Brief Contents

Contents

CHAPTER 1

What Is Chemistry?

1. SOME BASIC DEFINITIONS

Test Yourself

1.
  1. The moon takes up space, so it is matter.
  2. The invention itself may be matter, but the idea for it is not.

Test Yourself

1.
  1. Because of the fire, chemical changes are occurring in the fire place. This is a chemical change.
  2. Because water is just warmed in its liquid phase, the change is a physical change.

Test Yourself

1.
  1. The human body as a whole contains many parts that are visible to the naked eye. Hence, it is a heterogeneous mixture.
  2. Because amalgam contains metals dissolved in mercury, amalgam is a homogeneous mixture.

Exercises

1.
  a. A book has a mass and occupies space. Hence, it is matter.
  b. Hate is an emotion, which does not occupy space. Hence, it is not matter.
  c. Light does not have a mass nor occupies space. Hence, it is not matter.
  d. A car has a mass and occupies space. Hence, it is matter.
  e. A fried egg has a mass and occupies space. Hence, it is matter.

3.
  a. Physical characteristics of matter are shape, color, size, and temperature.
  b. Chemical properties are characteristics of matter that describe how matter changes form in the presence of other matter.
  c. Physical characteristics of matter are shape, color, size, and temperature.

5.
  a. Because water is going from a liquid phase to a gas phase, the change is a physical change.
  b. Because of the food, chemical changes are occurring in the digestive system. This is a chemical change.
  c. Because alcohol is going from a liquid phase to a solid phase, the change is a physical change.

7. An element is a fundamental chemical part of a substance; there are about 115 known elements. A compound is a combination of elements that acts as a different substance; there are over 50 million known substances.

9.
  a. When salt is mixed with pepper, both can be seen clearly. Hence, the mixture is heterogeneous.
  b. Sugar completely dissolves in water. Hence, the mixture is homogeneous.
  c. Pasta in water can be seen clearly. Hence, the mixture is heterogeneous.

11. b. Sugar dissolved in water is a solution; another name for a homogenous mixture is solution.

13. Iron is a metal because it is solid, is shiny, and conducts electricity and heat well.

15. Metals are typically shiny, conduct electricity and heat well, and are malleable and ductile; nonmetals are a variety of colors and phases, are brittle in the solid phase, and do not conduct heat or electricity well.

17. Carbon behaves as a metal because it conducts heat and electricity well. It is a nonmetal because it is black and brittle and cannot be made into sheets or wires.

2. CHEMISTRY AS A SCIENCE

Test Yourself

1.
   1. Dynamics is science.
   2. Aesthetics is not science.

Test Yourself

1.
   1. Because the statement describes a physical property of roses and violets, this statement is qualitative.
   2. This statement mentions a specific amount, so it is quantitative.

Exercises

1. Simply stated, the scientific method includes three steps: (1) stating a hypothesis, (2) testing the hypothesis, and (3) refining the hypothesis.

3. Scientists perform experiments to test their hypotheses because sometimes the nature of natural universe is not obvious.

5. A scientific law is a specific statement that is thought to be never violated by the entire natural universe. Everyday laws are arbitrary limits that society puts on its members.

7.
   a. yes
   b. no
   c. no

9.
   a. Because the statement describes a physical property of the *Titanic*, this statement is qualitative.
   b. This statement mentions a specific amount, so it is not qualitative.
   c. This statement mentions a specific amount, so it is not qualitative.

11. b. The statement describing the population of United States as a specific quantity is quantitative.

c. The statement describing the height of Mount Everest as a specific quantity is quantitative.

CHAPTER 2

Measurements

1. EXPRESSING NUMBERS

Test Yourself

1. 1. The number 23,070 is 2.307 times 10,000, or 2.307 times 10^{4}. In scientific notation, the number is 2.307×10^{4}.

2. The number 0.0009706 is 9.706 times 1/10,000, which is 9.706 times 10^{-4}. In scientific notation, the number is 9.706×10^{-4}.

Exercises

1. a. The number 56.9 is 5.69 times 10, or 5.69 times 10^{1}. In scientific notation, the number is 5.69×10^{1}.

b. The number 563,100 is 5.631 times 100,000 or 5.631 times 10^{5}. In scientific notation, the number is 5.631×10^{5}.

c. The number 0.0804 is 8.04 times 1/100 or 8.04 times 10^{-2}. In scientific notation, the number is 8.04×10^{-2}.

d. The number 0.00000667 is 6.67 times 1/1,000,000 or 6.67 times 10^{-6}. In scientific notation, the number is 6.67×10^{-6}.

3. a. The number 0.00656 is 6.56 times 1/1,000, or 6.56 times 10^{-3}. In scientific notation, the number is 6.56×10^{-3}.

b. The number 65,600 is 6.56 times 10,000, or 6.56 times 10^{4}. In scientific notation, the number is 6.56×10^{4}.

c. The number 4,567,000 is 4.567 times 1,000,000, or 4.567 times 10^{6}. In scientific notation, the number is 4.567×10^{6}.

d. The number 0.000005507 is 5.507 times 1/1,000,000, or 5.507 times 10^{-6}. In scientific notation, the number is 5.507×10^{-6}.

5. a. In scientific notation, the number is 1.381×10^{5}, or 1.381 times 10^{5}, or 1.381 times 100,000. The standard notation is 138,100.

b. In scientific notation, the number is 5.22×10^{-7}, or 5.22 times 10^{-7}, or 5.22 times 1/10,000,000. The standard notation is 0.000000522.

c. In scientific notation, the number is 9.998×10^{4}, or 9.998 times 10^{4}, or 9.998 times 10,000. The standard notation is 99,980.

7. a. In scientific notation, the number is 8.09×10^{0}, or 8.09 times 10^{0}, or 8.09 times 1. The standard notation is 8.09.

b. In scientific notation, the number is 3.088×10^{-5}, or 3.088 times 10^{-5}, or 3.088 times 1/100,000. The standard notation is 0.00003088.

c. In scientific notation, the number is -4.239×10^{2}, or -4.239 times 10^{2}, or -4.239 times 100. The standard notation is -423.9.

9. a. The number 72.44×10^{3} is 7.244 times 10,000, or 7.244 times 10^{4}. In scientific notation, the number is 7.244×10^{4}.

b. The number $9{,}943 \times 10^{-5}$ is 9.943 times 1/100, or 9.943 times 10^{-2}. In scientific notation, the number is 9.943×10^{-2}.

c. The number $588{,}399 \times 10^{2}$ is 5.88399 times 10,000,000, or 5.88399 times 10^{7}. In scientific notation, the number is 5.88399×10^{7}.

11. a. The number 345.1×10^{2} is 3.451 times 10,000, or 3.451 times 10^{4}. In scientific notation, the number is 3.451×10^{4}.

b. The number 0.234×10^{-3} is 2.34 times 1/10,000, or 2.34 times 10^{-4}. In scientific notation, the number is 2.34×10^{-4}.

c. The number $1{,}800 \times 10^{-2}$ is 1.8 times 10, or 1.8 times 10^{1}. In scientific notation, the number is 1.8×10^{1}.

13. a.
1 2 3 , 4 5 6 .78 = 1.2345678×10^{5}
↖ ↖ ↖ ↖ ↖
5 4 3 2 1

b.
9 8 , 4 9 0 = 9.849×10^{4}
↖ ↖ ↖ ↖
4 3 2 1

c.
0. 0 0 0 0 0 0 4 45 = 4.45×10^{-7}
↗ ↗ ↗ ↗ ↗ ↗ ↗
1 2 3 4 5 6 7

15. a. $456 \times (7.4 \times 10^{8}) = 3374.4 \times 10^{8} = 3.3744 \times 10^{11}$

b. $(3.02 \times 10^{5}) \div (9.04 \times 10^{15}) = 0.33407 \times 10^{-10} = 3.3407 \times 10^{-11}$

c. $0.0044 \times 0.000833 = 0.0000036652 \times 10^{-10} = 3.6652 \times 10^{-6}$

17. a. $45 \times 132 \div 882 = 6.7346 = 6.7346 \times 10^{0}$

b. $[(6.37 \times 10^{4}) \times (8.44 \times 10^{-4})] \div (3.2209 \times 10^{15}) = 16.691 \times 10^{-15} = 1.6691 \times 10^{-14}$

2. EXPRESSING UNITS

Test Yourself

1.
1. 2,500,000 L can be written as 2.5×10^6 L. Since 10^6 defines the mega-prefix, volume of the Olympic-sized swimming pool is about 2.5 ML.
2. If velocity is defined as a distance quantity divided by a time quantity, then velocity is 6.1 m/30 min. Dividing the numbers gives us 6.1/30 = 0.203, and dividing the units gives us meters/minute, or m/min. The velocity is 0.203 m/min.

Exercises

1. a. boxes of crayons
   b. grams of gold

3. a. seconds
   b. meters

5. a. 1,000 ×
   b. 1/1,000 ×
   c. 1,000,000 ×

7. a. milli-
   b. kilo-
   c. giga-

9.

Unit	Abbreviation
kilosecond	ks
milliliter	mL
megagram	Mg
centimeter	cm

11. a. Since 10^{-6} defines the micro-prefix, 3.44×10^{-6}can be written as 3.44 µs.
   b. 3,500 L can be written as 3.5×10^3 L. Since 10^3 defines the kilo-prefix, 3.5×10^3 L can be written as 3.5 kL.
   c. 0.045 m can be written as 4.5×10^{-2} m. Since 10^{-2} defines the centi-prefix, 4.5×10^{-2} m can be written as 4.5 cm.

13. a. Since the milli-prefix is defined by 10^{-3}, 43,600 mL can be written as $43{,}600 \times 10^{-3}$ L or 43.6 L.
   b. 0.0000044 m can be written as 4.4×10^{-6} m. Since 10^{-6} defines the micro-prefix, 4.4×10^{-6} m can be written as 4.4 µm.
   c. Since the milli-prefix is defined by 10^{-3}, 1,438 ms can be written as $1{,}438 \times 10^{-3}$ s or 1.438 s.

15. 1,000,000 or 10^6 is used to define the mega-prefix. Hence, the unit for one million bytes is 1 megabyte (Mb).

17. meters/second^2

3. SIGNIFICANT FIGURES

Test Yourself

1. The rectangle width is between 0.5 and 1.0 cm. The rectangle's width is past the first tick mark but not the second; if each tick mark represents 0.1, then the rectangle is at least 0.6 cm. The next place we have to estimate because there are no markings to guide us. It appears to be about3/10 th the distance between the two tick marks, so we will estimate the next place to be 1. Thus, the measured width of the rectangle is 0.63 cm. The measurement is reported to two significant figures.

Test Yourself

1.
    1. By rule 4, the first four zeros are not significant, but by rule 2 the zero between six and one is; therefore, this number has three significant figures.
    2. By rule 2, the zero after the decimal point is significant and by rule 3, zero at the end of the number with a decimal point is significant, so this measurement has five significant figures.

Test Yourself

1. The answer obtained is 21.875. The first number stops its significant figure in the thousands place after the decimal, the second number stops its significant figure in the hundreds place after the decimal, and the third number stops its significant figure in the tenth place after the decimal. Hence, we limit our final answer to the tenth place after the decimal. The final answer is 21.9.

Test Yourself

1.
    1. The answer obtained is 186.2336. The first number has three significant figures and the second number has four significant figures. Hence, we limit our answer to the three significant figures. The final answer is 186.
    2. The answer obtained is 0.2294. The first number has four significant figures and the second number has three significant figures. Hence, we limit our answer to three significant figures. The final answer is 0.229.

Exercises

1.
    a. The arrow is between 300 and 400, so the measurement is at least 300. The arrow is between the seventh and eighth small tick marks, so it is at least 370. We will have to estimate the last place. It looks like about half of the way across the space, so let us estimate the units place as 5. Combining the digits, we have a measurement of 375 psi (psi stands for "pounds per square inch" and is a unit of pressure, like air in a tire). We say that the measurement is reported to three significant figures.

    b. The rectangle width is between 1.0 and 1.5 cm. The rectangle's width is on the third tick mark; if each tick mark represents 0.1, then the rectangle is at least 1.3 cm. The next place we have to estimate because there are no markings to guide us. It appears to be on the third tick mark, so we will estimate the next place to be 0. Thus, the measured width of the rectangle is 1.30 cm. The measurement is reported to two significant figures.

3.
    a. By rule 1, a nonzero digit is significant. Therefore, 23 has two significant figures.

    b. By rule 3, zero at the end of any number with a decimal point is significant. Therefore, 23.0 has three significant figures.

    c. By rule 4, zeros at the beginning of a decimal number (i.e, leading zeros) are not significant. Therefore, 0.00023 has two significant figures.

    d. By rule 4, zeros at the beginning of a decimal number (i.e, leading zeros) are not significant. Therefore, 0.0002302 has four significant figures.

5.
    a. By rule 3, zeros at the end of a number without a decimal point (i.e., trailing zeros) are not significant. Therefore, 765,890 has five significant figures.

    b. By rule 3, zero at the end of any number with a decimal point is significant. Therefore, 765,890.0 has seven significant figures.

    c. By rule 3, zero at the end of any number with a decimal point is significant. Therefore, 1.2000×10^5 has five significant figures.

    d. By rule 3, zero at the end of any number with a decimal point is significant. Therefore, 0.0005060 has four significant figures.

7. a. 56.0 + 3.44 = 59.44. The first number stops its significant figure in the tenths place after the decimal, and the second number stops its significant figure in the hundredths place after the decimal. Hence, we limit our final answer to the tenths place after the decimal. The final answer is 59.4.

b. 0.00665 + 1.004 = 1.01065. The first number stops its significant figure in the ten thousandths place after the decimal, and the second number stops its significant figure in the thousandths place after the decimal. Hence, we limit our final answer to the thousandths place after the decimal. The final answer is 1.011.

c. 45.99 − 32.8 = 13.19. The first number stops its significant figure in the hundredths place after the decimal, and the second number stops its significant figure in the tenths place after the decimal. Hence, we limit our final answer to the tenths place after the decimal. The final answer is 13.2.

d. 45.99 − 32.8 + 75.02 = 88.21. The first number stops its significant figure in the hundredths place after the decimal, the second number stops its significant figure in the tenths place after the decimal, and the third number stops its significant figure in the hundredths place after the decimal. Hence, we limit our final answer to the tenths place after the decimal. The final answer is 88.2.

9. a. 56.7 × 66.99 = 3,798.33. The first number has three significant figures; the second number has 4 significant figures. Hence, the final answer is limited to three significant figures. The final answer is 3,800 or 3.80×10^3.

b. 1.000 ÷ 77 = 0.01298. The first number has four significant figures; the second number has two significant figures. Hence, the final answer is limited to two significant figures. The final answer is 0.013.

c. 1.000 ÷ 77.0 = 0.01298. The first number has four significant figures; the second number has three significant figures. Hence, the final answer is limited to three significant figures. The final answer is 0.0130.

d. 6.022 × 1.89 = 11.38158. The first number has four significant figures; the second number has three significant figures. Hence, the final answer is limited to three significant figures. The final answer is 11.4.

11. a. Scientific notation of 87,449 is 8.7449×10^4. With respect to four significant figures, 8.7449×10^4 can be written as 8.745×10^4.

b. Scientific notation of 0.000066600 is 6.6600×10^{-5}. With respect to five significant figures, 6.6600×10^{-5} stays the same.

13. a. 67,883 × 0.004321 = 293.322. With respect to three significant figures, 293.322 can be written as 293.

b. $(9.67 \times 10^3) \times 0.0055087 = 53.2691$. With respect to three significant figures, 53.2691 can be written as 53.3.

4. CONVERTING UNITS

Test Yourself

1. 1. Conversion factor: $\frac{10^{-6}\text{L}}{1\ \mu\text{L}}$

On multiplying 67.08 µL with the conversion factor, $\frac{10^{-6}\text{L}}{1\ \mu\text{L}} \times 67.08\ \mu\text{L} = 6.708 \times 10^{-5}\text{L}$

2. Conversion factor: $\frac{10^{-3}\text{km}}{1\ \text{m}}$

On multiplying 56.8 m with the conversion factor, $\frac{10^{-3}\text{km}}{1\ \text{m}} \times 56.8\ \text{m} = 5.68 \times 10^{-2}\ \text{km}$

Test Yourself

1. Conversion factor: $\frac{1{,}000\ \text{mm}}{1\ \text{m}}$

Multiply 0.0923 m^3 with the conversion factor three times, since the exponent is 3,

$0.0923\ \text{m}^3 \times \frac{1{,}000\ \text{mm}}{1\ \text{m}} \times \frac{1{,}000\ \text{mm}}{1\ \text{m}} \times \frac{1{,}000\ \text{mm}}{1\ \text{m}} = 9.23 \times 10^7\text{mm}^3$

Test Yourself

1. Conversion factor: $\frac{1\ \text{min}}{60\ \text{s}}$

On multiplying 0.203 m/min with the conversion factor, $0.203\frac{\text{m}}{\text{min}} \times \frac{1\ \text{min}}{60\ \text{s}} = 0.00338\ \text{m}/\text{s}$ or $3.38 \times 10^{-3}\ \text{m}/\text{s}$

Test Yourself

1. $607.8\ \text{kL} \times \frac{1{,}000\ \text{L}}{1\ \text{kL}} \times \frac{1{,}000\ \text{mL}}{1\ \text{L}} = 6.078 \times 10^8\text{mL}$

Test Yourself

1. $2.1\ \text{cm} \times 34.0\ \text{cm} \times 118\ \text{cm} \times \frac{1\ \text{m}}{100\ \text{cm}} \times \frac{1\ \text{m}}{100\ \text{cm}} \times \frac{1\ \text{m}}{100\ \text{cm}} = 8.4252 \times 10^{-3}\ \text{m}^3$

Considering significant figures, $8.4252 \times 10^{-3}\ \text{m}^3$ can be written as $8.4 \times 10^{-3}\ \text{m}^3$ or $0.0084\ \text{m}^3$.

Exercises

1. a. $\frac{1{,}000\ \text{mL}}{1\ \text{L}}$ and $\frac{1\ \text{L}}{1{,}000\ \text{mL}}$

b. $\frac{1{,}000{,}000\ \mu\text{s}}{1\ \text{s}}$ and $\frac{1\ \text{s}}{1{,}000{,}000\ \mu\text{s}}$

c. $\frac{1{,}000\ \text{m}}{1\ \text{km}}$ and $\frac{1\ \text{km}}{1{,}000\ \text{m}}$

3. a. $5.4\ \text{km} \times \frac{1{,}000\ \text{m}}{1\ \text{km}} = 5{,}\ 400\ \text{m}$

b. $0.665\ \text{m} \times \frac{1{,}000\ \text{mm}}{1\ \text{m}} = 665\ \text{mm}$

c. $0.665\ \text{m} \times \frac{1\ \text{km}}{1{,}000\ \text{m}} = 6.65 \times 10^{-4}\ \text{km}$

5. a. $17.8\ \mu g \times \frac{1\ g}{1{,}000{,}000\ \mu g} = 1.78 \times 10^{-5} g$

b. $7.22 \times 10^{2} kg \times \frac{1{,}000\ g}{1\ kg} = 7.22 \times 10^{5} g$

c. $0.00118\ g \times \frac{1{,}000{,}000{,}000\ ng}{1\ g} = 1.18 \times 10^{6} ng$

7. a. $9.44\ m^{2} \times \frac{100\ cm}{1\ m} \times \frac{100\ cm}{1\ m} = 94,\ 400\ cm^{2}$

b. $3.44 \times 10^{8} mm^{3} \times \frac{1\ m}{1{,}000\ mm} \times \frac{1\ m}{1{,}000\ mm} \times \frac{1\ m}{1{,}000\ mm} = 0.344\ m^{3}$

9. One is a unit of area, and the other is a unit of volume.

11. a. $45.0 \frac{m}{min} \times \frac{1\ min}{60\ s} = 0.75\ m \big/ s$

b. $0.000444 \frac{m}{s} \times \frac{1{,}000{,}000\ \mu m}{1\ m} = 444 \mu m \big/ s$

c. $60.0 \frac{km}{h} \times \frac{1\ h}{60\ min} \times \frac{1\ min}{60\ s} = 0.0166\ km \big/ s$ or $1.666 \times 10^{-2}\ km \big/ s$

13. a. $0.674\ kL \times \frac{1{,}000\ L}{1\ kL} \times \frac{1{,}000\ mL}{1\ L} = 674,\ 000\ mL$

b. $2.81 \times 10^{12} mm \times \frac{1\ m}{1{,}000\ mm} \times \frac{1\ km}{1{,}000\ m} = 2.81 \times 10^{6} km$

c. $94.5\ kg \times \frac{1{,}000\ g}{1\ kg} \times \frac{1{,}000\ mg}{1\ g} = 9.45 \times 10^{7} mg$

15. a. $6.77 \times 10^{14} ms \times \frac{1\ s}{1{,}000\ ms} \times \frac{1\ ks}{1{,}000\ s} = 6.77 \times 10^{8} ks$

b. $34,\ 550{,}000\ cm \times \frac{1\ m}{100\ cm} \times \frac{1\ km}{1000\ m} = 345.5\ km$

17. a. $88 \frac{ft}{s} \times \frac{1\ mi}{5{,}280\ ft} \times \frac{60\ s}{1\ min} \times \frac{60\ min}{1\ h} = 60\ mi \big/ h$ or $6.0 \times 10^{1} mi \big/ h$

b. $0.00667 \frac{km}{h} \times \frac{1{,}000\ m}{1\ km} \times \frac{1\ h}{60\ min} \times \frac{1\ min}{60\ s} = 0.00185\ m \big/ s$

19. Area in square millimeters

$$= 2.44\ cm \times 6.077\ cm \times \frac{10\ mm}{1\ cm} \times \frac{10\ mm}{1\ cm}$$

$$= 1,\ 482.788\ mm^{2}$$

Considering significant numbers, 1,482.788 mm^2 can be written as 1,480 mm^2 or 1.48×10^3 mm^2.

21. Area = ½ × 1.007 m × 0.665 m = 0.3348275 m^2

$$\text{Area} = 0.3348275\ m^{2} \times \frac{100\ cm}{1\ m} \times \frac{100\ cm}{1\ m} = 3,\ 348.275\ cm^{2}$$

Considering significant numbers, 3,348.275 cm^2 can be written as 3,350 cm^2 or 3.35×10^3 cm^2.

5. OTHER UNITS: TEMPERATURE AND DENSITY

Test Yourself

1. 1. Using the formula, $^\circ\mathrm{C} = (^\circ\mathrm{F} - 32) \times \frac{5}{9}$

$$^\circ\mathrm{C} = (0 - 32) \times \frac{5}{9} = -17.8\ ^\circ\mathrm{C}$$

2. Using the formula, $^\circ\mathrm{F} = \left(^\circ\mathrm{C} \times \frac{9}{5}\right) + 32$

$$^\circ\mathrm{F} = \left(212 \times \frac{9}{5}\right) + 32 = 413.6\ ^\circ\mathrm{F}$$

Considering significant numbers, 413.6°F can be written as 414°F.

Test Yourself

1. Using the formula, $^\circ\mathrm{C} = (^\circ\mathrm{F} - 32) \times \frac{5}{9}$

$$^\circ\mathrm{C} = (98.6 - 32) \times \frac{5}{9} = 37\ ^\circ\mathrm{C}$$

K = 37°C + 273.15 = 310.15 K
Considering significant numbers, 310.15 K can be written as 310.2 K.

Test Yourself

1. From the Table 2.2 in the text, density of iron is 7.87 $\mathrm{g/cm^3}$.
To calculate the mass of iron:

$$25.0\ \mathrm{cm^3} \times \frac{7.87\ \mathrm{g}}{1\ \mathrm{cm^3}} = 196.75\ \mathrm{g}$$

Considering significant numbers 196.75 g can be written as 197 g.

Test Yourself

1. From the Table 2.2 in the text, density of gold is 19.3 $\mathrm{g/cm^3}$.
To calculate the volume of gold:

$$3.78\ \mathrm{g} \times \frac{1\ \mathrm{cm^3}}{19.3\ \mathrm{g}} = 0.196\ \mathrm{cm^3}$$

Exercises

1. a. Using the formula, $^\circ\mathrm{C} = (^\circ\mathrm{F} - 32) \times \frac{5}{9}$

   $$^\circ\mathrm{C} = (255 - 32) \times \frac{5}{9} = 124\ ^\circ\mathrm{C}$$

   b. Using the formula, $^\circ\mathrm{C} = (^\circ\mathrm{F} - 32) \times \frac{5}{9}$

   $$^\circ\mathrm{C} = (-255 - 32) \times \frac{5}{9} = -159\ ^\circ\mathrm{C}$$

   c. Using the formula, $^\circ\mathrm{F} = \left(^\circ\mathrm{C} \times \frac{9}{5}\right) + 32$

   $$^\circ\mathrm{F} = \left(50 \times \frac{9}{5}\right) + 32 = 122\ ^\circ\mathrm{F}$$

   d. Using the formula, $^\circ\mathrm{F} = \left(^\circ\mathrm{C} \times \frac{9}{5}\right) + 32$

   $$^\circ\mathrm{F} = \left(-50 \times \frac{9}{5}\right) + 32 = -58\ ^\circ\mathrm{F}$$

3. a. K = 100.0 + 273.15 = 373 K
   b. K = −100.0 + 273.15 = 173 K
   c. °C = 100 − 273.15 = −173°C
   d. °C = 300 − 273.15 = 27°C
5. −273°C. This is the lowest possible temperature in degrees Celsius.
7. Using the formula, $^\circ\mathrm{C} = (^\circ\mathrm{F} - 32) \times \frac{5}{9}$

   $$^\circ\mathrm{C} = (136 - 32) \times \frac{5}{9} = 57.8^\mathrm{o}\mathrm{C}$$

   K = 57.8 + 273.15 = 331 K
9. g/mL, g/L, and kg/L (answers will vary)

11. Using the Table 2.2 in the text, density of iron is 7.87 g/cm^3.

    To calculate the mass of the iron sample:

    $$48.2\ \mathrm{cm}^3 \times \frac{7.87\ \mathrm{g}}{1\ \mathrm{cm}^3} = 379\ \mathrm{g}$$

13. To calculate the mass of hydrogen used by the airship:

    $$2.000 \times 10^8\,\mathrm{L} \times \frac{0.0899\ \mathrm{g}}{1\ \mathrm{L}} = 1.80\ \times 10^7\,\mathrm{g}$$

15. Using the Table 2.2 in the Text, density of gold is 19.3 g/cm^3.

    To calculate the mass of gold present in the ring:

    $$0.77\ \mathrm{cm}^3 \times \frac{19.3\ \mathrm{g}}{1\ \mathrm{cm}^3} = 15\ \mathrm{g}$$

17. Density = 11.34 g/cm^3

    To calculate the volume of lead:

    $$100.0\ \mathrm{g} \times \frac{1\ \mathrm{cm}^3}{11.34\ \mathrm{g}} = 8.818\ \mathrm{cm}^3$$

19. Density = 0.900 g/L

    To calculate the volume of neon:

    $$222\ \mathrm{g} \times \frac{1\ \mathrm{L}}{0.900\ \mathrm{g}} = 247\ \mathrm{L}$$

21. For iron, density = 7.87 g/cm^3.

    To calculate the volume of iron:

    $$100.0\ \mathrm{g} \times \frac{1\ \mathrm{cm}^3}{7.87\ \mathrm{g}} = 12.7\ \mathrm{cm}^3$$

    For gold, density = 19.3 g/cm^3.

    To calculate the volume of gold:

    $$75.0\ \mathrm{g} \times \frac{1\ \mathrm{cm}^3}{19.3\ \mathrm{g}} = 3.89\ \mathrm{cm}^3$$

    The 100 g of iron has the greater volume.

6. END-OF-CHAPTER MATERIAL

Additional Exercises

1. $0.00000000552 = 5.52 \times 10^{-9}$

$0.0000000006188 = 6.188 \times 10^{-10}$

$5.52 \times 10^{-9} \times 6.188 \times 10^{-10} = 3.42 \times 10^{-18}$

3. 602,200,000,000,000,000,000,000

5. $(4.506 \times 10^{4}) \times (1.003 \times 10^{2}) = (4.506 \times 1.003) \times 10^{2+4} = 4.520 \times 10^{6}$

7. The quantity is two; dozen is the unit.

9. $1\ \text{km} \times \frac{1{,}000\ \text{m}}{1\ \text{km}} \times \frac{1{,}000{,}000\ \mu\text{m}}{1\ \text{m}} = 1,\ 000{,}000{,}000\mu\text{m}$

11. $1\ \text{cL} \times \frac{1\ \text{L}}{100\ \text{cL}} \times \frac{1\ \text{ML}}{1{,}000{,}000\ \text{L}} = \frac{1}{100{,}000{,}000}\ \text{ML}$

13. $67.3\frac{\text{km}}{\text{h}} \times \frac{1{,}000\ \text{m}}{1\ \text{km}} \times \frac{1\ \text{h}}{60\ \text{min}} \times \frac{1\ \text{min}}{60\ \text{s}} = 18.7\ \text{m}\big/\text{s}$

15. $60.0\frac{\text{mi}}{\text{h}} \times \frac{1.602\ \text{km}}{1\ \text{mi}} = 96.1\ \text{km}\big/\text{h}$

17. $52.09 \times \frac{\text{km}}{\text{h}} \times \frac{1{,}000\ \text{m}}{1\ \text{km}} \times \frac{1\ \text{h}}{60\ \text{h}} \times \frac{1\ \text{min}}{60\ \text{s}} = 14.47\ \text{m}\big/\text{s}$

19. One Fahrenheit degree is 9/5ths the size of a Celsius degree.

21. Using the Table 2.2 in the text, density of mercury is 13.6 g/cm^3.

$1\ \text{cm}^3 = 1\ \text{mL}$

To calculate the mass of mercury:

$12.67\ \text{L} \times \frac{10^3\ \text{mL}}{1\ \text{L}} \times \frac{1\ \text{cm}^3}{1\ \text{mL}} \times \frac{13.6\ \text{g}}{1\ \text{cm}^3} = 1.72 \times 10^5\ \text{g}$

23. Using the Table 2.2 in the text, density of gold is 19.3 g/cm^3.

To calculate the volume of gold:

$2.884\ \text{kg} \times \frac{1{,}000\ \text{g}}{1\ \text{kg}} \times \frac{1\ \text{cm}^3}{19.3\ \text{g}} = 149.4\ \text{cm}^3\text{or mL}$

CHAPTER 3

Atoms, Molecules, and Ions

1. ATOMIC THEORY

Test Yourself

1. If the tin atom has 50 protons in its nucleus, then the atomic number is 50. It has 68 neutrons in the nucleus. So, its mass number is 50 + 68, or 118.

Test Yourself

1. This sodium atom has 11 protons.

Exercises

1. All matter is composed of atoms; atoms of the same element are the same, and atoms of different elements are different; atoms combine in whole number ratios to form compounds.
3. A proton is larger than an electron.
5. proton: 1+; electron: 1−; neutron: 0
7. Figure 3.X1. is missing.
9. The atomic number is the number of protons in the nucleus of an atom. Boron atom has 5 protons in its nucleus. So, its atomic number is five.
11. Isotopes are atoms of the same element but with different numbers of neutrons. ${}^{1}_{1}\mathrm{H}$ and ${}^{2}_{1}\mathrm{H}$ are examples.
13. a. ${}^{4}_{2}\mathrm{He}$ has 2 protons and 4 − 2, or 2 neutrons. ${}^{3}_{2}\mathrm{He}$ has 2 protons and 3 − 2, or 1 neutron. The two atoms have the same number of protons but different number of neutrons. So, they are isotopes.
    b. ${}^{56}_{26}\mathrm{Fe}$ has 26 protons and 56 − 26, or 30 neutrons. ${}^{56}_{25}\mathrm{Mn}$ has 25 protons and 56 − 25, or 31 neutrons. The two atoms have different number of protons and neutrons. So, they are not isotopes.
    c. ${}^{28}_{14}\mathrm{Si}$ has 14 protons and 28 − 14, or 14 neutrons. ${}^{31}_{15}\mathrm{P}$ has 15 protons and 31 − 15, or 16 neutrons. They have different number of protons and neutrons. So, they are not isotopes.
15. a. An oxygen atom with 8 protons has an atomic number of 8. It has 8 neutrons. So, the mass number is 8 + 8, or 16. Hence, it is represented as ${}^{16}_{8}\mathrm{O}$.
    b. A potassium atom with 19 protons has an atomic number of 19. It has 20 neutrons. So, its mass number is 19 + 20, or 39. Hence, it is represented as ${}^{39}_{19}\mathrm{K}$.
    c. A lithium atom with 3 protons has an atomic number of 3. It has 4 neutrons. So, the mass number is 3 + 4, or 7. Hence, it is represented as ${}^{7}_{3}\mathrm{Li}$.
17. Americium has an atomic number of 95. Its mass number is 241. So, the symbol for this isotope is ${}^{241}_{95}\mathrm{Am}$.
19. a. Na
    b. Ar
    c. N
    d. Rn
21. a. silicon
    b. manganese
    c. iron
    d. chromium

2. MOLECULES AND CHEMICAL NOMENCLATURE

Test Yourself

1.
  1. A molecule with one sulfur atom and four fluorine atoms is named as sulfur tetrafluoride.
  2. A molecule with two phosphorus atoms and five sulfur atoms is diphosphorus pentasulfide.

Test Yourself

1.
  1. The name *disulfur difluoride* indicates 2 sulfur atoms and 2 fluorine atoms. So, the compound is S_2F_2.
  2. The name *iodine pentabromide* indicates one iodine atom and five bromine atoms. So, the compound is IBr_5.

Exercises

1.
  a. Fe is not a molecule. It represents the element iron.
  b. PCl_3 is a molecule. It has one phosphorus atom and three chlorine atoms. Hence, PCl_3 has 4 atoms.
  c. P_4 is a molecule having four phosphorus atoms.
  d. Ar is not a molecule. It represents the element argon.

3. CO is a compound of carbon and oxygen; Co is the element cobalt.

5. Hydrogen - H_2
Oxygen - O_2
Nitrogen - N_2
Fluorine - F_2
Chlorine - Cl_2
Bromine - Br_2
Iodine - I_2

7. The stem of fluorine used in molecule names is *fluor-*. For example, CF_4 is named as carbon tetrafluoride.

9.
  a. The molecule has one phosphorus atom and three fluorine atoms. So, it is named phosphorus trifluoride.
  b. The molecule has one tellurium atom and two chlorine atoms. So, it is named tellurium dichloride.
  c. The molecule has two nitrogen atoms and three oxygen atoms. So, it is named dinitrogen trioxide.

11.
  a. The molecule with one xenon atom and two fluorine atoms is called xenon difluoride.
  b. The molecule with two oxygen atoms and two fluorine atoms is called dioxygen difluoride.
  c. The molecule with one sulfur atom and six fluorine atoms is called sulfur hexafluoride.

13.
  a. The molecule with two nitrogen atoms and one oxygen atom is called dinitrogen monoxide.
  b. The molecule with two nitrogen atoms and four oxygen atoms is called dinitrogen tetroxide.
  c. The molecule with two nitrogen atoms and five oxygen atoms is called dinitrogen pentoxide.

15.
  a. The name implies two nitrogen atoms and five oxygen atoms. So, the formula is N_2O_5.
  b. The name implies four boron atoms and three carbon atoms. So, the formula is B_4C_3.
  c. The name implies one phosphorus atom and five chlorine atoms. So, the formula is PCl_5.

17.
  a. The name implies two oxygen atoms and two chlorine atoms. So, the formula is O_2Cl_2.
  b. The name implies two nitrogen atoms and three sulfur atoms. So, the formula is N_2S_3.
  c. The name implies one xenon atom and four fluorine atoms. So, the formula is XeF_4.

19.
  a. The name implies one iodine atom and three fluorine atoms. So, the formula is IF_3.
  b. The name implies one xenon atom and three oxygen atoms. So, the formula is XeO_3.
  c. The name implies two sulfur atoms and ten fluorine atoms. So, the formula is S_2F_{10}.

3. MASSES OF ATOMS AND MOLECULES

Test Yourself

1. 1. Add one atomic mass of sulfur and two atomic masses of oxygen:

1 S mass	= 32.065 u
2 O masses = 2 × 15.999 u	= 31.998 u
Total	= 64.063 u = the molecular mass of SO_2

2. Add one atomic mass of phosphorus and three atomic masses of fluorine:

1 P mass	= 30.974 u
3 F masses = 3 × 18.998 u	= 56.994 u
Total	= 87.968 u = the molecular mass of PF_3

Exercises

1. The atomic mass unit is defined as one-twelfth of the mass of a carbon-12 atom. Its abbreviation is u.

3. We can estimate the mass of an atom, in whole number, by simply counting the total number of protons and neutrons in the nucleus (i.e., identify its mass number) and ignore the electrons.

a. The mass of hydrogen-1is1 u.

b. The mass of hydrogen-3 is 3 u.

c. The mass of hydrogen-56 is 56 u.

5. a. The atomic mass of lithium is the weighted average of the masses of the isotopes of lithium.

0.924 × 7.016 u	= 6.483 u
0.076 × 6.015 u	= 0.457 u
Sum	= 6.940 u = the atomic mass of lithium

b. The atomic mass of oxygen is the weighted average of the masses of the isotopes of oxygen.

0.9976 × 15.995 u	= 15.957 u
0.00035× 16.999 u	= 0.006 u
0.00205 × 17.999 u	= 0.037 u
Sum	= 16.000 u = the atomic mass of oxygen

7. The whole number mass of lithium-7 isotope is 7u and that of lihium-6 is 6u. So, the atomic mass of lithium is:

0.924 × 7 u	= 6.468 u
0.076 × 6 u	= 0.456 u
Sum	= 6.924 u = the atomic mass of lithium

9. a. The atomic mass of an oxygen atom is 15.999 u.

b. Oxygen exists as a diatomic molecule, O_2. So, the molecular mass of oxygen is 2 × 15.999 u = 31.998 u.

11. a. The atomic mass of fluorine is 18.998 u. So, the mass of F_2 is 2F masses = 2 × 18.998 u = 37.996 u.

b. The mass of CO is:

1 C mass	= 12.011 u
1 O mass	= 15.999 u
Total	= 28.010 u = the mass of CO

c. The mass of CO_2 is:

1 C mass	= 12.011 u
2 O masses = 2 × 15.999 u	= 31.998 u
Total	= 44.009 u = the mass of CO_2

13. a. The atomic mass of sodium is 22.990 u.

b. The mass of diboron trioxide is:

2 B masses = 2 × 10.811 u	= 21.622 u
3 O masses = 3 × 15.999 u	= 47.997 u
Total	= 69.619 u = the mass of B_2O_3

c. The mass of disulfur dichloride is:

2 S masses = 2 × 32.065 u	= 64.130 u
2 Cl masses = 2 × 35.453 u	= 70.906 u
Total	= 135.036 u = the mass of S_2Cl_2

15. a. The mass of germanium dioxide molecule is:

1 Ge mass	= 72.64 u
2 O masses = 2 × 15.999 u	= 31.998 u
Total	= 104.64 u = the mass of GeO_2

b. The mass of iodine trifluoride molecule is:

1 I mass	= 126.904 u
3 F masses = 3 × 18.998 u	= 56.994 u
Total	= 183.898 u = the mass of IF_3

c. The mass of a xenon hexafluoride molecule is:

1 Xe mass	= 131.293 u
6 F masses = 6 × 18.998 u	= 113.988 u
Total	= 245.281 u = the mass of XeF_6

4. IONS AND IONIC COMPOUNDS

Test Yourself

1.
1. This species has a negative charge (3−) on it, so it is an anion. Anions are named using the stem of the element name, then adding the suffix —*ide*. This is a phosphide anion.
2. This species has a positive charge (2+) on it, so it is a cation. For a cation, we simply use the element's name and the word 'ion' or 'cation' after it. This is a strontium cation.

Test Yourself

1.
1. We need only one of each ion to balance the charges. So, The proper ionic formula is FeS.
2. With Fe^{3+} and S^{2-}, neither charge is a perfect multiple of the other. So, we need to find the least common multiple of 2 and 3, which is 6. To get a total of 6+ we need two Fe^{3+} ions; to get 6–, we need three S^{2-} ions. The proper ionic formula is Fe_2S_3.

Test Yourself

1.
1. Using the names of the ions, this ionic compound is named scandium oxide.
2. Using the names of the ions, this ionic compound is named silver chloride.

Test Yourself

1.
1. We need three ammonium ions to balance the charge on the phosphate ion. So, the formula of the compound formed is $(NH_4)_3PO_4$. The name of the compound is ammonium phosphate.
2. We need three nitrite ions to balance the charge on the cobalt ion. So, the formula of the compound formed is $Co(NO_2)_3$. The name of the compound is cobalt (III) nitrite.

Exercises

1. Cations form by losing electrons.

3.
- a. Potassium takes a 1+ charge when it forms an ion.
- b. Oxygen takes a 2– charge when it forms an ion.
- c. Cobalt can form two ions; one with a 2+ charge and the other with a 3+ charge.

5.
- a. Silver takes a 1+ charge when it forms an ion.
- b. Gold can form two ions; one with a 1+ charge, and the other with a 3+ charge.
- c. Bromine takes a 1– charge when it forms an ion.

7.
- a. K^+ ion is the potassium ion.
- b. O^{2-} is the oxide ion.
- c. Co^{2+} is the cobalt(II) ion ; Co^{3+} is the cobalt(III) ion.

9.
- a. Ag^+ is the silver ion.
- b. Au^+ is the gold(I) ion; Au^{3+} is the gold(III) ion.
- c. Br^- is the bromide ion.

11.
- a. We need two chloride ions to balance the charge on the magnesium ion. So, the formula of the compound formed is $MgCl_2$. The name of the compound is magnesium chloride.
- b. We need one atom each to balance the charges. So, the formula of the compound formed is FeO. The name of the compound is iron(II) oxide.
- c. Both the ions have different charges and neither charge is an exact multiple of the other. So, we need to consider the least common multiple, 6 in this case. To get 6+, we need two iron(III) ions; and to get 6–, we need three oxide ions. The formula of the compound formed is Fe_2O_3. The name of the compound is iron(III) oxide.

13.
- a. We need two fluoride ions to balance the charge on a single copper ion. So, the formula of the compound formed is CuF_2. The name of the compound is copper(II) fluoride.
- b. We need one of each ion to balance the charges. So, the formula of the compound formed is CaO. The name of the compound is calcium oxide.
- c. We need three potassium ions to balance the charge on the phosphide ion. So, the formula of the compound formed is K_3P and the name of the compound is potassium phosphide.

15. a. We need two potassium ions to balance the charge on the sulfate ion. So, the formula of the compound formed is K_2SO_4. The name of the compound is potassium sulfate.

b. We need two ammonium ions to balance the charge on the sulfide ion. So, the formula of the compound formed is $(NH_4)_2S$. The name of the compound is ammonium sulfide.

c. We need three ammonium ions to balance the charge on the phosphate ion. So, the formula of the compound formed is $(NH_4)_3PO_4$. The name of the compound is ammonium phosphate.

