

PRACTICE

(Page 428)

Understanding Concepts

1. Evaporating water molecules cannot escape from the plastic bag. The saturation of the air inside the bag with water vapour prevents the towel from drying completely as a dynamic equilibrium becomes established.
2. Dissolution and crystallization continue to occur at the same rate, resulting in no visible changes.
3. Stirring removes solute as it dissolves, lowering the concentration in the region of the solute, allowing further dissolution to occur.

Making Connections

4. Hot moist air can escape from the clothes dryer through the exhaust vent, to be replaced by drier air.
5. Resealing carbonated soft drink bottles prevents carbon dioxide gas from leaving the solution. Equilibrium is quickly reestablished between dissolved CO_2 and gaseous CO_2 in the air space above the soft drink.

PRACTICE

(Page 437)

Understanding Concepts

6. $[\text{CO}_{2(g)}]_{\text{initial}} = \frac{2.0 \text{ mol}}{5.0 \text{ L}}$
 $[\text{CO}_{2(g)}]_{\text{initial}} = 0.40 \text{ mol/L}$

ICE Table for the Decomposition of Carbon Dioxide			
	$2 \text{ CO}_{2(g)} \rightleftharpoons$	$2 \text{ CO}_{(g)} +$	$\text{O}_{2(g)}$
Initial concentration (mol/L)	0.4	0	0
Change in concentration (mol/L)	-2x	+2x	+x
Equilibrium concentration (mol/L)	0.39		

At equilibrium:

$$[\text{CO}_{2(g)}] = 0.40 \text{ mol/L} - 2x$$

$$[\text{CO}_{2(g)}] = 0.39$$

$$2x = 0.01$$

$$[\text{CO}_{(g)}] = 2x$$

$$[\text{CO}_{(g)}] = 0.01 \text{ mol/L}$$

$$[\text{O}_{2(g)}] = x$$

$$x = 0.005$$

The concentration of carbon monoxide is 0.01 mol/L and the concentration of oxygen is 0.005 mol/L.

7. $[\text{NOCl}_{(g)}]_{\text{initial}} = \frac{2.0 \text{ mol}}{2.0 \text{ L}}$

$$[\text{NOCl}_{(g)}]_{\text{initial}} = 1.0 \text{ mol/L}$$

ICE Table for the Decomposition of NOCl			
	$2 \text{ NOCl}_{(g)} \rightleftharpoons$	$2 \text{ NO}_{(g)} +$	$\text{Cl}_{2(g)}$
Initial concentration (mol/L)	1.0		
Change in concentration (mol/L)	-2x	+2x	+x
Equilibrium concentration (mol/L)		0.032	

At equilibrium:

$$[\text{NO}_{(g)}] = 0.32 \text{ mol/L}$$

$$[\text{NO}_{(g)}] = 2x$$

$$x = 0.16 \text{ mol/L}$$

$$[\text{NOCl}_{(g)}] = 1.0 - 2x$$

$$= 1.0 - 0.32$$

$$[\text{NOCl}_{(g)}] = 0.68 \text{ mol/L}$$

$$[\text{Cl}_{2(g)}] = x$$

$$= 0.16 \text{ mol/L}$$

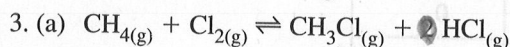
The concentration of $\text{NO}_{(g)}$ and $\text{Cl}_{2(g)}$ are 0.68 mol/L and 0.16 mol/L, respectively.

SECTION 7.1 QUESTIONS

(Page 437)

Understanding Concepts

- No observable changes occur.
 - Equilibrium is considered to be “dynamic” because changes are still occurring at the microscopic level.
 - The rates of the two opposing processes are equal.
- (Answers will vary. Possible examples include:
 - undissolved sugar at the bottom of cup of tea (dynamic equilibrium)
 - an ice cube floating in water at 0°C (dynamic equilibrium)
 - carbon dioxide gas in equilibrium with dissolved carbon dioxide in soft drinks (dynamic equilibrium)
 - preparing a saturated sugar solution when making “rock” candy (dynamic equilibrium)



(b) percent reaction = $\frac{\text{actual product yield}}{\text{theoretical product yield}} \times 100\%$

$$= \frac{1.4}{2.0} \times 100\%$$

percent reaction = 70%

The percent reaction of this equilibrium is 70%. This equilibrium favours the products.

