t=0$	$\text{H}_2(\text{g})$ I	$\text{I}_2(\text{g})$ I	$2 \text{HI}(\text{g})$ I
	$1 \times 10^{-3} \text{ M}$	$2 \times 10^{-3} \text{ M}$	
	$-\frac{1}{2}(1.87 \times 10^{-3})$	$-\frac{1}{2}(1.87 \times 10^{-3})$	$+1.87 \times 10^{-3}$
	6.5×10^{-5}	1.065×10^{-3}	1.87×10^{-3}

MOLE RATIO!

$K_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} = \frac{(1.87 \times 10^{-3})^2}{(6.5 \times 10^{-5})(1.065 \times 10^{-3})} = 50.5$

K_c at 448°C
 EQ constant

Feb 7-8:53 AM

$$2 \text{SO}_2(\text{g}) \rightleftharpoons 2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g})$$

$t=0$	$2 \text{SO}_2(\text{g})$ I	$2 \text{SO}_2(\text{g})$ I	$\text{O}_2(\text{g})$ I
	0.5 atm		
	-0.3	$+0.3$	$+0.15$
	0.2 atm	0.3	0.15

MOLE RATIO

$K_p = \frac{P(\text{SO}_2)^2 \cdot P(\text{O}_2)}{P(\text{SO}_2)^2} = \frac{(0.3)^2 (0.15)}{(0.2)^2} = 0.3375$

K_p at 1000K
 EQ constant

Feb 7-9:11 AM

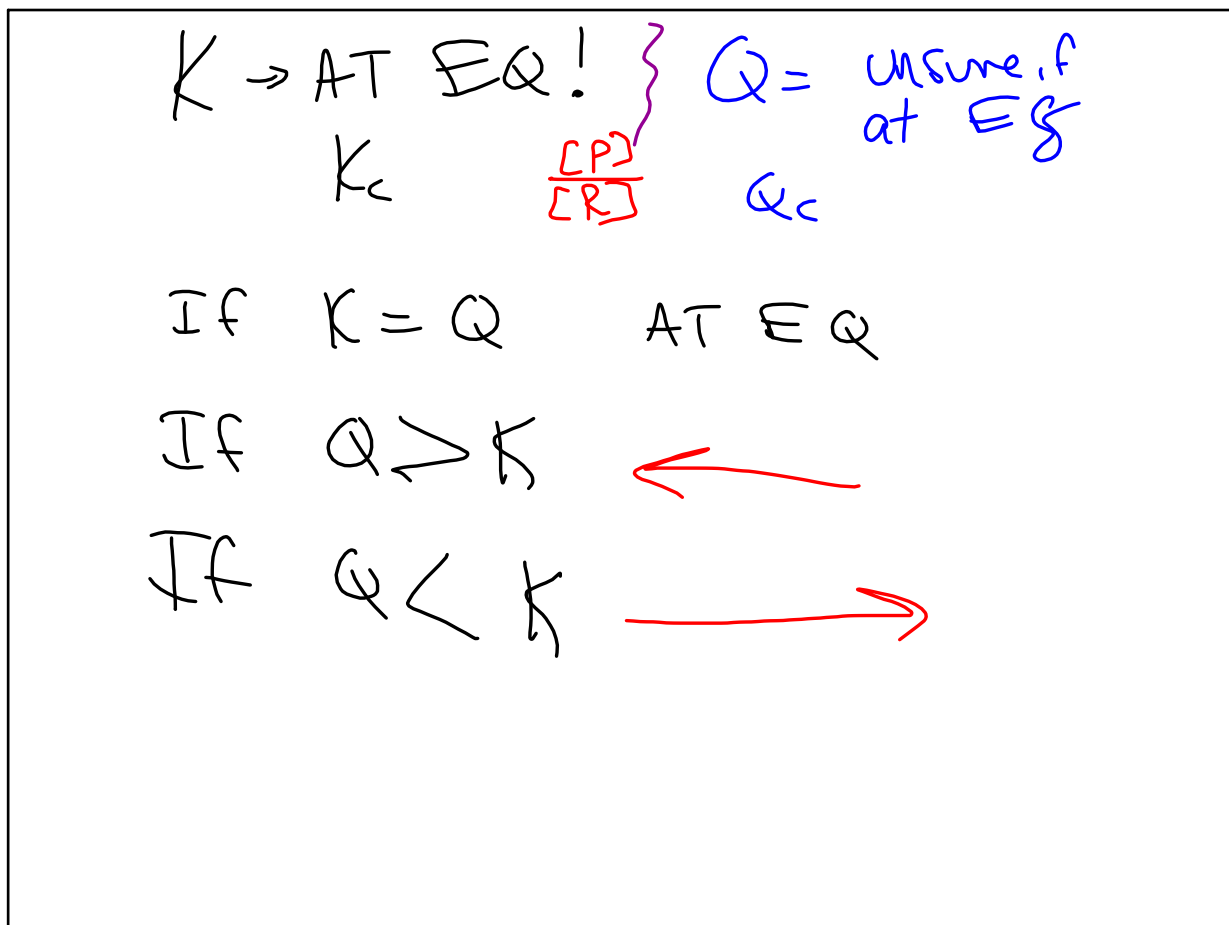
$$K_p = 0.3375 \quad \Delta n = 1$$

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

$$\frac{0.3375}{82.06} = K_c \left[\frac{(0.08206)(1000)}{82.06} \right]^1$$

$$K_c = 0.00411$$

Feb 7-9:19 AM



Feb 7-9:21 AM

$H_2 + I_2 \rightleftharpoons 2HI$

I			}	$\frac{M}{l} = \frac{\text{mole}}{l}$
Δ				
E	$1 \times 10^{-2} \text{ mole}$	$3 \times 10^{-2} \text{ mole}$		

AT EQ?

$$Q = \frac{(HI)^2}{(H_2)(I_2)} = \frac{\left(\frac{2 \times 10^{-2}}{2}\right)^2}{\left(\frac{1 \times 10^{-2}}{2}\right)\left(\frac{3 \times 10^{-2}}{2}\right)} = 1.3$$

$Q = 1.3 < K = 50.5$

→

molarity

Feb 7-9:24 AM

PSIS # 1-12
23-25

Feb 7-9:30 AM