17. a. We need two sulfate ions to balance the charge on the lead ion. The formula of the compound formed is $Pb(SO_4)_2$. The name of the compound is lead(IV) sulfate.

b. We need only one of each ion to balance the charges. So, the formula of the compound formed is NaI_3. The name of the compound is sodium triiodide.

c. We need two lithium ions to balance the charge on the dichromate ion. So, the formula of the compound formed is $Li_2Cr_2O_7$. The name of the compound is lithium dichromate.

19. a. We need two silver ions to balance the charge on the sulfite ion. So, the formula of the compound formed is Ag_2SO_3. The name of the compound is silver sulfite.

b. We need only one of each ion to balance the charges. So, the formula of the compound formed is $NaHCO_3$. The name of the compound is sodium hydrogen carbonate.

c. We need three ions of chlorate to balance the charge on the iron(III) ion. So, the formula of the compound formed is $Fe(ClO_3)_3$. The name of the compound is iron(III) chlorate.

21. SO_3 is sulfur trioxide, while SO_3^{2-} is the sulfite ion.

5. ACIDS

Test Yourself

1.
   1. As a binary acid, the name is hydro- + stem name + -ic acid. Because this acid contains a fluorine atom, the name is hydrofluoric acid
   2. Because this acid is derived from the nitrite ion, the name of the acid is the stem of the anion name + -ous acid. The name of this acid is nitrous acid.

Exercises

1.
   a. The acid is composed of hydrogen ions and the chlorate anion. Thus, the formula is $HClO_4(aq)$

   b. The acid is composed of hydrogen ions and the iodine anion. Thus, the formula is HI(aq).

3.
   a. As a binary acid, the name is hydro- + stem name + -ic acid. Because this acid contains a fluorine atom, the name is hydrofluoric acid.

   b. Because this acid is derived from the nitrate ion, the name of the acid is the stem of the anion name + -ic acid. The name of this acid is nitric acid.

   c. Because this acid is derived from the oxalate ion, the name of the acid is the stem of the anion name + -ic acid. The name of this acid is oxalic acid.

5. Oxalic acid ($H_2C_2O_4(aq)$), found in spinach and other green leafy vegetables. (answers will vary)

6. END-OF-CHAPTER MATERIAL

Additional Exercises

1. mass of one proton = 1.6×10^{-27} kg; mass of one electron = 9.1×10^{-31} kg

To calculate the number of electrons that makes the mass of one proton:

$$\frac{\text{mass of one proton}}{\text{mass of one electron}} = \frac{1.6 \times 10^{-27}\ \text{kg}}{9.1 \times 10^{-31}\ \text{kg}} = 1758.24 \approx 1800$$

Around 1800 electrons make the mass of one proton.

3. It is not strictly correct because of the existence of isotopes.

5. The three isotopes of hydrogen are hydrogen-1, with one proton and no neutron; hydrogen-2 or deuterium, with one proton and one neutron; hydrogen-3 or tritium, with one proton and two neutrons. Their atomic symbols are $^{1}_{1}H$, $^{2}_{1}H$, and $^{3}_{1}H$ respectively.

7. It is a metal.

9. The atomic number of americium is 95. So, it has 95 protons. The mass number of americium-241 is 241. Thus, the number of neutrons present in its nucleus is 241 – 95, or 146.

11. The atomic mass of ruthenium is the weighted average of the masses of its isotopes. It can be calculated as follows:

Ruthenium-96	0.0554 × 95.907 u	= 5.313 u
Ruthenium-98	0.0187 × 97.905 u	= 1.831 u
Ruthenium-99	0.1276 × 98.906 u	= 12.620 u
Ruthenium-100	0.1260 × 99.904 u	= 12.588 u
Ruthenium-101	0.1706 × 100.906 u	= 17.215 u
Ruthenium-102	0.3155 × 101.904 u	= 32.151 u
Ruthenium-104	0.1862 × 103.905 u	= 19.347 u
Total		= 101.065 u = the mass of Ru

13. Atomic mass of sodium is 22.99 u.

Mass of one sodium atom is:

$$22.99\ \text{u} \times \frac{1.6605 \times 10^{-24}\ \text{g}}{1\ \text{u}} = 3.817 \times 10^{-23}\ \text{g}$$

15. The mass of one molecule of water is:

2 masses H = 2 × 1.008	= 2.016 u
1 mass O	= 15.999 u
Sum	= 18.015 u

Mass of one H_2O molecule is:

$$18.015\ \text{u} \times \frac{1.6605 \times 10^{-24}\ \text{g}}{1\ \text{u}} = 2.991 \times 10^{-23}\ \text{g}$$

17. From the periodic table, we find that copper is a metal and iodine a non metal. They combine to form an ionic compound.

19.
- a. Since mercury exists as a diatomic ion, the overall charge on it will be 2+, with each ion contributing a 1+ charge. Thus, the formula for the mercury(I) ion is Hg_2^{2+}.
- b. A chloride ion has a 1– charge on it. So, we need two chloride ions to balance the charge on the mercury ion. Thus, the formula is Hg_2Cl_2.

21. The fluoride ion has 1– charge. So, we need two fluoride ions to balance the charge on the uranyl cation. Thus, the formula is UO_2F_2 and the name is uranyl fluoride. The sulfate ion has the same charge as uranyl cation. So, we need one of each to balance the charges. Thus, the formula is UO_2SO_4 and the name is uranyl sulfate.

Neither of the charges on phosphate ion and uranyl cation are exact multiples of each other. So, we need to take the least common multiple, which is 6 in this case. To get 6+ we need three uranyl ions; to get 6–, we need two phosphate ions. Thus, the formula is $(UO_2)_3(PO_4)_2$ and the name is uranyl phosphate.

CHAPTER 4

Chemical Reactions and Equations

1. THE CHEMICAL EQUATION

Test Yourself

1. Let us start by simply writing a chemical equation in terms of the formulas of the substances, remembering that both elemental hydrogen and nitrogen are diatomic.

$N_2 + H_2 \rightarrow NH_3$

The reactant side has two N atoms; the product side has one N atom. So, add 2 before NH3 to balance the number of N atoms on both the sides.

$N_2 + H_2 \rightarrow 2NH_3$

The reactant side has two H atoms; the product side has six H atoms. So, add 3 before H_2 to balance the number of H atoms on both the sides.

$N_2 + 3H_2 \rightarrow 2NH_3$

Hence, this is the balanced reaction.

Exercises

1. reactants: nitrogen and hydrogen; product: ammonia

3. reactants: magnesium hydroxide and nitric acid; products: magnesium nitrate and water

5. reactants: nitrogen and hydrogen; product: ammonia

Let us start by simply writing a chemical equation in terms of the formulas of the substances, remembering that both elemental hydrogen and nitrogen are diatomic.

$N_2 + H_2 \rightarrow NH_3$

The reactant side has two N atoms; the product side has one N atom. So, add 2 before NH3 to balance the number of N atoms on both the sides.

$N_2 + H_2 \rightarrow 2NH_3$

The reactant side has two H atoms; the product side has six H atoms. So, add 3 before H_2 to balance the number of H atoms on both the sides.

$N_2 + 3H_2 \rightarrow 2NH_3$

Hence, this is the balanced reaction.

7. reactants: magnesium hydroxide and nitric acid; products: magnesium nitrate and water

Let us start by simply writing a chemical equation in terms of the formulas of the substances.

$Mg(OH)_2 + HNO_3 \rightarrow Mg(NO_3)_2 + H_2O$

The number of Mg atoms on both the sides of the reaction is balanced. The number of N atoms on the reactant side is one; N atoms on the product side are two. So, add 2 before HNO3 to balance the number of N atoms on both the sides.

$Mg(OH)_2 + 2HNO_3 \rightarrow Mg(NO_3)_2 + H_2O$

There are eight O atoms on the reactant side and seven on the product side. So, add 2 before H_2O to balance the number of O atoms on both the sides.

$Mg(OH)_2 + 2HNO_3 \rightarrow Mg(NO_3)_2 + 2H_2O$

The number of H atoms on both the sides is balanced. Hence, this is the balanced equation.

9. The number of sodium and chlorine atoms on both the sides is balanced. There are three O atoms on the reactant side. However, because oxygen is a diatomic molecule, we can only get an even number of oxygen atoms at a time on the product side. To get an even number of O atoms on the reactant side, all the coefficients are multiplied by 2, except O_2.

$2NaClO_3 \rightarrow 2NaCl + O_2$

There are six O atoms on the reactant side; two O atoms on the product side. So, add 3 before O_2 to balance the number of O atoms on both the sides.

$2NaClO_3 \rightarrow 2NaCl + 3O_2$

This is the balanced reaction.

11. There are two atoms of Al on the product side; one Al atom on the reactant side. So, add 2 before Al to balance the number of Al atoms on both the sides.

$2Al + O_2 \rightarrow Al_2O_3$

There are three O atoms on the product side; two O atoms on the reactant side. But reactant side can have O atoms only in multiples of two. So, add 2 before Al_2O_3. Then the number of O atoms on the product side will be six. To balance the number of Al atoms again, add 4 before Al.

$4Al + O_2 \rightarrow 2Al_2O_3$

There are six O atoms on the product side; two O atoms on the reactant side. So, add 3 before O_2 to balance the number of O atoms on both the sides.

$4Al + 3O_2 \rightarrow 2Al_2O_3$

This is the balanced reaction.

13. The balanced chemical reaction is:

$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

2. TYPES OF CHEMICAL REACTIONS: SINGLE- AND DOUBLE-DISPLACEMENT REACTIONS

Test Yourself

1. Because chlorine is above iodine in the periodic table, a single-replacement reaction will occur. The products are FeCl2 and I2.

   $$\mathrm{FeI_2 + Cl_2 \rightarrow FeCl_2 + I_2}$$

Test Yourself

1. Since magnesium is above aluminum in the activity series, it will replace aluminum in the compound. The products of this single replacement reaction are Mg3(PO4)2 and Al.

Test Yourself

1. Thinking about the reaction as either swapping the cations, or swapping the anions, the products would be KNO3 and AgBr.

Test Yourself

1. According to the solubility rules, both Sr(NO3)2 and KCl are soluble. Now we consider what the double-replacement products would be by switching the cations (or the anions)— the products are SrCl2 and KNO3. However, the solubility rules predict that these two substances would also be soluble, so no precipitate would form. Thus, we predict no reaction in this case.

Exercises

1. One element replaces another element in a compound.
3. a. In this reaction, zinc replaces iron and results in the formation of zinc nitrate. Hence, the products are zinc nitrate and iron. The balanced reaction is:

   $$\mathrm{Zn{+}Fe\left(NO_3\right)_2 \rightarrow Zn\left(NO_3\right)_2 + Fe}$$

   b. In this reaction, fluorine replaces iodine. The products are iron(III) fluoride and iodine. The balanced reaction is:

   $$\mathrm{3F_2 + 2FeI_3 \rightarrow 3I_2 + 2FeF_3}$$

5. a. Tin replaces hydrogen from the acid in the reaction. The products are tin sulfate and hydrogen gas. The balanced reaction is:

   $$\mathrm{Sn{+}H_2SO_4 \rightarrow SnSO_4 + H_2}$$

   b. Aluminum replaces nickel in the reaction. The products are aluminum bromide and nickel. The balanced reaction is:

   $$\mathrm{2Al + 3NiBr_2 \rightarrow 2AlBr_3 + 3Ni}$$

7. a. Since chlorine lies above bromine in the periodic table, it cannot be replaced by bromine. Hence, the reaction does not occur.

   b. Since aluminum lies above iron in the activity series, it can replace iron from its nitrate. Hence, the balanced chemical equation is:

   $$\mathrm{Fe\left(NO_3\right)_3 + Al \rightarrow Al\left(NO_3\right)_3 + Fe}$$

9. a. Since chlorine lies above iodine in the periodic table, it can replace iodine from its compound. Hence, this single-replacement reaction will occur. The balanced chemical reaction is:

   $$\mathrm{2NaI + Cl_2 \rightarrow 2NaCl + I_2}$$

   b. Since gold is the last element in the activity series, it cannot replace any other metal from a compound as all metals lie above it in the activity series. So, no reaction occurs.

11. a. If we consider the double-replacement reaction to occur, by switching the cations (or anions), the products would be zinc hydroxide and sodium nitrate. The balanced chemical reaction is:

$$Zn(NO_3)_2 + 2NaOH \rightarrow Zn(OH)_2 + 2NaNO_3$$

b. Assuming that the double-replacement reaction will occur, by switching the cations (or anions), the products would be sodium chloride and hydrogen sulfide. The balanced reaction is:

$$2HCl + Na_2S \rightarrow 2NaCl + H_2S$$

13. a. Assuming that the double- replacement reaction occurs, by switching the cations (or anions), the products would be lead bromide and potassium nitrate. The balanced equation is:

$$Pb(NO_3)_2 + 2KBr \rightarrow PbBr_2 + 2KNO_3$$

b. Assuming that the double-replacement reaction occurs, by switching the cations (or anions), the products would be potassium carbonate and magnesium oxide. The balanced equation is:

$$K_2O + MgCO_3 \rightarrow K_2CO_3 + MgO$$

15. a. If we consider the double-replacement reaction to occur, by switching the cations (or anions), the products would be $PbBr_2$ and KNO_3. From, the solubility rules we find that potassium nitrate is soluble but lead bromide is not soluble. Thus, the reaction would occur and the balanced reaction is:

$$Pb(NO_3)_2 + 2KBr \rightarrow PbBr_2(s) + 2KNO_3$$

b. If we consider the double-replacement reaction to occur, by switching the cations (or anions), the products would be Na_2O and K_2CO_3. From solubility rules we find that all sodium and potassium compounds are soluble without any exception. Thus, no reaction occurs.

17. a. If we consider the double- replacement reaction to occur, by switching the cations (or anions), the products would be $Sr_3(PO_4)_2$ and KCl. From solubility rules we find that potassium chloride is soluble but strontium phosphate is insoluble. Thus, the reaction will occur. The balanced equation is

$$2K_3PO_4 + 3SrCl_2 \rightarrow Sr_3(PO_4)_2(s) + 6KCl$$

b. If we consider the double- replacement reaction to occur, by switching the cations (or anions), the products would be NaCl and $Mg(OH)_2$. From solubility rules we find that sodium chloride is soluble but magnesium hydroxide is insoluble. Thus, the reaction will occur. The balanced chemical reaction is:

$$2NaOH + MgCl_2 \rightarrow 2NaCl + Mg(OH)_2(s)$$

3. IONIC EQUATIONS: A CLOSER LOOK

Test Yourself

1. On dissociation nitrogen and hydrogen in ammonium ion stay together as the ammonium ion but sulfide ions go their own way. The chemical reaction that represents the dissociation of ammonium sulfide is:

$$(NH_4)_2S(s) \rightarrow 2NH_4{}^{+}(aq) + S^{2-}(aq)$$

Test Yourself

1. For any ionic compound that is aqueous, we will write the compound as separated ions. The compound that is solid (a precipitate) is written as it is, and not as its ions. The complete ionic equation is

$$Ca^{2+}(aq) + 2Cl^{-}(aq) + Pb^{2+}(aq) + 2NO_3^{-}(aq) \rightarrow Ca^{2+}(aq) + 2NO_3^{-}(aq) + PbCl_2(s)$$

Test Yourself

1. The complete ionic equation is

$$Ca^{2+}(aq) + 2Cl^{-}(aq) + Pb^{2+}(aq) + 2NO_3^{-}(aq) \rightarrow Ca^{2+}(aq) + 2NO_3^{-}(aq) + PbCl_2(s)$$

On cancelling the spectator ions, Ca^{2+} and $2NO3^{-}$, we get:

$$Pb^{2+}(aq) + 2Cl^{-}(aq) \rightarrow PbCl_2(s)$$

Exercises

1. In water, sodium bromide would dissociate into sodium ions and bromide ions. This can be represented by the following chemical equation:

$$NaBr(s) \xrightarrow{H_2O} Na^{+}(aq) + Br^{-}(aq)$$

3. In water ammonium phosphate dissociates into ammonium ions and phosphate ions. This can be represented by the following chemical equation:

$$(NH_4)_3PO_4(s) \xrightarrow{H_2O} 3NH_4{}^{+}(aq) + PO_4{}^{3-}(aq)$$

5. Iron(II) chloride and silver nitrate undergo double-replacement reaction. By switching the cations (or anions), the products obtained are iron(II) nitrate and silver chloride. From the solubility rules we find that iron nitrate is soluble but, silver chloride is insoluble. Hence, the precipitation reaction occurs. The balanced reaction is:

$$FeCl_2(aq) + 2AgNO_3(aq) \rightarrow Fe(NO_3)_2(aq) + 2AgCl(s)$$

In water iron(II) chloride, silver nitrate, and iron(II) nitrate dissociate into ions but silver chloride remains as such. The complete ionic equation is:

$$Fe^{2+}(aq) + 2Cl^{-}(aq) + 2Ag^{+}(aq) + 2NO_3{}^{-}(aq)$$
$$\rightarrow Fe^{2+}(aq) + 2NO_3{}^{-}(aq) + 2AgCl(s)$$

7. Let us assume that potassium chloride and sodium acetate undergoes the double-replacement reaction. By switching the cations (or anions), the products obtained are potassium acetate and sodium chloride. From the solubility rules we find that all compounds of potassium and sodium are soluble in water. Hence, no precipitation reaction occurs. Thus, the reaction can be represented as:

$$\mathrm{KCl(aq) + NaC_2H_3O_2(aq) \rightarrow NaCl(aq) + KC_2H_3O_2(aq)}$$

All compounds in the above chemical equation are soluble in water and dissociate into their respective ions. The complete ionic equation is:

$$\mathrm{K^+(aq) + Cl^-(aq) + Na^+(aq) + C_2H_3O_2{}^-(aq)}$$
$$\mathrm{\rightarrow Na^+(aq) + Cl^-(aq) + K^+(aq) + C_2H_3O_2{}^-(aq)}$$

9. Iron(II) chloride and silver nitrate undergo double-replacement reaction. By switching the cations (or anions), the products obtained are iron(II) nitrate and silver chloride. From the solubility rules we find that iron(II) nitrate is soluble but, silver chloride is insoluble. Hence, the precipitation reaction occurs. The balanced reaction is

$$\mathrm{FeCl_2(aq) + 2AgNO_3(aq) \rightarrow Fe\left(NO_3\right)_2(aq) + 2AgCl(s)}$$

In water iron(II) chloride, silver nitrate, and iron(II) nitrate dissociate into ions but silver chloride remains as such. The complete ionic equation is:

$$\mathrm{Fe^{2+}(aq) + 2Cl^-(aq) + 2Ag^+(aq) + 2NO_3{}^-(aq)}$$
$$\mathrm{\rightarrow Fe^{2+}(aq) + 2NO_3{}^-(aq) + 2AgCl(s)}$$

In this equation $\mathrm{Fe^{2+}}$ and $\mathrm{2NO3^-}$ are spectator ions. So, they are cancelled.

Hence, the net ionic equation is:

$$\mathrm{2Cl^-(aq) + 2Ag^+(aq) \rightarrow 2AgCl(s)}$$

11. Let us assume that potassium chloride and sodium acetate undergo double- replacement reaction. By switching the cations (or anions), the products obtained are potassium acetate and sodium chloride. From the solubility rules we find that all compounds of potassium and sodium are soluble in water. Hence, no precipitation reaction occurs. The reaction can be represented as

$$\mathrm{KCl(aq) + NaC_2H_3O_2(aq) \rightarrow NaCl(aq) + KC_2H_3O_2(aq)}$$

All compounds in the above chemical equation are soluble in water and dissociate into their respective ions. The complete ionic equation is:

$$\mathrm{K^+(aq) + Cl^-(aq) + Na^+(aq) + C_2H_3O_2{}^-(aq)}$$
$$\mathrm{\rightarrow Na^+(aq) + Cl^-(aq) + K^+(aq) + C_2H_3O_2{}^-(aq)}$$

In the above chemical equation, $\mathrm{K^+}$, $\mathrm{Cl^-}$, $\mathrm{Na^+}$, and $\mathrm{C2H3O2^-}$are spectator ions. So, they are cancelled.

Since all ions on both the sides of the equation are cancelled, no overall reaction occurs. This can be predicted from the solubility rules itself. According to the solubility rules, since both the predicted products in this reaction, sodium chloride and potassium acetate are soluble, a double-replacement reaction cannot occur.

13. In Exercise 9, the complete ionic equation is:

$$\mathrm{Fe^{2+}(aq) + 2Cl^-(aq) + 2Ag^+(aq) + 2NO_3{}^-(aq)}$$
$$\mathrm{\rightarrow Fe^{2+}(aq) + 2NO_3{}^-(aq) + 2AgCl(s)}$$

In this ionic equation $\mathrm{Fe^{2+}(aq)}$ and $\mathrm{NO3^-(aq)}$ ions remain the same on the reactant side as well as the product side. They do not undergo any chemical change. Hence, $\mathrm{Fe^{2+}(aq)}$ and $\mathrm{NO3^-(aq)}$ ions are the spectator ions. They are present but don't participate in the overall reaction.

In Exercise 10, the complete ionic equation is:

$$\mathrm{Ba^{2+}(aq) + 2Cl^-(aq) + 2Na^+(aq) + SO_4{}^{2-}(aq)}$$
$$\mathrm{\rightarrow BaSO_4(s) + 2Na^+(aq) + 2Cl^-(aq)}$$

In this ionic equation, $\mathrm{Na^+(aq)}$ and $\mathrm{Cl^-(aq)}$ ions remain the same on the reactant side as well as the product side. They do not undergo any chemical change. Hence, $\mathrm{Na^+(aq)}$ and $\mathrm{Cl^-(aq)}$ ions are the spectator ions. They are present but don't participate in the overall reaction.

4. COMPOSITION, DECOMPOSITION, AND COMBUSTION REACTIONS

Test Yourself

1. Single substance propane reacts to form two substances- propyne and hydrogen. This is a decomposition reaction.

Test Yourself

1. The products of the reaction are CO_2 and H_2O, so our unbalanced equation is:

   $$C_3H_6O + O_2 \rightarrow CO_2 + H_2O$$

   Balancing the above equation, we get:

   $$C_3H_6O + 4O_2 \rightarrow 3CO_2 + 3H_2O$$

Exercises

1. a. In this reaction, two different reactants react to produce two different products. This is not a composition reaction. A composition reaction produces a single substance from multiple reactants.
   b. In this reaction, two different reactants, calcium oxide and carbondioxide, react to produce a single product, calcium carbonate. This is a composition reaction.
3. a. In this reaction, two different reactants, sulfur dioxide and water, react to produce a single product, sulfur trioxide. This is a composition reaction.
   b. In this reaction, two different reactants, carbon and hydrogen, react to produce a single product, benzene. This is a composition reaction.
5. a. This is not a decomposition reaction. In this reaction, two different reactants react to produce two new compounds. It involves switching of cations (or anions) in reactants and looks like a double-replacement reaction. In a decomposition reaction, a single reactant produces multiple products.
   b. This is a decomposition reaction. In this reaction, a single reactant, calcium carbonate, produces two new products, calcium oxide and carbon dioxide.
7. a. In this reaction, two reactants combine to form a single product (composition reaction). This is not a decomposition reaction. A decomposition reaction produces multiple products from a single reactant.
   b. This is a decomposition reaction. A single reactant, H_2SO_3 produces two different products, H_2O and SO_2.
9. a. In this reaction, a compound of carbon and hydrogen combines with oxygen to form carbondioxide and water. This is a combustion reaction. Combustion reactions are the combination of some compound with oxygen to make oxides of the other elements as products.
   b. In this reaction, a compound of iron and sulfur, combines with oxygen to form oxides of iron and sulfur. This is a combustion reaction. Combustion reactions are the combination of some compound with oxygen to make oxides of the other elements as products
11. a. In this reaction, phosphorus combines with oxygen to form an oxide of phosphorus. Hence, this is a combustion reaction.
    b. In this reaction, a compound of aluminum and sulfur combines with oxygen to form oxides of aluminum and sulfur. Hence, this is a combustion reaction.
13. Yes, a composition reaction can be a combustion reaction, if the combining compounds include oxygen. For example,

    $$2H_2 + O_2 \rightarrow 2H_2O.$$

15. a. The products of the reaction are CO_2 and H_2O. So, the unbalanced reaction is:

$$C_4H_9OH + O_2 \rightarrow CO_2 + H_2O$$

On balancing the number of carbon, oxygen, and hydrogen atoms on both the sides of the above chemical equation, we get:

$$C_4H_9OH + 6O_2 \rightarrow 4CO_2 + 5H_2O$$

b. The nitrogen atoms in nitro methane will react with oxygen to make N_2; carbon atoms will react to make CO_2; and hydrogen atoms will react to make H_2O. So, the unbalanced reaction is:

$$CH_3NO_2 + O_2 \rightarrow CO_2 + H_2O + N_2$$

On balancing the reaction, we get:

$$4CH_3NO_2 + 3O_2 \rightarrow 4CO_2 + 6H_2O + 2N_2$$

5. NEUTRALIZATION REACTIONS

Test Yourself

1. First, we will write the chemical equation with the formulas of the reactants and the expected products; then we will balance the equation. The expected products are water and strontium sulfate. So, the initial chemical equation is:

$$H_2SO_4(aq) + Sr(OH)_2(aq) \rightarrow H_2O(\ell) + SrSO_4(aq)$$

We need to realize that there will be two water molecules. Then, the equation becomes:

$$H_2SO_4(aq) + Sr(OH)_2(aq) \rightarrow 2H_2O(\ell) + SrSO_4(aq)$$

This chemical equation is now balanced. It represents the neutralization reaction between sulfuric acid and strontium hydroxide.

Test Yourself

1. The products of the neutralization reaction are titanium nitrate and water. The chemical reaction is:

$$4HNO_3(aq) + Ti(OH)_4(s) \rightarrow 4H_2O(\ell) + Ti(NO_3)_4(aq)$$

The complete ionic equation is:

$$4H^+(aq) + 4\,NO_3{}^-(aq) + Ti(OH)_4(s) \rightarrow 4H_2O(\ell) + Ti^{4+}(aq) + 4\,NO_3{}^-(aq)$$

The nitrate ions are the only spectator ions here, so the net ionic equation is

$$4H^+(aq) + Ti(OH)_4(s) \rightarrow 4H_2O(\ell) + Ti^{4+}(aq)$$

Exercises

1. An acid is any compound that increases the amount of hydrogen ion (H^+) in an aqueous solution.

3. In a neutralization reaction, an *acid* and a *base* react to form *water* and a *salt*.

  a. The products are KCl (salt) and H2O.
  b. The products are K2SO4(salt) and H2O.
  c. The products are Ni3(PO4)2(salt) and H2O.

5. a. Hydrochloric acid combines with potassium hydroxide to give water and potassium chloride. The chemical reaction is:

$$HCl+KOH \rightarrow KCl + H_2O$$

This is the balanced chemical equation for this neutralization reaction.

b. Sulfuric acid combines with potassium hydroxide to give potassium sulfate and water. The chemical reaction is:

$$H_2SO_4 + KOH \rightarrow K_2SO_4 + H_2O$$

To balance the potassium atoms on the reactant side, we must take two KOH molecules. This would result in two water molecules. So, the balanced chemical equation is:

$$H_2SO_4 + 2KOH \rightarrow K_2SO_4 + 2H_2O$$

c. Phosphoric acid and nickel hydroxide combine to form nickel phosphate and water. The unbalanced chemical equation is:

$$H_3PO_4 + Ni(OH)_2 \rightarrow Ni_3(PO_4)_2 + H_2O$$

To balance the number of nickel atoms on both the sides, we must take three molecules of nickel hydroxide and to balance the number of phosphate ions on both the sides, two molecules of phosphoric acid must be taken. This will give six molecules of water. Hence, the balanced chemical equation is:

$$2H_3PO_4 + 3Ni(OH)_2 \rightarrow Ni_3(PO_4)_2 + 6H_2O$$

7. a. The products of this reaction are potassium iodide and water. The reaction can be represented as:

$$HI(aq) + KOH(aq) \rightarrow KI(aq) + H_2O(\ell)$$

The number of atoms of each kind in the reactant side and product side is balanced. Hence, the above reaction is the balanced chemical reaction.

b. The products of this reaction are barium sulfate and water. The unbalanced chemical reaction is:

$$H_2SO_4(aq) + Ba(OH)_2(aq) \rightarrow BaSO_4(s) + H_2O(\ell)$$

From the solubility rules we find that barium sulfate is insoluble. So, the label (s) is added to indicate the solid phase. The reaction yields two molecules of water. Hence, the balanced chemical reaction is:

$$H_2SO_4(aq) + Ba(OH)_2(aq) \rightarrow BaSO_4(s) + 2H_2O(\ell)$$

9. a. The products of this reaction are potassium iodide and water. The balanced chemical reaction is:

$$HI(aq) + KOH(aq) \rightarrow KI(aq) + H_2O(\ell)$$

The complete ionic equation is

$$H^+(aq) + I^-(aq) + K^+(aq) + OH^-(aq) \\ \rightarrow K^+(aq) + I^-(aq) + H_2O(\ell)$$

Here K^+ and I^- ions are the spectator ions and are cancelled. Hence, the net ionic equation is:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(\ell)$$

b. The products of this reaction are barium sulfate and water. The balanced chemical equation is:

$$H_2SO_4(aq) + Ba(OH)_2(aq) \rightarrow BaSO_4(s) + 2H_2O(\ell)$$

From the solubility rules we find that barium sulfate is insoluble. So, it can not dissociate into ions. Hence, the complete ionic equation is:

$$2H^+(aq) + SO_4{}^{2-}(aq) + Ba^{2+}(aq) + 2OH^-(aq) \\ \rightarrow BaSO_4(s) + 2H_2O(\ell)$$

Since there are no common ions (spectator ions) on the reactant and the product side, the equation above is the net ionic equation.

11. The products of this reaction are zinc chlorate and water. The balanced chemical reaction is:

$$2HClO_3(aq) + Zn(OH)_2(aq) \rightarrow Zn(ClO_3)_2(aq) + 2H_2O(\ell)$$

Separating the ionic compounds into their ions, the complete ionic equation is:

$$2H^+(aq) + 2ClO_3{}^-(aq) + Zn^{2+}(aq) + 2OH^-(aq) \\ \rightarrow Zn^{2+}(aq) + 2ClO_3{}^-(aq) + 2H_2O(\ell)$$

Since zinc and chlorate ions are spectator ions, they are cancelled. The net ionic equation is:

$$2H^+(aq) + 2\,OH^-(aq) \rightarrow 2H_2O(\ell)$$

13. Neutralization reaction between HCl(aq) and KOH(aq): HCl and KOH react to form KCl and H_2O. The balanced chemical equation is:

$$HCl\,(aq) + KOH\,(aq) \rightarrow KCl\,(aq) + H_2O\,(\ell)$$

The complete ionic equation is:

$$H^+(aq) + Cl^-(aq) + K^+(aq) + OH^-(aq) \rightarrow K^+(aq) + Cl^-(aq) + H_2O(\ell)$$

Since potassium and chloride ions are spectator ions, they are cancelled. The net ionic equation is:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(\ell)$$

Neutralization reaction between HNO_3(aq) and RbOH(aq): The products of the reaction between HNO_3(aq) and RbOH(aq) are $RbNO_3$(aq) and H_2O. The balanced chemical reaction is:

$$HNO_3(aq) + RbOH(aq) \rightarrow H_2O(\ell) + RbNO_3(aq)$$

The complete ionic equation is:

$$H^+(aq) + NO_3^-(aq) + Rb^+(aq) + OH^-(aq) \rightarrow H_2O(\ell) + Rb^+(aq) + NO_3^-(aq)$$

Since rubidium and nitrate ions are spectator ions, they are cancelled. The net ionic equation is:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(\ell)$$

Thus, we find that the net ionic equation for the neutralization reaction between HCl(aq) and KOH(aq) is the same as the net ionic equation for the neutralization reaction between HNO_3(aq) and RbOH(aq). The reason is that the salts formed after neutralization reaction, in both the cases, are soluble.

15. HCl and KOH react to form KCl and water. The balanced chemical equation is:

$$HCl\,(aq) + KOH\,(aq) \rightarrow KCl\,(aq) + H_2O\,(\ell)$$

The complete ionic equation is:

$$H^+(aq) + Cl^-(aq) + K^+(aq) + OH^-(aq) \rightarrow K^+(aq) + Cl^-(aq) + H_2O(\ell)$$

Using H_3O^+ ion in the place of H^+, the complete ionic equation becomes:

$$H_3O^+(aq) + Cl^-(aq) + K^+(aq) + OH^-(aq) \rightarrow K^+(aq) + Cl^-(aq) + 2H_2O(\ell)$$

On using hydronium ion, just an extra water molecule is added. In this ionic equation potassium and chloride ions are spectator ions; and are cancelled. The net ionic equation is:

$$H_3O^+(aq) + OH^-(aq) \rightarrow 2H_2O(\ell)$$

On using H^+, we get one molecule of water. Whereas, on using H_3O^+ we get two molecules of water.

6. OXIDATION-REDUCTION REACTIONS

Test Yourself

1. According to rule 3, the oxidation number of oxygen is −2 and that of hydrogen is +1. According to rule 4, the sum of the oxidation number on all the atoms must equal the charge on the species, so we have the simple algebraic equation

   $3 \times (+1) + x + 4 \times (-2) = 0$

   Here, x represents the oxidation number of phosphorus; +1 is for hydrogen, and −2 for oxygen. There are three hydrogen and four oxygen atoms. Evaluating the above equation, we get:

   $$3 + x - 8 = 0$$
   $$\text{or,}\quad x = 5$$

   Hence, the oxidation number of phosphorus in H3PO4 is +5, oxidation number of hydrogen is +1, and that of oxygen is −2.

Test Yourself

1. According to rule 1, atoms in their elemental state are assigned an oxidation number of 0. Because both the reactants are the elemental forms of their atoms, the C and O atoms as reactants have oxidation numbers of zero. In the product, the oxygen atoms have an oxidation number of −2, and carbon atoms have an oxidation number of +4.

   $$\begin{array}{lccc} & C + & O_2 \rightarrow & CO_2 \\ \text{ox\#} & 0 & 0 & +4\ -2 \end{array}$$

   We note that carbon is increasing its oxidation number from 0 to +4, so it is being oxidized; and oxygen is decreasing its oxidation number from 0 to −2, so it is being reduced. Since oxidation numbers are changing, this is a redox reaction. Note that the total number of electrons lost by the carbon atom (four), is gained by the oxygen atoms (four, two gained for each atom).

   C is being oxidized from 0 to +4; O is being reduced from 0 to −2.

Exercises

1. Since both the reactants are in their elemental form, potassium and bromine atoms have an oxidation number zero. In the product, potassium bromide, the potassium atom has an oxidation number +1; and the bromine atom has an oxidation number −1.

   $$\begin{array}{lccc} & 2K(s)\ + & Br_2(\ell) \rightarrow & 2KBr(s) \\ \text{ox\#} & 0 & 0 & +1\ -1 \end{array}$$

   We find that the oxidation numbers are changing. The oxidation number of potassium increases from 0 from 1, so it is oxidized; and oxidation number of bromine decreases from 0 to −1, so it is reduced. Since oxidation and reduction takes place in this reaction, this is a redox reaction. We note that the two electrons lost from the two potassium atoms, is gained by the two bromine atoms.

3. Since both the reactants are in their elemental form, calcium and oxygen atoms have an oxidation number zero. In calcium oxide, the calcium atom has an oxidation number +2 and the oxygen atom has an oxidation number −2.

   $$\begin{array}{lccc} & 2Ca(s)\ + & O_2(g) \rightarrow & 2CaO \\ \text{ox\#} & 0 & 0 & +2\ -2 \end{array}$$

   We find that the oxidation number of calcium increases from 0 to +2 and that of oxygen decreases from 0 to −2. So, calcium is oxidized and loses electrons, whereas, oxygen is reduced and gains electrons. Each calcium atom has lost two electrons, which in turn are gained by each oxygen atom.

5. Since both the reactants are in their elemental form, lithium and oxygen atoms have an oxidation number zero. In lithium peroxide, lithium atom has an oxidation number +1 and the oxygen atom has an oxidation number −1.

$$\begin{array}{lllll} & 2\text{Li(s)} & + \; O_2(g) & \rightarrow & Li_2O_2(s) \\ \text{ox\#} & 0 & 0 & & +1 \; -1 \end{array}$$

We find that the oxidation number of lithium increases from 0 to +1 and that of oxygen decreases from 0 to −1. So, lithium is oxidized and loses electrons, whereas, oxygen is reduced and gains electrons.

7. loss of electrons; increase in oxidation number

9. a. Since phosphorus is in its elemental form, its oxidation number is zero.

b. According to rule 3, oxygen is usually assigned an oxidation number −2. According to rule 4, the sum of the oxidation numbers of sulfur and oxygen must be zero (since SO_2 is neutral). If the oxidation number of sulfur is *x*, then:

$$x + 2(-2) = 0 \text{ or } x = 4$$

Thus, the oxidation number of sulfur is +4.

c. According to rule 3, oxygen is usually assigned an oxidation number −2. According to rule 4, the sum of the oxidation numbers of sulfur and oxygen must be −2. If the oxidation number of sulfur is *x*, then,

$$x + 2(-2) = -2 \text{ or } x = -2 + 4 = 2$$

Thus, the oxidation number of sulfur is +2.

d. The compound calcium nitrate can be separated into two parts: Ca^{2+} and $NO_3{}^-$. Considering these separately, the Ca atom has an oxidation number of +2, by rule 2. Now consider the $NO3^-$ ion. Oxygen is assigned an oxidation number of −2, and there are three of them. According to rule 4, the sum of the oxidation numbers on all the atoms must equal the charge on the species, so we have the simple algebraic equation:

$$x + 3(-2) = -1$$

where, *x* is the oxidation number of the nitrogen atom and −1 represents the charge on the species.

$$x + (-6) = -1 \text{ or } x = +5$$

Thus, the oxidation number of nitrogen is +5.

11. a. Oxygen is usually assigned an oxidation number −2. Let the oxidation number of carbon be *x*. According to rule 4,

$$x + 1(-2) = 0$$
$$x = 2$$

Hence, the oxidation number of carbon is +2.

b. Oxygen is usually assigned an oxidation number −2. Let the oxidation number of carbon be *x*. According to rule 4,

$$x + 2(-2) = 0$$
$$x = 4$$

Hence, the oxidation number of carbon is +4.

c. Chlorine is usually assigned an oxidation number −1. Let the oxidation number of nickel be *x*. According to rule 4,

$$x + 2(-1) = 0$$
$$x = 2$$

Hence, the oxidation number of nickel is +2.

d. Chlorine is usually assigned an oxidation number −1. Let the oxidation number of nickel be *x*. According to rule 4,

$$x + 3(-1) = 0$$
$$x = 3$$

Hence, the oxidation number of nickel is +3.

13. a. Oxygen is usually assigned an oxidation number −2. Hydrogen is assigned an oxidation number +1. Let the oxidation number of carbon be *x*. According to rule 4,

$$x + 2(+1) + 1(-2) = 0$$
$$x = 0$$

Hence, the oxidation number of carbon is 0.

b. Hydrogen is usually assigned an oxidation number +1. Hence, the oxidation number of nitrogen is −3 (rule 4).

c. This compound can be seperated into two parts. Rb^{1+} and $SO_4{}^{2-}$. According to rule 2, the oxidation number of rubidium is +1 (since charge on the rubidium ion is 1+). Now let us consider the sulfate ion. Oxygen has an oxidation number −2. Let the oxidation number of sulfur be *x*. According to rule 4,

$$x + 4(-2) = -2$$
$$x = 6$$

Hence, the oxidation number sulfur is +6.

d. This compound can be separated into two parts: Zn^{2+} and $C_2H_3O_2{}^{-}$. According to rule 2, the oxidation number of zinc is +2. Considering the acetate ion, oxygen has an oxidation number −2 and hydrogen has an oxidation number +1. Let the oxidation number of carbon be *x*. According to rule 4,

$$2x + 3(+1) + 2(-2) = -1$$
$$x = 0$$

Hence, the oxidation number of carbon is 0.

15. In the reactants, chlorine has an oxidation number zero (since it is in its elemental form). In nitrogen monoxide, oxygen has an oxidation number −2 (rule 3). So, oxidation number of nitrogen is +2 (rule 4). In the product, oxygen will retain the oxidation number −2, and chlorine will be assigned an oxidation number −1(rule 3). Let oxidation number of nitrogen be *x*. According to rule 4,

$$x - 2 - 1 = 0$$
$$x = 3$$

Hence, the oxidation number of the nitrogen atom is +3.

$$2NO + Cl_2 \rightarrow 2NOCl$$

ox# +2 −2 0 +3 −2 −1

We find that the oxidation number of nitrogen increases from +2 to +3; so it is oxidized. The oxidation number of chlorine decreases from 0 to −1; so it is reduced.

17. First, let us consider the reactants. In krypton fluoride, fluorine has an oxidation number −1(rule 3). So, krypton has an oxidation number +2 (rule 4). In water, hydrogen has an oxidation number +1 and oxygen has an oxidation number −2. In the products, Kr and oxygen are in their elemental form. So, the oxidation numbers of krypton and oxygen is zero. In hydrogen fluoride, fluorine has an oxidation number −1 and hydrogen has an oxidation number +1.

$$2KrF_2 + 2H_2O \rightarrow 2Kr + 4HF + O_2$$

ox# +2 −1 +1 −2 0 +1 −1 0

We find that, during the course of the reaction, the oxidation number of oxygen increases from −2 to 0, so it is oxidized; whereas oxidation number of krypton is decreases from +2 to 0, so it is reduced.

19. First let us consider the reactants. Potassium is in the elemental form. So, its oxidation number is zero. Magnesium chloride is made of one Mg^{2+} and two Cl^- ions. So, the oxidation number of magnesium is +2 and that of chlorine is −1. Now, let us consider the products. Potassium chloride is composed of a potassium ion, K^+ and a chloride ion, Cl^-. So, the oxidation number of potassium is +1 and that of chlorine is −1. Since magnesium is in its elemental form, its oxidation number is zero.

$$2K + MgCl_2 \rightarrow 2KCl + Mg$$

ox# 0 +2 −1 +1 −1 0

During the course of the reaction, the oxidation number of potassium increases from 0 to +1, so it is oxidized. The oxidation number of magnesium decreases from +2 to 0, so it is reduced.

7. END-OF-CHAPTER MATERIAL

Additional Exercises

1. On boiling, water doesn't undergo any chemical change. It only changes its phase from liquid to gas.

$$H_2O(\ell) \rightarrow H_2O(g)$$

3. The coefficients are not in their lowest whole number ratio.

5. No; zinc is lower in the activity series than aluminum.

7. In the products, the cation is pairing with the cation, and the anion is pairing with the anion.

9. On swapping the cations(or anions) of the reactants, we obtain the products, barium sulfate and silver chloride. From the solubility rules, we find that both the products are insoluble. So, the balanced chemical equation is:

$$BaCl_2(aq) + Ag_2SO_4(aq) \rightarrow BaSO_4(s) + 2AgCl(s)$$

The insoluble products are indicated by the label (s) for the solid phase. They cannot be separated into ions. The complete ionic equation is:

$$Ba^{2+}(aq) + 2Cl^-(aq) + 2Ag^+(aq) + SO_4{}^{2-}(aq) \rightarrow BaSO_4(s) + 2AgCl(s)$$

There are no spectator ions. Therefore, the net ionic equation is the same as the complete ionic equation.

