South Pasadena • AP Chemistry

Name KEY
Period ___ Date ___/___/___

16 • Chemical Equilibrium

STATION 1 • MASS ACTION EXPRESSIONS

Write the mass action expression for the equilibrium: $Fe_3O_4(s) + 4\ H_2(g) \rightleftharpoons 3\ Fe(s) + 4\ H_2O(g)$

$$K_c = \frac{[H_2O]^4}{[H_2]^4}$$

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

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

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

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

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STATION 2 • MANIPULATING K

Given:

① $H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$ FLIP $\quad K_c = 1 \times 10^{-14}$ INVERSE $\frac{1}{1 \times 10^{-14}}$

② $HCN(aq) \rightleftharpoons H^+(aq) + CN^-(aq)$ $\quad K_c = 4.0 \times 10^{-10}$

Calculate K_c for this reaction:

$HCN(aq) + OH^-(aq) \rightleftharpoons H_2O(l) + CN^-(aq)$ $\quad K_c = ???$

①' $\cancel{H^+} + OH^- \rightleftharpoons H_2O_{(l)}$ $\quad K_1' = \frac{1}{1 \times 10^{-14}}$

② $HCN \rightleftharpoons \cancel{H^+} + CN^-$ $\quad K_2 = 4.0 \times 10^{-10}$

$HCN + OH^- \rightleftharpoons H_2O_{(l)} + CN^-$ $\quad K_c = K_1' \times K_2 = \frac{4.0 \times 10^{-10}}{1.0 \times 10^{-14}}$

$= \boxed{4.0 \times 10^4}$

$K_c >> 1$

This reaction is PRODUCT -favored.

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STATION 3 • K_p & K_c

Consider the equilibrium: $2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$.

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

$[NO] = 0.52$ M; $[O_2] = 0.24$ M; $[NO_2] = 0.18$ M

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

$$K_c = \frac{[NO_2]^2}{[NO]^2[O_2]} = \frac{(.18)^2}{(.52)^2(.24)} = .49924 \approx \boxed{.50\ M^{-1}}$$

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.

$$K_p = \frac{(P_{NO_2})^2}{(P_{NO})^2(P_{O_2})} = K_c(RT)^{\Delta n} = (.50)(.0821 \times 373K)^{-1}$$

$$= \boxed{0.016}\ atm^{-1}$$

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STATION 4 • LE CHÂTELIER'S PRINCIPLE

Consider the gaseous equilibrium: Heat + $2CCl_4(g) + O_2(g) \rightleftharpoons 2COCl_2(g) + 2Cl_2(g)$ $\Delta H = +35$ kJ

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

3 mol gas ⇌ 4 mol gas

Add CCl_4	↑	↓	→	↑	↑
Remove Cl_2	↓	↓	→	↑	↓
Add $COCl_2$	↑	↑	←	↑	↓
Increase temperature	↓	↓	→	↑	↑
Reduce container volume	↑	↑	←	↓	↓
Add a catalyst	—	—		—	—
Remove O_2	↑	↓	←	↓	↓
Add He to increase pressure	—	—		—	—

A different equilibrium shifts toward the reactants ← when the temperature Heat 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: $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

If 0.200 mol $SO_3(g)$ is placed a 0.500 Liter container, it is found that 0.050 mole of $O_2(g)$ is in the container at equilibrium. Fill in the ICE box and determine the K_c for this reaction.

	2 SO_2	O_2	2 SO_3
Initial	0 M	0 M	$\frac{.200 \text{ mol}}{.500 L} = .400$ M
Change	+ .200 M	+ .100 M	− .200 M
Equilibrium	.200 M	$\frac{.050 \text{ mol}}{.500 L} = .100$ M	.200 M

$$K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]} = \frac{(.200)^2}{(.200)^2(.100)} = \boxed{10.0}$$

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STATION 6 • ANOTHER ICE BOX PROBLEM

At 985°C, the equilibrium constant for the reaction, $H_2(g) + CO_2(g) \rightleftarrows H_2O(g) + CO(g)$, is 1.63.

