

## Chemistry 30 – Equilibrium – Unit Homework

Topic	Textbook Reading	
Solutions and Molarity	Section 15.2	#14-20, 76-79
Reaction Rates and Collision Theory	Section 17.1	#4-9, 34, 40
Rate Determining Factors	Section 17.2	#11-13, 42-45, 47-49
Reversible Reactions	Section 18.1 (pp. 559-563)	
Equilibrium Constant	Section 18.1 (pp. 563-568)	
ICE Tables	Section 18.3 (pp. 575-576)	
Le Chatelier's Principle	Section 18.2	

### Reversible Reactions

1. What is chemical equilibrium?
2. Two colourless solutions are mixed in a stoppered flask. As the reaction proceeds, the resulting solution turns red, and a colourless gas is formed. After a few minutes, no more gas is evolved but the red colour remains. What two pieces of evidence are there that equilibrium has been established?
3. Sketch two concentration-time graphs: one for an irreversible reaction and one for a reversible reaction. How are they different?
4. Why can equilibrium not exist in an open system?

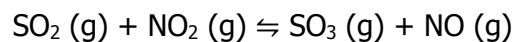
### Equilibrium Constant

5. Write equilibrium expressions for the following reversible reactions:
  - a.  $2 \text{NO}_2(\text{g}) \rightleftharpoons \text{N}_2\text{O}_4(\text{g})$
  - b.  $\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightleftharpoons 2 \text{NH}_3(\text{g})$
  - c.  $2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2 \text{SO}_3(\text{g})$
  - d.  $\text{Ca}(\text{NO}_3)_2(\text{s}) \rightleftharpoons \text{Ca}^{2+}(\text{aq}) + 2 \text{NO}_3^{-}(\text{aq})$
  - e.  $2 \text{HgO}(\text{s}) \rightleftharpoons 2 \text{Hg}(\text{l}) + \text{O}_2(\text{g})$
  - f.  $\text{NH}_4\text{Cl}(\text{s}) \rightleftharpoons \text{NH}_3(\text{g}) + \text{HCl}(\text{g})$
6. What condition(s) of a system must be changed to affect the value of the equilibrium constant of the system?
7. Why are solids and pure liquids left out of an equilibrium expression?
8. For the equilibrium system described by  $2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2 \text{SO}_3(\text{g})$ , the equilibrium concentrations of  $\text{SO}_2$ ,  $\text{O}_2$  and  $\text{SO}_3$  were 0.75 M, 0.30 M, and 0.15 M, respectively. Calculate the equilibrium constant,  $K_{\text{eq}}$ , for the reaction.
9.  $K_{\text{eq}} = 35$  at  $487^\circ\text{C}$  for the equilibrium system described by:

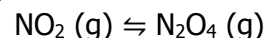


If the concentrations of the  $\text{PCl}_5$  and  $\text{PCl}_3$  are 0.015 M and 0.78 M, respectively, what is the concentration of the  $\text{Cl}_2$ ?

10. At a given temperature, analysis of an equilibrium mixture represent below is given as:  $[\text{SO}_2] = 4.0 \text{ M}$ ,  $[\text{NO}_2] = 0.50 \text{ M}$ ,  $[\text{SO}_3] = 3.0 \text{ M}$ ,  $[\text{NO}] = 2.0 \text{ M}$ . Find the value of  $K_{\text{eq}}$ , if the reaction is:

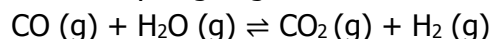


11. At  $55^\circ\text{C}$ ,  $K_{\text{eq}} = 1.15$  for the reaction:



- a. Write the equilibrium expression.
- b. Calculate the concentration of  $\text{N}_2\text{O}_4(\text{g})$  present in equilibrium with 0.50 mol/L of  $\text{NO}_2$ .

12. The following table gives some values for reactant and product equilibrium concentrations (in mol/L) at  $700^\circ\text{C}$  for the Shift reaction, an important method for the commercial production of hydrogen gas:

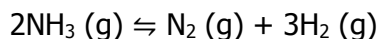


Trial	$[\text{CO}_2]$	$[\text{H}_2]$	$[\text{CO}]$	$[\text{H}_2\text{O}]$
1	0.600	0.600	0.266	0.266
2	0.600	0.800	0.330	0.286
3	2.00	2.00	0.877	0.877
4	1.00	1.50	0.450	0.655
5	1.80	2.00	0.590	1.20

Calculate  $K_{\text{eq}}$  for each of the five trials. Explain why the answers may be different for each trial.

