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7.4 Qualitative Changes in Equilibrium Systems Part 1

Le Chatelier's Principle is an important idea! It is important!!!

- Define what the Le Chatelier's principle.(p. 439)
- Define an equilibrium shift (p. 439)

ANALAYZE: Figure 2 (p. 440): Analyze the graph here are the steps:

ANALYZING GRAPHS

a. Look at the balanced equation
b. Show what is happening to each of the reactants and products during the reaction process.
c. This means- during the first third the reactant decrease for Fe and FeSCN
d. During the 2nd part, what is the pattern?
e. During the 3rd part, what is added and what happens to the equilibrium.

Analyze Figure 4 (p. 441). What happens when carbon dioxide is added? Analyze what happened when Carbon monoxide is added?

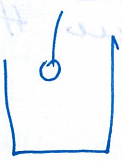
Collision Theory and Concentration Changes in an Equilibrium System, p. 441.

- Explain and apply how the collision theory explains the shifts in equilibrium.

Real Life Application, P. 441-442.

- Explain aqueous nitric acids are used and in what everyday products will you find these products.
- Explain the process of adding carbon monoxide gas in an equilibrium reaction. (p. 442)
- Explain in biological processes of hemoglobin in the blood (p. 442)

add more CO_2 – more reactions therefore the whole system shifts to the right because more collisions take place.

Collision Theory:  $CO_{2(g)} \rightleftharpoons CO + O_2$.

However, once too much product is formed the reaction reverses.

$3\ NO_2 + H_2O(l) \rightleftharpoons 2\ HNO_{3(aq)} + NO(g).$

nitric Oxide. not as (useful).

nitric acid.
- used in synthetic fertilizers.
explosives.
dyes.
perfumes.

---

* Why this matters? In this case a less useful NO_2 is reacted with a natural and abundant substance water to create a much more useful product, The Equilibrium constantly is manipulated to create the desired substance, HNO_3.

Another equilibrium is

$C_s + H_2O \rightleftharpoons H_2(g) + CO(g)$ -

hydrogen gas is used as fuel.

equilibrium used for energy uses.

Like C, CO(g) can be manipulated to create more H_2.

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

used as fuel.

The equilibrium shifts so more H_2 is created and used as fuel.

$Hb(aq) + O_2(g) \rightleftharpoons HbO_2(aq)$

hemoglobin. oxygen from lungs.

constant when O_2 is available. more Hemoglobin-oxygen complex, when less O_2, therefore it is realeased

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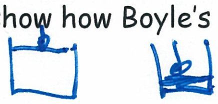
7.4 Part 2 Le Chatelier's Principle and Changes in Energy

Endothermic Reaction	Exothermic Reaction
Provide an example $N_2O_{4(g)} + energy \rightleftharpoons 2NO_{2(g)}$	Provide an example $2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_3 + energy$.
What happens when energy is added? Which direction will the system shift? energy added Shift to the right $\rightarrow$.	What happens when energy is added? energy removed. Shift to the right.

Le Chatelier's Principle and changes in Gas Volume (p. 443)

Define the following:

- ideal gas (p.443)
- partial pressure (p. 443)
  Use a diagram to show how Boyle's applies to molecules and which direction will it move (p. 444)  Shifts to the right.

Changing an Equilibrium System **without affecting** Equilibrium Position (p. 444)

Methods	Explanation
Catalysts lower activation energy.	Rates of reaction both reverse and forward are faster.
Inert Gas (does not react.	as long as forward + reverse reaction is unchanged.
State of Reactants	adding more solid does not change equilibrium.

Do and apply, P. 446, # 1 and 2.

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7.5 Calculating Equilibrium Concentrations
Review: concentration c= n/V= mol/L
n =
V= units

Calculating concentrations if you have a volume of 2.00L,.all are in gas state.
Given a PCl_5 = 0.00870 mol
Given a PCl_3= 0.298 mol
Given a Cl_2 = 0.00

Change is
PCl_5 is -0.001 mol/L
PCl_3 is +0.001 mol/L
Cl_2 is +0.001 mol/L
What is the balanced equation and create and solve the ICE table and determine the K value for this reaction.

Do and Apply, p. 454, Practice # 1 and 2.

7.6 Determine Ksp
Practice: Write a solubility product constant for the solid of CaF_2 and 2 F ions and Ca ion both in aqueous state.

Do and apply, p. 462.

If you finish you can start working on section 8.2 calculations, p.
Define the following:
Strong acid (p. 495)
Weak acid (p. 495)
Analyze table 2; Compare value of Ka for each one, position of ionization equilibrium, and concentration of HNO_3 Acid

Determine the Ka for $HClO_4$ into H and Cl04:

Define oxyacid (p. 496); organic acid (p. 497), strong base (p. 497) weak base (498); base ionization (Kb) determine the Kb for NH3 and water into a hydroxide ion and NH_4^+ ammonium ion.
Define autoionization of water (p. 499); ion-product constant for water (Kw)

7.4 Part 2.

Section 7.4: Qualitative Changes in Equilibrium Systems

Section 7.4 Questions, page 446

SG: ✓

1. (a) If the volume of the container is decreased, the pressure will increase, so the equilibrium is likely to shift toward the right because there are fewer product entities.
(b) If the temperature of the container is increased, the additional energy is likely to be absorbed by a shift of the equilibrium toward the left and increase the concentration of the reactants.
(c) As ethane is removed, the equilibrium will shift toward the right because there are fewer product entities.
(d) As hydrogen is added, the equilibrium will shift toward the right, reducing the amount of reactants in the mixture.

