Recitation Worksheet 7: Chemical Equilibrium I (13.1 – 13.5)

Name: key

MyID:

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 17th at 11:59 pm. Write your work on separate sheets of paper, convert to a PDF and upload to the "Recitation Worksheet 7 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.
   c. If you have access to a printer, print out the worksheet, write your answer in the answer boxes, and show your work on it when appropriate. Then convert it to a PDF and upload to **Gradescope** by Friday, March 17th at 11:59 pm. You do not need to upload anything to eLC. The pages must be in the correct order and have the same layout as the original, or Gradescope will not be able to read it.
   d. There is a **Gradescope App** available for both iOS and Android devices that allows you to scan and submit your printed work or you can submit your fillable PDF directly. Detailed instructions on how to access and use the app can be found on your CHEM 1212 class eLC page under content ⟶ Welcome module ⟶ Gradescope ⟶ Gradescope new mobile app.
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. Use the equation below to answer the following questions:

$$Fe(OH)_3(s) + 3\,H^+(aq) \rightleftharpoons Fe^{3+}(aq) + 3\,H_2O(l) \quad K = 9.1 \times 10^3$$

solids and liquids are not included in the equilibrium expression

A. The correct equilibrium constant expression for this reaction is

$\boxed{iv}$

i. $\frac{[FeOH_3][H^+]^3}{[Fe^{3+}][H_2O]^3}$

ii. $\frac{[Fe^{3+}][H_2O]^3}{[FeOH_3][H^+]^3}$

iii. $\frac{[Fe^{3+}][H_2O]}{[FeOH_3][H^+]}$

iv. $\frac{[Fe^{3+}]}{[H^+]^3}$

v. $\frac{[FeOH_3][H^+]}{[Fe^{3+}][H_2O]}$

$K = \frac{[Fe^{3+}]}{[H^+]^3}$

B. What is the equilibrium concentration for $[Fe^{3+}]$ if the equilibrium concentration for $[H^+] = 2.5 \times 10^{-8}$? Keep your answer to two significant figures and use scientific notation to report your answer.

$\boxed{1.4} \times 10^{\boxed{-19}}$ M

$K = 9.1 \times 10^3$

$9.1 \times 10^3 = \frac{[Fe^{3+}]}{[2.5 \times 10^{-8}]^3}$

$[Fe^{3+}] = 1.421875 \times 10^{-19}$

$\sim 1.4 \times 10^{-19}$ M

2. Use the reaction below to answer the following questions

$$N_2(g) + O_2(g) + Cl_2(g) \rightleftharpoons 2\,NOCl(g) \quad K_p = ?$$

Altering K does not follow the same rules as Hess's law

A. Calculate K_p at 25 °C using the set of equations provided. Keep your answer to two significant figures and use scientific notation to report your answer.

$\frac{1}{2} N_2(g) + O_2(g) \rightleftharpoons NO_2(g)$	$K_{p_1} = 1.0 \times 10^{-9}$
$NOCl(g) + \frac{1}{2} O_2(g) \rightleftharpoons NO_2Cl(g)$	$K_{p_2} = 1.1 \times 10^{2}$
$NO_2(g) + \frac{1}{2} Cl_2(g) \rightleftharpoons NO_2Cl(g)$	$K_{p_3} = 3.0 \times 10^{-1}$

$\boxed{7.4} \times 10^{\boxed{-24}}$

multiply by 2

① $N_2(g) + \cancel{2O_2(g)} \rightleftharpoons \cancel{2\,NO_2(g)}$ ②

$1\,O_2(g)$

Kp_1 ← kp here will be squared

Kp_2 ← take the inverse of kp & then square kp $(1/kp)^2$

reverse equation & multiply by 2

② $\cancel{2\,NO_2Cl(g)} \rightleftharpoons 2NOCl(g) + \cancel{O_2(g)}$ ①

Kp_3 ← kp will be squared

multiply by 2

③ $\cancel{2\,NO_2(g)} + Cl_2(g) \rightleftharpoons \cancel{2\,NO_2Cl\,(g)}$ ③

Add the three equations up

instead of adding K from each equation they will be multiplied

$N_2(g) + O_2(g) + Cl_2(g) \rightleftharpoons 2\,NOCl(g)$

K_{tot} $Kp_1 \times Kp_2 \times Kp_3 =$

$(1.0 \times 10^{-9})^2 \times \left(\frac{1}{1.1 \times 10^2}\right)^2 \times (3.0 \times 10^{-1})^2$

$= 7.4 \times 10^{-24}$

B. Calculate K_c for the same reaction. Keep your answer to two significant figures and use scientific notation to report your answer.