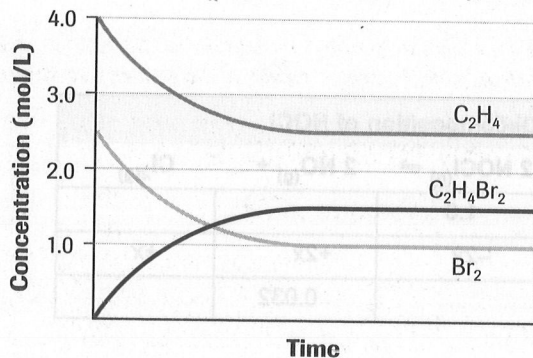
4. (a)

ICE Table for the Reaction $\text{C}_2\text{H}_4(g)$ with $\text{Br}_2(g)$			
	$\text{C}_2\text{H}_4(g)$	$+$	$\text{Br}_2(g) \rightleftharpoons \text{C}_2\text{H}_4\text{Br}_2(g)$
Initial concentration (mol/L)	4.00		2.50
Change in concentration (mol/L)	$-x$		$+x$
Equilibrium concentration (mol/L)	2.50		1.50

(b)

Reaction of Ethene

$$2.50 = 4 - x \therefore x = 1.50$$



$$\begin{aligned} \text{(c) percent reaction} &= \frac{\text{actual product yield}}{\text{theoretical product yield}} \times 100\% \\ &= \frac{1.5}{2.5} \times 100\% \end{aligned}$$

$$\text{percent reaction} = 60\%$$

5. (a) $< 50\%$
 $\text{N}_{2(g)} + 3 \text{H}_{2(g)} \rightleftharpoons 2 \text{NH}_{3(g)}$
- (b) $> 50\%$
 $\text{C}_{(s)} + \text{H}_2\text{O}_{(g)} \rightleftharpoons \text{CO}_{(g)} + \text{H}_{2(g)}$
- (c) $> 99\%$
 $2 \text{Ag}^+_{(aq)} + \text{Cu}_{(s)} \rightleftharpoons 2 \text{Ag}_{(s)} + \text{Cu}^{2+}_{(aq)}$
- (d) $> 50\%$
 $2 \text{SO}_{2(g)} + \text{O}_{2(g)} \rightleftharpoons 2 \text{SO}_{3(g)}$

6. The addition reactions of halogens and alkenes are generally quantitative, so for the reaction of bromine with ethene (ethylene) a $>99\%$ yield of products is predicted.

7. (a) $n_{\text{HI}} = 2.00 \text{ L} \times 0.00 \text{ mol/L}$

$$n_{\text{HI}} = 0.00 \text{ mol}$$

$$n_{\text{H}_2} = 2.00 \text{ L} \times 6.0 \text{ mol/L}$$

$$n_{\text{H}_2} = 12.0 \text{ mol}$$

$$n_{\text{I}_2} = 2.00 \text{ L} \times 4.0 \text{ mol/L}$$

$$n_{\text{I}_2} = 8.0 \text{ mol}$$

(b)

ICE Table for the Reaction of Hydrogen and Iodine			
	$\text{H}_{2(g)}$	$\text{I}_{2(g)}$	$\text{HI}_{(g)}$
Initial concentration (mol/L)	6.0	4.0	0.0
Change in concentration (mol/L)	$-x$ 3.6	$-x$ 3.6	$+2x$ 7.2
Equilibrium concentration (mol/L)	2.4	0.4	7.2

← read from graph.

At equilibrium, 14.4 mol of hydrogen iodide had formed. $4 - x = 0.4$
 $x = 3.6$

(c) Since iodine is the limiting reagent, the maximum possible yield of hydrogen iodide is calculated as

$$n_{\text{HI}} = 8.0 \text{ mol} \times \frac{2 \text{ mol}}{1 \text{ mol}}$$

$$n_{\text{HI}} = 16.0 \text{ mol}$$

$$\text{percent reaction} = \frac{14.4 \text{ mol}}{16.0 \text{ mol}} \times 100\%$$

$$\text{percent reaction} = 90\%$$

(d) At equilibrium, the rate of the forward reaction is equal to the rate of the reverse reaction. In this case, at equilibrium, the rate of reaction of hydrogen with iodine is equal to the rate of decomposition of hydrogen iodide.

8.