11. This is a balanced chemical reaction and all the reactants as well as products are soluble. So, they can disassociate into their respective ions. The complete ionic equation is:

$$Na^+(aq) + Cl^-(aq) + K^+(aq) + Br^-(aq) \rightarrow Na^+(aq) + Br^-(aq) + K^+(aq) + Cl^-(aq)$$

Since all the ions are spectator ions, there is no net ionic equation.

13. Yes. The reaction between hydrogen and chlorine to form a single product, hydrogen chloride, is an example.

$$H_2 + Cl_2 \rightarrow 2HCl$$

This is a composition reaction as two reactants react to give a single product.

In this reaction, since the reactants are in the elemental form, the oxidation numbers of hydrogen and chlorine are zeroes, respectively. In the product hydrogen chloride, hydrogen is assigned an oxidation number +1 and the oxidation number of chlorine is −1. The oxidation number of hydrogen increases from 0 to +1, so it is oxidized. The oxidation number of chlorine decreases from 0 to −1, so it is reduced. Hence, the above reaction is a redox reaction also.

15. Yes, decomposition reaction of hydrogen chloride into hydrogen and chlorine is a redox reaction.

$$2HCl \rightarrow H_2 + Cl_2$$

Since multiple products are obtained from a single reactant, this is a decomposition reaction.

In hydrogen chloride, the oxidation number of hydrogen is +1 and that of chlorine is −1. Considering the products, since they are in their elemental form, oxidation numbers of hydrogen and chlorine are zeroes, respectively. Thus, in the course of the reaction, the oxidation number of hydrogen decreases from +1 to 0, so it is reduced. The oxidation number of chlorine increases from −1 to 0, so it is oxidized. Hence, the above reaction is a redox reaction.

17. It does not increase the H^+ ion concentration; it is not a compound of H^+.

19. a. Since it is the elemental form of nitrogen, the oxidation number of nitrogen is zero.

b. Hydrogen is assigned +1 oxidation number. So, the oxidation number of nitrogen is −3.

c. Oxygen is assigned an oxidation number −2. So, the oxidation number of nitrogen is +2.

d. Oxygen is assigned an oxidation number −2. So, the sum of the oxidation numbers of two nitrogen atoms must be +2. Hence, the oxidation number of nitrogen is +1.

e. Oxygen is assigned an oxidation number −2. So, the sum of the oxidation numbers of two nitrogen atoms must be +4 (equal to the sum of the oxidation numbers of two oxygen atoms). Hence, the oxidation number of nitrogen is +4.

f. Oxygen is assigned an oxidation number −2. So, the sum of the oxidation numbers of two nitrogen atoms must be +8 (equal to the sum of the oxidation numbers of four oxygen atoms). Hence, the oxidation number of nitrogen is +4.

g. Oxygen is assigned an oxidation number −2. So, the sum of the oxidation numbers of two nitrogen atoms must be +10 (equal to the sum of the oxidation numbers of five oxygen atoms).

Hence,

$$2x + 5(-2) = 0$$
$$x = 5$$

The oxidation number of nitrogen is +5.

h. The compound sodium nitrate comprises of two ionic parts: Na^+ and $NO_3{}^-$. The oxidation number of sodium is the same as the charge on it. So, its oxidation number is +1. In nitrate ion, the oxidation number of oxygen is −2. The sum of the oxidation numbers of one nitrogen atom and three oxygen atoms must be equal to −1, the charge on the ion.

$$x + 3(-2) = -1$$
$$x = 5$$

The oxidation number of nitrogen is +5.

21. First, let us consider the reactant, copper chloride. It consists of two ions: Cu^+ and Cl^-. The oxidation number of chlorine atom is −1 and that of copper atom is +1. Considering the products, copper in the elemental form has an oxidation number zero. In copper chloride (made of two ions: Cu^{2+} and Cl^-), the oxidation number of chlorine is −1 and that of copper is +2.

$$\begin{array}{lccc} & 2CuCl(aq) & \rightarrow\ CuCl_2(aq) & +\ Cu(s) \\ \text{ox\#} & +1\ \ -1 & +2\ \ -1 & 0 \end{array}$$

During the course of the reaction, the oxidation number of copper increases from +1 to +2, so it is oxidized. Also, the oxidation number of copper is decreases from +1 to 0, so it is reduced. The same element copper is being reduced and oxidized. Hence, in this reaction, copper is the disproportionating element.

CHAPTER 5

Stoichiometry and the Mole

1. STOICHIOMETRY

Test Yourself

1. We use the balanced chemical equation to construct a conversion factor between H_2 and N_2. The number of molecules of N_2 goes on the bottom of our conversion factor so it cancels with our given amount, and the molecules of H_2 go on the top. Thus, the appropriate conversion factor is $\frac{3\text{ molecules }H_2}{1\text{ molecule }N_2}$. Starting with our given amount and applying the conversion factor, the result is

$$29\text{ molecules }N_2 \times \frac{3\text{ molecules }H_2}{1\text{ molecule }N_2} = 87\text{ molecules }H_2$$

We need 87 molecules of H_2 to react with 29 molecules of N_2.

Test Yourself

1. From the formula, we know that 1 molecule of $Fe_2(SO_4)_3$ has 3 atoms of S. Using this fact as the conversion factor:

$$777\text{ atoms S} \times \frac{1\text{ molecule }Fe_2(SO_4)_3}{3\text{ atoms S}} = 259\text{ molecules }Fe_2(SO_4)_3$$

We can make 259 molecules of $Fe_2(SO_4)_3$ from 777 atoms of S.

Exercises

1. The conversion factors that can be constructed from the pound cake recipe are:

$$\frac{1\text{ pound butter}}{1\text{ pound flour}} \text{ or } \frac{1\text{ pound butter}}{1\text{ pound eggs}} \text{ or } \frac{1\text{ pound butter}}{1\text{ pound sugar}} \text{ or }$$

$$\frac{1\text{ pound sugar}}{1\text{ pound eggs}} \text{ or } \frac{1\text{ pound sugar}}{1\text{ pound flour}} \text{ or } \frac{1\text{ pound flour}}{1\text{ pound eggs}}$$

3. $\frac{2\text{ molecules }H_2}{1\text{ molecule }O_2}$, $\frac{1\text{ molecule }O_2}{2\text{ molecules }H_2O}$, $\frac{2\text{ molecules }H_2}{2\text{ molecules }H_2O}$, and their reciprocals are the conversion factors that can be constructed.

5. a. The given equation would be balanced if we have ½ H_2 molecules. To have the whole number coefficient for H_2, we can multiply the coefficients of other molecules (reactants and products) by 2. Then the balanced chemical equation is

$$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$$

   b. The conversion factors that can be constructed from this balanced equation are:

$$\frac{2\text{ molecules }H_2O}{2\text{ atoms Na}} \text{ or } \frac{2\text{ molecules NaOH}}{2\text{ atoms Na}} \text{ or } \frac{1\text{ molecule }H_2}{2\text{ atoms Na}}$$

Several other conversion factors are possible.

The number of molecules of Na goes on the bottom of our conversion factor so that it cancels with our given amount, and the molecules of H_2 go on the top. Thus, the appropriate conversion factor is $\frac{1\text{ molecule }H_2}{2\text{ atoms Na}}$. Starting with our given amount and applying the conversion factor, the result is

$$332\text{ atoms Na} \times \frac{1\text{ molecule }H_2}{2\text{ atoms Na}} = 166\text{ molecules }H_2$$

Thus, 166 molecules of H_2 are produced when 332 atoms of Na react.

7. We need to calculate the number of molecules of H_2O produced when 75 molecules of H_2O_2 react. The number of molecules of H_2O_2 goes on the bottom of our conversion factor so that it cancels with our given amount, and the molecules of H_2O go on the top. The appropriate conversion factor is $\frac{\text{8 molecules } H_2O}{\text{5 molecules } H_2O_2}$. Starting with our given amount and applying the conversion factor, the result is

$$75 \text{ molecules } H_2O_2 \times \frac{\text{8 molecules } H_2O}{\text{5 molecules } H_2O_2} = 120 \text{ molecules } H_2O$$

Thus, 120 molecules of H_2O are produced when 75 molecules of H_2O_2 react.

9. We need to calculate the number of molecules of $Fe_2(SO_4)_3$ produced when 321 atoms of S react. The number of atoms of S goes on the bottom of our conversion factor so it cancels with our given amount, and the molecules of $Fe_2(SO_4)_3$ go on the top. The balanced chemical equation has SO_3 and not S. We know that one molecule of SO_3 is equivalent to 1 atom of S. Hence, the appropriate conversion factor is $\frac{\text{1 molecule } Fe_2(SO_4)_3}{\text{3 molecules } SO_3} \times \frac{\text{1 molecule } SO_3}{\text{1 atom S}}$. In this conversion factor the molecules of SO_3 cancel and the required units are left in it. Starting with our given amount and applying the conversion factor, the result is

$$321 \text{ atoms S} \times \frac{\text{1 molecule } Fe_2(SO_4)_3}{\text{3 molecules } SO_3} \times \frac{\text{1 molecules } SO_3}{\text{1 atom S}}$$
$$= 107 \text{ molecules } Fe_2(SO_4)_3$$

Thus, 107 molecules of $Fe_2(SO_4)_3$ are produced if 321 atoms of S are reacted.

11. We need to calculate the number of molecules of SO_3 needed to produce 145,000 molecules of $Fe_2(SO_4)_3$. The number of molecules of $Fe_2(SO_4)_3$ goes on the bottom of our conversion factor so it cancels with our given amount, and the molecules of SO_3 go on the top. The appropriate conversion factor is $\frac{\text{3 molecules } SO_3}{\text{1 molecule } Fe_2(SO_4)_3}$. Starting with our given amount and applying the conversion factor, the result is

$$145{,}000 \text{ molecules } Fe_2(SO_4)_3 \times \frac{\text{3 molecules } SO_3}{\text{1 molecule } Fe_2(SO_4)_3}$$
$$= 435{,}000 \text{ molecules } SO_3$$

Hence, we need 435,000 molecules of SO_3 to produce 145,000 molecules of $Fe_2(SO_4)_3$.

13. a. The independent conversion factors possible for the reaction $2H_2 + O_2 \rightarrow 2H_2O$ are:

$$\frac{\text{2 molecules } H_2}{\text{1 molecule } O_2}, \frac{\text{1 molecule } O_2}{\text{2 molecules } H_2O}, \text{ and } \frac{\text{2 molecules } H_2}{\text{2 molecules } H_2O}$$

b. The independent conversion factors possible for the reaction $H_2 + O_2 \rightarrow H_2O_2$ are:

$$\frac{\text{1 molecule } H_2}{\text{1 molecule } O_2}, \frac{\text{1 molecule } O_2}{\text{1 molecule } H_2O_2}, \text{ and } \frac{\text{1 molecule } H_2}{\text{1 molecule } H_2O_2}$$

The conversion factors are different because the stoichiometries of the balanced chemical reactions are different.

2. THE MOLE

Test Yourself

1. The definition of a mole is an equality that can be used to construct a conversion factor.

$$4.61 \times 10^{-2}\text{mol } O_2 \times \frac{6.022 \times 10^{23}\text{molecules } O_2}{1 \text{ mol } O_2} = 2.78 \times 10^{22}\text{molecules } O_2$$

Hence, 2.78×10^{22} molecules are present in 4.61×10^{-2} mol of O_2.

Test Yourself

1. To determine the molar mass, we simply add the atomic masses of the atoms in the molecular formula but express the total in grams per mole, not atomic mass units.

1 Ag	107.868
1 N	14.007
3 O = 3 × 15.999	47.997
Total	169.87 g/mol

The molar mass of $AgNO_3$ is 169.87 g/mol.

Test Yourself

1. Use the molar mass as a conversion factor between moles and grams. Because we want to cancel the mole unit and introduce the gram unit, we can use the molar mass as given:

$$33.7 \text{ mol } H_2O \times \frac{18 \text{ g } H_2O}{1 \text{ mol } H_2O} = 607 \text{ g } H_2O$$

The mass of 33.7 mol H_2O is 607 g.

Test Yourself

1. Use molar mass of H_2SO_4 as a conversion factor from mass to moles. The molar mass of H_2SO_4 is 98.08 g/mol. However, because we want to cancel the gram unit and introduce moles, we need to take the reciprocal of this quantity, or 1 mol/98.08 g:

$$35.6 \text{ g } H_2SO_4 \times \frac{1 \text{ mol } H_2SO_4}{98.08 \text{ g } H_2SO_4} = 0.363 \text{ mol } H_2SO_4$$

Test Yourself

1. Here we use density to convert from volume to mass and then use the molar mass to determine the number of moles. The molar mass of benzene is 78.114 g/mol.

$$17.9 \text{ mL } C_6H_6 \times \frac{0.879 \text{ g } C_6H_6}{1 \text{ mL } C_6H_6} \times \frac{1 \text{ mol } C_6H_6}{78.114 \text{ g } C_6H_6} = 0.201 \text{ mol } C_6H_6$$

Exercises

1. The definition of a mole is an equality that can be used to construct a conversion factor. The number of atoms present in 4.55 mol of Fe is

$$4.55 \text{ mol Fe} \times \frac{6.022 \times 10^{23} \text{ atoms Fe}}{1 \text{ mol Fe}} = 2.74 \times 10^{24} \text{ atoms Fe}$$

3. The definition of a mole is an equality that can be used to construct a conversion factor. The number of molecules present in 2.509 mol of H_2S is

$$2.509 \text{ mol } H_2S \times \frac{6.022 \times 10^{23} \text{ molecules } H_2S}{1 \text{ mol } H_2S}$$

$$= 1.511 \times 10^{24} \text{ molecules } H_2S$$

5. The definition of a mole is an equality that can be used to construct a conversion factor. The number of moles present in 3.55×10^{24} Pb atoms is

$$3.55 \times 10^{24} \text{ atoms Pb} \times \frac{1 \text{ mol Pb}}{6.022 \times 10^{23} \text{ atoms Pb}} = 5.90 \text{ mol Pb}$$

7. The definition of a mole is an equality that can be used to construct a conversion factor. The number of moles present in 1.00×10^{23} PF_3 molecules is

$$1.00 \times 10^{23} \text{ molecules } PF_3 \times \frac{1 \text{ mol } PF_3}{6.022 \times 10^{23} \text{ molecules } PF_3}$$

$$= 0.166 \text{ mol } PF_3$$

9. Molar mass is the mass of 1 mol of atoms (or molecules) of an element (or a compound).
   a. The molar mass of Si is 28.086 g/mol.
   b. The molar mass of SiH_4 can be calculated as follows:

1 Si	= 28.086
4 H = 4 × 1.008	= 4.032
Total	= 32.118 g/mol = molar mass of SiH_4

   c. The molar mass of K_2O can be calculated as follows:

2 K = 2 × 39.098	= 78.196
1 O	= 15.999
Total	= 94.195 g/mol = molar mass of K_2O

11. Molar mass is the mass of 1 mol of atoms (or molecules) of an element (or a compound).
   a. The molar mass of Al is 26.981 g/mol.
   b. The molar mass of Al_2O_3 can be calculated as follows:

2 Al = 2 × 26.981	= 53.962
3 O = 3 × 15.999	= 47.997
Total	= 101.959 g/mol = molar mass of Al_2O_3

   c. The molar mass of $CoCl_3$ can be calculated as follows:

1 Co	= 58.933
3 Cl = 3 × 35.453	= 106.359
Total	= 165.292 g/mol = molar mass of $CoCl_3$

13. Molar mass of Rb is 85.468 g/mol. The mass of 4.44 mol of Rb is

$$4.44 \text{ mol Rb} \times \frac{85.468 \text{ g Rb}}{1 \text{ mol Rb}} = 379 \text{ g Rb}$$

15. Molar mass of aluminum sulfate can be calculated as follows:

2 Al = 2 × 26.981	= 53.962
3 S = 3 × 32.065	= 96.195
12 O = 12 × 15.999	= 191.988
Total	= 342.145 g/mol = molar mass of $Al_2(SO_4)_3$

The mass of one mole of aluminum sulfate is 342.145 g. The mass of 12.34 mol of $Al_2(SO_4)_3$ is

$$12.34 \text{ mol } Al_2(SO_4)_3 \times \frac{342.145 \text{ g } Al_2(SO_4)_3}{1 \text{ mol } Al_2(SO_4)_3}$$

$$= 4{,}222 \text{ g } Al_2(SO_4)_3$$

17. Molar mass of carbon monoxide can be calculated as follows:

1 C	= 12.011
1 O	= 15.999
Total	= 28.01 g/mol = molar mass of CO

The molar mass of CO is 28.01 g/mol. The number of moles present in 45.6 g of CO is

$$45.6 \text{ g CO} \times \frac{1 \text{ mol CO}}{28.01 \text{ g CO}} = 1.63 \text{ mol CO}$$

19. Molar mass of sulfur hexafluoride can be calculated as follows:

1 S	= 32.065
6 F = 6 × 18.998	= 113.988
Total	= 146.053 g/mol = molar mass of SF_6

The molar mass of SF_6 is 146.053 g/mol. The number of moles present in 1.223 g of SF_6 is

$$1.223 \text{ g } SF_6 \times \frac{1 \text{ mol } SF_6}{146.053 \text{ g } SF_6} = 0.008374 \text{ mol } SF_6$$

21. Molar mass of mercury is 200.59 g/mol. Using density as the conversion factor, we get

$$54.8 \text{ mL Hg} \times \frac{13.6 \text{ g Hg}}{1 \text{ mL Hg}} \times \frac{1 \text{ mol Hg}}{200.59 \text{ g Hg}} = 3.72 \text{ mol Hg}$$

3. THE MOLE IN CHEMICAL REACTIONS

Test Yourself

1. The coefficients represent the number of moles that react, not just molecules. We would speak of this equation as “one mole of elemental nitrogen reacts with three moles of elemental hydrogen to produce two moles of ammonia.”

Test Yourself

1. According to the balanced chemical equation, the equivalence between water and oxygen is

   $13\text{ mol }O_2 \Leftrightarrow 10\text{ mol }H_2O$

   We can use this equivalence to construct the proper conversion factor. We start with what we are given and apply the conversion factor:

   $$154\text{ mol }O_2 \times \frac{10\text{ mol }H_2O}{13\text{ mol }O_2} = 118\text{ mol }H_2O$$

Exercises

1. This chemical equation means: “one mole of methane reacts with two moles of elemental oxygen to produce one mole of carbondioxide and two moles of water.”
3. This chemical equation involves 6.022×10^{23} molecules of CH_4 reacting with 1.2044×10^{24} (= $2 \times 6.022 \times 10^{23}$) molecules of O_2 to produce 6.022×10^{23} molecules of CO_2 and 1.2044×10^{24} (= $2 \times 6.022 \times 10^{23}$) molecules of H_2O.
5. According to the balanced chemical equation, the equivalences are:

   $2\text{ mol }C_2H_6 \Leftrightarrow 7\text{ mol }O_2 \Leftrightarrow 4\text{ mol }CO_2 \Leftrightarrow 6\text{ mol }H_2O$

7. The balanced chemical reaction is

   $C_5H_{12} + 8O_2 \rightarrow 5CO_2 + 6H_2O$

   To determine the moles of H_2O, the moles of O_2 should go to the bottom so that it cancels. Using the appropriate conversion factor and starting with the given amount, we have:

   $$5.8\text{ mol }O_2 \times \frac{6\text{ mol }H_2O}{8\text{ mol }O_2} = 4.4\text{ mol }H_2O$$

9. To determine the moles of Cu^{2+}, the moles of H^+ must go to the bottom in the conversion factor, so that it cancels. Using the appropriate conversion factor and starting with the given amount, we have:

   $$55.7\text{ mol }H^+ \times \frac{3\text{ mol }Cu^{2+}}{8\text{ mol }H^+} = 20.9\text{ mol }Cu^{2+}$$

11. To determine the moles of H_2O, the moles of NH_3 must go to the bottom in the conversion factor, so that it cancels. Using the appropriate conversion factor and starting with the given amount, we have:

    $$0.669\text{ mol }NH_3 \times \frac{6\text{ mol }H_2O}{4\text{ mol }NH_3} = 1.00\text{ mol }H_2O$$

13. The moles of KO_2 must go to the bottom in the conversion factor, so that it cancels. Using the appropriate conversion factor and starting with the given amount, the number of moles of potassium carbonate formed can be calculated as follows:

    $$6.88\text{ mol }KO_2 \times \frac{2\text{ mol }K_2CO_3}{4\text{ mol }KO_2} = 3.44\text{ mol }K_2CO_3$$

    The number of moles of oxygen gas liberated is

    $$6.88\text{ mol }KO_2 \times \frac{3\text{ mol }O_2}{4\text{ mol }KO_2} = 5.16\text{ mol }O_2$$

4. MOLE-MASS AND MASS-MASS CALCULATIONS

Test Yourself

1. We will do this in two steps: Convert the mass of H_2O to moles and then use the balanced chemical equation to find the number of moles of Al_2O_3 formed. The molar mass of H_2O is 18.00 g/mol, which we have to invert to get the appropriate conversion factor:

$$23.9\text{ g }H_2O \times \frac{1\text{ mol }H_2O}{18.00\text{ g }H_2O} = 1.33\text{ mol }H_2O$$

We can use this quantity to determine the number of moles of Al_2O_3 that will form. From the balanced chemical equation, we construct a conversion factor between the number of moles of H_2O and the number of moles of Al_2O_3:

$$\frac{1\text{ mol }Al_2O_3}{3\text{ mol }H_2O}$$

Applying this conversion factor to the quantity of H_2O, we get

$$1.33\text{ mol }H_2O \times \frac{1\text{ mol }Al_2O_3}{3\text{ mol }H_2O} = 0.443\text{ mol }Al_2O_3$$

Alternatively, we could have done this in one line:

$$23.9\text{ g }H_2O \times \frac{1\text{ mol }H_2O}{18.00\text{ g }H_2O} \times \frac{1\text{ mol }Al_2O_3}{3\text{ mol }H_2O} = 0.443\text{ mol }Al_2O_3$$

Test Yourself

1. First, using the balanced chemical equation, we will calculate the number of moles of N_2 needed to produce 2.17 mol of NH_3.

$$2.17\text{mol }NH_3 \times \frac{1\text{ mol }N_2}{2\text{ mol }NH_3} = 1.085\text{ mol }N_2$$

The molar mass of N_2 is 28.014 g/mol. Hence, the mass of 1.085 mol N_2 in grams is

$$1.085\text{ mol }N_2 \times \frac{28.014\text{ g }N_2}{1\text{ mol }N_2} = 30.40\text{ g }N_2$$

Hence, 30.40 grams of N_2 are needed to produce 2.17 mol of NH_3.

Test Yourself

1. The molar mass of H_2 is 2.016 g/mol and that of zinc is 65.409 g/mol. We can calculate the mass of H_2 produced in one step, as follows:

$$122\text{ g Zn} \times \frac{1\text{ mol Zn}}{65.409\text{ g Zn}} \times \frac{1\text{ mol }H_2}{1\text{ mol Zn}} \times \frac{2.016\text{ g }H_2}{1\text{ mol }H_2} = 3.76\text{ g }H_2$$

Hence, 3.76 g H_2 will be produced when 122 g of Zn are reacted.

Exercises

1. The molar mass of CO_2 is 44.0 g/mol. The mass of CO_2 produced by the combustion of 1.00 mol of CH_4 can be calculated as follows:

$$1.00 \text{ mol } CH_4 \times \frac{1 \text{ mol } CO_2}{1 \text{ mol } CH_4} \times \frac{44.0 \text{ g } CO_2}{1 \text{ mol } CO_2} = 44.0 \text{ g } CO_2$$

3. The molar mass of HgO is 216.59 g/mol. The mass of HgO required to produce 0.692 mol of O_2 can be calculated as follows:

$$0.692 \text{ mol } O_2 \times \frac{2 \text{ mol HgO}}{1 \text{ mol } O_2} \times \frac{216.59 \text{ g HgO}}{1 \text{ mol HgO}}$$

$$= 3.00 \times 10^2 \text{ g HgO}$$

5. The molar mass of Ag is 107.868 g/mol. The number of moles of Al produced from 10.87 g of Ag can be determined as follows:

$$10.87 \text{ g Ag} \times \frac{1 \text{ mol Ag}}{107.868 \text{ g Ag}} \times \frac{1 \text{ mol Al}}{3 \text{ mol Ag}} = 0.0336 \text{ mol Al}$$

7. The molar mass of $Ca(NO_3)_2$ is 164.086 g/mol. The number of moles of O_2 needed to prepare 1.00 g of $Ca(NO_3)_2$ can be calculated as follows:

$$1.00 \text{ g } Ca(NO_3)_2 \times \frac{1 \text{ mol } Ca(NO_3)_2}{164.086 \text{ g } Ca(NO_3)_2} \times \frac{3 \text{ mol } O_2}{1 \text{ mol } Ca(NO_3)_2}$$

$$= 0.0183 \text{ mol } O_2$$

9. The molar mass of $NaClO_3$ is 106.44 g/mol. First, let us calculate the number of moles of O_2 generated by the decomposition of 100.0 g of $NaClO_3$.

$$100.0 \text{ g } NaClO_3 \times \frac{1 \text{ mol } NaClO_3}{106.44 \text{ g } NaClO_3} \times \frac{3 \text{ mol } O_2}{2 \text{ mol } NaClO_3}$$

$$= 1.409 \text{ mol } O_2$$

Now we will calculate O_2 in terms of its mass. The molar mass of O_2 is 31.998 g/mol. Hence, the mass of O_2 generated by the decomposition of 100.0 g of $NaClO_3$ is

$$1.409 \text{ mol } O_2 \times \frac{31.998 \text{ g } O_2}{1 \text{ mol } O_2} = 45.1 \text{ g } O_2$$

In one step, we can calculate the mass of O_2 as follows:

$$100.0 \text{ g } NaClO_3 \times \frac{1 \text{ mol } NaClO_3}{106.44 \text{ g } NaClO_3} \times \frac{3 \text{ mol } O_2}{2 \text{ mol } NaClO_3}$$

$$\times \frac{31.998 \text{ g } O_2}{1 \text{ mol } O_2} = 45.1 \text{ g } O_2$$

11. The molar mass of Fe_2O_3 is 159.687 g/mol. The molar mass of Al_2O_3 is 101.959. The mass of Fe_2O_3 that must be reacted to generate 324 g of Al_2O_3 can be calculated in one step as follows:

$$324 \text{ g } Al_2O_3 \times \frac{1 \text{ mol } Al_2O_3}{101.959 \text{ g } Al_2O_3} \times \frac{1 \text{ mol } Fe_2O_3}{1 \text{ mol } Al_2O_3}$$

$$\times \frac{159.687 \text{ g } Fe_2O_3}{1 \text{ mol } Fe_2O_3} = 507 \text{ g } Fe_2O_3$$

13. The molar mass of H_2O is 18.0 g/mol. The molar mass of MnO_2 is 86.936 g/mol. The mass of MnO_2 produced when 445 g of H_2O are reacted can be calculated in one step as follows:

$$445 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.0 \text{ g } H_2O} \times \frac{2 \text{ mol } MnO_2}{1 \text{ mol } H_2O} \times \frac{86.936 \text{ g } MnO_2}{1 \text{ mol } MnO_2}$$

$$= 4.30 \times 10^3 \text{ g } MnO_2$$

15. The molar mass of ZnO is 81.408 g/mol. The molar mass of Mn_2O_3 is 157.873 g/mol. The mass of Mn_2O_3 formed along with 83.9 g of ZnO can be calculated as follows:

$$83.9 \text{ g ZnO} \times \frac{1 \text{ mol ZnO}}{81.408 \text{ g ZnO}} \times \frac{1 \text{ mol } Mn_2O_3}{1 \text{ mol ZnO}}$$

$$\times \frac{157.873 \text{ g } Mn_2O_3}{1 \text{ mol } Mn_2O_3} = 163 \text{ g } Mn_2O_3$$

17. The molar mass of CH_2S is 46.092 g/mol. The molar mass of HF is 20.0 g/mol. The mass of HF produced when 88.4 g of CH_2S are reacted can be calculated as follows:

$$88.4 \text{ g } CH_2S \times \frac{1 \text{ mol } CH_2S}{46.092 \text{ g } CH_2S} \times \frac{2 \text{ mol HF}}{1 \text{ mol } CH_2S} \times \frac{20.0 \text{ g HF}}{1 \text{ mol HF}}$$

$$= 76.7 \text{ g HF}$$

5. YIELDS

Test Yourself

1. A mass-mass calculation can be performed to determine the theoretical yield. We need the molar masses of N_2 (28.014 g/mol) and H_2 (2.016 g/mol). In three steps, the mass-mass calculation is

$$16.5 \text{ g } N_2 \times \frac{1 \text{ mol } N_2}{28.014 \text{ g } N_2} \times \frac{2 \text{ mol } NH_3}{1 \text{mol } N_2} \times \frac{17.031 \text{ g } NH_3}{1 \text{ mol } NH_3} = 20.1 \text{ g } NH_3$$

Thus, the theoretical yield is 20.1 g of NH_3. The actual yield is the amount that was actually made, which is 2.05 g of NH_3. To calculate the percent yield, we take the actual yield and divide it by the theoretical yield and multiply by 100:

$$\frac{2.05 \text{ g } NH_3}{20.1 \text{ g } NH_3} \times 100 = 10.2\%$$

Exercises

1. Theoretical yield is what you expect stoichiometrically from a chemical reaction; actual yield is what you actually get from a chemical reaction.
3. A mass-mass calculation can be performed to determine the theoretical yield. We need the molar masses of SiF_4 (104.078 g/mol) and SiO_2 (60.084 g/mol). In three steps, the mass-mass calculation is

$$2.339 \text{ g } SiO_2 \times \frac{1 \text{ mol } SiO_2}{60.084 \text{ g } SiO_2} \times \frac{1 \text{ mol } SiF_4}{1 \text{ mol } SiO_2} \times \frac{104.078 \text{ g } SiF_4}{1 \text{ mol } SiF_4} = 4.052 \text{ g } SiF_4$$

Thus, the theoretical yield is 4.052 g of SiF_4. The actual yield is the amount that was actually made, which is 2.675 g of SiF_4.

5. A mass-mass calculation can be performed to determine the theoretical yield. We need the molar masses of $NaHCO_3$ (84.006 g/mol) and Na_2CO_3 (105.988 g/mol). In three steps, the mass-mass calculation is

$$1.006 \text{ g } NaHCO_3 \times \frac{1 \text{ mol } NaHCO_3}{84.006 \text{ g } NaHCO_3} \times \frac{1 \text{ mol } Na_2CO_3}{2 \text{ mol } NaHCO_3} \times \frac{105.988 \text{ g } Na_2CO_3}{1 \text{ mol } Na_2CO_3} = 0.635 \text{ g } Na_2CO_3$$

Thus, the theoretical yield is 0.635 g of Na_2CO_3. The actual yield is the amount that was actually made, which is 0.0334 g of Na_2CO_3.

7. The theoretical yield is 4.052 g of SiF_4. The actual yield is the amount that was actually made, which is 2.675 g of SiF_4. To calculate the percent yield, we take the actual yield and divide it by the theoretical yield and multiply by 100:

$$\frac{2.675 \text{ g } SiF_4}{4.052 \text{ g } SiF_4} \times 100 = 66.02\%$$

9. The theoretical yield is 0.635 g of Na_2CO_3. The actual yield is the amount that was actually made, which is 0.0334 g of Na_2CO_3. To calculate the percent yield, we take the actual yield and divide it by the theoretical yield and multiply by 100:

$$\frac{0.0334 \text{ g } Na_2CO_3}{0.635 \text{ g } Na_2CO_3} \times 100 = 5.26\%$$

6. LIMITING REAGENTS

Test Yourself

1. We can perform two mass-mass calculations and determine the limiting reagent. We need the molar masses of MgO (40.304 g/mol), H_2S (34.081 g/mol), and MgS (56.37 g/mol)

$$22.7\text{ g MgO} \times \frac{1\text{ mol MgO}}{40.304\text{ g MgO}} \times \frac{1\text{ mol MgS}}{1\text{ mol MgO}} \times \frac{56.37\text{ g MgS}}{1\text{ mol MgS}} = 31.75\text{ g MgS}$$

$$17.9\text{ g H}_2\text{S} \times \frac{1\text{ mol H}_2\text{S}}{34.081\text{ g H}_2\text{S}} \times \frac{1\text{ mol MgS}}{1\text{ mol H}_2\text{S}} \times \frac{56.37\text{ g MgS}}{1\text{ mol MgS}} = 29.61\text{ g MgS}$$

The 29.61 g of MgS is the lesser quantity, so the associated reactant—17.9 g of H_2S —is the limiting reagent. To determine how much of the other reactant is left, we have to do one more mass-mass calculation to determine what mass of MgO reacted with the 17.9 g of H_2S and then subtract the amount reacted from the original amount.

$$17.9\text{ g H}_2\text{S} \times \frac{1\text{ mol H}_2\text{S}}{34.081\text{ g H}_2\text{S}} \times \frac{1\text{ mol MgO}}{1\text{ mol H}_2\text{S}} \times \frac{40.304\text{ g MgO}}{1\text{mol MgO}} = 21.2\text{ g MgO reacted}$$

The amount of MgO left is: 22.7 g MgO – 21.2 g MgO reacted = 1.5 g MgO left

Exercises

1. In this box there are not enough nitrogen molecules to use up all the hydrogen molecules. We run out of nitrogen molecules and can't make any more ammonia molecules, so the process stops when we run out of nitrogen molecules. Hence, nitrogen is the limiting reagent.
3. Yes; methane is the limiting reagent. Since oxygen is in excess, it is not the limiting reagent.
5. We can perform two mass-mass calculations and determine the limiting reagent. We need the molar masses of C (12.011 g/mol), H_2 (2.016 g/mol), and C_2H_2 (26.038 g/mol).

$$7.08\text{ g C} \times \frac{1\text{ mol C}}{12.011\text{ g C}} \times \frac{1\text{ mol C}_2\text{H}_2}{2\text{ mol C}} \times \frac{26.038\text{ g C}_2\text{H}_2}{1\text{ mol C}_2\text{H}_2}$$
$$= 7.67\text{ g C}_2\text{H}_2$$

$$4.92\text{ g H}_2 \times \frac{1\text{ mol H}_2}{2.016\text{ g H}_2} \times \frac{1\text{ mol C}_2\text{H}_2}{1\text{ molH}_2} \times \frac{26.038\text{ g C}_2\text{H}_2}{1\text{ mol C}_2\text{H}_2}$$
$$= 63.54\text{ g C}_2\text{H}_2$$

The 7.67 g of C_2H_2 is the lesser quantity, so the associated reactant—7.08 g of C—is the limiting reagent. To determine how much of the other reactant is left, we have to do one more mass-mass calculation to determine what mass of H_2 has reacted with the 7.08 g of C and then subtract the amount reacted from the original amount.

$$7.08\text{ g C} \times \frac{1\text{ molC}}{12.011\text{ g C}} \times \frac{1\text{ mol H}_2}{2\text{ mol C}} \times \frac{2.016\text{ g H}_2}{1\text{mol H}_2}$$
$$= 0.59\text{ g H}_2\text{ reacted}$$

The amount of the excess reagent, H_2, left over is: 4.92 g H_2 – 0.59 g H_2 reacted = 4.33 g H_2

7. We can perform two mass-mass calculations and determine the limiting reagent. We need the molar masses of P_4O_6 (219.89 g/mol), H_2O (18.0 g/mol), and H_3PO_4 (97.99 g/mol).

$$35.6\text{ g }P_4O_6 \times \frac{1\text{ mol }P_4O_6}{219.89\text{ g }P_4O_6} \times \frac{4\text{ mol }H_3PO_4}{1\text{ mol }P_4O_6} \times \frac{97.99\text{ g }H_3PO_4}{1\text{ mol }H_3PO_4} = 63.46\text{ g }H_3PO_4$$

$$4.77\text{ g }H_2O \times \frac{1\text{ mol }H_2O}{18.0\text{ g }H_2O} \times \frac{4\text{ mol }H_3PO_4}{6\text{ mol }H_2O} \times \frac{97.99\text{ g }H_3PO_4}{1\text{ mol }H_3PO_4} = 17.31\text{ g }H_3PO_4$$

The 17.31 g of H_3PO_4 is the lesser quantity, so the associated reactant—4.77 g of H_2O—is the limiting reagent. To determine how much of the other reactant is left, we have to do one more mass-mass calculation to determine what mass of P_4O_6 has reacted with the 7.08 g of H_2O and then subtract the amount reacted from the original amount.

$$4.77\text{ g }H_2O \times \frac{1\text{ mol }H_2O}{18.0\text{ g }H_2O} \times \frac{1\text{ mol }P_4O_6}{6\text{ mol }H_2O} \times \frac{219.89\text{ g }P_4O_6}{1\text{ mol }P_4O_6} = 9.71\text{g }P_4O_6\text{ reacted}$$

The amount of P_4O_6 (excess reagent) left over is: 35.6 g P_4O_6 – 9.71 g P_4O_6 reacted = 25.9 g P_4O_6

9. The balanced chemical reaction is

$$2NaCl(aq) + Pb(NO_3)_2(aq) \rightarrow PbCl_2(s) + 2NaNO_3(aq)$$

We can perform two mass-mass calculations and determine the limiting reagent. We need the molar masses of NaCl (58.44 g/mol), Pb $(NO_3)_2$ (331.21 g/mol), and $PbCl_2$ (278.11 g/mol).

$$2.88 \times \frac{1\text{ mol NaCl}}{58.44\text{ g NaCl}} \times \frac{1\text{ mol }PbCl_2}{2\text{ mol NaCl}} \times \frac{278.11\text{ g }PbCl_2}{1\text{ mol }PbCl_2} = 6.85\text{ g }PbCl_2$$

$$7.21\text{ g }Pb(NO_3)_2 \times \frac{1\text{ mol }Pb(NO_3)_2}{331.21\text{ g }Pb(NO_3)_2} \times \frac{1\text{ mol }PbCl_2}{1\text{ mol }Pb(NO_3)_2} \times \frac{278.11\text{ g }PbCl_2}{1\text{ mol }PbCl_2} = 6.05\text{ g }PbCl_2$$

The 6.05 g of $Pb(NO_3)_2$ is the lesser quantity, so the associated reactant—7.21 g of $Pb(NO_3)_2$—is the limiting reagent. To determine how much of the other reactant is left, we have to do one more mass-mass calculation to determine what mass of NaCl has reacted with the 7.21 g of $Pb(NO_3)_2$ and then subtract the amount reacted from the original amount.

$$7.21\text{ g }Pb(NO_3)_2 \times \frac{1\text{ mol }Pb(NO_3)_2}{331.21\text{ g }Pb(NO_3)_2} \times \frac{2\text{ mol NaCl}}{1\text{ mol }Pb(NO_3)_2} \times \frac{58.44\text{ g NaCl}}{1\text{ mol NaCl}} = 2.54\text{ g NaCl reacted}$$

The amount of NaCl (excess reagent) left over is: 2.88 g of NaCl – 2.54 g NaCl reacted = 0.34 g NaCl left over

Hence, 7.21 g of $Pb(NO_3)_2$ reacts with 2.88 g of NaCl to form 6.05 g of $PbCl_2$ precipitate and 0.34 g of NaCl is in excess.

7. END-OF-CHAPTER MATERIAL

Additional Exercises

1. From the balanced reaction, we can write the appropriate conversion factor. Starting with the given amount and using the conversion factor, we get

$$6.022\times10^{23}\text{molecules H}_2\times\frac{1\text{ molecule O}_2}{2\text{ molecules H}_2}=3.011\times10^{23}\text{ molecules O}_2$$

Hence, 3.011×10^{23} molecules of O_2 will react with 6.022×10^{23} molecules of H_2 to make water.

3. The molar mass of CO_2 is 44.01 g/mol. Inverting the unit and using it as a conversion factor, we can calculate the number of moles present in 6.411 kg of CO_2.

$$6.411\times10^{3}\text{g CO}_2\times\frac{1\text{ mol CO}_2}{44.01\text{ g CO}_2}=145.7\text{ mol CO}_2$$

Using Avogadro number, we can calculate the number of molecules in this quantity.

$$145.7\text{ mol CO}_2\times\frac{6.022\times10^{23}\text{ molecules CO}_2}{1\text{ mol CO}_2}=8.774\times10^{25}\text{ molecules CO}_2$$

Hence, 8.774×10^{25} molecules are present in 6.411 kg of CO_2.

5. The molar mass of CO_2 is 44.01 g/mol. Starting with the given amount and using the appropriate conversion factor, we get

$$7.22\times10^{20}\text{molecules CO}_2\times\frac{44.01\text{ g CO}_2}{6.022\times10^{23}\text{molecules CO}_2}\times\frac{1{,}000\text{ mg CO}_2}{1\text{ g CO}_2}=52.8\text{ mg CO}_2$$

7. The molar mass of H_2O is 18.0 g/mol. Starting with the given amount and using the appropriate conversion factor, we get

$$1\text{ molecule H}_2\text{O}\times\frac{18.0\text{ g H}_2\text{O}}{6.022\times10^{23}\text{ molecules H}_2\text{O}}=2.99\times10^{-23}\text{ g H}_2\text{O}$$

9. The molar mass of Ga is 69.72 g/mol. Starting with the given amount and using the appropriate conversion factor, we get

$$3.44\text{ mol Ga}\times\frac{69.72\text{ g Ga}}{1\text{ mol Ga}}\times\frac{1\text{ mL Ga}}{6.08\text{ g Ga}}=39.4\text{ mL Ga}$$

11. The molar mass of C_4H_{10} is 58.1 g/mol. The molar mass of CO_2 is 44.0 g/mol. The mass of carbondioxide produced during the reaction can be calculated as follows:

$$13.4\text{ g C}_4\text{H}_{10}\times\frac{1\text{ mol C}_4\text{H}_{10}}{58.1\text{ g C}_4\text{H}_{10}}\times\frac{8\text{ mol CO}_2}{2\text{ mol C}_4\text{H}_{10}}\times\frac{44.0\text{ g CO}_2}{1\text{ mol CO}_2}=40.6\text{ g CO}_2$$

The density of CO_2 is 1.96 g/L. So, the volume of 40.6 g CO_2 in liters is

$$40.6\text{ g CO}_2\times\frac{1\text{ L CO}_2}{1.96\text{ g CO}_2}=20.7\text{ L CO}_2$$

13. We need the molar masses of CuCl (99.0 g/mol), Cu(63.5 g/mol), and $CuCl_2$(134.5 g/mol). The mass of $CuCl_2$ formed is:

$$100\text{ g CuCl}\times\frac{1\text{ mol CuCl}}{99\text{ g CuCl}}\times\frac{1\text{mol CuCl}_2}{2\text{ mol CuCl}}\times\frac{134.5\text{ g CuCl}_2}{1\text{ mol CuCl}_2}=67.91\text{ g CuCl}_2$$

$$100\text{ g CuCl}\times\frac{1\text{ mol CuCl}}{99\text{ g CuCl}}\times\frac{1\text{mol Cu}}{2\text{ mol CuCl}}\times\frac{63.5\text{ g Cu}}{1\text{ mol Cu}}=32.09\text{ g Cu}$$

We find that the two masses add to 100.0 g, the initial amount of starting material, demonstrating the law of conservation of matter.

