Sample problem

K_P for the following reaction is 0.16 at 25 ˚C. What is the value of K_c?

$$2\ NOBr\ (g) \Leftrightarrow 2\ NO\ (g) + Br_2\ (g)$$

- $K_P = K_C(RT)^{\Delta n}$
- $0.16 = K_c(0.08206\ L\ atm\ mol^{-1}\ K^{-1} \times 298.15\ K)^{3-2}$
- **$K_c = 6.5 \times 10^{-3}$**

Sample problem

$K_c = 5.6 \times 10^{-12}$ at 500 K for the dissociation of iodine molecules to iodine atoms.

$$I_2\,(g) \Leftrightarrow 2\,I\,(g)$$

A mixture has $[I_2] = 0.020$ M and $[I] = 2.0 \times 10^{-8}$ M. Is the reaction at equilibrium at 500 K? If not, which way must the reaction proceed to reach equilibrium?

- $Q = \frac{[I]^2}{[I_2]} = \frac{(2.0\times10^{-8})^2}{0.020} = 2.0\times10^{-14}$
- $Q < K_c \rightarrow$ too many reactant particles
- Reaction will shift to the **right** (to make more product)

Sample problem

In aqueous solution Fe^{3+} ions react with I^- ions to give Fe^{2+} ions and I_3^- ions. Suppose the initial concentration of Fe^{3+} is 0.200 M, the initial I^- concentration is 0.300 M, and the equilibrium concentration of I_3^- ions is 0.0866 M. What is the value of K_C?

	2 Fe^{3+}	+ 3 I^-	⇔	2 Fe^{2+}	+ I_3^-
I	0.200 M	0.300 M		0	0
C	– 2x	– 3x		+ 2x	+ x
E	0.200 – 2x	0.300 – 3x		2x	x = 0.0866 M

$$K_c = \frac{[Fe^{2+}]^2[I_3^-]}{[Fe^{3+}]^2[I^-]^3} = \frac{(2 \times 0.0866)^2(0.0866)}{(0.200 - 2 \times 0.0866)^2(0.300 - 3 \times 0.0866)^3} = \mathbf{5.6 \times 10^4}$$

Sample problem

The equilibrium constant K_c (=55.64) for the reaction H_2 (g) + I_2 (g) ⇔ 2 HI (g) has been determined at 425 ˚C. If 0.130 mol each of H_2 and I_2 is placed in a 25.0-L flask at 425 ˚C, what are the concentrations of H_2, I_2 and HI when equilibrium has been achieved?

	H_2	+ I_2	⇔	2 HI
I	5.20×10^{-3} M	5.20×10^{-3} M		0
C	− x	− x		+ 2x
E	5.20×10^{-3} − x	5.20×10^{-3} − x		2x

[HI] = 8.20×10^{-3} M

$[H_2]$= $[I_2]$ = 1.10×10^{-3} M

$$K_c = \frac{[HI]^2}{[H_2]\,[I_2]} = \frac{(2x)^2}{(5.20\times10^{-3}-x)(5.20\times10^{-3}-x)} = \left(\frac{2x}{5.20\times10^{-3}-x}\right)^2 = 55.64$$

$$\frac{2x}{5.20\times10^{-3}-x} = 7.459 \rightarrow 2x = 0.0388 - 7.459x \rightarrow 9.459x = 0.0388 \rightarrow \boldsymbol{x = 4.10\times10^{-3}}$$

Equilibrium calculations: quadratic expression involving approximation

The reaction N_2 (g) + O_2 (g) ⇔ 2 NO (g) has K_c = 1.0 x 10^{-5} at 1500 K. Suppose a sample of air has [N_2] = 0.080 M and [O_2] = 0.020 M before any reaction occurs. Calculate the equilibrium concentration of reactants and products after the mixture has been heated to 1500 K.

	N_2	+ O_2	⇔	2 NO
I	0.080 M	0.020 M		0
C	– x	– x		+ 2x
E	0.080 – x	0.020 – x		2x

$$K_c = \frac{[NO]^2}{[N_2]\,[O_2]} = \frac{(2x)^2}{(0.080 - x)(0.020 - x)} = 1.0 \times 10^{-5}$$

$$4x^2 - (1.0 \times 10^{-5})x - 1.6 \times 10^{-8} = 0$$ Quadratic expression

Equilibrium calculations: quadratic expression involving approximation

Use approximation to update the table and make your life easier

$[N_2]$ and $[O_2]$ don't significantly change during the reaction.

	N_2	$+ O_2$	⇔	2 NO
I	0.080 M	0.020 M		0
C	~ 0	– x		+ 2x
E	~0.080 M	~0.020 M		2x

$$K_c = \frac{[NO]^2}{[N_2]\,[O_2]} = \frac{(2x)^2}{(0.080)(0.020)} = 1.0 \times 10^{-5}$$

[NO] = 1.3 x 10⁻⁴ M

$$x^2 = 4.0\times10^{-9} \rightarrow \boldsymbol{x = 6.3\times10^{-5}}$$

Sample problem

Calculate K for the reaction

$$SnO_2\ (s) + 2\ CO\ (g) \Leftrightarrow Sn\ (s) + 2\ CO_2\ (g)$$

given the following information:

- $SnO_2\ (s) + 2\ H_2\ (g) \Leftrightarrow Sn\ (s) + 2\ H_2O\ (g)$ K = 8.12
- $H_2\ (g) + CO_2\ (g) \Leftrightarrow H_2O\ (g) + CO\ (g)$ K = 0.771
- $\mathbf{SnO_2\ (s)} + \cancel{2\ H_2\ (g)} \Leftrightarrow \mathbf{Sn\ (s)} + \cancel{2\ H_2O\ (g)}$ $K_1 = 8.12$
- $\cancel{2H_2O\ (g)} + \mathbf{2\ CO\ (g)} \Leftrightarrow \cancel{2\ H_2\ (g)} + \mathbf{2\ CO_2\ (g)}$ $K_2 = (\frac{1}{0.771})^2$
- $\boldsymbol{K_{net}} = K_1 K_2 = (8.12)(\frac{1}{0.771})^2 = \mathbf{13.7}$

Sample problem

N_2O_3 decomposes to NO and NO_2 is an endothermic process (ΔH° = 40.5 kJ)

$$N_2O_3\,(g) \Leftrightarrow NO\,(g) + NO_2\,(g)$$

Predict the effect of the following changes on the equilibrium position.

- Adding more N_2O_3 (g) $\rightarrow$
- Adding more NO_2 (g) $\leftarrow$
- Increasing the volume of the reaction flask $\rightarrow$
  - System feels lower P (because V increased)
  - System needs to increase P by shifting to the side with more moles of gas
  - Reaction will shift to right (2 moles of gas on the right vs. 1 mole of gas on the left)
- Lowering the temperature $\leftarrow$
  - Reaction is endothermic so heat is a reactant
  - Lower T = less heat = less reactant