If 2.00 moles each of $H_2(g)$ and $CO_2(g)$ 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
	.880 M	.880 M	1.12 M	1.12 M

Answers

$$K_c = \frac{[H_2O][CO]}{[H_2][CO_2]} = 1.63 = \frac{x^2}{(2.00 - x)^2}$$

SQUARE ROOT BOTH SIDES

$$\sqrt{1.63} = 1.2767 = \frac{x}{2.00 - x}$$

$$x = (1.2767)(2.00 - x)$$

$$x = 2.55 - 1.2767x$$

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

$x = 1.12$ M

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STATION 7 • TEST QUOTIENT, Q

Consider the equilibrium: $2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$. $K_c = 0.499$

The system is set up with the following concentrations:

$[NO] = 0.50\ \underline{M}$; $[O_2] = 0.25\ \underline{M}$; $[NO_2] = 0.25\ \underline{M}$

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

Justify your prediction.

$$Q = \frac{[NO_2]^2}{[NO]^2[O_2]} = \frac{(.25\,M)^2}{(.50M)^2(.25M)} = 1.0$$

$Q > K_c$ TOO MANY PRODUCTS $\frac{[\text{PRODUCTS}]}{[\text{REACTANTS}]}$

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.

$$2HgO(s) \rightleftharpoons 2Hg(l) + O_2(g)$$

What is the equilibrium expression for this process?

(A) $K = \frac{[Hg]^2[O_2]}{[HgO]^2}$ **(B)** $K = \frac{[Hg][O_2]}{[HgO]}$

(C) $K = [Hg][O_2]$ **(D)** $\boxed{K = [O_2]}$

34. Consider this reaction.

$$2NO(g) + Cl_2(g) \rightleftharpoons 2NOCl(g) + \text{HEAT} \quad \Delta H = -78.38\text{ kJ}$$

What conditions of temperature and pressure will produce the highest yield of NOCl at equilibrium? means "SHIFT RIGHT"

3 moles gas ⇌ 2 moles gas

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

HIGH P

Questions **32** and **33** should both be answered with reference to this system. 1 mole gas ⇌ 2 mole gas

$$\text{Heat} + H_2(g) + I_2(s) \rightleftharpoons 2HI(g) \quad \Delta H = +51.8\text{ kJ}$$

SHIFT

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

I. increasing pressure **II.** increasing temperature

(A) I only **(B)** $\boxed{\text{II only}}$

(C) Both I and II **(D)** Neither I nor II

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

(A) $K = \frac{[HI]^2}{[H_2][I_2]}$ **(B)** $K = \frac{[H_2][I_2]}{[HI]^2}$

(C) $K = \frac{2[HI]}{[H_2][I_2]}$ **(D)** $\boxed{K = \frac{[HI]^2}{[H_2]}}$

STATION 9 • PREDICTING REACTIONS

$2HCl + 2KNO_2 \rightarrow 2KCl + \cancel{2HNO_2}$ $NO_2 + NO + H_2O$

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

(i) Balanced equation:

$2H^+ + 2NO_2^- \rightarrow NO + NO_2 + H_2O$

(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
N = +3	N=+2	N=+4

Note: Begin with: $HCl + KNO_2 \rightarrow KCl + HNO_2$
Recognize that HNO_2 decomposes into two gases (NO & NO_2)
To show the correct decomposition, you need 2 moles of HNO_2

$$2\,HNO_2 \rightarrow NO(g) + NO_2(g) + H_2O(l)$$

Molecular Equation:

$$2\,HCl + 2KNO_2 \rightarrow 2\,KCl + NO + NO_2 + H_2O$$

Ionic Equation:

$$2H^+ + 2\,Cl^- + 2\,K^+ + 2\,NO_2^- \rightarrow 2\,K^+ + 2\,Cl^- + NO + NO_2 + H_2O$$

Net Equation is shown above.