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. $2\ NO_2\ (g) \rightleftharpoons N_2O_4\ (g)$
   b. $N_2\ (g) + 3\ H_2\ (g) \rightleftharpoons 2\ NH_3\ (g)$
   c. $2\ SO_2\ (g) + O_2\ (g) \rightleftharpoons 2\ SO_3\ (g)$
   d. $Ca(NO_3)_2\ (s) \rightleftharpoons Ca^{2+}\ (aq) + 2\ NO_3^-\ (aq)$
   e. $2\ HgO\ (s) \rightleftharpoons 2\ Hg\ (l) + O_2\ (g)$
   f. $NH_4Cl\ (s) \rightleftharpoons NH_3\ (g) + HCl\ (g)$
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 $2\ SO_2\ (g) + O_2\ (g) \rightleftharpoons 2\ SO_3\ (g)$, the equilibrium concentrations of SO_2, O_2 and SO_3 were 0.75 M, 0.30 M, and 0.15 M, respectively. Calculate the equilibrium constant, K_{eq}, for the reaction.
9. K_{eq} = 35 at 487°C for the equilibrium system described by:

   $$PCl_5\ (g) \rightleftharpoons PCl_3\ (g) + Cl_2\ (g)$$

   If the concentrations of the PCl_5 and PCl_3 are 0.015 M and 0.78 M, respectively, what is the concentration of the Cl_2?
10. At a given temperature, analysis of an equilibrium mixture represent below is given as: $[SO_2]$ = 4.0 M, $[NO_2]$ = 0.50 M, $[SO_3]$ = 3.0 M, $[NO]$ = 2.0 M Find the value of K_{eq}, if the reaction is:

    $$SO_2\ (g) + NO_2\ (g) \leftrightharpoons SO_3\ (g) + NO\ (g)$$
11. At 55°C, K_{eq} = 1.15 for the reaction:

    $$NO_2\ (g) \leftrightharpoons N_2O_4\ (g)$$
    a. Write the equilibrium expression.
    b. Calculate the concentration of $N_2O_4(g)$ present in equilibrium with 0.50 mol/L of NO_2.
12. The following table gives some values for reactant and product equilibrium concentrations (in mol/L) at 700°C for the Shift reaction, an important method for the commercial production of hydrogen gas:

    $$CO\ (g) + H_2O\ (g) \rightleftharpoons CO_2\ (g) + H_2\ (g)$$

Trial	$[CO_2]$	$[H_2]$	$[CO]$	$[H_2O]$
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_{eq} for each of the five trials. Explain why the answers may be different for each trial.

ICE Tables

13. The decomposition of ammonia:

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

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.

a. What is the equilibrium concentration of hydrogen gas?

b. What is the equilibrium constant for the reaction?

14. For the following reversible reaction:

$$2NO\ (g) \leftrightharpoons N_2\ (g) + O_2\ (g),\ K_{eq} = 0.185$$

a. Write the expression for the equilibrium constant of this reaction.

b. 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:

$$A\ (g) \leftrightharpoons 3B\ (g) + 2C\ (g)$$

a. What are the equilibrium concentrations of A and C?

b. What is the equilibrium constant?

16. Gas X_2 reacts with gas Y_2:

$$X_2 + Y_2 \leftrightharpoons 2\ XY$$

0.50 mole each of X_2 and 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_{eq}= 64 at 525°C:

$$2\ HBr\ (g) \rightleftharpoons H_2\ (g) + Br_2\ (g)$$

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_{eq} = 10.0 at 600°C:

$$CO\ (g) + H_2O\ (g) \leftrightharpoons CO_2\ (g) + H_2\ (g)$$

Determine the equilibrium concentrations of all of the species if 0.100 mol/L of CO and H_2O 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.