$\boxed{1.8}$ X $10^{\boxed{-22}}$

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

$K_p = 7.4 \times 10^{-24}$

$R = 0.08204 \frac{L \cdot atm}{mol \cdot K}$

$T = 25 + 273.15 = 298K$

$\Delta n = 2 - 3 = 1$

$7.4 \times 10^{-24} = K_c\left(0.08206 \frac{L \cdot atm}{mol \cdot K} \times 298\right)^{(2-3)}$

$K_c = 1.81888364 \times 10^{-22}$

$\sim 1.8 \times 10^{-22}$

3. In which of the following reactions will $K_p = K_c$? Select all that apply. Insert letters without spaces in the answer box, example **ABCD**.

$\boxed{BD}$

~~A.~~ $4 KO_2(s) + 2 CO_2(g) \rightleftharpoons 2 K_2CO_3(s) + 3 O_2(g)$ $\Delta n = 3 - 2 = 1$

B. $3 Fe(s) + 4 H_2O(g) \rightleftharpoons Fe_3O_4(s) + 4 H_2(g)$ $\Delta n = 4 - 4 = 0$

~~C.~~ $2 H_2S(g) + CH_4(g) \rightleftharpoons CS_2(g) + 4 H_2(g)$ $\Delta n = 5 - 3 = 2$

D. $2 HI(g) \rightleftharpoons H_2(g) + I_2(g)$ $\Delta n = 2 - 2 = 0$

~~E.~~ $4 NH_3(g) + 5 O_2(g) \rightleftharpoons 4 NO(g) + 6 H_2O(g)$ $\Delta n = 10 - 9 = 1$

$K_p = K_c$ when $\Delta n = 0$

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

$K_p = K_c (RT)^0$

$\therefore K_p = K_c$

*4. For the reaction,

$$2 SO_2 (g) + O_2(g) \rightleftharpoons 2 SO_3(g) \quad K_c = 2.8 \times 10^2 \text{ at } 298 K$$

→ Do not forget to calculate the molarity of reactants & products

If a mixture contains 0.455 mol of SO_2, 0.183 mol of O_2, and 0.568 mol SO_3, are introduced into a 1.90 L vessel at 298 K. Which of the following statements is true?

$\boxed{C}$

A. The reaction is at equilibrium because $Q_c = K_c$

B. The reaction will proceed towards the reactants because $Q_c > K_c$

C. The reaction will proceed towards the products because $Q_c < K_c$

D. Not enough information is provided in the question

to determine if the reaction is at equilibrium you will need to calculate Q_c & compare it to K_c

- if $Q_c = K_c$ ∴ the reaction is at equilibrium
- if $Q_c < K_c$ ∴ the reaction proceeds towards the products (to the right)
- if $Q_c > K_c$ ∴ the reaction proceeds towards the reactants (to the left)

$[SO_2] = \frac{0.455}{1.90} = 0.23947868$

$[O_2] = \frac{0.183}{1.90} = 0.096315789$

$[SO_3] = \frac{0.568}{1.90} = 0.298947368$

$Q_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]} = \frac{[0.298947368]^2}{[0.23947868]^2[0.096315789]}$

$= 16.17992072$

$Q_c < K_c$ ∴ reaction proceeds towards products

5. You are given two elementary reactions below which are the reverse of one another and the reaction rates for elementary reaction are provided as well. Which of the following statements about an equilibrium mixture of X, Y, and Z is **false**?

At equilibrium

Rate of forward = Rate of reverse

$k_1[X][Y] = k_2[Z]^2$

$$X + Y \longrightarrow Z + Z \qquad \text{Rate} = k_1[X][Y]$$
$$Z + Z \longrightarrow X + Y \qquad \text{Rate} = k_2[Z]^2$$

$\boxed{D}$

~~A.~~ At equilibrium, k_1 and k_2 can have different values
~~B.~~ At equilibrium, $k_1[X][Y] = k_2[Z]^2$
~~C.~~ At equilibrium, X and Y still react to form Z
D. At equilibrium, the concentrations of X and Y must be the same
~~E.~~ At equilibrium, $\frac{[X][Y]}{[Z]^2}$ is a constant

