

Kildonan-East Collegiate
Chemistry 40S
Course Review

The questions below can be found in your textbook at the end of each chapter. These questions, by chapter, are suggested but you should choose carefully as not all necessarily need to be considered. For example, if you struggled with a unit, you may want to do the "Mastering Concepts" but if you want a more efficient review you could skip to the "Mastering Problems" section. The "Standardized Test Practice" are good preparation for assessment (exam). You may also wish to re-do old questions or worksheets. If you need more practice try the "Mixed Review" or "Cumulative Review". At this stage, reading should only be done for concepts which you have forgotten or did not fully learn. Your time is better used being selective and doing problems. Best of luck. Please book time with me if you need help. Questions? Email me: hpeltz@retsd.mb.ca

Chapter 5: 30,34,39,40,41,44,47,49,50,52,54,57,64,78,79(a-e),80
Standardized Test Practice (STP) 3,4,5,7,8,10

Chapter 6: 42,43,44,45,46,47,50,52,54,55,57,58,59,62,63,65,67
STP 2,3,5,6,7,8

Chapter 9: 99,108,110 STP 3

Chapter 17: 35,36,37,38,40,42,43,44,47,49,52,54,56,64,69,70,71
STP 1,7,8

Chapter 18: 33,43,46,50,52,53,57,59,60,61,62,63,64,68 (find [] at equilibrium, K_{sp})
STP 1,2,3,4,8,9

Chapter 19: 42,43,46,49,50,52,53,56,71,81,83,87,90,91,92 (given K_a find pH and % ioniz)
STP 4,7,10

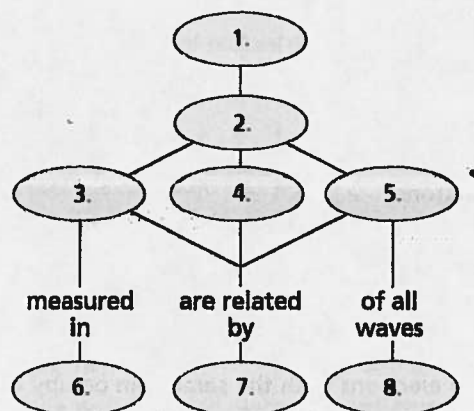
Chapter 20: 38,39,40,42,44,45,47,49,53,56,57,59,61
STP 1,2,3,4,5,6

Chapter 21: STP 1,2,4,5,7,9

Chapter 5 Assessment pages 146–148

Concept Mapping

29. Complete the concept map using the following terms: speed, $c = \lambda\nu$, electromagnetic waves, wavelength, characteristic properties, frequency, c , and hertz.



1. electromagnetic waves; 2. characteristic properties; 3. frequency; 4. wavelength; 5. speed; 6. hertz; 7. $c = \lambda\nu$; 8. c

Mastering Concepts

30. Define the following terms.

- a. frequency (5.1)

Frequency is the number of waves that pass a given point per second.

- b. wavelength (5.1)

Wavelength is the shortest distance between equivalent points on a continuous wave.

- c. quantum (5.1)

A quantum is the minimum amount of energy that can be lost or gained by an atom.

- d. ground state (5.2)

An atom's ground state is its lowest allowable energy state.

31. Why did scientists consider Rutherford's nuclear model of the atom incomplete? (5.1)

Typical answers will say that the model did not explain the following: how the atom's negatively charged electrons occupy the space around the nucleus; why the electrons are not drawn into the atom's positively charged nucleus; a rationale for the chemical properties of the elements.

32. Name one type of electromagnetic radiation. (5.1)

light, microwaves, X rays, radio waves

33. Explain how the gaseous neon atoms in a neon sign emit light. (5.1)

Electricity passing through the tube excites neon atoms to higher energy levels. As the excited atoms drop back to lower energy levels, they emit light.

34. What is a photon? (5.1)

a particle of electromagnetic radiation having a rest mass of zero and carrying a quantum of energy

35. What is the photoelectric effect? (5.1)

a phenomenon in which a metal emits electrons when light of a sufficient frequency shines on it

36. Explain Planck's quantum concept as it relates to energy lost or gained by matter. (5.1)

According to Planck, for a given frequency, ν , matter can emit or absorb energy only in discrete quanta that are whole-number multiples of $h\nu$.

37. How did Einstein explain the previously unexplainable photoelectric effect? (5.1)

He proposed that photons must have a certain minimum, or threshold, value to cause the ejection of a photoelectron.

Section 5.3 Assessment

page 141

24. State the Aufbau principle in your own words.

Electrons tend to occupy the lowest energy orbital available.

- 38.** Arrange the following types of electromagnetic radiation in order of increasing wavelength. (5.1)
- ultraviolet light
 - microwaves
 - radio waves
 - X rays
- d.** X rays, **a.** ultraviolet light, **b.** microwaves, **c.** radio waves
- 39.** What is the difference between an atom's ground state and an excited state? (5.2)
- An atom's ground state is its lowest energy state, while any energy state higher than the ground state is an excited state.
- 40.** According to the Bohr model, how do electrons move in atoms? (5.2)
- Electrons move in circular orbits around the nucleus.
- 41.** What does n designate in Bohr's atomic model? (5.2)
- The quantum number n specifies the electron's orbit.
- 42.** Why are you unaware of the wavelengths of moving objects such as automobiles and tennis balls? (5.2)
- Their wavelengths are too small to be seen.
- 43.** What is the name of the atomic model in which electrons are treated as waves? Who first wrote the electron wave equations that led to this model? (5.2)
- the quantum mechanical model of the atom; Erwin Schrödinger
- 44.** What is an atomic orbital? (5.2)
- a three-dimensional region around the nucleus describing an electron's probable location
- 45.** What is the probability that an electron will be found within an atomic orbital? (5.2)
- The probability is 90%.
- 46.** What does n represent in the quantum mechanical model of the atom? (5.2)
- n represents an orbital's principal quantum number, which indicates the relative size and energy of the orbital.
- 47.** How many energy sublevels are contained in each of the hydrogen atom's first three energy levels? (5.2)
- energy level 1 has one sublevel, energy level 2 has two sublevels, energy level 3 has three sublevels
- 48.** What atomic orbitals are related to a p sublevel? To a d sublevel? (5.2)
- p sublevel: x , y , and z orbitals; d sublevel: xy , xz , yz , $x^2 - y^2$, and z^2 orbitals
- 49.** Which of the following atomic orbital designations are incorrect? (5.2)
- 7f
 - 3f
 - 2d
 - 6p
- b, c** are incorrect
- 50.** What do the sublevel designations s, p, d, and f specify with respect to the atom's orbitals? (5.2)
- their shapes
- 51.** What do subscripts such as y and xz tell you about atomic orbitals? (5.2)
- their orientations
- 52.** What is the maximum number of electrons an orbital may contain? (5.2)
- two electrons
- 53.** Why is it impossible to know precisely the velocity and position of an electron at the same time? (5.2)
- The photon required to measure an electron's velocity or position changes both the position and velocity of the electron.

54. What shortcomings caused scientists to finally reject Bohr's model of the atom? (5.2)

Bohr's model did not explain the spectra of atoms having more than one electron and did not fully explain the chemical behavior of atoms.

55. Describe de Broglie's revolutionary concept involving the characteristics of moving particles. (5.2)

de Broglie proposed that all moving particles have wave characteristics.

56. How is an orbital's principal quantum number related to the atom's major energy levels? (5.2)

Because the orbital's principal quantum number indicates the orbital's relative size and energy, it also specifies the atom's major energy level.

57. Explain the meaning of the aufbau principle as it applies to atoms with many electrons. (5.3)

The aufbau principle describes the sequence in which an atom's orbitals are filled with electrons.

58. In what sequence do electrons fill the atomic orbitals related to a sublevel? (5.3)

Each orbital must contain a single electron before any orbital contains two electrons.

59. Why must the two arrows within a single block of an orbital diagram be written in opposite (up and down) directions? (5.3)

Two electrons occupying a single atomic orbital must have opposite spins.

60. How does noble-gas notation shorten the process of writing an element's electron configuration? (5.3)

The noble-gas notation uses the bracketed symbol of the preceding noble gas in the periodic table to represent an atom's inner electrons.

61. What are valence electrons? How many of a magnesium atom's 12 electrons are valence electrons? (5.3)

Valence electrons are the electrons in an atom's outermost orbitals; 2

62. Light is said to have a dual wave-particle nature. What does this statement mean? (5.3)

Light exhibits wavelike behavior in some situations and particlelike behavior in others.

63. Describe the difference between a quantum and a photon. (5.3)

A quantum is the minimum amount of energy that can be lost or gained by an atom, while a photon is a particle of light that carries a quantum of energy.

64. How many electrons are shown in the electron-dot structures of the following elements? (5.3)

a. carbon

4

b. iodine

7

c. calcium

2

d. gallium

3

Mastering Problems

Wavelength, Frequency, Speed, and Energy (5.1) Level 1

65. What is the wavelength of electromagnetic radiation having a frequency of 5.00×10^{12} Hz? What kind of electromagnetic radiation is this?

$$\lambda = c/\nu = \frac{3.00 \times 10^8 \text{ m/s}}{5.00 \times 10^{12} \text{ s}^{-1}} = 6.00 \times 10^{-5} \text{ m}$$

infrared radiation

66. What is the frequency of electromagnetic radiation having a wavelength of 3.33×10^{-8} m? What type of electromagnetic radiation is this?

$$\nu = c/\lambda = \frac{3.00 \times 10^8 \text{ m/s}}{3.33 \times 10^{-8} \text{ m}} = 9.01 \times 10^{15} \text{ s}^{-1}$$

UV radiation

67. The laser in a compact disc (CD) player uses light with a wavelength of 780 nm. What is the frequency of this light?

$$\lambda = 780 \text{ nm} \times \frac{1 \text{ m}}{10^9 \text{ nm}} = 7.8 \times 10^{-7} \text{ m}$$

$$\nu = c/\lambda = \frac{3.00 \times 10^8 \text{ m/s}}{7.8 \times 10^{-7} \text{ m}} = 3.8 \times 10^{14} \text{ s}^{-1}$$

68. What is the speed of an electromagnetic wave having a frequency of $1.33 \times 10^{17} \text{ Hz}$ and a wavelength of 2.25 nm?

$$\nu = 3.00 \times 10^8 \text{ m/s}$$

69. Use Figure 5-5 to determine each of the following types of radiation.

- a. radiation with a frequency of $8.6 \times 10^{11} \text{ s}^{-1}$
infrared

- b. radiation with a wavelength 4.2 nm
X ray

- c. radiation with a frequency of 5.6 MHz
AM radio

- d. radiation that travels at a speed of $3.00 \times 10^8 \text{ m/s}$
any electromagnetic wave

70. What is the energy of a photon of red light having a frequency of $4.48 \times 10^{14} \text{ Hz}$?

$$E_{\text{photon}} = h\nu = (6.626 \times 10^{-34} \text{ J}\cdot\text{s})(4.48 \times 10^{14} \text{ s}^{-1}) = 2.97 \times 10^{-19} \text{ J}$$

71. Mercury's atomic emission spectrum is shown below. Estimate the wavelength of the orange line. What is its frequency? What is the energy of an orange photon emitted by the mercury atom?

$$\lambda \approx 615 \text{ nm} = 6.15 \times 10^{-7} \text{ m}$$

$$\nu = c/\lambda = (3.00 \times 10^8 \text{ m/s})/(6.15 \times 10^{-7} \text{ m}) = 4.88 \times 10^{14} \text{ s}^{-1}$$

$$E_{\text{photon}} = h\nu = (6.626 \times 10^{-34} \text{ J}\cdot\text{s})(4.88 \times 10^{14} \text{ s}^{-1}) = 3.23 \times 10^{-19} \text{ J}$$

72. What is the energy of an ultraviolet photon having a wavelength of $1.18 \times 10^{-8} \text{ m}$?

$$\nu = c/\lambda = \frac{3.00 \times 10^8 \text{ m/s}}{1.18 \times 10^{-8} \text{ m}} = 2.54 \times 10^{16} \text{ s}^{-1}$$

$$E_{\text{photon}} = h\nu = (6.626 \times 10^{-34} \text{ J}\cdot\text{s})(2.54 \times 10^{16} \text{ s}^{-1}) = 1.68 \times 10^{-17} \text{ J}$$

73. A photon has an energy of $2.93 \times 10^{-25} \text{ J}$. What is its frequency? What type of electromagnetic radiation is the photon?

$$\nu = E_{\text{photon}}/h = (2.93 \times 10^{-25} \text{ J})/(6.626 \times 10^{-34} \text{ J}\cdot\text{s}) = 4.42 \times 10^8 \text{ s}^{-1}$$

TV or FM waves

74. A photon has an energy of $1.10 \times 10^{-13} \text{ J}$. What is the photon's wavelength? What type of electromagnetic radiation is it?

$$\nu = E_{\text{photon}}/h = (1.10 \times 10^{-13} \text{ J})/(6.626 \times 10^{-34} \text{ J}\cdot\text{s}) = 1.66 \times 10^{20} \text{ s}^{-1}$$

$$\lambda = c/\nu = (3.00 \times 10^8 \text{ m/s})/(1.66 \times 10^{20} \text{ s}^{-1}) = 1.81 \times 10^{-12} \text{ m}$$

an X ray or gamma radiation

75. How long does it take a radio signal from the *Voyager* spacecraft to reach Earth if the distance between *Voyager* and Earth is $2.72 \times 10^9 \text{ km}$?

$$\text{velocity} = \frac{\text{distance}}{\text{time}} = \frac{d}{t}$$

$$d = 2.72 \times 10^9 \text{ km} \times \frac{1000 \text{ m}}{1 \text{ km}} = 2.72 \times 10^{12} \text{ m}$$

$$t = d/\nu = (2.72 \times 10^{12} \text{ m})/(3.00 \times 10^8 \text{ m/s}) = 9070 \text{ s or } 151 \text{ min}$$

76. If your favorite FM radio station broadcasts at a frequency of 104.5 MHz, what is the wavelength of the station's signal in meters? What is the energy of a photon of the station's electromagnetic signal?

$$\nu = 104.5 \text{ MHz} \times \frac{10^6 \text{ Hz}}{\text{MHz}} = 1.045 \times 10^8 \text{ Hz}$$

$$\lambda = c/\nu = \frac{3.00 \times 10^8 \text{ m/s}}{1.045 \times 10^8 \text{ s}^{-1}} = 2.87 \text{ m}$$

$$E_{\text{photon}} = h\nu = (6.626 \times 10^{-34} \text{ J}\cdot\text{s})(1.045 \times 10^8 \text{ s}^{-1}) = 6.92 \times 10^{-26} \text{ J}$$

Electron Configurations (5.3)

Level 1

77. List the aufbau sequence of orbitals from 1s to 7p.

1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p

78. Write orbital notations and complete electron configurations for atoms of the following elements.

a. beryllium

Be $1s^2 2s^2$

$\uparrow\downarrow \uparrow\downarrow$

1s 2s

b. aluminum

Al $1s^2 2s^2 2p^6 3s^2 3p^1$

$\uparrow\downarrow \uparrow\downarrow \uparrow\downarrow \uparrow\downarrow \uparrow\downarrow \uparrow\downarrow \uparrow\downarrow \uparrow$

1s 2s 2p 3s 3p

c. nitrogen

N $1s^2 2s^2 2p^3$

$\uparrow\downarrow \uparrow\downarrow \uparrow \uparrow \uparrow$

1s 2s 2p

d. sodium

Na $1s^2 2s^2 2p^6 3s^1$

$\uparrow\downarrow \uparrow\downarrow \uparrow\downarrow \uparrow\downarrow \uparrow\downarrow \uparrow\downarrow \uparrow$

1s 2s 2p 3s

79. Use noble-gas notation to describe the electron configurations of the elements represented by the following symbols.

a. Mn

Mn $[\text{Ar}]4s^2 3d^5$

b. Kr

Kr $[\text{Ar}]4s^2 3d^{10} 4p^6$

c. P

P $[\text{Ne}]3s^2 3p^3$

d. Zn

Zn $[\text{Ar}]4s^2 3d^{10}$

e. Zr

Zr $[\text{Kr}]5s^2 4d^2$

f. W

W $[\text{Xe}]6s^2 4f^{14} 5d^4$

g. Pb

Pb $[\text{Xe}]6s^2 4f^{14} 5d^{10} 6p^2$

h. Ra

Ra $[\text{Rn}]7s^2$

i. Sm

Sm $[\text{Xe}]6s^2 4f^6$

j. Bk

Bk $[\text{Rn}]7s^2 5f^9$

Level 2

80. What elements are represented by each of the following electron configurations?

a. $1s^2 2s^2 2p^5$

F

b. $[\text{Ar}]4s^2$

Ca

c. $[\text{Xe}]6s^2 4f^4$

Nd

d. $[\text{Kr}]5s^2 4d^{10} 5p^4$

Te

e. $[\text{Rn}]7s^2 5f^{13}$

Md

f. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$

Br

81. Draw electron-dot structures for atoms of each of the following elements.

a. carbon

$\cdot\dot{\text{C}}\cdot$

b. arsenic

$\cdot\ddot{\text{As}}\cdot$

c. polonium



d. potassium



e. barium



82. An atom of arsenic has how many electron-containing orbitals? How many of the orbitals are completely filled? How many of the orbitals are associated with the atom's $n = 4$ principal energy?

An atom of arsenic ($1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^3$) has 18 electron-containing orbitals, 15 filled orbitals, and 4 orbitals associated with $n = 4$.

Mixed Review

Sharpen your problem-solving skills by answering the following.

