At equilibrium, a 1.0 liter container was found to contain 0.20 moles of A, 0.20 moles of B, 0.40 moles of C and 0.40 mole of D. If 0.10 moles of A and 0.10 moles of B are added to this system, what will be the new equilibrium concentration of A?

$$A_{(g)} + B_{(g)} \leftrightharpoons C_{(g)} + D_{(g)}$$

$K_c = \frac{[C][D]}{[A][B]}$

$K_c = \frac{(.40)^2}{(.20)^2} = 4$ equilibrium constant

w/ added mol of A & B

	A	B	C	D
I	.30	.30	.40	.40
C	−x	−x	+x	+x
E	.30−x	.30−x	.40+x	.40+x

final equil. moles

$K_c = \frac{(.40+x)^2}{(.30-x)^2} = 4$ $\quad \frac{.40+x}{.30-x} = 2$

$.40 + x = .60 - 2x$
$.40 + 3x = .60$
$3x = .20$
$x = 0.067$

at new equilibrium:
$[A] = .30 - x = .30 - 0.067 = \boxed{0.23\ M}$

A. 0.37 M

B. 0.47 M

C. 0.87 M

D. 0.23 M

E. 0.15 M

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$Ni(CO)_4(g) \leftrightharpoons Ni(s) + 4\ CO(g)$

Adding nickel to this reaction will cause the equilibrium to

Ni is a solid. Solids do not appear in an equilibrium expression & do not affect the equilibrium as long as some solid was already present.

a. shift toward products.

b. shift toward reactants.

c. remain unchanged.

d. change based on the amount added.

★ solids & liquids do not affect equilibrium

2

For the reaction Le Chatelier & volume (pressure)

$$2\ P_{2\,(g)} \leftrightharpoons P_{4\,(g)}$$

What happens if an inert gas is added to the container increasing the total pressure at constant temperature and volume?

A. The reaction shifts right (toward products). Constant T & V

B. The reaction shifts left (toward reactants).

C. Nothing.

Because the volume does not change, the concentrations do not change.

↳ therefore, has no effect on equilibrium

3

For the reaction

$$2\ P_{2\,(g)} \leftrightharpoons P_{4\,(g)}$$

Constant P & T

What happens if an inert gas is added to the container but the total pressure and temperature are kept constant?

A. The reaction shifts right (toward products).

B. The reaction shifts left (toward reactants).

C. Nothing.

The volume changes, changing the concentrations. The molarity decreases which also decreases the pressure. In order to re-establish equilibrium, the rxn shifts to the side with a greater number of moles of gas.

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adding an inert gas at constant pressure will increase total volume. therefore, the total # of moles per unit volume will decrease (molarity). the equilibrium shifts toward the direction where there is an increase in # of moles of gas.

decreases molarity

For which of the following systems at equilibrium will doubling the volume cause a shift to the right? (decreases the P)

The rxn will shift to the side with ↑ mol of gas

A. $H_{2(g)} + Cl_{2(g)} \leftrightharpoons 2\ HCl_{(g)}$ No change, 2 mol of gas on each side

B. $2\ CO_{(g)} + O_{2(g)} \leftrightharpoons 2\ CO_{2(g)}$ shifts left

C. $N_{2g)} + 3H_{2g)} \leftrightharpoons 2\ NH_{3(g)}$ shifts left

D. $PCl_{5(g)} \leftrightharpoons PCl_{3(g)} + Cl_{2g)}$ shifts right

1 mol 2 mol

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$N_2(g) + 3\ H_2(g) \rightleftarrows 2\ NH_3(g) \quad \Delta G° = -32.7\ kJ/mol$

Calculate ΔG for the reaction under the following conditions:

P N_2 = 2.00 atm
P H_2 = 7.00 atm
P NH_3 = 0.021 atm
T = 100 °C

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

A. Calculate Q.

reaction quotient

$$\frac{(.021)^2}{(2)(7)^3} = 6.43 \times 10^{-7}$$

B. Is Q_____ than/to K?

a) >
b) <
c) =

Since $\Delta G°$ is negative, K must be > 1

C. Calculate ΔG. $\Delta G = \Delta G° + RT \ln Q_P$

$$\Delta G = -32.7\frac{kJ}{mol} + \left(8.314 \times 10^{-3}\frac{kJ}{mol \cdot K} \times 373\ K\right) \ln 6.43 \times 10^{-7} \qquad \Delta G = -76.91\frac{kJ}{mol}$$

D. Does the reaction proceed spontaneously to the right?

a) yes
b) no

negative ΔG, spontaneous in the forward direction

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6

What is K if $\Delta G° = -19.0$ kJ for a reaction at 25 °C?

convert to K

$\Delta G° = -RT \ln K$

$\ln K = \frac{\Delta G°}{-RT}$

$\ln K = \frac{-19.0 \text{ kJ}}{-8.314 \times 10^{-3} \cdot 298} = 7.668 \rightarrow e^{7.668} =$

$K = 2.14 \times 10^{3}$

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For the following reaction find K_p at 25 °C and indicate whether K_p should increase or decrease as the temperature rises.

