Dr. Zellmer
Time: 7 PM Sun.
30 min

Chemistry 1220
Spring Semester 2023
Quiz IV

All Sections
February 12, 2023

Name KEY Rec. Instr.

1. (3 pts) The equilibrium constant for the following reaction is 70 at 350°C. A system at equilibrium has $[N_2] = 0.200$ M and $[NH_3] = 0.118$ M. What is the $[H_2]$?

$$N_2(g) + 3\,H_2(g) \rightleftarrows 2\,NH_3(g)$$

$$K_c = \frac{[NH_3]^2}{[N_2]\,[H_2]^3} = 70$$

$$[H_2] = \left(\frac{[NH_3]^2}{K_c\ [N_2]}\right)^{1/3}$$

$$[H_2] = \left(\frac{(0.118)^2}{(70)\,(0.200)}\right)^{1/3}$$

$$= 0.0998 \text{ M}$$

$$= 0.10 \text{ M}$$

2. (3 pts) The properly written heterogeneous K_c expression for the following reaction as written is:

$$NiCO_3\,(s) + 2\,H^+\,(aq) \rightleftarrows Ni^{2+}\,(aq) + CO_2\,(g) + H_2O\,(l)$$

Pure **solids** and **liquids** do **<u>NOT</u>** appear in the equilibrium constant expression (or in Q).

$$K_c = \frac{[Ni^{2+}]\,[CO_2]}{[H^+]^2}$$

$NiCO_3$ (s) & H_2O (l) do **<u>NOT</u>** appear in K

- gases & solutes in solution do appear in the expression for K.

3. (3 pts) At equilibrium, which of the following is(are) **TRUE**?

a) All chemical processes have ceased.

b) The rate constant for the forward reaction equals that of the reverse.

c)* The rate of the forward reaction equals that of the reverse.

d) Both the rate of the forward reaction equals that of the reverse and the rate constant for the forward reaction equals that of the reverse.

e)* The concentrations of reactants and products are constant.

At equilibrium:

forward rate = reverse rate
concentrations are constant & $Q_c = K_c$ (equil. constant expression is constant)

The rate constants of the forward and reverse reactions are NOT necessarily equal.

Remember the demo in class involving the aquariums (one empty and one full). Equilibrium was reached when the same amount of beads was transferred from one aquarium to the other during the same period of time (**rate of transfer** from one aquarium to the other was **equal**). The amount of beads in each aquarium was NOT the same. However, the amounts of beads in each aquarium remained constant once equilibrium was reached.

4. (4 pts) For the following reaction $K_P = 3.14 \times 10^{-10}$ at 727.0 °C .

$$5\ NO\ (g)\ +\ 5/2\ Br_2\ (g) \rightleftarrows 5\ NOBr\ (g) \qquad \textbf{Rxn 1}$$

What is the value of K_P for the following reaction? **Show all work or explain.**

$$2\ NOBr\ (g) \rightleftarrows 2\ NO\ (g)\ +\ Br_2\ (g) \qquad \textbf{Rxn 2}$$

When multiply a rxn by some factor the K for the new rxn is equal to original K raised to the factor the rxn was multiplied by. Remember if you multiply a reaction by a -1 that's the same as reversing it since you need to flip the action around to get rid of the negative coefficients you would get by multiplying by -1. (multiplying by -1 reverses the rxn)

Based on this K_{rev} for rev. rx is $1/K_{for}$ (K_{for}^{-1}, the reciprocal of K_{for} for the forward rx)

To get Rxn 2 you need to reverse the original reaction (rxn 1) and multiply it by 2/5 This is the same as multiplying reaction (1) by - 2/5 .

$$K_2 = K_1^{-2/5} = 1/(K_1^{2/5}) = (3.14 \times 10^{-10})^{-2/5} = 6.3\underline{\mathbf{2}}74 \times 10^3 = 6.33 \times 10^3 \qquad (3\ \text{s.f.})$$

