Department of Chemistry
University of Texas at Austin

Name:_______________

Solving Rice Tables and Equilibria Problems – Supplemental Worksheet KEY

1. If the K_p for the following reaction is 2.4×10^{-9} and the initial concentration of CO_2 is 2 atm, what are the partial pressures of the substances at equilibrium? Hint: make necessary assumptions to solve.

$$C\ (s) + CO_2\ (g) \leftrightarrow 2CO\ (g)$$

R	C +	CO_2	$\leftrightarrow$	2CO
I	---	2 atm		0 atm
C	---	-x		+2x
E	---	2 - x		2x

$$K_p = \frac{(P_{CO})^2}{(P_{CO2})}$$

$$2.4 \times 10^{-9} = \frac{(2x)^2}{(2-x)}$$

The small K (less than 1×10^{-3}) tells us that the reactants will not change much so we can ignore x here!

$$2.4 \times 10^{-9} = \frac{(2x)^2}{(2-x)}$$

$$2.4 \times 10^{-9} = \frac{4x^2}{2}$$

$$2.4 \times 10^{-9} = 2x^2$$

$$1.2 \times 10^{-9} = x^2$$

$$3.5 \times 10^{-5} = x$$

At equilibrium:

$P_{CO2} = 2\text{-}x \approx 2$ atm

$P_{CO} = 2x = 6.9 \times 10^{-5}$ atm

Name:________________

2. Here is a general reaction with a K value of 2.8×10^{-7}:

$$A\,(aq) + B\,(aq) \leftrightarrow 2C(aq)$$

Initially you are given 4M of substance A and 4M of substance B. Set-up an equilibrium expression to solve for the equilibrium concentrations of each substance. Hint: you can solve this one all the way through!

R	A	+	B	↔	2C
I	4M		4M		0M
C	-x		-x		+2x
E	4 - x		4 - x		2x

$$K_c = \frac{[C]^2}{[A][B]}$$

$$K_c = \frac{[2x]^2}{[4-x][4-x]}$$

$$2.8 \times 10^{-7} = \frac{[2x]^2}{[4-x]^2}$$

Now take the square root of everything!

$$5.29 \times 10^{-4} = \frac{[2x]}{[4-x]}$$

$$21.2 \times 10^{-4} - 0.000529x = 2x$$

$$0.00212 - 2.000529x = 0$$

$$x = 0.00106\ M$$

Here, x is so small compared to the initial concentrations of A and B that we could have ignored it, but it was fairly easy to solve for it.

[A] = 4M - x = 4M - 0.00106 M = 3.99894 M ≈ 4M

[B] = 4M - x = 4M - 0.00106 M = 3.99894 M ≈ 4M

[C] = 2x = 0.00212 M

Name:________________

3. Given the K_C at 298K is 0.0059 for the following reaction and the initial concentration of N_2O_4 is 3.5M, set-up an equilibrium expression to solve for the equilibrium concentrations of the products and reactants. Hint: you will need a graphing calculator or program to solve fully, but you can make an assumption and still be approximately close. (Challenge: solve for K_p!)

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

R	N_2O_4 (g)	$\leftrightarrow$	$2NO_2$ (g)
I	3.5M		0
C	-x		+2x
E	3.5M - x		2x

$$K_c = \frac{[NO_2]^2}{[N_2O_4]}$$

$$0.0059 = \frac{[2x]^2}{[3.5 - x]}$$

$$0.0059 \approx \frac{[2x]^2}{[3.5]}$$

*** Here we can assume that 3.5-x ≈ 3.5. Because K is less than one, the reaction favors the reactants AND furthermore, because 3.5 is a huge starting concentration, the change in reactants "x" will be small compared to such a 3.5. We normally avoid making assumptions unless K < 1.0 x 10^{-3} but for such a large initial concentration it is reasonable.

$$0.02065 \approx 4x^2$$

$$0.0051625 \approx x^2$$

$$0.072\ M \approx x$$

So at equilibrium $[N_2O_4]$ ≈ 3.5 M and $[NO_2]$ = 2x ≈ 0.14 M

Challenge: To find K_p multiply K_C by $(RT)^{\Delta n}$

Δn = (gas moles of products) - (gas moles of reactants)

Δn = (2 mol) - (1 mol) = 1

$K_p = (K_C*(RT)^{\Delta n}) = (0.0059*(0.0821*298)^{2-1}) = (0.0059*(25.5658)^1)$

