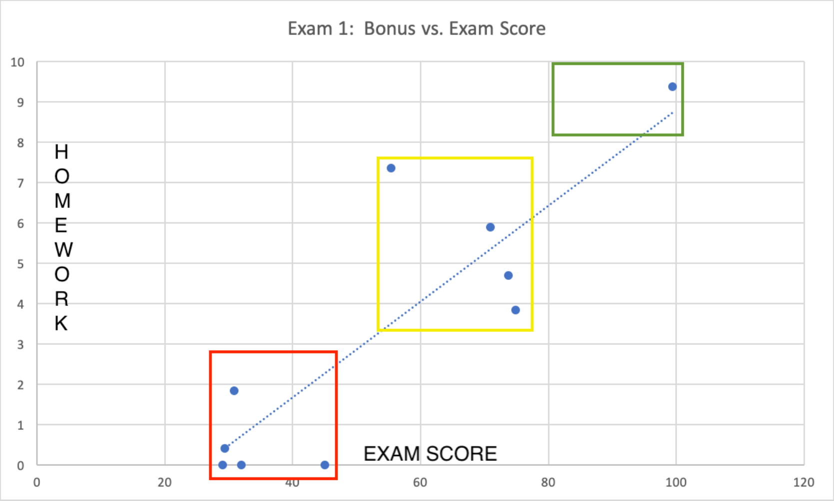
1412 - E2
CH 13 - Equilibrium
Notes

13


Exam 1: Bonus vs. Exam Score
HOMEWORK
EXAM SCORE
10
9
8
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6
5
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3
2
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20
40
60
80
100
120

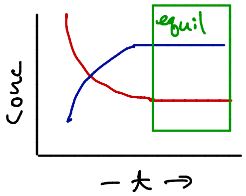
CHAPTER 13
CHEMICAL EQUILIBRIA

Introduction [13.1]

Basic idea of this chapter is to set up equations which all one to calculate the concentrations of reactants and products after a reaction has “run it’s course” (after the rate is no longer changing)

One-way versus Equilibrium Reactions

- For some reactions virually all reactants are converted to product (“one way”)
- For other reactions, only part of the reactants are converted to products (“equilibrium”)

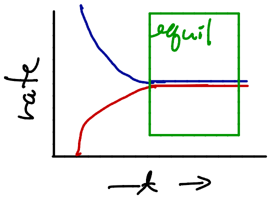


One-way $A \rightarrow B$

Dynamic Equilibrium $A \rightleftharpoons B$

e.g., PHASE CHANGES are often in equilibrium

$Br_2(l) \rightleftharpoons Br_2(g)$



Reaction Quotients & Equilibrium Constants [13.2]

$$aA + bB \rightleftharpoons cC + dD$$

molar
concentrations

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

$$Q_c = \frac{[PRODUCTS]^p}{[REACTANTS]^r}$$

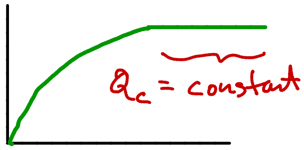
… for example…

$4\ HN_3(g) + 7\ O_2(g) \rightarrow 4\ NO_2(g) + 6\ H_2O(g)$

$$Q_c = \frac{[NO_2]^4[H_2O]^6}{[NH_3]^4[O_2]^7}$$

Reaction Quotient vs. Equilibrium Constant

- At equilibrium, the ‘REACTION QUOTIENT’ is referred to as the EQUILIBRIUM CONSTANT, which is assigned the variable “Kc” (rather than Qc)



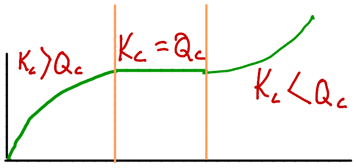
$Q_c = \overbrace{\text{constant}}$ $\rightsquigarrow$ Equilibrium Constant = K_c

Caveat!!!