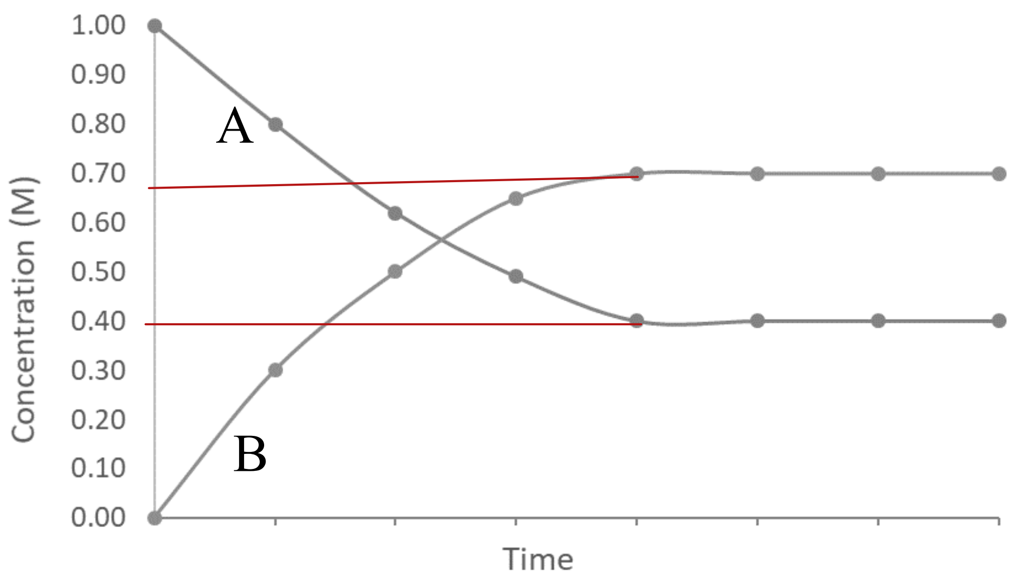
K1 & K2 can have different values but when multiplied by their respective conc. the rates will be equal to each other

the concentration of reactants is not equal at equilibrium

6. A graph for the reaction A(g) $\leftrightharpoons$ 2B(g) shows the change in concentration over time. What is the equilibrium constant for this reaction?

$\boxed{B}$

A. 1.8
B. 1.2
C. 4.4
D. 0.82
E. 0.57



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

$$= \frac{[0.70]^2}{[0.40]} = 1.225 \sim 1.2$$

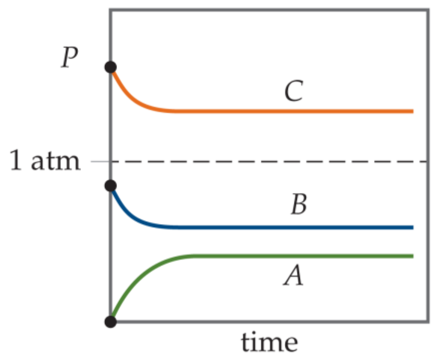
7. For the reaction A(g) + B(g) ⇌ C(g), $K_p < 1$. Which of the following charts below describes the approach to equilibrium of a mixture of B(g) and C(g)? → Since $K_p < 1$ this means that the reaction is reactant favored & ∴ C will produce A&B

the conc. of C decreases to form A&B

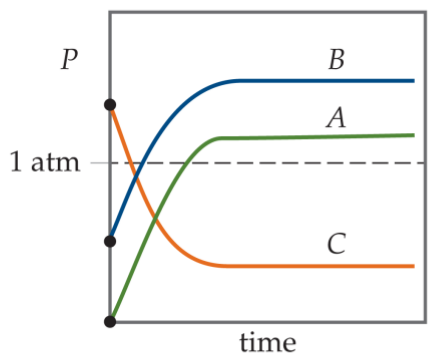
the conc. of A& B increases

B

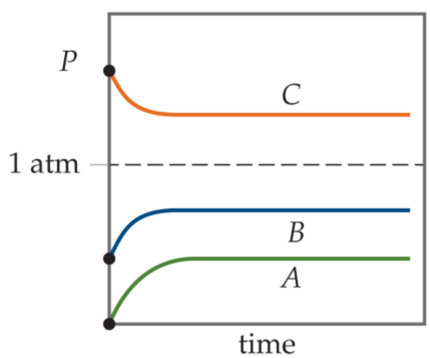
A.



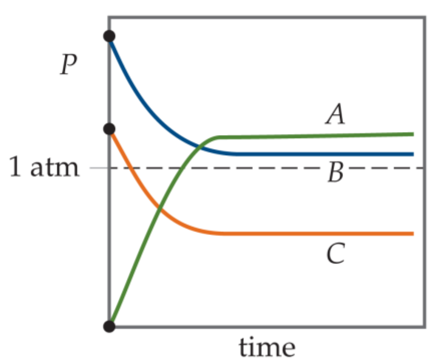
B.



