3/8/2019

What about solids + liquids?

K_c, molar concs (ignore 'em!)

K_p, gas pressures (atm, ignore units).

$$(g), (aq). \quad // \quad \underbrace{(s), (l)}_{\text{effective conc/pressure} = 1}$$

ex: $2CO(g) \rightleftharpoons CO_2(g) + C(s)$

$$K_c = \frac{[CO_2]\cancel{[C]}^{1}}{[CO]^2} = \frac{[CO_2]}{[CO]^2}$$

* $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$

$$K_p = \frac{\cancel{(P_{CaO})}^{1}(P_{CO_2})}{\cancel{(P_{CaCO_3})}^{1}} = P_{CO_2}$$

reverse

$CaO(s) + CO_2(g) \rightleftharpoons CaCO_3(s).$

$$K_p = \frac{\cancel{(P_{CaCO_3})}^{1}}{\cancel{(P_{CaO})}^{1}(P_{CO_2})} = \frac{1}{P_{CO_2}}$$

$$HF(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + F^-(aq)$$

$$K_c = \frac{[H_3O^+][F^-]}{[HF]\cancel{[H_2O]}^{1}} = \frac{[H_3O^+][F^-]}{[HF]}$$

Calculating K_c from eq^m concs

ex: What's K_c for $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$

if $[H_2] = [I_2] = 0.11M$ and $[HI] = 0.78M$ @ EQ^m

$$K_c = \frac{[HI]^2}{[H_2][I_2]} = \frac{(0.78)^2}{(0.11)(0.11)} = 50. \text{ or } 5.0 \times 10^1$$

What happens if we know INITIAL concs, but only know some of the final eq^m concs? Can we find K_c? (YES!)

- use stoichiometry + (R)ICE charts.

(Rxn) init / change / eq^m

Note: no matter the initial concentrations of H_2, I_2, or HI — at equilibrium the ratio of HI squared to $H_2 \times I_2$ is always a constant (the equilibrium constant)

Typo! The squared should be outside of the [ ]

TABLE 15.1 Initial and Equilibrium Concentrations for the Reaction $H_2(g) + I_2(g) \rightleftharpoons 2\,HI(g)$ at 445 °C

Initial Concentrations			Equilibrium Concentrations			Equilibrium Constant
$[H_2]$	$[I_2]$	$[HI]$	$[H_2]$	$[I_2]$	$[HI]$	$K_c = \frac{[HI^2]}{[H_2][I_2]}$
0.50	0.50	0.0	0.11	0.11	0.78	$\frac{(0.78)^2}{(0.11)(0.11)} = 50$
0.0	0.0	0.50	0.055	0.055	0.39	$\frac{(0.39)^2}{(0.055)(0.055)} = 50$
0.50	0.50	0.50	0.165	0.165	1.17	$\frac{(1.17)^2}{(0.165)(0.165)} = 50$
1.0	0.50	0.0	0.53	0.033	0.934	$\frac{(0.934)^2}{(0.53)(0.033)} = 50$
0.50	1.0	0.0	0.033	0.53	0.934	$\frac{(0.934)^2}{(0.033)(0.53)} = 50$

ex: $A(g) \rightleftharpoons 2B(g)$

$[A]_0 = 1.00M$
$[B]_0 = 0.00M$

$[A]_{eq} = 0.75M$
$K_c = ?$

Rxn:	$A(g)$	$\rightleftharpoons$ $2B(g)$
Initial	1.00	0
Change	-0.25	$\boxed{+2\times0.25}$ +0.50
Equilibrium	0.75	0.50

$$K_c = \frac{[B]^2}{[A]_{eq}}$$

$$= \frac{(0.50^2)}{(0.75)} = 0.33$$

p 690
Ex: 15.5, 15.6 Do!