Kc changes with temperature

Shifting toward Equilibrium: Kc vs. Qc

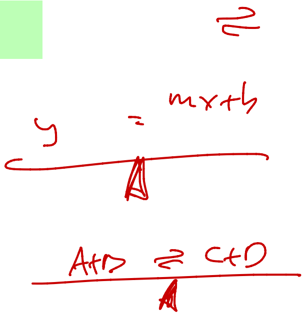
- One way to know if a reaction is at equilibrium is to compare its calculated Qc value to the reactions known Kc value (this assumes Kc is known)
  - ↳ if Kc > Qc, then the reaction will shift to the right (the measured situation does not have enough product)
  - ↳ if Kc < Qc, then the reaction will shift to the left (the measured situation has too much product)
- Remember, Qc will change until it equals Kc; that is, the reaction will keep working toward equilibrium



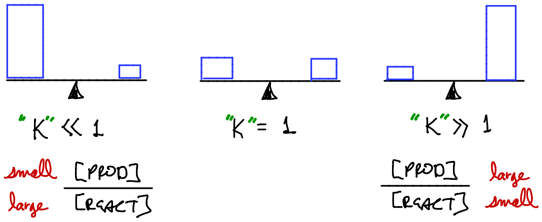
memory device

$\xrightarrow[\text{SHFT RT}]{K > Q}$ $\quad K = Q \quad$ $\xleftarrow[\text{SHIFT LFT}]{K < Q}$

Visual aid: Relative Sizes of "K"



↳ If K = 1, then [PRODUCTS]^p is equal to [REACTANTS]^r
↳ If K < 1, then [PRODUCTS]^p is less than [REACTANTS]^r
↳ If K > 1, then [PRODUCTS]^p is more than [REACTANTS]^r



- In industry, companies tend to want to make product, and so desire reactions for which K is as large as possible, which means little reactant is left over (unreacted). This translates to more product to sell and less waste.

(EX) K vs. Q – Reaction Shift
¿If Kc for a reaction is 65.0, but the current concentrations of reactant and products give Q = 38.1, what will happen to the reaction mixture: it will shift to the right; it will shift to the left; or it will remain as it is?

$K_c > Q$ ∴ shift rxn to right

65.0 38.1

— Need to know next three sections to be able to write expressions for “K” —

Activity vs. Concentration

- Activity and Concentration are related, but they are not exactly the same…

- ACTIVITY is the “effective concentration” at specified conditions
  ↳ activities are dimensionless
  ↳ for DILUTE SOLUTIONS, activity ~ [conc]
  ↳ for PURE SOLIDS AND LIQUIDS (consensed phases), activity = 1
  ↳ for BULK SOLVENTS, assume activity = 1

NOTE:
recognizing the above unity activity situations can greatly simply some of the calculations to come… factors of “1” can oft be ignored, simplying many equations which would otherwise be more complicated

Homogeneous Equilibria

- All reactants and products are present in the same phase
- Two types
  - ↳ all liquid solution
  - ↳ all gas solution
- Review
  - ↳ (s) = solid, activity = 1
  - ↳ (l) = liquid, activity = 1
  - ↳ (g) = gas
  - ↳ (aq) = aqueous, activity = [conc] if dilute

Partial Pressure vs. Concentration

- Because Partial Pressure is proportional to [CONC], equilibrium constants for solutions containing gases can be derived using pressure

$$PV = nRT \rightsquigarrow P = \frac{n}{V}RT \rightsquigarrow PV = MRT \rightsquigarrow P \propto M$$

Generic expressions: Can write GAS solutions in terms of Partial Pressure or Concentration

$$aA(g) \rightleftharpoons bB(g) + cC(g)$$

$$K_p = \frac{(P_B)^b(P_C)^c}{(P_A)^a} \qquad K_c = \frac{[B]^b[C]^c}{[A]^a}$$

Caveat!!!

$$K_p \neq K_c$$

(EX) Expressions for K
¿Write expressions for Kc and/or Kp for the following reactions?

$2HgO(s) \rightleftharpoons 2Hg(l) + O_2(g)$ $K_c = [O_2]$ $K_p = P_{O_2}$

$2NH_3(g) + H_2SO_4(l) \rightleftharpoons (NH_4)_2SO_4(s)$

$$K_c = \frac{1}{[NH_3]^2} = [NH_3]^{-2}$$

$$K_p = \frac{1}{(P_{NH_3})^2} = (P_{NH_3})^{-2}$$

(EX) Expressions for K
¿Write expressions for Kc and/or Kp for the following reactions?