$K_p = 0.144$

Name:_______________

4. At some temperature, the K_p for the following reaction is 0.26. If you began with 0.1 atm of NO, 0.3 atm of Cl_2 and 0 atm of NOCl. What would the partial pressures be for each gas at equilibrium? Set-up the equilibrium expression in terms of "x" and describe how you could find the exact partial pressures. Hint: you will need a graphing calculator or program to solve fully.

$$2NO\ (g) + Cl_2\ (g) \leftrightarrow 2NOCl\ (g)$$

R	2NO	+	Cl_2	$\leftrightarrow$	2NOCl
I	0.1 atm		0.3 atm		0 atm
C	-2x		-x		+2x
E	0.1 - 2x		0.3 - x		2x

$$K_p = \frac{(P_{NOCl})^2}{(P_{NO})^2(P_{Cl_2})}$$

$$0.26 = \frac{(2x)^2}{(0.1 - 2x)^2(0.3 - x)}$$

$$0.0059 = \frac{[2x]^2}{[3.5 - x]}$$

$$0.0059 \approx \frac{[2x]^2}{[3.5]}$$

We would need a calculator to figure this out. The K is small but still big enough (bigger than 1×10^{-3}) where we should not ignore the x!

$0.26 = (4x^2)/(0.01 - 0.4x + 4x^2)(0.3-x)$
$0.26 = (4x^2)/(0.003 - 0.13x + 1.6x^2 - 4x^3)$
$0.00078 - 0.0338x + 0.416x^2 - 1.04x^3 = 4x^2$
$0.00078 - 0.0338x - 3.584x^2 - 1.04x^3 = 0$

The solutions could be x = -3.44 atm, -0.0203 atm or 0.0108 atm found on wolfram online or with a graphing calculator.

The positive solution is the only one that makes sense so the equilibrium partial pressures would be:
$P_{NO} = 0.1 - 2x = 0.0785$ atm
$P_{Cl2} = 0.3 - x = 0.2892$ atm
$P_{NOCl} = 2x = 0.0215$ atm

Name:________________

5. Here is a general reaction with a K value of 1.6×10^{-6}:

$$2A\ (aq) + 3B\ (aq) \leftrightarrow 2C(aq)$$

Initially you are given 0.1M of substance A and 0.2M of substance B. Set-up an equilibrium expression to solve for the equilibrium concentrations of each substance in terms of x. Hint: do not actually solve!

R	2A	+	3B	$\leftrightarrow$	2C
I	0.1M		0.2M		0M
C	-2x		-3x		+2x
E	0.1 - 2x		0.2 - 3x		2x

$$K_c = \frac{[C]^2}{[A]^2[B]^3}$$

$$1.6\times10^{-6} = \frac{[2x]^2}{[0.1-2x]^2[0.2-3x]^3}$$

If you wanted, now you could use a calculator or program to solve for each concentration.

6. Here is a general reaction with a K value of 144:

$$A_2\ (aq) + B_2\ (aq) \leftrightarrow 2C\ (aq)$$

If the initial concentrations for A_2 and B_2 are 0.7 M, find the final concentration of C. Hint: make necessary assumptions to solve.

R	A_2 (aq) +	B_2 (aq) $\leftrightarrow$	2C (aq)
I	0.7M	0.7M	0
C	-x	-x	+2x
E	0.7-x	0.7-x	2x

In this situation, you ***cannot*** assume that x is negligible because the value of K > 1.0×10^{-3}

$$K_c = \frac{[C]^2}{[A][B]}$$

$$144 = \frac{[2x]^2}{[0.7-x][0.7-x]}$$

$$144 = \frac{[2x]^2}{[0.7-x]^2}$$

$$\sqrt{144} = \sqrt{\frac{[2x]^2}{[0.7-x]^2}}$$

Name:________________

$$12 = \frac{2x}{0.7 - x}$$
$$12(0.7 - x) = 2x$$
$$8.4 - 12x = 2x$$
$$8.4 = 14x$$
$$x = 0.6M$$
$$[A_2] = [B_2] = 0.7 - x = 0.1M$$
$$[C] = 2x = 1.2M$$

Name:________________

7. Given that the molar solubility of $PbSO_4$ is 1.59 x 10^{-4} M, what is the K_{sp} of $PbSO_4$?

R	$PbSO_4$ (s) ↔	Pb^{2+} (aq) +	SO_4^{2-} (aq)
I	----	0	0
C	----	+x	+x
E	----	x	x

The molar solubility of the salt is equal to the value of change in Pb^{2+} ion concentration x because of the 1:1 ratio between the $PbSO_4$ solid and its Pb^{2+} ion.

$$K_{sp} = [Pb^{2+}][SO_4^{2-}] = [x][x] = x^2$$

$$K_{sp} = (1.59 \times 10^{-4})^2$$

$$K_{sp} = 2.53 \times 10^{-8}$$

Name:________________

8. In the previous problem, if we had placed the $PbSO_4$ solid into a solution containing 0.5M $(NH_4)_2SO_4$ what concentration of Pb^{2+} ion will be in solution at equilibrium? You will need a calculator to solve completely.