ICE Table for the Decomposition of $\text{PCl}_{5(g)}$			
	$\text{PCl}_{5(g)}$	$\text{PCl}_{3(g)}$	$\text{Cl}_{2(g)}$
Initial concentration (mol/L)	1.00	0	0
Change in concentration (mol/L)	$-0.100 -x$	$+0.100 +x$	$+0.100 +x$
Equilibrium concentration (mol/L)	0.90	0.100	0.100

$$\begin{aligned} 0 + x &= 0.100 \\ x &= 0.100 \end{aligned}$$

9. (a) The concentrations of $\text{PCl}_3(\text{g})$ and $\text{Cl}_2(\text{g})$ are both 0.100 mol/L.

ICE Table for the Production of Methanol			
	$\text{CO}(\text{g}) +$	$2 \text{H}_2(\text{g}) \rightleftharpoons$	$\text{CH}_3\text{OH}(\text{g})$
Initial concentration (mol/L)	0.100	0.200	0
Change in concentration (mol/L)	$-x$ -0.040	$-2x$ -0.080	$+x$ +0.040
Equilibrium concentration (mol/L)	0.060	0.120	0.040

$$x = 0.04$$

The concentrations of carbon monoxide and methanol are 0.060 mol/L and 0.0400 mol/L, respectively.

(b) percent reaction = $\frac{\text{actual product yield}}{\text{theoretical product yield}} \times 100\%$

$$= \frac{0.0400}{0.100} \times 100\%$$

$$\text{percent reaction} = 40\%$$

10. (a) The saturated solution is in a state of dynamic equilibrium. Even though the appearance of the system did not change, the evidence suggests that some radioactive iodine became incorporated into the original sample of iodine. Furthermore, the radioactivity of the solution implies that some radioactive iodine had dissolved.

Making Connections

11. (a) Immediately following a fructose-rich meal, a net amount of diffusion of fructose will occur into the (surrounding) intestinal cells because the concentration of fructose in the intestinal tract is higher than the concentration of fructose in the intestinal cells. Diffusion into the cells continues until the concentration of fructose inside the intestinal cells and the intestinal tract are equal. At this point, a dynamic equilibrium is reached; the rate of fructose diffusing into the cells equals the rate of fructose diffusing out of the cells. Thus, a certain amount of fructose remains in the intestinal tract (unabsorbed).
- (b) Cells use active transport mechanisms to pump nutrient molecules such as glucose and amino acids (digested protein). These systems require the use of cell membrane proteins and cell energy in the form of adenosine triphosphate (ATP). Using energy to actively pump certain nutrients against a concentration gradient ensures that intestinal cells absorb the majority of nutrient molecules.

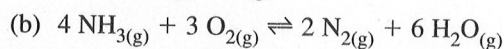
7.2 EQUILIBRIUM LAW IN CHEMICAL REACTIONS

PRACTICE

(Page 442)

Understanding Concepts

1. (a) $K = \frac{[\text{SO}_3(\text{g})]^2}{[\text{SO}_2(\text{g})]^2[\text{O}_2(\text{g})]}$



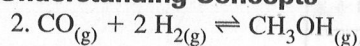
$$K = \frac{[\text{N}_2(\text{g})]^2[\text{H}_2\text{O}(\text{g})]^6}{[\text{NH}_3(\text{g})]^4[\text{O}_2(\text{g})]^3}$$

(c) $K = \frac{[\text{NO}(\text{g})]^2[\text{Br}_2(\text{g})]}{[\text{NOBr}(\text{g})]^2}$

PRACTICE

(Page 444)

Understanding Concepts

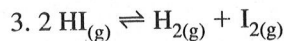


$$K = \frac{[\text{CH}_3\text{OH}_{(g)}]}{[\text{CO}_{(g)}][\text{H}_2_{(g)}]^2}$$

$$= \frac{[0.021]}{[0.079][0.158]^2}$$

$$K = 10.6$$

The value of the equilibrium constant for this reaction at 327°C is 10.6.



$$K = \frac{[\text{H}_{2(g)}][\text{I}_{2(g)}]}{[\text{HI}_{(g)}]^2}$$

$$= \frac{[1.065][0.065]}{[1.870]^2}$$

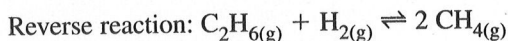
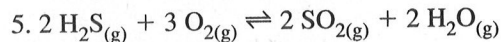
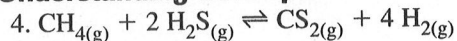
$$K = 0.020$$

The value of the equilibrium constant for this reaction at 440°C is 0.020.