83. What is the frequency of electromagnetic radiation having a wavelength of 1.00 m?

$$\begin{aligned}\nu &= c/\lambda = (3.00 \times 10^8 \text{ m/s})/(1.00 \text{ m}) \\ &= 3.00 \times 10^8 \text{ s}^{-1}\end{aligned}$$

84. What is the maximum number of electrons that can be contained in an atom's orbitals having the following principal quantum numbers?

a. 3

$$\begin{aligned}n &= 3 \\ \text{number of electrons} &= 2n^2 = 2(3)^2 \\ &= 18 \text{ electrons}\end{aligned}$$

b. 4

$$\begin{aligned}n &= 4 \\ \text{number of electrons} &= 2n^2 = 2(4)^2 \\ &= 32 \text{ electrons}\end{aligned}$$

c. 6

$$\begin{aligned}n &= 6 \\ \text{number of electrons} &= 2n^2 = 2(6)^2 \\ &= 72 \text{ electrons}\end{aligned}$$

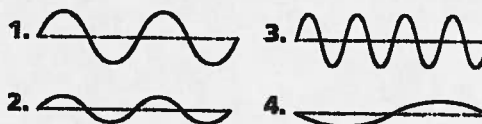
d. 7

$$\begin{aligned}n &= 7 \\ \text{number of electrons} &= 2n^2 = 2(7)^2 \\ &= 98 \text{ electrons}\end{aligned}$$

85. What is the wavelength with a frequency of 5.77×10^{14} Hz?

$$\begin{aligned}\lambda &= c/\nu = (3.00 \times 10^8 \text{ m/s})/(5.77 \times 10^{14} \text{ s}^{-1}) \\ &= 5.20 \times 10^{-7} \text{ m}\end{aligned}$$

86. Using the waves shown below, identify the wave or waves with the following characteristics.



a. longest wavelength

longest wavelength: 4

b. greatest frequency

greatest frequency: 3

c. largest amplitude

largest amplitude: 1 and 3

d. shortest wavelength

shortest wavelength: 3

87. How many orientations are possible for the orbitals related to each of the following sublevels?

a. s

1

b. p

3

c. d

5

d. f

7

Standardized Test Practice Chapter 5

page 149

Use these questions and the test-taking tip to prepare for your standardized test.

1. Cosmic rays are high-energy radiation from outer space. What is the frequency of a cosmic ray that has a wavelength of 2.67×10^{-13} m when it reaches Earth? (The speed of light is 3.00×10^8 m/s.)
- $8.90 \times 10^{-22} \text{ s}^{-1}$
 - $3.75 \times 10^{12} \text{ s}^{-1}$
 - $8.01 \times 10^{-5} \text{ s}^{-1}$
 - $1.12 \times 10^{21} \text{ s}^{-1}$

d

$$c = \lambda \nu$$

$$\nu = c/\lambda$$

$$\nu = (3.00 \times 10^8 \text{ m/s}) / (2.67 \times 10^{-13} \text{ m})$$

$$= 1.12 \times 10^{21} \text{ s}^{-1}$$

2. Wavelengths of light shorter than about 4.00×10^{-7} m are not visible to the human eye. What is the energy of a photon of ultraviolet light having a frequency of $5.45 \times 10^{16} \text{ s}^{-1}$? (Planck's constant is $6.626 \times 10^{-34} \text{ J} \cdot \text{s}$.)
- $3.61 \times 10^{-17} \text{ J}$
 - $1.22 \times 10^{-50} \text{ J}$
 - $8.23 \times 10^{49} \text{ J}$
 - $3.81 \times 10^{-24} \text{ J}$

a

$$E_{\text{photon}} = h\nu$$

$$E_{\text{photon}} = (6.626 \times 10^{-34} \text{ J} \cdot \text{s})(5.45 \times 10^{16} \text{ s}^{-1})$$

$$= 3.61 \times 10^{-17} \text{ J}$$

Interpreting Charts Use the periodic table and the chart below to answer questions 3–6.

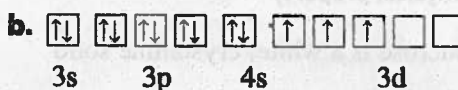
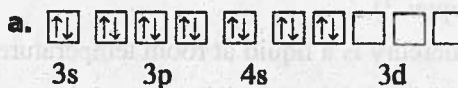
Electron Configurations for Selected Transition Metals

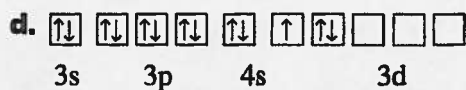
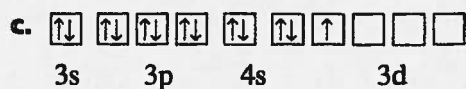
Element	Symbol	Atomic number	Electron configuration
Vanadium	V	23	[Ar]4s ² 3d ³
Yttrium	Y	39	[Kr]5s ² 4d ¹
_____	_____	_____	[Xe]6s ² 4f ¹⁴ 5d ⁶
Scandium	Sc	21	[Ar]4s ² 3d ¹
Cadmium	Cd	48	_____

3. Using noble-gas notation, the ground-state electron configuration of Cd is _____.
- [Kr]4d¹⁰4f²
 - [Ar]4s²3d¹⁰
 - [Kr]5s²4d¹⁰
 - [Xe]5s²4d¹⁰
- c
4. The element that has the ground-state electron configuration [Xe]6s²4f¹⁴5d⁶ is _____.
- La
 - Ti
 - W
 - Os
- d
5. The complete electron configuration of a scandium atom is _____.
- 1s²2s²2p⁶3s²3p⁶4s²3d¹
 - 1s²2s²2p⁷3s²3p⁷4s²3d¹
 - 1s²2s²2p⁵3s²3p⁵4s²3d¹
 - 1s²2s¹2p⁷3s¹3p⁷4s²3d¹

a

6. Which of the following is the correct orbital diagram for the third and fourth principal energy levels of vanadium?





b

7. Which of the following orbitals has the highest energy?

- a. 4f
- b. 5p
- c. 6s
- d. 3d

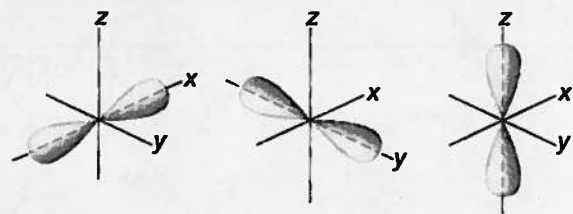
a

8. What is the electron-dot structure for indium?

- a. $\cdot\text{In}$
- b. $\cdot\text{In}\cdot$
- c. $\cdot\overset{\cdot}{\text{In}}\cdot$
- d. $\cdot\underset{\cdot}{\text{In}}\cdot$

c

9. The picture below shows all of the orbitals related to one type of sublevel. The type of sublevel to which these orbitals belong is _____.



- a. s
- b. p
- c. d
- d. f

b

10. What is the maximum number of electrons related to the fifth principal energy level of an atom?

- a. 10
- b. 20
- c. 25
- d. 50

d

number of electrons in a principal energy level

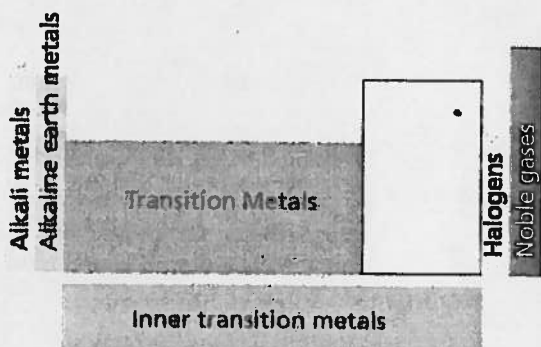
$$= 2n^2$$

$$2n^2 = 2(5^2) = 50 \text{ electrons}$$

34. Identify each of the elements in problem 31 as a representative element or a transition element. (6.1)

a. rep.; b. rep.; c. rep.; d. trans.; e. rep.; f. trans.

35. Sketch a simplified periodic table and use labels to identify the alkali metals, alkaline earth metals, transition metals, inner transition metals, noble gases, and halogens. (6.1)



36. A shiny solid element also is ductile. What side of the periodic table is it likely to be found? (6.1)

properties describe a metal; left of the stair step line

37. What are the general properties of a metalloid? List three metalloid elements. (6.1)

Metalloids have properties intermediate between metals and nonmetals. (B, Si, Ge, As, Sb, Te, Po, At) are metalloids.

38. What is the purpose of the heavy stair-step line on the periodic table? (6.1)

The line separates metals from nonmetals. Most elements bordering the line are metalloids.

39. Describe the two types of numbering used to identify groups on the periodic table. (6.1)

One system uses 1A–8A for representative elements, and 1B–8B for transition elements. The other system numbers the columns 1–18 left to right.

40. Give the chemical symbol of each of the following elements. (6.1)

a. the two elements that are liquids at room temperature

Br, Hg

b. the noble gas with the greatest atomic mass

Rn

c. any metal from group 4A

Sn or Pb

d. any inner transition metal

elements 58–71 or 90–103

41. Why do the elements chlorine and iodine have similar chemical properties? (6.2)

They have the same valence electron configuration (s^2p^5).

42. How are the numbers of valence electrons of the group A elements related to the group number? (6.2)

The number of valence electrons equals the group number for group A elements.

43. How is the energy level of an atom's valence electrons related to the period it is in on the periodic table? (6.2)

The energy level of an atom's valence electrons equals its period number.

44. How many valence electrons do each of the noble gases have? (6.2)

All noble gases have eight valence electrons, except for helium, which has two.

45. What are the four blocks of the periodic table? (6.2)

s-, p-, d-, and f-block

46. In general, what electron configuration has the greatest stability? (6.2)

ns^2np^6 , where n is the energy level

47. Determine the group, period, and block in which each of the following elements is located on the periodic table. (6.2)
- $[\text{Kr}]5s^24d^1$
3B, period 5, d-block
 - $[\text{Ar}]4s^23d^{10}4p^3$
5A, period 4, p-block
 - $[\text{He}]2s^22p^6$
8A, period 2, p-block
 - $[\text{Ne}]3s^23p^1$
3A, period 3, p-block
48. Categorize each of the elements in problem 47 as a representative element or a transition metal. (6.2)
- trans. metal, b. rep., c. rep., d. rep.
49. Explain how an atom's valence electron configuration determines its place on the periodic table. (6.2)
- Elements in a given column have the same number of valence electrons. The energy level of an atom's valence electrons determines its period.
50. Write the electron configuration for the element fitting each of the following descriptions. (6.2)
- the metal in group 5A
Bi: $[\text{Xe}]6s^24f^{14}5d^{10}6p^3$
 - the halogen in period 3
Cl: $[\text{Ne}]3s^23p^5$
 - the alkali metal in period 2
Li: $[\text{He}]2s^1$
 - the transition metal that is a liquid at room temperature
Hg: $[\text{Xe}]6s^24f^{14}5d^{10}$
51. Explain why the radius of an atom cannot be measured directly. (6.3)
- because the boundaries of an atom are indistinct
52. Given any two elements within a group, is the element with the larger atomic number likely to have a larger or smaller atomic radius than the other element? (6.2)
- larger
53. Which elements are characterized as having their d orbitals fill with electrons as you move left-to-right across a period? (6.2)
- transition metals
54. Explain why is it harder to remove an inner shell electron than a valence electron from an atom. (6.3)
- There are fewer shielding electrons between inner electrons and the nucleus. Thus, the inner electrons are more tightly bound to the nucleus by attractive electrostatic forces.
55. An element forms a negative ion when ionized. On what side of the periodic table is the element located? Explain. (6.3)
- Elements on the right side of periodic table gain electrons to gain a stable octet.
56. Of the elements magnesium, calcium, and barium, which forms the ion with the largest radius? The smallest? What periodic trend explains this? (6.3)
- Ba^{2+} is the largest; Mg^{2+} is the smallest; ionic size increases down a group.
57. What is ionization energy? (6.3)
- Ionization energy is the energy needed to remove an electron from a neutral atom in its gaseous state.
58. Explain why each successive ionization of an electron requires a greater amount of energy. (6.3)
- With each removed electron, there are fewer electrons to shield the remaining electrons from the electrostatic force of attraction of the nucleus. The increased nuclear attraction makes it more difficult to remove subsequent electrons.

59. Which group has the highest ionization energies? Explain why. (6.3)

The group 8A elements have the highest ionization energies because their electron configurations are the most stable.

60. Define an ion. (6.3)

An ion is an atom or a bonded group of atoms with a positive or negative charge.

61. How does the ionic radius of a nonmetal compare with its atomic radius? Explain why the change in radius occurs. (6.3)

The ionic radius of a nonmetal is larger than its neutral atom. Nonmetals tend to gain electrons in the atom's current energy level; these additional electrons repel each other and increase the size of the ion.

62. Explain why atomic radii decrease as you move left-to-right across a period. (6.3)

Atomic radii decrease left-to-right because the nuclear charge increases as the shielding of inner core electrons remains constant. The increased attraction of the nucleus for its electrons pulls the electrons inward, resulting in a decreased atomic size.

63. Which element in each pair has the larger ionization energy? (6.3)

a. Li, N

N

b. Kr, Ne

Ne

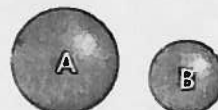
c. Cs, Li

Li

64. Explain the octet rule. (6.3)

The ns^2np^6 electron configuration, known as the octet configuration, contains eight electrons and generally has the lowest energy and is the most stable. Atoms gain, lose, or share electrons in order to obtain the stable octet configuration.

65. Use the illustration of spheres A and B to answer each of the following questions. Explain your reasoning for each answer. (6.3)



- a. If A is an ion, and B is an atom of the same element, is the ion a positive or negative ion?

The ion is negative. A negative ion is always larger than its atom.

- b. If A and B represent the atomic radii of two elements in the same period, what is their correct order (left-to-right)?

A is to the left of B. Atomic radius in a period decreases left-to-right.

- c. If A and B represent the ionic radii of two elements in the same group, what is their correct order (top-to-bottom)?

A is below B. Ionic radius increases down a group.

66. How many valence electrons do elements in each of the following groups have? (6.3)

a. group 8A

8

b. group 3A

3

c. group 1A

1

67. Na^+ and Mg^{2+} ions each have ten electrons surrounding their nuclei. Which ion would you expect to have the larger radius? Why? (6.3)

Na^+ has the larger radius. The greater nuclear charge of Mg^{2+} produces an increased inward pull on its ten electrons and results in a smaller radius.

84. Convert the following mass measurements as indicated. (Chapter 2)

a. 1.1 cm to meters

$$1.1 \text{ cm} \times \frac{1 \text{ m}}{100 \text{ cm}} = 1.1 \times 10^{-2} \text{ m}$$

b. 76.2 pm to millimeters

$$76.2 \text{ pm} \times \frac{1 \text{ m}}{10^{12} \text{ pm}} \times \frac{10^3 \text{ mm}}{1 \text{ m}} \\ = 7.62 \times 10^{-8} \text{ mm}$$

c. 11 Mg to kilograms

$$11 \text{ Mg} \times \frac{10^6 \text{ g}}{1 \text{ Mg}} \times \frac{1 \text{ kg}}{10^3 \text{ g}} = 1.1 \times 10^4 \text{ kg}$$

d. 7.23 micrograms to kilograms

$$7.23 \text{ } \mu\text{g} \times \frac{1 \text{ g}}{10^6 \text{ } \mu\text{g}} \times \frac{1 \text{ kg}}{10^3 \text{ g}} = 7.23 \times 10^{-9} \text{ kg}$$

85. How is the energy of a quantum of emitted radiation related to the frequency of the radiation? (Chapter 5)

The energy of a quantum equals the frequency times Planck's constant.

86. What element has the ground-state electron configuration of $[\text{Ar}]4s^23d^6$? (Chapter 5).

iron

Standardized Test Practice Chapter 6

page 177

Use these questions and the test-taking tip to prepare for your standardized test.

1. Periodic law states that elements show a _____.

- repetition of their physical properties when arranged by increasing atomic radius
- repetition of their chemical properties when arranged by increasing atomic mass
- periodic repetition of their properties when arranged by increasing atomic number
- periodic repetition of their properties when arranged by increasing atomic mass

c

2. Elements in the same group of the periodic table have the same _____

- number of valence electrons
- physical properties
- number of electrons
- electron configuration

a

3. All of the following are true EXCEPT _____.

- atomic radius of Na < atomic radius of Mg
- electronegativity of C > electronegativity of B
- ionic radius of Br^- > atomic radius of Br
- first ionization energy of K > first ionization energy of Rb

a

4. Which of the following is NOT true of an atom obeying the octet rule?

- obtains a full set of eight valence electrons
- acquires the valence configuration of a noble gas
- possesses eight electrons in total
- has a s^2p^6 valence configuration

c

5. What is the group, period, and block of an atom with the electron configuration $[\text{Ar}]4s^23d^{10}4p^4$?

- group 4A, period 4, d-block
- group 6A, period 3, p-block
- group 4A, period 4, p-block
- group 6A, period 4, p-block

d

6. Moving down a group on the periodic table, which two atomic properties follow the same trend?

- atomic radius and ionization energy
- ionic radius and atomic radius
- ionization energy and ionic radius
- ionic radius and electronegativity

b

Interpreting Tables Use the periodic table and the table at the bottom of the page to answer questions 7 and 8.

