Recitation Worksheet 8: Chemical Equilibrium II (13.6 – 13.7, 17.8 – 17.9)

Name: Key

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Instructions:

1. Please enter your first and last name as it appears on the eLC classlist (do not use a nickname).
2. Your UGA myID is a combination of letters and numbers (example: Dr. Abdelrahman MyID is ema88805). **Do not use your 81x number.**
   a. If you do not have access to a printer, type your answers in the worksheet PDF and then upload it to **Gradescope** by Friday, March 24th at 11:59 pm. Write your work on separate sheets of paper, convert to a PDF and upload to the "Recitation Worksheet 8 Dropbox" on eLC.
   b. If you are using an app to annotate the worksheet, make sure the pages are in the correct order and have the same layout as the original or Gradescope will not be able to read it.
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3. Answers must be written in the corresponding answer box, or no credit will be awarded.
4. The instructions for uploading worksheets to Gradescope can be found in the Content area of eLC in the Welcome Module.

1. Phosgene, a colorless toxic gas ($COCl_2$) used in the synthesis of dyes and resins can be synthesized using carbon monoxide and chorine gases represented by the equilibrium reaction below

$$CO(g) + Cl_2(g) \rightleftharpoons COCl_2(g) \quad K_c = 1.2 \times 10^3 \text{ at } 395\ °C$$

If 4.20 g CO(g) and 10.6 g Cl_2(g) are injected in a 3.25 L flask, what is the equilibrium concentration of (please use scientific notation and keep your answers to 3 significant figures):

A. $Cl_2(g)$ $\boxed{5.72} \times 10^{\boxed{-3}}$ M

① Calculate the initial concentrations of CO & Cl_2

$$\text{Molarity of CO} = \frac{4.20\,g\,CO \times \frac{1\,mol\,CO}{28.01\,g\,CO}}{3.25\,L} = 0.04613736852$$

$$\text{Molarity of } Cl_2 = \frac{10.6\,g\,Cl_2 \times \frac{1\,mol\,Cl_2}{70.90\,g\,Cl_2}}{3.25\,L} = 0.04599546854$$

② Set up an ICE table to calculate the equilibrium concentrations

Calculating Q_c

$= \frac{[COCl_2]}{[CO][Cl_2]}$

$= \frac{[0]}{[0.0461][0.0460]} = 0$

$Q_c < K_c$ ∴ reaction proceeds to the right

numbers are cut off due to limited space

	$CO(g)$ +	$Cl_2(g) \rightleftharpoons$	$COCl_2(g)$
I	0.0461	0.0460	0
C	$-x$	$-x$	$+x$
E	$0.0461-x$	$0.0460-x$	x

③ Plug the equilibrium values from the ICE table into the equilibrium constant expression

$K_c = \frac{[COCl_2]}{[CO][Cl_2]}$

$1.2 \times 10^3 = \frac{x}{(0.0461-x)(0.0460-x)}$

$1.2 \times 10^3 (2.1221097 \times 10^{-3} - 9.213283406 \times 10^{-2} + x^2) = x$

$2.546531693 - 1.105594009 \times 10^2 x + 1.2 \times 10^3 x^2 = x$

$1.2 \times 10^3 x^2 - 1.111659 4009 \times 10^2 x + 2.546531693 = 0$

$x = 0.0402728289$ or $x = 0.05\cancel{2}69334352$

B. $CO(g)$ $\boxed{5.86}$ X 10 $\boxed{-3}$ M

$[CO]_{eq} =$

$0.04613736852 -$
0.0402728289
$= 0.00586454462$ M

$[Cl_2]_{eq} = 0.04599546554 -$
$0.0402728289 =$
0.00572264164 M

* Keep all your digits & round at the end to 3 sig figs

C. $COCl_2(g)$ $\boxed{4.03}$ X 10 $\boxed{-2}$ M to check if your equilibrium concentrations are correct plug the equilibrium values you calculated back into the equilibrium constant expression

$$K_c = \frac{[COCl_2]}{[CO][Cl_2]} = \frac{[0.0403]}{[0.00586][0.00572]}$$

$$= 1.20 \times 10^3$$

↓

matches Kc from the question

2. A mixture of 0.250 mol $H_2(g)$ and 0.250 mol $I_2(g)$ in a 4.05 L flask is represented by the reaction below

$$H_2(g) + I_2(g) \rightleftharpoons 2\,HI(g) \quad K_c = 50.2$$

Calculate the equilibrium concentrations of (please use scientific notation and keep your answers to 3 significant figures):