PRACTICE

(Page 444)

Understanding Concepts



$$K' = \frac{[\text{CH}_{4(g)}]^2}{[\text{H}_2_{(g)}][\text{C}_2\text{H}_{6(g)}]}$$

$$= \frac{1}{K}$$

$$= \frac{1}{9.5 \times 10^{-13}}$$

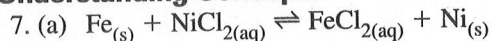
$$K' = 1.1 \times 10^{12}$$

The value of the equilibrium constant for the reverse reaction is 1.1×10^{12} .

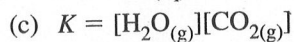
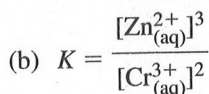
PRACTICE

(Page 447)

Understanding Concepts



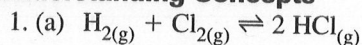
$$K = \frac{[\text{Fe}^{2+}_{(aq)}]}{[\text{Ni}^{2+}_{(aq)}]}$$



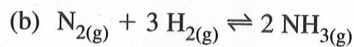
SECTION 7.2 QUESTIONS

(Page 448)

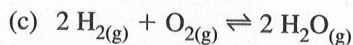
Understanding Concepts



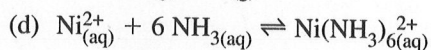
$$K = \frac{[\text{HCl}_{(g)}]^2}{[\text{H}_{2(g)}][\text{Cl}_{2(g)}]}$$



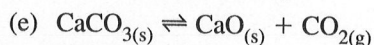
$$K = \frac{[\text{NH}_{3(g)}]^2}{[\text{N}_{2(g)}][\text{H}_{2(g)}]^3}$$



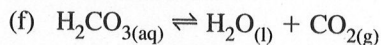
$$K = \frac{[\text{H}_2\text{O}_{(g)}]^2}{[\text{H}_{2(g)}]^2[\text{O}_{2(g)}]}$$



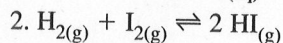
$$K = \frac{[\text{Ni}(\text{NH}_3)_6^{2+}_{(aq)}]}{[\text{Ni}_{(aq)}^{2+}][\text{NH}_{3(aq)}]^6}$$



$$K = [\text{CO}_{2(g)}]$$



$$K = \frac{[\text{CO}_{2(g)}]}{[\text{H}_2\text{CO}_{3(aq)}]}$$



$$K = \frac{[\text{HI}_{2(g)}]^2}{[\text{H}_{2(g)}][\text{I}_{2(g)}]}$$

$$= \frac{[3.116]^2}{[0.442][0.442]}$$

$$K = 49.7$$

The value of the equilibrium constant at 485°C is 49.7.

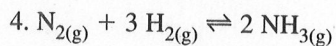


$$K = \frac{[\text{PCl}_{3(g)}][\text{Cl}_{2(g)}]}{[\text{PCl}_{5(g)}]}$$

$$= \frac{[0.014][0.014]}{[4.3 \times 10^{-4}]}$$

$$K = 0.46$$

The value of the equilibrium constant at 200°C is 0.46.



$$K = \frac{[\text{NH}_{3(g)}]^2}{[\text{N}_{2(g)}][\text{H}_{2(g)}]^3}$$

$$K = 8.00 \times 10^{-7}$$

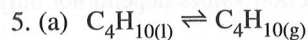
$$[\text{NH}_3(\text{g})]^2 = 8.00 \times 10^{-7} \times [\text{N}_2(\text{g})][\text{H}_2(\text{g})]^3$$

$$= 8.00 \times 10^{-7} [1.50][0.50]^3$$

$$[\text{NH}_3(\text{g})]^2 = 1.5 \times 10^{-7}$$

$$[\text{NH}_3(\text{g})] = 3.9 \times 10^{-4} \text{ mol/L}$$

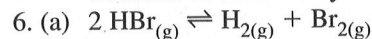
The equilibrium concentration of ammonia is $3.9 \times 10^{-4} \text{ mol/L}$.



$$K = [\text{C}_4\text{H}_{10(\text{g})}]$$

(b) The amount of butane in the gas phase varies directly with the temperature. As the temperature increases, for example, more butane will be found in the gas phase. However, the butane concentration at a particular temperature is a constant. Since the value of the equilibrium constant depends only on $[\text{C}_4\text{H}_{10(\text{g})}]$, it also is constant.

