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AP Chemistry Unit 7 Study Guide

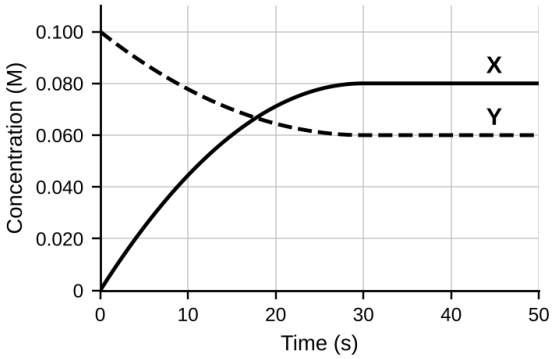
Equilibrium (Topics 7.1 to 7.12)

Show all work with units and correct significant figures. Leave pure solids and liquids out of every Q and K expression. When you explain a shift, compare Q to K.

Equations	$K_c = [C]^c[D]^d / ([A]^a[B]^b)$	$K_p = (P_C)^c(P_D)^d / ((P_A)^a(P_B)^b)$	Q < K: net forward Q > K: net reverse	$K_{sp} = [M^{y+}]^x[X^{x-}]^y$ for $M_xX_y(s)$
	Reverse a reaction: $K_{new} = 1/K$	Multiply coefficients by n: $K_{new} = K^n$	Add reactions: $K_{overall} = K_1 \times K_2$	Only temperature changes the value of K

Part 1. Reversible Reactions and Dynamic Equilibrium (Topics 7.1 to 7.2)

1. A sealed flask starts with 0.100 M $N_2O_4(g)$ and no $NO_2(g)$. The reaction $N_2O_4(g) \rightleftharpoons 2NO_2(g)$ is followed over time. Curves X and Y are shown.

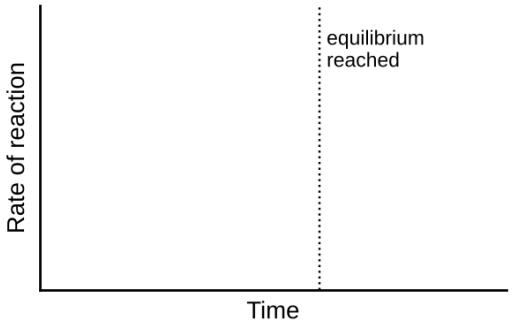


(a) Which curve, X or Y, represents NO_2? Justify your answer with the change in each concentration.

(b) At what time is equilibrium first reached? Describe the forward and reverse rates at 40 s.

(c) Calculate K_c for this reaction.

2. On the axes, sketch the rate of the forward reaction and the rate of the reverse reaction for the flask in Question 1, from the moment N_2O_4 is added until after equilibrium. Label each curve.



3. At one moment during the reaction A(g) ⇌ B(g), the forward rate is 3.2×10^{-3} M/s and the reverse rate is 1.1×10^{-3} M/s. Is the system at equilibrium? In which direction is there a net change, and is Q greater than or less than K? Explain.

4. Decide whether each statement is true or false. Rewrite each false statement so it is true.

Statement	T or F	Correction (if false)
At equilibrium, the concentrations of reactants and products are equal.		
At equilibrium, the forward and reverse reactions stop.		
In a sealed bottle of water at constant temperature, evaporation stops once the vapor pressure stays constant.		
Equilibrium can be reached starting from only reactants or from only products.		

Part 2. Writing Q and K Expressions (Topic 7.3)

5. Write the equilibrium expression asked for in each row. Then explain why $CaCO_3(s)$ and $CaO(s)$ do not appear in the second expression, even though both must be present for the system to be at equilibrium.

Reaction	Write	Expression
$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$	K_p	
$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$	K_p	
$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$	K_c	
$Ag_2CrO_4(s) \rightleftharpoons 2Ag^+(aq) + CrO_4^{2-}(aq)$	K_{sp}	

Part 3. Calculating K and Q (Topics 7.4, 7.5, and 7.7)

6. For $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ at 430 °C, an equilibrium mixture contains $[H_2] = 0.0110$ M, $[I_2] = 0.0110$ M, and $[HI] = 0.0786$ M. Calculate K_c.