A. $H_2(g)$

$\boxed{1.36} \times 10^{\boxed{-2}}$ M

① Conc of H_2 & I_2

$$[H_2] = [I_2] = \frac{0.250 \text{ mol}}{4.05 \text{ L}} = 0.0617283950\underline{6}\text{ M}$$

Q=0 ∴ reaction proceeds towards products

numbers are truncated due to limited space

	$H_2(g)$	+ $I_2(g)$	$\rightleftharpoons$ 2 HI(g)
I	0.0617	0.0617	0
C	$-x$	$-x$	$+2x$
E	$0.0617-x$	$0.0617-x$	$2x$

② Plug the equilibrium conc. in Kc & solve for x

$$K_c = \frac{[HI]^2}{[H_2][I_2]}$$

$$50.2 = \frac{(2x)^2}{(0.0617-x)(0.0617-x)}$$

$$\sqrt{50.2} = \sqrt{\frac{4x^2}{(0.0617-x)^2}}$$

taking the square root of both sides

$$7.085195834 = \frac{2x}{0.0617-x}$$

$$2x = 7.0851958\,(0.0617-x)$$

$$2x = 0.4373577675 - 7.085195834x$$

$$9.085195834x = 0.4373577675$$

$$\therefore x = 0.04813960816$$

B. $I_2(g)$

$\boxed{1.36} \times 10^{\boxed{-2}}$ M

C. HI(g)

$\boxed{9.63}$ X 10 $^{\boxed{-2}}$ M

③ $[H_2] = [I_2] = 0.0617\underline{2}839506 - 0.04813960816$
$= 0.013\underline{5}887869$
$= 0.0136\,M$

$[HI] = 2x = 2(0.04813960816)$
$= 0.0963\,M$

check your answers:

$k_c = \frac{[HI]^2}{[H_2][I_2]} = \frac{[0.0963]^2}{[0.0136][0.0136]} = 50.1$ close to 50.2

Le Châtlier's principle

3. Which of the statements is true regarding the equilibrium reaction $A(s) \rightleftharpoons B(s) + 2\,C(g) + \frac{1}{2}\,D(g)$ $\Delta H = 0$?

$\boxed{E}$

A. The addition of excess solid A(s) at equilibrium will not have any effect on the equilibrium mixture. ✓
B. If the volume of the equilibrium mixture is decreased by doubling the pressure, the system shifts to the left to reestablish equilibrium. ✓
C. Increasing or decreasing the temperature will not affect the equilibrium constant, Kc, for this reaction. ✓
D. If Ne(g) is introduced to the equilibrium reaction at constant volume, it will have no effect on the equilibrium position.
E. All of the statements above are true. ✓

A Additions of solids & liquids at equilibrium has no effect on the equilibrium reaction

B Using the ideal gas law $PV = nRT$, $P \propto \frac{1}{V}$ & $P \propto n$
if $V\downarrow$ ∴ $P\uparrow$ & $\uparrow n$
∴ the system shifts to the left (fewer moles) to reestablish equilibrium

C $\Delta H = 0$ ∴ temperature has no effect on the equilibrium constant

D Introducing an inert gas at constant volume increases the total pressure but has no effect on the concentrations or partial pressures of the reactants or products

4. For the hypothetical equilibrium reaction

$$A(g) + B(g) \rightleftharpoons C(g) + D(g) \quad K_c = 4.0$$

If you start with 0.373 M A(g), 0.396 M B(g), 0.552 M C(g), and 3.95 M D(g), what is the equilibrium concentration of C(g) and D(g)? Keep your answers to 3 significant figures.