7. It can be predicted that silicon will experience a large jump in ionization energy after its _____.

- a. second ionization
- b. third ionization
- c. fourth ionization
- d. fifth ionization

c

8. Which of the following requires the most energy?

- a. second ionization of Li
- b. fourth ionization of N
- c. first ionization of Ne
- d. third ionization of Be

a

9. Niobium (Nb) is a(n) _____.

- a. nonmetal
- b. transition metal
- c. alkali metal
- d. halogen

b

10. It can be predicted that element 118 would have properties similar to a(n) _____.

- a. alkali earth metal
- b. halogen
- c. metalloid
- d. noble gas

d

Successive Ionization Energies for the Period 2 Elements

Element	Valence electrons	Ionization energy (kJ/mol)*								
		1st	2nd	3rd	4th	5th	6th	7th	8th	9th
Li	1	520	7300							
Be	2	900	1760	14 850						
B	3	800	2430	3660	25 020					
C	4	1090	2350	4620	6220	37 830				
N	5	1400	2860	4580	7480	9440	53 270			
O	6	1310	3390	5300	7470	10 980	13 330	71 330		
F	7	1680	3370	6050	8410	11 020	15 160	17 870	92 040	
Ne	8	2080	3950	6120	9370	12 180	15 240	20 000	23 070	115 380

* mol is an abbreviation for mole, a quantity of matter.

d. sulfurous acid



98. Write the formula for each of the following.

a. silicon dioxide



b. bromous acid



c. chlorine trifluoride



d. hydrobromic acid



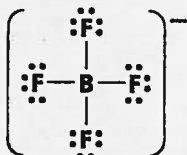
Lewis Structures (9.3)

99. Draw the Lewis structure for each of these molecules or ions.

a. H_2S



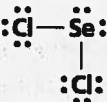
b. BF_4^-



c. SO_2

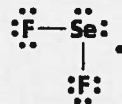


d. SeCl_2

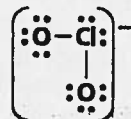


100. Draw the Lewis structure for each of these molecules or ions.

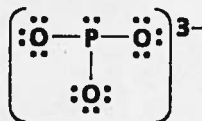
a. SeF_2



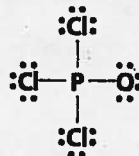
b. ClO_2^-



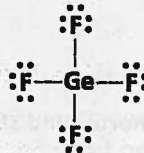
c. PO_3^{3-}



d. POCl_3



e. GeF_4

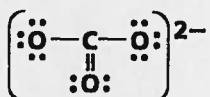
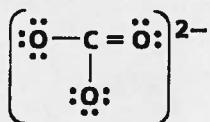
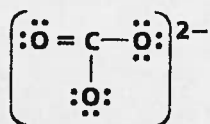


101. Which of the following elements are capable of forming molecules in which an atom has an expanded octet? Explain your answer.

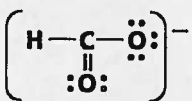
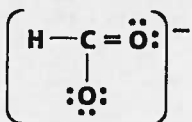
- B
- C
- P
- O
- Se

P and Se because they are period 3 and higher and have a d sublevel available

102. Draw three resonance structures for the polyatomic ion CO_3^{2-} .



103. Draw two resonance structures for the polyatomic ion CHO_2^- .



104. Draw the Lewis structure for each of the following molecules that have central atoms that do not obey the octet rule.

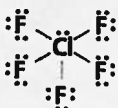
- a. PCl_5



- b. BF_3



- c. ClF_5



- d. BeH_2



Molecular Shape (9.4)

105. Predict the molecular shape and bond angle, and identify the hybrid orbitals for each of the following. Drawing the Lewis structure may help you.

- a. SCl_2

bent, 104.5° , sp^3

- b. NH_2Cl

trigonal pyramidal, 107° , sp^3

- c. HOF

bent, 104.5° , sp^3

- d. BF_3

trigonal planar, 120° , sp^2

106. For each of the following, predict the molecular shape.

- a. COS

linear

- b. CF_2Cl_2

tetrahedral

107. Identify the expected hybrid on the central atom for each of the following. Drawing the Lewis structure may help you.

- a. XeF_4

sp^3d^2

- b. TeF_4

sp^3d

- c. KrF_2

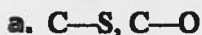
sp^3d

- d. OF_2

sp^3

Electronegativity and Polarity (9.5)

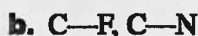
108. For each pair, indicate the more polar bond by circling the negative end of its dipole.



O is circled because it has the greatest electronegativity. C—O is the more polar bond. There is a greater electronegativity difference between C and O.

electronegativity S = 2.58
electronegativity C = 2.55
difference 0.03

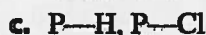
electronegativity O = 3.44
electronegativity C = 2.55
difference 0.89



F is circled because it has the greatest electronegativity. C—F is the more polar bond. There is a greater electronegativity difference between C and F.

electronegativity F = 3.98
electronegativity C = 2.55
difference 1.43

electronegativity N = 3.04
electronegativity C = 2.55
difference 0.49



Cl is circled because it has the greatest electronegativity. P—Cl is the more polar bond. There is a greater electronegativity difference between Cl and P.

electronegativity H = 2.20
electronegativity P = 2.19
difference 0.01

electronegativity Cl = 3.16
electronegativity P = 2.19
difference 0.97

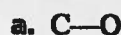
109. For each of the bonds listed, tell which atom is more negatively charged.

- a. C—H
b. C—N
c. C—S
d. C—O

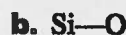
The most negatively charged atom has the greatest electronegativity. Use Figure 9-15.

- a. C b. N c. S d. O

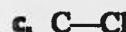
110. Predict which of the following bonds is the most polar.



electronegativity O = 3.44
electronegativity C = 2.55
difference 0.89



electronegativity O = 3.44
electronegativity Si = 1.90
difference 1.54



electronegativity Cl = 3.16
electronegativity C = 2.55
difference 0.61



electronegativity Br = 2.96
electronegativity C = 2.55
difference 0.41

Si—O is the most polar because it has the greatest electronegativity difference.

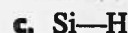
111. Rank the following bonds according to increasing polarity.



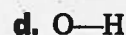
electronegativity C = 2.55
electronegativity H = 2.20
difference 0.35



electronegativity N = 3.04
electronegativity H = 2.20
difference 0.84



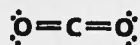
electronegativity H = 2.20
electronegativity Si = 1.90
difference 0.30



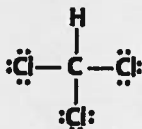
electronegativity O = 3.44
electronegativity H = 2.20
difference 1.24

2. Which of the following compounds contains at least one pi bond?

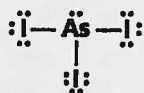
a. CO_2



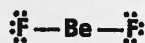
b. CHCl_3



c. AsI_3



d. BeF_2



CO_2 is the only choice that contains a double bond, so it is the only one containing a pi bond.

a

3. The Lewis structure for silicon disulfide is _____.

a. $\text{Si}:\ddot{\text{S}}:$

b. $\ddot{\text{S}}:\text{Si}:\ddot{\text{S}}$

c. $\ddot{\text{S}}:\text{Si}:\ddot{\text{S}}$

d. $:\text{S}:\text{Si}:\text{S}:$

b

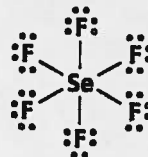
4. The central selenium atom in selenium hexafluoride forms an expanded octet. How many electron pairs surround the central Se atom?

a. 4

b. 5

c. 6

d. 7



c

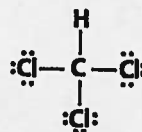
5. Chloroform (CHCl_3) was one of the first anesthetics used in medicine. The chloroform molecule contains 26 valence electrons in total. How many of these valence electrons take part in covalent bonds?

a. 26

b. 13

c. 8

d. 4



c

6. Which is the strongest type of intermolecular bond?

a. ionic bond

b. dipole-dipole force

c. dispersion force

d. hydrogen bond

d

26. $[\text{NO}] = 0.00446\text{M}$ and $[\text{H}_2] = 0.00282\text{M}$
 $[\text{NO}] = 0.00446\text{M}$
 $[\text{H}_2] = 0.00282\text{M}$
 $k = 2.90 \times 10^2 \text{ L}^2/(\text{mol}^2 \cdot \text{s})$
 $\text{Rate} = k [\text{NO}]^2 [\text{H}_2]$
 $= [2.90 \times 10^2 \text{ L}^2/(\text{mol}^2 \cdot \text{s})](0.00446\text{M})^2(0.00282\text{M})$
 $= [2.90 \times 10^2 \text{ L}^2/(\text{mol}^2 \cdot \text{s})]$
 $(0.00446 \text{ mol/L})^2(0.00282 \text{ mol/L})$
 $\text{Rate} = 1.63 \times 10^{-5} \text{ mol}/(\text{L} \cdot \text{s})$

Section 17.4 Assessment

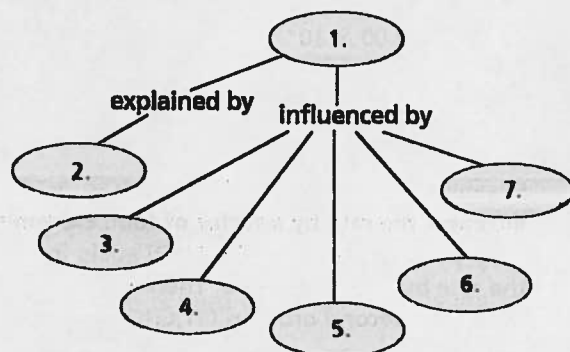
page 549

27. How can the rate law for a chemical reaction be used to determine an instantaneous reaction rate?
 Insert actual experimental rate constant and reactant concentrations into the rate law, then calculate.
28. Compare and contrast an elementary chemical reaction with a complex chemical reaction.
 An elementary reaction occurs in a single step. A complex reaction consists of two or more elementary steps.
29. What is a reaction mechanism? An intermediate?
 A reaction mechanism is the complete sequence of elementary steps that make up a complex reaction. An intermediate is a substance produced in one elementary step and consumed in a subsequent step.
30. **Thinking Critically** How can you determine whether a product of one of the elementary steps in a complex reaction is an intermediate?
 The product is consumed in a subsequent elementary step.
31. **Communicating** How would you explain the significance of the rate-determining step in a chemical reaction?
 Answers will vary but may include a description of how the rate-determining step limits the reaction rate, an analogy such as the "weakest link," and/or a real-world example.

Chapter 17 Assessment pages 554–556

Concept Mapping

32. Complete the following concept map using the following terms: surface area, collision theory, temperature, reaction rates, concentration, reactivity, catalyst.



1. reaction rates; 2. collision theory;
 3–7. surface area, temperature, concentration, reactivity, catalyst

Mastering Concepts

33. For a specific chemical reaction, assume that the change in free energy (ΔG) is negative. What does this information tell you about the rate of the reaction? (17.1)
 Although this information tells you the reaction is spontaneous under the specified conditions, it tells you nothing about the reaction rate.
34. How would you express the rate of the chemical reaction $\text{A} \rightarrow \text{B}$ based on the concentration of reactant A? How would that rate compare with the reaction rate based on the product B? (17.1)
 The rate may be expressed as the decrease in [A] per unit time: $\text{Rate} = \frac{\Delta[\text{A}]}{\Delta t}$.
 Numerically, the two rates would be equal; however, $\frac{\Delta[\text{A}]}{\Delta t}$ would be negative and $\frac{\Delta[\text{B}]}{\Delta t}$ would be positive.
35. What does the activation energy for a chemical reaction represent? (17.1)
 Activation energy represents the minimum amount of energy that reacting particles must have to form the activated complex.

- 36.** What is the role of the activated complex in a chemical reaction? (17.1)
The activated complex is a transition state between reactants and products.
- 37.** Suppose two molecules that can react collide. Under what circumstances do the colliding molecules not react? (17.1)
The molecules do not react if they collide with insufficient energy or lack favorable orientations at the instant of impact.
- 38.** How is the activation energy for a chemical reaction related to whether or not a collision between molecules initiates a reaction? (17.1)
The activation energy defines the minimum energy particles must have in order for collisions to result in the formation of the activated complex.
- 39.** In the activated complex for a chemical reaction, what bonds are broken and what bonds are formed? (17.1)
Reactant bonds are being broken and product bonds are being formed.
- 40.** If $A \rightarrow B$ is exothermic, how does the activation energy for the forward reaction compare with the activation energy for the reverse reaction ($A \leftarrow B$)? (17.1)
The activation energy for the forward reaction is less than the activation energy for the reverse reaction.
- 41.** What role does the reactivity of the reactants play in determining the rate of a chemical reaction? (17.2)
Generally, the greater the reactivity of the reactants, the faster the reaction.
- 42.** Explain why a crushed solid reacts with a gas more quickly than a large chunk of the same solid. (17.2)
The finely divided solid presents a larger surface area than the large chunk.
- 43.** What do you call a substance that increases the rate of a chemical reaction without being consumed in the reaction? (17.2)
a catalyst
- 44.** In general, what is the relationship between reaction rate and reactant concentration? (17.2)
Increasing the concentration increases the rate, and decreasing the concentration decreases the rate.
- 45.** In general, what is the relationship between reaction rate and temperature? (17.2)
Increasing the temperature increases the rate, and decreasing the temperature decreases the rate.
- 46.** Distinguish between a homogeneous catalyst and a heterogeneous catalyst. (17.2)
A homogeneous catalyst exists in the same physical state as the reaction it catalyzes; a heterogeneous catalyst exists in a different physical state from the reaction it catalyzes.
- 47.** Explain how a catalyst affects the activation energy for a chemical reaction. (17.2)
A catalyst lowers the activation energy.
- 48.** Use the collision theory to explain why increasing the concentration of a reactant usually increases the reaction rate. (17.2)
Increasing the concentration of a reactant increases the frequency of collisions between reacting particles.
- 49.** Use the collision theory to explain why increasing the temperature usually increases the reaction rate. (17.2)
Increasing the temperature increases collision frequency and collision energy.
- 50.** In a chemical reaction, what relationship does the rate law describe? (17.3)
The rate law defines the mathematical relationship between the reaction rate and the concentration of the reactants.

51. What is the name of the proportionality constant in the mathematical expression that relates reaction rate and reactant concentration? (17.3)

The specific rate constant

52. What does the order of a reactant tell you about the way the concentration of that reactant appears in the rate law? (17.3)

The reactant's order defines the exponent to which the reactant's concentration is raised in the rate law.

53. Why does the specific rate constant for a chemical reaction often double for each increase of 10 K? (17.3)

For many reactions, an increase of 10 K approximately doubles the number of collisions having sufficient energy to initiate reaction.

54. Explain why the rates of most chemical reactions decrease over time. (17.3)

As reactants are consumed, their concentrations decrease over time.

55. In the method of initial rates used to determine the rate law for a chemical reaction, what is the significance of the word initial? (17.3)

The initial rate is the instantaneous rate at the stated concentrations; however, the rate begins to decrease the instant the reaction starts.

56. If a reaction has three reactants and is first order in one, second order in another, and third order in the third, what is the overall order of the reaction? (17.3)

sixth order

57. What do you call the slowest of the elementary steps that make up a complex reaction? (17.4)

the rate-determining step or rate-limiting step

58. What is an intermediate in a complex reaction? (17.4)

An intermediate is a substance produced in one elementary step and consumed in a subsequent elementary step.

59. Distinguish between an elementary step, a complex reaction, and a reaction mechanism. (17.4)

A simple reaction occurs in a single step; a complex reaction is made up of two or more elementary steps; a reaction mechanism consists of the complete sequence of elementary steps that make up a complex reaction.

60. Under what circumstances is the rate law for the reaction $2A + 3B \rightarrow$ products correctly written as $\text{Rate} = k[A]^2[B]^3$? (17.3)

The rate law is correct if the reaction occurs in a single step. The rate law is possible, but unlikely, if the reaction occurs in multiple steps.

61. How does the activation energy of the rate-determining step in a complex reaction compare with the activation energies of the other elementary steps? (17.4)

The rate-determining step has the highest activation energy.

Mastering Problems

A Model for Reaction Rates (17.1)

Level 1

62. In the gas-phase reaction $I_2 + Cl_2 \rightarrow 2ICl$, the $[I_2]$ changes from $0.400M$ at time = 0 to $0.300M$ at time = 4.00 min. Calculate the average reaction rate in moles I_2 consumed per liter per minute.

$[I_2]$ is consumed. Average reaction rate must be positive.

average reaction rate

$$= -\frac{[I_2] \text{ at time } t_2 - [I_2] \text{ at time } t_1}{t_2 - t_1} = -\frac{\Delta[I_2]}{\Delta t}$$

$$= -\frac{\Delta[I_2]}{\Delta t} = -\frac{0.300M - 0.400M}{4.00 \text{ min} - 0.00 \text{ min}} = -\frac{-0.100M}{4.00 \text{ min}}$$

$$= 0.0250 \text{ mol}/(L \cdot \text{min})$$

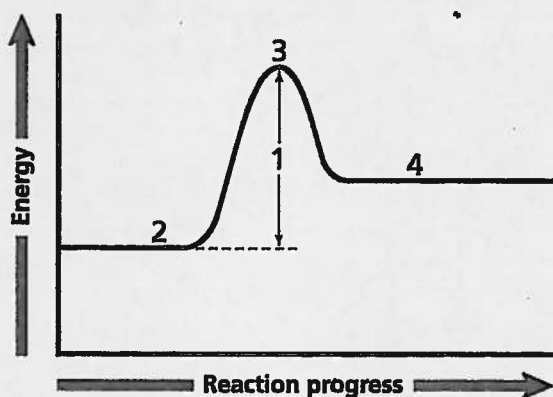
The average reaction rate is 0.0250 mole I_2 consumed per liter per minute.

63. If a chemical reaction occurs at a rate of 2.25×10^{-2} moles per liter per second at 322 K, what is the rate expressed in moles per liter per minute?

$$\begin{aligned} \text{rate} &= 2.25 \times 10^{-2} \text{ mol}/(\text{L}\cdot\text{s}) \times \frac{60 \text{ sec}}{1 \text{ min}} \\ &= 1.35 \text{ mol}/(\text{L}\cdot\text{min}) \end{aligned}$$

Level 2

64. On the accompanying energy level diagram, match the appropriate number with the quantity it represents.



- a. reactants
2
- b. activated complex
3
- c. products
4
- d. activation energy
1

65. Given the following data for the decomposition of hydrogen peroxide, calculate the average reaction rate in moles H_2O_2 consumed per liter per minute for each time interval.