SG: ✓

2. (a) Prediction: As the concentration of chloride ions is increased, the equilibrium will shift toward the product, $CuCl_4^{2-}$(aq).
(b) Increasing the concentration of chloride ions caused the colour to change from blue toward green, which indicates that the concentration of $CuCl_4^{2-}$(aq) increased due to an equilibrium shift toward the reaction product.
(c) Independent variable: [Cl^-(aq)]; dependent variable: [$CuCl_4^{2-}$(aq)]; controls: total volume, mixing procedure.

3. If the temperature is increased, the smell of ammonia will increase. Because the reaction is endothermic, addition of thermal energy will shift the equilibrium to the right and increase the concentration of the product, ammonia.

4. (a) For an exothermic reaction, the equilibrium will shift toward the reactants, so the value of *K* will decrease.
(b) For an endothermic reaction, the equilibrium will shift toward the products, so the value of *K* will increase.
(c) If the value of *K* increases, the reaction is exothermic. An increase in *K* indicates that the equilibrium has shifted toward the products. An exothermic reaction shifts toward the products when thermal energy is removed.

Section 7.5: Quantitative Changes in Equilibrium Systems

Tutorial 1 Practice, page 452

/ **1. (a) Given:** $[HI(g)] = 0.14$ mol/L; $[H_2(g)] = 0.040$ mol/L; $[I_2(g)] = 0.010$ mol/L; $K = 0.020$

Required: Q; direction of reaction to reach equilibrium

SG

Solution: $$Q = \frac{[H_2(g)][I_2(g)]}{[HI(g)]^2} = \frac{(0.040)(0.010)}{(0.14)^2}$$

$$Q = 0.020$$

Statement: The value of Q is 0.020. Q is equal to K, so the reaction is at equilibrium.

(b) Given: $[HI(g)] = 0.20$ mol/L; $[H_2(g)] = 0.15$ mol/L; $[I_2(g)] = 0.090$ mol/L; $K = 0.020$

Required: Q; direction of reaction to reach equilibrium

Solution: $$Q = \frac{[H_2(g)][I_2(g)]}{[HI(g)]^2} = \frac{(0.15)(0.090)}{(0.20)^2}$$

$$Q = 0.34$$

Statement: The value of Q is 0.34. Q is greater than K, so the product concentrations are greater than at equilibrium; the reaction will shift toward the left, more HI.

/ **2.** Time 1:

SG

Given: $[N_2O_4(g)] = 0.80$ mol/L; $[NO_2(g)] = 1.55$ mol/L; $K = 0.87$

Required: Q; direction of reaction to reach equilibrium

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

$$Q = \frac{[NO_2(g)]^2}{[N_2O_4(g)]} = \frac{(1.55)^2}{0.80}$$

$$Q = 3.0$$

Statement: The value of Q is 3.0. Q is greater than K, so the product concentrations are greater than at equilibrium; the reaction will shift toward the left, more N_2O_4.

Time 2:

Given: $[N_2O_4(g)] = 1.66$ mol/L; $[NO_2(g)] = 1.20$ mol/L; $K = 0.87$

Required: Q; direction of reaction to reach equilibrium

Solution: $$Q = \frac{[NO_2(g)]^2}{[N_2O_4(g)]} = \frac{(1.20)^2}{1.66}$$

$$Q = 0.867$$

Statement: The value of Q is 0.867. Q is approximately equal to K, so the reaction is at equilibrium.

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3. (a) The equilibrium constant equation for the reaction represented by

$$2\ Ag_2O(s) \rightleftharpoons 4\ Ag(s) + O_2(g)$$

is $K = [O_2(g)]$, because the other substances are solids.

(b) Given: $K = 2.5 \times 10^{-3}$; $[O_2(g)]_{instantaneous} = 5.0 \times 10^{-2}$

Required: Q; direction of reaction to reach equilibrium

Solution: $Q = [O_2(g)]$

$$Q = 5.0 \times 10^{-2}$$

Statement: The value of Q is greater than K, so the reaction will shift to the left, toward the reactant.

Tutorial 2 Practice, page 454

1. Given: Volume, $V = 250$ mL $= 0.25$ L; $n_{\text{initial } I_2(g)} = 0.50$ mol; $n_{\text{initial } Br_2(g)} = 0.50$ mol; $K = 1.2 \times 10^2$

Required: $[I_2(g)]_{equilibrium}$; $[Br_2(g)]_{equilibrium}$

Analysis: $c = \dfrac{n}{V}$

Solution:

Step 1. Calculate concentrations, c, for in mol/L from the given amounts of all entities.

Calculate $[I_2(g)]_{initial}$:

$$c = \frac{n_{initial}}{V}$$

$$= \frac{0.50 \text{ mol}}{0.25 \text{ L}}$$

$$c = 2.0 \text{ mol/L}$$

Using the same formula,

$[Br_2(g)]_{initial} = 2.0$ mol/L

Since there is no iodine monobromide gas initially, $[IBr(g)]_{initial}$ is 0.0 mol/L.

Step 2. Write the balanced equation for the equilibrium reaction system.

$$I_2(g) + Br_2(g) \rightleftharpoons 2\ IBr(g)$$

Step 3. Determine the equilibrium law equation.

$$K = \frac{[IBr(g)]^2}{[I_2(g)][Br_2(g)]}$$

Step 4. Use an ICE table to determine the relationship between the equilibrium concentrations of the reactant and the products.

	$I_2(g)$ +	$Br_2(g)$ $\rightleftharpoons$	$2\ IBr(g)$
I	2.0	2.0	0.0
C	$-x$	$-x$	$+2x$
E	$2.0 - x$	$2.0 - x$	$2x$