Predicting direction of change

let's say we have $A(g) \rightleftharpoons B(g)$
$K_c = 1.45$ (25°c)
$[A]_0 = [B]_0 = 1.00M$

Q: Which way will rxn proceed?
$A \longrightarrow B$
$A \longleftarrow B$

We need to calculate the
Reaction Quotient, Q

$aA + bB \rightleftharpoons cC + dD$

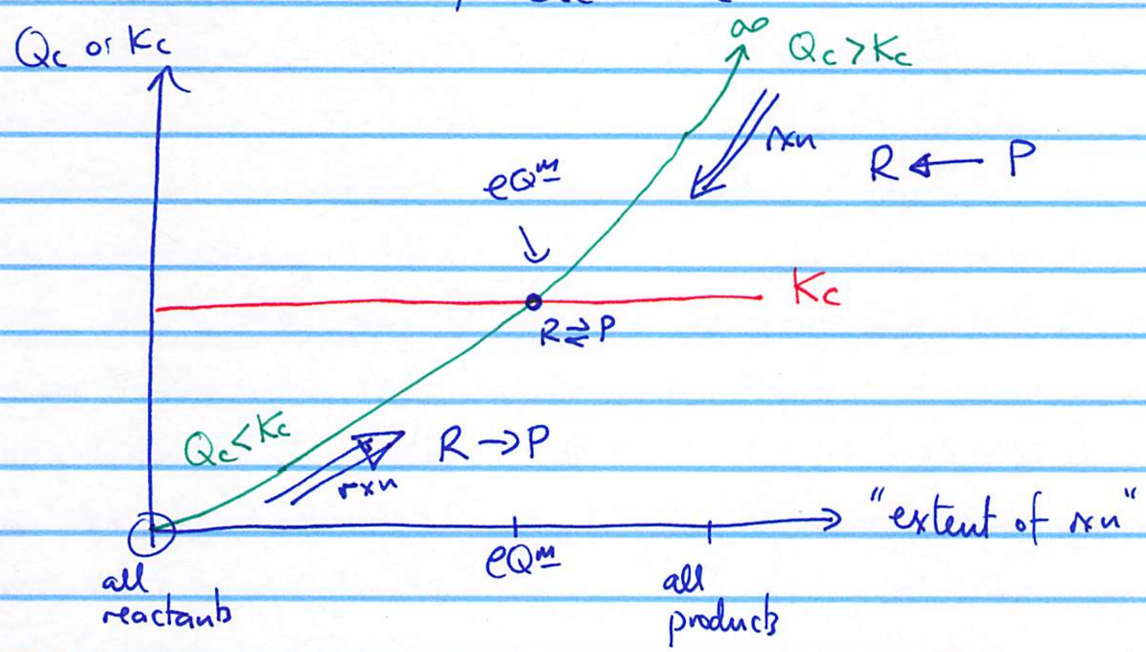
$$Q_c = \frac{[C]^c [D]^d}{[A]^a [B]^b} \longleftarrow \text{use current concs! (init)}$$