15. The balanced chemical equation is

$$2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$$

From the balanced equation we can construct an appropriate conversion factor and find the mass of CO_2 produced. We are given 1 gallon of gasoline, i.e. 2,801 g of C_8H_{18} (the molar mass of C_8H_{18} is 114.2 g/mol). The amount of CO_2 (the molar mass of CO_2 is 44.0 g/mol) produced from it can be calculated as follows:

$$2{,}801\text{ g }C_8H_{18} \times \frac{1\text{ mol }C_8H_{18}}{114.2\text{ g }C_8H_{18}} \times \frac{16\text{ mol }CO_2}{2\text{ mol }C_8H_{18}} \times \frac{44.0\text{ g }CO_2}{1\text{ mol }CO_2} = 8{,}633\text{ g }CO_2$$

17. The percent yield is given by:

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

Therefore, actual yield can be given by,

$$\text{actual yield} = \frac{\text{percent yield}\times\text{theoretical yield}}{100} = \frac{88.40 \times 19.98\text{ g}}{100} = 17.66\text{ g}$$

19. We can perform three mass-mass calculations and determine the limiting reagent. We need the molar masses of P_4 (123.88 g/mol), NaOH (40.0 g/mol), H_2O (18.0 g/mol), and Na_2HPO_4 (142.0 g/mol).

$$35.0\text{ g }P_4 \times \frac{1\text{ mol }P_4}{123.88\text{ g }P_4} \times \frac{3\text{ mol }Na_2HPO_4}{2\text{ mol }P_4} \times \frac{142.0\text{ g }Na_2HPO_4}{1\text{ mol }Na_2HPO_4} = 60.18\text{ g }Na_2HPO_4$$

$$12.7\text{ g NaOH} \times \frac{1\text{ mol NaOH}}{40.0\text{ g NaOH}} \times \frac{3\text{ mol }Na_2HPO_4}{6\text{ mol NaOH}} \times \frac{142.0\text{ g }Na_2HPO_4}{1\text{ mol }Na_2HPO_4} = 22.54\text{ g }Na_2HPO_4$$

$$9.33\text{ g }H_2O \times \frac{1\text{ mol }H_2O}{18.0\text{ g }H_2O} \times \frac{3\text{ mol }Na_2HPO_4}{6\text{ mol }H_2O} \times \frac{142.0\text{ g }Na_2HPO_4}{1\text{ mol }Na_2HPO_4} = 36.80\text{ g }Na_2HPO_4$$

The 22.54 g of Na_2HPO_4 is the lesser quantity, so the associated reactant—12.7 g of NaOH —is the limiting reagent. To determine how much of the other reactants are left, we have to do two more mass-mass calculations.

$$12.7\text{ g NaOH} \times \frac{1\text{ mol NaOH}}{40.0\text{ g NaOH}} \times \frac{2\text{ mol }P_4}{6\text{ mol NaOH}} \times \frac{123.88\text{ g }P_4}{1\text{ mol }P_4} = 13.11\text{ g }P_4$$

The amount of P_4 (excess reagent) left over is: 35.0 g P_4 – 13.11 g P_4 reacted = 21.9 g P_4

$$12.7\text{ g NaOH} \times \frac{1\text{ mol NaOH}}{40.0\text{ g NaOH}} \times \frac{6\text{ mol }H_2O}{6\text{ mol NaOH}} \times \frac{18.0\text{ g }H_2O}{1\text{ mol }H_2O} = 5.72\text{ g }H_2O$$

The amount of H_2O (excess reagent) left over is: 9.33 g H_2O – 5.72 g H_2O reacted = 3.61 g H_2O

21. We can perform two mass-mass calculations and determine the limiting reagent. We need the molar masses of C_3H_8 (44.1 g/mol), O_2 (32.0 g/mol), CO_2(44.0 g/mol), and H_2O(18.0 g/mol). Let us predict the limiting reagent by determining the product CO_2.

$$26.3\text{ g } C_3H_8 \times \frac{1\text{ mol } C_3H_8}{44.1\text{ g } C_3H_8} \times \frac{3\text{ mol } CO_2}{1\text{ mol } C_3H_8} \times \frac{44.0\text{ g } CO_2}{1\text{ mol } CO_2} = 78.7\text{ g } CO_2$$

$$21.8\text{ g } O_2 \times \frac{1\text{ mol } O_2}{32.0\text{ g } O_2} \times \frac{3\text{ mol } CO_2}{5\text{ mol } O_2} \times \frac{44.0\text{ g } CO_2}{1\text{ mol } CO_2} = 18.0\text{ g } CO_2$$

The 18.0 g of CO_2 is the lesser quantity, so the associated reactant—21.8 g of O_2—is the limiting reagent. To determine how much of the other reactant is left, we have to do one more mass-mass calculation:

$$21.8\text{ g } O_2 \times \frac{1\text{ mol } O_2}{32.0\text{ g } O_2} \times \frac{1\text{ mol } C_3H_8}{5\text{ mol } O_2} \times \frac{44.1\text{ g } C_3H_8}{1\text{ mol } C_3H_8} = 6.0\text{ g } C_3H_8\text{ reacted}$$

The amount of C_3H_8 (excess reagent) left over is: 26.3 g C_3H_8− 6.0 g C_3H_8 reacted = 20.3 g C_3H_8 left

Let us predict the limiting reagent by determining the product H_2O.

$$26.3\text{ g } C_3H_8 \times \frac{1\text{ mol } C_3H_8}{44.1\text{ g } C_3H_8} \times \frac{4\text{ mol } H_2O}{1\text{ mol } C_3H_8} \times \frac{18.0\text{ g } H_2O}{1\text{ mol } H_2O} = 42.9\text{ g } H_2O$$

$$21.8\text{ g } O_2 \times \frac{1\text{ mol } O_2}{32.0\text{ g } O_2} \times \frac{4\text{ mol } H_2O}{5\text{ mol } O_2} \times \frac{18.0\text{ g } H_2O}{1\text{ mol } H_2O} = 9.81\text{ g } H_2O$$

The 9.81 g of H_2O is the lesser quantity, so the associated reactant—21.8 g of O_2—is the limiting reagent.

Both products predict that O_2 is the limiting reagent; 20.3 g of C_3H_8 are left over.

CHAPTER 6

Gases

1. KINETIC THEORY OF GASES

Exercises

1. Gases consist of tiny particles of matter that are in constant motion. Gas particles are constantly colliding with each other and the walls of a container; these collisions are elastic; that is, there is no net loss of energy from the collisions. Gas particles are separated by large distances, with the size of a gas particle tiny compared to the distances that separate them. There are no interactive forces (i.e., attraction or repulsion) between the particles of a gas. The average speed of gas particles is dependent on the temperature of the gas.

3. An ideal gas is a gas that exactly follows the statements of the kinetic theory of gases. Ideal gases do not exist, but the kinetic theory allows us to model them.

2. PRESSURE

Test Yourself

1. Using the pressure equivalences, we construct a conversion factor between torr and atmospheres: $\frac{1\text{ atm}}{760\text{ torr}}$. Thus,

   $1{,}022\text{ torr} \times \frac{1\text{ atm}}{760\text{ torr}} = 1.345\text{ atm}$

Test Yourself

1. Using the pressure equivalences, we construct a conversion factor between torr and atmospheres: $\frac{760\text{ torr}}{1\text{ atm}}$. Thus,

   $0.859\text{ atm} \times \frac{760\text{ torr}}{1\text{ atm}} = 653\text{ torr}$

Exercises

1. Pressure is force per unit area. It is caused by gas particles hitting the walls of its container.
3. One pascal is defined as 1 N/m^2. Pressure is force per unit area. Thus, pressure in pascals is:

   $\frac{16.7\text{ N}}{2.44\text{ m}^2} = 6.84\text{ N/m}^2 = 6.84\text{ Pa}$

5. Because the atmospheric pressure at sea level is variable, it is not a consistent unit of pressure.
7. Using the pressure equivalences, we construct a conversion factor between mmHg and atmospheres: $\frac{1\text{ atm}}{760\text{ mmHg}}$. Thus,

   $889\text{ mmHg} \times \frac{1\text{ atm}}{760\text{ mmHg}} = 1.17\text{ atm}$

9. Using the pressure equivalences, we construct a conversion factor between torr and atmospheres: $\frac{760\text{ torr}}{1\text{ atm}}$. Thus,

   $2.443\text{ atm} \times \frac{760\text{ torr}}{1\text{ atm}} = 1{,}857\text{ torr}$

11. Using the pressure equivalences, we construct a conversion factor between mmHg and torr: 1 torr = 1 mmHg. Therefore,

    334 torr = 334 mmHg.
13. Using the pressure equivalences, we construct a conversion factor between pascals and torr: $\frac{101{,}325\text{ pascals}}{760\text{ torr}}$. Thus,

    $1\text{ torr} \times \frac{101{,}325\text{ pascals}}{760\text{ torr}} = 133\text{ pascals}$

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3. GAS LAWS

Test Yourself

1. $P_1 = 334$ torr; $P_2 = 102$ torr; $V_1 = 37.8$ mL; $V_2 = ?$

   Substituting these values into Boyle's law, $P_1V_1 = P_2V_2$, we get:

$$(334\text{ torr})(37.8\text{ mL}) = (102\text{ torr})V_2$$

$$\frac{(334\text{ torr})(37.8\text{ mL})}{102\text{ torr}} = V_2$$

$$V_2 = 124\text{ mL}$$

Test Yourself

1. $P_1 = 308$ torr ; $P_2 = 1.55$ atm; $V_1 = 456$ mL; $V_2 = ?$

   To have the same units of pressure, convert P_2 from atm to torr:

$$1.55\text{ atm} \times \frac{760\text{ torr}}{1\text{ atm}} = 1{,}178\text{ torr}$$

   Substituting these values into Boyle's law, $P_1V_1 = P_2V_2$, we get:

$$(308\text{ torr})(456\text{ mL}) = (1{,}178\text{ torr })V_2$$

$$\frac{(308\text{ torr})(456\text{ mL})}{1{,}178\text{ torr}} = V_2$$

$$V_2 = 119\text{ mL}$$

Test Yourself

1. $V_1 = 3.77$ L; $T_1 = 255$ K; $T_2 = 123$ K; $V_2 = ?$

   Substituting these values into Charles's law, $\frac{V_1}{T_1} = \frac{V_2}{T_2}$, we get:

$$\frac{3.77\text{ L}}{255\text{ K}} = \frac{V_2}{123\text{ K}}$$

$$\frac{(3.77\text{ L})(123\text{ K})}{255\text{ K}} = V_2$$

$$V_2 = 1.82\text{ L}$$

Test Yourself

1. $V_1 = 623$ mL ; $T_1 = 255°C$; $V_2 = 277$ mL; $T_2 = ?$

   Converting T_1 from °C to K:

   $T_1 = 255°C + 273.15 \text{ K} = 528 \text{ K}$

   Substituting these values into Charles's law, $\frac{V_1}{T_1} = \frac{V_2}{T_2}$, we get:

   $$\frac{623 \text{ mL}}{528 \text{ K}} = \frac{277 \text{ mL}}{T_2}$$

   $$\frac{(528 \text{ K})(277 \text{ mL})}{623 \text{ mL}} = T_2$$

   $$T_2 = 235 \text{ K}$$

   Converting T_2 from K to °C:

   $T_2 = 235 - 273.15 = -38°C$

Exercises

1. A gas law is a simple mathematical formula that allows one to predict the physical properties of a gas. The units of changing properties (volume, pressure, etc.) must be the same. All temperatures must be expressed in the absolute temperature scale (Kelvin).
3. pressure, volume.
5. amount of gas and temperature
7. $P_1 = 1.445$ atm; $V_1 = 1.009$ L; $P_2 = ?$; $V_2 = 0.556$ L

   Since temperature and amount are held constant, Boyle' law is used.

   Substituting these values into Boyle's law, $P_1V_1 = P_2V_2$, we get:

   $$(1.445 \text{ atm})(1.009 \text{ L}) = (0.556 \text{ L})P_2$$

   $$\frac{(1.445 \text{ atm})(1.009 \text{ L})}{0.556 \text{ atm}} = P_2$$

   $$P_2 = 2.62 \text{ atm}$$

9. $P_1 = 4.33$ atm; $V_1 = 5.88$ L; $P_2 = 0.506$ atm; $V_2 = ?$

   Since temperature and amount are held constant, Boyle' law is used.

   Substituting these values into Boyle's law, $P_1V_1 = P_2V_2$, we get:

   $$(4.33 \text{ atm})(5.88 \text{ L}) = (0.506 \text{ atm})V_2$$

   $$\frac{(4.33 \text{ atm})(5.88 \text{ L})}{0.506 \text{ atm}} = V_2$$

   $$V_2 = 50.3 \text{ L}$$

11. $P_1 = 779$ torr; $V_1 = 638$ mL; $P_2 = 0.335$ atm; $V_2 = ?$

    To have the same units of pressure, convert P_2 from atm to torr:

    $$0.335 \text{ atm} \times \frac{760 \text{ torr}}{1 \text{ atm}} = 255 \text{ torr}$$

    Since temperature and amount are held constant, Boyle' law is used.

    Substituting these values into Boyle's law, $P_1V_1 = P_2V_2$, we get:

    $$(779 \text{ torr})(638 \text{ mL}) = (255 \text{ torr})V_2$$

    $$\frac{(779 \text{ torr})(638 \text{ mL})}{255 \text{ torr}} = V_2$$

    $$V_2 = 1,950 \text{ mL}$$

    Converting V_2 from mL to L:

    $$1{,}950 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 1.95 \text{ L}$$

13. $V_1 = 67.5$ mL; $T_1 = 315$ K; $V_2 = ?$; $T_2 = 244$ K

    Since pressure and amount are held constant, Charles's law is used.

    Substituting these values into Charles's law, $\frac{V_1}{T_1} = \frac{V_2}{T_2}$, we get:

    $$\frac{67.5 \text{ mL}}{315 \text{ K}} = \frac{V_2}{244 \text{ K}}$$

    $$\frac{(67.5 \text{ mL})(244 \text{ K})}{315 \text{ K}} = V_2$$

    $$V_2 = 52.3 \text{ mL}$$

15. $V_1 = 655$ mL; $T_1 = 295$ K ; $V_2 = 577$ mL; $T_2 = ?$

Since pressure and amount are held constant, Charles's law is used.

Substituting these values into Charles's law, $\frac{V_1}{T_1} = \frac{V_2}{T_2}$, we get:

$$\frac{655\text{ mL}}{295\text{ K}} = \frac{577\text{ mL}}{T_2}$$

$$\frac{(295\text{ K})(577\text{ mL})}{655\text{ mL}} = T_2$$

$$T_2 = 260\text{ K}$$

17. $V_1 = 685$ mL; $T_1 = 29$°C; $V_2 = 1.006$ L; $T_2 = ?$

To have the same units of volume, convert V_2 from L to mL:

$$1.006\text{ L} \times \frac{1{,}000\text{ mL}}{1\text{ L}} = 1{,}006\text{ mL}$$

Converting T_1 from °C to K:

$T_1 = 29 + 273.15 = 302$ K

Since pressure and amount are held constant, Charles's law is used.

Substituting these values into Charles's law, $\frac{V_1}{T_1} = \frac{V_2}{T_2}$, we get:

$$\frac{685\text{ mL}}{302\text{ K}} = \frac{1{,}006\text{ mL}}{T_2}$$

$$\frac{(1{,}006\text{ mL})(302\text{ K})}{685\text{ mL}} = T_2$$

$$T_2 = 444\text{ K}$$

Converting T_2 from K to °C:

$T_2 = 444 - 273.15 = 171^{\text{o}}\text{C}$

4. OTHER GAS LAWS

Test Yourself

1. V_1 =12.8 L; n_1 = 3.00 × 10^{20} particles; V_2 = ?; n_2 = 8.22 × 10^{18} particles

   Since temperature and pressure are held constant, Avogadro's law can be used.

   Substituting these values in the Avogadro's law, $\frac{V_1}{n_1} = \frac{V_2}{n_2}$, we get:

$$\frac{12.8\ \text{L}}{3.00\ \times\ 10^{20}\ \text{particles}} = \frac{V_2}{8.22\ \times\ 10^{18}\ \text{particles}}$$

$$V_2 = \frac{12.8\ \text{L}}{3.00\ \times\ 10^{20}\text{particles}} \times 8.22\ \times\ 10^{18}\ \text{particles}$$

$$V_2 = 0.351\ \text{L}$$

Test Yourself

1. P_1 = 662 torr; V_1 = 46.7 mL; T_1 = 266 K; P_2 = 409 torr; V_2 = ?; T_2 = 371 K

   Substituting the values in the combined gas law, $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$, we get:

$$\frac{(662\ \text{torr})(46.7\ \text{mL})}{266\ \text{K}} = \frac{(409\ \text{torr})V_2}{371\ \text{K}}$$

$$\frac{(662\ \text{torr})(46.7\ \text{mL})(371\ \text{K})}{(266\ \text{K})(409\ \text{torr})} = V_2$$

$$V_2 = 105\ \text{mL}$$

Exercises

1. The pressure of a gas is proportional to its absolute temperature.
3. P_1 = 602 torr; T_1 = 356 K; P_2 = ?; T_2 = 277 K

   Since volume and amount are held constant, Gay-Lussac's law can be used.

   Substituting these values in the Gay-Lussac's law, $\frac{P_1}{T_1} = \frac{P_2}{T_2}$, we get:

$$\frac{602\ \text{torr}}{356\ \text{K}} = \frac{P_2}{277\ \text{K}}$$

$$P_2 = \frac{(602\ \text{torr})(277\ \text{K})}{356\ \text{K}}$$

$$P_2 = 468\ \text{torr}$$

5. V_1 =1.55 L; n_1 = 3.45 × 10^{22} atoms; V_2 = ?; n_2 = 6.00 × 10^{23} atoms

   Since temperature and pressure are held constant, Avogadro's law can be used.

   Substituting these values in the Avogadro's law, $\frac{V_1}{n_1} = \frac{V_2}{n_2}$, we get:

$$\frac{1.55\ \text{L}}{3.45\ \times\ 10^{22}\ \text{atoms}} = \frac{V_2}{6.00\ \times\ 10^{23}\ \text{atoms}}$$

$$V_2 = \frac{1.55\ \text{L}}{3.45\ \times\ 10^{22}\ \text{atoms}} \times 6.00\ \times\ 10^{23}\ \text{atoms}$$

$$V_2 = 27.0\ \text{L}$$

7. $V_1 = 6.72$ L; $n_1 = 3.88$ mol; $V_2 = ?$; $n_2 = 6.10$ mol

Since temperature and pressure are held constant, Avogadro's law can be used.

Substituting these values in the Avogadro's law, $\frac{V_1}{n_1} = \frac{V_2}{n_2}$, we get:

$$\frac{6.72\ \text{L}}{3.88\ \text{mol}} = \frac{V_2}{6.10\ \text{mol}}$$

$$V_2 = \frac{6.72\ \text{L}}{3.88\ \text{mol}} \times 6.10\ \text{mol}$$

$$V_2 = 10.6\ \text{L}$$

9. The combined gas law is

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \quad \text{at constant } n$$

Substituting the given values in the above expression, we can solve the above table:

$$\frac{(334\ \text{torr})(56.9\ \text{mL})}{266\ \text{K}} = \frac{(722\ \text{torr})V_2}{334\ \text{K}}$$

a. $$\frac{(334\ \text{K})(334\ \text{torr})(56.9\ \text{mL})}{(722\ \text{torr})(266\ \text{K})} = V_2 = 33.1\ \text{mL}$$

$$\frac{(2.33\ \text{atm})(0.976\ \text{L})}{443\ \text{K}} = \frac{(1.223\ \text{L})P_2}{355\ \text{K}}$$

b. $$\frac{(355\ \text{K})(2.33\ \text{atm})(0.976\ \text{L})}{(1.223\ \text{L})(443\ \text{L})} = P_2 = 1.49\ \text{atm}$$

c. We have to convert the unit of temperature of T_1 from °C to K:

$$T_1 = 23 + 273.15\ \text{K} = 296\ \text{K}$$

$$\frac{(889\ \text{torr})(3.66\ \text{L})}{296\ \text{K}} = \frac{(2.19\ \text{L})(739\ \text{torr})}{T_2}$$

$$T_2 = \frac{(2.19\ \text{L})(739\ \text{torr})(296\ \text{K})}{(889\ \text{torr})(3.66\ \text{L})} = 147\ \text{K}$$

$$\text{or, } T_2 = 147 - 273.15°\text{C} = -126°\text{C}$$

These values can be tabulated as follows:

$V_1 =$	$P_1 =$	$T_1 =$	$V_2 =$	$P_2 =$	$T_2 =$
56.9 mL	334 torr	266 K	33.1 mL	722 torr	334 K
0.976 L	2.33 atm	443 K	1.223 L	1.49 atm	355 K
3.66 L	889 torr	23°C	2.19 L	739 torr	147 K, or −126°C

11. $P_1 = 3.008$ atm; $V_1 = 78.9$ mL; $T_1 = 56$°C; $P_2 = 2.55$ atm; $V_2 = 35.6$ mL; $T_2 = ?$

Converting T_1 from °C to K:

$T_1 = 56 + 273.15$ K $= 329$ K

Substituting the values in the combined gas law, $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$, we get:

$$\frac{(3.008\ \text{atm})(78.9\ \text{mL})}{329\ \text{K}} = \frac{(2.55\ \text{atm})(35.6\ \text{mL})}{T_2}$$

$$T_2 = \frac{(35.6\ \text{mL})(2.55\ \text{atm})(329\ \text{K})}{(3.008\ \text{atm})(78.9\ \text{mL})}$$

$$T_2 = 126\ \text{K}$$

$$\text{or, } T_2 = 126 - 273.15°\text{C} = -147°\text{C}$$

13. $P_1 = 7.11$ atm; $V_1 = 90.7$ mL; $T_1 = 66$°C; $P_2 = 14.33$ atm; $V_2 = ?$; $T_2 = 33$°C

Converting T_1 and T_2 from °C to K:

$$T_1 = 66 + 273.15\ \text{K} = 339\ \text{K}$$

$$T_2 = 33 + 273.15\ \text{K} = 306\ \text{K}$$

Substituting the values in the combined gas law, $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$, we get:

$$\frac{(7.11\ \text{atm})(90.7\ \text{mL})}{339\ \text{K}} = \frac{(14.33\ \text{atm})V_2}{306\ \text{K}}$$

$$V_2 = \frac{(306\ \text{K})(7.11\ \text{atm})(90.7\ \text{mL})}{(14.33\ \text{atm})(339\ \text{K})}$$

$$V_2 = 40.6\ \text{mL}$$

5. THE IDEAL GAS LAW AND SOME APPLICATIONS

Test Yourself

1. Substituting the given values in the ideal gas law, $PV = nRT$, we get:

$$(0.692\ \text{atm})V = (0.0997\ \text{mol})\left(0.08205\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(333\ \text{K})$$

$$V = \frac{(0.0997\ \text{mol})\left(0.08205\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(333\ \text{K})}{0.692\ \text{atm}}$$

$$V = 3.94\ \text{L}$$

Test Yourself

1. Pressure is given in units of torr. We can either convert this to atmospheres or use the value of the ideal gas constant that includes the torr unit. We will take the second option. Substituting into the ideal gas law, $PV = nRT$, we get:

$$(23.44\ \text{torr})V = (0.00554\ \text{mol})(\ 62.36\ \frac{\text{L}\cdot\text{torr}}{\text{mol}\cdot\text{K}}\ )(557\ \text{K})$$

$$V = \frac{(0.00554\ \text{mol})(\ 62.36\ \frac{\text{L}\cdot\text{torr}}{\text{mol}\cdot\text{K}}\ )(557\ \text{K})}{23.44\ \text{torr}}$$

$$V = 8.21\ \text{L}$$

Test Yourself

1. Here, we have a stoichiometry problem, where we need to find the number of moles of HCl produced. Then, we can use the ideal gas law, with the given temperature and volume, to determine the pressure of gas produced. First, the number of moles of HCl is calculated:

$$3.44\ \text{g Cl}_2 \times \frac{1\ \text{mol Cl}_2}{70.91\ \text{g Cl}_2} \times \frac{2\ \text{mol HCl}}{1\ \text{mol Cl}_2} = 0.0970\ \text{mol HCl}$$

Now that we know the number of moles of gas, we can use the ideal gas law, $PV = nRT$, to determine the pressure, given the other conditions:

$$P(4.55\ \text{L}) = (0.0970\ \text{mol})\left(0.08205\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(455\ \text{K})$$

$$P = \frac{(0.0970\ \text{mol})\left(0.08205\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(455\ \text{K})}{4.55\ \text{L}}$$

$$P = 0.796\ \text{atm}$$

Test Yourself

1. We can use the molar volume, 22.4 L/mol, as a conversion factor, the mol units cancel and L unit is introduced. It is a one-step conversion:

$$4.87\ \text{mol} \times \frac{22.4\ \text{L}}{1\ \text{mol}} = 109\ \text{L}$$

Test Yourself

1. Here, we have a stoichiometry problem. We need to find the number of moles of HCl produced. Then, we can use the ideal gas law, with the given temperature and volume, to determine the pressure of gas produced. First, the number of moles of HCl is calculated:

$$3.44 \text{ g } Cl_2 \times \frac{1 \text{ mol } Cl_2}{70.91 \text{g } Cl_2} \times \frac{2 \text{ mol HCl}}{1 \text{ mol } Cl_2} = 0.0970 \text{ mol HCl}$$

At STP, the molar volume is 22.4 L/mol. We can use the molar volume, 22.4 L/mol, as a conversion factor:

$$0.097 \text{ mol } \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 2.17 \text{ L}$$

Hence, the volume of HCl generated is 2.17 L.

Test Yourself

1. If we assume exactly 1 mol of CO_2, then we know its mass is 44.0 g. Using the ideal gas law, $PV = nRT$, we can calculate the volume:

$$(0.0079 \text{ atm})V = (1 \text{ mol})(\ 0.08205 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\ )(227 \text{ K})$$

$$V = \frac{(1 \text{ mol})(\ 0.08205 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\ )(227 \text{ K})}{0.0079 \text{ atm}}$$

$$V = 2{,}360 \text{ L}$$

Knowing the molar mass and the molar volume, we can determine the density of N_2 under these conditions, using the formula, $\text{density} = \frac{\text{molar mass}}{\text{molar volume}}$.

$$d = \frac{44.0 \text{ g}}{2{,}360 \text{ L}} = 0.019 \text{ g/L}$$

Hence, density of CO_2 is 0.019 g/L.

Exercises

1. The ideal gas law is $PV = nRT$. R is the ideal gas law constant, which relates the other four variables.
3. Substituting the values in ideal gas law, $PV = nRT$, we get:

$$(2.09 \text{ atm})(3.91 \text{ L}) = n(\ 0.08205 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\ )(305 \text{ K})$$

$$n = \frac{(2.09 \text{ atm})(3.91 \text{ L})}{(\ 0.08205 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\ )(305 \text{ K})}$$

$$n = 0.327 \text{ mol}$$

5. First, we have to convert the unit of temperature to kelvins: 188 + 273 = 461 K
Substituting the values in the ideal gas law, $PV = nRT$, we get:

$$P(0.577 \text{ L}) = (0.0555 \text{ mol})(\ 0.08205 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\ )(461 \text{ K})$$

$$P = \frac{(0.0555 \text{ mol})(\ 0.08205 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\ )(461 \text{ K})}{0.577 \text{ L}}$$

$$P = 3.64 \text{ atm}$$

7. First, we have to convert the unit of temperature into kelvins: 123 + 273.15 = 396 K
The number of moles of He in 7.55 g of He (molar mass = 4.0 g/mol):

$$7.55 \text{ g} \times \frac{1 \text{ mol}}{4.0 \text{ g}} = 1.89 \text{ mol He}$$

Converting the unit of volume from mL to L, using the conversion factor $\frac{1 \text{ L}}{1000 \text{ mL}}$, we get:

$$5{,}520 \text{ mL} \times \frac{1 \text{ L}}{1{,}000 \text{ mL}} = 5.520\ L$$

Substituting the values in the ideal gas law, $PV = nRT$, we get:

$$P(5.520 \text{ L}) = (1.89 \text{ mol})(\ 62.36 \frac{\text{L} \cdot \text{torr}}{\text{mol} \cdot \text{K}}\ )(396 \text{ K})$$

$$P = \frac{(1.89 \text{ mol})(\ 62.36 \frac{\text{L} \cdot \text{torr}}{\text{mol} \cdot \text{K}}\ )(396 \text{ K})}{5.520 \text{ L}}$$

$$P = 8{,}460 \text{ torr}$$

9. Substituting the values in the ideal gas law, $PV = nRT$, we get:

$$(0.772\ \text{atm})(18.95\ \text{L}) = n(\ 0.08205\,\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\ )(295\ \text{K})$$

$$n = \frac{(0.772\ \text{atm})(18.95\ \text{L})}{(\ 0.08205\,\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\ )(295\ \text{K})}$$

$$n = 0.604\ \text{mol Ne}$$

The molar mass of neon is 20.18 g/mol. Hence, the mass of neon present in the sample having 0.604 mol Ne is

$$0.604\ \text{mol}\ \times \frac{20.18\ \text{g}}{1\ \text{mol}} = 12.2\ \text{g}$$

11. First, we have to determine the number of moles of O_2 formed in the reaction. The molar mass of O_2 is 32.00 g/mol and that of HgO is 216.59 g/mol. By using proper conversion factor, we can calculate the number of moles of O_2 formed.

$$3.009\ \text{g HgO}\ \times \frac{1\ \text{mol HgO}}{216.59\ \text{g HgO}} \times \frac{1\ \text{mol O}_2}{2\ \text{mol HgO}} = 0.00695\ \text{mol O}_2$$

The unit of temperature is in degree Celsius. Converting the unit into kelvins: 122 + 273.15 = 395 K

Substituting the values in the ideal gas law, $PV = nRT$, we get:

$$(744\ \text{torr})V = (0.00695\ \text{mol})(\ 62.36\,\frac{\text{L}\cdot\text{torr}}{\text{mol}\cdot\text{K}}\ )(395\ \text{K})$$

$$V = \frac{(0.00695\ \text{mol})(\ 62.36\,\frac{\text{L}\cdot\text{torr}}{\text{mol}\cdot\text{K}}\ )(395\ \text{K})}{744\ \text{torr}}$$

$$V = 0.230\ \text{L}$$

13. We can use the molar volume, 22.4 L/mol at STP as a conversion factor, the mol units cancel and L unit is introduced. It is a one-step conversion:

$$17.88\ \text{mol}\ \times \frac{22.4\ \text{L}}{1\ \text{mol}} = 401\ \text{L}$$

15. First, we need to determine the amount of CO_2 produced from 100.0 g of C_8H_{18}.

$$100.0\ \text{g}\ C_8H_{18} \times \frac{1\ \text{mol}\ C_8H_{18}}{114.23\ \text{g}\ C_8H_{18}} \times \frac{16\ \text{mol}\ CO_2}{2\ \text{mol}\ C_8H_{18}}$$

$$= 7.003\ \text{mol}\ CO_2$$

Now, to determine the volume of CO_2 produced (7.003 mol), we can use the molar volume, 22.4 L/mol at STP as a conversion factor, the mol units cancel and L unit is introduced. It is a one-step conversion:

$$7.003\ \text{mol}\ \times \frac{22.4\ \text{L}}{1\ \text{mol}} = 157\ \text{L}$$

17. a. The molar mass of He is 4.0 g/mol. Its molar volume at STP is 22.4 L/mol. The density can be calculated by dividing molar mass by molar volume.

$$d = \frac{4.00\ \text{g}}{22.4\ \text{L}} = 0.179\ \text{g/L}$$

b. The molar mass of Ne is 20.18 g/mol. Its molar volume at STP is 22.4 L/mol. The density can be calculated by dividing molar mass by molar volume.

$$d = \frac{20.18\ \text{g}}{22.4\ \text{L}} = 0.901\ \text{g/L}$$

c. The molar mass of Ar is 39.95 g/mol. Its molar volume at STP is 22.4 L/mol. The density can be calculated by dividing molar mass by molar volume.

$$d = \frac{39.95\ \text{g}}{22.4\ \text{L}} = 1.78\ \text{g/L}$$

d. The molar mass of Kr is 83.80 g/mol. Its molar volume at STP is 22.4 L/mol. The density can be calculated by dividing molar mass by molar volume.

$$d = \frac{83.80\ \text{g}}{22.4\ \text{L}} = 3.74\ \text{g/L}$$

19. If we assume exactly 1 mol of SF_6, then we know its mass 146.05 g. Using the ideal gas law, $PV = nRT$, we can calculate the volume:

$$(788\ \text{torr})V = (1\ \text{mol})(\ 62.36\,\frac{\text{L}\cdot\text{torr}}{\text{mol}\cdot\text{K}}\ )(335\ \text{K})$$

$$V = \frac{(1\ \text{mol})(\ 62.36\,\frac{\text{L}\cdot\text{torr}}{\text{mol}\cdot\text{K}}\ )(335\ \text{K})}{788\ \text{torr}}$$

$$V = 26.5\ \text{L}$$

Knowing the molar mass and the molar volume, we can determine the density of SF_6 under these conditions:

$$d = \frac{146.05\ \text{g}}{26.5\ \text{L}} = 5.51\ \text{g/L}$$

6. GAS MIXTURES

Test Yourself

1. Dalton's law of partial pressures states that the total pressure is equal to the sum of the partial pressures:

$$P_{tot} = P_{N_2} + P_{o_2}$$

$$760 \text{ torr} = 608 \text{ torr} + P_{O_2}$$

Hence, $P_{O_2} = 760 \text{ torr} - 608 \text{ torr} = 152 \text{ torr}$

Test Yourself

1. Because gases act independently of each other, we can determine the resulting final pressures using Boyle's law and then add the two resulting pressures together to get the final pressure. The total final volume is 2.00 L + 3.00 L = 5.00 L. First, we use Boyle's law to determine the final pressure of He

$$(0.75 \text{ atm})(2.00 \text{ L}) = P_2(5.00 \text{ L})$$

$$P_2 = \frac{(0.75 \text{ atm})(2.00 \text{ L})}{5.00 \text{ L}}$$

$$P_2 = 0.3 \text{ atm} = \text{ partial pressure of He}$$

Now we do that same thing for the Ne:

$$(0.35 \text{ atm})(3.00 \text{ L}) = P_2(5.00 \text{ L})$$

$$P_2 = \frac{(0.35 \text{ atm})(3.00 \text{ L})}{5.00 \text{ L}}$$

$$P_2 = 0.21 \text{ atm} = \text{ partial pressure of Ne}$$

The total pressure is the sum of the two resulting partial pressures: $P_{\text{tot}} = 0.3 \text{ atm} + 0.21 \text{ atm} = 0.51 \text{ atm}$

Test Yourself

1. We need to take into account that the total pressure includes the vapor pressure of water. According to Table 6.2, the vapor pressure of water at 50°C is 92.59 torr. According to Dalton's law of partial pressures, the total pressure equals the sum of the pressures of the individual gases, so,

$$833 \text{ torr} = P_{CO_2} + P_{H_2O} = P_{CO_2} + 92.59 \text{ torr}$$

We solve by subtracting:

$$P_{CO_2} = 740 \text{ torr}$$

Now, we can use the ideal gas law to determine the number of moles (remembering to convert temperature to kelvins, making it 323 K):

$$(740 \text{ torr})(3.50 \text{ L}) = n\left(62.36 \ \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(323 \text{ K})$$

$$n = \frac{(740 \text{ torr})(3.50 \text{ L})}{\left(62.36 \ \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(323 \text{ K})}$$

$n = 0.129$ mol CO_2 generated.

Test Yourself

1. According to Dalton's law, the total pressure is the sum of the partial pressures:

   $P_{tot} = 0.65\text{ atm} + 1.30\text{ atm} = 1.95\text{ atm}$

   The mole fractions are the ratios of the partial pressure of each component and the total pressure:

   $$\chi_{O_2} = \frac{0.65\text{ atm}}{1.95\text{ atm}} = 0.33$$

   $$\chi_{N_2} = \frac{1.30\text{ atm}}{1.95\text{ atm}} = 0.67$$

   The sum of the mole fractions is exactly 1.

Exercises

1. According to Dalton's law, the total pressure is the sum of the partial pressures:

   $P_{tot} = 0.78\text{ atm} + 0.33\text{ atm} + 1.59\text{ atm} = 2.70\text{ atm}$

3. According to Dalton's law, the total pressure is the sum of the partial pressures:

   $P_{tot} = P_{Ne} + P_{He}$

   Converting P_{He} from atm to torr to maintain uniformity of units:

   $$0.228\text{ atm} \times \frac{760\text{ torr}}{1\text{ atm}} = 173\text{ torr}$$

   Therefore, $335\text{ torr} = P_{Ne} + 173\text{ torr}$

   $P_{Ne} = 162\text{ torr}$

   Converting torr to atm:

   $$162\text{ torr} \times \frac{1\text{ atm}}{760\text{ torr}} = 0.213\text{ atm}$$

   $P_{Ne} = 162\text{ torr or } 0.213\text{ atm}$

5. The molar mass of Ar is 39.95 g/mol and that of He is 4.00 g/mol.

   Number of moles in 56.7 g of Ar = (56.7 / 39.95) mol = 1.42 mol

   Number of moles in 33.9 g of He = (33.9 / 4.00) mol = 8.48 mol

   We need to convert the unit of temperature into kelvins: 33 + 273.15 = 306 K

   For calculating the partial pressure of Ar, we substitute the values in ideal gas law:

   $$(P_{Ar})(3.55\text{ L}) = (1.42\text{ mol})\left(0.08205\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(306\text{ K})$$

   Solving the equation we get, $P_{Ar} = 10.0\text{ atm}$

   For calculating the partial pressure of He, we substitute the values in ideal gas law:

   $$(P_{He})(3.55\text{ L}) = (8.48\text{ mol})\left(0.08205\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(306\text{ K})$$

   Solving the equation we get, $P_{He} = 60.0\text{ atm}$

   According to Dalton's law, the total pressure is the sum of the partial pressures:

   $P_{tot} = 10.0\text{ atm} + 60.0\text{ atm} = 70.0\text{ atm}$

7. We need to take into account that the total pressure includes the vapor pressure of water. From the Table 6.2 in the text, the vapor pressure of water at 20°C is 17.54 torr. According to Dalton's law of partial pressures, the total pressure equals the sum of the pressures of the individual gases, so

   $688\text{ torr} = P_{O_2} + P_{H_2O} = P_{O_2} + 17.54\text{ torr}$

   We solve by subtracting:

   $P_{O_2} = 670.46\text{ torr}$

   Now we can use the ideal gas law to determine the number of moles (remembering to convert temperature to kelvins, making it 293 K):

   $$(670.46\text{ torr})(5.00\text{ L}) = n\left(62.36\ \frac{\text{L}\cdot\text{torr}}{\text{mol}\cdot\text{K}}\right)(293\text{ K})$$

   $n = 0.183\text{ mol } O_2$ collected

9. We need to take into account that the total pressure includes the vapor pressure of water. From the Table 6.2 in the text, the vapor pressure of water at 5°C is 6.54 torr.

   Convert total pressure from atm to torr:

   $$0.112\text{ atm} \times \frac{760\text{ torr}}{1\text{ atm}} = 85.1\text{ torr}$$

   According to Dalton's law of partial pressures, the total pressure equals the sum of the pressures of the individual gases, so

   $85.1\text{ torr} = P_{CO} + P_{H_2O} = P_{CO} + 6.54\text{ torr}$

   We solve by subtracting:

   $P_{CO} = 78.56\text{ torr}$

   Now we can use the ideal gas law to determine the number of moles (remembering to convert temperature to kelvins, making it 278 K):

   $$(78.56\text{ torr})(25.00\text{ L}) = n\left(62.36\ \frac{\text{L}\cdot\text{torr}}{\text{mol}\cdot\text{K}}\right)(278\text{ K})$$

   $n = 0.113$ mol CO collected

11. We need to take into account that the total pressure includes the vapor pressure of water. From the Table 6.2 in the text, the vapor pressure of water at 25°C is 23.77 torr.

Converting total pressure from atm to torr:

$$0.495 \text{ atm} \times \frac{760 \text{ torr}}{1 \text{ atm}} = 376 \text{ torr}$$

According to Dalton's law of partial pressures, the total pressure equals the sum of the pressures of the individual gases, so

$$376 \text{ torr} = P_{NO} + P_{H_2O} = P_{NO} + 23.77 \text{ torr}$$

We solve by subtracting:

$$P_{NO} = 352.23 \text{ torr}$$

Converting volume from mL to L:

$$75.0 \text{ mL} \times \frac{1 \text{ L}}{1{,}000 \text{ mL}} = 0.075 \text{ L}$$

Now we can use the ideal gas law to determine the number of moles (remembering to convert temperature to kelvins, making it 298 K):

$$(352.23 \text{ torr})(0.075 \text{ L}) = n\left(62.36 \ \frac{\text{L} \cdot \text{torr}}{\text{mol} \cdot \text{K}}\right)(298 \text{ K})$$

$n = 0.00142$ mol NO collected

The molar mass of NO is 30.0 g/mol. Using this unit as a conversion factor we can calculate the amount of NO in grams:

$$0.00142 \text{ mol NO} \times \frac{30.0 \text{ g NO}}{1 \text{ mol NO}} = 0.0426 \text{ g NO}$$

13. The molar mass of He is 4.0 g/mol, H_2 is 2.0 g/mol.

Number of moles of He is 44.5 g of He:

$$44.5 \text{ g He} \times \frac{1 \text{ mol He}}{4.0 \text{ g He}} = 11.1 \text{ mol He}$$

Number of moles of H_2 in 8.83 g of H_2:

$$8.83 \text{ g } H_2 \times \frac{1 \text{ mol } H_2}{2.0 \text{ g } H_2} = 4.42 \text{ mol } H_2$$

The total number of moles = 11.1 + 4.42 = 15.52 mol

The mole fractions are simply the ratio of each mole amount and the total number of moles.

$$\chi_{He} = \frac{11.1 \text{ mol}}{15.52 \text{ mol}} = 0.715$$

$$\chi_{Ne} = \frac{4.42 \text{ mol}}{15.52 \text{ mol}} = 0.285$$

The sum of the mole fractions equals exactly 1.

15. According to Dalton's law, the total pressure is the sum of the partial pressures:

$P_{tot} = 4.56 \text{ atm} + 2.66 \text{ atm} = 7.22 \text{ atm}$

The mole fractions are the ratios of the partial pressure of each component and the total pressure.

$$\chi_{F_2} = \frac{4.56 \text{ atm}}{7.22 \text{ atm}} = 0.632$$

$$\chi_{Cl_2} = \frac{2.66 \text{ atm}}{7.22 \text{ atm}} = 0.368$$

The sum of the mole fractions is exactly 1.