In Openstax ...

- $PbCl_2(s) \rightleftharpoons Pb^{2+}(aq) + 2Cl^-(aq)$ $K_c = [Pb^{2+}][Cl^-]^2$
- $CaO(s) + CO_2(g) \rightleftharpoons CaCO_3(s)$ $K_c = 1/[CO_2]$
- $C(s) + 2S(g) \rightleftharpoons CS_2(g)$ $K_c = [CS_2]/[S]^2$
- $Br_2(l) \rightleftharpoons Br_2(g)$ $K_c = [Br_2]$
- $CaO(s) + CO_2(g) \rightleftharpoons CaCO_3(s)$ $K_p = 1/P_{CO_2}$
- $C(s) + 2S(g) \rightleftharpoons CS_2(g)$ $K_p = P_{CS_2}/(P_S)^2$

Converting Kp-to-Kc, and vice versa

Kc(forward) vs. Kc(reverse)

$$A \underset{r}{\overset{f}{\rightleftharpoons}} B \qquad K_f = \frac{[B]}{[A]} \xrightarrow{\text{flip}} K_r = \frac{[A]}{[B]} = \frac{1}{K_f}$$

$$K_{c,reverse} = \frac{1}{K_{c,forward}}$$

Kc vs. Kp

$$K_p \underset{g}{=} K_c (RT)^{\left(\text{total moles PROD side} - \text{total moles REACT side}\right)}$$

BCE coefficients

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

(EX) Calculate $\boxed{Kp}$ from Kc
¿The equilibrium constant for the Haber Process at 500 °C is 0.286. What is Kp??

$$\underbrace{N_2(g) + 3H_2(g)}_{4\ mol} \rightleftharpoons \underbrace{2NH_3(g)}_{2\ mol}$$

$K_c = 0.286$

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

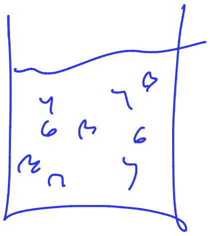
$$K_p = K_c(RT)^{\Delta n} = (0.286)\left[(0.0821)(773)\right]^{(2-4)} = \boxed{7.10 \times 10^{-5}}$$

How Kc changes with a Coefficient Change

$$K_{p,\underset{new}{\cancel{original}}} = K^{n}_{p,\underset{orig}{\cancel{new}}}$$

n = multiplier factor

orig × n = new



e.g. $2A \rightleftharpoons 2B + C$ $\quad K_c = \dfrac{(3)^2 \cdot 4}{(2)^2} = 9$

$\downarrow \times \frac{1}{2}$

$A \rightleftharpoons B + \frac{1}{2}C$ $\quad K_c = \underbrace{\dfrac{3 \cdot (4)^{1/2}}{2}}_{\text{concentrations don't change b/c chemist uses different coeff.}} = \underbrace{3}_{\text{K}_c \text{ changes so [conc] can stay the same}}$

(EX) cont'd
In a previous example, K was determined to be 24.5,

$$2\,N_2O \rightleftharpoons 2\,N_2 + O_2$$

$K_c = 24.5$

¿What is Kc for the reaction…

$$N_2O \rightleftharpoons N_2 + ½\,O_2$$

The multiplication factor is ½. Don't start over and perform RICE. Rather…

$$K_{new} = (K_{orig})^{n\text{-}factor} = (24.5)^{1/2}$$

$$\boxed{K_{new} = 4.95}$$

(EX) Calculate Kc from RICE
¿10.0 moles of N2O are placed into a 2.00 L container, where it decomposes. At equilibrium, 2.20 moles of N2O remain. What is Kc?

(EX) Calculate Kc from RICE

¿10.0 moles of N2O are placed into a 2.00 L container, where it decomposes. At equilibrium, 2.20 moles of N2O remain. What is Kc?