Decomposition of H_2O_2	
Time (min)	$[\text{H}_2\text{O}_2]$ (M)
0	2.50
2	2.12
5	1.82
10	1.48
20	1.00

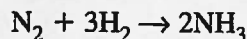
$$\begin{aligned} \text{First interval: } & - \frac{\Delta[\text{H}_2\text{O}_2]}{\Delta t} \\ &= - \frac{2.12 \text{ mol/L} - 2.50 \text{ mol/L}}{2 \text{ min} - 0 \text{ min}} = 0.2 \text{ mol}/(\text{L}\cdot\text{min}) \end{aligned}$$

$$\begin{aligned} \text{Second interval: } & - \frac{\Delta[\text{H}_2\text{O}_2]}{\Delta t} \\ &= - \frac{1.82 \text{ mol/L} - 2.12 \text{ mol/L}}{5 \text{ min} - 2 \text{ min}} = 0.1 \text{ mol}/(\text{L}\cdot\text{min}) \end{aligned}$$

$$\begin{aligned} \text{Third interval: } & - \frac{\Delta[\text{H}_2\text{O}_2]}{\Delta t} \\ &= - \frac{1.48 \text{ mol/L} - 1.82 \text{ mol/L}}{10 \text{ min} - 5 \text{ min}} = 0.07 \text{ mol}/(\text{L}\cdot\text{min}) \end{aligned}$$

$$\begin{aligned} \text{Fourth interval: } & - \frac{\Delta[\text{H}_2\text{O}_2]}{\Delta t} \\ &= - \frac{1.00 \text{ mol/L} - 1.48 \text{ mol/L}}{20 \text{ min} - 10 \text{ min}} = 0.05 \text{ mol}/(\text{L}\cdot\text{min}) \end{aligned}$$

66. At a given temperature and for a specific time interval, the average rate of the following reaction is 1.88×10^{-4} moles N_2 consumed per liter per second.



Express the reaction rate in moles H_2 consumed per liter per second and in moles NH_3 produced per liter per second.

$$\begin{aligned} & 1.88 \times 10^{-4} \text{ mol-N}_2/(\text{L}\cdot\text{s}) \times \frac{3 \text{ mol H}_2}{1 \text{ mol-N}_2} \\ &= 5.64 \times 10^{-4} \text{ mol H}_2/(\text{L}\cdot\text{s}) \end{aligned}$$

$$\begin{aligned} & 1.88 \times 10^{-4} \text{ mol-N}_2/(\text{L}\cdot\text{s}) \times \frac{2 \text{ mol NH}_3}{1 \text{ mol-N}_2} \\ &= 3.76 \times 10^{-4} \text{ mol NH}_3/(\text{L}\cdot\text{s}) \end{aligned}$$

Factors Affecting Reaction Rates (17.2)

Level 1

67. Estimate the rate of the reaction described in problem 63 at 332 K. Express the rate in moles per liter per second.

The temperature has increased by 10 K, which should cause the reaction rate to increase by about a factor of two. Therefore, approximate rate at 332 K = $[2.25 \times 10^{-2} \text{ mol}/(\text{L}\cdot\text{s})] \times 2 = 4.50 \times 10^{-2} \text{ mol}/(\text{L}\cdot\text{s})$.

Level 2

68. Estimate the rate of the reaction described in problem 63 at 352 K and with $[\text{I}_2]$ doubled (assume the reaction is first order in I_2).

The temperature has increased by 30 K, which should cause the reaction rate to increase by about a factor of 2^3 , and the $[\text{I}_2]$ has doubled, which should increase the rate by a factor of 2. Therefore, approximate rate at 352 K = $[2.25 \times 10^{-2} \text{ mol}/(\text{L}\cdot\text{s})] \times 2^3 \times 2 = 0.360 \text{ mol}/(\text{L}\cdot\text{s})$.

Reaction Rate Laws (17.3)

Level 1

69. Nitrogen monoxide gas and chlorine gas react according to the equation $2\text{NO} + \text{Cl}_2 \rightarrow 2\text{NOCl}$. Use the following data to determine the rate law for the reaction by the method of initial rates. Also, calculate the value of the specific rate constant.

2NO + Cl ₂ Reaction Data		
Initial [NO] (M)	Initial [Cl ₂] (M)	Initial Rate (mol/(L·min))
0.50	0.50	1.90×10^{-2}
1.00	0.50	7.60×10^{-2}
1.00	1.00	15.20×10^{-2}

Examining the first two trials, doubling [NO] increases the rate by a factor of four; therefore, the reaction is second order in NO. Examining the second and third trials, doubling $[\text{Cl}_2]$ doubles the rate; therefore, the reaction is first order in Cl_2 .

$$\text{Rate} = k[\text{NO}]^2[\text{Cl}_2]$$

$$k = \frac{\text{Rate}}{[\text{NO}]^2[\text{Cl}_2]} = 0.152 \text{ L}^2/(\text{mol}^2\cdot\text{min})$$

Level 2

70. Use the following data to determine the rate law and specific rate constant for the reaction $2\text{ClO}_2(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{ClO}_3^-(\text{aq}) + \text{ClO}_2^-(\text{aq}) + \text{H}_2\text{O}(\text{l})$.

2ClO ₂ (aq) + 2OH ⁻ (aq) Reaction Data		
Initial [ClO ₂] (M)	Initial [OH ⁻] (M)	Initial Rate (mol/(L·min))
0.0500	0.200	6.90
0.100	0.200	27.6
0.100	0.100	13.8

Examining the first two trials, doubling $[\text{ClO}_2]$ increases the rate by a factor of four; therefore, the reaction is second order in ClO_2 . Examining the second and third trials, halving $[\text{OH}^-]$ halves the rate; therefore, the reaction is first order in OH^- .

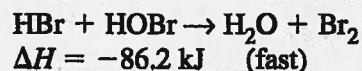
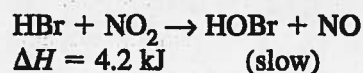
$$\text{Rate} = k[\text{ClO}_2]^2[\text{OH}^-]$$

$$k = \frac{\text{Rate}}{[\text{ClO}_2]^2[\text{OH}^-]} = 1.38 \times 10^4 \text{ L}^2/(\text{mol}^2\cdot\text{min})$$

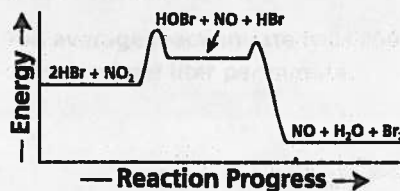
Instantaneous Reaction Rates and Reaction Mechanisms (17.4)

Level 1

71. The gas-phase reaction $2\text{HBr} + \text{NO}_2 \rightarrow \text{H}_2\text{O} + \text{NO} + \text{Br}_2$ is thought to occur by the following mechanism.



Draw the energy diagram that depicts this reaction mechanism. On the diagram, show the energy of the reactants, energy of the products, and relative activation energies of the two elementary steps.



83. Write an advertisement that explains why Company A's lawn care product (fertilizer or weed killer) works better than the competition's because of the smaller sized granules. Include applicable diagrams.

Student answers will vary, but they will emphasize that, for the same mass, Company A's product presents a greater surface area and therefore reacts more quickly to kill weeds.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

84. Classify each of the following elements as a metal, nonmetal, or metalloid. (Chapter 6)
- molybdenum
metal
 - bromine
nonmetal
 - arsenic
metalloid
 - neon
nonmetal
 - cerium
metal
85. Balance the following equations. (Chapter 10)
- $$\text{Sn}(s) + \text{NaOH}(aq) \rightarrow \text{Na}_2\text{SnO}_2 + \text{H}_2$$

$$\text{Sn}(s) + 2\text{NaOH}(aq) \rightarrow \text{Na}_2\text{SnO}_2 + \text{H}_2$$
 - $$\text{C}_8\text{H}_{18}(l) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l)$$

$$2\text{C}_8\text{H}_{18} + 25\text{O}_2 \rightarrow 16\text{CO}_2 + 18\text{H}_2\text{O}$$
 - $$\text{Al}(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{Al}_2(\text{SO}_4)_3(aq) + \text{H}_2(g)$$

$$2\text{Al} + 3\text{H}_2\text{SO}_4 \rightarrow \text{Al}_2(\text{SO}_4)_3 + 3\text{H}_2$$
86. What mass of iron(III) chloride is needed to prepare 1.00 L of a 0.255M solution? (Chapter 15)
- $$\text{molar mass} = 55.85 \text{ g/mol} + 3(35.45 \text{ g/mol}) = 162.20 \text{ g/mol FeCl}_3$$
- $$1.00 \cancel{\text{L}} \times \frac{0.255 \text{ mol-FeCl}_3}{1 \cancel{\text{L}}} \times \frac{162.20 \text{ g FeCl}_3}{1 \text{ mol-FeCl}_3} = 41.4 \text{ g FeCl}_3$$
87. ΔH for a reaction is negative. Compare the energy of the products and the reactants. Is the reaction endothermic or exothermic? (Chapter 16)
- When ΔH for a reaction is negative, the energy of the products is less than the energy of the reactants, and the reaction is exothermic.

Standardized Test Practice

Chapter 17

page 557

Use these questions and the test-taking tip to prepare for your standardized test.

1. The rate of a chemical reaction is all of the following EXCEPT _____.
- the speed at which a reaction takes place
 - the change in concentration of a reactant per unit time
 - the change in concentration of a product per unit time
 - the amount of product formed in a certain period of time

d

6. How long will it take for half of the original amount of SO_2Cl_2 to decompose at the average reaction rate?

- a. 285 min
b. 335 min
c. 385 min
d. 500 min

$$\text{Avg rate} = \frac{\Delta \text{qty}}{\Delta t}$$

$$1.30 \times 10^{-3} \text{ M/min} = -\frac{0.50\text{M} - 1.00\text{M}}{t - 0.00} = 385 \text{ min}$$

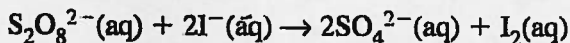
c

7. Which of the following does NOT affect reaction rate?

- a. catalysts
b. surface area of reactants
c. concentration of reactants
d. reactivity of products

d

8. The reaction between persulfate ($\text{S}_2\text{O}_8^{2-}$) and iodide (I^-) ions is often studied in student laboratories because it occurs slowly enough for its rate to be measured:



This reaction has been experimentally determined to be first order in $\text{S}_2\text{O}_8^{2-}$ and first order in I^- . Therefore, the overall rate law for this reaction is

- a. Rate = $k[\text{S}_2\text{O}_8^{2-}]^2[\text{I}^-]$
b. Rate = $k[\text{S}_2\text{O}_8^{2-}][\text{I}^-]$
c. Rate = $k[\text{S}_2\text{O}_8^{2-}][\text{I}^-]^2$
d. Rate = $k[\text{S}_2\text{O}_8^{2-}]^2[\text{I}^-]^2$

b

9. The rate law for the reaction $\text{A} + \text{B} + \text{C} \rightarrow$ products is: Rate = $k[\text{A}]^2[\text{C}]$

If $k = 6.92 \times 10^{-5} \text{ L}^3/(\text{mol}^3 \cdot \text{s})$, $[\text{A}] = 0.175\text{M}$, $[\text{B}] = 0.230\text{M}$, and $[\text{C}] = 0.315\text{M}$, what is the instantaneous reaction rate?

- a. $6.68 \times 10^{-7} \text{ mol}/(\text{L} \cdot \text{s})$
b. $8.77 \times 10^{-7} \text{ mol}/(\text{L} \cdot \text{s})$
c. $1.20 \times 10^{-6} \text{ mol}/(\text{L} \cdot \text{s})$
d. $3.81 \times 10^{-6} \text{ mol}/(\text{L} \cdot \text{s})$

$$\text{Rate} = k[\text{A}]^2[\text{C}]$$

$$= (6.92 \times 10^{-5} \text{ L}^3/\text{mol}^3 \cdot \text{s})(0.175\text{M})^2(0.315\text{M})$$

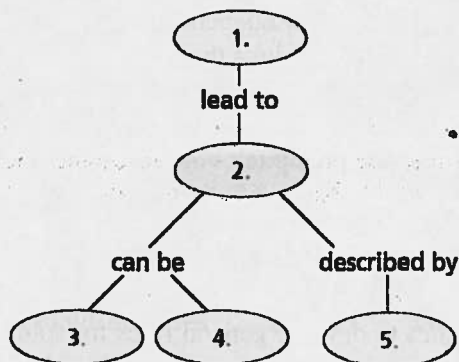
$$= 6.68 \times 10^{-7} \text{ mol}/(\text{L} \cdot \text{s})$$

a

Chapter 18 Assessment pages 590–592

Concept Mapping

25. Fill in the spaces on the concept map with the following phrases: equilibrium constant expressions, reversible reactions, heterogeneous equilibria, homogeneous equilibria, chemical equilibria.



1. reversible reactions; 2. chemical equilibria;
3. homogeneous equilibria; 4. heterogeneous equilibria; 5. equilibrium constant expressions

Mastering Concepts

26. Describe an equilibrium in everyday life that illustrates a state of balance between two opposing processes. (18.1)
situations such as vehicle traffic across a bridge
27. Given the fact that the concentrations of reactants and products are not changing, why is the word *dynamic* used for describing chemical equilibrium? (18.1)
Reactants are continuing to form products and products are continuing to form reactants.
28. How can you indicate in a chemical equation that a reaction is reversible? (18.1)
Use double arrows.
29. Although the general equation for a chemical reaction is reactants \rightarrow products, explain why this equation is not complete for a system at equilibrium. (18.1)
The equation should have double arrows.
30. Explain the difference between a homogeneous equilibrium and a heterogeneous equilibrium. (18.1)
homogeneous equilibrium: all reactants and products are in the same physical state;
heterogeneous equilibrium system: reactants and products exist in more than one physical state.
31. What is an equilibrium position? (18.1)
a particular set of equilibrium concentrations
32. Explain how to use the law of chemical equilibrium in writing an equilibrium constant expression. (18.1)
the ratio of product concentrations to reactant concentrations, with each concentration raised to the power corresponding to its coefficient in the balanced equation
33. Why does a numerically large K_{eq} mean that the products are favored in an equilibrium system? (18.1)
A numerically large K_{eq} means that product concentrations, which are in the numerator, are greater than reactant concentrations, which are in the denominator.
34. Why should you pay attention to the physical states of all reactants and products when writing equilibrium constant expressions? (18.1)
The concentrations of pure liquids or solids are omitted from the equilibrium constant expression.
35. How can an equilibrium system contain small and unchanging amounts of products yet have large amounts of reactants? What can you say about the relative size of K_{eq} for such an equilibrium? (18.1)
Such an equilibrium can exist if the small amounts of products react so quickly that the rate of the reverse reaction is quickly equal to the rate of the forward reaction. K_{eq} must have a small numerical value.
36. Describe the opposing processes in the physical equilibrium that exists in a closed container half-filled with liquid ethanol. (18.1)
the evaporation and condensation of ethanol

- 37.** What is meant by a stress on a reaction at equilibrium? (18.2)
 any change in concentration, volume (pressure), or temperature which causes the equilibrium to shift.
- 38.** How does Le Châtelier's principle describe an equilibrium's response to a stress? (18.2)
 It will shift in the direction that relieves the stress.
- 39.** Why does removing a product cause an equilibrium to shift in the direction of the products? (18.2)
 In order to restore the equilibrium amount of product, the equilibrium shifts to produce more product.
- 40.** When an equilibrium shifts toward the reactants in response to a stress, how is the equilibrium position changed? (18.2)
 The equilibrium position shifts to the left.
- 41.** Use Le Châtelier's principle to explain how a shift in the equilibrium $\text{H}_2\text{CO}_3(\text{aq}) \rightleftharpoons \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$ causes a soft drink to go flat when its container is left open to the atmosphere. (18.2)
 Because $\text{CO}_2(\text{g})$ continually escapes from the open container, the equilibrium shifts to the right until $\text{H}_2\text{CO}_3(\text{aq})$ is depleted.
- 42.** How is K_{eq} changed when heat is added to an equilibrium in which the forward reaction is exothermic? Explain using Le Châtelier's principle. (18.2)
 K_{eq} is lowered. Adding heat increases the rate of the heat-absorbing reaction, which is the direction of the reactants.
- 43.** Changing the volume of the system alters the equilibrium position of this equilibrium.
 $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$
 But a similar change has no effect on this equilibrium.
 $\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightleftharpoons 2\text{HCl}(\text{g})$
 Explain. (18.2)
 The number of gaseous particles reacting in the first equilibrium is different from the number of gaseous particles produced; therefore, the system can shift in response to a change in volume (pressure). The second equilibrium cannot shift in response to a change in volume because the number of gaseous particles reacting equals the number produced.
- 44.** How might the addition of a noble gas to the reaction vessel affect this equilibrium?
 $2\text{N}_2\text{H}_4(\text{g}) + 2\text{NO}_2(\text{g}) \rightleftharpoons 3\text{N}_2(\text{g}) + 4\text{H}_2\text{O}(\text{g})$
 Assume that the volume of the reaction vessel does not change. (18.2)
 Addition of a noble gas does not change the partial pressures or concentrations of the reactants or products, so the equilibrium is not affected.
- 45.** When an equilibrium shifts to the right, what happens to the following? (18.2)
 a. the concentrations of the reactants
 decrease
 b. the concentrations of the products
 increase
- 46.** How would each of the following changes affect the equilibrium position of the system used to produce methanol from carbon monoxide and hydrogen? (18.2)
 $\text{CO}(\text{g}) + 2\text{H}_2(\text{g}) \rightleftharpoons \text{CH}_3\text{OH}(\text{g}) + \text{heat}$
 a. adding CO to the system
 The equilibrium shifts to the right.
 b. cooling the system
 The equilibrium shifts to the right.

- c. adding a catalyst to the system

The equilibrium does not shift.

- d. removing CH_3OH from the system

The equilibrium shifts to the right.

- e. decreasing the volume of the system

The equilibrium shifts to the right.

47. Why is the concentration of a solid not included as part of the solubility product constant? (18.3)

The concentration of a solid is a constant value.

48. What does it mean to say that two solutions have a common ion? Give an example that supports your answer. (18.3)

They both contain the same ion. $\text{NaCl}(\text{aq})$ and $\text{KCl}(\text{aq})$ both contain $\text{Cl}^-(\text{aq})$.

49. Explain the difference between Q_{sp} and K_{sp} . (18.3)

K_{sp} describes a saturated solution. Q_{sp} describes a solution with any concentration of ions.

50. Explain why a common ion lowers the solubility of an ionic compound. (18.3)

If the concentration of either ion produced by the dissolving compound is increased by a common ion present in the solution, the concentration of the compound's other ion must decrease.

51. Describe the solution that results when two solutions are mixed and Q_{sp} is found to equal K_{sp} . Does a precipitate form?