A. C(g)

$\boxed{0.357}$ M

① Determine Q_c and compare to K_c to determine the direction of reaction

$$Q_c = \frac{[C][D]}{[A][B]} = \frac{[0.552][3.95]}{[0.373][0.396]} = 14.8$$

$Q_c > K_c$ ∴ the reaction proceeds towards reactants

② Set up an ICE table

	A(g) +	B(g) ⇌	C(g) +	D(g)
I	0.373	0.396	0.552	3.95
C	$+x$	$+x$	$-x$	$-x$
E	$0.373+x$	$0.396+x$	$0.552-x$	$3.95-x$

③ $$K_c = \frac{[C][D]}{[A][B]}$$

$$4.0 = \frac{[0.552-x][3.95-x]}{[0.373+x][0.396+x]}$$

$$4.0 = \frac{2.18 - 4.50x + x^2}{0.148 + 0.769x + x^2}$$

B. D (g)

$\boxed{3.76}$ M

$$4.0(0.148 + 0.769x + x^2) = 2.18 - 4.50x + x^2$$
$$0.592 + 3.08x + 4.0x^2 = 2.18 - 4.50x + x^2$$
$$3.0x^2 + 7.58x - 1.59 = 0$$
$$x = 0.195 \text{ or } x = \cancel{-2.72}$$

$[A] = 0.373 + 0.195 = 0.568$
$[B] = 0.396 + 0.195 = 0.591$

$[C] = 0.552 - 0.195 = 0.357$ M
$[D] = 3.95 - 0.195 = 3.76$ M

Check your answer

$$K_c = \frac{[C][D]}{[A][B]} = \frac{[0.357][3.76]}{[0.568][0.591]} = 3.998 \sim 4.0$$

5. If lead metal is added to a 0.100 M Cr^{3+}(aq) solution. What are the concentrations of Pb^{2+}(aq), Cr^{2+}(aq), and Cr^{3+}(aq) when the reaction is at equilibrium? <u>(Please use scientific notation and keep your answers to 3 significant figures)</u>:

solids are not included in the equilibrium constant expression

Pb(s) + 2 Cr^{3+}(aq) ⇌ Pb^{2+}(aq) + 2 Cr^{2+}(aq) $K_c = 3.2 \times 10^{-10}$

A. Pb^{2+}(aq)

$\boxed{9.28}$ X 10 $\boxed{-5}$ M

$$Pb(s) + 2\,Cr^{3+}(aq) \rightleftharpoons Pb^{2+}(aq) + 2\,Cr^{2+}(aq)$$

I	-	0.100	0	0
C	-	$-2x$	$+x$	$+2x$
E	-	$0.100-2x$	x	$2x$

① $K_c = \dfrac{[Pb^{2+}][Cr^{2+}]^2}{[Cr^{3+}]^2}$

$3.2\times10^{-10} = \dfrac{(x)(2x)^2}{(0.100-\cancel{2x})^2}$

$\dfrac{C}{K} = \dfrac{0.100}{3.2\times10^{-10}} >> 100$ ∴ x is dropped in the denominator

B. $Cr^{2+}(aq)$

$\boxed{1.86} \times 10^{\boxed{-4}}$ M

② $3.2\times10^{-10} = \dfrac{(x)(2x)^2}{(0.100)^2}$

$(0.100)^2 \times 3.2\times10^{-10} = 4x^2 \times x$

$3.2\times10^{-12} = 4x^3$

∴ $x^3 = 8.0\times10^{-13}$ → take the cube root of both sides

∴ $x = 9.28317767\times10^{-5}$

C. $Cr^{3+}(aq)$

$\boxed{9.98} \times 10^{\boxed{-2}}$ M

$[Pb^{2+}] = x = 9.28\times10^{-5}$ M

$[Cr^{2+}] = 2x = 2(9.28\times10^{-5}) = 1.86\times10^{-4}$ M

$[Cr^{3+}] = 0.100 - 2(9.28\times10^{-5}) = 9.98\times10^{-2}$ M

6. The equilibrium reaction for the synthesis of 1,2 dichloroethane is given below

$$C_2H_4(g) + Cl_2(g) \rightleftharpoons C_2H_4Cl_2(g) \quad \Delta H = -1244 \text{ kJ/mol}$$