## ICE Tables

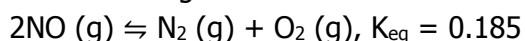
13. The decomposition of ammonia:



A vessel contains 1.00 mol/L of ammonia and is allowed to proceed to equilibrium. When the reaction stops, the concentration of ammonia is 0.87 mol/L.

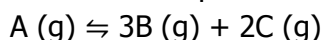
- What is the equilibrium concentration of hydrogen gas?
- What is the equilibrium constant for the reaction?

14. For the following reversible reaction:



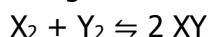
- Write the expression for the equilibrium constant of this reaction.
- Create an ICE table to show what will happen to the concentrations if 1.45 moles of **each** nitrogen gas and oxygen gas are introduced into a container that has a volume of 6.00 L. Don't solve!

15. A mixture of 3.31 mol of A, 4.33 mol of B and 5.95 mol of C is added into a 1.00-L container and allowed to reach equilibrium. At equilibrium, the number of moles of B is found to be 6.16. The reaction proceeds as follows:



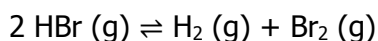
- What are the equilibrium concentrations of A and C?
- What is the equilibrium constant?

16. Gas  $\text{X}_2$  reacts with gas  $\text{Y}_2$ :



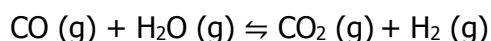
0.50 mole each of  $\text{X}_2$  and  $\text{Y}_2$  are placed in a 1.0 litre vessel and allowed to reach equilibrium at a given temperature. The equilibrium concentration of XY is found to be 0.025 mol/L. What is the equilibrium constant?

17. In the reversible reaction, where  $K_{\text{eq}} = 64$  at  $525^\circ\text{C}$ :



If 0.500 mol/L of HBr is put into a fixed volume container and allowed to reach equilibrium, what are the equilibrium concentrations of all of the species?

18. For the following reversible reaction, where  $K_{\text{eq}} = 10.0$  at  $600^\circ\text{C}$ :



Determine the equilibrium concentrations of all of the species if 0.100 mol/L of CO and  $\text{H}_2\text{O}$  are put into a fixed volume container and allowed to reach equilibrium.

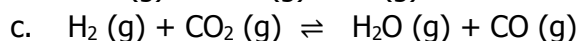
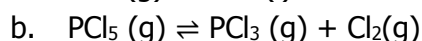
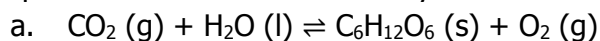
## Le Chatelier's Principle

19. What is Le Chatelier's Principle?

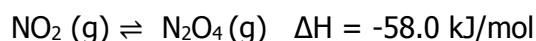
20. What factors can alter the equilibrium position in a chemical reaction?

21. Does changing the amount of a solid or liquid in a system have an effect on the equilibrium? Why or why not?

22. Balance each of the following reactions. Then, predict what will happen when the reaction pressure is increased in each system.



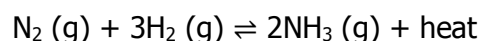
23. Balance the reaction below.



Predict the effect of each of the following changes on this system at equilibrium (drive forward reaction, drive reverse reaction, no effect).

- Add  $\text{N}_2\text{O}_4$
- Remove  $\text{NO}_2$
- Decrease the pressure
- Decrease the temperature

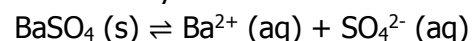
24. Consider the reaction:



Determine the effects of the following stresses on the equilibrium. State whether the equilibrium will shift to the left, to the right or not at all.

- Increase  $[\text{N}_2]$
- Decrease  $[\text{NH}_3]$
- Decrease  $[\text{H}_2]$
- Increase temperature
- Decrease total pressure
- Increase pressure

25. Consider the solution of barium sulfate, which has low solubility:



Determine how the equilibrium will shift in each situation:

- Add  $\text{Ba}^{2+}$  ions

b. Remove  $\text{SO}_4^{2-}$  ions

c. Add  $\text{BaSO}_4$  (s)

26. Consider the reaction:

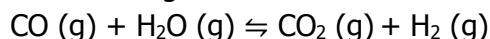


$$K_{\text{eq}} = 0.87 \text{ at } 55^\circ\text{C}$$

What is the effect of each of these changes upon the concentration of  $\text{N}_2\text{O}_4$  at equilibrium?