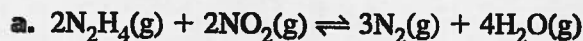
The new solution is saturated. No precipitate will form.

Mastering Problems

The Equilibrium Constant Expression (18.1)

Level 1

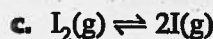
52. Write equilibrium constant expressions for these homogeneous equilibria.



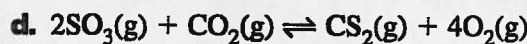
$$K_{\text{eq}} = \frac{[\text{N}_2]^3[\text{H}_2\text{O}]^4}{[\text{N}_2\text{H}_4]^2[\text{NO}_2]^2}$$



$$K_{\text{eq}} = \frac{[\text{NbCl}_3][\text{NbCl}_5]}{[\text{NbCl}_4]^2}$$

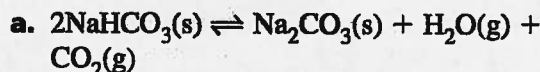


$$K_{\text{eq}} = \frac{[\text{I}]^2}{[\text{I}_2]}$$

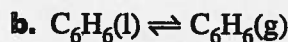


$$K_{\text{eq}} = \frac{[\text{CS}_2][\text{O}_2]^4}{[\text{SO}_3]^2[\text{CO}_2]}$$

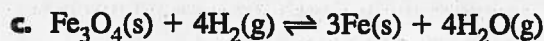
53. Write equilibrium constant expressions for these heterogeneous equilibria.



$$K_{\text{eq}} = [\text{H}_2\text{O}][\text{CO}_2]$$



$$K_{\text{eq}} = [\text{C}_6\text{H}_6(\text{g})]$$



$$K_{\text{eq}} = \frac{[\text{H}_2\text{O}]^4}{[\text{H}_2]^4}$$

54. Pure water has a density of 1.00 g/mL at 297 K. Calculate the molar concentration of pure water at this temperature.

$$\frac{1.00 \text{ g H}_2\text{O}}{1 \text{ mL}} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 55.5 \text{ mol/L}$$

Level 2

55. Calculate K_{eq} for the following equilibrium when $[SO_3] = 0.0160 \text{ mol/L}$, $[SO_2] = 0.00560 \text{ mol/L}$, and $[O_2] = 0.00210 \text{ mol/L}$.



$$K_{eq} = \frac{[SO_2]^2[O_2]}{[SO_3]^2} = \frac{(0.00560)^2(0.00210)}{(0.0160)^2} = 2.57 \times 10^{-4}$$

56. K_{eq} for this reaction is 3.63.



The data in the table show the concentrations of the reactants and product in two different reaction mixtures at the same temperature. Does the data provide evidence that both reactions are at equilibrium?

Concentrations of A, B, and C		
A (mol/L)	B (mol/L)	C (mol/L)
0.500	0.621	0.700
0.250	0.525	0.250

Calculate K_{eq} using the two sets of data.

$$K_{eq} = \frac{[C]}{[A][B]^2} = 3.63$$

$$\frac{(0.700)}{(0.500)(0.621)^2} = 3.63$$

$$\frac{(0.250)}{(0.250)(0.525)^2} = 3.63$$

Both reactions are at equilibrium.

57. When solid ammonium chloride is put in a reaction vessel at 323 K, the equilibrium concentrations of both ammonia and hydrogen chloride are found to be 0.0660 mol/L. $NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$. Calculate K_{eq} .
- $$K_{eq} = [NH_3][HCl] = (0.0660)(0.0660) = 4.36 \times 10^{-3}$$

58. Suppose you have a cube of pure manganese metal measuring 5.25 cm on each side. You find that the mass of the cube is 1076.6 g. What is the molar concentration of manganese in the cube?

$$\text{Volume of the cube} = (5.25 \text{ cm})^3 = 145 \text{ cm}^3$$

$$145 \text{ cm}^3 \times \frac{1 \text{ L}}{1000 \text{ cm}^3} = 0.145 \text{ L}$$

$$\text{Molar mass of manganese} = 54.94 \text{ g/mol}$$

$$1076.6 \text{ g Mn} \times \frac{1 \text{ mol Mn}}{54.94 \text{ g Mn}} = 19.596 \text{ mol Mn}$$

$$\frac{19.596 \text{ mol}}{0.145 \text{ L}} = 135 \text{ mol/L Mn}$$

Le Châtelier's Principle (18.2)

Level 1

59. Use Le Châtelier's principle to predict how each of the following changes would affect this equilibrium.



- adding $H_2O(g)$ to the system
shift to the left
- removing $CO(g)$ from the system
shift to the right
- adding $H_2(g)$ to the system
shift to the right
- adding something to the system to absorb $CO_2(g)$
shift to the left

60. How would increasing the volume of the reaction vessel affect these equilibria?

- $NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$
shift to the right
- $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$
no shift

61. How would decreasing the volume of the reaction vessel affect these equilibria?
- $2\text{N}_2\text{H}_4(\text{g}) + 2\text{NO}_2(\text{g}) \rightleftharpoons 3\text{N}_2(\text{g}) + 4\text{H}_2\text{O}(\text{g})$
shift to the left
 - $2\text{H}_2\text{O}(\text{g}) \rightleftharpoons 2\text{H}_2(\text{g}) + \text{O}_2(\text{g})$
shift to the left
62. How would these equilibria be affected by increasing the temperature?
- $4\text{NH}_3(\text{g}) + 5\text{O}_2(\text{g}) \rightleftharpoons 4\text{NO}(\text{g}) + 6\text{H}_2\text{O}(\text{g}) + \text{heat}$
shift to the left, decreasing K_{eq}
 - heat + $\text{NaCl}(\text{s}) \rightleftharpoons \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq})$
shift to the right, increasing K_{eq}

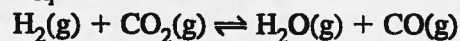
Level 2

63. Ethylene (C_2H_4) reacts with hydrogen to form ethane (C_2H_6).
- $$\text{C}_2\text{H}_4(\text{g}) + \text{H}_2(\text{g}) \rightleftharpoons \text{C}_2\text{H}_6(\text{g}) + \text{heat}$$
- How would you regulate the temperature of this equilibrium in order to do the following?
- increase the yield of ethane
lower the temperature
 - decrease the concentration of ethylene
lower the temperature
 - increase the amount of hydrogen in the system
raise the temperature
64. How would simultaneously decreasing the temperature and volume of the system affect these equilibria?
- heat + $\text{CaCO}_3(\text{s}) \rightleftharpoons \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$
shift to the left
 - $4\text{NH}_3(\text{g}) + 5\text{O}_2(\text{g}) \rightleftharpoons 4\text{NO}(\text{g}) + 6\text{H}_2\text{O}(\text{g}) + \text{heat}$
Decreasing the temperature causes a shift to the right and decreasing the volume causes a shift to the left.

Calculations Using K_{eq} (18.3)

Level 1

65. K_{eq} is 1.60 at 933 K for this reaction.



Calculate the equilibrium concentration of hydrogen when $[\text{CO}_2] = 0.320 \text{ mol/L}$, $[\text{H}_2\text{O}] = 0.240 \text{ mol/L}$, and $[\text{CO}] = 0.280 \text{ mol/L}$.

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

$$1.60 = \frac{(0.240)(0.280)}{[\text{H}_2](0.320)}$$

$$[\text{H}_2] = 0.131\text{M}$$

66. At 2273 K, $K_{\text{eq}} = 6.2 \times 10^{-4}$ for the reaction
- $$\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{NO}(\text{g})$$
- If $[\text{N}_2] = 0.05200 \text{ mol/L}$ and $[\text{O}_2] = 0.00120 \text{ mol/L}$, what is the concentration of NO at equilibrium?

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

$$6.2 \times 10^{-4} = \frac{[\text{NO}]^2}{(0.05200)(0.00120)}$$

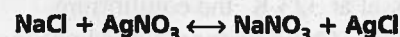
$$[\text{NO}]^2 = 3.9 \times 10^{-8}$$

$$[\text{NO}] = \sqrt{3.9 \times 10^{-8}} = 2.0 \times 10^{-4}\text{M}$$

Calculations Using K_{sp} (18.3)

Level 1

67. Calculate the ion product to determine if a precipitate will form when 125 mL 0.00500M sodium chloride is mixed with 125 mL 0.00100M silver nitrate solution.



Because $K_{\text{sp}} = 1.8 \times 10^{-10}$, a precipitate of AgCl may form.



Ions are diluted when one volume of NaCl solution is mixed with one volume of AgNO_3 solution resulting in the following concentrations.

$$[\text{Na}^+] \text{ in solution} = \frac{1}{2} (0.00500\text{M NaCl})$$

$$= 0.00250\text{M NaCl}$$

$$[\text{Cl}^-] \text{ in solution} = \frac{1}{2} (0.00500\text{M}) = 0.00250\text{M}$$

$$[\text{Ag}^+] \text{ in solution} = \frac{1}{2} (0.00100M \text{ Ag}^+ = 0.00050M$$

$$[\text{NO}_3^-] \text{ in solution} = \frac{1}{2} (0.00100M \text{ NO}_3^-]$$

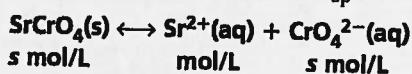
$$= 0.00050M$$

$$Q_{\text{sp}} = [\text{Ag}^+][\text{Cl}^-] = (0.00050)(0.00250) = 1.25 \times 10^{-6}$$

$$K_{\text{sp}} = 1.8 \times 10^{-10}$$

Because $Q_{\text{sp}} > K_{\text{sp}}$, a precipitate forms.

68. Calculate the molar solubility of strontium chromate in water at 298 K if $K_{\text{sp}} = 3.5 \times 10^{-5}$.



$$s \text{ mol/L} \quad \text{mol/L} \quad s \text{ mol/L}$$

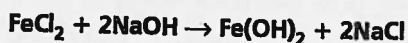
$$K_{\text{sp}} = [\text{Sr}^{2+}][\text{CrO}_4^{2-}]$$

$$3.5 \times 10^{-5} = (s)(s) = s^2$$

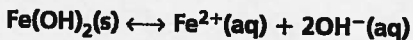
$$s = \sqrt{3.5 \times 10^{-5}} = 5.9 \times 10^{-3} \text{ mol/L}$$

Level 2

69. Will a precipitate form when 1.00 L of 0.150M iron(II) chloride solution is mixed with 2.00 L of 0.0333M sodium hydroxide solution? Explain your reasoning and show your calculations.



Because $K_{\text{sp}} = 4.9 \times 10^{-17}$, a precipitate of $\text{Fe}(\text{OH})_2$ may form.



Ions are diluted when one volume of FeCl_2 solution is mixed with two volumes of NaOH solution resulting in the following ion dilutions.

$$[\text{Fe}^{2+}] \text{ in solution} = \frac{1}{3} (0.150M \text{ Fe}^{2+})$$

$$= 0.0500M \text{ Fe}^{2+}$$

$$[\text{Cl}^-] \text{ in solution} = \frac{1}{3} (0.300M \text{ Cl}^-) = 0.100M \text{ Cl}^-]$$

$$[\text{Na}^+] \text{ in solution} = \frac{2}{3} (0.0333M \text{ Na}^+)$$

$$= 0.0222M \text{ Na}^+$$

$$[\text{OH}^-] \text{ in solution} = \frac{2}{3} (0.0333M \text{ OH}^-)$$

$$= 0.0222M \text{ OH}^-$$

$$Q_{\text{sp}} = [\text{Fe}^{2+}][\text{OH}^-]^2 = (0.0500)(0.0222)^2$$

$$= 2.46 \times 10^{-5}$$

$$K_{\text{sp}} = 4.9 \times 10^{-17}$$

Because $Q_{\text{sp}} \gg K_{\text{sp}}$, a precipitate forms.

Mixed Review

Sharpen your problem-solving skills by answering the following.

70. How many moles per liter of silver chloride will be in a saturated solution of AgCl ?

$$K_{\text{sp}} = 1.8 \times 10^{-10}$$



$$[\text{AgCl}] = [\text{Ag}^+] = [\text{Cl}^-] = s$$

$$[\text{Ag}^+][\text{Cl}^-] = 1.8 \times 10^{-10}$$

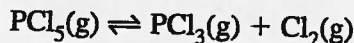
$$(s)(s) = 1.8 \times 10^{-10}$$

$$s = \sqrt{1.8 \times 10^{-10}}$$

$$= 1.3 \times 10^{-5} \text{ mol/L}$$

A saturated solution will contain 1.3×10^{-5} mol/L AgCl .

71. A 6.00-L vessel contains an equilibrium mixture of 0.0222 mol PCl_3 , 0.0189 mol PCl_5 , and 0.1044 mol Cl_2 . Calculate K_{eq} for the following reaction.



$$[\text{PCl}_3] = \frac{0.0222 \text{ mol}}{6.00 \text{ L}} = 0.00370 \text{ mol/L}$$

$$[\text{Cl}_2] = \frac{0.1044 \text{ mol}}{6.00 \text{ L}} = 0.0174 \text{ mol/L}$$

$$[\text{PCl}_5] = \frac{0.0189 \text{ mol}}{6.00 \text{ L}} = 0.00315 \text{ mol/L}$$

$$K_{\text{eq}} = \frac{[\text{PCl}_3][\text{Cl}_2]}{[\text{PCl}_5]} = \frac{(0.00370)(0.0174)}{(0.00315)}$$

$$= 2.04 \times 10^{-2}$$

Standardized Test Practice Chapter 18

page 593

Use these questions and the test-taking tip to prepare for your standardized test.

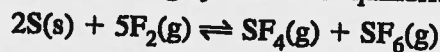
- A system reaches chemical equilibrium when _____.
 - no new product is formed by the forward reaction
 - the reverse reaction no longer occurs in the system
 - the concentration of reactants in the system is equal to the concentration of products
 - the rate at which the forward reaction occurs equals the rate of the reverse reaction

d
- A value of K_{eq} greater than 1 means that _____.
 - more reactants than products exist at equilibrium
 - more products than reactants exist at equilibrium
 - the rate of the forward reaction is high at equilibrium
 - the rate of the reverse reaction is high at equilibrium

b
- The hydrogen sulfide produced as a byproduct of petroleum refinement can be used to produce elemental sulfur: $2\text{H}_2\text{S}(\text{g}) + \text{SO}_2(\text{g}) \rightarrow 3\text{S}(\text{l}) + 2\text{H}_2\text{O}(\text{g})$
The equilibrium constant expression for this reaction is _____.
 - $K_{eq} = \frac{[\text{H}_2\text{O}]}{[\text{H}_2\text{S}][\text{SO}_2]}$
 - $K_{eq} = \frac{[\text{H}_2\text{S}]^2[\text{SO}_2]}{[\text{H}_2\text{S}]^2}$
 - $K_{eq} = \frac{[\text{H}_2\text{O}]^2}{[\text{H}_2\text{S}]^2[\text{SO}_2]}$
 - $K_{eq} = \frac{[\text{S}]^3[\text{H}_2\text{O}]^2}{[\text{H}_2\text{S}]^2[\text{SO}_2]}$

c

- The following system is in equilibrium:



The equilibrium will shift to the right if _____.

- the concentration of SF_4 is increased
- the concentration of SF_6 is increased
- the pressure on the system is increased
- the pressure on the system is decreased

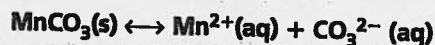
c

Concentration Data for the Equilibrium System
 $\text{MnCO}_3(\text{s}) \rightleftharpoons \text{Mn}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq})$ at 298 K

Trial	$[\text{Mn}^{2+}]_0$ (M)	$[\text{CO}_3^{2-}]_0$ (M)	$[\text{Mn}^{2+}]_{eq}$ (M)	$[\text{CO}_3^{2-}]_{eq}$ (M)
1	0.0000	0.00400	5.60×10^{-9}	4.00×10^{-3}
2	0.0100	0.0000	1.00×10^{-2}	2.24×10^{-9}
3	0.0000	0.0200	1.12×10^{-9}	2.00×10^{-2}

Interpreting Tables Use the table to answer questions 5–7.

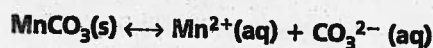
- The K_{sp} for MnCO_3 is _____.
 - 2.24×10^{-11}
 - 4.00×10^{-11}
 - 1.12×10^{-9}
 - 5.60×10^{-9}



$$K_{sp} = [\text{Mn}^{2+}][\text{CO}_3^{2-}] = (5.60 \times 10^{-9})(4.00 \times 10^{-3}) = 2.24 \times 10^{-11}$$

a

- What is the molar solubility of MnCO_3 at 298 K?
 - $4.73 \times 10^{-6}\text{M}$
 - $6.32 \times 10^{-2}\text{M}$
 - $7.48 \times 10^{-5}\text{M}$
 - $3.35 \times 10^{-5}\text{M}$



$$\text{solubility} = s = [\text{Mn}^{2+}] = [\text{CO}_3^{2-}]$$

$$(s)(s) = s^2 = 2.24 \times 10^{-11}$$

$$s = \sqrt{2.24 \times 10^{-11}} = 4.73 \times 10^{-6}\text{M}$$

a

7. A 50.0-mL volume of $3.00 \times 10^{-6} M$ K_2CO_3 is mixed with 50.0 mL of $MnCl_2$. A precipitate of $MnCO_3$ will form only when the concentration of the $MnCl_2$ solution is greater than _____.
- $7.47 \times 10^{-6} M$
 - $1.49 \times 10^{-5} M$
 - $2.99 \times 10^{-5} M$
 - $1.02 \times 10^{-5} M$

Because the volume of solution is doubled,

$$[CO_3^{2-}] \text{ in the mixed solution} = \frac{3.00 \times 10^{-6} M}{2}$$

$$= 1.50 \times 10^{-6} M$$

$$[Mn^{2+}][CO_3^{2-}] = [Mn^{2+}](1.50 \times 10^{-6})$$

$$= 2.24 \times 10^{-11}$$

$[Mn^{2+}] = 1.49 \times 10^{-5} M = [Mn^{2+}]$ in the mixed solution. To form a precipitate, the concentration of the original solution of $MnCl_2 = 2(x)$
 $= 2(1.49 \times 10^{-5} M) = 2.99 \times 10^{-5} M$

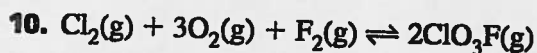
c

8. Which of the following statements about the common ion effect is NOT true?
- The effects of common ions on an equilibrium system can be explained by Le Châtelier's principle.
 - The decreased solubility of an ionic compound due to the presence of a common ion is called the common ion effect.
 - The addition of NaCl to a saturated solution of AgCl will produce the common ion effect.
 - The common ion effect is due to a shift in equilibrium towards the aqueous products of a system.

d

9. If the forward reaction of a system in equilibrium is endothermic, increasing the temperature of the system will _____.
- shift the equilibrium to the left
 - shift the equilibrium to the right
 - decrease the rate of the forward reaction
 - decrease the rate of the reverse reaction

b



The formation of perchloryl fluoride (ClO_3F) from its elements has an equilibrium constant of 3.42×10^{-9} at 298 K. If $[Cl_2] = 0.563 M$, $[O_2] = 1.01 M$, and $[ClO_3F] = 1.47 \times 10^{-5} M$ at equilibrium, what is the concentration of F_2 ?