2 mol (g) 1 mol (g)

To maximize the amount of $C_2H_4Cl_2(g)$ produced, which of these strategies might be applied? Assume the reaction is at equilibrium. Select all that apply. Insert letters without spaces in the answer box, example **ABCD.**

$\boxed{BCD}$

~~A.~~ Increasing the reaction volume
B. Removing $C_2H_4Cl_2(g)$ from the reaction mixture as it forms ✓
C. Adding $Cl_2(g)$ ✓
D. Lowering the reaction temperature ✓
~~E.~~ None of the choices above will increase the amount of $C_2H_4Cl_2(g)$ produced

(A) ↑ in volume ↓ Pressure ↓ no. of moles ∴ the reaction proceeds towards the side with greater number of moles (to the left) ∴ ↓ $C_2H_4Cl_2$ production

(B) Removing C_2H_4Cl (g) from the reaction ↓ its conc ∴ the reaction proceeds to the right to produce more C_2H_4Cl

(C) ↑ $Cl_2(g)$ would cause the reaction to proceed towards the right ↑ $C_2H_4Cl_2$ (g) production

(D) ↓ in reaction temperature cause the reaction to shift to the right to produce more heat ∴ ↑ $C_2H_4Cl_2$ production

7. You are given the information below for the equilibrium reaction between sulfur dioxide and oxygen to produce sulfur trioxide at 25 °C

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

	$2\,SO_2(g)$	$O_2(g)$	$2\,SO_3(g)$
ΔG°_f (kJ/mol)	-300.2	0.0	-371.1

If $SO_2(g)$ has a partial pressure is 1.0×10^{-4} atm, $O_2(g)$ is 0.20 atm, and $SO_3(g)$ is 0.10 atm, calculate $\mathbf{\Delta G_{rxn}}$. Keep your answer to three significant figures.

$\boxed{-104}$ kJ/mol

(1) Calculate ΔG°_{rxn}

$\Delta G^\circ_{rxn} = \Sigma n \Delta G^\circ_{products} - \Sigma n \Delta G^\circ_{reactants}$

$= (2 \times -371.1) - [(2 \times -300.2) + 0.0]$

$= -141.8$ kJ/mol

(2) Calculate Q_p

$$Q_p = \frac{(P_{SO_3})^2}{(P_{SO_2})^2 (O_2)} = \frac{(0.10)^2}{(1.0 \times 10^{-4})^2 (0.20)}$$

$= 5.0 \times 10^6$

3 $\Delta G_{rxn} = \Delta G^\circ_{rxn} + RT \ln Q$

$= \left(-141.8 \frac{kJ}{mol}\right) + \left[\left(8.314 \frac{J}{mol \cdot K} \times \frac{1 kJ}{1000 J}\right) \times 298 \times \ln 5.0 \times 10^6\right)\right]$

$= -104 \ kJ/mol$

8. The equilibrium mixture at 1000 K for the reaction $CO_2(g) + H_2(g) \rightleftharpoons CO(g) + H_2O(g)$ contains 0.276 atm $H_2(g)$, 0.276 atm $CO_2(g)$, 0.224 atm $CO(g)$, and 0.224 atm $H_2O(g)$. Determine:

A. The equilibrium constant, K_p. Keep your answer to three significant figures.

$\boxed{0.659}$

$K_p = \frac{(PCO)(PH_2O)}{(PCO_2)(PH_2)}$

$= \frac{(0.224)(0.224)}{(0.276)(0.276)} = 0.659$

B. ΔG°_{rxn} at 1000 K. Keep your answer to three significant figures.

$\boxed{3.47}$ kJ/mol

$\Delta G^\circ_{rxn} = -RT \ln K_p$

$= -8.314 \frac{J}{mol \cdot K} \times 1000 \times \ln 0.659$

$= 3.47 \times 10^3 \frac{J}{mol}$

$= 3.47 \frac{kJ}{mol}$

C. Does the reaction proceed towards products or reactants at 1000 K if a mixture contains 0.0750 atm $CO_2(g)$, 0.095 atm $H_2(g)$, 0.0340 atm CO(g), and 0.0650 atm $H_2O(g)$?

