

3. (a) no change
 (b) right
 (c) left
 (d) no change (liquid volumes are largely unaffected by pressure changes)
4. (a) $\text{N}_{2(\text{g})} + 3\text{H}_{2(\text{g})} \rightleftharpoons 2\text{NH}_{3(\text{g})} + 92 \text{ kJ}$
 (b) Changes that can increase the yield of ammonia include: adding N_2 or H_2 , removing NH_3 , increasing the pressure, lowering the temperature
5. (a) $2\text{SO}_{2(\text{g})} + \text{O}_{2(\text{g})} \rightleftharpoons 2\text{SO}_{3(\text{g})} + \text{energy}$ ΔH is negative (energy is released)
- (b) Changes that can increase the yield of sulfur trioxide include: adding SO_2 or O_2 , removing SO_3 , increasing the pressure, lowering the temperature
- (c) This reaction is carried out at high temperatures because at lower temperatures the yield of SO_3 is relatively low. High temperatures increase pressure of the reaction mixture, thereby shifting the equilibrium to the right.
- (d) A vanadium pentoxide catalyst is required to increase the yield of SO_3 at lower, more manageable temperatures.

Careers in Chemistry

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Answers to Making Career Connections

- Anesthesiologists can also be found in the recovery room watching patients emerging from the effects of the anesthesia, in a doctor's office or dental office assisting with pain management, assisting with critical care and trauma medicine, occasionally assisting in stabilizing patients during transfer, as well as in situations involving non-surgical treatments, diagnostic procedures, and research laboratories.
- Beyond ensuring patient comfort, an anesthesiologist must also make crucial medical judgments to protect and regulate a patient's life functions. Anesthesiologists, must know, understand, and evaluate a patient's medical condition, monitor body functions, and be aware of any interactions that could possibly occur between the anesthetics and a patient's current medication and diet.
- Local anesthetics: are used to control pain in a specific area of the body and are commonly used in minor medical procedures (e.g., receiving stitches or hernia repair).
 Regional anesthetics: also known as a regional block, are used to block pain in a specific part of the body (e.g., a hand or a leg)
 General anesthetics: are used to stop pain from being felt during a medical procedure such as surgery, by rendering a person unconscious.

Chapter 7 Review Answers

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Answers to Knowledge/Understanding Questions

- The rate of a reaction is the time required for a particular amount of product to form. The extent of a reaction is the relative concentration of products to reactants at equilibrium.
- The state of equilibrium involves dynamic change at the microscopic level, even though there is no change at the macroscopic level. Reactant and product molecules are constantly moving in the forward and reverse reactions.
- In a reaction that goes to completion, the products are favoured. The numerator in the equilibrium equation, which represents the concentration of products, must be large compared to the denominator. Therefore, K_c is very large.
- (a) $-\Delta H$ and $+\Delta S$ contribute to a favourable reaction.
 (b) $+\Delta H$ and $-\Delta S$ contribute to an unfavourable reaction.
- Factors that can affect the equilibrium of a reaction are temperature, enthalpy, and entropy.
- The solubility of a solid in a liquid increases as the temperature rises because of the larger entropy associated with the molecules in a liquid. The motion of molecules in a liquid is much more random and chaotic than in a solid. On the other hand, the solubility of a gas decreases as the temperature increases because of the decrease in entropy as the gaseous molecules react to form a liquid.
- A decrease in temperature will shift the equilibrium to the right. The lower temperature relieves the heat produced in the exothermic reaction. A decrease in volume will have no effect on the equilibrium because there are the same numbers of molecules on either side of the reaction equation.

8. An increase in pressure will result in a decrease in volume. Since the reactant side of the equation has fewer molecules than the product side, the equilibrium will shift to the left.
9. (a) The increase in pressure will increase the concentration of reactants, while the concentration of products will decrease. The equilibrium will shift to the right, where there are fewer molecules.
- (b) Once equilibrium is re-established, CH_4 and H_2O will show an increase in concentration.