- $9.18 \times 10^0 M$
- $3.73 \times 10^{-10} M$
- $1.09 \times 10^{-1} M$
- $6.32 \times 10^{-2} M$

$$K_{eq} = \frac{[ClO_3F]^2}{[F_2][O_2]^3[Cl_2]}$$

Solve for $[F_2]$.

$$[F_2] = \frac{[ClO_3F]^2}{K_{eq} \times [O_2]^3[Cl_2]}$$

Substitute known values.

$$[F_2] = \frac{(1.47 \times 10^{-5})^2}{(3.42 \times 10^{-9})(1.01)^3(0.563)} = 1.09 \times 10^{-1} M$$

c

35. Explain the difference between the equivalence point and the end point of a titration.

Equivalence point is the pH at which the moles of H^+ ions from the acid equal the moles of OH^- ions from the base. The end point is the point at which the indicator used in a titration changes color.

36. Predict the results of two experiments: A small amount of base is added to an unbuffered solution with a pH of 7 and the same amount of base is added to a buffered solution with a pH of 7.

The pH of the unbuffered solution increases more than the pH of the buffered solution.

37. **Thinking Critically** When a salt is dissolved in water, how can you predict whether or not a salt hydrolysis reaction occurs?

If the salt has a weak acid parent and/or a weak base parent, hydrolysis occurs.

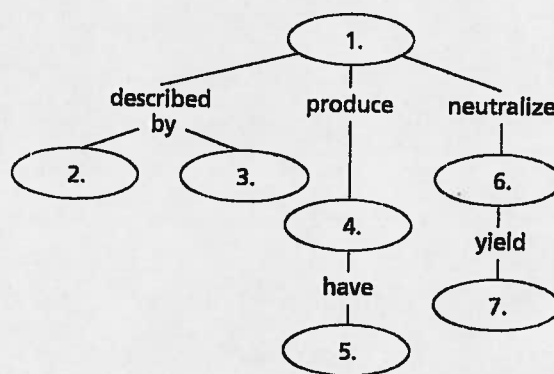
38. **Designing an Experiment** Describe how you would design and carry out a titration in which you use $0.250M HNO_3$ to determine the molarity of a cesium hydroxide solution.

Place a measured volume of $CsOH$ solution into a flask. Add an indicator. Fill a buret with the $0.250M HNO_3$ solution. Record the initial buret reading. Add HNO_3 solution slowly to the $CsOH$ solution until the end point. Record the final buret reading. Calculate the volume of HNO_3 added. Use the volume and molarity of HNO_3 and the volume of $CsOH$ to calculate the molarity of the $CsOH$ solution.

Chapter 19 Assessment pages 630–632

Concept Mapping

39. Use the following words and phrases to complete the concept map: acidic solutions, acids, bases, Arrhenius model, $pH < 7$, a salt plus water, Brønsted-Lowry model.



1. acids; 2. Arrhenius model; 3. Brønsted-Lowry model; 4. acidic solutions; 5. $pH < 7$; 6. bases; 7. a salt plus water

Mastering Concepts

40. An aqueous solution tastes bitter and turns litmus blue. Is the solution acidic or basic? (19.1)
The solution is basic.
41. An acidic solution reacts with magnesium carbonate to produce a gas. What is the formula for the gas? (19.1)
The formula is CO_2 .
42. In terms of ion concentrations, distinguish between acidic, neutral, and basic solutions. (19.1)
acidic, $[H^+] > [OH^-]$; neutral, $[H^+] = [OH^-]$;
basic, $[H^+] < [OH^-]$
43. Write a balanced chemical equation that represents the self-ionization of water. (19.1)
 $H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$

44. How did Arrhenius describe acids and bases? (19.1)

An acid contains hydrogen and ionizes to produce H^+ in aqueous solution. A base contains an OH group and dissociates to produce OH^- in aqueous solution.

45. Table sugar ($C_{12}H_{22}O_{11}$) contains 22 hydrogen atoms per molecule. Does this make table sugar an Arrhenius acid? Explain your answer. (19.1)

No, none of the hydrogens ionize to produce H^+ .

46. Classify each of the following compounds as an Arrhenius acid or an Arrhenius base. (19.1)

a. H_2S

acid

b. $RbOH$

base

c. $Mg(OH)_2$

base

d. H_3PO_4

acid

47. Explain the difference between a monoprotic acid, a diprotic acid, and a triprotic acid. Give an example of each. (19.1)

a monoprotic acid: can donate one H^+ (HCl); a

diprotic acid: can donate two H^+ (H_2SO_4); a

triprotic acid: can donate three H^+ (H_3PO_4)

48. Why does acid rain dissolve statues made of marble ($CaCO_3$)? Write the formula equation for the reaction between sulfuric acid and calcium carbonate. (19.1)

Acidic solutions react with and dissolve carbonates. $H_2SO_4(aq) + CaCO_3(s) \rightarrow CaSO_4(aq) + H_2O(l) + CO_2(g)$

49. Ammonia contains three hydrogen atoms per molecule. However, an aqueous ammonia solution is basic. Explain using the Brønsted-Lowry model of acids and bases. (19.1)

NH_3 accepts a hydrogen ion to form the ammonium ion.

50. Identify the conjugate acid-base pairs in the equilibrium equation $HC_2H_3O_2 + H_2O \rightleftharpoons H_3O^+ + C_2H_3O_2^-$. (19.1)

acid: $HC_2H_3O_2$; conjugate base: $C_2H_3O_2^-$; base: H_2O ; conjugate acid: H_3O^+

51. Gaseous HCl molecules interact with gaseous NH_3 molecules to form a white smoke made up of solid NH_4Cl particles. Explain whether or not this is an acid-base reaction according to both the Arrhenius model and the Brønsted-Lowry model. (19.1)

It is a Brønsted-Lowry acid-base reaction because HCl donates a hydrogen ion to NH_3 . This reaction does not occur in water. Thus, it does not follow the Arrhenius model.

52. Explain the difference between a strong acid and a weak acid. (19.2)

In dilute aqueous solution, a strong acid ionizes completely; a weak acid ionizes slightly.

53. Why are strong acids and bases also strong electrolytes? (19.2)

Because strong acids and bases ionize completely, the large number of ions produced causes the solutions to have high electrical conductivity.

54. State whether each of the following acids is strong or weak. (19.2)

a. acetic acid

weak

b. hydroiodic acid

strong

c. hydrofluoric acid

weak

d. phosphoric acid

weak

55. State whether each of the following bases is strong or weak. (19.2)

a. rubidium hydroxide

strong

b. methylamine

weak

c. ammonia

weak

d. sodium hydroxide

strong

56. How would you compare the strengths of two weak acids (19.2)

a. experimentally?

Compare the conductivities of equimolar solutions of the acids.

b. by looking up information in a table or a handbook?

Compare acid ionization constants of the two acids.

57. Figure 19-7 shows the conductivity of two acids. Explain how you could distinguish between solutions of two bases, one containing NaOH and the other NH_3 . (19.2)

The conductivity of the strong base NaOH is greater than the conductivity of the weak base NH_3 .

58. Explain why the base ionization constant (K_b) is a measure of the strength of a base. (19.2)

The greater the value of K_b , the greater the $[\text{OH}^-]$.

59. Explain why a weak acid has a strong conjugate base. Give an equation that illustrates your answer. (19.2)

$\text{HF}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{F}^-(\text{aq})$ HF has little tendency to donate a H^+ to H_2O (weak acid). Therefore, F^- has a strong tendency to accept a H^+ from H_3O^+ (strong base).

60. Explain why a weak base has a strong conjugate acid. Give an equation that illustrates your answer. (19.2)

$\text{CH}_3\text{NH}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{CH}_3\text{NH}_3^+(\text{aq}) + \text{OH}^-(\text{aq})$
 CH_3NH_2 has little tendency to accept a H^+ from H_2O . Therefore, CH_3NH_3^+ has a strong tendency to donate a H^+ to OH^- .

61. Explain how you can calculate K_a for a weak acid if you know the concentration of the acid and its pH. (19.2)

Obtain $[\text{H}^+]$ from the pH. $[\text{X}^-] = [\text{H}^+]$. $[\text{HX}] =$ the initial concentration minus $[\text{H}^+]$. Use the acid's ionization constant expression to calculate K_a .

62. What is the relationship between the pOH and the hydroxide-ion concentration of a solution? (19.3)

$\text{pOH} = -\log [\text{OH}^-]$

63. Solution A has a pH of 2.0. Solution B has a pH of 5.0. Which solution is more acidic? Based on the hydrogen-ion concentrations in the two solutions, how many times more acidic? (19.3)

Solution A is 10^3 , or 1000 times more acidic than B.

64. If the concentration of hydrogen ions in an aqueous solution decreases, what must happen to the concentration of hydroxide ions? Why? (19.3)

$[\text{OH}^-]$ increases. $[\text{H}^+][\text{OH}^-] = K_w$

65. Explain why pure water has a very slight electrical conductivity. (19.3)

Pure water ionizes to a slight degree producing H_3O^+ and OH^- ions.

66. Why is the pH of pure water 7.0 at 298 K? (19.3)

$\text{pH} = -\log [\text{H}^+]$, and $[\text{H}^+]$ in pure water at 298 K is $1.0 \times 10^{-7}\text{M}$.

67. What is a standard solution in an acid-base titration? (19.4)

the solution of known concentration

68. How do you recognize the end point in an acid-base titration? (19.4)

the color of the acid-base indicator changes

69. Give the name and formula of the acid and the base from which each salt was formed. (19.4)

a. NaCl

hydrochloric acid, HCl; sodium hydroxide, NaOH

b. KClO₃

chloric acid, HClO₃; potassium hydroxide, KOH

c. NH₄NO₂

nitrous acid, HNO₂; ammonium hydroxide, NH₄OH

d. CaS

hydrosulfuric acid, H₂S; calcium hydroxide, Ca(OH)₂

70. How does buffering a solution change the solution's behavior when a base is added? When an acid is added? (19.4)

Buffering minimizes the pH change in both instances.

71. Why are some aqueous salt solutions acidic or basic? (19.4)

Cations of some salts donate H⁺ to water, forming acidic solutions. Anions of other salts accept H⁺ from water, producing basic solutions.

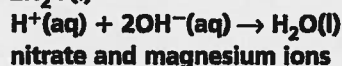
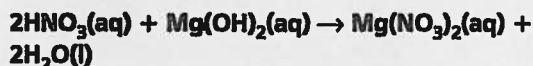
72. How does knowing the equivalence point in an acid-base titration help you choose an indicator for the titration? (19.4)

The color change should occur at a pH near the equivalence point.

73. An aqueous solution causes bromthymol blue to turn blue and phenolphthalein to turn colorless. What is the approximate pH of the solution? (19.4)

between 7.6 and 8.5

74. In the net ionic equation for the neutralization reaction between nitric acid and magnesium hydroxide, what ions are left out of the equation? (19.4)

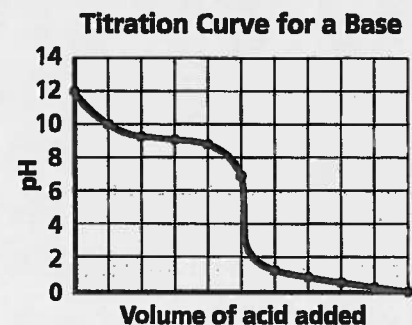


nitrate and magnesium ions

75. Describe two ways you might detect the end point of an acid-base titration experimentally. (19.4)

Use a pH meter or an acid-base indicator.

76. What is the approximate pH of the equivalence point in the titration pH curve? (19.4)



approximately pH 5

77. Define both the equivalence point and the end point of an acid-base titration. Why should you choose an indicator so that the two points are nearly the same pH? (19.4)

The equivalence point is the point at which mol H⁺ ions equals mol OH⁻ ions. The end point is the point at which the indicator changes color. An indicator detects the equivalence point, so the two should occur at the same pH.

78. Explain how you can predict whether an aqueous salt solution is acidic, basic, or neutral by evaluating the strengths of the salt's acid and base parents. (19.4)

An aqueous salt solution is acidic if the salt's acid parent is strong and its base parent is weak, basic if the salt's acid parent is weak and its base parent is strong, neutral if both the acid parent and base parent are strong.

79. In a hypochlorous acid/hypochlorite-ion buffer, what chemical species reacts when an acid is added to the solution? (19.4)

hypochlorite ion

80. Arrange the three buffers in order of increasing pH values. In order of increasing buffer capacity.
- 1.0M HClO/1.0M NaClO
 - 0.10M HClO/0.10M NaClO
 - 0.010M HClO/0.010M NaClO

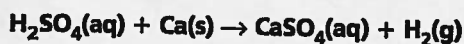
All three solutions have the same pH. the buffer capacity: c, b, a

Mastering Problems

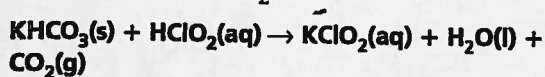
Equations for Acid and Base Reactions (19.1)

Level 1

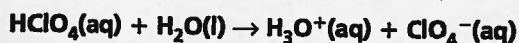
81. Write a balanced formula equation for the reaction between sulfuric acid and calcium metal.



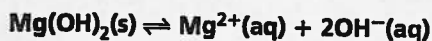
82. Write a balanced formula equation for the reaction between potassium hydrogen carbonate and chlorous acid (HClO₂).



83. Write the balanced chemical equation for the ionization of perchloric acid (HClO₄) in water.



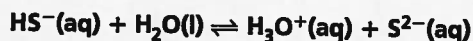
84. Write the balanced chemical equation for the dissociation of solid magnesium hydroxide in water.



Weak Acids and Bases (19.2)

Level 1

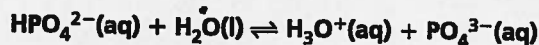
85. Write the equation for the ionization reaction and the acid ionization constant expression for the HS⁻ ion in water.



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{S}^{2-}]}{[\text{HS}^-]}$$

Level 2

86. Write the equation for the ionization reaction and the acid ionization constant expression for the third ionization of phosphoric acid in water.



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{PO}_4^{3-}]}{[\text{HPO}_4^{2-}]}$$

K_w, pH, and pOH (19.3)

Level 1

87. Given the concentration of either hydrogen ion or hydroxide ion, use the ion product constant of water to calculate the concentration of the other ion at 298 K.

a. $[\text{H}^+] = 1.0 \times 10^{-4} \text{M}$

$$K_w = [\text{H}^+][\text{OH}^-]$$

$$1.0 \times 10^{-14} = (1.0 \times 10^{-4})[\text{OH}^-]$$

$$\frac{1.0 \times 10^{-14}}{1.0 \times 10^{-4}} = \frac{(1.0 \times 10^{-4})[\text{OH}^-]}{1.0 \times 10^{-4}}$$

$$[\text{OH}^-] = 1.0 \times 10^{-10} \text{M}$$

b. $[\text{OH}^-] = 1.3 \times 10^{-2} \text{M}$

$$K_w = [\text{H}^+][\text{OH}^-]$$

$$1.0 \times 10^{-14} = [\text{H}^+](1.3 \times 10^{-2})$$

$$\frac{1.0 \times 10^{-14}}{1.3 \times 10^{-2}} = \frac{[\text{H}^+](1.3 \times 10^{-2})}{1.3 \times 10^{-2}}$$

$$[\text{H}^+] = 7.7 \times 10^{-13} \text{M}$$

88. Calculate the pH at 298 K of solutions having the following ion concentrations.

a. $[\text{H}^+] = 1.0 \times 10^{-4} \text{M}$

$$\text{pH} = -\log[\text{H}^+]$$

$$\text{pH} = -\log(1.0 \times 10^{-4}) = -(-4.00) = 4.00$$

b. $[\text{H}^+] = 5.8 \times 10^{-11} \text{M}$

$$\text{pH} = -\log[\text{H}^+]$$

$$\text{pH} = -\log(5.8 \times 10^{-11}) = -(-10.24) = 10.24$$

Level 2

89. Calculate the pOH and pH at 298 K of solutions having the following ion concentrations.

a. $[\text{OH}^-] = 1.0 \times 10^{-12} M$

$$\text{pOH} = -\log[\text{OH}^-]$$

$$\text{pOH} = -\log(1.0 \times 10^{-12}) = -(-12.00) = 12.00$$

$$\text{pH} + \text{pOH} = 14.00$$

$$\text{pH} + 12.00 = 14.00$$

$$\text{pH} = 14.00 - 12.00 = 2.00$$

b. $[\text{OH}^-] = 1.3 \times 10^{-2} M$

$$\text{pOH} = -\log[\text{OH}^-]$$

$$\text{pOH} = -\log(1.3 \times 10^{-2}) = -(-1.89) = 1.89$$

$$\text{pH} + \text{pOH} = 14.00$$

$$\text{pH} + 1.89 = 14.00$$

$$\text{pH} = 14.00 - 1.89 = 12.11$$

90. Calculate the pH of each of the following strong acid or strong base solutions at 298 K.

a. $2.6 \times 10^{-2} M \text{ HCl}$

$$\text{pH} = -\log[\text{H}^+] = -\log(2.6 \times 10^{-2}) = 1.59$$

b. $0.28 M \text{ HNO}_3$

$$\text{pH} = -\log[\text{H}^+] = -\log(0.28) = 0.55$$

c. $7.5 \times 10^{-3} M \text{ NaOH}$

$$\text{pOH} = -\log[\text{OH}^-] = -\log(7.5 \times 10^{-3}) = 2.13$$

$$\text{pH} + \text{pOH} = 14.00$$

$$\text{pH} + 2.13 = 14.00$$

$$\text{pH} = 14.00 - 2.13 = 11.87$$

d. $0.44 M \text{ KOH}$

$$\text{pOH} = -\log[\text{OH}^-] = -\log(0.44) = 0.36$$

$$\text{pH} + \text{pOH} = 14.00$$

$$\text{pH} + 0.36 = 14.00$$

$$\text{pH} = 14.00 - 0.36 = 13.64$$

Calculations Using K_a (19.3)