7. END-OF-CHAPTER MATERIAL

Additional Exercises

1. Pressure is defined as force per unit area. One pascal is defined as 1 N/m^2.

   Converting force from kN to N:

   $$4.88\ \text{kN} \times \frac{1{,}000\ \text{N}}{1\ \text{kN}} = 4{,}880\ \text{N}$$

   Converting area from cm^2 to m^2:

   $$235\ \text{cm}^2 \times \frac{1\ \text{m}^2}{(100\ \text{cm})^2} = 0.0235\ \text{m}^2$$

   Therefore,

   $$\text{pressure} = \frac{\text{force}}{\text{area}} = \frac{4{,}880\ \text{N}}{0.0235\ \text{m}^2} = 208{,}000\ \text{N/m}^2 \text{ or } 208{,}000\ \text{Pa}$$

3. P_1 = 822 torr; V_1 = 667 mL; T_1 = 67°C; P_2 = 2.98 atm; V_2 = 1.334 L; T_2 = ?

   Because the units are not consistent, convert V_1 from mL to L and convert P_1 from torr to atm.

   $$V_1 = 667\ \text{mL} \times \frac{1\ \text{L}}{1{,}000\ \text{mL}} = 0.667\ \text{L}$$

   $$P_1 = 822\ \text{torr} \times \frac{1\ \text{atm}}{760\ \text{torr}} = 1.081\ \text{atm}$$

   Converting T_1 from °C to K:

   T_1 = 67 + 273.15 = 340 K

   Substituting the values in the combined gas law, $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$, we get:

   $$\frac{(1.081\ \text{atm})(0.667\ \text{L})}{340\ \text{K}} = \frac{(2.98\ \text{atm})(1.334\ \text{L})}{T_2}$$

   $$\frac{(2.98\ \text{atm})(1.334\ \text{L})(340\ \text{K})}{(1.081\ \text{atm})(0.667\ \text{L})} = T_2$$

   $$T_2 = 1{,}874\ \text{K}$$

5. $$\frac{P_1V_1}{n_1} = \frac{P_2V_2}{n_2}$$

7. At constant temperature and pressure, $\frac{V_1}{n_1} = \frac{V_2}{n_2}$. Therefore,

   $$\frac{22.4\ \text{L}}{6.022 \times 10^{23}\text{particles}} = \frac{V_2}{1\ \text{particle}}$$

   $$V_2 = \frac{22.4\ \text{L} \times 1\ \text{particle}}{6.022 \times 10^{23}\text{particles}}$$

   $$V_2 = 3.72 \times 10^{-23}\ \text{L}$$

   1 particle of gas occupies 3.72×10^{-23} L.

9. The two values of R in terms of atm and torr are:

   $$R = 0.08205\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}$$

   $$R = 62.36\ \frac{\text{L}\cdot\text{torr}}{\text{mol}\cdot\text{K}}$$

   Since both values are equal to R, we can write:

   $$0.08205\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} = 62.36\ \frac{\text{L}\cdot\text{torr}}{\text{mol}\cdot\text{K}}$$

   Cancelling certain units and rearranging the equation, we get:

   $$1\ \text{atm} = \frac{62.36}{0.08205}\text{torr}$$

   or, 1 atm = 760 torr

   Yes, the number makes sense.

11. a. The initial volume can be calculated using the ideal gas law. However, the temperature must be taken in kelvins: 22 + 273.15 = 295 K

    Substituting the given values in the ideal gas law, we get:

    $$(760\ \text{torr})V = (1.000\ \text{mol})\left(62.36\ \frac{\text{L}\cdot\text{torr}}{\text{mol}\cdot\text{K}}\right)(295\ \text{K})$$

    $$V = \frac{(1.000\ \text{mol})\left(62.36\ \frac{\text{L}\cdot\text{torr}}{\text{mol}\cdot\text{K}}\right)(295\ \text{K})}{760\ \text{torr}}$$

    $$V = 24.2\ \text{L}$$

    b. The final volume can be calculated using combined gas law. However, the final temperature must be taken in kelvins: −20 + 273.15 = 253 K

    Substituting the values in the combined gas law, we get

    $$\frac{(760\ \text{torr})(24.2\ \text{L})}{295\ \text{K}} = \frac{(5.00\ \text{torr})V_2}{253\ \text{K}}$$

    $$V_2 = \frac{(760\ \text{torr})(24.2\ \text{L})(253\ \text{K})}{(295\ \text{K})(5.00\ \text{torr})}$$

    $$V_2 = 3{,}155\ \text{L}$$

13. First, we must convert the temperature into kelvins: 22 + 273.15 = 295 K

If we assume exactly 1 mol of air, then we know its mass is 28.8 g. Using the ideal gas law, we can calculate the volume:

$$(1.00\text{ atm})V = (1\text{ mol})\left(0.08205\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(295\text{ K})$$

$$V = \frac{(1\text{ mol})\left(0.08205\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(295\text{ K})}{1.00\text{ atm}}$$

$$V = 24.2\text{ L (molar volume)}$$

Knowing the molar mass and the molar volume, we can determine the density of air under these conditions:

$$d = \frac{28.8\text{ g}}{24.2\text{ L}} = 1.19\text{ g/L}$$

15. First, we must convert the temperature into kelvins: 460 + 273.15 = 733 K

If we assume exactly 1 mol of CO_2, then we know its mass is 44.0 g. Using the ideal gas law, we can calculate the volume:

$$(91.8\text{ atm})V = (1\text{ mol})\left(0.08205\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(733\text{ K})$$

$$V = \frac{(1\text{ mol})\left(0.08205\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(733\text{ K})}{91.8\text{ atm}}$$

$$V = 0.655\text{ L (molar volume)}$$

Knowing the molar mass and the molar volume, we can determine the density of CO_2 under these conditions:

$$d = \frac{44.0\text{ g}}{0.655\text{ L}} = 67.2\text{ g/L}$$

17. Here we have a stoichiometry problem where we need to find the amount of H_2 produced, when 23.8 g of Fe has been reacted.

$$23.8\text{ g Fe} \times \frac{1\text{ mol Fe}}{55.8\text{ g Fe}} \times \frac{1\text{ mol H}_2}{1\text{ mol Fe}} \times \frac{2.02\text{ g H}_2}{1\text{ mol H}_2}$$
$$= 0.862\text{ g H}_2$$

This gives the theoretical yield of H_2.

To calculate the actual yield of H_2 we use the ideal gas law. We need to take into account that the total pressure includes the vapor pressure of water. The vapor pressure of water from the Table 6.2 in the text at 25°C is 23.77 torr. According to Dalton's law of partial pressures, the total pressure equals the sum of the pressures of the individual gases, so

$$733\text{ torr} = P_{H_2} + P_{H_2O} = P_{H_2} + 23.77\text{ torr}$$

We solve by subtracting,

$$P_{H_2} = 709.23\text{ torr}$$

Converting volume of H_2 collected from mL to L:

$$446\text{ mL} \times \frac{1\text{ L}}{1{,}000\text{ mL}} = 0.446\text{ L}$$

Now we can use the ideal gas law to determine the number of moles of H_2 produced (remembering to convert temperature to kelvins, making it 298 K):

$$(709.23\text{ torr})(0.446\text{ L}) = n\left(62.36\ \frac{\text{L}\cdot\text{torr}}{\text{mol}\cdot\text{K}}\right)(298\text{ K})$$

$$n = \frac{(709.23\text{ torr})(0.446\text{ L})}{\left(62.36\ \frac{\text{L}\cdot\text{torr}}{\text{mol}\cdot\text{K}}\right)(298\text{ K})}$$

$$n = 0.0170\text{ mol}$$

Calculating the mass of H_2 collected:

$$0.0176\text{ mol H}_2 \times \frac{2.02\text{ g H}_2}{1\text{ mol H}_2} = 0.0355\text{ g H}_2$$

Percent yield is given by actual yield divided by theoretical yield. So,

$$\text{percent yield} = \frac{0.0355\text{ g}}{0.862\text{ g}} \times 100 = 4.12\%$$

CHAPTER 7

Energy and Chemistry

1. ENERGY

Test Yourself

1. We recognize that with a capital C, the Calories unit is actually kilocalories. To determine the number of joules, we convert first from kilocalories to calories (using the definition of the kilo-prefix) and then from calories to joules (using the relationship between calories and joules). So

$$110\text{ kcal} \times \frac{1{,}000\text{ cal}}{1\text{ kcal}} \times \frac{4.184\text{ J}}{1\text{ cal}} = 460{,}000\text{ J}$$

Exercises

1. Energy is the ability to do work. Work is a form of energy.
3. Using the relationship 1 cal = 4.184 J, we get

$$422\text{ J} \times \frac{1\text{ cal}}{4.184\text{ J}} = 101\text{ cal}$$

5. Using the relationship 1 cal = 4.184 J, we get

$$55.69\text{ cal} \times \frac{4.184\text{ J}}{1\text{ cal}} = 233.0\text{ J}$$

7. A Calorie is actually a kilocalorie, or 1,000 calories.
9. The total energy of an isolated system does not increase or decrease.

2. WORK AND HEAT

Test Yourself

1. First we need to determine the change in volume, ΔV. A change is always the final value minus the initial value:
   $\Delta V = V_{\text{final}} - V_{\text{initial}} = 1.33\ \text{L} - 0.66\ \text{L} = 0.67\ \text{L}$

   Now we can use the definition of work to determine the work done:

   $w = -P_{\text{ext}}\Delta V = (0.775\ \text{atm})(0.67\ \text{L}) = -0.520\ \text{L atm}$

   Now we construct a conversion factor from the relationship between liter·atmospheres and joules:

   $$-0.52\ \text{L}\cdot\text{atm} \times \frac{101.32\ \text{J}}{1\ \text{L}\cdot\text{atm}} = -53\ \text{J}$$

Test Yourself

1. First we need to determine ΔT. A change is always the final value minus the initial value:
   $\Delta T = 144.5°\text{C} - 17.8°\text{C} = 126.7°\text{C}$

   Now we can use the expression for q, substitute for all variables, and solve for heat:

   $$q = (76.50\ \text{g})\left(0.233\ \frac{\text{J}}{\text{g}\cdot°\text{C}}\right)(126.7°\text{C}) = 2{,}260\ \text{J}$$

Test Yourself

1. We can use the equation for q to solve for mass. Note that ΔT is given directly as 99.9°C and specific heat of gold is 0.129 J/g·°C. Substituting,

   $1,\ 377\ \text{J} = m(0.129\ \text{J}/\text{g}\ °\text{C})(99.9°\text{C})$

   $$m = \frac{1{,}377\ \text{J}}{(0.129\ \text{J}/\text{g}\cdot°\text{C})(99.9°\text{C})}$$

   $m = 107\ \text{g}$

Exercises

1. Work is a force acting through a distance or a volume changing against some pressure.

3. First we need to determine the change in volume, ΔV. A change is always the final value minus the initial value:
   $\Delta V = V_{\text{final}} - V_{\text{initial}} = 12.60\ \text{L} - 3.00\ \text{L} = 9.60\ \text{L}$

   Now we can use the definition of work to determine the work done:

   $w = -P_{\text{ext}}\Delta V = -(0.888\ \text{atm})(9.60\ \text{L}) = -8.53\ \text{L atm}$

   Now we construct a conversion factor from the relationship between liter·atmospheres and joules:

   $$-8.53\ \text{L}\cdot\text{atm} \times \frac{101.32\ \text{J}}{1\ \text{L}\cdot\text{atm}} = -864\ \text{J}$$

5. First we need to determine the change in volume, ΔV. A change is always the final value minus the initial value:

$\Delta V = V_{\text{final}} - V_{\text{initial}} = 0.97\text{ L} - 3.45\text{ L} = -2.48\text{ L}$

Now we can use the definition of work to determine the work done:

$w = -P_{\text{ext}}\Delta V = -(0.985\text{ atm})(-2.48\text{ L}) = 2.44\text{ L atm}$

Now we construct a conversion factor from the relationship between liter·atmospheres and joules:

$$2.44\text{ L}\cdot\text{atm} \times \frac{101.32\text{ J}}{1\text{ L}\cdot\text{atm}} = 248\text{ J}$$

7. When heat is positive, the total energy of the system is increasing.

9. First we need to determine ΔT. A change is always the final value minus the initial value:

$\Delta T = 177.9°\text{C} - 25.6°\text{C} = 152.3°\text{C}$

Heat capacity of Fe is 0.449 J/g·°C. Now we can use the expression for q, substitute for all variables, and solve for heat:

$$q = (55.6\text{ g})\left(0.449\frac{\text{J}}{\text{g}\cdot°\text{C}}\right)(152.3°\text{C}) = 3{,}800\text{ J}$$

11. First we need to determine ΔT. A change is always the final value minus the initial value:

$\Delta T = 298\text{ K} - 355\text{ K} = -57\text{ K or }°\text{C}$

The heat capacity of H_2O is 4.184 J/g·°C. Now we can use the expression for q, substitute for all variables, and solve for heat:

$$q = (245\text{ g})\left(4.184\frac{\text{J}}{\text{g}\cdot°\text{C}}\right)(-57°\text{C}) = -58{,}400\text{ J}$$

13. We can use the equation for q to solve for specific heat. The mass of metal sample is 36.8 g. We need to determine ΔT.

$\Delta T = 98.2°\text{C} - 22.9°\text{C} = 75.3°\text{C}$

Substituting the values in the equation for heat:

$452\text{ J} = (36.8\text{ g})c(75.3°\text{C})$

$$c = \frac{452\text{ J}}{(36.8\text{ g})(75.3°\text{C})}$$

$c = 0.163\text{ J / g °C}$

15. We can use the equation for q to solve for ΔT. The mass of Hg is 56.2 g and specific heat of Hg is 0.139 J/g·°C. Substituting the values in the equation for heat:

$336\text{ J} = (56.2\text{ g})(0.139\text{ J / g °C})\Delta T$

$$\Delta T = \frac{336\text{ J}}{(56.2\text{ g})(0.139\text{ J / g}\cdot°\text{C})}$$

$\Delta T = 43.0°\text{C}$

17. We can use the equation for q to solve for mass of Al. The specific heat of Al is 0.900 J/g·°C. We need to determine ΔT.

$\Delta T = 35.6°\text{C} - 23.5°\text{C} = 12.1°\text{C}$

Substituting the values in the equation for heat:

$187.9\text{ J} = m(0.900\text{ J / g °C})(12.1°\text{C})$

$$m = \frac{187.9\text{ J}}{(0.900\text{ J / g}\cdot°\text{C})(12.1°\text{C})}$$

$m = 17.3\text{ g}$

The molar mass of Al is 26.98 g/mol. Reversing the unit and using it as a conversion factor, we get

$$17.3\text{ g} \times \frac{1\text{ mol}}{26.98\text{ g}} = 0.640\text{ mol}$$

3. ENTHALPY AND CHEMICAL REACTIONS

Test Yourself

1. The thermochemical equation is

   $N_2(g) + O_2(g) \rightarrow 2NO(g)\ \Delta H = 181\ \text{kJ}$

Test Yourself

1. By definition, a chemical reaction that has a positive Δ*H* is endothermic, meaning that this much energy—in this case, 42 kJ—is absorbed by the reaction.

Test Yourself

1. Because we are given Δ*T* directly, we can determine the heat of the reaction, which is equal to Δ*H*:

   $$q = (750.0\ \text{g})\left(4.184\ \frac{\text{J}}{\text{g}\cdot{}^\circ\text{C}}\right)(28.4^\circ\text{C})$$

   Solving for *q*, we get

   $$q = 89,\ 100\ \text{J}$$

   Therefore, $\Delta H = -89{,}100\ \text{J}$.
   According to the stoichiometry of the reaction, exactly 0.10 mol of CH4 will form, so this quantity of heat is for 0.10 mol. For 1 mol of CH4, we need to scale up the heat by a factor of ten:

   $$q = 89,\ 100\ \text{J} \times 10 = 891,\ 000\ \text{J for 1 mol}\ CH_4$$

   On a molar basis, the change in enthalpy is

   $$\Delta H = -891,\ 000\ \text{J} = -891\ \text{kJ}$$

Exercises

1. under conditions of constant pressure
3. The balanced thermochemical equation is

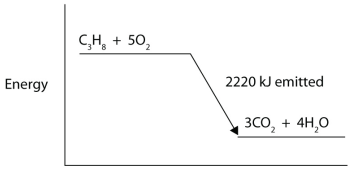
   $H_2(g) + Cl_2(g) \rightarrow 2HCl(g) \qquad \Delta H = -184\ \text{kJ}$

5. The balanced thermochemical equation is

   $N_2(g) + O_2(g) \rightarrow 2NO(g) \qquad \Delta H = 180\ \text{kJ}$

7. By definition, a chemical reaction that has a negative Δ*H* is exothermic, meaning that this much energy—in this case, −184 kJ —is given off by the reaction.
9. By definition, a chemical reaction that absorbs energy is endothermic, meaning that this much energy—in this case, 180 kJ —is positive.

11.



13. The balanced thermochemical reaction is

$$KOH(aq) + HNO_3(aq) \rightarrow KNO_3(aq) + H_2O(l)$$

We assume that heat capacity and density of the solution is same as that of pure water. The amount of solution is 250 mL, and with a density of 1.00 g/mL, we thus have 250 g of solution. Specific heat of water is 4.184 J/g·°C . The change in temperature, ΔT is: 35.9°C – 22.5°C = 13.4°C

Substituting the values in the equation for q which is equal to ΔH:

$$q = (250\ g)\left(4.184\ \frac{J}{g \cdot °C}\right)(13.4°C)$$

$$q = 14,\ 000\ J = 14.0\ kJ$$

Therefore, $\Delta H = -14{,}000\ J = -14.0\ kJ$

According to the stoichiometry of the reaction, exactly 0.25 mol of KNO3 will form, so this quantity of heat is for 0.25 mol. For 1 mol of KNO3, we need to scale up the heat by a factor of four:

$q = 14.0\ kJ \times 4 = 56.0\ kJ$ for 1 mol KNO3

On a molar basis, the change in enthalpy is

$\Delta H = -56.0\ kJ$ for 1 mol of reactants.

15. The balanced thermochemical reaction is

$$C_2H_5OH(g) + 3O2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$$

The specific heat of water is 4.184 J/g·°C . The change in temperature, ΔT is: 65.6°C – 24.6°C = 41.0°C

Substituting the values in the equation for q which is equal to ΔH:

$$q = (400.0\ g)\left(4.184\ \frac{J}{g \cdot °C}\right)(41.0°C)$$

Solving for q, we get

$$q = 68,\ 600\ J = 68.6\ kJ$$

Therefore, $\Delta H = -68{,}600\ J = -68.6\ kJ$

According to the stoichiometry of the reaction, exactly 0.050 mol of C2H5OH is burnt, so this quantity of heat is for 0.050 mol. For 1 mol of C2H5OH, we need to scale up the heat by a factor of 20:

$q = 68.6\ kJ \times 20 = 1{,}370\ kJ$ for 1 mol C2H5OH

On a molar basis, the change in enthalpy is

$\Delta H = -1{,}370$ kJ/ mol of ethanol.

4. STOICHIOMETRY CALCULATIONS USING ENTHALPY

Test Yourself

1. The balanced thermochemical equation relates the energy change to moles, not grams, so we first convert the amount of H_2 to moles and then use the thermochemical equation to determine the energy change:

$$1.00 \text{ g } H_2 \times \frac{1 \text{ mol } H_2}{2.02 \text{ g } H_2} \times \frac{-91.8 \text{ kJ}}{3 \text{ mol } H_2} = -15.1 \text{ kJ}$$

Test Yourself

1. The amount of N_2 reacted when 100.0 kJ of energy is supplied, can be calculated as follows:

$$100.0 \text{ kJ} \times \frac{1 \text{ mol } N_2}{180.6 \text{ kJ}} \times \frac{28.0 \text{ g } N_2}{1 \text{ mol } N_2} = 15.5 \text{ g } N_2$$

Exercises

1. 1 mol of PCl_3 ⇔ 1 mol of Cl_2 ⇔ 1 mol of PCl_5 ⇔ − 87.9 kJ

3. We use the thermochemical equation to determine the energy change:

$$17.8 \text{ mol } CH_4 \times \frac{-890.1 \text{ kJ}}{1 \text{ mol } CH_4} = -15{,}800 \text{ kJ}$$

When 17.8 mol of $CH_4(g)$ react, 15,800 kJ of energy is given off.

5. We use the thermochemical equation to determine the energy change:

$$23.09 \text{ mol } C_6H_6 \times \frac{49.0 \text{ kJ}}{1 \text{ mol } C_6H_6} = 1{,}130 \text{ kJ}$$

When 23.09 mol of $C_6H_6(\ell)$ are formed, 1,130 kJ of energy is absorbed.

7. The balanced thermochemical equation relates the energy change to moles, not grams, so we first convert the amount of $C_6H_{12}O_6$ to moles and then use the thermochemical equation to determine the energy change:

$$100.0 \text{ g } C_6H_{12}O_6 \times \frac{1 \text{ mol } C_6H_{12}O_6}{180.16 \text{ g } C_6H_{12}O_6} \times \frac{-2{,}799 \text{ kJ}}{1 \text{ mol } C_6H_{12}O_6}$$
$$= -1{,}554 \text{ kJ}$$

Thus, 1,554 kJ of energy is given off when 100.0 g of $C_6H_{12}O_6$ react.

9. The balanced thermochemical equation relates the energy change to moles, not grams, so we first convert the amount of CO_2 to moles and then use the thermochemical equation to determine the energy change:

$$85.2 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.0 \text{ g } CO_2} \times \frac{566 \text{ kJ}}{2 \text{ mol } CO_2} = 548 \text{ kJ}$$

Thus, 548 kJ of energy is absorbed when 85.2 g of CO_2 are reacted

11. The amount of $NaHCO_3$ decomposed by 256 kJ of energy can be calculated as follows:

$$256 \text{ kJ} \times \frac{2 \text{ mol } NaHCO_3}{91.5 \text{ kJ}} \times \frac{84.0 \text{ g } NaHCO_3}{1 \text{ mol } NaHCO_3} = 470 \text{ g } NaHCO_3$$

13. The amount of SO_3 needed to generate 1,566 kJ of energy, can be calculated as follows:

$$1{,}566 \text{ kJ} \times \frac{3 \text{ mol } SO_3}{570.2 \text{ kJ}} \times \frac{80.0 \text{ g } SO_3}{1 \text{ mol } SO_3} = 660 \text{ g } SO_3$$

5. HESS'S LAW

Test Yourself

1. We will start by writing chemical reactions that put the correct number of moles of the correct substance on the proper side. We need PbCl2 as the reactant. So, we reverse the first reaction and also change the sign on ΔH.

   $PbCl_2 \rightarrow Pb + Cl_2 \Delta H = 223\ kJ$

   We need Pb and PbCl4 as products. The second reaction has PbCl4 on the proper side, so let us include this reaction as it is.

   $PbCl_2 + Cl2 \rightarrow PbCl_4\ \ \Delta H = -87\ kJ$

   We combine these two reactions, modified as stated:

   $$\begin{array}{ll} PbCl_2 \rightarrow Pb + Cl_2 & \Delta H = +223\ kJ \\ \underline{PbCl_2 + Cl_2 \rightarrow PbCl_4} & \Delta H = -87\ kJ \end{array}$$

   $$PbCl_2 + PbCl_2 + Cl_2 \rightarrow Pb + Cl_2 + PbCl_4 \quad \Delta H = +136\ kJ$$

   Here Cl2 cancels and we are left with 2PbCl2 on reactant side; one Pb and one PbCl4 on product side.

   $2PbCl_2 \rightarrow Pb + PbCl_4$

   ΔH of this reaction is the sum of the two ΔH values:

   $\Delta H = +223 - 87 = +136\ kJ$

Exercises

1. If chemical equations are combined, their energy changes are also combined.
3. Since the equation is reversed the sign on the ΔH must be changed. So,

   $\Delta H = 65.6\ kJ$
5. Since the coefficients of the first equation are doubled, the ΔH value is also doubled. So,

   $\Delta H = 2(-65.6\ kJ) = -131.2\ kJ$
7. Since the first equation is reversed and the coefficients are doubled, the sign on ΔH is changed and its value is doubled.

   So, $\Delta H = 2\ (+44\ kJ) = 88\ kJ$

9. We will start by writing chemical reactions that put the correct number of moles of the correct substance on the proper side. We need Cu and Cl2 as reactants. The first reaction has Cu and Cl2 on the proper side. So, we include the reaction as it is written.

$$2Cu+Cl_2 \rightarrow 2CuCl \quad \Delta H = -274\ \text{kJ}$$

We need CuCl2 as product. The second reaction has CuCl2 on the proper side. So, we include the reaction as it is written.

$$2CuCl+Cl_2 \rightarrow 2CuCl_2 \quad \Delta H = -166\ \text{kJ}$$

We combine these two reactions, modified as stated:

$$\begin{array}{rl} 2Cu + Cl_2 \rightarrow 2CuCl & \Delta H = -274\ \text{kJ} \\ \underline{2CuCl + Cl_2 \rightarrow 2CuCl_2} & \Delta H = -166\ \text{kJ} \\ 2Cu + Cl_2 + 2CuCl + Cl_2 \rightarrow 2CuCl + 2CuCl_2 & \Delta H = -440\ \text{kJ} \end{array}$$

Here 2CuCl cancels and we are left with 2Cu and 2Cl2 on reactant side; 2CuCl2 on product side.

$$2Cu+2Cl_2 \rightarrow 2CuCl_2$$

The required reaction is not yet obtained. Δ*H* of this reaction is the sum of the two Δ*H* values:

$$\Delta H = -274 + (-166) = -440\ \text{kJ}$$

If we divide the coefficients of the above chemical equation by 2, we obtain the required reaction. Also the Δ*H* value should be divided by 2.

$$Cu+Cl_2 \rightarrow CuCl_2$$

Δ*H* of this reaction is:

$$\Delta H = (-440)/2 = -220\ \text{kJ}$$

11. We will start by writing chemical reactions that put the correct number of moles of the correct substance on the proper side. We need Fe2O3 as product. The first reaction has Fe2O3 on the proper side. But, it has two moles of Fe2O3 and the required reaction needs only one mole. So, we write the first equation as given and divide the coefficients by 2; the Δ*H* is also divided by 2.

$$2Fe+3/2\,O_2 \rightarrow Fe_2O_3 \quad \Delta H = -825\ \text{kJ}$$

We need SO3 as the product. The second reaction has SO3 on the proper side. But, it has two moles of SO3 and the required reaction needs three moles. So, we write the second equation as given and multiply the coefficients by 3/2. The Δ*H* is also multiplied by 3/2.

$$3S+9/2\,O_2 \rightarrow 3SO_3 \quad \Delta H = -1,188\ \text{kJ}$$

We need Fe2(SO4)3 as the reactant. So, we reverse the third reaction to get Fe2(SO4)3 on the product side. The sign on Δ*H* should be changed.

$$Fe_2(SO_4)_3 \rightarrow 2Fe + 3S + 6O_2 \Delta H = 2,583\ \text{kJ}$$

We combine these three reactions, modified as stated:

$$\begin{array}{rl} 2Fe + 3/2O_2 \rightarrow Fe_2O_3 & \Delta H = -825\ \text{kJ} \\ 3S + 9/2\ O_2 \rightarrow 3SO_3 & \Delta H = -1,188\ \text{kJ} \\ \underline{Fe_2(SO_4)_3 \rightarrow 2Fe + 3S + 6O_2} & \Delta H = +2,583\ \text{kJ} \\ 2Fe + 3/2O_2 + 3S + 9/2\ O_2 + Fe_2(SO_4)_3 \rightarrow Fe_2O_3 + 3SO_3 + 2Fe + 3S + 6O_2 & \Delta H = +570\ \text{kJ} \end{array}$$

Here 6O2, 3S, and 2Fe are cancelled. We are left with Fe2(SO4)3 on the reactant side; Fe2O3 and 3SO3 on the product side.

$$Fe_2(SO_4)_3 \rightarrow Fe_2O_3 + 3SO_3$$

This is the required reaction. The Δ*H* of this reaction is the sum of the three Δ*H* values:

$$\Delta H = -825 + (-1,188) + 2,583 = 570\ \text{kJ}$$

6. FORMATION REACTIONS

Test Yourself

1. This is not a proper formation reaction because oxygen is not written as a diatomic molecule.

Test Yourself

1. There is one mole of the substance as product; so the coefficients of the reactants may have to be fractional to balance the reaction.

   $$Ca(s) + C(s) + 3/2O_2(g) \rightarrow CaCO_3(s)$$

Test Yourself

1. There will be two formation reactions. The one for the product will be written as a formation reaction, while the one for the reactant will be written in reverse. Furthermore, the formation reaction for SO_3 and SO_2 will be multiplied by two because there are two moles of SO_3 and SO_2 in the balanced chemical equation. We must note that no formation reaction is required for O_2 as it is in its standard state. The formation reactions are as follows:

   $$2 \times \left[SO_2(g) \rightarrow S(s) + O_2(g)\right]$$

   $$2 \times \left[S(s)+\frac{3}{2}O_2(g) \rightarrow SO_3(g)\right]$$

Test Yourself

1. The enthalpies of formation are multiplied by the number of moles of each substance in the chemical equation, and the total enthalpy of formation for reactants is subtracted from the total enthalpy of formation of the products:

   $$\Delta H_{rxn} = [(1\ \text{mol})(-393.5\ \text{kJ/mol}) + (1\ \text{mol})(0\ \text{kJ/mol})]$$
   $$-[(1\ \text{mol})(-110.5\ \text{kJ/mol}) + (1\ \text{mol})(-285.8\ \text{kJ/mol})]$$

   $$\Delta H_{rxn} = +2.8\ \text{kJ}$$

Exercises

1. A formation reaction is a reaction that produces one mole of a substance from its elements. Example: $C(s) + O_2(g) \rightarrow CO_2(g)$
3. a. Formation reaction; a single substance is formed.
   b. It is not the formation of a single substance, so it is not a formation reaction.
5. a. Formation reaction; a single substance is formed.
   b. It is not the formation of a single substance, so it is not a formation reaction.
7. a. $3/2H_2(g) + P(s) + 2O_2(g) \rightarrow H_3PO_4(s)$
   b. $2Na(s) + 1/2O_2(g) \rightarrow Na_2O(s)$
   c. $3C(s) + 1/2O_2(g) + 4H_2(g) \rightarrow C_3H_7OH(\ell)$
9. a. $12C(s) + 11H_2(g) + 11/2O_2(g) \rightarrow C_{12}H_{22}O_{11}(s)$
   b. $Zn(s) + N_2(g) + 3O_2(g) \rightarrow Zn\left(NO_3\right)_2(s)$
   c. $Al(s) + 3/2O_2(g) + 3/2H_2(g) \rightarrow Al(OH)_3(s)$
11. There will be three formation reactions. The one for the products will be written as formation reactions, while the one for the reactant will be written in reverse. The formation reactions are as follows:

    $$MgCO_3(s) \rightarrow Mg(s) + C(s) + 3/2O_2(g)$$

    $$Mg(s) + 1/2O_2(g) \rightarrow MgO(s)$$

    $$C(s) + O_2(g) \rightarrow CO_2(g)$$

13. There will be three formation reactions. The one for the products will be written as formation reactions, while the one for the reactant will be written in reverse. Furthermore, the formation reaction for CuCl will be multiplied by two because there are two moles of CuCl in the balanced chemical equation. The formation reactions are as follows:

$$2 \times \left[\mathrm{CuCl(s)} \rightarrow \mathrm{Cu(s)} + 1 \big/ 2\mathrm{Cl_2(g)}\right]$$

$$\mathrm{Cu(s)} \rightarrow \mathrm{Cu(s)}$$

$$\mathrm{Cu(s)} + \mathrm{Cl_2(g)} \rightarrow \mathrm{CuCl_2(s)}$$

15. The enthalpy of formation for $CH_2O(g)$ is —115.90 kJ/mol, O_2 is 0 kJ/mol, $CO_2(g)$ is −393.51 kJ/mol, and $H_2O(\ell)$ is −285.83 kJ/mol. The enthalpy change for this reaction is

$$\begin{aligned}\Delta H_{rxn} &= [(1\ \mathrm{mol})(-393.51\ \mathrm{kJ/mol}) \\ &\quad +(1\ \mathrm{mol})(-285.83\ \mathrm{kJ/mol})] \\ &\quad -[(1\ \mathrm{mol})(-115.90\ \mathrm{kJ/mol}) + (1\ \mathrm{mol})(0\ \mathrm{kJ/mol})]\end{aligned}$$

$$\Delta H_{\mathrm{rxn}} = -563.44\ \mathrm{kJ}$$

17. The enthalpy of formation for Mg(s) is 0 kJ/mol, $N_2O_5(g)$ is 11.30 kJ/mol, MgO(s) is −601.60 kJ/mol, and $NO_2(g)$ is 33.1 kJ/mol. The enthalpy change for this reaction is

$$\begin{aligned}\Delta H_{\mathrm{rxn}} &= \left[(1\ \mathrm{mol})(-601.60\ \mathrm{kJ/mol}) + (2\ \mathrm{mol})(33.1\ \mathrm{kJ/mol})\right] \\ &\quad - \left[(1\ \mathrm{mol})(0\ \mathrm{kJ/mol}) + (1\ \mathrm{mol})(11.3\ \mathrm{kJ/mol})\right]\end{aligned}$$

$$\Delta H_{\mathrm{rxn}} = -546.7\ \mathrm{kJ}$$

7. END-OF-CHAPTER MATERIAL

Additional Exercises

1. First we need to determine the change in volume, ΔV. A change is always the final value minus the initial value:

$$\begin{aligned}\Delta V &= V_{\text{final}} - V_{\text{initial}}\\ &= 72.0\ \text{mL} - 124\ \text{mL}\\ &= -52\ \text{mL}\\ &= -0.052\ \text{L}\end{aligned}$$

The pressure is given in torr, converting it into atmospheres,

$$822\ \text{torr} \times \frac{1\ \text{atm}}{760\ \text{torr}} = 1.08\ \text{atm}$$

Now we can use the definition of work to determine the work done:

$$w = -P_{\text{ext}}\Delta V = -(1.08\ \text{atm})(-0.052\ \text{L}) = 0.0562\ \text{L atm}$$

Now we construct a conversion factor from the relationship between liter·atmospheres and joules:

$$0.0562\ \text{L} \cdot \text{atm} \times \frac{101.32\ \text{J}}{1\ \text{L} \cdot \text{atm}} = 5.7\ \text{J}$$

3. First we need to convert the work done from joules to liter·atmospheres:

$$156\ \text{J} \times \frac{1\ \text{L} \cdot \text{atm}}{101.32\ \text{J}} = 1.54\ \text{L} \cdot \text{atm}$$

Work is added to gas. Hence, w becomes negative.
The formula for work w will be

$$-w = -P_{\text{ext}}\Delta V$$

$$w = P_{\text{ext}}\Delta V$$

Now we can determine the change in volume, ΔV using the definition of work:

$$w = P_{\text{ext}}\Delta V$$

or, $1.54\ \text{L·atm} = (1.67\ \text{atm})\ \Delta V$
or, $\Delta V = 0.922\ \text{L}$
ΔV is always the final volume minus the initial volume:

$$0.992 = \text{V}_{\text{f}} - 3.77\ \text{L}$$

or, $V_{\text{f}} = 4.69\ \text{L}$

5. We are already given that ΔT= 334°C. Specific heat of Hg is 0.139 J/g·°C. Now we can use the expression for q, substitute for all variables, and solve for heat:

$$\begin{aligned}q &= (1{,}744\ \text{g})\left(0.139\ \frac{\text{J}}{\text{g} \cdot {}^\circ\text{C}}\right)(334^\circ\text{C})\\ &= 80{,}970\ \text{J or } 80.97\ \text{kJ}\end{aligned}$$

7. Specific heat of Al is 0.900 J/g·°C. We can use the expression for q, substitute for all variables, and solve for ΔT:

$$188\ \text{J} = (45.6\ \text{g})\left(0.900\ \frac{\text{J}}{\text{g} \cdot {}^\circ\text{C}}\right)(\Delta T)$$

Solving for ΔT we get:

$$\Delta T = 4.58^\circ\text{C}$$

We know, $\Delta T = T_{\text{f}} - T_{\text{i}}$
Therefore, $T_{\text{f}} = \Delta T + T_{\text{i}} = 4.58^\circ\text{C} + 87.3^\circ\text{C} = 91.9^\circ\text{C}$

9. The balanced chemical equation for this process is:

$$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(s) + 6O_2(g)$$

The enthalpy change for this reaction can be calculated using the enthalpy of formation of reactants and products:

$$\begin{aligned}\Delta H_{rxn} &= [(1\ \text{mol})(-1{,}277\ \text{kJ/mol}) + (6\ \text{mol})(0\ \text{kJ/mol})]\\ &\quad - [(6\ \text{mol})(-393.51\ \text{kJ/mol})\\ &\quad + (6\ \text{mol})(-285.83\ \text{kJ/mol})]\end{aligned}$$

$$\Delta H_{\text{rxn}} = -2{,}799\ \text{kJ}$$

Hence, the balanced thermochemical equation is

$$\begin{aligned}&6CO_2(g) + 6H_2O(\ell)\\ &\quad \rightarrow C_6H_{12}O_6(s) + 6O_2(g) \qquad \Delta H = 2{,}799\ \text{kJ}\end{aligned}$$

11. The balanced chemical equation for this process is:

$2Al_2O_3(s) \rightarrow 4Al(s) + 3O_2(g)$

The enthalpy change for this reaction can be calculated using the enthalpy of formation of reactants and products:

$$\Delta H_{rxn} = [(4\ mol)(0\ kJ/mol) + (3\ mol)(0\ kJ/mol)] - [(2\ mol)(-1,675.7\ kJ/mol)]$$

$$\Delta H_{rxn} = 3,351.4\ kJ$$

Hence, the balanced thermochemical equation is

$$2Al_2O_3(s) \rightarrow 4Al(s) + 3O_2(g) \quad \Delta H = 3,351.4\ kJ$$

13. The balanced equation for the combustion of methane is:

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(\ell)$

$$\begin{aligned}\Delta H_{rxn} &= [(1\ mol)(-393.51\ kJ/mol) \\ &\quad +(2\ mol)(-285.83\ kJ/mol)] \\ &\quad -[(1\ mol)(-74.87\ kJ/mol) + (2\ mol)(0\ kJ/mol)] \\ &= -891\ kJ/mol\end{aligned}$$

The balanced thermochemical equation for the combustion of methane is:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(\ell)\ \Delta H = -891 kJ/mol$$

On combustion of 1 mol of CH_4, 891 kJ of energy is given off. At constant pressure $\Delta H = q$. From the equation for heat, we can calculate the mass of H_2O that can be heated from 22°C to 80°C.

The change in temperature, $\Delta T = 80°C - 22°C = 58°C$

Substituting the values in the equation of heat:

$$891,000\ J = m\left(4.184 \frac{J}{g \cdot °C}\right)(58°C)$$

$$m = 3,670\ g$$

15. The unknown reaction has $PbCl_2$ as a reactant and $PbCl_4$ as the product. So, if we reverse the first equation and add to the third equation, we will obtain the unknown (second) equation.

$$\begin{array}{ll} PbCl_2(s) \rightarrow Pb(s) + Cl_2(g) & \Delta H = 359\ kJ \\ \underline{Pb(s) + 2Cl_2(g) \rightarrow PbCl_4(l)} & \Delta H = -329\ kJ \\ \underline{PbCl_2(s) + Cl_2(g) \rightarrow PbCl_4(l)} & \Delta H = 30\ kJ \end{array}$$

On adding the two equations, Pb and Cl_2 cancel. We are left with $PbCl_2$ and Cl_2 as reactants; $PbCl_4$ as product. Enthalpy of the unknown reaction is the sum of the enthalpies of the above two reactions.

17. The enthalpy for the reaction is the difference between enthalpy of formation of diamond and graphite:

$$\begin{aligned}\Delta H_{rxn} &= [(1\ mol)(1.897\ kJ/mol)] - [(1\ mol)(0\ kJ/mol)] \\ &= 1.897\ kJ\end{aligned}$$

The positive sign on enthalpy of reaction means that it is an endothermic reaction.

CHAPTER 8

Electronic Structure

1. LIGHT

Test Yourself

1. We use the equation that relates the wavelength and frequency of light with its speed. We have

$$3.00 \times 10^8 \text{m/s} = \lambda\left(1.55 \times 10^{10} \text{s}^{-1}\right)$$

$$\lambda = 0.0194 \text{ m}$$

Test Yourself

1. Using the formula for the energy of light, we have

$$4.156 \times 10^{-20} \text{ J} = (6.626 \times 10^{-34} \text{ J} \cdot \text{s})\ \nu$$

Joules are in the numerator and the denominator, so they cancel, leaving us with seconds in the denominator (i.e. per second), the unit of frequency. So

$$\nu = 6.27 \times 10^{13} \text{ s}^{-1}$$

Exercises

1. Light has a wavelength and a frequency.
3. We use the equation that relates the wavelength and frequency of light with its speed. We have

$$3.00 \times 10^8 \text{m/s} = \left(7.33 \times 10^{-5} \text{m}\right)\nu$$

Solving the equation for ν, we get

$$\nu = 4.09 \times 10^{12} \text{ s}^{-1}$$

5. We use the equation that relates the wavelength and frequency of light with its speed. We have

$$3.00 \times 10^8 \text{m/s} = \left(733 \times 10^{-9} \text{ m}\right)\nu$$

Solving the equation for ν, we get

$$\nu = 4.09 \times 10^{14} \text{ s}^{-1}$$

7. We use the equation that relates the wavelength and frequency of light with its speed. We have

$$3.00 \times 10^8 \text{m/s} = \lambda\left(8.19 \times 10^{14} \text{s}^{-1}\right)$$

Solving the equation for λ, we get

$$\lambda = 3.66 \times 10^{-7} \text{ m}$$

9. Hz is nothing but s^{-1}. We use the equation that relates the wavelength and frequency of light with its speed. We have

$$3.00 \times 10^8 \text{m/s} = \lambda\left(1.009 \times 10^6 \text{s}^{-1}\right)$$

Solving the equation for λ, we get

$$\lambda = 297 \text{ m}$$

11. Using the formula for the energy of light, we have

$$E = (6.626 \times 10^{-34} \text{ J} \cdot \text{s})(5.55 \times 10^{13} \text{ s}^{-1})$$

Seconds are in the numerator and the denominator, so they cancel, leaving us with joules, the unit of energy. So

$$E = 3.68 \times 10^{-20} \text{ J}$$

13. We use the equation that relates the wavelength and frequency of light with its speed. We have

$$3.00 \times 10^{8}\text{m/s} = \left(5.88 \times 10^{-4}\ \text{m}\right)\nu$$

Solving the equation for ν, we get

$$\nu = 5.10 \times 10^{11}\ \text{s}^{-1}$$

The energy of a photon with this frequency can be calculated using the formula for the energy of light:

$$E = (6.626 \times 10^{-34}\ \text{J}\cdot\text{s})(5.10 \times 10^{11}\ \text{s}^{-1})$$

Seconds are in the numerator and the denominator, so they cancel, leaving us with joules, the unit of energy. So

$$E = 3.38 \times 10^{-22}\ \text{J}$$

2. QUANTUM NUMBERS FOR ELECTRONS

Test Yourself

1.
    1. The principal quantum number *n* must be an integer, which it is here. The quantum number *l* must be less than *n*, which it is. The m_l quantum number must be between $-l$ and *l*, which it is. But the spin quantum number is 1, which is not allowed. Spin must be either +1/2 or −1/2, so this set of quantum numbers is not allowed.
    2. The principal quantum number *n* must be an integer, which it is here. The quantum number *l* must be less than *n*, which it is. The m_l quantum number must be between $-l$ and *l*, which it is. The spin quantum number is −1/2, which is allowed. Because this set of quantum numbers follows all restrictions, it is possible.