7. For $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$, the equilibrium partial pressures are $P_{PCl5} = 0.875$ atm, $P_{PCl3} = 0.463$ atm, and $P_{Cl2} = 0.463$ atm. Calculate K_p.

8. Use $K_c = 51.1$ for $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ from Question 6. A flask at the same temperature contains $[H_2] = 0.020$ M, $[I_2] = 0.020$ M, and $[HI] = 0.200$ M. Calculate Q, then state which direction the reaction proceeds and whether [HI] increases or decreases.

9. For each reaction, describe what the equilibrium mixture mostly contains.

Reaction	K	Mostly reactants, mostly products, or large amounts of both?
$2H_2(g) + O_2(g) \rightleftharpoons 2H_2O(g)$	3×10^{81}	
$N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$ at 25 °C	4×10^{-31}	
$CH_3COOH(aq) + CH_3OH(aq) \rightleftharpoons CH_3COOCH_3(aq) + H_2O(l)$	4	

Part 4. Manipulating K (Topic 7.6)

10. For $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$, $K_c = 0.50$ at a certain temperature. Find K_c for each reaction at the same temperature.

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

(b) $\frac{1}{2}N_2(g) + (3/2)H_2(g) \rightleftharpoons NH_3(g)$

(c) $4NH_3(g) \rightleftharpoons 2N_2(g) + 6H_2(g)$

11. Use the two reactions below, measured at the same high temperature.

(1) $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$ $K_1 = 2.5 \times 10^{-3}$

(2) $2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$ $K_2 = 1.6 \times 10^{2}$

(a) Calculate K for $N_2(g) + 2O_2(g) \rightleftharpoons 2NO_2(g)$.

(b) Calculate K for $NO_2(g) \rightleftharpoons \frac{1}{2}N_2(g) + O_2(g)$.

Part 5. Finding Equilibrium Concentrations (Topic 7.7)

12. For $CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$, $K_c = 4.00$ at a certain temperature. A flask starts with 0.200 M CO and 0.200 M H_2O and no products. Complete the ICE table and find every equilibrium concentration.

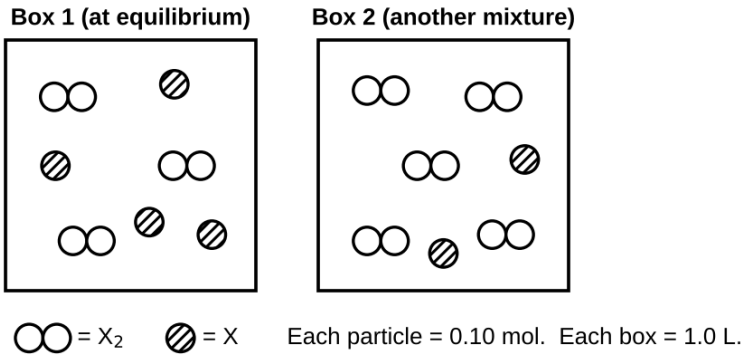
	CO	H_2O	CO_2	H_2
Initial (M)				
Change (M)				
Equilibrium (M)				

13. At 1500 K, $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$ has $K_c = 1.0 \times 10^{-5}$. A flask starts with 0.80 M N_2 and 0.20 M O_2. Calculate [NO] at equilibrium. Explain why you can treat x as small.

14. A rigid container at constant temperature is filled with NOCl(g) at 1.00 atm. When $2NOCl(g) \rightleftharpoons 2NO(g) + Cl_2(g)$ reaches equilibrium, $P_{Cl2} = 0.120$ atm. Find the other equilibrium pressures and calculate K_p.

Part 6. Particle Representations of Equilibrium (Topic 7.8)

15. The diagrams show the reaction $X_2(g) \rightleftharpoons 2X(g)$ at one temperature.



(a) Use Box 1 to calculate K_c.

