Quiz IV General Chemistry II Lecture Dr. Hahn 20 pts 3/1/13 F form A 9:30 am quiz # _____

Name Key (print) Name __________________ (sign)
Please show all work for full credit & to get partial credit. (suggestion: A guess is better than no answer.)

1. For the following overall reaction (not reaction mechanism step, the overall reaction), Given the concentrations and rates, give the order of the reactant by circling the order for the reagent given. You should assume an irreversible reaction. (note: I made up these reactions to illustrate the point so the reactions as given may not go experimentally as written.) (3 pts each, 6 pts total)

$Cl_2 + 2\,O_3 \rightarrow 2\,ClO + 2\,O_2$

double Cl_2 – rate no change
double O_3 – rate doubles
rate = $k[Cl_2]^0[O_3]$

$[Cl_2]$	$[O_3]$	rate
2	2	4
4	2	4
2	4	8

order of the [Cl] is (zero) (one) (circle one)
order of the $[O_3]$ is (zero) (one) (circle one)

2. For the following reaction give the equation for the equilibrium constant (K) in concentrations of the reagents given. You should assume that the gases dissolve so that you can give K in concentrations of the units of molarity. (6 pts)

$C(s) + H_2O\,(g) \rightarrow CO\,(g) + H_2\,(g)$

C is solid – leave out of K expression

$$K = \frac{[CO][H_2]}{[H_2O]}$$

3. For the reaction given, set up the ICE table for a reaction in which the reactant gases are mixed in a constant volume of an inert solvent with no products present initially: (1 pt each blank, 9 pts total)

$CO\,(g) + Cl_2\,(g) \rightarrow COCl_2\,(g)$

If the initial concentration of the CO (g) is 0.415 M, and the intial concentration of Cl_2 (g) is 0.543 M. If at equilibrium, the concentration of the CO Cl_2 is 0.098 M. Show the initial, change and equilibrium concentrations for all 3 reagents .

	[CO]	$[Cl_2]$	$[COCl_2]$
Initial	0.415	0.543	0
Change	−0.098	− 0.098	+ 0.098
Equilibrium	0.317	0.445	0.098

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1. For the following overall reaction (not reaction mechanism step, the overall reaction), Given the concentrations and rates, give the order of the reactant by circling the order for the reagent given. You should assume an irreversible reaction. (note: I made up these reactions to illustrate the point so the reactions as given may not go experimentally as written.) (3 pts each, 6 pts total)

$H_2 + 2\,Br_2 \rightarrow 2\,HBr + 2\,Br^-$

double Br_2 - rate doubles
double H_2 - rate not effected

$[H_2]$	$[Br_2]$	rate
3	3	1
3	6	2
6	3	1

$\text{rate} = k[H_2]^0[Br_2]^1$

order of the $[H_2]$ is (zero) (one) (circle one)
order of the $[Br_2]$ is (zero) (one) (circle one)

2. For the following reaction give the equation for the equilibrium constant (K) in concentrations of the reagents given. You should assume that the gases dissolve so that you can give K in concentrations of the units of molarity. (6 pts)

$2\,BrNO\,(g) \rightarrow 2\,NO\,(g) + Br_2\,(g)$

$$K = \frac{[NO]^2[Br_2]}{[BrNO]^2}$$

using $aA + bB \rightleftharpoons cC + dD$

$$K = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

3. For the reaction given, set up the ICE table for a reaction in which the reactant gases are mixed in a constant volume of an inert solvent with no products present initially: (1 pt each blank, 9 pts total)

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

If the initial concentration of the N_2 (g) is 0.985 M, and the intial concentration of H_2 (g) is 0.996 M. If at equilibrium, the concentration of the NH_3 is 0.057 M. Show the initial, change and equilibrium concentrations for all 3 reagents .

0.0285 0.0855

	$[N_2]$	$[H_2]$	$[NH_3]$
Initial	0.985	0.996	0
Change	$-\frac{0.057}{2}$	$-\left(\frac{0.057}{2}\right)\times 3$	+0.057
Equilibrium	0.957	0.911	0.057

for +0.057 NH_3 → ÷2 for N_2 →
÷2 × 3 for H_2

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H_2 (g) + Cl (g) → HCl (g) + H (g)

$[H_2]$	[Cl]	rate
9	9	4
9	18	4
18	9	4

double [Cl] rate no change

double $[H_2]$ rate no change

rate $= k[H_2]^0[Cl]^0$

order of the $[H_2]$ is (zero) (one) (circle one)

order of the [Cl] is (zero) (one) (circle one)