2.00 L
10.0 mols → 2.20 mols @ equil

EAST

$M = \frac{mol}{L} = \frac{10.0}{2.00} = 5.00$

$\frac{2.20}{2.00} = 1.10$

	$2\,N_2O$	⇌	$2\,N_2$	+	O_2
R	$2\,N_2O$	$\rightleftharpoons$	$2\,N_2$	+	O_2
I	5.00		0		0
C	−3.90		+3.90		$\frac{1}{2}(3.90)$
E	**1.10**		3.90		1.95

← × K

← × K

← EQ ← K

$$K_c = \frac{[N_2]^2[O_2]}{[N_2O]^2} = \frac{(3.90)^2 \,|\, 1.95}{(1.10)^2} = \boxed{24.5}$$ unitless

Wed. Oct 2

(EX) Calc EQ concentration from Kc

¿What are the equilibrium concentrations for A,B,C,D if one started with 0.400 mol of A & B in a 2.00 L container, and knowing Kc = 49?

$\frac{0.400\text{ mol}}{2.00\text{ L}}$

R	A	+	B	→	C	+	D
I	0.200		0.200		0		0
C	0.175 $-x$		0.175 $-x$		0.175 $+x$		0.175 $+x$
E	$0.200-x$ 0.025		$0.200-x$ 0.025		x 0.175		x 0.175

$$K_c = 49 = \frac{\overset{x}{[C]}\,\overset{x}{[D]}}{\underset{0.200-x}{[A]}\,\underset{0.200-x}{[B]}}$$

$$K_c = 49 = \frac{x \cdot x}{(0.200-x)(0.200-x)} = \sqrt{\frac{(x)^{\cancel{2}}}{(0.200-x)^{\cancel{2}}}} = \sqrt{49}$$

$$\frac{x}{0.200-x} = 7$$

$x = 7(0.200-x)$

$x = 1.400 - 7x$

$8x = 1.400$

$$\boxed{x = 0.175}$$

⬇

(EX) Calc EQ concentration from Kc
¿What are the equilibrium concentrations for A,B,C,D if one started with 0.400 mol of A & B in a 2.00 L container, and knowing Kc = 49?

$M_A = M_B = 0.400 \text{ mol}/2.00\text{L} = 0.200$

R	A	+	B	$\longrightarrow$	C	+	D
I	0.200		0.200		0		0
C	$-x$		$-x$		$+x$		$+x$
E	$0.2-x$		$0.2-x$		x		x

$K_c = 49$

$\Downarrow$

$$K_c = 49 = \frac{(x)(x)}{(0.2-x)(0.2-x)} = \frac{x^2}{(0.2-x)^2}$$

$$\sqrt{49} = \sqrt{\frac{x^2}{(0.2-x)^2}} \xrightarrow[\text{square}]{\text{perfect}} 7 = \frac{x}{0.2-x}$$

$7(0.2-x) = x$

$1.4 - 7x = x$

$1.4 = 8x$

$\boxed{0.175 = x}$

$\longleftarrow$

$A = 0.200 - 0.175 = 0.025M$

$B =$ " $= 0.025M$

$C = x = 0.175$

$D = x = 0.175$

The Quadratic Equation

for $ax^2 + bx + c = 0$

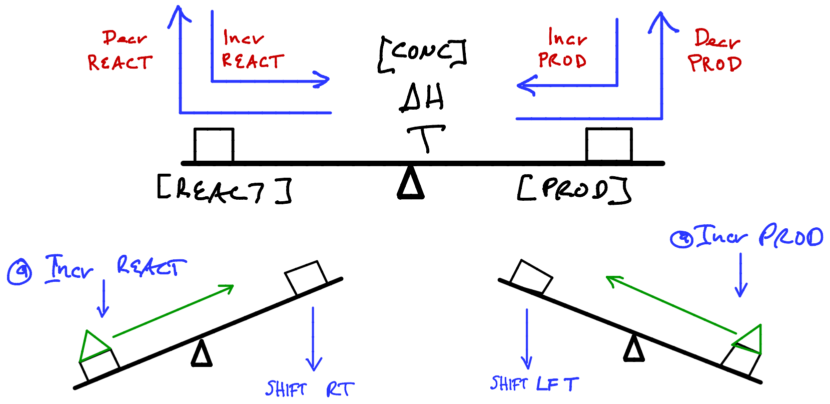
$$\boxed{x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}}$$

