Name:_____ Date:_____

Practice AP Questions

1.

$$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)} \qquad \Delta H = -92.4 \text{ kJ}$$

When the reaction above took place at a temperature of 570 K, the following equilibrium concentrations were measured.

$[NH_3] = 0.20$ mol/L
$[N_2] = 0.50$ mol/L
$[H_2] = 0.20$ mol/L

A. Write the expression for K and calculate its value.

$$K = \frac{[NH_3]^2}{[N_2][H_2]^3} = \frac{[.20]^2}{[.50][.20]^3} = 10$$

B. What is the value of K_p for the reaction?

$$K_p = K(RT)^{\Delta n} \qquad K_p = 10\left[\left(.0821 \frac{L\,atm}{mol\,K}\right)(570K)\right]^{-2}$$

$$K_p = 4.7 \times 10^{-3}$$

C. Is this reaction exothermic or endothermic?

exothermic b/c sign is negative

D. Describe how the concentration of H_2 will be affected by each of the following changes to the system at equilibrium.

i. The temperature is increased.

T↑ favors a shift to the left so H_2 will increase

ii. The volume of the reaction chamber is increased.

increase in volume favors the side with more moles which is to the left so H_2 will increase

iii. N_2 gas is added to the reaction chamber.

N_2 is a reactant so the reaction will shift left and H_2 will decrease

2. 250.0 grams of solid copper (II) nitrate is placed in an empty 4.0 liter flask. Upon heating, the flask to 250°C, some of the solid decomposes into solid copper(II) oxide, gaseous nitrogen (IV) oxide and oxygen gas. At equilibrium, the pressure is measured and found to be 5.50 atmospheres.

A. Write the balanced equation for the reaction.

$$2Cu(NO_3)_{2(s)} \rightarrow 2CuO_{(s)} + 4NO_{2(g)} + O_{2(g)}$$

B. Calculate the number of moles of oxygen gas present in the flask at equilibrium.

$PV = nRT$

$(5.50\ atm)(4.0L) = n(0.0821 \frac{L\,atm}{mol\,K})(523K)$

$n = .51$ mol gas

$.51\ mol \times \frac{1\ mol\ Oxygen}{5\ mol\ total\ gas} = .10\ mol\ O_2$

C. Calculate the number of grams of solid copper (II) nitrate that remained in the flask at equilibrium.

$.10\ mol\ O_2 \times \frac{2\ mol\ Cu(NO_3)_2}{1\ mol\ O_2} \times \frac{187.57g\ Cu(NO_3)_2}{1\ mol\ Cu(NO_3)_2} = 38g\ Cu(NO_3)_2$

mass that remains

$250.0g - 38.0g = 212g\ Cu(NO_3)_2$

D. Write the equilibrium expression for K_p and calculate the value of the equilibrium constant.

$K_p = (P_{NO_2})^4 (P_{O_2})$

$\frac{4\ mol\ NO_2}{5\ total\ moles\ gas} \times 5.50\ atm = 4.40\ atm\ NO_2$

$5.50\ atm - 4.40\ atm\ NO_2 = 1.10\ atm\ O_2$

$K_p = (4.40\ atm)^4 (1.10\ atm) = 412$

E. If 420.0 grams of the copper (II) nitrate had been placed into the empty flask at 250°C, what would the total pressure have been at equilibrium?

Because the temp was kept constant, as was the size of the flask, and because some of the original 250.0g of $Cu(NO_3)_2$ was left as solid in the flask at equilibrium, any extra $Cu(NO_3)_2$ introduced into flask would remain as solid – there would be no change in the pressure.

Review work

1. At 1100 K, $K_p = 0,.25$ for the reaction

$$2SO_{2(g)} + O_{2(g)} \rightarrow 2SO_{3(g)}$$

What is the value of K at this temperature?

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

$K = \frac{K_p}{(RT)^{\Delta n}}$ $\Delta n = 2-3 = -1$

$K = (0.25)(0.0821)(1100) = 23$

2. At a particular temperature, $K_p = 0.025$ for the reaction

$$N_2O_{4(g)} \rightarrow 2NO_{2(g)}$$

a. A flask containing only N_2O_4 at an initial pressure of 4.5 atm is allowed to reach equilibrium. Calculate the equilibrium partial pressures of the gases.

b. A flask containing only NO_2 at an initial pressure of 9.0 atm is allowed to reach equilibrium. Calculate the equilibrium partial pressures of the gases.

c. From your answers to parts a and b, does it matter from which direction and equilibrium position is reached?

a)