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5. (5 pts) Given the following two equilibrium reactions,

(1) $2\ NO\ (g) + Br_2\ (g) \rightleftarrows 2\ NOBr\ (g)$ $K_1 = 2.00$

$$K_1 = \frac{[NOBr]^2}{[NO]^2\ [Br_2]}$$

(2) $NO\ (g) \rightleftarrows \frac{1}{2}\ N_2\ (g) + \frac{1}{2}\ O_2\ (g)$ $K_2 = 1.45 \times 10^{15}$

$$K_2 = \frac{[N_2]^{1/2}\ [O_2]^{1/2}}{[NO]}$$

What is the equilibrium constant, K_3, for the reaction below,

(3) $N_2\ (g) + O_2\ (g) + Br_2\ (g) \rightleftarrows 2\ NOBr\ (g)$ K_3

K_{rev} for rev. rx is $1/K_{for}$ (reciprocal of K_{for} for the forward rx)

Also, when multiply a rxn by some factor the K for the new rxn is equal to original K raised to the factor the rxn was multiplied by (multiplying by -1 reverses the rxn)

Proof

$a\ A + b\ B \rightleftarrows x\ X + y\ Y$

$$K_1 = \frac{[X]^x\ [Y]^y}{[A]^a\ [B]^b}$$

Multiply by n

$(a \bullet n)\ A + (b \bullet n)\ B \rightleftarrows (x \bullet n)\ X + (y \bullet n)\ Y$

$$K_2 = \frac{[X]^{xn}\ [Y]^{yn}}{[A]^{an}\ [B]^{bn}} = \left(\frac{[X]^x\ [Y]^y}{[A]^a\ [B]^b}\right)^n = K_1^{\ n}$$

Multiply by -1 reverses rxn.
$-a\ A + -b\ B \rightleftarrows -x\ X + -y\ Y$

want (+) coeff - like in math switch reactants & products to other sides

$x\ X + y\ Y \rightleftarrows a\ A + b\ B$

***** **continued on next page** *****

5. (cont.)

When you add reactions together to get another reaction the K for the overall reaction equals the K's from the added reactions multiplied together.

Rxn 1 + Rxn 2 = Rxn 3

$K_1 * K_2 = K_3$

Proof

(1) $a\,A + b\,B \rightleftarrows x\,X + y\,Y$ and (2) $c\,C + d\,D \rightleftarrows r\,R + s\,S$

$$K_1 = \frac{[X]^x\,[Y]^y}{[A]^a\,[B]^b} \qquad K_2 = \frac{[R]^r\,[S]^s}{[C]^c\,[D]^d}$$

Add the two equations (reactions) to get a third equation (reaction)

$a\,A + b\,B + c\,C + d\,D \rightleftarrows x\,X + y\,Y + r\,R + s\,S$

$$K_3 = \frac{[X]^x\,[Y]^y\,[R]^r\,[S]^s}{[A]^a\,[B]^b\,[C]^c\,[D]^d} = \left(\frac{[X]^x\,[Y]^y}{[A]^a\,[B]^b}\right) * \left(\frac{[R]^r\,[S]^s}{[C]^c\,[D]^d}\right)$$

$K_3 = K_1 * K_2$

To get reaction 3 for this question,

(3) $N_2\,(g) + O_2\,(g) + Br_2\,(g) \rightleftarrows 2\,NOBr\,(g)$

must first reverse rx 2 and multiply it by 2 and then add it to rxn 1

$(2)_{rev*2}$ $N_2\,(g) + O_2\,(g) \rightleftarrows 2\,NO\,(g)$ $K_{rev*2} = K_2^{-2}$