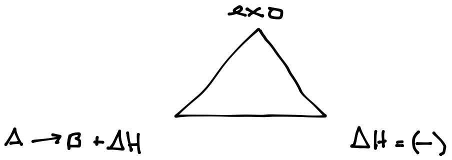
Shifting Equilibria [13.3]

$$K_c = \frac{[PRODUCTS]^p}{[REACTANTS]^r}$$

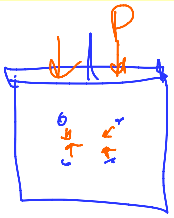
If any one changes, another must too...
... to RESTORE K

Rock-the-Boat: LeChatlier's Principle

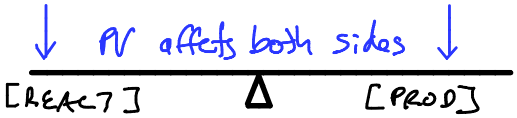


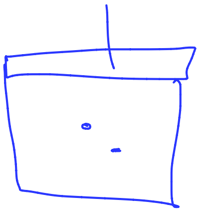


Gases



$$P = (n/V)\,RT$$ [conc]





- sys wants to keep same amount of "CROWDEDNESS"
  - ⓐ ↑P, ↓V → more crowded → SHFT to side that gives SMALLER # particles
  - ⓑ ↓P, ↑V → less crowded → SHFT to side that gives LARGER # particles

Catalyst

No affect on K

(EX) Predict Shift
¿In a closed container at 500 °C, predict the effect if . . .

$$\underbrace{N_2(g) + 3\,H_2(g)}_{4} \rightarrow \underbrace{2\,NH_3(g)}_{2} + \Delta H \qquad \Delta H = -92 \text{ kJ/mol rxn}$$

ⓐ Force more ↑H_2 into SYS	SHFT RT →	↑ (HABER)
ⓑ Remove ↓NH_3	SHFT RT →	
ⓒ Remove ↓N_2	SHFT LFT ←	
ⓓ ↑Pressure by ↓Vol	SHFT RT → (to ↓ crowding)	
ⓔ ↑Temp	SHFT LFT ← (rxn is EXO)	
ⓕ add catalyst	NO EFFECT	

The Haber Process Dilemma
(a case study of when equilibrium and kinetics work against each other)

$N_2(g) + 3\,H_2(g) \rightarrow 2\,NH_3(g)$ $\Delta H = -92$ kJ/mol rxn

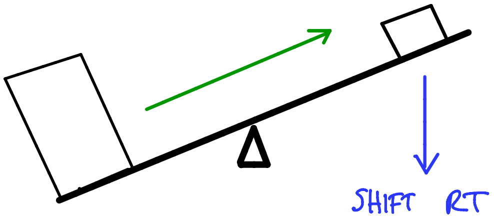
↳ Add heat shifts EQ LEFT, but ...

↳ Remove heat makes RXN uneconomically slow (slow @ <450°C)

⇓

Remedy

- ○ Remove product to drive RXN RT
- ○ High Temp replaces product rapidly



Equilibrium Calculations [13.4]

Determining Relative Changes — The “I” in “RICE”

- Ratio per COEFF in BCE
- Sign per whether or not material is lost (−) or gained (+)

$$\begin{array}{llllll} \text{R} & C_3H_8(g) & + \; 5O_2(g) & \rightleftharpoons & 3CO_2(g) & + \; 4H_2O(g) \\ \text{I} & +x & +5x & & -3x & -4x \end{array}$$

Three Basic Types of Equilibrium Problems

1 eq/1 unk (p732)

- Calc K_c from $[\text{REACT}, \text{PROD}]^{EQ}$
- Calc $[\text{REACT}, \text{PROD}]^{EQ}$ from K_c ; all [CONC] known but one.
- Calc $[\text{REACT}, \text{PROD}]^{EQ}$ from K_c ; RICE (≠ RICEICE)

(EX) Calc [conc] for 1 Unknown from Kc [ex 13.7]

¿NO[redacted] is produced by reaction of N2 and O2 at high temperature. At 2000 K, Kc = 4.1E–4. Find the [NO] in the equilibrium mixture with air at 1 atm, at this temperature, knowing that in air, [N2] = 0.036 M and [O2] = 0.0089 M?