	N_2O_4	$\leftrightarrow$	$2NO_2$
I	4.5		0
C	$-x$		$+2x$
E	$4.5-x$		$2x$

$K_p = 0.25$

$K_p = \frac{(P_{NO_2})^2}{P_{N_2O_4}} = \frac{(2x)^2}{4.5-x} = 0.025$

insignificant

$\frac{4x^2}{4.5} = 0.025$ $x = 0.17$

$P_{N_2O_4} = 4.5$ atm

$P_{NO_2} = 2(0.17) = 0.34$ atm

5% rule

$\frac{0.17}{4.5} \times 100 = 3.8\%$

b)

	N_2O_4	$\leftrightarrow$	$2NO_2$
I	0		9.0
C	$+x$		$-2x$
E	x		$9.0-2x$

$K_p = \frac{(9.0-2x)^2}{x} = 0.025$

insignificant

$\frac{(9.0)^2}{x} = 0.025$ $x = 3240$ does not make sense

$\frac{(9.0-2x)^2}{x} = 0.025$

$4x^2 - 36.025x + 81 = 0$

$\frac{-b \pm \sqrt{b^2-4ac}}{2a}$

$x = 4.67$ does not make sense

$x = 4.34$

$P_{N_2O_4} = 4.34$ atm $P_{NO_2} = 9-2x = 0.32$ atm

c) no - we get the same equilibrium either way

3. At a particular temperature, K = 4.0 x 10^{-7} for the reaction
$N_2O_{4(g)} \rightarrow 2NO_{2(g)}$

In an experiment, 1.0 mole N_2O_4 is placed in a 10.0 L vessel. Calculate the concentrations of N_2O_4 and NO_2 when this reaction reaches equilibrium.

	N_2O_4	$\longleftrightarrow$	$2NO_2$
I	0.10		0
C	$-x$		$+2x$
E	$0.10-x$		$2x$

$K = \frac{[NO_2]^2}{[N_2O_4]} = 4.0 \times 10^{-7}$

$K = \frac{(2x)^2}{(0.10-x)} = 4.0 \times 10^{-7}$

$\frac{4x^2}{0.10} = 4.0 \times 10^{-7}$

$x = 1.0 \times 10^{-4}$ M

$0.10 \times 0.05 = 0.005$ ✓

$[N_2O_4] = 0.10$

$[NO_2] = 2x = 2.0 \times 10^{-4}$ M

4. At a particular temperature, K = 2.0 x 10^{-6} for the reaction
$2CO_{2(g)} \rightarrow 2CO_{(g)} + O_{2(g)}$

If 2.0 mole CO_2 is initially placed into a 5.0 L vessel, calculate the equilibrium concentrations of all species.

	$2CO_2$	$\longleftrightarrow$	$2CO$	$+ O_2$
I	0.40		0	0
C	$-2x$		$+2x$	$+x$
E	$0.40-2x$		$2x$	x

$K = \frac{[CO]^2[O_2]}{[CO_2]^2} = 2.0 \times 10^{-6}$

$2.0 \times 10^{-6} = \frac{(2x)^2(x)}{(0.40-2x)^2}$

$2.0 \times 10^{-6} = \frac{4x^3}{(0.40)^2}$ $\quad x = 4.3 \times 10^{-3}$ M

check assumption $0.40 \times 0.05 = 0.02$ ✓

$[CO_2] = 0.40 - 2(4.3 \times 10^{-3}) = 0.39$ M

$[CO] = 2x = 8.6 \times 10^{-3}$ M

$[O_2] = 4.3 \times 10^{-3}$ M

5. A sample of solid ammonium chloride was placed in an evacuated container and then heated so that it decomposed to ammonia gas and hydrogen chloride gas. After heating, the partial pressure of NH_3 in the container was found to be 2.2 atm. Calculate K_p at this temperature for the decomposition reaction.
$NH_4Cl_{(s)} \rightarrow NH_{3(g)} + HCl_{(g)}$

	NH_4Cl	$\longrightarrow$	NH_3	$+$	HCl
I	–		0		0
C	–		2.2		2.2
E	–		2.2		2.2

$K_p = (P_{NH_3})(P_{HCl}) = (2.2)^2 = 4.8$

6. In which direction will the position of the equilibrium

$$2HI_{(g)} \rightarrow H_{2(g)} + I_{2(g)}$$

be shifted for each of the following changes. (make sure you understand why!)

a. $H_{2(g)}$ is added ←
b. $I_{2(g)}$ is removed →
c. $HI_{(g)}$ is removed ←
d. Some $Ar_{(g)}$ is added same
e. The volume of the container is doubled same
f. The temperature is decreased (the reaction is exothermic) →

7. Hydrogen for the use in ammonia production is produced by the reaction

$$CH_{4(g)} + H_2O_{(g)} \rightarrow CO_{(g)} + 3H_{2(g)}$$

What will happen to a reaction at equilibrium if

a. $H_2O_{(g)}$ is removed? ←
b. The temperature is increased? (the reaction is endothermic) →
c. An inert gas is added? same
d. $CO_{(g)}$ is removed? →
e. The volume of the container is tripled? →