a. $CO_2\ (g) + H_2O\ (l) \rightleftharpoons C_6H_{12}O_6\ (s) + O_2\ (g)$

b. $PCl_5\ (g) \rightleftharpoons PCl_3\ (g) + Cl_2(g)$

c. $H_2\ (g) + CO_2\ (g) \rightleftharpoons H_2O\ (g) + CO\ (g)$

23. Balance the reaction below.

$$NO_2\ (g) \rightleftharpoons N_2O_4\ (g) \quad \Delta H = -58.0\ kJ/mol$$

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

a. Add N_2O_4

b. Remove NO_2

c. Decrease the pressure

d. Decrease the temperature

24. Consider the reaction:

$$N_2\ (g) + 3H_2\ (g) \rightleftharpoons 2NH_3\ (g) + heat$$

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.

a. Increase $[N_2]$

b. Decrease $[NH_3]$

c. Decrease $[H_2]$

d. Increase temperature

e. Decrease total pressure

f. Increase pressure

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

$$BaSO_4\ (s) \rightleftharpoons Ba^{2+}\ (aq) + SO_4^{2-}\ (aq)$$

Determine how the equilibrium will shift in each situation:

a. Add Ba^{2+} ions

b. Remove SO_4^{2-} ions
c. Add $BaSO_4$ (s)

26. Consider the reaction:

$$N_2O_4 (g) + \text{heat} \leftrightharpoons 2\ NO_2 (g)$$

K_{eq} = 0.87 at 55°C

What is the effect of each of these changes upon the concentration of N_2O_4 at equilibrium?

a. Increasing the temperature
b. Decreasing the pressure
c. Adding more NO_2 (g) to the system,
d. Adding He gas to the container (fixed volume)

27. For the following reversible reaction:

$$CO (g) + H_2O (g) \leftrightharpoons CO_2 (g) + H_2 (g)$$

a. Write the expression for the equilibrium constant of this reaction.
b. Find the equilibrium constant for this expression. At 588°C, the equilibrium concentrations of the compounds are:

[CO]	2.00 mol/L	[CO_2]	3.52 mol/L
[H_2O]	0.241 mol/L	[H_2]	4.30 mol/L

c. If the concentration of water in the system is increased to 3.50 mol/L, which way will the equilibrium shift?
d. 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:

$$CO (g) + H_2O (g) \leftrightharpoons CO_2 (g) + H_2 (g)$$

a. Write the expression for the equilibrium constant of this reaction.
b. Find the equilibrium constant for this expression. The equilibrium concentrations at 861°C are:

[CO]	2.00 mol/L
[H_2O]	0.241 mol/L
[CO_2]	3.52 mol/L
[H_2]	4.30 mol/L

c. If a 10.0 L vessel has 2.50 mol of CO and H_2O and 5.00 mol of CO_2 and H_2 at 861°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°C, K_{eq} is 0.0156 for:

$$2\ HI \leftrightharpoons H_2 + I_2$$

0.800 mol of each HI, I_2 and H_2 are placed in a 2.00 L container and allowed to come to equilibrium at 520°C. Which way will the reaction proceed to equilibrium?

30. For the reversible reaction:

$$COBr_2 (g) \leftrightharpoons CO (g) + Br_2 (g)$$

K_{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:

$$FeO (s) + CO (g) \leftrightharpoons Fe (s) + CO_2 (g)$$

K_{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 CO_2, which of the solids will you have more of when equilibrium is reached?

32. One mole of 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 H_2.

$$2\ NH_3 \leftrightharpoons N_2 + 3\ H_2$$

a. Calculate the concentration of each species at equilibrium.
b. Calculate the equilibrium constant for this system at this temperature and pressure.
c. Which way would the equilibrium shift if 0.600 mol of $H_{2(g)}$ were added to the flask?
d. 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.

a. C_2H_5OH (aq) + CH_3COOH (aq) $\leftrightharpoons$ $CH_3COOC_2H_5$ (aq) + H_2O (g), K_{eq} = 2.30
[C_2H_5OH] = 0.450 M, [CH_3COOH] = 1.21 M, [$CH_3COOC_2H_5$] = 0.986 M, [H_2O] = 3.21 M
b. $2NH_3$ (g) $\leftrightharpoons$ N_2 (g) + $3H_2$ (g), $K_{eq} = 2.63 \times 10^{-9}$
[NH_3] = 10.3 M, [N_2] = 0.221 M, [H_2] = 0.01 M