T = 2000 K

R	N_2	+	O_2	→	$2NO$
I	0.036		0.0089		0
C	$-x$		$-x$		$+2x$
E	0.036 ~~0.036 − x~~		0.0089 ~~0.0089 − x~~		$\boxed{2x}$

✓

$$K_c = 4.1E-4 = \frac{[NO]^2}{[N_2][O_2]}$$

✓ ✓

LECTURE STOP

⬆

(EX) Calc [conc] for 1 Unknown from Kc [ex 13.7]
¿NOx is produced by reaction of N2 and O2 at high temperature. At 2000 K, Kc = 4.1E–4. Find the [NO] in the equilibrium mixture with air at 1 atm, at this temperature, knowing that in air, [N2] = 0.036 M and [O2] = 0.0089 M?

R	$N_2(g)$ +	$O_2(g)$ →	$NO(g)$
I			
C			
E	0.036	0.0089	$\boxed{}$

Kc = 4.1E–4

$$K_c = Q_c = \frac{[NO]^2}{[N_2][O_2]}$$

$$[NO]^2 = K_c[N_2][O_2]$$

$$[NO] = \sqrt{K_c[N_2][O_2]} = \sqrt{(4.1\times10^{-4})(0.036)(0.0089)}$$

$$[NO] = \boxed{3.6\times10^{-4}\ M}$$

Mon, Oct 7

(EX) EQ Shift Calculation: Quantitative

¿If a system at equilibrium has the following concentrations, what is the affect of halving the volume?

problem which ilustrates a number of the equilibrium concepts

@ EQ : $[NO_2] = 6.00\text{E-}2\,M$; $[N_2O_4] = 0.770\,M$

(0.0600 M)

new [INIT]

R	$N_2O_{4\,(g)}$	$\rightleftharpoons$	$2\,NO_{2\,(g)}$
I	2 (0.770)		2 (0.0600)
C	+ x		− 2x
E	1.54 + x		0.120 − 2x

halving VOL, double [ ]

$$K_c = \frac{[NO_2]^2}{[N_2O_4]} = 4.66\text{E-}3 \quad (0.00466)$$

halving VOL ⇒ SHFT LFT

(to ↓ crowding ; ↓ # part)

MAX value is 0.0600

(can't lose more than start with)

⇓

$$K = 4.66\text{E}-3 = \frac{(0.120 - 2x)^2}{1.54 + x}$$

Quadratic

$$x^2 - 0.120x + 1.80\text{E}-3 = 0$$

MAX = 0.0600 ← ~~x = 0.104~~ $\quad x = 0.017$

$$[NO_2] = (0.120 - 2x)$$
$$= 0.120 - 2(0.017)$$
$$= 0.120 - 0.034$$
$$= 0.086\,M$$

$$[N_2O_4] = (1.54 + x)$$
$$= 1.54 + 0.017$$
$$= 0.156\,M$$

Recall : new [INIT] = 0.120 M

orig [INIT] = 0.600

new [EQ] = 0.086

dec

incr

incr less than DBL

Recall : new [INIT] : 1.54 M

orig [INIT] = 0.770

new [EQ] = 1.56

incr

incr

incr more than DBL

Approximate Solutions: when K is very small (makes problem much easier)

- If K is very small, the Quadratic Eq can oft be avoided
- 5% rule is invoked

(EX) The 5% Rule Illustrated

¿What are the equilibrium concentratons in a 0.15 M HCN solution?

R	$HCN(aq)$	=	$H^+(aq)$	+	$CN^-(aq)$
I	0.15		0		0
C	$-x$		$+x$		$+x$
E	$0.15-x$		x		x

Kc = 4.9E–10

$$K_c = 4.9\times10^{-10} = \frac{(x)(x)}{(0.15-x)} \rightarrow 4.9\times10^{-10} = \frac{x^2}{0.15}$$

very small amt ... neglect

Will discuss next section, how you know when you can and can't ignore a value

$$x = \sqrt{(0.15)(4.9\times10^{-10})}$$

$$\boxed{x = 8.6\times10^{-6}\,M}$$

CHECK 5% RULE ←

(see next page)

The 5% Rule — How does it work?