reaction quotient

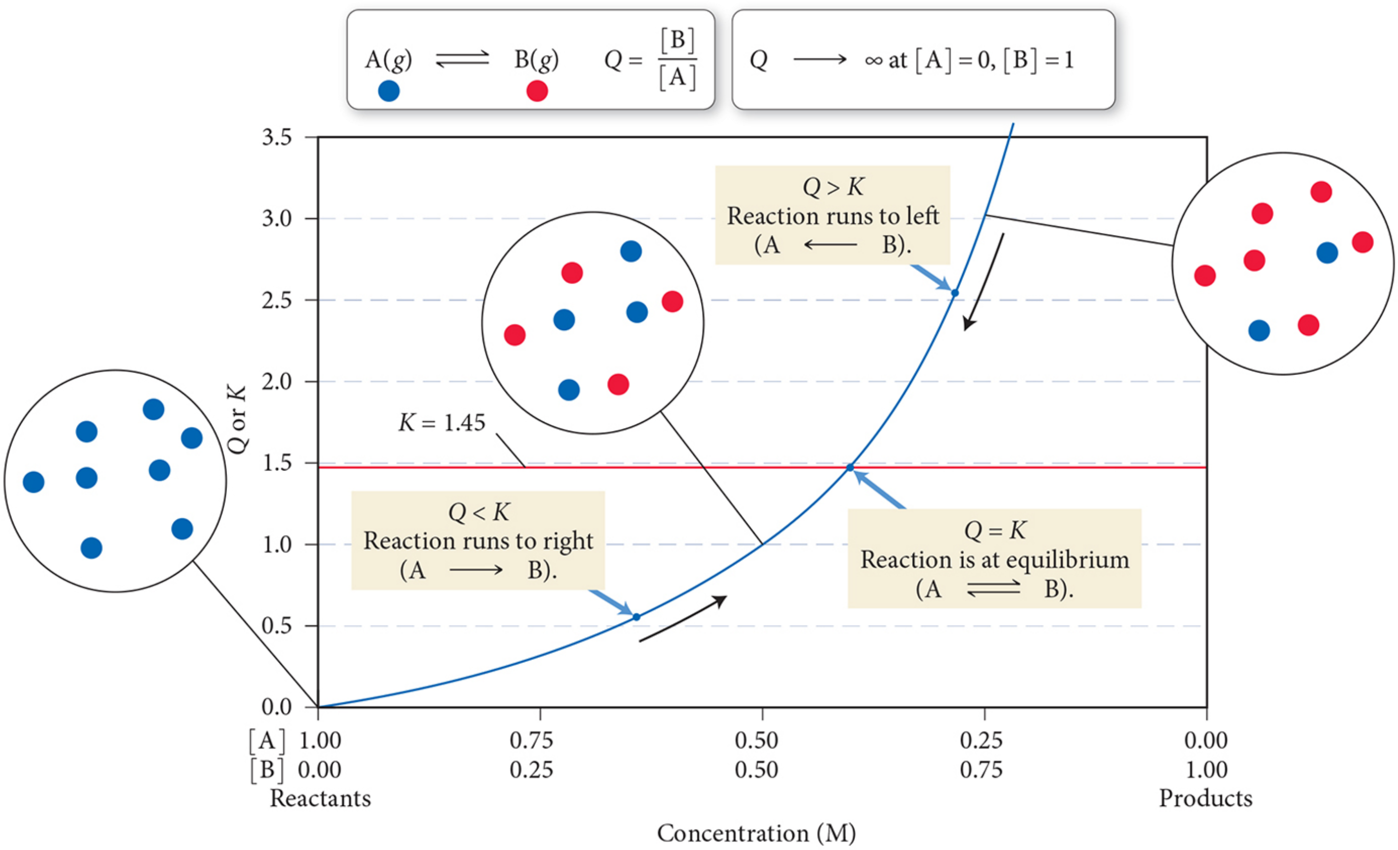
ex: @ beginning of rxn: $[C] = [D] = 0$, so $Q_c = 0$

@ end of rxn: $[A] = [B] = 0$, so $Q_c = \infty$

@ eQm, $Q_c = K_c$



Q, K, and the Direction of a Reaction



our last situation: $A(g) \rightleftharpoons B(g)$; $K_c = 1.45$

$[A]_0 = [B]_0 = 1.00M$

$Q_c = ?$

which way will it shift?

$Q_c \neq K_c$ (not @ eqm)

$$Q_c = \frac{[B]}{[A]} = \frac{1.00}{1.00} = 1.00$$

$Q_c < K_c$: Shift to RHS

↖ instantaneous/init values.

$\left(\frac{[B]\uparrow}{[A]\downarrow}, Q = K \right)$

$A \rightarrow B$

Yes, in theory you can memorize the direction of shift as follows:

- Q < K, shift to RHS
- Q = K, no shift (@ eqm)
- Q > K, shift to LHS

But if you understand that in order for the reaction to come to equilibrium the value of Q must increase or decrease to equalize K — and that only happens by causing more A to be converted to B or vice versa — then you will never forget on an exam!

Understanding takes a lot more time, energy, and effort than memorizing. Unfortunately that's the price for not forgetting. 😀