C.



D.



→ in a saturated solution the dissolved solute is in dynamic equilibrium with the undissolved solute

8. You prepare a saturated solution of $PbBr_2$, and the equilibrium equation is written as follows:

$$PbBr_2(s) \rightleftharpoons Pb^{2+}(aq) + 2\ Br^{-}(aq)$$

If solid $PbBr_2$ is added to this solution, which of the following statements is **true**?

$\boxed{C}$

A. The concentration of Pb^{2+} will increase more than the concentration of Br-
B. The concentration of Br^- will increase more than the concentration of Pb^{2+}
C. The concentration of Pb^{2+} and Br^- will be the same as they were before more $PbBr_2$ was added
D. The concentration of Pb^{2+} and Br^- will increase after more $PbBr_2$ was added
E. The value of K_c will decrease

the addition of solute to a saturated solution will not increase the conc. of the ions in the solution

9. A mixture of 1.00 g H_2 and 1.06 g H_2S in a 0.500 L flask comes to equilibrium at 1670 K:

$$2\ H_2(g) + S_2(g) \rightleftharpoons 2\ H_2S(g)$$

The equilibrium amount of $S_2(g)$ found is 8.00 x 10^{-6} mol.

A. What is the equilibrium concentration of H_2S? Keep your answer to two significant figures and use scientific notation to report your answer.

$\boxed{6.2}$ X $10^{\boxed{-2}}$ M

$[H_2] = \frac{1.00 g \times \frac{1 mol}{2.02 g}}{0.500 L} = 0.990099$ M

$[H_2S] = \frac{1.06 g \times \frac{1 mol}{34.10 g}}{0.500 L} = 0.0621700 88$ M

the equilibrium concentration

$[S_2] = \frac{8.00 \times 10^{-6} mol}{0.500 L} = 1.60 \times 10^{-5}$ M

to determine the direction of the reaction we use Q & K

$Q = \frac{[H_2S]^2}{[H_2]^2[S_2]} = \frac{[0.0622]^2}{[0.990]^2[0]} = \infty$

∴ the equilibrium reaction proceeds towards products

	2 $H_2(g)$	+ $S_2(g)$	⇌ 2 $H_2S(g)$
I	0.990	0	0.0622
C	$(2\times +1.60\times10^{-5})$	$+1.60\times10^{-5}$	$(-1.60\times10^{-5})\times2$
E	0.990	$+1.6\times10^{-5}$	0.0622

B. What is the value of K_c? Keep your answer to three significant figures.

$\boxed{246}$

$$K_c = \frac{[H_2S]^2}{[H_2]^2[S_2]} = \frac{[0.0622]^2}{[0.990]^2[1.60 \times 10^{-5}]} = 246.4748 \sim 246$$

C. What is the value of K_p? Keep your answer to three significant figures.

$\boxed{1.80}$

$$K_p = K_c (RT)^{\Delta n}$$
$$= 246 (0.08206 \times 1670)^{2-3}$$
$$= 1.7950937$$
$$\sim 1.80$$

Substitute in $K = [0.280]^2[0.0140]$
$= 1.0987 \times 10^{-5}$
$\sim 1.1 \times 10^{-5}$ ~ the solution is at equilibrium ∴ saturated ∴ no more Ag_2SO_4 will dissolve

Molar mass of Ag_2SO_4 = 311.799 g/mol

10. You are a given a solution of silver sulfate and the ionization of silver sulfate in water can be represented by the equation below

$$Ag_2SO_4(s) \rightleftharpoons 2\ Ag^+(aq) + SO_4^{2-}(aq) \quad K = 1.1 \times 10^{-5} \text{ at } 298\ K$$

If the solution a 1.5 L solution contains 6.55 g Ag_2SO_4, and you attempt to dissolve additional solid silver sulfate in the solution will it dissolve?

$\boxed{B}$

A. Yes
B. No

$K = [Ag^+]^2[SO_4^{2-}]$

* In order to determined if additional $AgSO_4$ will dissolve, you will have to determine the conc of Ag^+ ions & SO_4^{2-} then substitute the conc in the equilibrium expression (to determine Q & compare to K)

* if the solution is at equilibrium determined by K then the solution is saturated & no additional $AgSO_4$ will dissolve (Q = K)