This is a **common ion** problem type. The initial concentration of SO_4^{2-} is 0.5M.

R	$PbSO_4$ (s) ↔	Pb^{2+} (aq) +	SO_4^{2-} (aq)
I	----	0	0.5M
C	----	+x	+x
E	----	x	0.5 + x

$K_{sp} = 2.53 \times 10^{-8} = [Pb^{2+}][SO^{2-}{}_4]$

(the equilibrium constant that we found in the previous problem is valid in this situation too because K is a constant for a reaction at a given temperature.)

$2.53 \times 10^{-8} = [x][0.5 + x]$

$2.53 \times 10^{-8} = (0.5x + x^2)$

$0 = x^2 + 0.5x - 2.53 \times 10^{-8}$

$x = 5.06 \times 10^{-8}$ (very small!)

$[Pb^{2+}] = x = 5.06 \times 10^{-8}$ M

9. Given that the K_{sp} at 298 K is 9.8×10^{-11} for the dissociation of $CuCl_2$, set-up an equilibrium expression to solve for the molar solubility of the salt.

R	$CuCl_2$ (s) $\leftrightarrow$	Cu^{2+} (aq) +	$2Cl^-$ (aq)
I	----	0	0
C	----	+x	+2x
E	----	x	2x

$$K_{sp} = [Cl^-]^2[Cu^{2+}]$$
$$9.8 \times 10^{-11} = (2x)^2 x$$
$$9.8 \times 10^{-11} = 4x^3$$
$$x = 2.904 \times 10^{-4}\ M$$

Here "x" represents the molar solubility of the whole salt because of the 1:1 ratio between the $CuCl_2$ solid and the Cu^{2+} ion. Aka: salt particles dissolving give off a ratio of one Cu^{2+} ion to every one $CuCl_2$ salt unit.

Name:________________

10. Here is a general reaction:

$$Ax\ (aq) \leftrightarrow By\ (aq) + Cz\ (aq)$$

a. In this general form of an aqueous reaction, how would one set up the equilibrium expression, K?

$$K = \frac{[By][Cz]}{[Ax]}$$

b. Say we have 0.5 M of Ax and 0.2 M of Cz at initial conditions. How would you set up the RICE table and K expression? Set these up and do not solve yet.

R	A_x (aq) ↔	B_y (aq) +	C_z (aq)
I	0.5	0	0.2
C	-x	+x	+x
E	0.5-x	+x	0.2+x

$$K = \frac{[x][0.2 + x]}{[0.5 - x]}$$

c. Let's say we had found our K to be 3.9x10^{-4} . Then we increased the concentration of A_x to 0.7 M. Would the reaction shift left or right?

This reaction would shift to right (the products side) in order to reestablish equilibrium due to the increase in concentration on the reactants side of the equation.

d. What will the new equilibrium concentrations be for this situation? Hint: make necessary assumptions and/or use technology to aid your solving.

R	A_x (aq) ↔	B_y (aq) +	C_z (aq)
I	**0.7**	0	0.2
C	-x	+x	+x
E	0.7-x	+x	0.2+x

$$K = \frac{[x][0.2 + x]}{[0.7 - x]}$$

$$3.9 \times 10^{-4} = \frac{[x][0.2 + x]}{[0.7 - x]}$$

***Here because K < 1.0 x 10^{-3} you may ignore the subtraction x from the reactants because the reactants experience a small change (we do not eliminate the addition of x in products).

$$3.9 \times 10^{-4} = \frac{[x][0.2 + x]}{[0.7]}$$

$$2.73 \times 10^{-4} - (0.20039)x - x^2 = 0$$

Name:_______________

$$0.00135\ M = x$$
$$[Ax] = 0.7 - x = 0.69865\ M \approx 0.7M$$
$$[By] = x = 0.00135M$$
$$[Cz] = 0.2 + x = 0.20135\ M$$