- Increasing the temperature
- Decreasing the pressure
- Adding more  $\text{NO}_2$  (g) to the system,
- Adding He gas to the container (fixed volume)

27. For the following reversible reaction:



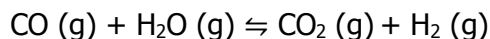
- Write the expression for the equilibrium constant of this reaction.
- Find the equilibrium constant for this expression. At  $588^\circ\text{C}$ , the equilibrium concentrations of the compounds are:

[CO]	2.00 mol/L	[CO <sub>2</sub> ]	3.52 mol/L
[H <sub>2</sub> O]	0.241 mol/L	[H <sub>2</sub> ]	4.30 mol/L

- If the concentration of water in the system is increased to 3.50 mol/L, which way will the equilibrium shift?
- If the equilibrium concentration of water is 1.58 M, what are the equilibrium concentrations of the other species? (The initial concentrations are as shown above, and 3.50 mol/L for water.)

### Reaction Quotient

28. For the reversible reaction:



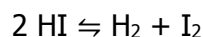
- Write the expression for the equilibrium constant of this reaction.
- Find the equilibrium constant for this expression. The equilibrium concentrations at  $861^\circ\text{C}$  are:

[CO]	2.00 mol/L
[H <sub>2</sub> O]	0.241 mol/L
[CO <sub>2</sub> ]	3.52 mol/L
[H <sub>2</sub> ]	4.30 mol/L

- If a 10.0 L vessel has 2.50 mol of CO and  $\text{H}_2\text{O}$  and 5.00 mol of  $\text{CO}_2$  and  $\text{H}_2$  at  $861^\circ\text{C}$ , which way will the reaction shift?

d. Using the values from part c, create an ICE table for this system, but do not solve it.

29. At  $520^\circ\text{C}$ ,  $K_{\text{eq}}$  is 0.0156 for:



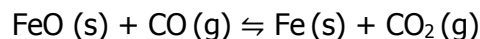
0.800 mol of each HI,  $\text{I}_2$  and  $\text{H}_2$  are placed in a 2.00 L container and allowed to come to equilibrium at  $520^\circ\text{C}$ . Which way will the reaction proceed to equilibrium?

30. For the reversible reaction:



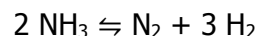
$K_{\text{eq}}$  is 0.190 at a certain temperature. If you combine 2.00 mol of each substance in a 1.00L flask, which way will the reaction proceed to reach equilibrium?

31. For the reversible reaction:



$K_{\text{eq}}$  is 0.560 at a certain temperature. If you start with 0.5 g of both FeO and Fe in a 5.00 L reaction vessel, then add 4.30 mol of CO and 2.12 mol of  $\text{CO}_2$ , which of the solids will you have more of when equilibrium is reached?

32. One mole of  $\text{NH}_3$  was injected into a 1 L flask at a certain temperature. The equilibrium mixture was then analyzed and found to contain 0.300 moles of  $\text{H}_2$ .



- Calculate the concentration of each species at equilibrium.
- Calculate the equilibrium constant for this system at this temperature and pressure.
- Which way would the equilibrium shift if 0.600 mol of  $\text{H}_{2(\text{g})}$  were added to the flask?
- How would the equilibrium constant be affected if the pressure of this system were suddenly increased?

33. Determine which way the reaction will shift given the initial concentrations.

- $\text{C}_2\text{H}_5\text{OH}(\text{aq}) + \text{CH}_3\text{COOH}(\text{aq}) \rightleftharpoons \text{CH}_3\text{COOC}_2\text{H}_5(\text{aq}) + \text{H}_2\text{O}(\text{g})$ ,  $K_{\text{eq}} = 2.30$   
[ $\text{C}_2\text{H}_5\text{OH}$ ] = 0.450 M, [ $\text{CH}_3\text{COOH}$ ] = 1.21 M,  
[ $\text{CH}_3\text{COOC}_2\text{H}_5$ ] = 0.986 M, [ $\text{H}_2\text{O}$ ] = 3.21 M
- $2\text{NH}_3(\text{g}) \rightleftharpoons \text{N}_2(\text{g}) + 3\text{H}_2(\text{g})$ ,  $K_{\text{eq}} = 2.63 \times 10^{-9}$   
[ $\text{NH}_3$ ] = 10.3 M, [ $\text{N}_2$ ] = 0.221 M, [ $\text{H}_2$ ] = 0.01 M