2. For the following reaction give the equation for the equilibrium constant (K) in concentrations of the reagents given. You should assume that the gases dissolve so that you can give K in concentrations of the units of molarity. (6 pts)

CO_3^{-2} (aq) + H_2O (l) → HCO_3^- (aq) + OH^- (aq)

$$K = \frac{[HCO_3^-][OH^-]}{[CO_3^{-2}]}$$

H_2O is a liquid leave it out of equilibrium expression

3. For the reaction given, set up the ICE table for a reaction in which the reactant gases are mixed in a constant volume of an inert solvent with no products present initially: (1 pt each blank, 9 pts total)

2 SO_2 (g) + O_2 (g) → 2 SO_3 (g)

If the initial concentration of the SO_2 (g) is 0.117 M, and the initial concentration of O_2 (g) is 0.223 M. If at equilibrium, the concentration of the SO_3 is 0.043 M. Show the initial, change and equilibrium concentrations for all 3 reagents.

0.0215

	$[SO_2]$	$[O_2]$	$[SO_3]$
Initial	0.117	0.223	0
Change	−0.043	− 0.043/2	+0.043
Equilibrium	0.074	0.202	0.043

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1. For the following overall reaction (not reaction mechanism step, the overall reaction), Given the concentrations and rates, give the order of the reactant by circling the order for the reagent given. You should assume an irreversible reaction. (note: I made up these reactions to illustrate the point so the reactions as given may not go experimentally as written.) (3 pts each, 6 pts total)

$2\ NH_3\ (aq) + OCl^-\ (aq) \rightarrow N_2H_4\ (aq) + H_2O\ (l) + Cl^-\ (aq)$

$[NH_3]$	$[OCl^-]$	rate
2	4	3
4	4	6
2	8	3

double $[NH_3]$ double rate

double OCl^-, no effect rate

$\text{rate} = k[NH_3][OCl^-]^0$

order of the $[NH_3]$ is (zero) (one) (circle one)

order of the $[OCl^-]$ is (zero) (one) (circle one)

2. For the following reaction give the equation for the equilibrium constant (K) in concentrations of the reagents given. You should assume that the gases dissolve so that you can give K in concentrations of the units of molarity. (6 pts)

using $aA + bB \rightleftarrows cC + dD$

$N_2(g) + 3H_2\ (g) \rightarrow 2\ NH_3\ (g)$

$$K = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

$$K = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

3. For the reaction given, set up the ICE table for a reaction in which the reactant gases are mixed in a constant volume of an inert solvent with no products present initially: (1 pt each blank, 9 pts total)

$N_2\ (g) + O_2\ (g) \rightarrow 2\ NO\ (g)$

If the initial concentration of the N_2 (g) is 0.789 M, and the intial concentration of O_2 (g) is 0.899 M. If at equilibrium, the concentration of the NO is 0.230 M. Show the initial, change and equilibrium concentrations for all 3 reagents .

0.115

	$[N_2]$	$[O_2]$	[NO]
Initial	0.789	0.899	0
Change	− 0.230/2	− 0.230/2	+ 0.230
Equilibrium	0.674	0.784	0.230

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$Cl_2 + 2 O_3 \rightarrow 2 ClO + 2 O_2$

$[Cl_2]$	$[O_3]$	rate
2	2	4
4	2	4
2	4	8

order of the [Cl] is (zero) (one) (circle one)
order of the $[O_3]$ is (zero)(one) (circle one)

2. For the following reaction give the equation for the equilibrium constant (K) in concentrations of the reagents given. You should assume that the gases dissolve so that you can give K in concentrations of the units of molarity. (6 pts)

$C(s) + H_2O (g) \rightarrow CO (g) + H_2 (g)$

K =

3. For the reaction given, set up the ICE table for a reaction in which the reactant gases are mixed in a constant volume of an inert solvent with no products present initially: (1 pt each blank, 9 pts total)

$CO (g) + Cl_2 (g) \rightarrow COCl_2 (g)$

If the initial concentration of the CO (g) is 0.415 M, and the intial concentration of Cl_2 (g) is 0.543 M. If at equilibrium, the concentration of the CO Cl_2 is 0.098 M. Show the initial, change and equilibrium concentrations for all 3 reagents .