- The basics idea is that the calculated concentration is acceptable IF that calculated value varies less than 5% of the starting amount.
- Depending on the situation, your employer might accept a 10% variance, or maybe want no more than a 1% variance… but for this class, we'll go with 5%

How to calculate %Variance

QUESTION: Is the variance less than 5%?
ANSWER: Calculate variance, then compare it to 5%.

$$\%\ variance\ = \frac{calculated\ amount}{initial\ amount} x\ 100$$

(example next page)

(EX) …continuing from the previous example problem…
¿Having calculated equilibrium concentratoins in a 0.15 M HCN solution using the 5% rule, is the final values acceptable?

$$\%\,variance = \frac{calculated\ amount}{initial\ amount\ (or\ C_a)} \times 100 = \frac{x}{0.15} \times 100 = \frac{8.6\times10^{-6}}{0.15} \times 100 = 0.006\ \%$$

OK: $0.006\ \% < 5\ \%$

Alternate mathematical version (easier math)

$Is\ C_a\ greater\ than\ 20\,x?$

$$C_a > 20\,x$$

$0.15 > 20(8.6\text{E}{-6})$

$0.15 > 0.0002$ $\therefore$ good assumption

$\frac{1}{20} > \frac{x}{C_a}$

"BEFORE-the-calculation" best guess as to whether or not the 5% rule will work
(from source other than the textbook)

$$Is\ C_a\ greater\ than\ 100\ K_a?$$

$$C_a > 100\ K_a$$

$0.15 > 100(4.9E-10)$

$0.15 > 4.9E-8$

∴ anticipated
good assumption

Before-the-fact: compare Ca to Ka. $C_a > 100\ K_a$

After-the-fact: compare Ca to x $C_a > 20\ x$

Select Review Questions

(EX) Calc % Decomposition of Product <hard>
¿For the following reaction and associated data, what % of HBr is decomposed at equilibrium, at 1297 K, starting with pure HBr?

For the rxn: $H_2(g) + Br_2(g) \rightleftharpoons 2\,HBr(g)$

Kc values are: $K_c = 35{,}000$ @ 1495 K

$K_c = 160{,}000$ @ 1297 K

Tme ↓ ↑ Kc

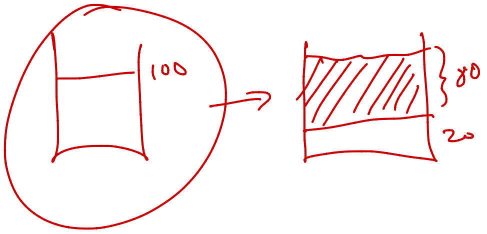
exo

$\frac{1}{2}H_2(g) + \frac{1}{2}Br_2(g) \rightleftharpoons HBr(g)$ $K_{c\,(new)} =$

$$\%\,Decomp = \frac{disappeared}{C_a} \times 100$$

part

whole



EAST

(EX) Calc % Decomposition of Product <hard>

¿For the following reaction and associated data, what % of HBr is decomposed at equilibrium, at 1297 K, starting with pure HBr?

For the rxn: $1\,H_2(g) + Br_2(g) \rightleftharpoons 2\,HBr(g)$ } all gas rxn.

K_c values are: $K_c = 35{,}000$ @ 1495 K } ↑ΔH= unfavorable (lowers K_c)

$K_c = 160{,}000$ @ 1297 K } ∴ must be EXOthermic

$\times 1/2$

R	$\frac{1}{2}H_2(g)$	+ $\frac{1}{2}Br_2(g)$	$\rightleftharpoons$ $HBr(g)$
I	0	0	1
C	x	x	$-x$
E	x	x	$1-x$

$$K_c = \left(K_{c,\text{orig}}\right)^{1/2}$$

$$= \sqrt{160{,}000} = 400$$

$$K_c = \frac{1-x}{(x)^{1/2}(x)^{1/2}} = \frac{1-x}{x} = 400$$

$$1 - x = 400x$$

$$1 = 401x$$

$$x = 0.00249$$

= amt dissoc

$$\Rightarrow$$

$$\% = \frac{C_{diss}}{C_{init}} \times 100$$

$$= \frac{0.00249}{1} \times 100$$

$$= \boxed{25\,\%}$$

$$\%\ \text{decomp} = \frac{C_{diss}}{C_{init}} \times 100$$

(EX) Addn of material to system at Equilibrium

¿H2 and I2 are mixed at 229 °C in a sealed 1.00 L container. Equilibrium is established at [H2] = 0.080 M, [I2] = 0.060 M, and [HI] = 0.490 M. If 0.300 mol of HI is then added, what concentrations are present when equilibrium is re-established?