Exercises

1. A continuous spectrum is a range of light frequencies or wavelengths; a line spectrum shows only certain frequencies or wavelengths.

3. Using the equation obeyed by the lines in the spectrum of hydrogen gas:

$$\frac{1}{\lambda} = (109{,}700\ \text{cm}^{-1})\left(\frac{1}{4} - \frac{1}{9}\right)$$

$$\frac{36}{5(109{,}700\ \text{cm}^{-1})} = \lambda$$

$\lambda = 6.56 \times 10^{-5}$ cm or 6.56×10^{-7} m or 656 nm

5. The principal quantum number is restricted to being a positive whole number.

7. The absolute value of m_l must be less than or equal to *l*: $|m_l| \leq l$.

9. *l* can be 0, 1, 2, 3, or 4.

11. m_l can be −3, −2, −1, 0, 1, 2, or 3.

13. An *s* orbital is spherical in shape.

15.
    a. The principal quantum number *n* must be an integer, which it is here ($n = 4$). The quantum number *l* must be less than *n*, which it is ($l = 1 < 4$). The absolute value of the quantum number, $|m_l|$ must be less than or equal to *l*, but here $|m_l| = 2$ which is not less than 1. So, this set of quantum numbers is not allowed.

    b. The principal quantum number *n* must be an integer, which it is here ($n = 2$). The quantum number *l* must be less than *n*, which it is ($l = 0 < 2$). The m_l quantum number must be between $-l$ and *l*, which it is ($m_l = 0$). The spin quantum number is −1/2, which is allowed. Because this set of quantum numbers follows all restrictions, it is allowed.

3. ORGANIZATION OF ELECTRONS IN ATOMS

Test Yourself

1. The first two electrons go into the 1*s* shell-subshell combination. Two additional electrons can go into the 2*s* shell-subshell, but now this subshell is filled with the maximum number of electrons. The $n = 2$ shell has a *p* subshell also, which can hold a maximum of six electrons. So, six additional electrons can go into the 2*p* shell-subshell. Now this subshell is also filled with the maximum number of electrons. We have assigned ten electrons so far and we are left with one electron. There are no more subshells in the $n = 2$ shell, so the last electron should go to $n = 3$ shell. It is therefore assigned to the 3*s* shell-subshell combination of quantum numbers.

Test Yourself

1. The first two electrons occupy the 1*s* subshell. The next two occupy the 2*s* subshell, while the next six electrons occupy the 2*p* subshell. This gives us 10 electrons so far, with 2 electrons left. These two electrons goes into the $n = 3$ shell, *s* subshell. Thus, the electron configuration of Mg is $1s^2 2s^2 2p^6 3s^2$.

Test Yourself

1. We can accommodate the 56 electrons in the subshells in the following order:
   Ba: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2$

Test Yourself

1. With 37 electrons, the electron configuration of Rb is
   Rb: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1$
   The first immediate noble gas is Kr, which has an electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$. Using the electron configuration of Kr to represent the first 36 electrons, the abbreviated electron configuration of Rb is
   Rb: $[\text{Kr}]5s^1$

Exercises

1. The H atom has a single electron which would go into the $n = 1$ shell, which has only a single *s* subshell with one orbital (because m_l can equal only 0). Thus, the electron goes in the 1*s* shell-subshell combination. In the H atom, the electron has $m_l = 0$ (the only possible value) and an m_s of either +1/2 or −1/2. Thus, the possible sets of four quantum numbers for the electron in hydrogen atom are:
   {1, 0, 0, 1/2} and [1, 0, 0, −1/2}
3. Sodium has 11 electrons. The first two electrons occupy the 1*s* subshell. The next two occupy the 2*s* subshell, while the next six electrons occupy the 2*p* subshell. Thus, 1*s*, 2*s*, and 2*p* subshells are completely filled. This gives us 10 electrons so far, with 1 electron left. This last electron goes into the $n = 3$ shell, *s* subshell. This subshell can accommodate one more electron, so 3*s* is partially filled (or unfilled).
5. For $n = 2$, the possible values of *l* are 0 and 1 — *s* and *p* subshells. 2*s* shell-subshell can accommodate 2 electrons and 2*p* shell-subshell combination can accommodate 6 electrons. The maximum number of electrons in the entire $n = 2$ shell is: 2 + 6 = 8 electrons.
7. a. The first two electrons occupy the 1*s* subshell. The next two occupy the 2*s* subshell, while the next six electrons occupy the 2*p* subshell. This gives us 10 electrons so far. The remaining four electrons go to $n = 3$. Two electrons go to 3*s* subshell and the remaining two go to 3*p* subshell. Hence, the electron configuration for Si is: $1s^2 2s^2 2p^6 3s^2 3p^2$
   b. Following the order of filling subshells with electrons, the electron configuration of scandium is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$
9. a. Following the order of filling subshells with electrons, the electron configuration of cadmium with 48 electrons is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10}$
   b. The first two electrons occupy the 1*s* subshell. The next two occupy the 2*s* subshell, while the next six electrons occupy the 2*p* subshell. This gives us 10 electrons so far. The remaining two electrons go to 3s subshell. Hence, the electron configuration for magnesium is: $1s^2 2s^2 2p^6 3s^2$

11. a. The electron configuration for Si is: $1s^2 2s^2 2p^6 3s^2 3p^2$

The first immediate noble gas is Ne, which has an electron configuration of $1s^2 2s^2 2p^6$. Using the electron configuration of Ne to represent the first 10 electrons, the abbreviated electron configuration of Si is: $[\mathrm{Ne}]3s^2 3p^2$

b. The electron configuration for Sc is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$ The first immediate noble gas is Ar, which has an electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^6$. Using the electron configuration of Ar to represent the first 18 electrons, the abbreviated electron configuration of Sc is: $[\mathrm{Ar}]4s^2 3d^1$

13. a. The electron configuration of cadmium with 48 electrons is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10}$

The first immediate noble gas is Kr, which has an electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$. Using the electron configuration of Kr to represent the first 36 electrons, the abbreviated electron configuration of Cd is : $[\mathrm{Kr}]5s^2 4d^{10}$

b. The electron configuration of Mg with 12 electrons is: $1s^2 2s^2 2p^6 3s^2$.

The first immediate noble gas is Ne, which has an electron configuration of $1s^2 2s^2 2p^6$. Using the electron configuration of Ne to represent the first 10 electrons, the abbreviated electron configuration of Mg is: $[\mathrm{Ne}]3s^2$

4. ELECTRONIC STRUCTURE AND THE PERIODIC TABLE

Test Yourself

1. a. Ti is located in the second column of the *d* block. We would expect that its electron configuration should end with d^2. The electron configuration of titanium is $[\mathrm{Ar}]4s^23d^2$
   b. Cl is located in the fifth column of the *p* block. We would expect that its electron configuration should end with p^5. The electron configuration of chlorine is $[\mathrm{Ne}]3s^23p^5$

Exercises

1. the first two columns
3. the *s* block
5. The elements in the second column of the periodic table belong to the second column of the *s* block. So, the valence *s* subshell has two electrons. Hence, the valence shell electron configuration is ns^2.
7. ns^2np^1
9. a. Sr is located in the second column of the *s* block. We would expect that its electron configuration should end with s^2. Hence, strontium's electron configuration is $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^2$.
   b. Sulfur is located in the fourth column of the *p* block. We would expect that its electron configuration should end with p^4. Hence, electron configuration of S is $1s^22s^22p^63s^23p^4$.
11. a. V is located in the third column of the *d* block in the periodic table. We would expect that its electron configuration should end with d^3. Hence, electron configuration of vanadium is $1s^22s^22p^63s^23p^64s^23d^3$.
    b. Ar is located in the sixth column of the periodic table. We would expect that its electron configuration should end with p^6. Hence, electron configuration of argon is $1s^22s^22p^63s^23p^6$.
13. a. Ge is located in the second column of the *p* block. We would expect that its electron configuration should end with p^2. Hence, electron configuration of Ge is $1s^22s^22p^63s^23p^64s^23d^{10}4p^2$.
    b. C is located in the second column of the *p* block. We would expect that its electron configuration should end with p^2. Hence, electron configuration of C is $1s^22s^22p^2$.

5. PERIODIC TRENDS

Test Yourself

1. Br is to the right of Ca on the periodic table, so it is smaller because as you go across the row, the atoms get smaller.

Test Yourself

1. Because C is to the left of F, it is easy to remove an electron from it; thus, C has the lower IE.

Test Yourself

1. As and Br are in the same row on the periodic table, but Br is farther to the right. Therefore, Br will have the highest magnitude of EA.

Exercises

1. The chemical equation for the ionization of sodium atom is:

   $Na(g) \rightarrow Na^{+}(g) + e^{-} \Delta H = IE$ (answers will vary)
3. As you go across, atomic radii decrease; as you go down, atomic radii increase.
5. a. Cs is below Na on the periodic table, so it is larger because as you go down the column, the atoms get larger.
   b. Bi is below N on the periodic table, so it is larger because as you go down the column, the atoms get larger.
7. a. K is to the left of Cl on the periodic table, so it is larger because as you go across the row, the atoms get smaller.
   b. Ba is to the left of Bi on the periodic table, so it is larger because as you go across the row, the atoms get smaller.
9. a. Because S is to the right of Na on the periodic table, it is difficult to remove an electron from it; thus, S has the higher IE.
   b. Because Br is to the right of Ge on the periodic table, it is difficult to remove an electron from it; thus, Br has the higher IE
11. a. Because Li is above Cs on the periodic table, it is difficult to remove an electron from it; thus, Li has the higher IE.
    b. Because O is to above Se on the periodic table, it is difficult to remove an electron from it; thus, O has the higher IE.
13. Each successive IE is larger than the previous. We note that the third IE is much larger than the first and the second IE, so it must correspond to the removal of an electron from the *p* subshell of an inner shell; and the first two electrons must be from the *s* subshell. Two valence *s* electrons indicate that the element should belong to second column. Since it is a third row element, it is Mg.
15. The electron configuration of Ca is [Ar] $4s^2$. There are two valence electrons- located in 4*s* subshell. So, the third electron has to be removed from the *p* subshell of the inner shell, $n = 3$. Hence, the third IE shows a large jump in Ca.
17. a. C and F are in the same row on the periodic table, but F is farther to the right. Therefore, F should have the larger magnitude of EA.
    b. Al and Cl are in the same row on the periodic table, but Cl is farther to the right. Therefore, Cl should have the larger magnitude of EA.

6. END-OF-CHAPTER MATERIAL

Additional Exercises

1. We use the equation that relates the wavelength and frequency of light with its speed. We have

$$3.00 \times 10^8 \text{m/s} = (1.00\ \text{m})\nu$$

$$\nu = 3.00 \times 10^8\ \text{s}^{-1}$$

3. We use the equation that relates the wavelength and frequency of light with its speed. We have

$$3.00 \times 10^8 \text{m/s} = (1.00\ \text{m})\nu$$

$$\nu = 3.00 \times 10^8\ \text{s}^{-1}$$

The energy of a photon with this frequency can be calculated using the formula for the energy of light:

$$E = (6.626 \times 10^{-34}\ \text{J s})(3.00 \times 10^8\ \text{s}^{-1})$$

Seconds are in the numerator and the denominator, so they cancel, leaving us with joules, the unit of energy. So

$$E = 1.99 \times 10^{-25}\ \text{J or } 1.99 \times 10^{-22}\ \text{kj}$$

5. We use the equation that relates the wavelength and frequency of light with its speed. We have

For $\lambda = 400$ nm: $3.00 \times 10^8 \text{m/s} = \left(400 \times 10^{-9}\ \text{m}\right)\nu$

Or, $\nu = 7.5 \times 10^{14}\ \text{s}^{-1}$

And,

$$E = (6.626 \times 10^{-34}\ \text{J s})(7.5 \times 10^{14}\ \text{s}^{-1}) = 4.97 \times 10^{-19}\ \text{J}$$

For $\lambda = 700$ nm: $3.00 \times 10^8 \text{m/s} = \left(700 \times 10^{-9}\ \text{m}\right)\nu$

Or, $\nu = 4.29 \times 10^{14}\ \text{s}^{-1}$

And,

$$E = (6.626 \times 10^{-34}\ \text{J s})(4.29 \times 10^{14}\ \text{s}^{-1}) = 2.84 \times 10^{-19}\ \text{J}$$

Hence, the energy range for visible light photons is 4.97×10^{-19} J to 2.84×10^{-19} J

7. Using the equation obeyed by the lines in the spectrum of hydrogen gas:

When, $n = 7$,

$$\frac{1}{\lambda} = (109{,}700\ \text{cm}^{-1})\left(\frac{1}{4} - \frac{1}{49}\right)$$

$$\frac{196}{45(109{,}700\ \text{cm}^{-1})} = \lambda$$

$\lambda = 3.97 \times 10^{-5}$ cm, or 3.97×10^{-7} m or 397 nm

When, $n = 8$,

$$\frac{1}{\lambda} = (109{,}700\ \text{cm}^{-1})\left(\frac{1}{4} - \frac{1}{64}\right)$$

$$\frac{256}{60(109{,}700\ \text{cm}^{-1})} = \lambda$$

$\lambda = 3.89 \times 10^{-5}$ cm, or 3.89×10^{-7} m or 389 nm

9.

n	l	m_l	m_s
5	0	0	1/2 or −1/2
5	1	−1, 0, 1	1/2 or −1/2
5	2	−2, −1, 0, 1, 2	1/2 or −1/2
5	3	−3, −2, −1, 0, 1, 2, 3	1/2 or −1/2
5	4	−4, −3, −2, −1, 0, 1, 2, 3, 4	1/2 or −1/2

11. a. Since they are *d* block elements, the valence electrons go into *d* subshell. The electron configurations are predicted to end in $3d^1$, $3d^2$, $3d^3$, $3d^4$, $3d^5$, $3d^6$, $3d^7$, $3d^8$, $3d^9$, and $3d^{10}$.

b. Cr and Cu are exceptions.

13. The element 117 would belong to seventh row and next-to-last column on the periodic table. The element 117’s valence shell electron configuration should be $7s^27p^5$.

15. Because P is to the left of O and below O on the periodic table, it is easy to remove an electron from it; thus, O has the higher IE.

17. As is to the left of Cl and below Cl on the periodic table, Cl has the smaller radius. As you go across the row, the atoms get smaller and as you go down the column, the atoms get larger.

19. H has only one IE: $H \rightarrow H^+ + e^-$

21. Na looses and electron from its outermost shell to become Na^+. Hence, Na^+ is smaller.

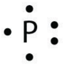
CHAPTER 9

Chemical Bonds

1. LEWIS ELECTRON DOT DIAGRAMS

Test Yourself

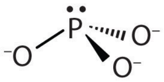
1. 1. The valence electron configuration for phosphorus is $3s^2 3p^3$. So it would have five dots around the symbol for phosphorus, two of them paired to represent the 3*s* electrons:



2. The valence electron configuration for argon is $3s^2 3p^6$. So it would have eight dots around the symbol for argon.

Test Yourself

1. The valence electron configuration for thallium is $6s^2 5d^{10} 6p^1$. In the highest-numbered shell, the $n = 6$ shell, there are three electrons. In Tl^+ ion, one electron is removed from $n = 6$ shell, so only two electrons are remaining in it. Thus, its electron dot diagram is as follows



Exercises

1. The first two electrons in a valence shell are *s* electrons, which are paired.

3. The second column of the periodic table

5. a. The valence electron configuration for strontium is $5s^2$. So there are two electrons in the valence shell. Thus, its electron dot diagram is as follows.

Sr:

b. The valence electron configuration for silicon is $3s^2 3p^2$, so there are four electrons in the valence shell. Thus, its electron dot diagram is as follows.

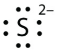
7. a. The valence electron configuration for titanium is $4s^2 3d^2$. In the highest numbered shell, $n = 4$ shell, there are two electrons. Thus, its electron dot diagram is as follows.

Ti:

b. The valence electron configuration for phosphorus is $3s^2 3p^3$. So it would have five dots around the symbol for phosphorus, two of them paired to represent the 3*s* electrons:

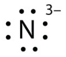
9. a. Having lost its two original valence electrons, the Lewis electron dot diagram is just Mg^{2+}.

b. The S^{2-} ion has gained two electrons in its valence shell, so its Lewis electron dot diagram is as follows:



11. a. The valence electron configuration for iron is $4s^2 3d^6$. In the highest-numbered shell, the $n = 4$ shell, there are two electrons. Having lost the two valence electrons, the electron dot diagram is just Fe^{2+}.

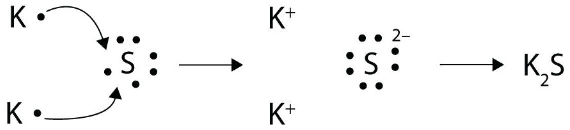
b. Having gained three electrons the Lewis electron dot diagram of N^{3-} is as follows:



2. ELECTRON TRANSFER: IONIC BONDS

Test Yourself

1. The S atom needs two electrons to complete its valence octet, but the K atom supplies only one electron. Thus we need two K atoms to donate the two electrons needed by a sulfur atom. The transfer process looks as follows:



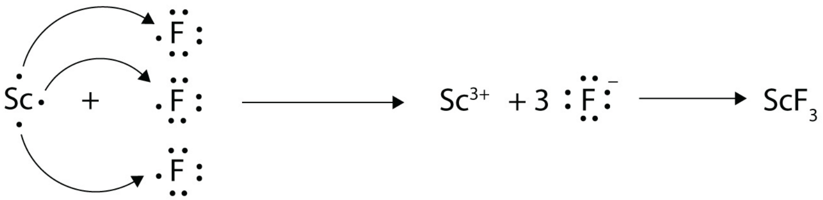
The oppositely charged ions attract each other to make K2S.

Exercises

1. The K^{2+} ion is unlikely to form because the K^+ ion already satisfies the octet rule and is rather stable.
3. The electron configuration of Ba is [Xe]6s^2. It has two electrons in its valence shell, so it has to lose two electrons to have a complete octet in its valence shell.
5. The electron configuration of Se is [Ar] 3d^{10}4$s^2$4p^4. It has six electrons in the highest numbered shell, $n = 4$ shell, so it has to gain two more electrons to have a complete octet in its valence shell.
7. A K atom has one valence electron, while a Cl atom has seven electrons. A Cl atom needs one more to complete its octet, while a K atom has one electron to lose. The transfer process looks as follows:

The two oppositely charged ions attract to form KCl.

9. A scandium atom has three valence electrons, while a fluorine atom has seven electrons. An F atom needs only one more to complete its octet, while an Sc atom has three electrons to lose. Thus we need three F atoms to accept the three electrons from one Sc atom. The transfer process looks as follows:



The oppositely charged ions attract to form ScF3.

11. MgO because the ions have a higher magnitude charge
13. MgO because the ions are smaller

3. COVALENT BONDS

Test Yourself

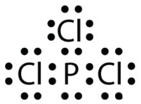
1. Cl atoms have seven electrons in their valence shell:

The two atoms can share their unpaired electrons to make a covalent bond.

Test Yourself

1. The P atom has the following Lewis electron dot diagram:

It has five unpaired electrons, each of which can make a covalent bond by sharing electrons with a Cl atom. The electron dot diagram of PCl3 is as follows:



Test Yourself

1. The central atom is a C atom, with an O atom and an S atom as surrounding atoms. We have a total of 4 + 6 + 6 = 16 valence electrons. Following the rules for Lewis electron dot diagrams for elements gives us

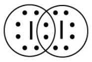
The S and O atoms have complete octets around them, but the C atom has only four electrons around it. The way to solve this dilemma is to make a double bond between carbon and O atom as well as carbon and S atom.

The O atom and S atom still have eight electrons around them, but now the C atom also has a complete octet. This is an acceptable Lewis electron dot diagram for COS.

Exercises

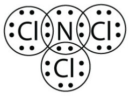
1. two

3. Iodine atoms have seven electrons in their valence shell. The two atoms can share their unpaired electrons to have a complete octet. The Lewis electron dot diagram of I_2 is:



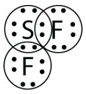
By circling the electrons around each atom, we can now see that both the I atoms complete an octet, so this is an acceptable Lewis electron dot diagram.

5. N atom has five electrons in its valence shell and the chlorine atom has seven electrons. The central atom is the N atom, with three Cl atoms as surrounding atoms. We have a total of 5 + 7 + 7 + 7 = 26 valence electrons. Following the rules for Lewis electron dot diagrams for compounds gives us:

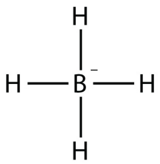


By circling the electrons around each atom, we can now see that the N atom and each Cl atom complete an octet, so this is an acceptable Lewis electron dot diagram.

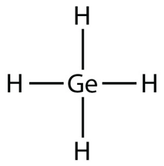
7. a. S atom has six electrons in its valence shell and the F atom has seven electrons. The central atom is the S atom, with two F atoms as surrounding atoms. We have a total of 6 + 7 + 7 = 20 valence electrons. Following the rules for Lewis electron dot diagrams for compounds gives us:



b. B atom has three electrons in its valence shell and the H atom has only one electron. The central atom is the B atom, with four H atoms as surrounding atoms. There is a negative sign on the species, so we have an extra electron to consider. Thus, we have a total of 3 + 1 + 1 + 1 + 1 + 1 = 8 valence electrons. Following the rules for Lewis electron dot diagrams for compounds gives us:



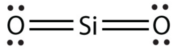
9. a. Ge atom has four electrons in its valence shell and the H atom has only one electron. The central atom is the Ge atom, with four H atoms as surrounding atoms. We have a total of 4 + 1 + 1 + 1 + 1 = 8 valence electrons. Following the rules for Lewis electron dot diagrams for compounds gives us:



b. Cl and F atoms have seven electrons each in their valence shell. The central atom is the Cl atom, with one F atom as the surrounding atom. We have a total of 7 + 7 = 14 valence electrons. Following the rules for Lewis electron dot diagrams for compounds gives us:

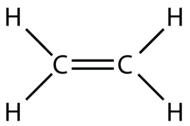
11. a. The central atom is a Si atom, with two O atoms as surrounding atoms. We have a total of 4 + 6 + 6 = 16 valence electrons. Following the rules for Lewis electron dot diagrams for compounds gives us

The O atoms complete the octet around them, but the Si atom has only four electrons around it. The way to solve this dilemma is to make two double bonds between Si and O atoms.



The O atoms still have eight electrons around them, but now the Si atom also completes an octet. This is an acceptable Lewis electron dot diagram for SiO_2.

b. The two carbon atoms are the two central atoms with two H atoms as the surrounding atoms for each C atom. We have a total of 4 + 4 + 1 + 1 +1 + 1 = 12 valence electrons. Each H atom has two electrons around it (its valence shell is full), but each C atom has only seven electrons around it. The way to solve this dilemma is to make a double bond between the two carbon atoms.

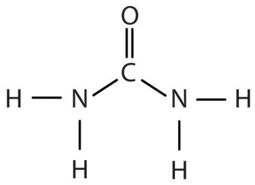


Now the C atoms complete an octet. This is an acceptable Lewis electron dot diagram for C_2H_4.

13. a. The C atom has four valence electrons and the S atoms have six valence electrons each. We have a total of 4 + 6 + 6 = 16 valence electrons. Following the rules for Lewis electron dot diagrams for compounds gives us an octet around S atom but C atom doesn't have an octet around it. The way to solve this dilemma is to make a double bond between the carbon and each of the S atoms. So the Lewis electron dot diagram for CS_2 is:

$$S{=}C{=}S$$

b. We assume that N and C atoms are central atoms. Each N atom has two H atoms and the C atom around it; while the C atom has one O atom and two N atoms around it. Following the rules for Lewis electron dot diagrams for compounds gives us an octet around each N atom and the O atom, and two electrons around each H atom; but the C atom doesn't have an octet around it. The way to solve this dilemma is to make a double bond between the carbon and the O atom. Now the carbon atoms also have an octet around them.



4. OTHER ASPECTS OF COVALENT BONDS

Test Yourself

1. Using Figure 9.4, we can calculate the difference of the electronegativities of the atoms involved in the bond.
    1. For the Rb—F bond, the difference in the electronegativities is 4.0 − 0.8 = 3.2. Thus we predict that this bond will be likely ionic.
    2. For the P—Cl bond, the difference in electronegativities is 3.0 − 2.1 = 0.9, so we predict that this bond will be definitely polar covalent.

Test Yourself

1. Here, we are breaking a C—C triple bond and two H—H single bonds and making a C—C single bond and four C—H single bonds. Bond breaking is endothermic, while bond making is exothermic. For the bond breaking:

1 C≡C:	+837 kJ/mol
2 H—H:	2(+436 kJ/mol)
Total:	+1,709 kJ/mol

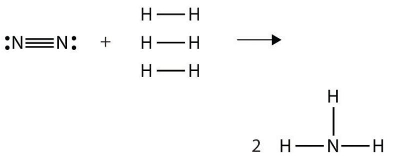
For the bond making:

1 C—C:	−348 kJ/mol
4 C—H:	4(−414 kJ/mol)
Total	−2,004 kJ/mol

Overall, the energy change is +1,709 + (−2,004) = −295 kJ/mol.

Exercises

1. H—H; it is nonpolar because the two atoms have the same electronegativities.
3. a. Since the Cl atom has more electrons in its nucleus than that of an H atom, it attracts the bonding pair of electrons so much more strongly that the electrons remain closer to the Cl atom than to the H atom; so the Cl side of the molecule takes on a partial negative charge.
    b. Since the S atom has more electrons in its nucleus than that of an H atom, it attracts the bonding pair of electrons so much more strongly that the electrons remain closer to the S atom than to the H atom; so the S side of the molecule takes on a partial negative charge.
5. a. For the C—H bond, the difference in the electronegativities is 2.5 − 2.1 = 0.4. Thus we predict that this bond will be slightly polar covalent.
    b. For the C—F bond, the difference in electronegativities is 2.5 − 4.0 = 1.5, so we predict that this bond will be definitely polar covalent.
    c. For the K—F bond, the difference in the electronegativities is 4.0 − 0.8 = 3.2. Thus we predict that this bond will be likely ionic.
7. The bond energy of C—H bond is 414 kJ/mol and that of C—C bond is 348 kJ/mol. So, C—H bond is stronger.
9. We can draw Lewis electron dot diagrams for each substance to see what bonds are broken and what bonds are formed:



Here, we are breaking a N—N triple bond and three H—H single bonds and making six N—H single bonds. Bond breaking is endothermic, while bond making is exothermic. For the bond breaking:

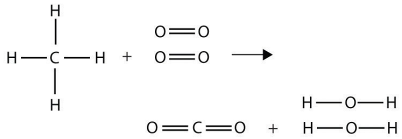
1 N≡N:	+946 kJ/mol
3 H—H:	3(+436 kJ/mol)
Total:	+2,254 kJ/mol

For the bond making:

6 N—H:	6(−389 kJ/mol)
Total	−2,334 kJ/mol

Overall, the energy change is +2,254 + (−2,334) = −80 kJ/mol.

11. We can draw Lewis electron dot diagrams for each substance to see what bonds are broken and what bonds are formed:



Here, we are breaking four C—H single bonds and two O—O double bonds and making two C—O double bonds and four O—H single bonds. Bond breaking is endothermic, while bond making is exothermic. For the bond breaking:

4 C—H:	4(+414 kJ/mol)
2 O=O:	2(+498 kJ/mol)
Total:	+2,652 kJ/mol

For the bond making:

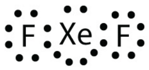
2 C=O:	2(−799 kJ/mol)
4 O—H:	4(−463 kJ/mol)
Total	−3,450 kJ/mol

Overall, the energy change is +2,652 + (−3,450) = −798 kJ/mol.

5. VIOLATIONS OF THE OCTET RULE

Test Yourself

1. With one Xe atom and two F atoms, this molecule has 8 + 7 + 7 = 22 valence electrons. The central Xe atom makes a single bond with each F atom, so it has two bonding electron pairs and three lone electron pairs while each F atom has an octet around it. A Lewis electron dot diagram for this molecule is as follows:

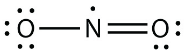


The Xe atom has an expanded valence shell with more than eight electrons around it.

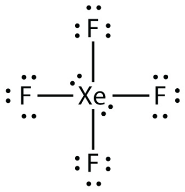
Exercises

1. There is no way all electrons can be paired if there are an odd number of them.

3. a. With one N atom and two O atoms, this molecule has 5 + 6 + 6 = 17 valence electrons, so it is an odd-electron molecule. A Lewis electron dot diagram for this molecule is as follows:

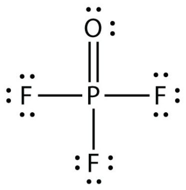


   b. With one Xe atom and four F atoms, this molecule has 8 + 7 + 7 + 7 + 7 = 36 valence electrons. The central Xe atom makes a single bond with each F atom, so it has four bonding electron pairs and two lone electron pairs while each F atom has an octet around it. A Lewis electron dot diagram for this molecule is as follows:



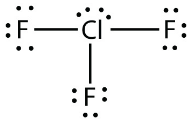
The Xe atom has an expanded valence shell with more than eight electrons around it.

5. a. The central P atom makes a single bond with each F atom and a double bond with the O atom, so it has five bonding electron pairs. A Lewis electron dot diagram for this molecule is as follows:



The P atom has an expanded valence shell with more than eight electrons around it.

   b. The central Cl atom makes a single bond with each F atom, so it has three bonding electron pairs and two lone electron pairs while each F atom has an octet around it. A Lewis electron dot diagram for this molecule is as follows:



The Cl atom has an expanded valence shell with more than eight electrons around it.

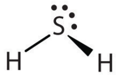
6. MOLECULAR SHAPES

Test Yourself

1. The C atom has four electron groups bonded to four surrounding atoms, so the molecular shape is tetrahedral.

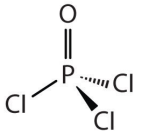
Exercises

1. Electron pairs repel each other.
3. a. For H_2S, the electron dot diagram is as follows:



Although the electron groups are oriented in the shape of a tetrahedron, the shape of the molecule is *bent* or *angular*.

b. For $POCl_3$, the electron dot diagram is as follows:



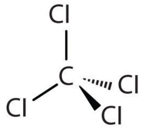
The P atom has four electron groups bonded to four surrounding atoms, so the electron group geometry and the molecular shape is tetrahedral.

5. a. For HCN, the electron dot diagram is as follows:

$$H-C\equiv N:$$

The C atom has two electron groups bonded to two surrounding atoms, so the electron group geometry and the molecular shape is linear.

b. For CCl_4, the electron dot diagram is as follows:



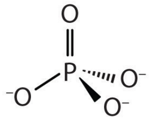
The C atom has four electron groups oriented in the shape of a tetrahedron, bonded to four surrounding atoms, so the electron group geometry and the molecular shape is tetrahedral.

7. a. For CN^-, the electron dot diagram is as follows:

$$:C\equiv N:^-$$

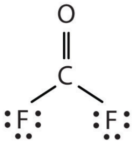
The geometry of CN^- is linear.

b. For PO_4^{3-}, the electron dot diagram is as follows:



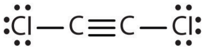
The geometry of the molecule is tetrahedral.

9. a. For COF_2, the electron dot diagram is as follows:



The geometry of the molecule is trigonal planar.

b. For C_2Cl_2, the electron dot diagram is as follows:



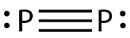
$$:\underset{\cdot\cdot}{\overset{\cdot\cdot}{Cl}}-C\equiv C-\underset{\cdot\cdot}{\overset{\cdot\cdot}{Cl}}:$$

The geometry of the molecule is linear about each central atom(C atom).

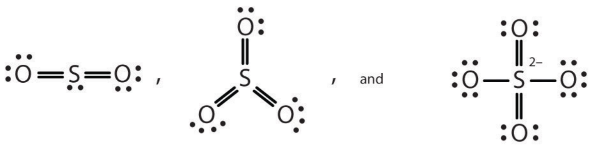
7. END-OF-CHAPTER MATERIAL

Additional Exercises

1. Iron and copper have *d* electrons that typically are not shown on Lewis electron dot diagrams. Other than *d* electrons, they have same configuration, so they have same Lewis electron dot diagram.
3. LiF
5. It would be like N_2:



7. The Lewis electron dot diagrams are as follows:



9. For the O—H bond, the difference in the electronegativities is 3.5 − 2.1 = 1.4.
   For the N—H bond, the difference in the electronegativities is 3.0 − 2.1 = 0.9.
   Since, the difference in the electronegativities of an S—O bond is larger, the S—O bond is more polar.
11. In this reaction, we are breaking eight C—H single bonds, two C—C single bonds, and five O—O double bonds and making six C—O double bonds and eight O—H single bonds. Bond breaking is endothermic, while bond making is exothermic. For the bond breaking:

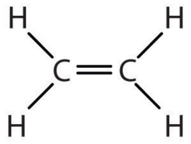
8 C—H:	8(+414 kJ/mol)
2 C—C:	2(+348 kJ/mol)
5 O=O:	5(+498 kJ/mol)
Total:	+6,498 kJ/mol

For the bond making:

6 C=O:	6(−799 kJ/mol)
8 O—H:	8(−463 kJ/mol)
Total	−8,498 kJ/mol

Overall, the energy change is +6,498 + (−8,498) = −2,000 kJ

13. The two carbon atoms are the two central atoms with two H atoms as the surrounding atoms for each C atom. The Lewis electron dot diagram for C_2H_4 is as follows:



Each central C atom has three electron groups around it because the double bond counts as one electron group. The three electron groups repel each other to adopt a trigonal planar shape. So, the molecule is trigonal planar about both central C atoms.

CHAPTER 10

Solids and Liquids

1. INTERMOLECULAR FORCES

Test Yourself

1.
   1. This molecule has an H atom bonded to an F atom, so it will experience hydrogen bonding.
   2. Although this molecule does not experience hydrogen bonding, the Lewis electron dot diagram indicates that it is linear, so it has a permanent dipole. The most significant force in this substance is dipole-dipole interaction.

Exercises

1. dispersion force
3. An H atom must be bonded to an N, O, or F atom.
5.
   a. The most significant intermolecular force for He would be dispersion forces.
   b. The Lewis electron dot diagram and VSEPR indicate that $CHCl_3$ is tetrahedral, but it has one C—H bond which is less polar than the other three C—Cl bonds. So it has a permanent dipole. The most significant force in this substance is dipole-dipole interaction.
   c. This molecule has an H atom bonded to an O atom and F atom, so it will experience hydrogen bonding.
7.
   a. Although this molecule does not experience hydrogen bonding, the Lewis electron dot diagram indicates that it is linear, so it has a permanent dipole. The most significant force in this substance is dipole-dipole interactions.
   b. This molecule has an H atom bonded to an N atom, so it will experience hydrogen bonding.
   c. This molecule is non polar, so the most significant intermolecular force for this substance would be dispersion forces.

2. PHASE TRANSITIONS: MELTING, BOILING, AND SUBLIMING

Test Yourself

1. The ΔH_{fus} of C_6H_6 is 9.95 kJ/mol. However, our quantity is given in units of grams, not moles, so the first step is to convert grams to moles using the molar mass of C_6H_6, which is 78.1 g/mol. Then we can use ΔH_{fus} as a conversion factor. Because the substance is freezing, the process is exothermic, so the energy change will have a negative sign.

   $$108\text{ g }C_6H_6 \times \frac{1\text{ mol }C_6H_6}{78.1\text{ g}} \times \frac{9.95\text{ kJ}}{1\text{ mol}} = -13.8\text{ kJ}$$

Test Yourself

1. The ΔH_{vap} of C_2H_5OH is 38.6 kJ/mol. Boiling is an endothermic process, so the ΔH value will be positive. To determine the magnitude of the energy change, we must first convert the amount of C_2H_5OH to moles. Then we can use ΔH_{vap} as a conversion factor.

   $$822\text{ g }C_2H_5OH \times \frac{1\text{ mol }C_2H_5OH}{46.07\text{ g}} \times \frac{38.6\text{ kJ}}{1\text{ mol}} = 689\text{ kJ}$$

Exercises

1. Melting is the phase change from a solid to a liquid, whereas solidification is the phase change from a liquid to a solid.
3. The molecules have enough energy to move about each other but not enough to completely separate from each other.
5. The ΔH_{fus} of Hg is 2.29 kJ/mol. However, our quantity is given in units of grams, not moles, so the first step is to convert grams to moles using the molar mass of Hg, which is 200.59 g/mol. Then we can use ΔH_{fus} as a conversion factor. Because the substance is melting, the process is endothermic, so the energy change will have a positive sign.

   $$78.0\text{ g Hg} \times \frac{1\text{ mol Hg}}{200.59\text{ g}} \times \frac{2.29\text{ kJ}}{\text{mol}} = 0.890\text{ kJ or }890\text{ J}$$

7. The ΔH_{vap} of Br_2 is 15.4 kJ/mol. Boiling is an endothermic process, so the ΔH value will be positive. To determine the magnitude of the energy change, we must first convert the amount of Br_2 to moles. Then we can use ΔH_{vap} as a conversion factor.

   $$111\text{ g }Br_2 \times \frac{1\text{ mol }Br_2}{159.8\text{ g}} \times \frac{15.4\text{ kJ}}{\text{mol}} = 10.7\text{ kJ}$$

9. a. Temperature does not change during a phase change.
   b. The process of a liquid becoming a gas is called boiling; the process of a solid becoming a gas is called sublimation.
11. Melting is the phase change from a solid to a liquid.

    $Na(s) \rightarrow Na(\ell)$

13. Sublimation is the phase change from a solid to a gas.

    $CO_2(s) \rightarrow CO_2(g)$

15. $$\begin{aligned}\Delta H_{sub} &= \Delta H_{fus} + \Delta H_{vap}\\ &= 6.01 + 40.68\\ &= 46.69\text{ kJ/mol}\end{aligned}$$

3. PROPERTIES OF LIQUIDS

Test Yourself

1. In the Figure 10.8, we extend a line from the point corresponding to 400 torr on the *y*-axis across to the curve for ethanol and then drop it down to the *x*-axis to read the associated temperature. It looks like the point on the ethanol vapor pressure curve corresponds to a temperature of about 65°C, so we conclude that the boiling point of ethanol at 400 torr is approximately 65°C.

Exercises

1. Evaporation occurs when a liquid becomes a gas at temperatures below that liquid's boiling point, whereas boiling is the conversion of a liquid to a gas at the liquid's boiling point.
3. the temperature at which the vapor pressure of a liquid is 760 torr
5. In the Figure 10.8, we extend a line from the point corresponding to 400 torr on the *y*-axis across to the vapor pressure curves. At the lowest temperature the line corresponds to diethylether.
7. In the Figure 10.8, we extend a line from the point corresponding to 200 torr on the *y*-axis across to the curve for ethanol and then drop it down to the *x*-axis to read the associated temperature. The ethanol vapor pressure curve corresponds to a temperature of about 48°C, so we conclude that the boiling point of ethanol at 200 torr is approximately 48°C.
9. Surface tension is an imbalance of attractive forces between liquid molecules at the surface of a liquid.
11. Adhesion must be greater than cohesion.

4. SOLIDS

Test Yourself

1.
    1. I_2 is a covalently bonded molecular compound. In the solid state, it would form molecular crystals.
    2. $Ca(NO_3)_2$ is an ionic compound, so it would exist as ionic crystals in the solid state.

Exercises

1. At the atomic level, a crystalline solid has a regular arrangement of atoms, whereas an amorphous solid has a random arrangement of atoms.
3. The oppositely charged ions are very strongly held together, so ionic crystals have high melting points. Ionic crystals are also brittle because any distortion of the crystal moves same-charged ions closer to each other, so they repel.
5. The covalent network solid is essentially one molecule, making it very hard and giving it a very high melting point.
7. ionic solids, covalent network solids
9. molecular solids
11.
    a. Mercury is a metal, so it would exist as a metallic solid in the solid state.
    b. PH_3 is a covalently bonded molecular compound. In the solid state, it would form a molecular solid.
    c. A combination of a metal and a nonmetal makes an ionic compound, so CaF_2 would exist as ionic crystals in the solid state.
13.
    a. SO_3 is a covalently bonded molecular compound. In the solid state, it would form a molecular solid.
    b. Br_2 is a molecule with covalent bonding. In the solid state, it would form a molecular solid.
    c. Na_2SO_3 is an ionic compound, so it would exist as ionic crystals in the solid state.
15.
    a. H_2S is a covalently bonded molecular compound. In the solid state, it would form a molecular solid.
    b. In the solid state, Si would form a molecular solid.
    c. CsF is an ionic compound, so it would exist as ionic crystals in the solid state.

5. END-OF-CHAPTER MATERIAL

Additional Exercises

1. dispersion forces < dipole-dipole interactions < hydrogen bonding < ionic bonding
3. Yes, but one intermolecular force usually dominates.
5. The ΔH_{fus} of H_2O is 6.01 kJ/mol. Using ΔH_{fus} as a conversion factor,

$$1.55\ \text{kJ} \times \frac{1\ \text{mol}}{6.01\ \text{kJ}} \times \frac{18.0\ \text{g H}_2\text{O}}{1\ \text{mol}} = 4.64\ \text{g H}_2\text{O}$$

7. Water in the vapor phase has no space to evaporate into.
9. liquid krypton because it would have stronger dispersion forces (due to higher molar mass)

CHAPTER 11

Solutions

1. SOME DEFINITIONS

Test Yourself

1. Either by mass or by moles, the obvious minor component is HCl, so it is the solute. CH_3OH—the majority component—is the solvent.

Test Yourself

1. Both C_3H_7OH and H_2O experience hydrogen bonding, so C_3H_7OH is soluble in water.

Exercises

1. The solvent is the majority component of a solution, whereas the solute is the minority component of a solution.
3. By mass the minor component is CO_2, so it is the solute. H_2O —the majority component—is the solvent.
5. The solubility of NaCl is 36.1 g per 100 g of H_2O. Since the solution contains more than 36.1 g of NaCl (45.0 g), the solution is supersaturated.
7. concentrated because there is a lot of solute.
9. AgCl or $CaCO_3$
11. Br_2 is a nonpolar compound. Of the two solvents, $C_6H_6$is nonpolar and CH_3OH is polar, so Br_2 would be expected to be more soluble in C_6H_6.
13. When *n* is small the intermolecular forces due to the polar end —OH dominate, so it is soluble in water which is polar. The nonpolar end dominates intermolecular forces when *n* is large, so it becomes insoluble.