* if Q < K ∴ additional Ag_2SO_4 will dissolve

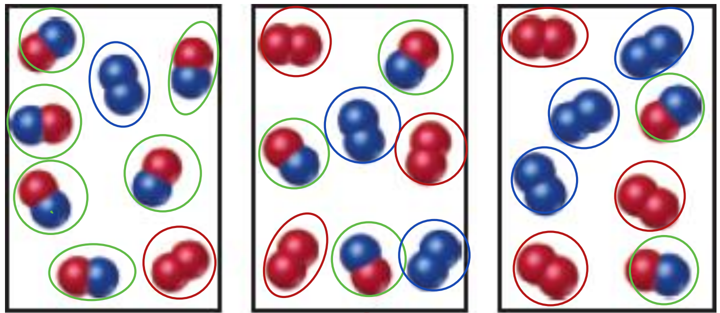
Conc of Ag^+ & SO_4^{2-} ions

$$\frac{6.55\cancel{g}\ Ag_2SO_4}{1.5L\ solution} \times \frac{1 mol\ Ag_2SO_4}{311.799\cancel{g}\ Ag_2SO_4} \times \frac{2\ Ag^+ \text{ ions}}{Ag_2SO_4} = 0.0280094 \sim 0.280\ M\ Ag^+$$

$$\frac{6.55\cancel{g}\ Ag_2SO_4}{1.5L\ solution} \times \frac{1 mol\ Ag_2SO_4}{311.799\cancel{g}\ Ag_2SO_4} \times \frac{SO_4^{2-} \text{ ions}}{Ag_2SO_4} = 0.0140047 \sim 0.0140\ SO_4^{2-} \text{ ions}$$

11. The diagram below represents the reaction between $A_2(g) + B_2(g) \rightleftharpoons 2\ AB(g)$, A_2 molecules are represented by red spheres, B_2 molecules are represented by blue spheres and the reaction has an equilibrium constant, $K_c = 1.5$. Which of the following diagrams a – c represents:

$\frac{[6]^2}{[1][1]} = 36$



a b c

$\frac{[2]^2}{[3][3]} = 0.44$

$\frac{[3]^2}{[3][2]} = 1.5$

A. The reaction at equilibrium.

$\boxed{b}$

$K_c = \frac{[AB]^2}{[A_2][B_2]}$

B. $Q_c < K_c$

$\boxed{c}$ $0.44 < 1.5$

C. $Q_c > K_c$

$\boxed{a}$ $36 > 1.5$

12. In the Ostwald process for oxidizing ammonia, a variety of products is possible – N_2, N_2O, NO, and NO_2 – depending on the conditions. One possibility is:

← keep the same

$$NH_3(g) + \frac{5}{4} O_2(g) \rightleftharpoons NO(g) + \frac{3}{2} H_2O\ (g)\quad K_{p_1} = 2.11 \times 10^{19} \text{ at } 700\text{ K}$$

For the decomposition of NO_2 at 700 K,

← flip $\frac{1}{Kp_2}$

$$NO_2(g) \rightleftharpoons NO(g) + \frac{1}{2} O_2(g)\quad K_{p_2} = 0.524 \text{ at } 700\text{ K}$$

What is K_p for the oxidation of $NH_3(g)$ to $NO_2(g)$? Keep your answer to three significant figures and use scientific notation to report your answer.

$\boxed{4.03}$ X $10^{\boxed{19}}$

$$NH_3(g) + \frac{5}{4} O_2(g) \rightleftharpoons \cancel{NO(g)} + \frac{3}{2} H_2O(g)$$

$$\cancel{NO(g)} + \frac{1}{2} O_2(g) \rightleftharpoons NO_2(g)$$

$$NH_3(g) + \frac{7}{4} (g) \rightleftharpoons NO_2 + \frac{3}{2} (H_2O)$$

$K_p = 3.893 \times 10^{19}$

$\sim 3.89 \times 10^{19}$

$\therefore Kp = Kp_1 \times \frac{1}{Kp_2}$

13. The decomposition of N_2O_4 to NO_2 is represented by the equilibrium reaction equation below

$$N_2O_4(g) \rightleftharpoons 2\ NO_2\ (g)$$

Initial concentration — change

If 1.75 moles of N_2O_4 is injected in a 1.5 L vessel and at equilibrium 25% N_2O_4 is dissociated, what is the equilibrium constant for this reaction? <u>Keep your answer to three significant figures.</u>

$\boxed{0.389}$

	$N_2O_4(g)$	$\rightleftharpoons 2\ NO_2\ (g)$
I	1.17	0
C	−0.292	+2(0.292)
E	0.875	0.583

- change in conc of $N_2O_4 = \frac{25}{100} \times 1.17 = 0.29167$
- $K = \frac{[NO_2]^2}{[N_2O_4]} = \frac{[0.583]^2}{[0.875]}$

$= 0.389$