(1) $2\,NO\,(g) + Br_2\,(g) \rightleftarrows 2\,NOBr\,(g)$ K_1

$N_2\,(g) + O_2\,(g) + Br_2\,(g) \rightleftarrows 2\,NOBr\,(g)$ K_3

$K_3 = K_1 * K_{rev*2} = K_1 * K_2^{-2} = (2.00)(1.45 \times 10^{15})^{-2} = 9.51 \times 10^{-31}$

6. (4 pts) The equilibrium constant K_c for the following reaction at 1100°C is 6.80×10^{51}. What is K_p?

$$B(\mathbf{s}) + 3/2\ F_2(g) \rightleftarrows BF_3(g)$$

use $K_p = K_c (RT)^{\Delta n}$ (Δn = change in moles of **gas**)

$\Delta n = 1 - 3/2 = -1/2$ Do **NOT** include the **solid** (if you do then $\Delta n = -3/2$)

$$K_c = \frac{[BF_3]}{[F_2]^{3/2}} \qquad Kp = \frac{P_{BF3}}{(P_{F2})^{3/2}}$$

R = 0.08206 L•atm/mol•K **NOT** 8.314 J/mol•K

T = 1100°C + 273 = 1373 K (**must** be in **kelvin**)

$K_c = 6.80 \times 10^{51}$

$$K_p = K_c (RT)^{-1/2}$$

$$= (6.80 \times 10^{51}) \{(0.08206)(1373)\}^{-1/2}$$

$$K_p = 6.40 \times 10^{50}$$

7. (3 pts) For the following reaction $K_C = 25.8$. The reaction is started with 1.000 mole of AB_3, 2.000 moles of AB_2 and 1.000 mole of B_2 in a 2.000-L container. Determine if the reaction is at equilibrium or not? If not, which direction will it proceed to reach equilibrium? **Show work and explain!**

$$2\ AB_2(g) + B_2(g) \rightleftarrows 2\ AB_3(g)$$

Determine reaction quotient, Q, and compare to K.

Q = K system at equlibirium

Q < K not at equilibrium - not enough product & too much reactant
shifts to **right** **(forward)** (------>) **toward** **product** to reach equil. (when Q = K)

Q > K not at equilibrium - too much product & not enough reactant
shifts to **left** **(reverse)** (<------) **toward** **reactant** to reach equil. (when Q = K)

$$Q_c = \frac{[AB_3]^2}{[AB_2]^2\,[B_2]} = \frac{(0.5000)^2}{(1.000)^2\,(0.5000)} = 0.5000$$

$Q_c = 0.5000$ $K_c = 25.8$

$Q_c < K_c$ **numerator** ([prod]) is **too** **small**, need to **inc numerator** (**products**)

∴ rxn is **NOT** at **equilibrium**: Too much reactant and not enough product so the reaction proceeds to the **right** **toward products** (**forward** direction)

8. (5 pts) The following reaction is started with 2.000 moles of SO_3 in a 2.000-L container. When equilibrium is reached there are 1.645 moles of SO_3 in the container. What is the value of the **equilibrium constant**, K_C? **(Show the ICE table. You can use numbers or variables in your ICE table.)**

$$2\,SO_3(g) \rightleftarrows 2\,SO_2(g) + O_2(g)$$

$$K_C = \frac{[SO_2]^2 \bullet [O_2]}{[SO_3]^2}$$

Concentration of **Products** in **Numerator** Concentration of **Reactants** in **Denominator**

Use **coefficients** in **balanced equation** as the **exponents** of the concentrations.

Pure solids (s) and liquids (ℓ) do not appear in the expression for K or Q.

The solution to the actual problem given is:

<u>Use ICE (equil.) table</u> (in Molarity when dealing w. K_C)

	$2\,SO_3(g)$	$\rightleftarrows$	$2\,SO_2(g)$	+	$O_2(g)$
initial	1.000 M		0 M		0 M
change	-2x		+ 2x		+ x
equil	1.0000 - 2x		2x		x

you know rxn must proceed to right (in forward direction) since starting with no SO_2 or O_2. Conc. of SO_3 is 1.000 M since you have 2 mol/2 L. Must use Molarity in ICE table.

$$K_C = \frac{[SO_2]^2\,[O_2]}{[SO_3]^2} = \frac{(2x)^2\,(x)}{(1.000 - 2x)^2} = \frac{4x^3}{(1.000 - 2x)^2}$$

Can solve for x from the information given for the equilibrium concentration of SO_3.