$NH_4HS(s) \rightleftharpoons H_2S(g) + NH_3(g)$ $\Delta H° = 83.47$ kJ and $\Delta G° = 17.5$ kJ at 25 °C.

$\Delta G° = -RT \ln K_p$

$\ln K_p = \frac{\Delta G°}{-RT}$

$\ln K_p = \frac{17.5 \text{ kJ}}{-8.314 \times 10^{-3} \frac{\text{kJ}}{\text{mol} \cdot \text{K}} \times 298 \text{K}} = -7.063$

$K_p = 8.56 \times 10^{-4}$; Since the rxn is endothermic, the addition of heat will shift the rxn to the right & K will increase.

↳ since there will be an increase in products

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According to Le Châtelier's principle, if the volume of the vessel containing the equilibrium system shown below is decreased, there will be an increase in the concentration of H_2 and a decrease in the concentration of H.

$H_2(g) \rightleftharpoons 2\ H(g)$

$\downarrow V, \uparrow P$

the rxn shifts to the left
(Fewer moles of gas)

if volume decreases,
then molarity increases
so must shift to
direction where there are
less moles in order to
decrease molarity

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If additional SCN^- is added to the equilibrium system shown below, Le Châtelier's principle predicts a net reaction from left to right, causing the red color to become darker.

$Fe^{3+}(aq) + SCN^-(aq) \rightleftharpoons FeSCN^{2+}(aq)$

yellow colorless red

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$K > Q$ rxn proceeds forward
$K < Q$ rxn proceeds in reverse
$K = Q$ equilibrium

At an elevated temperature, $K_p = 0.19$ for the reaction

$$2\,NOCl(g) \rightleftharpoons 2\,NO(g) + Cl_2(g)$$

If the initial partial pressures of NOCl, NO, and Cl_2 are 0.50 atm, 0.25 atm, and 0.45 atm, respectively, a net ________ reaction must occur in order to achieve equilibrium.

$$Q_p = \frac{P_{NO}^2 \times P_{Cl_2}}{P_{NOCl}^2}$$

A. Forward
B. Reverse
C. Not enough information is given
D. The reaction is at equilibrium

$$Q_p = \frac{(.25)^2(.45)}{(.50)^2} = 0.11$$

$K_p > Q_p$ rxn will shift right

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For the reaction shown below the value of K_p is greater than the value of K_c, because Δn = 1.

$$N_2O_4(g) \rightleftharpoons 2\,NO_2(g)$$

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

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Which of the following statements about a catalyst is true?

→ catalyst does not affect equilibrium

A. A catalyst decreases the position of the equilibrium in a reaction.
B. A catalyst increases the pressure of a reaction.
C. A catalyst is consumed in a chemical reaction.
D. A catalyst provides a lower energy pathway for a reaction.

(catalyst affects rate of rxn)

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→ spontaneous (large Keq)

The hexaammine cobalt(III) ion is very unstable in acidic aqueous solution:

$$[Co(NH_3)_6]^{3+}(aq) + 6\ H_3O^+(aq) \leftrightharpoons [Co(H_2O)_6]^{4+}(aq) + 6\ NH_4^+(aq)$$

However, solutions of hexaammine cobalt(III) can be stored in acidic solution for months without noticeable decomposition. Which statement below about the equilibrium constant and the activation energy for the reaction is true?

large activation varrier ←

A. $K_{eq} < 10^3$ and E_a is very small.
B. $K_{eq} > 10^3$ and E_a is very small.
C. $K_{eq} < 10^3$ and E_a is very large.
D. $K_{eq} > 10^3$ and E_a is very large.

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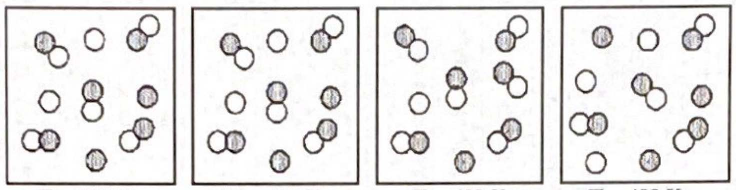
Picture (1) represents the equilibrium mixture for the reaction A + B ⇌ 2 AB at 298 K.



$K = \frac{[AB]^2}{[A][B]}$

If this reaction is exothermic, which picture (2)-(4) represents the equilibrium mixture at 400 K? the rxn will shift left with ↑T

A. picture (2) [A][B] & [AB] no change
B. picture (3) [A][B] ↓, [AB] ↑
C. picture (4) [A][B] ↑, [AB] ↓
D. none of these

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exothermic - reactants favored
endothermic - products favored