$\boxed{I}$

I. The reaction proceeds towards products
II. The reaction proceeds towards reactants
III. The reaction is at equilibrium

$$Q_p = \frac{(P_{CO})(P_{H_2O})}{(P_{CO_2})(P_{H_2O})} = \frac{(0.0340)(0.0650)}{(0.0750)(0.0950)}$$

$$= 0.310$$

$Q_p < K_p$ ∴ the reaction proceeds to the right (toward products)

9. Two equations can be written for the dissolution of $Mg(OH)_2(s)$ in acidic solution

Equation 1: $Mg(OH)_2(s) + 2\,H^+(aq) \rightleftharpoons Mg^{2+}(aq) + 2\,H_2O(l)$
Equation 2: $\frac{1}{2}\,Mg(OH)_2(s) + H^+(aq) \rightleftharpoons \frac{1}{2}\,Mg^{2+}(aq) + H_2O(l)$

Which of the statements below is/are correct regarding the relationship between the two equations? Select all that apply. Insert letters without spaces in the answer box, example **ABCD.**

$\boxed{AC}$

A. If $\Delta G°_{rxn}$ for equation 1 is –95.5 kJ/mol therefore $\Delta G°_{rxn}$ for equation 2 will be –47.8 kJ/mol
B. If $\Delta G°_{rxn}$ for equation 1 is –95.5 kJ/mol therefore $\Delta G°_{rxn}$ for equation 2 will be 9120 kJ/mol
C. K_1, the equilibrium constant for equation 1 is the square of K_2, the equilibrium constant for equation 2
D. K_1, the equilibrium constant for equation 1 is the square root of K_2, the equilibrium constant for equation 2
E. None of the above statements is/are correct

Equation 1: $Mg(OH)_2(s) + 2\,H^+(aq) \rightleftharpoons Mg^{2+}(aq) + 2\,H_2O(l)$ $\Delta G°_{rxn} = -95.5$ kJ/mol
Equation 2: $\frac{1}{2}\,Mg(OH)_2(s) + H^+(aq) \rightleftharpoons \frac{1}{2}\,Mg^{2+}(aq) + H_2O(l)$ $\Delta G°_{rxn} = -95.5 \text{ kJ/mol} \times \frac{1}{2}$
$= -47.8$ kJ/mol

↓ Equation 2 is equation 1 multiplied by $\frac{1}{2}$

Equation 1: $Mg(OH)_2(s) + 2\,H^+(aq) \rightleftharpoons Mg^{2+}(aq) + 2\,H_2O(l)$ K_1
Equation 2: $\frac{1}{2}\,Mg(OH)_2(s) + H^+(aq) \rightleftharpoons \frac{1}{2}\,Mg^{2+}(aq) + H_2O(l)$ K_2

$K_2 = (K_1)^{1/2}$
∴ $K_2 = \sqrt{K_1}$
∴ $K_1 = (K_2)^2$

10. One of the processes used in the production of chlorine gas is the Deacon process given by the equation below

$$4\,HCl(g) + O_2(g) \rightleftharpoons 2\,H_2O(g) + 2\,Cl_2(g) \quad \Delta_rH° = -114 \text{ kJ/mol}$$

If the reaction mixture is brought to equilibrium at 400 °C, which of the changes below will cause the system to increase the amount of $Cl_2(g)$? Select all that apply. Insert letters without spaces in the answer box, example **ABCD.**

$\boxed{ABD}$

A. Cooling the reaction below 400 °C
B. Additional moles of HCl gas added to the system
C. Increasing the volume of the system
D. Increasing the pressure of the system
E. Addition of a catalyst

Ⓐ ↓ in reaction temperature cause the reaction to shift to the right to produce more heat ∴ ↑ Cl_2 production

Ⓑ Addition of HCl ↓Q $\left(\downarrow Q_c = \frac{[H_2O]^2[Cl_2]}{\uparrow[HCl]^4[O_2]}\right)$, $Q_c < K_c$, therefore the reaction shifts to the right producing more Cl_2.