Answers to Inquiry Questions

10. $[\text{I}_2] = 5.6 \times 10^{-4} \text{ mol/L}$

The volume is constant. The concentration of HI and H_2 is calculated using mole ratios. $[\text{I}_2]:[\text{H}_2]$ are in a 1:1 ratio. Therefore, $[\text{H}_2] = 5.6 \times 10^{-4} \text{ mol/L}$. $[\text{I}_2]:[\text{HI}]$ are in a 1:2 ratio. Therefore, $[\text{HI}] = 5.6 \times 10^{-4} \text{ mol/L} \times 2 = 1.12 \times 10^{-3} \text{ mol/L}$.

11.

Concentration (mol/L)	$2\text{SO}_{2(g)}$	+	$\text{O}_{2(g)}$	\rightleftharpoons	$2\text{SO}_{3(g)}$
Initial	1.7		1.7		0
Change	$-2x$		$-x$		$+2x$
Equilibrium	$1.7 - 2x$		$1.7 - x$		$2x$

$$K_c = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2[\text{O}_2]}$$

use the approximation method

$$3.6 \times 10^{-3} = \frac{(2x)^2}{(1.7)^2(1.7)}$$

$$x = 0.066$$

$$[\text{SO}_3] = 2(0.066) = 0.13 \text{ mol/L}$$

12. [ethanol] = $30.0 \text{ g} \div 46 \text{ g/mol} = 0.65 \text{ mol}$
 [propanoic acid] = $40.0 \text{ g} \div 74 \text{ g/mol} = 0.54 \text{ mol}$

Concentration (mol/L)	$\text{C}_2\text{H}_5\text{OH}$	+	$\text{C}_2\text{H}_5\text{COOH}$	\rightleftharpoons	$\text{C}_2\text{H}_5\text{COOC}_2\text{H}_5$	+	H_2O
Initial	0.65		0.54		0		
Change	$-x$		$-x$		$+x$		
Equilibrium	$0.65 - x$		$0.54 - x$		x		

$$7.5 = \frac{x}{(0.65 - x)(0.54 - x)}$$

$$7.5x^2 - 9.94x + 2.64 = 0$$

$$x = 0.38 \text{ or } 0.96$$

$$x = 0.38 \text{ mol}$$

$$\text{Mass of ethyl propanoate} = 0.38 \text{ mol} \times 102 \text{ g/mol} = 37.4 \text{ g}$$

13. $[\text{CO}] = [\text{H}_2\text{O}] = 0.50 \text{ mol}/10\text{L} = 0.05 \text{ mol/L}$

Concentration (mol/L)	$\text{CO}_{(g)}$	+	$\text{H}_2\text{O}_{(g)}$	\rightleftharpoons	$\text{H}_2_{(g)}$	+	$\text{CO}_{2(g)}$
Initial	0.05		0.05		0		0
Change	$-x$		$-x$		$+x$		$+x$
Equilibrium	$0.05 - x$		$0.05 - x$		x		x

$$K_c = \frac{[\text{H}_2][\text{CO}_2]}{[\text{CO}][\text{H}_2\text{O}]}$$

$$8.3 = \frac{x^2}{(0.05 - x)^2}$$

$$2.88 = \frac{x}{(0.05 - x)}$$

$$x = 0.037$$

$$[\text{H}_2] = [\text{CO}_2] = 0.037 \text{ mol/L}$$

$$[\text{CO}] = [\text{H}_2\text{O}] = 0.05 - 0.037 = 0.013 \text{ mol/L}$$

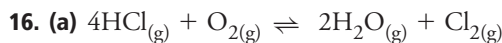
14. Change in concentration of $\text{S}_{8(g)} = 9.2 \times 10^{-3} \text{ mol/L} - 2.3 \times 10^{-3} \text{ mol/L} = 0.0069 \text{ mol/L}$. Concentration of $\text{S}_{2(g)}$ at equilibrium = $0.0069 \text{ mol/L} \times 4 = 0.028 \text{ mol/L}$.