	[CO]	$[Cl_2]$	$[COCl_2]$
Initial			
Change			
Equilibrium			

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4. For the following overall reaction (not reaction mechanism step, the overall reaction), Given the concentrations and rates, give the order of the reactant by circling the order for the reagent given. You should assume an irreversible reaction. (note: I made up these reactions to illustrate the point so the reactions as given may not go experimentally as written.) (3 pts each, 6 pts total)

$H_2 + 2\ Br_2 \rightarrow 2\ HBr + 2\ Br^-$

$[H_2]$	$[Br_2]$	rate
3	3	1
3	6	2
6	3	1

order of the $[H_2]$ is (zero) (one) (circle one)
order of the $[Br_2]$ is (zero)(one) (circle one)

5. For the following reaction give the equation for the equilibrium constant (K) in concentrations of the reagents given. You should assume that the gases dissolve so that you can give K in concentrations of the units of molarity. (6 pts)

$2\ BrNO\ (g) \rightarrow 2\ NO\ (g) + Br_2\ (g)$

K =

6. For the reaction given, set up the ICE table for a reaction in which the reactant gases are mixed in a constant volume of an inert solvent with no products present initially: (1 pt each blank, 9 pts total)

$N_2(g) + 3H_2\ (g) \rightarrow 2\ NH_3\ (g)$

If the initial concentration of the N_2 (g) is 0.985 M, and the intial concentration of H_2 (g) is 0.996 M. If at equilibrium, the concentration of the NH_3 is 0.057 M. Show the initial, change and equilibrium concentrations for all 3 reagents .

	$[N_2]$	$[H_2]$	$[NH_3]$
Initial			
Change			
Equilibrium			

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7. For the following overall reaction (not reaction mechanism step, the overall reaction), Given the concentrations and rates, give the order of the reactant by circling the order for the reagent given. You should assume an irreversible reaction. (note: I made up these reactions to illustrate the point so the reactions as given may not go experimentally as written.) (3 pts each, 6 pts total)

H_2 (g) + Cl (g) → HCl (g) + H (g)

$[H_2]$	[Cl]	rate
9	9	4
9	18	4
18	9	4

order of the $[H_2]$ is (zero) (one) (circle one)
order of the [Cl] is (zero)(one) (circle one)

8. For the following reaction give the equation for the equilibrium constant (K) in concentrations of the reagents given. You should assume that the gases dissolve so that you can give K in concentrations of the units of molarity. (6 pts)

CO_3^{-2} (aq) + H_2O (l) → HCO^{3-} (aq) + OH^- (aq)

K =

9. For the reaction given, set up the ICE table for a reaction in which the reactant gases are mixed in a constant volume of an inert solvent with no products present initially: (1 pt each blank, 9 pts total)

2 SO_2 (g) + O_2 (g) → 2 SO_3 (g)

If the initial concentration of the SO_2 (g) is 0.117 M, and the intial concentration of O_2 (g) is 0.223 M. If at equilibrium, the concentration of the SO_3 is 0.043 M. Show the initial, change and equilibrium concentrations for all 3 reagents .

	$[SO_2]$	$[O_2]$	$[SO_3]$
Initial			
Change			
Equilibrium			

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10. For the following overall reaction (not reaction mechanism step, the overall reaction), Given the concentrations and rates, give the order of the reactant by circling the order for the reagent given. You should assume an irreversible reaction. (note: I made up these reactions to illustrate the point so the reactions as given may not go experimentally as written.) (3 pts each, 6 pts total)

$2\ NH_3\ (aq)\ +\ OCl^-\ (aq) \rightarrow N_2H_4\ (aq)\ +\ H_2O\ (l)\ +\ Cl^-\ (aq)$

$[NH_3]$	$[OCl^-]$	rate
2	4	3
4	4	6
2	8	3

order of the $[NH_3]$ is (zero) (one) (circle one)
order of the $[OCl^-]$ is (zero)(one) (circle one)

11. For the following reaction give the equation for the equilibrium constant (K) in concentrations of the reagents given. You should assume that the gases dissolve so that you can give K in concentrations of the units of molarity. (6 pts)

$N_2(g) + 3H_2\ (g) \rightarrow 2\ NH_3\ (g)$

K =

12. For the reaction given, set up the ICE table for a reaction in which the reactant gases are mixed in a constant volume of an inert solvent with no products present initially: (1 pt each blank, 9 pts total)

$N_2\ (g) + O_2\ (g) \rightarrow 2\ NO\ (g)$

If the initial concentration of the N_2 (g) is 0.789 M, and the intial concentration of O_2 (g) is 0.899 M. If at equilibrium, the concentration of the NO is 0.230 M. Show the initial, change and equilibrium concentrations for all 3 reagents .

	$[N_2]$	$[O_2]$	[NO]
Initial			
Change			
Equilibrium			