Ⓒ $PV = nRT$ $\quad \downarrow P = \frac{\downarrow n}{\uparrow V} RT$ → ↑ volume, ↓ pressure, ↓ no. of moles ∴ reaction shifts towards the side with more moles (left side ∴ ↓ Cl_2)

Ⓓ ↑ pressure ↑ no. of moles ∴ reaction shifts to the side with fewer moles (right side ∴ ↑ moles of Cl_2).

11. The free energy change of the reaction A(g) ⟶ B(g) is zero under certain conditions. The standard free energy of the reaction is -42.5 kJ. Which statement is true about the reaction? Select all that apply. Insert letters without spaces in the answer box, example **ABCD.** $\Delta G^\circ = -42.5$ kJ/mol

$\boxed{AB}$

$\Delta G = 0$

A. The concentration of the product is greater than the concentration of the reactant. $\Delta G^\circ_{rxn} < 0$ ∴ $K > 1$ & the reaction is product favored

B. The reaction is at equilibrium. $\Delta G_{rxn} = 0$

~~C.~~ The concentration of the reactant is greater than the concentration of the product. same as A

~~D.~~ None of the above statements is correct.

$$\Delta G = \Delta G^\circ_{rxn} + RT\ln Q$$
$$\Delta G^\circ_{rxn} + RT\ln Q = 0$$
$$\therefore \Delta G^\circ_{rxn} = -RT\ln Q$$

12. For the reaction $N_2O_4(g) \rightleftharpoons 2\ NO_2(g)$, ΔH° = +57.2 kJ/mol and K = 0.113 at 25 °C, what is the value of

A. K at 0 °C. Keep your answer to three significant figures and use scientific notation.

$\boxed{1.36}$ X 10 $\boxed{-2}$

heat + N2O4(g) ⇌ 2 NO2(g)

↓ temp of reaction, reaction shifts to the left

$\downarrow K = \frac{[\text{products}]}{[\text{reactants}]\uparrow}$

Using the Van't Hoff Equation:

$K_1 = ?$ $\quad T_1 = 0°C + 273.15 = 273 K$

$K_2 = 0.113$ $\quad T_2 = 25 + 273.15 = 298 K$

$\Delta H^\circ = 57.2 \frac{kJ}{mol} \times \frac{1000 J}{1 kJ} = 5.72 \times 10^4 \frac{J}{mol}$

$$\ln\left(\frac{K_2}{K_1}\right) = \frac{\Delta H^\circ}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$$\ln\left(\frac{0.113}{K_1}\right) = \frac{5.72 \times 10^4 \frac{J}{mol}}{8.314 \frac{J}{mol \cdot K}}\left(\frac{1}{273} - \frac{1}{298}\right)$$

$$\ln\left(\frac{0.113}{K_1}\right) = 2.114205052$$ → taking the inverse 'e' of both sides

$$\frac{0.113}{K_1} = e^{2.114205052}$$

$$\frac{0.113}{K_1} = 8.282998596$$

$\therefore K_1 = 0.013642$
$\approx 1.36 \times 10^{-2}$

B. At what temperature will K = 1.00? Keep your answer to three significant figures.

$\boxed{329}$ K

Using the Van't Hoff Equation:

$$\ln\left(\frac{K_2}{K_1}\right) = \frac{\Delta H^\circ}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$K_1 = 1.00$, $T_1 = ?$
$K_2 = 0.113$ $T_2 = 298$
$\Delta H = 5.72 \times 10^4 \frac{J}{mol}$
$R = 8.314 \frac{J}{mol \cdot K}$

$$\ln\left(\frac{0.113}{1.00}\right) = \frac{5.72 \times 10^4 \frac{J}{mol}}{8.314 \frac{J}{mol \cdot K}}\left(\frac{1}{T_1} - \frac{1}{298}\right)$$

$$-2.18036746 = 6.8799615 \times 10^3 \left(\frac{1}{T_1} - \frac{1}{298}\right)$$

$$-3.16915648 \times 10^{-4} = \frac{1}{T_1} - \frac{1}{298}$$

$$\frac{1}{T_1} = 3.03878906 \times 10^{-3}$$

$$\therefore T_1 = 329.078$$

$$\sim 329 \text{ K}$$