(b) Is Box 2 at equilibrium? If not, which way will it shift? Support your answer with Q.

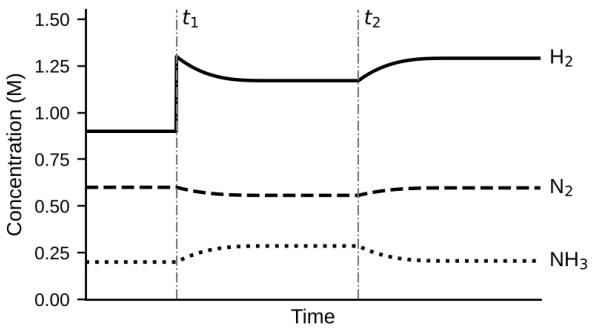
Part 7. Le Châtelier's Principle (Topics 7.9 to 7.10)

16. Consider $N_2O_4(g) \rightleftharpoons 2NO_2(g)$, $\Delta H° = +57\ kJ/mol_{rxn}$. Complete the table for each change made to a mixture at equilibrium.

Change	Shift (left, right, or none)	Value of K (increases, decreases, or no change)
Add more N_2O_4		
Remove some NO_2		
Raise the temperature		
Decrease the volume at constant temperature		
Add argon gas at constant volume		

17. For the same mixture, the volume is suddenly cut in half at constant temperature, so every concentration doubles. Show what happens to Q_c, and use it to explain the direction of the shift.

18. The graph shows $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ ($\Delta H° < 0$) as two different changes are made at t_1 and t_2.



(a) What change was made at t_1? Explain why $[N_2]$ decreases after t_1.

(b) At t_2 the temperature was changed. Was it raised or lowered? Justify your answer.

19. A solution contains the equilibrium $CoCl_4^{2-}(aq) + 6H_2O(l) \rightleftharpoons Co(H_2O)_6^{2+}(aq) + 4Cl^-(aq)$. $CoCl_4^{2-}$ is blue and $Co(H_2O)_6^{2+}$ is pink. Circle the color change for each action, then answer the question.

(a) Concentrated HCl is added (more Cl^-). more blue / more pink

(b) $AgNO_3$ is added, which removes Cl^- as AgCl(s). more blue / more pink

(c) Water is added until the volume doubles. more blue / more pink

(d) The solution turns blue when it is heated. Is the forward reaction endothermic or exothermic? Explain.

Part 8. Solubility Equilibria and the Common-Ion Effect (Topics 7.11 to 7.12)

20. The molar solubility of PbI_2 in pure water is 1.3×10^{-3} M. Calculate K_{sp} for PbI_2.

21. AgCl has $K_{sp} = 1.8 \times 10^{-10}$ and Ag_2CrO_4 has $K_{sp} = 1.1 \times 10^{-12}$.

(a) Calculate the molar solubility of each salt in pure water.

(b) A student says AgCl is more soluble because its K_{sp} is larger. Do you agree? Explain.

22. Equal volumes of 2.0×10^{-4} M $AgNO_3$ and 2.0×10^{-4} M NaCl are mixed. Will AgCl precipitate? ($K_{sp} = 1.8 \times 10^{-10}$) Show your reasoning with Q.

23. Calculate the molar solubility of AgCl in 0.10 M NaCl. How does it compare with its solubility in pure water? Explain the difference using Le Châtelier's principle.

24. The molar solubility of CaF_2 in 0.020 M NaF is measured as 9.8×10^{-8} M. Calculate K_{sp} for CaF_2.

25. A student finds K_{sp} for $Ca(OH)_2$ by titrating the OH^- in a saturated solution, then sets $s = [OH^-]/2$ and uses $K_{sp} = 4s^3$. By mistake, the student made the saturated solution with 0.010 M $CaCl_2$ instead of pure water. Will the calculated K_{sp} be too high, too low, or correct? Explain.