$$K_c = \frac{[\text{S}_2]^4}{[\text{S}_8]}$$

$$K_c = \frac{(0.028)^4}{2.3 \times 10^{-3}}$$

$$K_c = 2.5 \times 10^{-4}$$

15. According to the second law of thermodynamics, the total entropy of the universe is increasing. Heat released by the machine will radiate into the surroundings, causing an increase in the random motion of particles, thereby increasing the entropy of the surroundings. ΔS_{surr} is increasing. In nature, a state of increased entropy is favoured, so energy will spontaneously move from the machine to the surroundings, increasing the state of disorder. The energy will not be recycled back into the machine, causing its motion to eventually stop.



(b)
$$K_c = \frac{[\text{H}_2\text{O}]^2[\text{Cl}_2]}{[\text{HCl}]^4[\text{O}_2]} = \frac{(7.8 \times 10^{-3})^2(3.6 \times 10^{-3})^2}{(2.7 \times 10^{-2})^4(8.6 \times 10^{-2})} = 1.7 \times 10^{-2}$$

17. (a) $\text{SO}_3(\text{g}) + 6\text{HF}(\text{g}) \rightleftharpoons \text{SF}_6(\text{g}) + 3\text{H}_2\text{O}(\text{g})$
 (b) $[\text{SO}_3] = 2.9 \text{ mol} \div 4.7 \text{ L} = 0.62 \text{ mol/L}$
 $[\text{HF}] = 9.1 \text{ mol} \div 4.7 \text{ L} = 1.94 \text{ mol/L}$

Concentration (mol/L)	$\text{SO}_3(\text{g})$	+	$6\text{HF}(\text{g})$	\rightleftharpoons	$\text{SF}_6(\text{g})$	+	$3\text{H}_2\text{O}(\text{g})$
Initial	0.62		1.94		0		0
Change	$-x$		$-6x$		$+x$		$+3x$
Equilibrium	$0.62 - x$		$1.94 - 6x$		x		$3x$

(c) When the equilibrium conditions are put into the equilibrium expression for the reaction, the equation contains terms to the 6th power. This proves to be very difficult to solve using simple mathematics.

18. For Experiment 1

$$K_c = \frac{[\text{SO}_3]^2}{[\text{O}_2][\text{SO}_2]^2} = \frac{(3.50)^2}{(1.25)(1.5)^2} = 4.36$$

For Experiment 2

$$K_c = \frac{[\text{SO}_3]^2}{[\text{O}_2][\text{SO}_2]^2} = \frac{(0.260)^2}{(0.0450)(0.590)^2} = 4.32$$

Within experimental error, these values may be considered equal.

19. $4\text{NH}_3(\text{g}) + 5\text{O}_2(\text{g}) \rightleftharpoons 4\text{NO}(\text{g}) + 6\text{H}_2\text{O}(\text{g})$

$$K_c = \frac{[\text{NO}]^4(2.0)^6}{(3.0)^5[\text{NH}_3]^4}$$

$$[\text{NO}] = [\text{NH}_3]$$

$$K_c = 0.26$$

Answers to Communication Questions

20. A student answer might read: Energy is conserved in a chemical reaction. Energy gained by the system is equal to the energy lost by the surroundings, or vice versa. Therefore, the combination of a reaction system and its surroundings cannot gain any additional energy; the best you can do, in other words, is break even. However, since the entropy of the universe is always increasing, you cannot even break even.
21. (a) This graph should show the concentration of H_2 decreasing with time, and the concentration of HI increasing with time.
 (b) The graph of $[\text{I}_2]$ and $[\text{HI}]$ with time should look similar to the graph sketched in (a). The concentrations of I_2 and H_2 should decrease at the same rate, and that of HI should increase.
 (c) The value of Q_c will increase until it reaches the same value of K_c .
22. Adding a small crystal to a saturated solution of copper(II) sulfate disrupts the equilibrium of the system. CuSO_4 can no longer dissolve and the reaction shifts to adjust to the addition of the solid. Thus, the ions come out of solution to form a solid.