2. QUANTITATIVE UNITS OF CONCENTRATION

Test Yourself

1. To use the definition of molarity, both quantities must be converted to the proper units. First, convert the volume units from milliliters to liters:

$$235\ \text{mL} \times \frac{1\ \text{L}}{1{,}000\ \text{mL}} = 0.235\ \text{L}$$

Now we convert the amount of solute to moles, using the molar mass of $C_6H_{12}O_6$, which is 180.2 g/mol:

$$66.2\ \text{g}\ C_6H_{12}O_6 \times \frac{1\ \text{mol}\ C_6H_{12}O_6}{180.2\ \text{g}\ C_6H_{12}O_6} = 0.367\ \text{mol}\ C_6H_{12}O_6$$

Now we can use the definition of molarity to determine the molar concentration:

$$\text{M} = \frac{0.367\ \text{mol}\ C_6H_{12}O_6}{0.235\ \text{L}} = 1.56\ \text{M}\ C_6H_{12}O_6$$

Test Yourself

1. We know the volume and the molarity; we can use the definition of molarity to mathematically solve for the amount in moles. Converting mL to L:

$$225\ \text{mL} \times \frac{1\ \text{L}}{1{,}000\ \text{mL}} = 0.225\ \text{L}$$

Substituting the quantities into the definition of molarity:

$$1.44\ \text{M} = \frac{\text{mol}\ CaCl_2}{0.225\ \text{L}}$$

$$\text{mol}\ CaCl_2 = (1.44)(0.255\ \text{L}) = 0.324\ \text{mol}$$

Test Yourself

1. Using the definition of molarity, we have

$$0.570\ \text{M} = \frac{0.872\ \text{mol}}{\text{L}}$$

$$\text{L} = \frac{0.872\ \text{mol}}{0.570\ \text{M}}$$

The volume is 1.53 L.

Test Yourself

1. Using the definition of mass percentage, we have

$$\%\ \text{m/m} = \frac{1.67\ \text{g}\ H_2O_2}{55.5\ \text{g sample}} \times 100 = 3.01\%\ H_2O_2$$

Test Yourself

1. Use the definition of parts per million to determine the concentration. Substituting

$$\frac{0.551 \times 10^{-3}\ \text{g As}}{348\ \text{g solution}} \times 1{,}000{,}000 = 1.58\ \text{ppm}$$

Test Yourself

1. First, use the density of H_2O to determine the mass of the sample:

$$3{,}450\ \text{mL} \times \frac{1.00\ \text{g}}{\text{mL}} = 3{,}450\ \text{g}$$

Now we can use the definition of ppm:

$$335.0\ \text{ppm} = \frac{\text{mass solute}}{3{,}450\ \text{g solution}} \times 1{,}000{,}000$$

$$\text{mass solute} = \frac{(335.0\ \text{ppm})(3450\ \text{g solution})}{1{,}000{,}000} = 1.16\ \text{g}$$

Exercises

1. Molarity is moles per liter, whereas molality is moles per kilogram of solvent.

3. To use the definition of molarity, both quantities must be converted to the proper units. First, convert the volume units from milliliters to liters:

$$345\ \text{mL} \times \frac{1\ \text{L}}{1{,}000\ \text{mL}} = 0.345\ \text{L}$$

Now we convert the amount of solute to moles, using the molar mass of $NaNO_3$, which is 85.0 g/mol:

$$13.4\ \text{g NaNO}_3 \times \frac{1\ \text{mol NaNO}_3}{85.0\ \text{g NaNO}_3} = 0.158\ \text{mol NaNO}_3$$

Now we can use the definition of molarity to determine the molar concentration:

$$M = \frac{0.158\ \text{mol NaNO}_3}{0.345\ \text{L}} = 0.458\ \text{M NaNO}_3$$

5. Using the definition of molarity, we have

$$2.55\ \text{M} = \frac{\text{mol MgCl}_2}{0.0331\ \text{L}}$$

$$\text{mol MgCl}_2 = 2.55\ \text{M} \times 0.0331\ \text{L} = 0.0844\ \text{mol}$$

7. Using the definition of molarity, we have

$$0.556\ \text{M} = \frac{0.882\ \text{mol}}{\text{L}}$$

$$\text{L} = \frac{0.882\ \text{mol}}{0.556\ \text{M}} = 1.59\ \text{L}$$

9. First we need to convert the amount of solute to moles, using the molar mass of $Al(NO_3)_3$, which is 213.0 g/mol:

$$26.7\ \text{g Al(NO}_3)_3 \times \frac{1\ \text{mol Al(NO}_3)_3}{213.0\ \text{g Al(NO}_3)_3} = 0.125\ \text{mol Al(NO}_3)_3$$

Using the definition of molarity, we have

$$0.333\ \text{M} = \frac{0.125\ \text{mol}}{\text{L}}$$

$$\text{L} = \frac{0.125\ \text{mol}}{0.333\ \text{M}} = 0.375\ \text{L}$$

11. In the given solution, the concentration of Mg^{2+} ion is 0.66 M and that of NO_3^- is 1.32 M because there is one Mg^{2+} ion and two NO_3^- ions per formula unit of the salt. The total ion concentration is the sum of the individual ion concentrations.

    Total ion concentration = 0.66 M + 1.32 M = 1.98 M

13. In the given solution, the concentration of $C_2H_3O_2^-$ is twice the concentration of Ca^{2+} because there is one Ca^{2+} ion and two $C_2H_3O_2^-$ ions per formula unit of the salt. So, the concentration of Ca^{2+} is 0.277 M which is half the concentration of $C_2H_3O_2^-$.

3. DILUTIONS AND CONCENTRATIONS

Test Yourself

1. Using the dilution equation, we have

   $(0.885\ \mathrm{M})(76.5\ \mathrm{mL}) = (0.500\ \mathrm{M})V_2$

   $V_2 = 135.4\ \mathrm{mL}$

Exercises

1. Dilution is a decrease in a solution's concentration, whereas concentration is an increase in a solution's concentration.
3. Using the dilution equation, we have

   $(1.88\ \mathrm{M})(3.45\ \mathrm{mL}) = M_2(134\ \mathrm{mL})$

   $M_2 = 0.484\ \mathrm{M}$

5. Using the dilution equation, we have

   $(2.25\ \mathrm{M})(1.00\ \mathrm{mL}) = (1.00\ \mathrm{M})V_2$

   $V_2 = 2.25\ \mathrm{mL}$

7. Using the dilution equation, we have

   $(0.875\ \mathrm{M})(665\ \mathrm{mL}) = (1.45\ \mathrm{M})V_2$

   $V_2 = 401\ \mathrm{mL}$

9. Using the dilution equation, we have:

   $(0.332\ \mathrm{M})(75.0\ \mathrm{mL}) = (0.250\ \mathrm{M})V_2$

   $V_2 = 99.6\ \mathrm{mL}$

   In going from 75.0 mL to 99.6 mL, 99.6 − 75.0 = 24.6 mL of solvent must be added.

4. CONCENTRATIONS AS CONVERSION FACTORS

Test Yourself

1. This is a one-step conversion, but the concentration must be written as the reciprocal for the units to work out:

$$0.0773 \text{ mol } CH_2O \times \frac{1 \text{ L}}{0.0444 \text{ mol}} = 1.74 \text{ L of solution}$$

Test Yourself

1. This is a two-step conversion, first using concentration as a conversion factor to determine the number of moles and then the molar mass of H_2SO_4 (98.1 g/mol) to convert to mass:

$$1.08 \text{ L} \times \frac{0.0578 \text{ mol } H_2SO_4}{1 \text{ L solution}} \times \frac{98.1 \text{ g } H_2SO_4}{1 \text{ mol } H_2SO_4} = 6.12 \text{ g } H_2SO_4$$

Test Yourself

1. Here, we must first convert the mass of $H_2C_2O_4$ to moles before using the balanced chemical equation and then the definition of molarity as a conversion factor:

$$17.9 \text{ g } H_2C_2O_4 \times \frac{1 \text{ mol } H_2C_2O_4}{90.0 \text{ g } H_2C_2O_4} \times \frac{2 \text{ mol NaOH}}{1 \text{ mol } H_2C_2O_4} \times \frac{1 \text{ L solution}}{0.512 \text{ mol NaOH}} = 0.777 \text{ L}$$

Test Yourself

1. First we need to determine the number of moles of KOH that reacted. We will convert the volume to liters and then use the concentration of the solution as a conversion factor:

$$54.06 \text{ mL} \times \frac{1 \text{ L}}{1{,}000 \text{ L}} \times \frac{0.0987 \text{ mol KOH}}{\text{L}} = 0.00534 \text{ mol KOH}$$

Now we will use the balanced chemical equation to determine the number of moles of H_3PO_4 (aq) that were present in the initial aliquot:

$$0.00534 \text{ mol KOH} \times \frac{1 \text{ mol } H_3PO_4}{3 \text{ mol KOH}} = 0.00178 \text{ mol } H_3PO_4$$

Then we determine the concentration of H_3PO_4 in the original solution. Converting 25.00 mL into liters (0.02500 L), we use the definition of molarity directly:

$$M = \frac{\text{mol}}{\text{L}} = \frac{0.00178 \text{ mol } H_3PO_4}{0.02500 \text{ L}} = 0.0712 \text{ M } H_3PO_4$$

Test Yourself

1. We are given the mass of solution and we need to first find the amount of H_2O_2 that reacted. Then use the balanced chemical equation to find the number of moles of O_2 produced and finally convert the number of moles into mass.

$$258 \text{ g solution} \times \frac{3 \text{ g } H_2O_2}{100 \text{ g solution}} \times \frac{1 \text{ mol } H_2O_2}{34.02 \text{ g } H_2O_2} \times \frac{5 \text{ mol } O_2}{5 \text{ mol } H_2O_2} \times \frac{32.0 \text{ g } O_2}{1 \text{ mol } O_2} = 7.28 \text{ g } O_2$$

Exercises

1. Using concentration as a conversion factor,

$$3.44\ \text{L} \times \frac{0.753\ \text{mol CaCl}_2}{1\ \text{L}} = 2.59\ \text{mol CaCl}_2$$

3. This is a one-step conversion, but the concentration must be written as the reciprocal for the units to work out:

$$0.822\ \text{mol} \times \frac{1\ \text{L}}{0.665\ \text{mol}} = 1.24\ \text{L of solution}$$

5. This is a two-step conversion, first using concentration as a conversion factor to determine the number of moles and then the molar mass of $CoCl_2$ (129.83 g/mol) to convert to mass (we must remember to take volume in liters):

$$0.0245\ \text{L} \times \frac{0.755\ \text{mol CoCl}_2}{1\ \text{L solution}} \times \frac{129.83\ \text{g CoCl}_2}{1\ \text{mol CoCl}_2} = 2.40\ \text{g CoCl}_2$$

7. This is a two-step conversion, first using the molar mass of NiF_2 (96.69 g/mol) to convert to moles and then the concentration as a conversion factor to determine the volume of the solution:

$$9.04\ \text{g NiF}_2 \times \frac{1\ \text{mol NiF}_2}{96.69\ \text{g NiF}_2} \times \frac{1\ \text{L solution}}{0.332\ \text{mol NiF}_2}$$

$$= 0.282\ \text{L solution}$$

9. Here, we must first use the balanced chemical equation and then the definition of molarity as a conversion factor:

$$5.33\ \text{mol CaCO}_3 \times \frac{2\ \text{mol HCl}}{1\ \text{mol CaCO}_3} \times \frac{1\ \text{L solution}}{3.44\ \text{mol HCl}}$$

$$= 3.10\ \text{L solution}$$

11. First we need to determine the number of moles of NaOH that reacted. We will convert the volume to liters and then use the concentration of the solution as a conversion factor:

$$26.7\ \text{mL} \times \frac{1\ \text{L}}{1{,}000\ \text{mL}} \times \frac{0.554\ \text{mol NaOH}}{1\ \text{L}} = 0.0148\ \text{mol NaOH}$$

Now we will use the balanced chemical equation to determine the number of moles of H_2SO_4 (aq) that were present in the initial aliquot:

$$0.0148\ \text{mol NaOH} \times \frac{1\ \text{mol H}_2\text{SO}_4}{2\ \text{mol NaOH}} = 0.0074\ \text{mol H}_2\text{SO}_4$$

Then we determine the volume of the original solution, we use the definition of molarity directly:

$$0.905\ \text{M} = \frac{0.0074\ \text{mol H}_2\text{SO}_4}{\text{Volume of the solution}}$$

Volume of the solution = 0.00817 L = 8.17 mL

13. The balanced chemical equation is

$$2\text{HCl} + \text{Ca(OH)}_2 \rightarrow \text{CaCl}_2 + 2\text{H}_2\text{O}$$

First we need to determine the number of moles of HCl that reacted. We will convert the volume to liters and then use the concentration of HCl as a conversion factor:

$$23.77\ \text{mL} \times \frac{1\ \text{L}}{1{,}000\ \text{mL}} \times \frac{0.1505\ \text{mol HCl}}{\text{L}} = 0.003577\ \text{mol HCl}$$

Now we will use the balanced chemical equation to determine the number of moles of $Ca(OH)_2$ that were present in the initial aliquot:

$$0.003577\ \text{mol HCl} \times \frac{1\ \text{mol Ca(OH)}_2}{2\ \text{mol HCl}} = 0.001789\ \text{mol Ca(OH)}_2$$

Then we determine the concentration of the original solution. We will convert the volume to liters and then use the definition of molarity directly:

$$\text{M} = \frac{0.001789\ \text{mol Ca(OH)}_2}{0.01500\ \text{L}}$$

Concentration of the original solution = 0.1193 M

15. First we need to determine the number of moles of HNO_3 required for dissolving the given sample of copper. We will convert the volume to liters and then use the concentration of HNO_3 as a conversion factor:

$$4.667\ \text{mL} \times \frac{1\ \text{L}}{1{,}000\ \text{mL}} \times \frac{0.0997\ \text{mol HNO}_3}{1\ \text{L}} = 0.000465\ \text{mol HNO}_3$$

Now we will use the balanced chemical equation to determine the number of moles of Cu that can be dissolved in the given solution, and then convert it into mass using molar mass of Cu(63.55 g/mol) as a conversion factor:

$$0.000465\ \text{mol HNO}_3 \times \frac{1\ \text{mol Cu}}{4\ \text{mol HNO}_3} \times \frac{63.55\ \text{g Cu}}{1\ \text{mol Cu}}$$

$$= 0.00739\ \text{g Cu}$$

$$= 7.39\ \text{mg Cu}$$

17. We are given the amount of product formed in grams, so first we need to convert it to moles and then use the balanced equation to determine the amount of H_2O_2 reacted. Finally using mass percentage as the conversion factor, we need to determine the amount of solution needed.

$$66.3\ \text{g O}_2 \times \frac{1\ \text{mol O}_2}{32.0\ \text{g O}_2} \times \frac{2\ \text{mol H}_2\text{O}_2}{1\ \text{mol O}_2} \times \frac{34.02\ \text{g H}_2\text{O}_2}{1\ \text{mol H}_2\text{O}_2}$$

$$\times \frac{100\ \text{g solution}}{3\ \text{g H}_2\text{O}_2} = 4{,}700\ \text{g or } 4.70\ \text{kg solution}$$

5. COLLIGATIVE PROPERTIES OF SOLUTIONS

Test Yourself

1. We need to determine the number of moles of each substance, add them together to get the total number of moles, and then divide to determine the mole fraction of CH_3OH. The number of moles of CH_3OH is as follows:

$$33.8\text{ g }CH_3OH \times \frac{1\text{ mol }CH_3OH}{32.0\text{ g }CH_3OH} = 1.06\text{ mol }CH_3OH$$

The number of moles of H_2O is as follows:

$$50.0\text{ g }H_2O \times \frac{1\text{ mol }H_2O}{18.0\text{ g }H_2O} = 2.78\text{ mol }H_2O$$

The total number of moles is

1.06 mol + 2.78 mol = 3.84 mol

Now we can calculate the mole fraction of CH_3OH:

$$\chi_{CH_3OH} = \frac{1.06\text{ mol}}{3.84\text{ mol}} = 0.276$$

Test Yourself

1. Here we need the mole fraction of H_2O. The number of moles of H_2O is as follows:

$$50.0\text{ g }H_2O \times \frac{1\text{ mol }H_2O}{18.0\text{ g }H_2O} = 2.78\text{ mol }H_2O$$

The number of moles of $C_6H_{12}O_6$ is as follows:

$$33.8\text{ g }C_6H_{12}O_6 \times \frac{1\text{ mol }C_6H_{12}O_6}{180.2\text{ g }C_6H_{12}O_6} = 0.188\text{ mol }C_6H_{12}O_6$$

So the total number of moles is

2.78 mol + 0.188 mol = 2.968 mol

Now we can calculate the mole fraction of C_6H_6:

$$\chi_{H_2O} = \frac{2.78\text{ mol}}{2.968\text{ mol}} = 0.937$$

Now we can use Raoult's law to determine the vapor pressure in equilibrium with the solution:

$P_{soln} = (0.937)(25.7\text{ torr}) = 24.1\text{ torr}$

Test Yourself

1. Using the equation for the boiling point elevation,

$\Delta T_b = (6.95\ m)(0.512°C/m) = 3.56°C$

We have calculated the change in the boiling point temperature, not the final boiling point temperature. If the boiling point goes up by 3.56°C, we need to add this to the normal boiling point of H_2O to get the new boiling point of the solution:

$T_{BP} = 100.0°C + 3.56°C = 103.6°C$

Test Yourself

1. We use the equation to calculate the change in the freezing point and then subtract this number from the normal freezing point of $C_{10}H_8$to get the freezing point of the solution:

   $\Delta T_f = (3.05\ m)(6.8°C/m) = 20.74°C$

   Now we subtract this number from the normal freezing point of $C_{10}H_8$, which is 80.2°C:

   $80.2 - 20.74 = 59.5°C$

   The freezing point of the solution is 59.5°C.

Test Yourself

1. First we need to convert our temperature to kelvins:

$$T = 55 + 273 = 328\ \text{K}$$

$$\Pi = \left(0.0522\ \frac{\text{mol}}{1\ \text{L}}\right)\left(0.08205\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(328\ \text{K})$$

$$\Pi = 1.40\ \text{atm}$$

Exercises

1. boiling point elevation, freezing point depression, vapor pressure depression
3. We need to determine the number of moles of each substance, add them together to get the total number of moles, and then divide to determine the mole fraction of each substance. The number of moles of C_6H_6 is as follows:

$$45.0\ \text{g}\ C_6H_6 \times \frac{1\ \text{mol}\ C_6H_6}{78.1\ \text{g}\ C_6H_6} = 0.576\ \text{mol}\ C_6H_6$$

The number of moles of $C_6H_5CH_3$ is as follows:

$$60.0\ \text{g}\ C_6H_5CH_3 \times \frac{1\ \text{mol}\ C_6H_5CH_3}{92.1\ \text{g}\ C_6H_5CH_3} = 0.651\ \text{mol}\ C_6H_5CH_3$$

The total number of moles is

$0.576\ \text{mol} + 0.651\ \text{mol} = 1.227\ \text{mol}$

Now we can calculate the mole fraction of C_6H_6:

$$\chi_{C_6H_6} = \frac{0.576\ \text{mol}}{1.227\ \text{mol}} = 0.469$$

Next we can calculate the mole fraction of $C_6H_5CH_3$:

$$\chi_{C_6H_5CH_3} = \frac{0.651\ \text{mol}}{1.227\ \text{mol}} = 0.531$$

5. We need to determine the number of moles of each substance, add them together to get the total number of moles, and then divide to determine the mole fraction of each substance. The number of moles of NaCl is as follows:

$$36.5\ \text{g NaCl} \times \frac{1\ \text{mol NaCl}}{58.5\ \text{g NaCl}} = 0.623\ \text{mol NaCl}$$

The number of moles of O_2 is as follows:

$$63.5\ \text{g}\ H_2O \times \frac{1\ \text{mol}\ H_2O}{18.0\ \text{g}\ H_2O} = 3.53\ \text{mol}\ H_2O$$

The total number of moles is

$0.623\ \text{mol} + 3.53\ \text{mol} = 4.153\ \text{mol}$

Now we can calculate the mole fraction of NaCl:

$$\chi_{NaCl} = \frac{0.623\ \text{mol}}{4.153\ \text{mol}} = 0.150$$

Next we can calculate the mole fraction of H_2O:

$$\chi_{H_2O} = \frac{3.53\ \text{mol}}{4.153\ \text{mol}} = 0.850$$

7. Here we need the mole fraction of C_6H_6. The number of moles of C_6H_6 is as follows:

$$45.0\text{ g }C_6H_6 \times \frac{1\text{ mol }C_6H_6}{78.1\text{ g }C_6H_6} = 0.576\text{ mol }C_6H_6$$

The number of moles of $C_{10}H_8$ is as follows:

$$12.0\text{ g }C_{10}H_8 \times \frac{1\text{ mol }C_{10}H_8}{128.2\text{ g }C_{10}H_8} = 0.0936\text{ mol }C_{10}H_8$$

So the total number of moles is

0.576 mol + 0.0936 mol = 0.6696 mol

Now we can calculate the mole fraction of C_6H_6:

$$\chi_{C_6H_6} = \frac{0.576\text{ mol}}{0.6696\text{ mol}} = 0.860$$

Now we can use Raoult's law to determine the vapor pressure in equilibrium with the solution:

$P_{soln} = (0.860)(76.5\text{ torr}) = 65.8\text{ torr}$

9. Here we need the mole fraction of Hg. The number of moles of Hg is as follows:

$$3.77\text{ g Hg} \times \frac{1\text{ mol Hg}}{200.59\text{ g Hg}} = 0.0188\text{ mol Hg}$$

The number of moles of Au is as follows:

$$0.775\text{ g Au} \times \frac{1\text{ mol Au}}{196.97\text{ g Au}} = 0.00394\text{ mol Au}$$

So the total number of moles is

0.0188 mol + 0.00394 mol = 0.0227 mol

Now we can calculate the mole fraction of Hg:

$$\chi_{Hg} = \frac{0.0188\text{ mol}}{0.0227\text{ mol}} = 0.828$$

Now we can use Raoult's law to determine the vapor pressure in equilibrium with the solution:

$P_{soln} = (0.828)(32.97\text{ torr}) = 27.26\text{ torr}$

11. First we need to convert the mass of $C_6H_{12}O_6$ into number of moles using the reverse of the unit of its molar mass (180.2 g/mol) as the conversion factor.

$$25.0\text{ g }C_6H_{12}O_6 \times \frac{1\text{ mol }C_6H_{12}O_6}{180.2\text{ g }C_6H_{12}O_6} = 0.139\text{ mol }C_6H_{12}O_6$$

Using the definition of molality:

$$\text{molality} = \frac{0.139\text{ mol }C_6H_{12}O_6}{0.1000\text{ kg }H_2O} = 1.39\ m$$

Using the equation for the boiling point elevation,

$\Delta T_b = (1.39\ m)(0.512°C/m) = 0.712°C$

We have calculated the change in the boiling point temperature, not the final boiling point temperature. If the boiling point goes up by 0.712°C, we need to add this to the normal boiling point of H_2O to get the new boiling point of the solution:

$T_{BP} = 100.0°C + 0.712°C = 100.71°C$

13. First we need to determine the mass of solvent. We can use the molar mass of CCl_4 (153.8 g/mol) as the conversion factor.

$$2\text{ mol }CCl_4 \times \frac{153.8\text{ g }CCl_4}{1\text{ mol }CCl_4} = 307.6\text{ g }CCl_4$$

Using the definition of molality:

$$\text{molality} = \frac{1\text{ mol }CBr_4}{0.3076\text{ kg }CCl_4} = 3.25\ m$$

Using the equation for the boiling point elevation,

$\Delta T_b = (3.25\ m)(4.95°C/m) = 16.1°C$

We have calculated the change in the boiling point temperature, not the final boiling point temperature. If the boiling point goes up by 16.1°C, we need to add this to the normal boiling point of CCl_4 to get the new boiling point of the solution:

$T_{BP} = 76.8°C + 16.1°C = 92.9°C$

15. First we need to convert the mass of $C_{10}H_{16}O$ into number of moles using the reverse of the unit of its molar mass (152.24 g/mol) as the conversion factor.

$$123.0\text{ g }C_{10}H_{16}O \times \frac{1\text{ mol }C_{10}H_{16}O}{152.24\text{ g }C_{10}H_{16}O} = 0.808\text{ mol }C_{10}H_{16}O$$

Using the definition of molality:

$$\text{molality} = \frac{0.808\text{ mol }C_{10}H_{16}O}{0.355\text{ kg }C_6H_6} = 2.28\ m$$

We use the equation to calculate the change in the freezing point and then subtract this number from the normal freezing point of C_6H_6 to get the freezing point of the solution:

$\Delta T_f = (2.28\ m)(4.90°C/m) = 11.2°C$

Now we subtract this number from the normal freezing point of C_6H_6, which is 5.51°C:

$5.51 - 11.2 = -5.69°C$

The freezing point of the solution is − 5.69°C.

17. First we need to convert the mass of $C_8H_{17}OH$ into number of moles using the reverse of the unit of its molar mass (130.2 g/mol) as the conversion factor.

$$7.22\text{ g }C_6H_{17}OH \times \frac{1\text{ mol }C_6H_{17}OH}{130.2\text{ g }C_6H_{17}OH} = 0.0554\text{ mol }C_6H_{17}OH$$

Using the definition of molality:

$$\text{molality} = \frac{0.0554\text{ mol }C_6H_{17}OH}{0.0453\text{ kg }C_6H_{12}} = 1.22\ m$$

We use the equation to calculate the change in the freezing point and then subtract this number from the normal freezing point of C_6H_{12} to get the freezing point of the solution:

$\Delta T_f = (1.22\ m)(20.2°C/m) = 24.6°C$

Now we subtract this number from the normal freezing point of C_6H_{12} which is 6.4°C:

$6.4°C - 24.6°C = -18.2°C$

The freezing point of the solution is − 18.2°C.

19. First we need to convert our temperature to kelvins:

$T = 25 + 273 = 298\ \text{K}$

We need to determine the number of moles of $C_6H_{12}O_6$ dissolved. We can use the reverse of the unit of its molar mass (180.2 g/mol) as the conversion factor.

$$25.0\ \text{g}\ C_6H_{12}O_6 \times \frac{1\ \text{mol}\ C_6H_{12}O_6}{180.2\ \text{g}\ C_6H_{12}O_6} = 0.139\ \text{mol}\ C_6H_{12}O_6$$

We need to calculate the molarity of the solution. Using the equation of molarity directly:

$$\text{molarity} = \frac{0.139\ \text{mol}}{0.100\ \text{L}} = 1.39\ \text{M}$$

Now we can substitute into the equation for osmotic pressure, recalling the value for *R*:

$$\Pi = \left(1.39\ \frac{\text{mol}}{\text{L}}\right)\left(0.08205\ \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(298\ \text{K})$$

$\Pi = 33.9\ \text{atm}$

21. We substitute the given values into the equation for osmotic pressure, recalling the value for *R*:

$$1.00\ \text{atm} = \text{M}\left(0.08205\ \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(298\ \text{K})$$

Molarity = 0.0409 M

6. COLLIGATIVE PROPERTIES OF IONIC SOLUTES

Test Yourself

1. When $Fe(NO_3)_3$ dissolves, it separates into one Fe^{3+} ion and three NO_3^- ions:
   $Fe(NO_3)_3 \rightarrow Fe^{3+}(aq) + 3NO_3^-(aq)$
   Because it breaks up into four ions, its van't Hoff factor is 4.

Test Yourself

1. For $CaCl_2$, we need to remember to include the van't Hoff factor, which is 3.
   $\Delta T_b = (3)(0.887\ m)(0.512°C/m) = 1.36°C$
   This represents the change in the boiling point, which is increasing. So we have to add this change to the normal boiling point of water, 100.0°C:
   $100.0 + 1.36 = 101.36°C$

Exercises

1. Ionic solutes separate into more than one particle when they dissolve, whereas molecular solutes do not.
3. First we need to convert the mass of $NaNO_3$ into number of moles using the reverse of the unit of its molar mass (85.0 g/mol) as the conversion factor.

   $$15.6\text{ g NaNO}_3 \times \frac{1\text{ mol NaNO}_3}{85.0\text{ g NaNO}_3} = 0.184\text{ mol NaNO}_3$$

   Using the definition of molality:

   $$\text{molality} = \frac{0.184\text{ mol solute}}{0.1\text{ kg solvent}} = 1.84\ m$$

   $NaNO_3$ is made up of two ions, so the van't Hoff factor is 2. Using the equation for the boiling point elevation,
   $\Delta T_b = 2(1.84\ m)(0.512°C/m) = 1.88°C$
   We have calculated the change in the boiling point temperature, not the final boiling point temperature. If the boiling point goes up by 1.88°C, we need to add this to the normal boiling point of H_2O to get the new boiling point of the solution:
   $T_{BP} = 100.0°C + 1.88°C = 101.9°C$
5. First we need to convert the mass of $CaCl_2$ into number of moles using the reverse of the unit of its molar mass (111 g/mol) as the conversion factor.

   $$345\text{ g CaCl}_2 \times \frac{1\text{ mol CaCl}_2}{111\text{ g CaCl}_2} = 3.11\text{ mol CaCl}_2$$

   Using the definition of molality:

   $$\text{molality} = \frac{3.11\text{ mol solute}}{1.550\text{ kg solvent}} = 2.01\ m$$

   $CaCl_2$ is made up of three ions, so the van't Hoff factor is 3. We use the equation to calculate the change in the freezing point and then subtract this number from the normal freezing point of H_2O to get the freezing point of the solution:
   $\Delta T_f = 3(2.01\ m)(1.86°C/m) = 11.2°C$
   Now we subtract this number from the normal freezing point of H_2O which is 0.00°C:
   $0.00°C - 11.2°C = -11.2°C$
   The freezing point of the solution is – 11.2°C.
7. First we need to convert the mass of NaCl into number of moles using the reverse of the unit of its molar mass (58.45 g/mol) as the conversion factor.

   $$3.5\text{ g NaCl} \times \frac{1\text{ mol NaCl}}{58.45\text{ g NaCl}} = 0.060\text{ mol NaCl}$$

   Using the definition of molarity:

   $$\text{molarity} = \frac{0.060\text{ mol}}{0.0965\text{ L}} = 0.622\text{ M}$$

   NaCl is made up of two ions, so the van't Hoff factor is 2.

   $$\Pi = (2)\left(0.622\ \frac{\text{mol}}{\text{L}}\right)\left(0.08205\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(298\text{ K})$$

   $$\Pi = 30.4\text{ atm}$$

9. To calculate vapor pressure depression according to Raoult's law, the mole fraction of solvent particles must be recalculated to take into account the increased number of particles formed on ionization. Here we need the mole fraction of H_2O. The number of moles of H_2O is as follows:

$$199.5 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.0 \text{ g } H_2O} = 11.1 \text{ mol } H_2O$$

The number of moles of KBr is as follows:

$$36.4 \text{ g KBr} \times \frac{1 \text{ mol KBr}}{119 \text{ g KBr}} = 0.306 \text{ mol KBr}$$

KBr is made up of two ions, so one mole of KBr dissociates into one mole of K^+ and one mole of $Br^—$. Hence, the total number of moles is

11.1 mol + 0.306 mol + 0.306 mol = 11.7 mol

Now we can calculate the mole fraction of H_2O:

$$\chi_{H_2O} = \frac{11.1 \text{ mol}}{11.7 \text{ mol}} = 0.949$$

Now we can use Raoult's law to determine the vapor pressure in equilibrium with the solution:

P_{soln} = (0.9490)(32.55 torr) = 30.86 torr

Any two-ion salt should have the same effect.

7. END-OF-CHAPTER MATERIAL

Additional Exercises

1. majority component is the solvent: ethyl alcohol; minority component is the solute: water
3. Saturated means all the possible solute that can dissolve is dissolved, whereas concentrated implies that a lot of solute is dissolved.
5. We can use molarity as a conversion factor to determine the number of moles of $FeCl_2$ present in the solution, then use molar mass of $FeCl_2$ as a conversion factor to determine its mass:

$$0.445\ \text{L} \times \frac{0.0812\ \text{mol FeCl}_2}{1\ \text{L}} \times \frac{126.75\ \text{g FeCl}_2}{1\ \text{mol FeCl}_2} = 4.58\ \text{g FeCl}_2$$

7. Using molar mass of $Ca(OH)_2$ (74.1 g/mol) we can calculate the number of moles of $Ca(OH)_2$. Then we can use the definition of molarity to calculate the volume of the solution:

$$100.0\ \text{g Ca(OH)}_2 \times \frac{1\ \text{mol Ca(OH)}_2}{74.1\ \text{g Ca(OH)}_2} \times \frac{1\ \text{L}}{0.225\ \text{mol Ca(OH)}_2}$$

$$= 6.00\ \text{L}$$

9. The density of water is assumed to be 1 g/mL. We can use the definition of ppm:

$$1.0\ \text{ppm} = \frac{\text{mass solute}}{1{,}920\ \text{g solution}} \times 1{,}000{,}000$$

Solving the above equation for the mass of solute, we get:

Mass of solute = 0.001920 g (or) 1.92 mg

Hence, 1.92 mg fluoride ions are being ingested per day.

11. We can use the definition of ppm:

For 2.0 ppm: $2.0\ \text{ppm} = \dfrac{\text{mass solute}}{8.50 \times 10^{20}\ \text{kg solution}} \times 1{,}000{,}000$

Mass of solute = 1.7×10^{15} kg

For 4.0 ppm: $4.0\ \text{ppm} = \dfrac{\text{mass solute}}{8.50 \times 10^{20}\ \text{kg solution}} \times 1{,}000{,}000$

Mass of solute = 3.4×10^{15} kg

Hence, 1.7×10^{15} to 3.4×10^{15} kg of uranium is thought to be present in the earth crust.

13. We are given the amount of product formed in grams, so first we need to convert it to moles and then use the balanced equation to determine the amount of H_2O_2 reacted. Finally using mass percentage as the conversion factor, we need to determine the amount of solution needed.

$$35.7\ \text{g O}_2 \times \frac{1\ \text{mol O}_2}{32.0\ \text{g O}_2} \times \frac{2\ \text{mol H}_2\text{O}_2}{1\ \text{mol O}_2} \times \frac{34.02\ \text{g H}_2\text{O}_2}{1\ \text{mol H}_2\text{O}_2}$$

$$\times \frac{100\ \text{g solution}}{3\ \text{g H}_2\text{O}_2} = 2{,}530\ \text{g}$$

15. The solution of $MgCl_2$ has a freezing point of −2.60°C, but the freezing point of pure water is 0.00°C. So, the change in freezing point temperature is

$\Delta Tf = 0.00°\text{C} — (-2.60°\text{C}) = 2.60°\text{C}$

We use the equation to calculate the change in the freezing point including the van't Hoff factor *i*.

Substituting:

$2.60°\text{C} = i(0.500\ m)(1.86°\text{C}/m)$

The true van't Hoff factor:

$i = 2.80$

$MgCl_2$ is an ionic salt composed of three ions, so the ideal van't Hoff factor of this ionic compound is 3.

The true value is less than the ideal value because not all ions behave as independent particles.

17. $C_6H_{12}O_6$ is a molecular compound, so its ideal van't Hoff factor is 1. NaCl is an ionic salt composed of two ions, so its ideal van't Hoff factor is 2. $Au(NO_3)_3$ is an ionic salt composed of four ions, so its ideal van't Hoff factor is 4. $Al_2(SO_4)_3$ is composed of five ions, so the ideal van't Hoff factor is 5.

The order of the given solutions in the order of increasing boiling point is:

$0.10\ m\ C_6H_{12}O_6 < 0.06\ m\ \text{NaCl} < 0.4\ m\ Au(NO_3)_3 < 0.4\ m\ Al_2(SO_4)_3$

CHAPTER 12

Acids and Bases

1. ARRHENIUS ACIDS AND BASES

Test Yourself

1. 1. This formula has an OH in it, but we recognize that potassium is present as K^+ cations. As such, this is an ionic compound of the OH^- ion and is an Arrhenius base.
   2. This compound is an ionic compound between H^+ ions and $SO4^{2-}$ ions, so it is an Arrhenius acid.
   3. It is neither an acid nor a base. (In fact, it is the formula for ethane, an organic compound.)

Test Yourself

1. The general reaction is as follows:

   $$HCl + Mg(OH)_2 \rightarrow H_2O + \text{salt}$$

   Because the base has two OH^-ions in its formula, we need two H^+ ions to react, making two H2O molecules as product. The remaining ions, Mg^{2+} and Cl^-, make the salt magnesium chloride (MgCl2). The balanced chemical reaction is as follows:

   $$2HCl + Mg(OH)_2 \rightarrow 2H_2O + MgCl_2$$

Exercises

1. a compound that increases the H^+ concentration in water
3. sour taste, react with metals, and turn litmus red
5. a. This formula has an OH in it, but we recognize that sodium is present as Na^+ cations. As such, this is an ionic compound of the OH^- ion and is an Arrhenius base.
   b. Although this formula has an OH in it, we do not recognize the remaining part of the molecule as a cation. It is neither an acid nor a base. (In fact, it is the formula for ethanol, an organic compound.)
   c. This compound is an ionic compound between H^+ ions and $PO4^{3-}$ ions, so it is an Arrhenius acid.
7. The general reaction is as follows:

   $$H_2C_2O_4 + KOH \rightarrow H_2O + \text{salt}$$

   Because the acid has two H^+ ions in its formula, we need two OH^- ions to react, making two H2O molecules as product. The remaining ions, K^+ and $C2O4^{2-}$, make the salt potassium oxalate (K2C2O4). The balanced chemical reaction is as follows:

   $$H_2C_2O_4 + 2KOH \rightarrow 2H_2O + K_2C_2O_4$$

9. The general reaction is as follows:

   $$HCl + Fe(OH)_3 \rightarrow H_2O + \text{salt}$$

   The acid has one H^+ ion in its formula; the base has three OH^- ions in its formula. Hence, we need three H^+ ions, making three H2O molecules as product. The remaining ions, Fe^{3+} and Cl^-, make the salt iron(III) chloride (FeCl3). The balanced chemical reaction is as follows:

   $$3HCl + Fe(OH)_3 \rightarrow 3H_2O + FeCl_3$$

11. HCl and Ca(OH)2
    The balanced chemical reaction showing the reactants and products is:

    $$2HCl + Ca(OH)_2 \rightarrow CaCl_2 + 2H_2O$$

13. H2SO4 and Ba(OH)2
    The balanced chemical reaction showing the reactants and products is:

    $$H_2SO_4 + Ba(OH)_2 \rightarrow BaSO_4 + 2H_2O$$

2. BRØNSTED-LOWRY ACIDS AND BASES

Test Yourself

1. The $Al(H_2O)_6^{3+}$ ion is losing an H^+ to become $Al(H_2O)_5(OH)^{2+}$; it is the proton donor and the Brønsted-Lowry acid. The H_2O molecule is accepting the H^+ ion to become H_3O^+ (hydronium ion), so it is the Brønsted-Lowry base.

Test Yourself

1. One pair is H_2O and OH^-, where H_2O has one more H^+ and is the conjugate acid, while OH^- has one less H^+ and is the conjugate base. The other pair consists of NH_2^+ and NH_3, where NH_3 is the conjugate acid (it has an additional proton) and NH_2^+ is the conjugate base.

Exercises

1. A Brønsted-Lowry acid is a proton donor. It does not necessarily increase the H^+ concentration in water.
3. The dissociation of hydrogen bromide in water:

   $$HBr + H_2O \rightarrow H_3O^+ + Br^-$$

   The HBr molecule is losing an H^+ to become Br^-; it is the proton donor. The H_2O molecule is accepting the H^+ ion to become H_3O^+; it is the proton acceptor.
5. Since pyridine acts as a Brønsted-Lowry base in water, it accepts the proton from water.

   $$C_5H_5N + H_2O \rightarrow C_2H_5NH^+ + OH^-$$

   The H_2O molecule is losing an H^+ to become OH^-; it is the proton donor. The C_5H_5N molecule is accepting the H^+ ion to become $C_5H_5NH^+$; it is the proton acceptor.
7. The H_3PO_4 molecule is losing an H^+ to become $H_2PO_4^-$; it is the proton donor and the Brønsted-Lowry acid. The OH^- ion is accepting the H^+ ion to become H_2O, so it is the Brønsted-Lowry base.
9. The $HC_2H_3O_2$ looses one H^+ ion to become $C_2H_3O_2^-$; C_2H_5N accepts the H^+ proton to become $C_5H_5NH^+$.

   $$HC_2H_3O_2 + C_5H_5N \rightarrow C_2H_3O_2^- + C_5H_5NH^+$$
11. Conjugate acids have one more H^+. Hence, conjugate acids of H_2O and NH_3 are H_3O^+ and NH_4^+, respectively.
13. Conjugate bases have one less H^+. Hence, conjugate bases of HSO_4^- and H_2O are SO_4^{2-} and OH^-, respectively.
15. One pair is HSO_4^- and SO_4^{2-}, where HSO_4^- has one more H^+ and is the conjugate acid, while SO_4^{2-} has one less H^+ and is the conjugate base. The other pair consists of $PO_4^{3-}$and HPO_4^{2-}, where HPO_4^{2-} is the conjugate acid (it has an additional proton) and PO_4^{3-} is the conjugate base.
17. One pair is NH_3 and NH_2^-, where NH_3 has one more H^+ and is the conjugate acid, while NH_2^- has one less H^+ and is the conjugate base. The other pair consists of $C_6H_5O^-$and C_6H_5OH, where C_6H_5OH is the conjugate acid (it has an additional proton) and $C_6H_5O^-$is the conjugate base.

3. ACID-BASE TITRATIONS

Test Yourself

1. In liters, the volume is 0.01809 L. We calculate the number of moles of titrant:

   # moles NaOH = (0.01809 L)(0.2235 M) = 0.004043 mol NaOH

   Using the balanced chemical equation, we can determine the number of moles of $H_2C_2O_4$ present in the analyte:

   $$0.004043 \text{ mol NaOH} \times \frac{1 \text{ mol } H_2C_2O_4}{2 \text{ mol NaOH}} = 0.002022 \text{ mol } H_2C_2O_4$$

   Then we convert this to a mass using the molar mass of $H_2C_2O_4$:

   $$0.002022 \text{ mol } H_2C_2O_4 \times \frac{90.035 \text{ g } H_2C_2O_4}{1 \text{ mol } H_2C_2O_4} = 0.1820 \text{ g } H_2C_2O_4$$

Exercises

1. a chemical reaction performed in a quantitative fashion
3. False; a base can be a titrant, or the reaction being performed may not even be an acid-base reaction.
5. The balanced equation for the reaction above is:

   $$HCl + NaOH \rightarrow H_2O + NaCl$$

   In liters, the volume is 0.05560 L. We calculate the number of moles of titrant:

   # moles HCl = (0.05560 L)(0.2221 M) = 0.01235 mol HCl

   Using the balanced equation we can calculate the number of moles of NaOH:

   $$0.01235 \text{ mol HCl} \times \frac{1 \text{ mol NaOH}}{1 \text{ mol HCl}} = 0.01235 \text{ mol NaOH}$$

   Then we convert this to a mass using the molar mass of NaOH:

   $$0.01235 \text{ mol NaOH} \times \frac{40.0 \text{ g NaOH}}{1 \text{ mol NaOH}} = 0.4940 \text{ g NaOH}$$

7. The balanced equation for the reaction above is:

   $$Ca(OH)_2 + 2HBr \rightarrow CaBr_2 + 2H_2O$$

   In liters, the volume is 0.04566 L. We calculate the number of moles of titrant:

   # moles HBr = (0.04566 L)(0.1126 M) = 0.005141 mol HBr

   Using the balanced chemical equation, we can determine the number of moles of $Ca(OH)_2$ present in the analyte:

   $$0.005141 \text{ mol HBr} \times \frac{1 \text{ mol } Ca(OH)_2}{2 \text{ mol HBr}} = 0.002571 \text{ mol } Ca(OH)_2$$

   Then we find the concentration of $Ca(OH)_2$ solution (volume of $Ca(OH)_2$ is converted to L, 0.02500 L):

   $$\text{Molarity} = \frac{0.002571 \text{ mol } Ca(OH)_2}{0.02500 \text{ L } Ca(OH)_2} = 0.1028 \text{ M } Ca(OH)_2$$

4. STRONG AND WEAK ACIDS AND BASES AND THEIR SALTS

Test Yourself

1. 1. Because RbOH is listed in Table 12.2, it is a strong base.
   2. Because HNO2is not listed in Table 12.2, it is a weak acid.