Level 1

91. A $8.6 \times 10^{-3} M$ solution of H_3PO_4 has a $\text{pH} = 2.30$. What is K_a for H_3PO_4 ?

$$K_a = \frac{[\text{H}^+][\text{H}_2\text{PO}_4^-]}{[\text{H}_3\text{PO}_4]}$$

$$[\text{H}^+] = [\text{H}_2\text{PO}_4^-] = \text{antilog}(-2.30) = 5.0 \times 10^{-3} M$$

$$[\text{H}_3\text{PO}_4] = 8.6 \times 10^{-3} M - 5.0 \times 10^{-3} M \\ = 3.6 \times 10^{-3} M$$

$$K_a = \frac{(5.0 \times 10^{-3})(5.0 \times 10^{-3})}{3.6 \times 10^{-3}} = 6.9 \times 10^{-3}$$

92. What is K_a for a solution of chloroacetic acid ($\text{C}_2\text{H}_3\text{ClO}_2$) which has a concentration of $0.112 M$ and a pH of 1.92 ?

$$K_a = \frac{[\text{H}^+][\text{C}_2\text{H}_2\text{ClO}_2^-]}{[\text{C}_2\text{H}_3\text{ClO}_2]}$$

$$[\text{H}^+] = [\text{C}_2\text{H}_2\text{ClO}_2^-] = \text{antilog}(-1.92) = 1.2 \times 10^{-2} M$$

$$[\text{C}_2\text{H}_3\text{ClO}_2] = 0.112 M - 1.2 \times 10^{-2} M = 0.100 M$$

$$K_a = \frac{(1.2 \times 10^{-2})(1.2 \times 10^{-2})}{0.100} = 1.4 \times 10^{-3}$$

Neutralization Reactions (19.4)

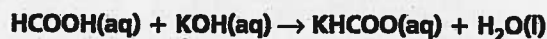
Level 1

93. Write formula equations for the following acid-base neutralization reactions.

a. sulfuric acid + sodium hydroxide

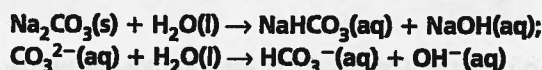


b. methanoic acid + potassium hydroxide

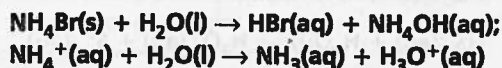


94. Write formula equations and net ionic equations for the hydrolysis of the following salts in water.

a. sodium carbonate



b. ammonium bromide



2. At the equivalence point of a strong acid-strong base titration, what is the approximate pH?

- a. 3
- b. 5
- c. 7
- d. 9

c

3. Hydrogen bromide (HBr) is a strong, highly corrosive acid. What is the pOH of a 0.0375M HBr solution?

- a. 12.57
- b. 12.27
- c. 1.73
- d. 1.43

$$\text{pH} = -\log(0.0375) = 1.426$$

$$\text{pOH} = 14.00 - 1.426 = 12.57$$

a

4. A compound that accepts H^+ ions is

- a. an Arrhenius acid
- b. an Arrhenius base
- c. a Brønsted-Lowry acid
- d. a Brønsted-Lowry base

d

Interpreting Tables Use the table to answer questions 5–7.

5. Which of the following acids is the strongest?

- a. formic acid
- b. cyanoacetic acid
- c. lutidinic acid
- d. barbituric acid

c

6. What is the acid dissociation constant of propanoic acid?

- a. 1.4×10^{-5}
- b. 2.43×10^0
- c. 3.72×10^{-3}
- d. 7.3×10^4

$$K_a = \frac{[\text{H}^+][\text{C}_3\text{H}_5\text{O}_2^-]}{[\text{HC}_3\text{H}_5\text{O}_2]}$$

$$[\text{H}^+] = [\text{C}_3\text{H}_5\text{O}_2^-] = \text{antilog}(-2.43) = 3.7 \times 10^{-3}M$$

$$[\text{HC}_3\text{H}_5\text{O}_2] = 1.00M - 3.7 \times 10^{-3}M = 0.10M$$

Substitute the values of the concentrations into the K_a expression.

$$K_a = \frac{(3.7 \times 10^{-3})^2}{(0.10)} = 1.4 \times 10^{-5}$$

a

7. What is the pH of a 0.400M solution of cyanoacetic acid?

- a. 2.059
- b. 1.22
- c. 2.45
- d. 1.42

$$K_a = \frac{[\text{H}^+][\text{C}_3\text{H}_2\text{NO}_2^-]}{[\text{HC}_3\text{H}_2\text{NO}_2]} = 3.55 \times 10^{-3}$$

Determine $[\text{H}^+]$.

$$[\text{H}^+] = [\text{C}_3\text{H}_2\text{NO}_2^-] = x$$

$$[\text{HC}_3\text{H}_2\text{NO}_2] = 0.400M - [\text{H}^+]$$

Assume $[\text{H}^+]$ is negligibly small compared to 0.400M because $\text{HC}_3\text{H}_2\text{NO}_2$ is a weak acid. Thus, $[\text{HC}_3\text{H}_2\text{NO}_2] = 0.400M$.

$$\frac{x^2}{0.400} = 3.55 \times 10^{-3}$$

$$x^2 = (0.400)(3.55 \times 10^{-3})$$

$$x = \sqrt{(0.400)(3.55 \times 10^{-3})} = 3.77 \times 10^{-2}M = [\text{H}^+]$$

$$\text{pH} = -\log 3.77 \times 10^{-2} = 1.423$$

d

Ionization Constants and pH Data for Several Weak Organic Acids

Acid	Ionization equation	pH of 1.000M solution	K_a
Formic	$\text{HCHO}_2 \rightleftharpoons \text{H}^+ + \text{CHO}_2^-$	1.87	1.78×10^{-4}
Cyanoacetic	$\text{HC}_3\text{H}_2\text{NO}_2 \rightleftharpoons \text{H}^+ + \text{C}_3\text{H}_2\text{NO}_2^-$?	3.55×10^{-3}
Propanoic	$\text{HC}_3\text{H}_5\text{O}_2 \rightleftharpoons \text{H}^+ + \text{C}_3\text{H}_5\text{O}_2^-$	2.43	?
Lutidinic	$\text{H}_2\text{C}_7\text{H}_3\text{NO}_4 \rightleftharpoons \text{H}^+ + \text{C}_7\text{H}_4\text{NO}_4^-$	1.09	7.08×10^{-3}
Barbituric	$\text{HC}_4\text{H}_3\text{N}_2\text{O}_3 \rightleftharpoons \text{H}^+ + \text{C}_4\text{H}_3\text{N}_2\text{O}_3^-$	2.01	9.77×10^{-5}

8. Which of the following is NOT a characteristic of a base?

- a. bitter taste
- b. ability to conduct electricity
- c. reactivity with some metals
- d. slippery feel

c

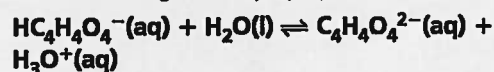
9. Diprotic succinic acid ($\text{H}_2\text{C}_4\text{H}_4\text{O}_4$) is an important part of the process that converts glucose to energy in the human body. What is the K_a expression for the second ionization of succinic acid?

a. $K_a = \frac{[\text{H}_3\text{O}^+][\text{HC}_4\text{H}_4\text{O}_4^-]}{[\text{H}_2\text{C}_4\text{H}_4\text{O}_4]}$

b. $K_a = \frac{[\text{H}_3\text{O}^+][\text{C}_4\text{H}_4\text{O}_4^{2-}]}{[\text{HC}_4\text{H}_4\text{O}_4^-]}$

c. $K_a = \frac{[\text{H}_2\text{C}_4\text{H}_4\text{O}_4]}{[\text{H}_3\text{O}^+][\text{HC}_4\text{H}_4\text{O}_4^-]}$

d. $K_a = \frac{[\text{H}_2\text{C}_4\text{H}_4\text{O}_4]}{[\text{H}_3\text{O}^+][\text{C}_4\text{H}_4\text{O}_4^{2-}]}$



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{C}_4\text{H}_4\text{O}_4^{2-}]}{[\text{HC}_4\text{H}_4\text{O}_4^-]}$$

b

10. A solution of 0.600M HCl is used to titrate 15.00 mL of KOH solution. The endpoint of the titration is reached after the addition of 27.13 mL of HCl. What is the concentration of the KOH solution?

- a. 9.00M
- b. 1.09M
- c. 0.332M
- d. 0.0163M

Calculate moles of HCl used.

$$27.13 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.02713 \text{ L HCl}$$

$$0.02713 \text{ L} \times \frac{0.600 \text{ mol}}{1 \text{ L}} = 0.0163 \text{ mol HCl}$$



The ratio of mol HCl : mol KOH is 1 : 1.

Thus, mol KOH = 0.0163 mol

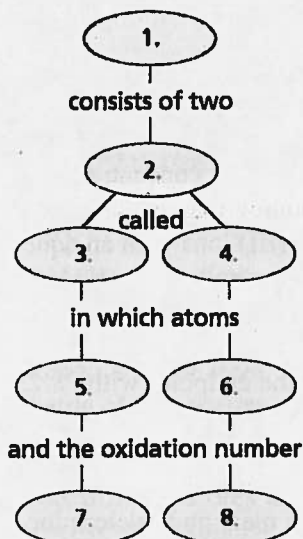
$$\frac{0.0163 \text{ mol KOH}}{0.0150 \text{ L KOH}} = 1.09\text{M}$$

b

Chapter 20 Assessment pages 658–660

Concept Mapping

32. Complete the concept map using the following terms: decreases, half-reactions, gain electrons, reduction, lose electrons, redox reaction, oxidation, increases.



1. redox reaction; 2. half-reactions; 3. oxidation;
4. reduction; 5. lose electrons; 6. gain electrons;
7. increases; 8. decreases

Mastering Concepts

33. What is the main characteristic of oxidation–reduction reactions? (20.1)
All oxidation–reduction reactions involve the transfer of electrons.
34. In terms of electrons, what happens when an atom is oxidized? When an atom is reduced? (20.1)
Electrons are lost; electrons are gained.
35. What is the oxidation number of alkaline earth metals in their compounds? Of alkali metals? (20.1)
alkaline earth = +2 alkali metals = +1

36. How does the oxidation number change when an element is oxidized? When it is reduced? (20.1)

Oxidation number increases; oxidation number decreases.

37. Explain why oxidation and reduction must occur simultaneously. (20.1)

If an atom loses an electron, some other species must gain the electron.

38. Identify the oxidizing agent and the reducing agent in the following equation. Explain your answer. (20.2)



oxidizing agent, Ag^+ ; reducing agent, Fe; Ag^+ is reduced; Fe is oxidized.

39. How can you tell that the redox equation in question 38 is not balanced? (20.2)

The total charge on the left-hand side does not equal the total charge on the right-hand side.

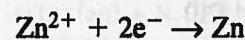
40. How does the change in oxidation number in an oxidation process relate to the number of electrons lost? How does the change in oxidation number in a reduction process relate to the number of electrons gained? (20.2)

The change in oxidation number equals the number of electrons lost in oxidation, gained in reduction.

41. Before you attempt to balance the equation for a redox reaction, why do you need to know whether the reaction takes place in acidic or basic solution? (20.3)

The type of solution provides the H^+ and OH^- ions needed to balance the redox equation because they are readily available.

42. Does the following equation represent a reduction or an oxidation process? Explain your answer. (20.3)



reduction; Electrons are gained and the oxidation number for Zn decreases.

43. What term is used for the type of reaction represented in question 42? (20.3)

half-reaction

Mastering Problems

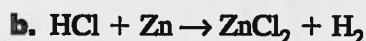
Oxidation and Reduction (20.1)

Level 1

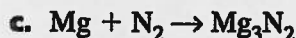
44. Identify the species oxidized and the species reduced in each of these redox equations.



Ga is oxidized, Br_2 is reduced.

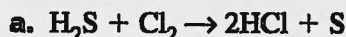


Zn is oxidized, HCl is reduced.

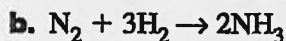


Mg is oxidized, N_2 is reduced.

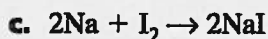
45. Identify the oxidizing agent and the reducing agent in each of these redox equations.



Cl_2 is the oxidizing agent, H_2S is the reducing agent.



N_2 is the oxidizing agent, H_2 is the reducing agent.

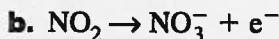


I_2 is the oxidizing agent, Na is the reducing agent.

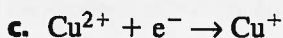
46. Identify each of these half-reactions as either oxidation or reduction.



oxidation



oxidation



reduction

47. Determine the oxidation number of the bold element in these substances and ions.



+6



+6

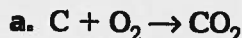


+3

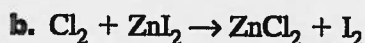


+5

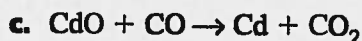
48. Determine the net change of oxidation number of each of the elements in these redox equations.



C, +4; O, -2

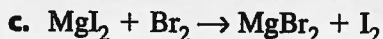
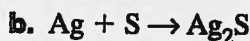
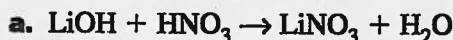


I, +1; Cl, -1; Zn, no change



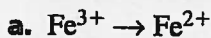
C, +2; Cd, -2; O, no change

49. Which of these equations does **not** represent a redox reaction? Explain your answer.

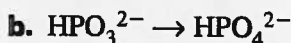


Choice a is not redox because none of the atoms in the reaction undergo a change in oxidation number.

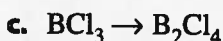
50. Identify each of these half-reactions as either oxidation or reduction.



reduction



oxidation



reduction

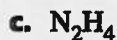
51. Determine the oxidation number of nitrogen in each of these molecules or ions.



-3



-3



-2



+5



+1



+3

52. Determine the oxidation number of each element in these compounds or ions.



Fe, +2; Cr, +3; O, -2



Au, +3; Se, +6; O, -2



Ni, +2; C, +2; N, -3

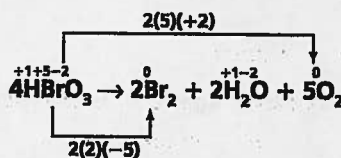
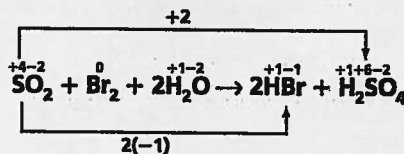
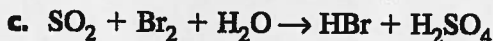
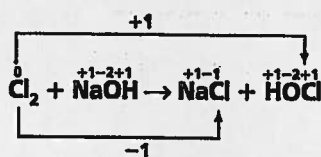
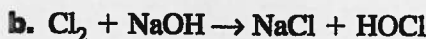
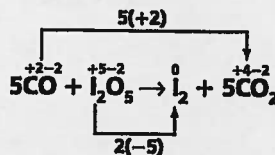
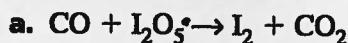
53. Explain how the sulfite ion (SO_3^{2-}) differs from sulfur trioxide (SO_3).

SO_3^{2-} is a polyatomic ion and the oxidation number of sulfur is +4. SO_3 is a compound and the oxidation number of S in this compound is +6.