(c) The concentration of butane gas depends only on the temperature and not on the amount of liquid butane. If the volume of liquid butane decreases, there's more space above the liquid for gaseous butane to occupy. The number of moles of butane may have increased but the mol/L remains the same.



(b)
$$K = \frac{[\text{H}_{2(\text{g})}][\text{Br}_{2(\text{g})}]}{[\text{HBr}_{(\text{g})}]^2}$$

(c)
$$n_{\text{HBr}_{(\text{g})}} = C_{\text{HBr}_{(\text{g})}} \times 2.00 \text{ L}$$

$$= 0.100 \text{ mol/L} \times 2.00 \text{ L}$$

$$n_{\text{HBr}_{(\text{g})}} = 0.200 \text{ mol}$$

0.200 mol of $\text{HBr}_{(\text{g})}$ is present at equilibrium.

(d) $n_{\text{HBr}} = 0.100 \text{ mol/L} \times 2.00 \text{ L}$

$$n_{\text{HBr}} = 0.200 \text{ mol}$$

0.200 mol of hydrogen bromide is present at equilibrium.

$$n_{\text{HBr}_{\text{reacted}}} = n_{\text{HBr}_{\text{initial}}} - n_{\text{HBr}_{\text{equilibrium}}}$$

$$= 1.00 \text{ mol} - 0.200 \text{ mol}$$

$$n_{\text{HBr}_{\text{reacted}}} = 0.800 \text{ mol}$$

0.800 mol of hydrogen bromide reacted.

(e) Given the 2:1 ratio, 0.400 mol of both hydrogen and bromine are produced.

(f)
$$[\text{H}_2] = \frac{0.400 \text{ mol}}{2.00 \text{ L}}$$

$$= 0.200 \text{ mol/L}$$

$$[\text{H}_2] = [\text{Br}_2]$$

The concentrations of hydrogen and bromine are both 0.200 mol/L, while the hydrogen bromide concentration is 0.100 mol/L.

(g)
$$K = \frac{[\text{H}_{2(\text{g})}][\text{Br}_{2(\text{g})}]}{[\text{HBr}_{(\text{g})}]^2}$$

$$= \frac{[0.20][0.20]}{[0.100]^2}$$

$$K = 4.0$$

The value of the equilibrium constant is 4.00.

7. (a) One could specify either the value of the equilibrium constant or the extent of reaction (as percent reaction).
 (b) Both methods measure the extent to which reactants become products in a closed system. However, percent reaction describes the extent of the reaction as the yield of product at equilibrium compared to the maximum possible yield. The value of the equilibrium constant gives the ratio of products to reactants at equilibrium. Furthermore, the percent reaction is variable, for a given reaction, depending on the concentration of reactants; the equilibrium constant is independent of concentration.
 (c) Both methods are effective in describing the extent of the reaction. However, they also have their limitations. The value of the equilibrium constant is valid only for a given temperature. Percent reaction values depend not only on temperature but on concentration as well.

Applying Inquiry Skills

8. Analysis

$$(a) K = \frac{[\text{SO}_{3(g)}]^2}{[\text{SO}_{2(g)}]^2[\text{O}_{2(g)}]}$$

For trial 1...

$$K = \frac{[3.50 \times 10^{-2}]^2}{[1.50 \times 10^{-1}]^2[1.26 \times 10^{-2}]}$$

$$K = 4.32$$

For trial 2...

$$K = \frac{[2.60 \times 10^{-3}]^2}{[5.90 \times 10^{-2}]^2[4.50 \times 10^{-4}]}$$

$$K = 4.32$$

For trial 3...

$$K = \frac{[3.6 \times 10^{-3}]^2}{[1.00 \times 10^{-2}]^2[3.0 \times 10^{-2}]}$$

$$K = 4.32$$

Making Connections

9. The small value of the equilibrium constant suggests that the equilibrium strongly favours oxygen. Only a negligible amount of ozone is produced by this reaction.

7.3 QUALITATIVE CHANGES IN EQUILIBRIUM SYSTEMS

PRACTICE

(Page 457)

Understanding Concepts

1. The equilibrium will shift to the
 (a) left
 (b) right
 (c) left
 (d) ~~right~~ left
 (e) right