

16 • Chemical Equilibrium**STATION 1 • MASS ACTION EXPRESSIONS**Write the mass action expression for the equilibrium: $\text{Fe}_3\text{O}_4(\text{s}) + 4 \text{H}_2(\text{g}) \rightleftharpoons 3 \text{Fe}(\text{s}) + 4 \text{H}_2\text{O}(\text{g})$

$$K_c = \frac{[\text{H}_2\text{O}]^4}{[\text{H}_2]^4}$$

Write the mass action expression for the equilibrium: $\text{Ca}(\text{OH})_2(\text{s}) \rightleftharpoons \text{Ca}^{2+}(\text{aq}) + 2 \text{OH}^-(\text{aq})$

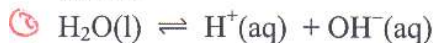
$$K_c = [\text{Ca}^{2+}][\text{OH}^-]^2$$

The equilibrium constant, $K_c = 7.9 \times 10^{-6}$, is this equilibrium system reactant or product-favored. REACTANT-

$$K_c \ll 1 \quad \frac{[\text{PRODUCTS}]}{[\text{REACTANTS}]}$$

FAVORED**16 • Chemical Equilibrium****STATION 2 • MANIPULATING K**

Given:

FLIP

$K_c = 1 \times 10^{-14}$

INVERSE $\frac{1}{1 \times 10^{-14}}$

$K_c = 4.0 \times 10^{-10}$

Calculate K_c for this reaction:

$K_c = ???$



$K_1 = \frac{1}{1 \times 10^{-14}}$

$K_2 = 4.0 \times 10^{-10}$

$K_c = K_1 \times K_2 = \frac{4.0 \times 10^{-10}}{1.0 \times 10^{-14}}$

$K_c >> 1 \quad = \boxed{4.0 \times 10^4}$

This reaction is PRODUCT-favored.

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STATION 3 • K_p & K_c

Consider the equilibrium: $2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$.

At 100°C , the equilibrium concentrations for this system are:

$$[\text{NO}] = 0.52 \text{ M}; \quad [\text{O}_2] = 0.24 \text{ M}; \quad [\text{NO}_2] = 0.18 \text{ M}$$

Write the expression for K_c and calculate its value at this temperature?

$$K_c = \frac{[\text{NO}_2]^2}{[\text{NO}]^2 [\text{O}_2]} = \frac{(0.18)^2}{(0.52)^2 (0.24)} = .49924 \approx \boxed{.50 \text{ M}^{-1}}$$

What is Δn for this system? -1 $2 - 3 = -1$ (moles of product gas) - (moles of reactant gas)

Write the expression for K_p and calculate its value at this temperature.

$$K_p = \frac{(P_{\text{NO}_2})^2}{(P_{\text{NO}})^2 (P_{\text{O}_2})} = K_c (RT)^{\Delta n} = (.50)(.0821 \times 373)^{-1} = \boxed{0.016} \text{ atm}^{-1}$$

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STATION 4 • LE CHÂTELIER'S PRINCIPLE

Consider the gaseous equilibrium: $2\text{CCl}_4(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{COCl}_2(\text{g}) + 2\text{Cl}_2(\text{g})$ $\Delta H = +35 \text{ kJ}$

Predict the effect each change would have on the concentrations of the each substance.

Add CCl_4	$\uparrow$	$\downarrow$	$\rightarrow$	$\uparrow$	$\uparrow$
Remove Cl_2	$\downarrow$	$\downarrow$	$\rightarrow$	$\uparrow$	$\downarrow$
Add COCl_2	$\uparrow$	$\uparrow$	$\rightarrow$	$\uparrow$	$\downarrow$
Increase temperature	$\downarrow$	$\downarrow$	$\rightarrow$	$\uparrow$	$\uparrow$
Reduce container volume	$\uparrow$	$\uparrow$	$\rightarrow$	$\downarrow$	$\downarrow$
Add a catalyst	—	—	—	—	—
Remove O_2	$\uparrow$	$\downarrow$	$\rightarrow$	$\downarrow$	$\downarrow$
Add He to increase pressure	—	—	—	—	—

3 mol gas $\rightleftharpoons$ 4 mol gas

A different equilibrium shifts toward the reactants when the temperature is increased. From this observation, you know that the reaction is Exothermic (exothermic / endothermic).

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STATION 5 • ICE BOX PROBLEM

Consider the equilibrium: $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})$

If 0.200 mol $\text{SO}_3(\text{g})$ is placed in a 0.500 Liter container, it is found that 0.050 mole of $\text{O}_2(\text{g})$ is in the container at equilibrium. Fill in the ICE box and determine the K_c for this reaction.