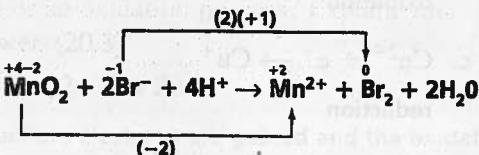
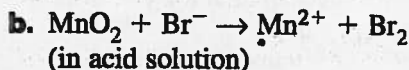
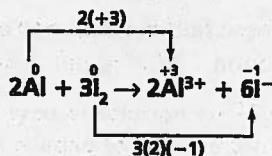
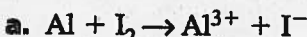
Balancing Redox Equations (20.2)

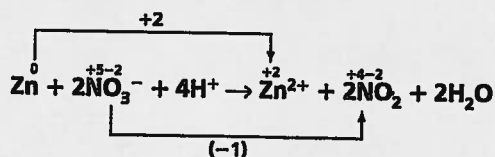
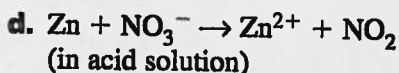
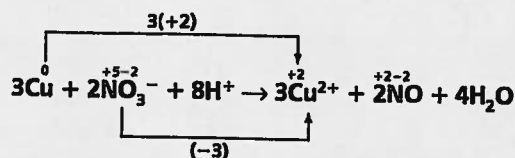
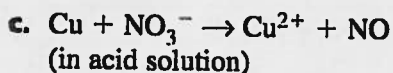
Level 1

54. Use the oxidation-number method to balance these redox equations.

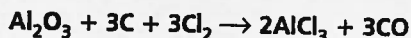
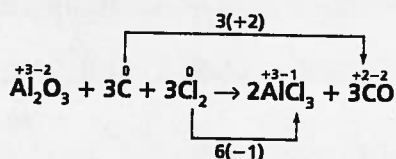
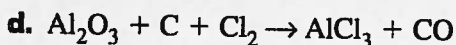
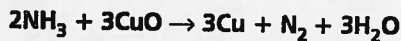
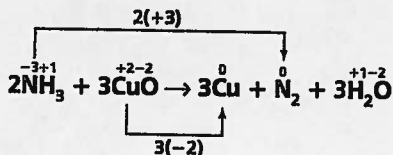
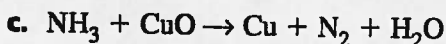
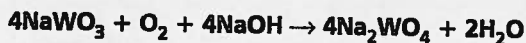
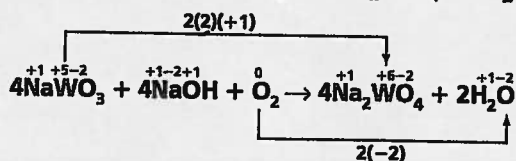
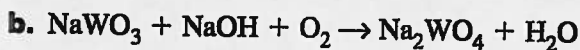
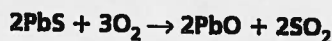
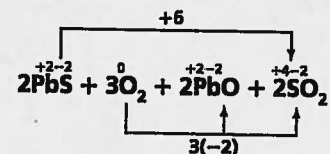
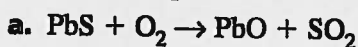


55. Use the oxidation-number method to balance the following ionic redox equations.

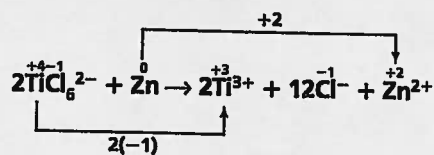
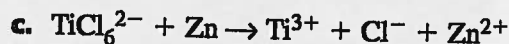
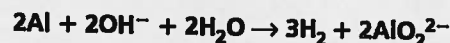
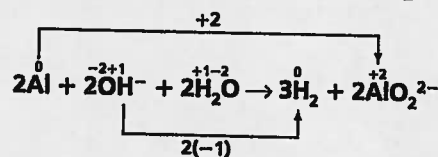
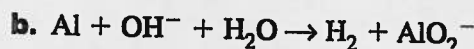
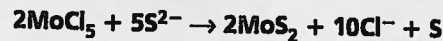
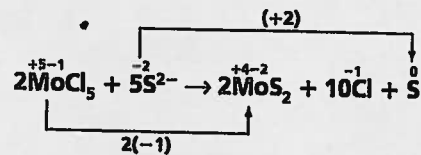
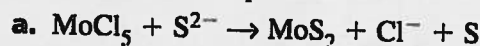




56. Use the oxidation-number method to balance these redox equations.

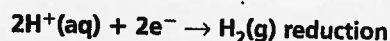
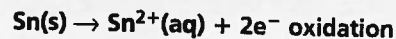
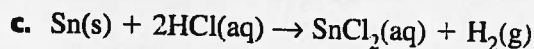
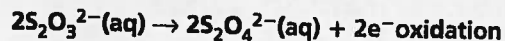
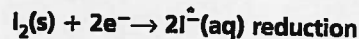
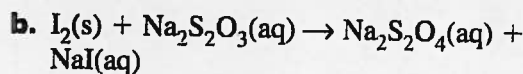
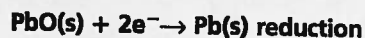
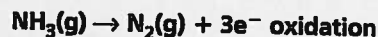
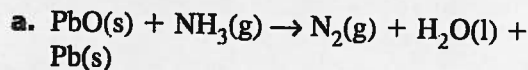


57. Use the oxidation-number method to balance these ionic redox equations.

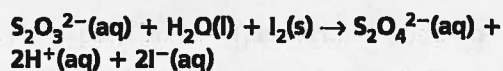
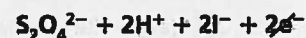
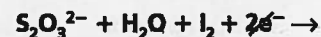
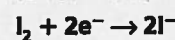
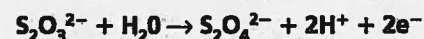
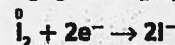
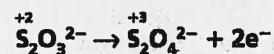
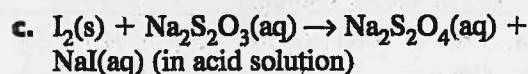
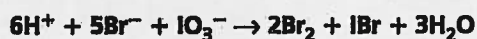
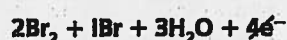
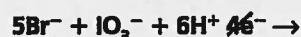
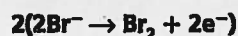
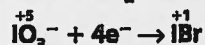
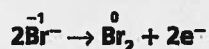
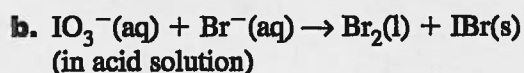
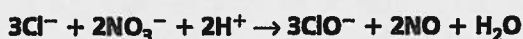
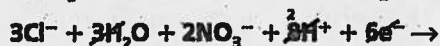
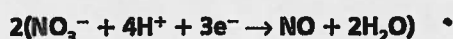
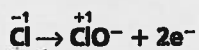
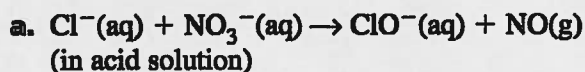


Half-Reactions (20.3)

58. Write the oxidation and reduction half-reactions represented in each of these redox equations. Write the half-reactions in net ionic form if they occur in aqueous solution.

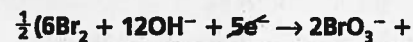
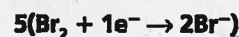
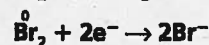
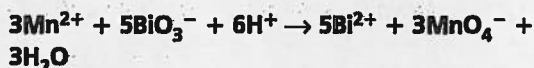
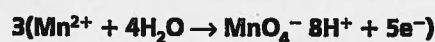
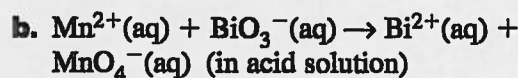
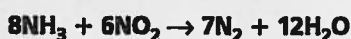
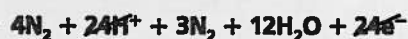
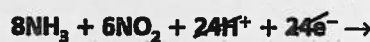
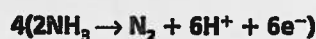
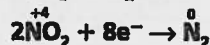
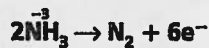
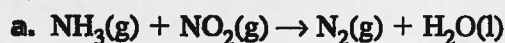


59. Use the half-reaction method to balance these equations. Add water molecules and hydrogen ions (in acid solutions) or hydroxide ions (in basic solutions) as needed. Keep balanced equations in net ionic form.

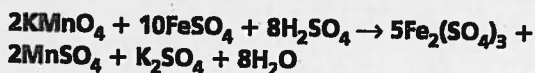
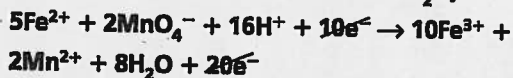
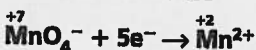
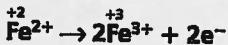
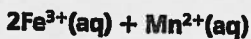
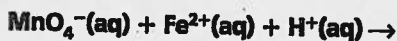
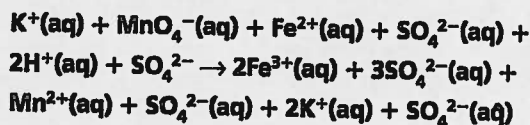
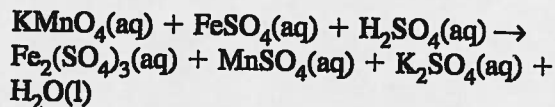


Level 2

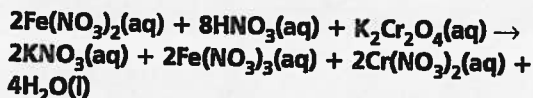
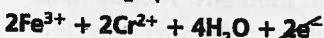
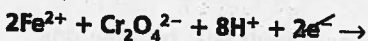
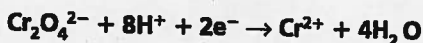
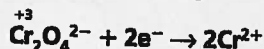
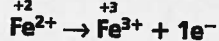
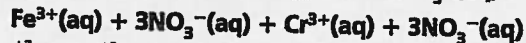
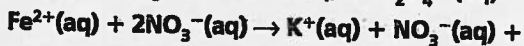
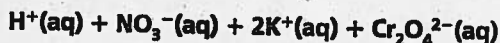
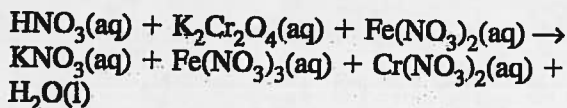
60. Use the half-reaction method to balance these equations for redox reactions. Add water molecules and hydrogen ions (in acid solutions) or hydroxide ions (in basic solutions) as needed.



61. Balance the following redox chemical equation. Rewrite the equation in full ionic form, then derive the net ionic equation and balance by the half-reaction method. Give the final answer as it is shown below but with the balancing coefficients.



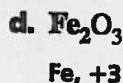
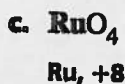
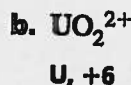
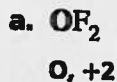
62. Balance this equation in the same manner as in question 61 above.



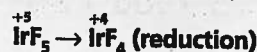
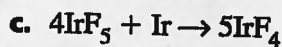
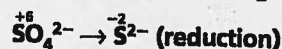
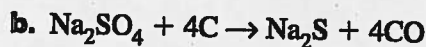
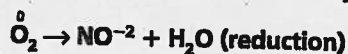
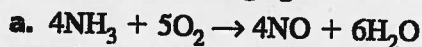
Mixed Review

Sharpen your problem-solving skills by answering the following.

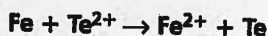
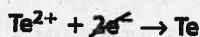
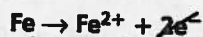
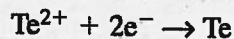
63. Determine the oxidation number of the bold element in each of the following examples.



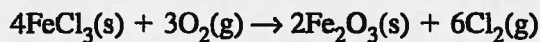
64. Identify the reducing agents in these equations.



65. Write a balanced ionic redox equation using the following pairs of redox half-reactions.



76. When iron(III) chloride reacts in an atmosphere of pure oxygen the following occurs:



If 45.0 g of iron(III) chloride reacts and 20.5 g of iron(III) oxide is recovered, determine the percent yield. (Chapter 12)

$$45.0 \text{ g FeCl}_3 \times \frac{1 \text{ mol FeCl}_3}{162.20 \text{ g FeCl}_3} \times \frac{2 \text{ mol Fe}_2\text{O}_3}{4 \text{ mol FeCl}_3} \times \frac{159.69 \text{ g Fe}_2\text{O}_3}{1 \text{ mol Fe}_2\text{O}_3} = 22.15 \text{ g Fe}_2\text{O}_3$$

$$\frac{20.5 \text{ g Fe}_2\text{O}_3}{22.15 \text{ g Fe}_2\text{O}_3} \times 100 = 92.5\% \text{ yield}$$

Standardized Test Practice Chapter 20

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Use these questions and the test-taking tip to prepare for your standardized test.

- The reducing agent in a redox reaction is all of the following EXCEPT _____.
 - the substance oxidized
 - the electron acceptor
 - the reducer of another substance
 - the electron donor

b
- The oxidation numbers of the elements in CuSO_4 are _____.
 - $\text{Cu} = +2, \text{S} = +6, \text{O} = -2$
 - $\text{Cu} = +3, \text{S} = +5, \text{O} = -2$
 - $\text{Cu} = +2, \text{S} = +2, \text{O} = -1$
 - $\text{Cu} = +2, \text{S} = 0, \text{O} = -2$

a
- For the reaction $\text{X} + \text{Y} \rightarrow \text{XY}$, the element that will be reduced is the one that is _____.
 - more reactive
 - more massive
 - more electronegative
 - more radioactive

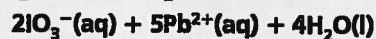
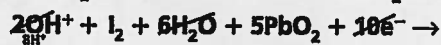
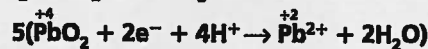
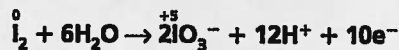
c

4. The net ionic reaction between iodine and lead(IV) oxide is shown below:



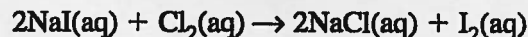
If the reaction takes place in acidic solution, the balanced equation is _____.

- $\text{I}_2(\text{s}) + 5\text{PbO}_2(\text{s}) + 4\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{IO}_3^-(\text{aq}) + 5\text{Pb}^{2+}(\text{aq}) + 8\text{OH}^-(\text{aq})$
- $\text{I}_2(\text{s}) + 5\text{PbO}_2(\text{s}) + \text{H}^+(\text{aq}) \rightarrow 2\text{IO}_3^-(\text{aq}) + 5\text{Pb}^{2+}(\text{aq}) + \text{H}_2\text{O}(\text{l})$
- $\text{I}_2(\text{s}) + 5\text{PbO}_2(\text{s}) + 4\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{IO}_3^-(\text{aq}) + 5\text{Pb}^{2+}(\text{aq}) + 8\text{H}^+(\text{aq})$
- $\text{I}_2(\text{s}) + 5\text{PbO}_2(\text{s}) + 8\text{H}^+(\text{aq}) \rightarrow 2\text{IO}_3^-(\text{aq}) + 5\text{Pb}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$



d

5. The reaction between sodium iodide and chlorine is shown below:

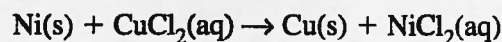


The oxidation state of Na remains unchanged because _____.

- Na^+ is a spectator ion
- Na^+ cannot be reduced
- Na is an uncombined element
- Na^+ is a monatomic ion

a

6. The reaction between nickel and copper(II) chloride is shown below:



The half reactions for this redox reaction are _____.

- $\text{Ni} \rightarrow \text{Ni}^{2+} + 2\text{e}^-$, $\text{Cl}^- \rightarrow 2\text{Cl}^- + 2\text{e}^-$
- $\text{Ni} \rightarrow \text{Ni}^+ + \text{e}^-$, $\text{Cu}^+ + \text{e}^- \rightarrow \text{Cu}$
- $\text{Ni} \rightarrow \text{Ni}^{2+} + 2\text{e}^-$, $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$
- $\text{Ni} \rightarrow \text{Ni}^{2+} + 2\text{e}^-$, $2\text{Cu}^+ + 2\text{e}^- \rightarrow \text{Cu}$

c

Standardized Test Practice Chapter 21

Page 695

Use these questions and the test-taking tip to prepare for your standardized test.

1. A salt bridge is essential to a voltaic cell for all of the following reasons EXCEPT _____.
- it allows ions to move from the solution of one cell to the other
 - it causes electric current to flow between the two electrodes of a cell
 - it relieves the buildup of positive charge on the anode side of the cell
 - it allows electrons to move from the solution of one cell to the other
- d

Using Tables Use the table to answer questions 2–5.

Selected Standard Reduction Potentials at 25°C, 1 atm, and 1M Ion Concentration	
Half-Reaction	E^0 (V)
$\text{Mg}^{2+} + 2e^- \rightarrow \text{Mg}$	-2.372
$\text{Al}^{3+} + 3e^- \rightarrow \text{Al}$	-1.662
$\text{Pb}^{2+} + 2e^- \rightarrow \text{Pb}$	-0.1262
$\text{Ag}^+ + e^- \rightarrow \text{Ag}$	0.7996
$\text{Hg}^{2+} + 2e^- \rightarrow \text{Hg}$	0.851

2. Which of the metal ions is most easily reduced?
- Mg^{2+}
 - Hg^{2+}
 - Ag^+
 - Al^{3+}
- b

3. On the basis of the standard reduction potentials shown above, which of the following standard cell notations below correctly represents its voltaic cell?

- $\text{Ag} | \text{Ag}^+ || \text{Al}^{3+} | \text{Al}$
- $\text{Mg} | \text{Mg}^{2+} || \text{H}^+ | \text{H}_2$
- $\text{H}_2 | \text{H}^+ || \text{Pb}^{2+} | \text{Pb}$
- $\text{Pb} | \text{Pb}^{2+} || \text{Al}^{3+} | \text{Al}$

b

4. A voltaic cell consists of a magnesium bar dipping into a 1M Mg^{2+} solution and a silver bar dipping into a 1M Ag^+ solution. What is the standard potential of this cell?

- 1.572 V
- 3.172 V
- 0.773 V
- 3.971 V

b

5. Assuming standard conditions, which of the following cells will produce a potential of 2.513 V?

- $\text{Al} | \text{Al}^{3+} || \text{Hg}^{2+} | \text{Hg}$
- $\text{H}_2 | \text{H}^+ || \text{Hg}^{2+} | \text{Hg}$
- $\text{Mg} | \text{Mg}^{2+} || \text{Al}^{3+} | \text{Al}$
- $\text{Pb} | \text{Pb}^{2+} || \text{Ag}^+ | \text{Ag}$

a

6. Which of the following statements is NOT true of batteries?

- Batteries are compact forms of voltaic cells.
- Secondary batteries also are known as storage batteries.
- Batteries can consist only of a single cell.
- The redox reaction in a rechargeable battery is reversible.

c

7. The corrosion, or rusting, of iron is an example of a naturally occurring voltaic cell. To prevent corrosion, sacrificial anodes are sometimes attached to rust-susceptible iron. Sacrificial anodes must _____ .
- be more likely to be reduced than iron
 - have a higher reduction potential than iron
 - be more porous and abraded than iron
 - lose electrons more easily than iron
- d
8. A strip of metal X is immersed in a 1M solution of X^+ ions. When this half-cell is connected to a standard hydrogen electrode, a voltmeter reads a positive reduction potential. Which of the following is true of the X electrode?
- It accepts electrons more readily than H^+ ions.
 - It is undergoing oxidation.
 - It is adding positive X^+ ions to its solution.
 - It acts as the anode in the cell.
- a
9. To electroplate an iron fork with silver, which of the following is true?
- The silver electrode must have more mass than the fork.
 - The iron fork must act as the anode in the cell.
 - Electrical current must be applied to the iron fork.
 - Iron ions must be present in the cell solution.
- c