$(1.000 - 2x) = 1.645 \text{ mol}/2.000 \text{ L} = 0.8225 \text{ M}$

$x = 0.088\underline{\mathbf{7}}5 \text{ M}$

$$K_C = \frac{4x^3}{(1.000 - 2x)^2} = \frac{4\,(0.08875)^3}{\{1.000 - 2(0.08875)\}^2} = 4.1\underline{\mathbf{3}}33 \times 10^{-3} = \underset{(3\text{ s.f.})}{4.13 \times 10^{-3}}$$

9. (4 pts) For the following reaction K_C equals 7.10×10^{-4}, at 25 °C.

$$CaCrO_4\,(s) \rightleftarrows Ca^{2+}\,(aq) + CrO_4^{2-}\,(aq)$$

What are the **equilibrium** concentrations of Ca^{2+} and CrO_4^{2-} if solid $CaCrO_4$ is placed in water to form a saturated solution at 25°C? (**Show the ICE table.**)

Use ICE (equil.) table (in Molarity)

	$CaCrO_4$ (s)	$\rightleftarrows$	Ca^{2+} (aq)	+	CrO_4^{2-} (aq)
initial	C		0 M		0 M
change	-x		+ x		+ x
equil	constant		+ x		+ x

$$K_c = [Ca^{2+}] \bullet [CrO_4^{2-}] = 7.10 \times 10^{-4}$$

$CaCrO_4$ (s) doesn't appear in K (pure solids and liquids do not appear in K)

$$x^2 = 7.10 \times 10^{-4}$$

$$x = (7.10 \times 10^{-4})^{1/2} = 2.6\underline{6}45 \times 10^{-2}\ M$$

$$[Ca^{2+}] = [CrO_4^{2-}] = 2.66 \times 10^{-2}\ M \quad (3\ s.f.)$$

10. (3 pts) A plot of ln(r) vs. ln[A] has a slope of -2.5 and an intercept of -10.55. Determine the **rate constant** and **order** of the reaction for the rate law, $r = k[A]^n$.

$r = k\ [A]^n$ k = rate constant n = order

Convert this to a linear equation.

Take the log of both sides, use the rules of logarithms and rearrange to make it look like an equation for a straight line. You can use either log or ln.

$$\log(r) = \log\{k[A]^n\} = \log(k) + \log\{[A]^n\} = \log(k) + n \log[A]$$

$$\log(r) = n*\log[A] + \log(k) \qquad \text{or} \qquad \ln(r) = n*\ln[A] + \ln(k)$$

$$y = m*x + b$$

Using the equation from (a) how do you graphically obtain k and n (i.e. what do you plot as x and y and what are n and k related to on the graph)?

The data you have is concentration, [A], and the corresponding rate. You take the log of each and plot log(r) vs log[A] (i.e. log(r) is on the y axis and log[A] is on the x axis) or ln(r) and ln[A].

if using log: $\log(k) = \text{y-intercept}$ $k = 10^{(\text{y-int})}$

if using ln: $\ln(k) = \text{y-intercept}$ $k = e^{(\text{y-int})}$

n = slope (The slope can be either positive or negative depending on the order.)

$n = -2.5$ $\ln(k) = -10.55$ $k = e^{-10.55} = 2.6 \times 10^{-5}\ M^{3.5} \bullet s^{-1}$

(Assuming molarity and seconds for units.)

This was the problem I presented in lecture and told everyone to do and ask the TAs or me whether you did it correctly. Only a few people asked.