	2SO_2	O_2	2SO_3
Initial	0 M	0 M	$\frac{0.200 \text{ mol}}{0.500 \text{ L}} = 0.400 \text{ M}$
Change	+0.200 M	+0.100 M	-0.200 M
Equilibrium	0.200 M	$\frac{0.050 \text{ mol}}{0.500 \text{ L}} = 0.100 \text{ M}$	0.200 M

$$K_c = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2 [\text{O}_2]} = \frac{(0.200)^2}{(0.200)^2 (0.100)} = \boxed{10.0}$$

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STATION 6 • ANOTHER ICE BOX PROBLEM

At 985°C, the equilibrium constant for the reaction, $\text{H}_2(\text{g}) + \text{CO}_2(\text{g}) \rightleftharpoons \text{H}_2\text{O}(\text{g}) + \text{CO}(\text{g})$, is 1.63.

If 2.00 moles each of $\text{H}_2(\text{g})$ and $\text{CO}_2(\text{g})$ are placed in a 1.00-Liter container and allowed to come to equilibrium, determine the equilibrium concentrations of the four chemicals.

	H_2	CO_2	H_2O	CO
Initial	2.00 M	2.00 M	0 M	0 M
Change	-x	-x	+x	+x
Equilibrium	2.00 - x	2.00 - x	x	x

$$\begin{array}{cccc} .880 \text{ M} & .880 \text{ M} & 1.12 \text{ M} & 1.12 \text{ M} \end{array}$$

$$K_c = \frac{[\text{H}_2\text{O}][\text{CO}]}{[\text{H}_2][\text{CO}_2]} = 1.63 = \frac{x^2}{(2.00-x)^2}$$

SQUARE ROOT BOTH SIDES

$$\begin{aligned} \sqrt{1.63} &= 1.2767 = \frac{x}{2.00-x} \\ x &= (1.2767)(2.00-x) \\ x &= 2.55 - 1.2767x \end{aligned}$$

$$\frac{2.2767x}{2.2767} = \frac{2.55}{2.2767}$$

$$x = 1.12 \text{ M}$$

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STATION 7 • TEST QUOTIENT, Q

Consider the equilibrium: $2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$. $K_c = 0.499$

The system is set up with the following concentrations:

$$[\text{NO}] = 0.50 \text{ M}; \quad [\text{O}_2] = 0.25 \text{ M}; \quad [\text{NO}_2] = 0.25 \text{ M}$$

This reaction will SHIFT LEFT (shift right, shift left, remain unchanged).

Justify your prediction.

$$Q = \frac{[\text{NO}_2]^2}{[\text{NO}]^2 [\text{O}_2]} = \frac{(0.25 \text{ M})^2}{(0.50 \text{ M})^2 (0.25 \text{ M})} = 1.0$$

$$Q > K_c$$

TOO MANY
PRODUCTS

[PRODUCTS]
[REACTANTS]

SHIFT TOWARD REACTANTS TO REACH
EQUILIBRIUM.

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STATION 8 • NCHO PROBLEMS

33. Mercury(II) oxide, HgO , is decomposed upon heating according to this equation.



What is the equilibrium expression for this process?

- (A) $K = \frac{[\text{Hg}]^2 [\text{O}_2]}{[\text{HgO}]^2}$ (B) $K = \frac{[\text{Hg}][\text{O}_2]}{[\text{HgO}]}$
(C) $K = [\text{Hg}][\text{O}_2]$ (D) $K = [\text{O}_2]$

34. Consider this reaction.



What conditions of temperature and pressure will produce the highest yield of NOCl at equilibrium?

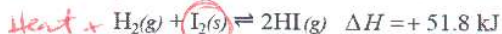
	T	P
(A)	high	high
(B)	high	low
(C)	low	high
(D)	low	low

means
"SHIFT RIGHT"

3 mols
gas $\rightleftharpoons$ 2 mols
gas

high P

Questions 32 and 33 should both be answered with reference to this system. $1 \text{ mole gas} \rightleftharpoons 2 \text{ mols gas}$



32. Which would increase the equilibrium quantity of $\text{HI}(\text{g})$? Assume the system has reached equilibrium with all three components present.

I. increasing pressure

II. increasing temperature

(A) I only

(B) II only

(C) Both I and II

(D) Neither I nor II

33. What is the equilibrium constant expression for this system?

(A) $K = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$

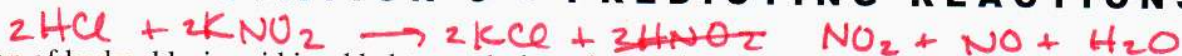
(B) $K = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2}$

(C) $K = \frac{2[\text{HI}]}{[\text{H}_2][\text{I}_2]}$

(D) $K = \frac{[\text{HI}]^2}{[\text{H}_2]}$

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STATION 9 • PREDICTING REACTIONS



A solution of hydrochloric acid is added to a solution of potassium nitrite.

(i) Balanced equation:



(ii) What are the oxidation states of the N atom before and after the reaction?

O = -2 in most
of its
compounds

NO_2^-	NO	NO_2
$\text{N} = +3$	$\text{N} = +2$	$\text{N} = +4$

Note: Begin with: $\text{HCl} + \text{KNO}_2 \rightarrow \text{KCl} + \text{HNO}_2$

Recognize that HNO_2 decomposes into two gases (NO & NO_2)

To show the correct decomposition, you need 2 moles of HNO_2



Molecular Equation:



Ionic Equation:



Net Equation is shown above.