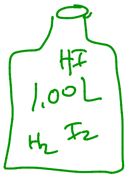
problem which ilustrates a number of the equilibrium concepts

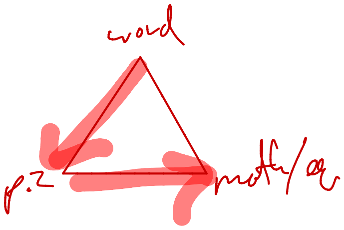
(source: W9, Ex17-11)

	R	$H_2(g)$ +	$I_2(g)$ $\rightleftharpoons$	$2\,HI(g)$
(ORIG)	I	0.080	0.060	0.490
	C	0	0	+0.300
("OOE")	~~E~~ I	$\boxed{0.080}$	$\boxed{0.060}$	$\boxed{0.790}$
	C			
	E			

$K_c = \boxed{}$

T = 229°C
(K)





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NOTE: The following is a single question that illustrates many of the equilibrium concepts.
(source: W9, Ex 17-11, p682)

	$H_2(g)$ +	$I_2(g)$ ⇌	$2\,HI(g)$	R
mol >	0.080	0.060	0.490	E
	∅	∅	+ 0.300	
	0.080	0.060	0.790	I
	+x	+x	−2x	C
	(0.080+x)	(0.060+x)	(0.790−2x)	E

$$K = \frac{(0.049)^2}{(0.080)(0.060)} = \boxed{50}$$

⇓

$$K = \underbrace{\frac{(0.790-2x)}{(0.080+x)(0.060+x)}}_{1\text{ eq }/\text{ 2 unks}} = \boxed{50}$$

!!!

thus, must calc K from other data

$$46x^2 + 102x - 0.38 = 0$$

↓ QUAD

$x = 0.032$ or $-\cancel{0.25}$

X cannot be less than ∅

$[H_2] = 0.112$ M (more than orig amt 0.080)

$[I_2] = 0.092$ M (more than orig amt 0.060)

$[HI] = 0.726$ M (less than "orig + stress" amt 0.790)

Cheat Sheet Notes

$$\begin{tikzcd}
 & K_p = K_c (RT)^{\Delta n} & \\
K_{c,new} = K_c^{\,n} & \boxed{K_c} \arrow[u] \arrow[l] \arrow[d] & \\
 & K_{c,rev} = 1/K_c &
\end{tikzcd}$$

Why "K" must change
w/ changing input...

K		[PROD, REACT]		INPUT
10	=	$\boxed{3}$	+	7
14	=	$\boxed{3}$	+	11
5	=	$\boxed{3}$	+	2

...must chg K so true [PROD, REACT] is maintained

If INPUT is changed...

van Hoff

$$\ln \frac{K'}{K} = \frac{\Delta H^\circ}{R}\left(\frac{1}{T} - \frac{1}{T'}\right)$$

cc

$$\ln \frac{k'}{k} = \frac{E_a}{R}\left(\frac{1}{T} - \frac{1}{T'}\right)$$

$$\ln \frac{P'}{P} = \frac{\Delta H_v}{R}\left(\frac{1}{T} - \frac{1}{T'}\right)$$

$$\left.\begin{array}{l} P_T = \sum P_i \\ P_i = \chi_i P_T \\ PV = nRT \end{array}\right\} \text{gas only}$$

$$\left.\begin{array}{l} \Delta P = \chi_{\text{SOLUTE}} \, P^\circ_{\text{SOLVENT}} \\ P_i = \chi_{(\text{in liq})} \, P_i^\circ \\ C_i = k P_i \end{array}\right\} \text{gas/liq barrier}$$