Answers to Making Connections Questions

23. A student answer might read: The second law of thermodynamics says that the entropy of the universe is always increasing. Western societies lead lives of high entropy. By using numerous products, we also produce large quantities of waste (high entropy). Consumer products originate through industrial processes that make use of the Earth's energy resources (low entropy matter). The impact of this is seen through the buildup of garbage on our streets as our landfills cannot keep up with the demand.
24. In 1 L of air, there is 0.80 L of N_2 and 0.20 L of O_2
 Concentration of $\text{N}_2 = (0.80 \text{ L} \div 22.4 \text{ L/mol}) \div 1 \text{ L} = 0.0357 \text{ mol/L}$
 Concentration of $\text{O}_2 = (0.20 \text{ L} \div 22.4 \text{ L/mol}) \div 1 \text{ L} = 0.0089 \text{ mol/L}$

	$\text{N}_2(\text{g})$	+	$\text{O}_2(\text{g})$	\rightleftharpoons	$2\text{NO}(\text{g})$
Initial	0.0357		0.0089		0
Change	$-x$		$-x$		$+2x$
Equilibrium	$0.0357 - x$		$0.0089 - x$		$2x$

$$K_c = \frac{[\text{NO}]^2}{[\text{O}_2][\text{N}_2]}$$

$$4.7 \times 10^{-31} = \frac{(2x)^2}{(0.0357 - x)(0.0089 - x)}$$

Using the approximation method

$$x = 6.11 \times 10^{-18}$$

$$[\text{NO}] = 1.2 \times 10^{-17}$$

Although this concentration of NO appears to be very small, additional amounts of NO are received from automobile and industrial pollution.

25. (a) A decrease in pressure, or the removal of heat from the system.
 (b) A catalyst lowers the activation energy and increases the rate of the reaction, but it does not change the equilibrium position.
 (c) This reaction must be carried out at relatively high temperatures to improve the rate at which NO is produced. Although the reaction appears to be favoured by low temperature, high temperatures are required for NO to be produced at an optimum rate.
 (d) Since there are more gaseous product molecules, the product side exerts a greater pressure than the reactant side. By lowering the pressure, the equilibrium will shift to the right.
 (e) The gaseous products are cooled to shift the reaction equilibrium to the right, favouring the production of NO_2 . The extracted heat is used to increase the temperature of the surroundings.

- (f) The NO is recycled back into the reaction. Uses of HNO_3 include explosives, polymers, and fertilizers.
26. shifts left; shifts right; shifts left; no change; no change
27. (a) Conditions of low pressure and high temperature will maximize the production of polystyrene.
- (b) Conditions of low pressure and high temperature are difficult to achieve in industrial applications. To maximize the yield, the polystyrene must be extracted as the reaction continues, keeping the equilibrium shifted to the right.
- (c) The products of the reaction between steam and carbon are carbon monoxide and hydrogen gas.
- (d) The super-heated steam supplies energy to a reaction that is endothermic. This will shift the equilibrium to the right, favouring the production of polystyrene.
- (e) Super-heated steam at 600°C will increase the reaction temperature, without having to heat the reaction vessel itself. If the steam were only at 100°C , additional heat would be needed to maximize polystyrene production.
- (f) The iron oxide catalyst is used to increase the reaction rate.
- (g) If the temperature is raised, possible degradation of the catalyst can occur; as well, the pressure of the system would likely increase, thereby shifting the equilibrium back to the left.
- (h) The ethyl benzene is evaporated from the reaction mixture, condensed, and re-used in further polystyrene production.