Test Yourself

1. This is an ionic compound of H$^+$ ions and N3$^-$ ions. When an ionic compound dissolves, it separates into its constituent ions:

   $$HN_3 \rightarrow H^+(aq) + N_3^-(aq)$$

   Because HN3is not listed in Table 12.2, this reaction does not proceed 100% to products.

Test Yourself

1. 1. C5H5NH$^+$ is derived from a weak base; Cl$^-$ is derived from a strong acid. Therefore the solution is acidic, and (C5H5NH)Cl is an acidic salt.
   2. Na$^+$ is derived from a strong base; SO3$^-$ is derived from a weak acid. Therefore the solution is basic, and Na2SO3 is a basic salt.

Exercises

1. A strong acid is 100% ionized in aqueous solution, whereas a weak acid is not 100% ionized.

3. a. Because HF is not listed in Table 12.2, it is a weak acid.
   b. Because HCl is listed in Table 12.2, it is a strong acid.
   c. Because HC2O4 is not listed in Table 12.2, it is a weak acid.

5. a. This is an ionic compound of H$^+$ ions and NO3$^-$ ions. When an ionic compound dissolves, it separates into its constituent ions:

   $$HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$$

   Because HNO3 is listed in Table 12.2, this reaction proceeds 100% to products.

   b. This is an ionic compound of H$^+$ ions and NO2$^-$ ions. When an ionic compound dissolves, it separates into its constituent ions:

   $$HNO_2(aq) \rightarrow H^+(aq) + NO_2^-(aq)$$

   Because HNO2 is not listed in Table 12.2, this reaction does not proceed 100% to products.

   c. This is an ionic compound of H$^+$ ions and I3$^-$ ions. When an ionic compound dissolves, it separates into its constituent ions:

   $$HI_3(aq) \rightarrow H^+(aq) + I_3^-(aq)$$

   Because HI3 is not listed in Table 12.2, this reaction does not proceed 100% to products.

7. a. HCl acts as the proton donor; C5H5N acts as the proton acceptor.

   $$HCl + C_5H_5N \rightarrow Cl^- + C_5H_5NH^+$$

   b. H2C2O4 acts as the proton donor; NH3 acts as the proton acceptor.

   $$H_2C_2O_4 + 2NH_3 \rightarrow C_2O_4{}^{2-} + 2NH_4{}^+$$

   c. HNO2 acts as the proton donor; C7H9N acts as the proton acceptor.

   $$HNO_2 + C_7H_9N \rightarrow NO_2^- + C_7H_9NH^+$$

9. a. Na$^+$ is derived from a strong base; Br$^-$ is derived from a strong acid. Therefore the solution is neutral, so NaBr is a neutral salt.
   b. Fe^{2+} is derived from a weak base; NO3$^-$ is derived from a strong acid. Therefore the solution is acidic, and Fe(NO3)2 is an acidic salt.
   c. Fe^{3+} is derived from a weak base; NO3$^-$ is derived from a strong acid. Therefore the solution is acidic, and Fe(NO3)3 is an acidic salt.

11. a. Na$^+$ is derived from a strong base; NO2$^-$ is derived from a weak acid. Therefore the solution is basic, and NaNO2 is a basic salt.
    b. Na$^+$ is derived from a strong base; NO3$^-$ is derived from a strong acid. Therefore the solution is neutral, and NaNO3 is a neutral salt.
    c. NH4$^+$ is derived from a weak base; NO3$^-$ is derived from a strong acid. Therefore the solution is acidic, and NH4NO3 is an acidic salt.

13. a. K^+ is derived from a strong base; SO_3^{2-} is derived from a weak acid. Hence, SO_3^{2-} undergoes hydrolysis.

$$SO_3^{2-} + H_2O \rightarrow HSO_3^{-} + OH^{-}$$

b. K^+ is derived from a strong base; I^- is derived from a strong acid. Hence, there is no reaction.

c. NH_4^+ is derived from a weak base; ClO_3^- is derived from a strong acid. Hence, NH_4^+ undergoes hydrolysis.

$$NH_4^{+} + H_2O \rightarrow NH_3 + H_3O^{+}$$

15. $NH_4^+ + H_2O \rightarrow NH_3 + H_3O^+$; $NO_2^- + H_2O \rightarrow HNO_2 + OH^-$; it is not possible to determine whether the solution will be acidic or basic.

17. greater than 0.015 M because there are two OH^- ions per formula unit of $Mg(OH)_2$

5. AUTOIONIZATION OF WATER

Test Yourself

1. Using the expression and known value for K_W, $K_W = [H^+][OH^-] = 1.0 \times 10^{-14} = (1.0 \times 10^{-9})[H^+]$

$$\left[H^+\right] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-9}} = 1.0 \times 10^{-5}\ M$$

Test Yourself

1. We begin by determining $[H^+]$. The concentration of the solute is 0.00032 M, but because H_2SO_4 is a strong acid, there are two H^+ ions in solution for every formula unit dissolved, so the actual $[H^+]$ is two times this, or 2 × 0.00032 M = 0.00064 M. Now we can use the K_W expression:

$$[H^+][OH^-] = 1.0 \times 10^{-14} = (0.00064\ M)[OH^-]$$

$$[OH^-] = \frac{1.0 \times 10^{-14}}{0.00064} = 1.6 \times 10^{-11}\ M$$

Test Yourself

1. Because the acid is only 0.44% ionized, we can determine $[OH^-]$ from the concentration of the acid. Recall that 0.44% is 0.0044 in decimal form:

$$[OH^-] = 0.0044 \times 0.0222 = 0.0000977\ M$$

With this $[OH^-]$, then $[H^+]$ can be calculated as follows:

$$[H^+] = \frac{1.0 \times 10^{-14}}{0.0000977} = 1.02 \times 10^{-10}$$

Exercises

1. $[H^+]$ varies with the amount of acid or base in a solution.

3. $[H^+] = \dfrac{K_W}{[OH^-]}$

5. $H_2O + H_2O \rightarrow H_3O^+ + OH^-$; H_2O/H_3O^+ and H_2O/OH^-

7. Using the expression and known value for K_W, $K_W = [H^+][OH^-] = 1.0 \times 10^{-14} = (1.0 \times 10^{-3})\ [OH^-]$

$$[OH^-] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-3}} = 1.0 \times 10^{-11}\ M$$

9. Using the expression and known value for K_W, $K_W = [H^+][OH^-] = 1.0 \times 10^{-14} = (7.92 \times 10^{-5})\ [OH^-]$

$$[OH^-] = \frac{1.0 \times 10^{-14}}{7.92 \times 10^{-5}} = 1.26 \times 10^{-10}\ M$$

11. Using the expression and known value for K_W, $K_W = [H^+][OH^-] = 1.0 \times 10^{-14} = (1.0 \times 10^{-5})\ [H^+]$

$$[H^+] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-5}} = 1.0 \times 10^{-9}\ M$$

13. Using the expression and known value for K_W, $K_W = [H^+][OH^-] = 1.0 \times 10^{-14} = (3.77 \times 10^{-4})\ [H^+]$

$$[H^+] = \frac{1.0 \times 10^{-14}}{3.77 \times 10^{-4}} = 2.65 \times 10^{-11}\ M$$

15. We begin by determining $[H^+]$. The concentration of the solute is 0.344 M, but because HNO_3 is a strong acid, there is one H^+ ion in solution for every formula unit dissolved, so the actual $[H^+]$ is one times this, or 1 × 0.344 M = 0.344 M. Now we can use the K_W expression:

$$[H^+][OH^-] = 1.0 \times 10^{-14} = (0.344\ M)[OH^-]$$

$$[OH^-] = \frac{1.0 \times 10^{-14}}{0.344} = 2.91 \times 10^{-14} M$$

17. We begin by determining [OH⁻]. The concentration of the solute is 0.00338 M, but because KOH is a strong base, there is one OH⁻ ion in solution for every formula unit dissolved, so the actual [OH⁻] is one times this, or 1 × 0.00338 M = 0.00338 M. Now we can use the K_W expression:

$$[H^+][OH^-] = 1.0 \times 10^{-14} = (0.00338\ M)[H^+]$$

$$[H^+] = \frac{1.0 \times 10^{-14}}{0.00338} = 2.96 \times 10^{-12} M$$

19. Because the acid is only 0.445% ionized, we can determine $[H^+]$ from the concentration of the acid. Recall that 0.445% is 0.00445 in decimal form:

$$[H^+] = 0.00445 \times 0.307 = 0.00137\ M$$

With this $[H^+]$, then $[OH^-]$ can be calculated as follows:

$$[OH^-] = \frac{1.0 \times 10^{-14}}{0.00137} = 7.30 \times 10^{-12}\ M$$

6. THE PH SCALE

Test Yourself

1.
   1. With a pH greater than 7, human blood is basic.
   2. With a pH greater than 7, household ammonia is basic.
   3. With a pH of less than 7, cherries are acidic.

Test Yourself

1. We need to evaluate the expression

   $[H^+] = 10^{-10.36}$

   $[H^+] = 4.4 \times 10^{-11}$ M

   From this, $[OH^-]$ can be determined:

   $$[OH^-] = \frac{1.0 \times 10^{-14}}{4.4 \times 10^{-11}} = 2.3 \times 10^{-4} \text{ M}$$

Test Yourself

1. Because the sum of pH and pOH equals 14, we have

   pH + 12.04 = 14

   pH = 1.96

   Now we evaluate the following two expressions:

   $[H^+] = 10^{-1.96}$

   $[OH^-] = 10^{-12.04}$

   So

   $[H^+] = 0.0110$ M

   $[OH^-] = 9.1 \times 10^{-13}$ M

Exercises

1. pH is the negative logarithm of $[H^+]$ and is equal to 14 – pOH.
3. pH < 7
5. pH = 7 for a neutral solution

   Now evaluate, $[H^+] = 10^{-7}$

   $[H^+] = 1.0 \times 10^{-7}$ M
7. Every entry above pure water is acidic.
9. $[H^+] = 3.44 \times 10^{-4}$ M

   Now evaluate, pH = —log $[3.44 \times 10^{-4}]$

   pH = 3.46

11. $[OH^-] = 6.22 \times 10^{-7}$ M

Now evaluate, pOH = —log $[6.22 \times 10^{-7}]$

pOH = 6.21

Because the sum of pH and pOH equals 14, we have

pH + 6.21 = 14

pH = 7.79

13. $[H^+] = 3.44 \times 10^{-4}$ M

Now evaluate, pH = —log $[3.44 \times 10^{-4}]$

pH=3.46

Because the sum of pH and pOH equals 14, we have

pOH + 3.46 = 14

pOH = 10.54

15. $[OH^-] = 6.22 \times 10^{-7}$ M

Now evaluate, pOH = —log $[6.22 \times 10^{-7}]$

pOH = 6.21

17. Because the sum of pH and pOH equals 14, we have
0.77 + pOH = 14

pOH = 13.23

Now we evaluate the following two expressions:

$[H^+] = 10^{-0.77}$

$[OH^-] = 10^{-13.23}$

So

$[H^+] = 0.17$ M

$[OH^-] = 5.89 \times 10^{-14}$ M

7. BUFFERS

Test Yourself

1.
    1. NaCl is a salt from a strong base, so the combination of these two solutes would not make a buffer solution.
    2. H_3PO_4 is a weak acid, while NaH_2PO_4 is the salt derived from the weak acid. The combination of these two solutes would make a buffer solution.
    3. NH_3 is a weak base, while $(NH_4)_3PO_4$ is the salt made from NH_3. The combination of these two solutes would make a buffer solution.
    4. NaOH is a strong base, so the combination of these two solutes would not make a buffer solution.

Exercises

1. A buffer is the combination of a weak acid or base and a salt of that weak acid or base.
3.
    a. HCl is a strong acid, not a weak acid, so the combination of these two solutes would not make a buffer solution.
    b. HNO_2 is a weak acid and $NaNO_2$ is the salt derived from the weak acid. The combination of these two solutes would make a buffer solution.
    c. HNO_3 is a strong acid, not a weak acid, so the combination of these two solutes would not make a buffer solution.
    d. NH_3 is a weak base, while NH_4NO_3 is the salt made from NH_3. The combination of these two solutes would make a buffer solution.
5. 3b. HNO_2 and $NaNO_2$

    Strong base: $HNO_2 + OH^- \rightarrow H_2O + NO_2^-$

    Strong acid: $NO_2^- + H^+ \rightarrow HNO_2$

    3d. NH_4NO_3 and NH_3

    Strong base: $NH_4^+ + OH^- \rightarrow NH_3 + H_2O$

    Strong acid: $NH_3 + H^+ \rightarrow NH_4^+$
7. Buffers can be made from three combinations: (1) H_3PO_4 and $H_2PO_4^-$, (2) $H_2PO_4^-$ and HPO_4^{2-}, and (3) HPO_4^{2-} and PO_4^{3-}. (Technically, a buffer can be made from any two components.)
9. Phosphate buffer have more solute dissolved in them. Hence, the phosphate buffer should have the larger capacity.

8. END-OF-CHAPTER MATERIAL

Additional Exercises

1. The reaction is as follows:

$$Zn + HCl \rightarrow H_2 + ZnCl_2$$

Insert 2 before HCl to get the balanced equation:

$$Zn + 2HCl \rightarrow H_2 + ZnCl_2$$

3. The O^{2-} ion would come from H_2O, which is not considered a classic acid in the Arrhenius sense.

5. The general reaction is as follows:

$$Al(OH)_3 + H_2SO_4 \rightarrow H_2O + \text{salt}$$

The base has three OH^- ions in its formula; the acid has two H^+ ions in its formula. This results in two molecules of $Al(OH)_3$ and three molecules of H_2SO_4 making six H_2O molecules as product. The remaining ions, Al^{3+} and SO_4^{2-}, make the salt aluminum sulfate ($Al_2(SO_4)_3$). The balanced chemical reaction is as follows:

$$2Al(OH)_3 + 3H_2SO_4 \rightarrow 6H_2O + Al_2(SO_4)_3$$

7. $C_8H_{10}N_4O_2$ acts as a Brønsted-Lowry base by accepting a proton from H_2O.

$$C_8H_{10}N_4O_2 + H_2O \rightarrow C_8H_{10}N_4O_2H^+ + OH^-$$

9. As a strong acid or base, an amphiprotic substance reacts 100% as an acid or a base, so it cannot be a base or an acid at the same time.

11. if the salt produced is an acidic salt

13. $NH_3 + NH_3 \rightarrow NH_4^+ + NH_2^-$

15. pOH > 7

17. We begin by determining $[H^+]$. The concentration of the solute is 12 M, but because HCl is a strong acid, there is one H^+ ion in solution for every formula unit dissolved, so the actual $[H^+]$ is one times this, or 1 × 12 M = 12 M. Now we can use the K_W expression:

$$[H^+][OH^-] = 1.0 \times 10^{-14} = (12\text{ M})[OH^-]$$

$$[OH^-] = \frac{1.0 \times 10^{-14}}{12} = 8.33 \times 10^{-16}\text{M}$$

Now we evaluate the following two expressions:

$$pH = -\log[12] = -1.08$$

$$pOH = -\log[8.33 \times 10^{-16}] = 15.08$$

CHAPTER 13

Chemical Equilibrium

1. CHEMICAL EQUILIBRIUM

Test Yourself

1. As an equilibrium situation, a double arrow is used. The equilibrium equation is written as follows:

   $2H_2 + O_2 \rightleftharpoons 2H_2O$

Exercises

1. The situation when the forward and reverse chemical reactions occur, leading to no additional net change in the reaction position; $H_2 + I_2 \rightleftharpoons 2HI$ (answers will vary)

3. As an equilibrium situation, a double arrow is used. The equilibrium equation is written as follows:

   $H_2 + Cl_2 \rightleftharpoons 2HCl$

5. As an equilibrium situation, a double arrow is used. The equilibrium equation is written as follows:

   C (gra) ⇌ C (dia); C (dia) ⇌ C (gra)

2. THE EQUILIBRIUM CONSTANT

Test Yourself

1. We start by writing the K_{eq} expression. Using the *products over reactants* approach, the K_{eq} expression is as follows:

$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]}$$

Substituting for the equilibrium $[H_2]$ and [HI] and for the given value of K_{eq}:

$$0.40 = \frac{(0.060)^2}{[H_2](0.90)}$$

$$[H_2] = \frac{(0.060)^2}{(0.40)(0.90)}$$

$$[H_2] = 0.010\ M$$

Test Yourself

1. We start by writing the K_{eq} expression from the balanced chemical equation:

$$K_{eq} = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

Substituting for the known equilibrium concentrations and the K_{eq}, this becomes

$$\begin{aligned} 13.7 &= \frac{[NH_3]^2}{(0.055)(1.62)^3} \\ [NH_3]^2 &= (13.7)(0.055)(1.62)^3 \\ [NH_3] &= 1.79\ M \end{aligned}$$

Test Yourself

1. Write the K_P expression for this reaction:

$$K_P = \frac{P^2_{HCl}}{P_{H_2} \cdot P_{Cl_2}}$$

Then substitute the equilibrium partial pressures into the expression and evaluate:

$$K_P = \frac{(2.98)^2}{(0.44)(0.22)} = 91.7$$

Test Yourself

1. Before we use the relevant equation, we need to do two things: convert the temperature to the kelvins and determine Δn. Converting the temperature:

   $T = 25 + 273 = 298$ K

   To determine the change in the number of moles of gas, take the number of moles of gaseous products and subtract the number of moles of gaseous reactants. There are 2 mol of gas as product and 1 mol of gas of reactant:

   $\Delta n = 2 - 1 = 1$ mol

   Now we can substitute into our equation, using $R = 0.08205$ L·atm/mol·K:

   $K_P = (98.3)[(0.08205)(298)]^1$

   $K_P = 2.40 \times 10^3$

Exercises

1. the relationship between the concentrations of reactants and products of a chemical reaction at equilibrium

3. a. $K_{eq} = \frac{[HCl]^2}{[H_2][Cl_2]}$

   b. $K_{eq} = \frac{[N_2O_3]}{[NO][NO_2]}$

5. a. $K_P = \frac{P^2_{H_2O}}{P^2_{H_2} P_{O_2}}$

   b. $K_P = \frac{P^2_{H_2O} P_{O_2}}{P^2_{H_2O_2}}$

7. We start by writing the K_{eq} expression from the balanced chemical equation:

   $$K_{eq} = \frac{[PBr_5]}{[PBr_3][Br_2]}$$

   Substituting for the known equilibrium concentrations and the K_{eq}, this becomes

   $$1.65 = \frac{(0.55)}{[PBr_3](2.05)}$$

   $$[PBr_3] = \frac{(0.55)}{(1.65)(2.05)}$$

   $$[PBr_3] = 0.163 \text{ M}$$

9. We start by writing the K_{eq} expression from the balanced chemical equation:

   $$K_{eq} = \frac{[HCl]^2[CH_2Cl_2]}{[Cl_2]^2[CH_4]}$$

   Substituting for the known equilibrium concentrations and the K_{eq}, this becomes

   $$2.30 = \frac{[HCl]^2(0.175)}{(0.150)^2(0.250)}$$

   $$[HCl]^2 = \frac{(2.30)(0.150)^2(0.250)}{0.175}$$

   $$[HCl] = 0.272 \text{ M}$$

11. $K_P = \frac{P^2_{N_2O_5}}{P^4_{NO_2} P_{O_2}}$

13. Write the K_P expression for this reaction:

   $$K_P = \frac{P_{COBr_2}}{P_{Br_2} \cdot P_{CO}}$$

   Then substitute the equilibrium partial pressures into the expression and evaluate:

   $$4.08 = \frac{P_{COBr_2}}{(0.235)(0.666)}$$

   $$P_{COBr_2} = (0.235)(0.666)(4.08)$$

   $$P_{COBr_2} = 0.639 \text{ atm}$$

15. Before we use the relevant equation, we need to determine Δn.

   To determine the change in the number of moles of gas, take the number of moles of gaseous products and subtract the number of moles of gaseous reactants. There are 2 mol of gas as product and 3 mol of gas of reactant:

   $\Delta n = 2 - 3 =$ —1 mol

   Now we can substitute into our equation, using $R =$ 0.08205 L·atm/mol·K:

   $K_P = (1.76 \times 10^{-3})[(0.08205)(298)]^{-1}$

   $K_P = 7.20 \times 10^{-5}$

17. Before we use the relevant equation, we need to do two things: convert the temperature to the kelvins and determine Δn. Converting the temperature:

$T = 660 + 273 = 933$ K

To determine the change in the number of moles of gas, take the number of moles of gaseous products and subtract the number of moles of gaseous reactants. There are 1 mol of gas as product and 2 mol of gas of reactant:

$\Delta n = 1 - 2 = -1$ mol

Now we can substitute into our equation, using R = 0.08205 L·atm/mol·K:

$5.205 \times 10^{-3} = K_{eq}\,[(0.08205)(933)]^{-1}$

$$K_{eq} = \frac{5.205 \times 10^{-3}}{[(0.08205)(933)]^{-1}}$$

$$K_{eq} = 3.98 \times 10^{-1}$$

19. Equation is heterogeneous. *Do not include the concentrations of pure solids and pure liquids in* K*eq* *expressions* for heterogeneous equations.

$$K_{eq} = \frac{[\mathrm{NaCl}]}{[\mathrm{NaOH}][\mathrm{HCl}]}$$

21. The equation is heterogeneous. Only partial pressures for gas-phase substances or concentrations in solutions are included in the expressions of equilibrium constants.

$$K_P = P_{CO_2}$$

3. SHIFTING EQUILIBRIA: LE CHATELIER'S PRINCIPLE

Test Yourself

1. 1. If Br_2 is removed, there is now less reactants, so the reaction will shift toward reactants.
   2. If $COBr_2$ is added, there is now more products, so the reaction will shift toward reactants.

Test Yourself

1. According to Le Chatelier's principle, if pressure is decreased, then the equilibrium shifts to the side with the greater number of moles of gas. This particular reaction shows a total of 3 mol of gas as reactants and 2 mol of gas as products, so the reaction shifts toward the reactants side.

Test Yourself

1. Because energy is listed as a reactant, it is being absorbed, so the reaction is endothermic. If the temperature is decreasing, a reactant is being removed from the equilibrium, so the equilibrium shifts to minimize the removal of a reactant: it shifts toward reactants.

Exercises

1. When an equilibrium is stressed, the equilibrium shifts to minimize that stress.
3. a. Because energy is listed as a reactant, it is being absorbed, so the reaction is endothermic. If the temperature is decreasing, a reactant is being removed from the equilibrium, so the equilibrium shifts to minimize the removal of a reactant: it shifts toward reactants.
   b. If pressure is increased, then the equilibrium shifts to the side with the lesser number of moles of gas. This particular reaction shows a total of 1 mol of gas as reactants and 2 mol of gas as products, so the reaction shifts toward the reactants side.
   c. If HI is removed, there is now less products, so the reaction will shift toward products.
5. a. If SO_3 is removed, there is now less products, so the reaction will shift toward products.
   b. If O_2 is added, there is now more reactants, so the reaction will shift toward products.
   c. Because energy is listed as a product, it is being produced, so the reaction is exothermic. If the temperature is decreasing, a product is being removed from the equilibrium, so the equilibrium shifts to minimize the removal of a product: it shifts toward products.
7. In the reaction energy is being produced, so the reaction is exothermic. When the temperature is decreased, a product will be removed from the equilibrium, so the equilibrium shifts to minimize the removal of a product: it shifts toward the products side, which maximizes the amount of NH_3.

   By increasing the pressure, the equilibrium shifts to the side with the smaller number of moles of gas. This particular reaction shows a total of 4 mol of gas as reactants and 2 mol of gas as products, so the reaction shifts toward the products side, which maximizes the amount of NH_3.

   By removing NH_3, there will be less number of products. The equilibrium shifts toward the products side, which maximizes the amount of NH_3.

4. CALCULATING EQUILIBRIUM CONCENTRATIONS FROM K_{EQ} VALUES

Test Yourself

1. The ICE chart is set up like this. First, the initial values:

	$N_2H_2(g)$	⇌	$N_2(g)$	+	$H_2(g)$
I	0.075		0		0
C					
E					

For x M of N_2H_2 lost, x M of N_2 and x M of H_2 is produced. These expressions go into the change row:

	$N_2H_2(g)$	⇌	$N_2(g)$	+	$H_2(g)$
I	0.075		0		0
C	$-x$		$+x$		$+x$
E					

At equilibrium, the resulting concentrations will be combination of the initial amount and the changes:

	$N_2H_2(g)$	⇌	$N_2(g)$	+	$H_2(g)$
I	0.075		0		0
C	$-x$		$+x$		$+x$
E	0.075 — x		$+x$		$+x$

The expressions in the equilibrium row go into the K_{eq} expression:

$$K_{eq} = \frac{(x)(x)}{(0.075\ - x)} = 0.052 = \frac{(x)(x)}{(0.075\ - x)}$$

$$0.052(0.075 - x) = x^2$$

$$x^2 + 0.052x - 0.0039 = 0$$

$$x = \frac{-0.052 \pm \sqrt{(0.052)^2 - 4(1)(-0.0039)}}{2(1)}$$

$$x = \frac{-0.052 \pm 0.135}{2}$$

$$x = 0.042 \text{ or } x = -0.094$$

Because x is the final concentration of both N_2 and H_2, it cannot be negative. Thus $x = 0.042$.
$[N_2H_2] = 0.075$ — $x = 0.075 - 0.042 = 0.033$ M
$[N_2] = x = 0.042$ M
$[H_2] = x = 0.042$ M

Exercises

1. I = initial concentrations; C = change in concentrations; E = equilibrium concentrations
3. The ICE chart is set up like this. First, the initial values:

	$3O_2(g)$	⇌	$2O_3(g)$
I	0.075		0
C			
E			

For $3x$ of M O_2 lost, $2x$ M of O_3 is produced. These expressions go into the change row:

	$3O_2(g)$	⇌	$2O_3(g)$
I	0.075		0
C	$-3x$		$+2x$
E			

At equilibrium, the resulting concentrations will be combination of the initial amount and the changes:

	$3O_2(g)$	⇌	$2O_3(g)$
I	0.075		0
C	$-3x$		$+2x$
E	$0.075 - 3x$		$+2x$

5. The ICE chart is set up like this. First, the initial values:

	CH_4	+	$2O_2$	⇌	CO_2	+	$2H_2O$
I	0.0060		0.055		0		0
C							
E							

Since H_2O is in liquid form, they do not appear in the change row. x M of CH_4 and $2x$ M of O_2 is lost; x M of CO_2 is produced. These expressions go into the change row:

	CH_4	+	$2O_2$	⇌	CO_2	+	$2H_2O$
I	0.0060		0.055		0		0
C	$-x$		$-2x$		$+x$		—
E							

At equilibrium, the resulting concentrations will be combination of the initial amount and the changes:

	CH_4	+	$2O_2$	⇌	CO_2	+	$2H_2O$
I	0.0060		0.055		0		0
C	$-x$		$-2x$		$+x$		—
E	$0.0060 - x$		$0.055 - 2x$		$+x$		—

7. First set up the ICE chart.

	$HCN(g)$	⇌	$HNC(g)$
I	2.00		0
C	x		$+x$
E	2.00 — x		$+x$

The expressions in the equilibrium row go into the K_{eq} expression:

$$K_{eq} = \frac{x}{2.00 - x} = 4.50 = \frac{x}{2.00 - x}$$

$$9.00 - 4.50x = x$$

$$5.50x = 9.00$$

$$x = 1.636$$

[HCN] = 2.00 — x = 2.00 − 1.636 = 0.364 M

[HNC] = x = 1.64 M

9. First set up the ICE chart.

	N_2O_3	$\rightleftharpoons$	NO	NO_2
I	0.0663		0	0
C	$-x$		$+x$	$+x$
E	$0.0663 — x$		$+x$	$+x$

The expressions in the equilibrium row go into the K_{eq} expression:

$$K_{eq} = \frac{(x)(x)}{0.0663 - x} = 2.50 = \frac{x^2}{0.0663 - x}$$

$$x^2 = 2.50(0.0663 - x)$$

$$x^2 = 0.16575 - 2.50x$$

$$x^2 + 2.50x - 0.16575 = 0$$

$$x = \frac{-2.50 \pm \sqrt{2.50^2 - 4(1)(-0.16575)}}{2(1)}$$

$$x = \frac{-2.50 \pm 2.6292}{2}$$

$$x = 0.0646 \text{ or } x = -2.5646$$

Because x is the final concentration of both NO and NO_2, it cannot be negative.

Hence, $x = 0.0646$.

$[N_2O_3] = 0.0663 — x = 0.0663 - 0.0646 = 0.0017$ M

$[NO] = x = 0.0646$ M

$[NO_2] = x = 0.0646$ M

11. First set up the ICE chart.

	H_2S	$\rightleftharpoons$	H_2	S(s)
I	0.882		0	0
C	$0.882 — x$		$+x$	–
E	$0.882 — x$		$+x$	–

The expressions in the equilibrium row go into the K_{eq} expression:

$$K_{eq} = \frac{x}{0.882 - x} = 0.055 = \frac{x}{0.882 - x}$$

$$0.055(0.882 - x) = x$$

$$0.04851 - 0.055x = x$$

$$1.055x = 0.04851$$

$$x = 0.0460$$

$[H_2S] = 0.882 — x = 0.882 - 0.0460 = 0.836$ M

$[H_2] = x = 0.836$ M

5. SOME SPECIAL TYPES OF EQUILIBRIA

Test Yourself

1. $HPO4^{2-}$ acts as a weak acid by separating into an H^+ ion and $PO4^{3-}$ ion:

$$HPO_4{}^{2-}(aq) \rightleftharpoons H^+(aq) + PO_4{}^{3-}(aq)$$

K_a is written in terms of the concentrations of products divided by concentrations of reactants:

$$K_a = \frac{[H^+][PO_4^{3-}]}{[HPO_4^{2-}]}$$

Test Yourself

1. The balanced equilibrium equation is

$$HCN(aq) \rightleftharpoons H^+(aq) + CN^-(aq)$$

0.500M

Set up the ICE chart.

	HCN	⇌	H^+	CN^-
I	0.500		0	0
C	$-x$		x	x
E	0.500 — x		x	x

The expressions in the equilibrium row go into the K_a expression:

$$K_a = \frac{(x)(x)}{0.500 - x} = 6.2 \times 10^{-10} = \frac{x^2}{0.500 - x}$$

For small values, $0.500 - x \approx 0.500$

$$6.2 \times 10^{-10} = \frac{x^2}{0.500}$$

$$x = 1.761 \times 10^{-5}$$

$[H^+] = 1.761 \times 10^{-5}$ M

pH = — log $[H^+]$ = — log $[1.761 \times 10^{-5}]$ = 4.75

Test Yourself

1. To determine the K_b for $PO4^{3-}$, we need to know the K_a of its conjugate acid. The conjugate acid of $PO4^{3-}$ is $HPO4^{2-}$. We know that $K_a = 2.2 \times 10^{-13}$.

$(2.2 \times 10^{-13})K_b = 1.0 \times 10^{-14}$

$$K_b = \frac{1.0 \times 10^{-14}}{2.2 \times 10^{-13}} = 4.5 \times 10^{-2}$$

Test Yourself

1. Recall that when an ionic compound dissolves, it separates into its individual ions. For Ag_2SO_4, the ionization reaction is as follows:

   $$Ag_2SO_4(s) \rightleftharpoons 2Ag^{+}(aq) + SO_4{}^{2-}(aq)$$

   Hence the K_{sp} expression is
   $K_{sp} = [Ag^{+}]^2[SO_4{}^{2-}]$

Test Yourself

1. The chemical equation for the dissolving of $BaSO_4$ is

   $$BaSO_4(s) \rightleftharpoons Ba^{2+}(aq) + SO_4{}^{2-}(aq)$$

   The K_{sp} expression is as follows:
   $K_{sp} = [Ba^{2+}][SO_4{}^{2-}]$

	$BaSO_4$ (s)	$\rightleftharpoons$	Ba^{2+}(aq)	+	$SO_4{}^{2-}$(aq)
I			0		0
C	$-x$		$+x$		$+x$
E			$+x$		$+x$

   Substituting the equilibrium values into the expression:
   $(x)(x) = 1.1 \times 10^{-10}$
   $x^2 = 1.1 \times 10^{-10}$
   $x = 1.0 \times 10^{-5}$
   $[Ba^{2+}] = 1.0 \times 10^{-5}$ M
   $[SO_4{}^{2-}] = 1.0 \times 10^{-5}$ M

Test Yourself

1. The equilibrium equation for the dissolving of $Mg(OH)_2$ is

   $$Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2OH^{-}(aq)$$

   The K_{sp} expression is as follows:
   $K_{sp} = [Mg^{2+}][OH^{-}]^2$

	$Mg(OH)_2$	$\rightleftharpoons$	Mg^{2+}	+	$2OH^{-}$
I			0		0
C	$-x$		$+x$		$+2x$
E			$+x$		$+2x$

   Substituting the equilibrium values into the expression:
   $(x)(2x)^2 = 5.6 \times 10^{-12}$
   $4x^3 = 5.6 \times 10^{-12}$
   $x = 1.1 \times 10^{-4}$
   $[Mg^{2+}] = x = 1.1 \times 10^{-4}$ M
   $[OH^{-}] = 2x = 2.2 \times 10^{-4}$ M

Exercises

1. The K_{sp} is a special type of the K_{eq} and applies to compounds that are only slightly soluble.
3. HF acts as a weak acid by separating into an H^+ ion and F^- ion:

$$HF(aq) \rightleftharpoons H^+(aq) + F^-(aq)$$

5. For the equation $HF(aq) \rightleftharpoons H^+(aq) + F^-(aq)$,

$$K_{eq} = \frac{[H^+][F^-]}{[HF]}$$

7. The balanced equilibrium equation:

$$HClO_2(aq) \rightleftharpoons H^+(aq) + ClO_2^-(aq)$$

0.100M

Set up the ICE chart.

	$HClO_2$	$\rightleftharpoons$	H^+	ClO_2^-
I	0.100		0	0
C	$-x$		$+x$	$+x$
E	$0.100 - x$		$+x$	$+x$

The expressions in the equilibrium row go into the K_a expression:

$$K_a = \frac{(x)(x)}{0.100 - x} = 1.1 \times 10^{-2} = \frac{x^2}{0.100 - x}$$

For small values, $0.100 - x \approx 0.100$

$$1.1 \times 10^{-2} = \frac{x^2}{0.100}$$

$x = 0.0332$

$[H^+] = x = 0.0332$ M
$[ClO_2^-] = x = 0.0332$ M
$[HClO_2] = 0.100 - x = 0.100 - 0.0332 = 0.0668$ M

9. The balanced equilibrium equation:

$$HNO_2(aq) \rightleftharpoons H^+(aq) + NO_2^-(aq)$$

1.00 M

Set up the ICE chart.

	HNO_2	$\rightleftharpoons$	H^+	NO_2^-
I	1.00		0	0
C	$-x$		$+x$	$+x$
E	$1.00 - x$		$+x$	$+x$

The expressions in the equilibrium row go into the K_a expression:

$$K_a = \frac{(x)(x)}{1.00 - x} = 5.6 \times 10^{-4} = \frac{x^2}{1.00 - x}$$

For small values, $1.00 - x \approx 1.00$

$$5.6 \times 10^{-4} = \frac{x^2}{1.00}$$

$x = 0.0236$

$[H^+] = x = 0.0236$ M
$pH = -\log[H^+] = -\log[0.0236]$
$= 1.63$

11. First step of dissociation:

$$H_3PO_4(aq) \rightleftharpoons H^+(aq) + H_2PO_4^-(aq)$$

$$K_a = \frac{[H^+][H_2PO_4^-]}{[H_3PO_4]}$$

Second step of dissociation:

$$H_2PO_4^-(aq) \rightleftharpoons H^+(aq) + HPO_4^{2-}(aq)$$

$$K_a = \frac{[H^+][HPO_4^{2-}]}{[H_2PO_4^-]}$$

Third step of dissociation:

$$HPO_4^{2-}(aq) \rightleftharpoons H^+(aq) + PO_4^{3-}(aq)$$

$$K_a = \frac{[H^+][PO_4^{3-}]}{[HPO_4^{2-}]}$$

13. To determine the K_b for NO_2^-, we need to know the K_a of its conjugate acid. The conjugate acid of NO_2^- is HNO_2. We know that $K_a = 5.6 \times 10^{-4}$.
$(5.6 \times 10^{-4})K_b = 1.0 \times 10^{-14}$

$$K_b = \frac{1.0 \times 10^{-14}}{5.6 \times 10^{-4}} = 1.8 \times 10^{-11}$$

15. $[OH^-][H^+] = 1.0 \times 10^{-14}$

$$[OH^-] = \frac{1.0 \times 10^{-14}}{3.23 \times 10^{-6}} = 3.10 \times 10^{-9}\text{ M}$$

17. $[OH^-][H^+] = 1.0 \times 10^{-14}$

$$[H^+] = \frac{1.0 \times 10^{-14}}{2.09 \times 10^{-2}} = 4.78 \times 10^{-13}\text{ M}$$

19. When an ionic compound dissolves, it separates into its individual ions. For $Mg(OH)_2$, the ionization reaction is as follows:

$$Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2OH^-(aq)$$

Hence the K_{sp} expression is
$K_{sp} = [Mg^{2+}][OH^-]^2$

21. The chemical equilibrium equation for the dissolving of $SrSO_4$ is

$$SrSO_4(s) \rightleftharpoons Sr^{2+}(aq) + SO_4^{2-}(aq)$$

The K_{sp} expression is as follows:
$K_{sp} = [Sr^{2+}][SO_4^{2-}]$

	$SrSO_4$	$\rightleftharpoons$	Sr^{2+}	+	SO_4^{2-}
I			0		0
C	$-x$		$+x$		$+x$
E			$+x$		$+x$

Substituting the equilibrium values into the expression:
$(x)(x) = 3.8 \times 10^{-4}$
$x^2 = 3.8 \times 10^{-4}$
$x = 1.9 \times 10^{-2}$
$[Sr^{2+}] = x = 1.9 \times 10^{-2}$ M
$[SO_4^{2-}] = x = 1.9 \times 10^{-2}$ M

23. The chemical equilibrium equation for the dissolving of $Ca(OH)_2$ is

$$Ca(OH)_2(s) \rightleftharpoons Ca^{2+}(aq) + 2OH^-(aq)$$

The K_{sp} expression is as follows:
$K_{sp} = [Ca^{2+}][OH^-]^2$

	$Ca(OH)_2$	$\rightleftharpoons$	Ca^{2+}	+	$2OH^-$
I			0		0
C	$-x$		$+x$		$+2x$
E			$+x$		$+2x$

Substituting the equilibrium values into the expression:
$(x)(2x)^2 = 5.0 \times 10^{-6}$
$4x^3 = 5.0 \times 10^{-6}$
$x = 0.011$
$[Ca^{2+}] = x = 0.011$ M
$[OH^-] = 2x = 0.022$ M

6. END-OF-CHAPTER MATERIAL

Additional Exercises

1. They are reciprocals of each other.
3. In the reaction energy is being produced, so the reaction is exothermic. When the temperature is decreased, a product will be removed from the equilibrium, so the equilibrium shifts to minimize the removal of a product: it shifts toward the products side, which maximizes the amount of PCl_5.

   By increasing the pressure, the equilibrium shifts to the side with the smaller number of moles of gas. This particular reaction shows a total of 2 mol of gas as reactants and 1 mol of gas as products, so the reaction shifts toward the products side, which maximizes the amount of PCl_5.

   By adding PCl_3, there will be more number of reactants. The equilibrium shifts toward the products side, which maximizes the amount of PCl_5.

   By adding Cl_2, there will be more number of reactants. The equilibrium shifts toward the products side, which maximizes the amount of PCl_5.

   By removing PCl_5, there will be less number of products. The equilibrium shifts toward the products side, which maximizes the amount of PCl_5.
5. favor products because the numerator of the ratio for the K_{eq} is larger than the denominator
7. Consider $HX(aq) \rightleftharpoons H^+(aq) + X^-(aq)$:

   $$K_a = \frac{[H^+][X^-]}{[HX]}$$

   Consider $X^-(aq) + H_2O(\ell) \rightleftharpoons HX(aq) + OH^-(aq)$:

   $$K_b = \frac{[HX][OH^-]}{[X^-]}$$

   $$K_a \times K_b = \frac{[H^+][X^-]}{[HX]} \times \frac{[HX][OH^-]}{[X^-]} = [H^+][OH^-] = K_w$$

9. The chemical equilibrium equation for the dissolving of AgCl is

   $$AgCl \rightleftharpoons Ag^+(aq) + Cl^-(aq)$$

   The K_{sp} expression is as follows:
   $K_{sp} = [Ag^+][Cl^-]$

	AgCl	$\rightleftharpoons$	Ag^+	+	Cl^-
I			0		0
C	$-x$		$+x$		$+x$
E			$+x$		$+x$

   Substituting the equilibrium values into the expression:
   $(x)(x) = 1.8 \times 10^{-10}$
   $x^2 = 1.8 \times 10^{-10}$
   $x = 1.3 \times 10^{-5}$
   $[Ag^+] = 1.3 \times 10^{-5}$ mol/L
   $[Cl^-] = 1.3 \times 10^{-5}$ mol/L
   Solubility of AgCl is 1.3×10^{-5} mol/L.
11. $K_{eq} = K_P$ when the number of moles of gas on both sides of the reaction are the same.
13. The chemical equation for the dissolving of $Mg(OH)_2$ is

   $$Mg(OH)_2 \rightleftharpoons Mg^{2+}(aq) + 2OH^-(aq)$$

   From the Table 13.2, K_{sp} of saturated solution of $Mg(OH)_2$ is 5.6×10^{-12}.
   The K_{sp} expression is as follows:
   $K_{sp} = [Mg^{2+}][OH^-]^2$

	$Mg(OH)_2$	$\rightleftharpoons$	Mg^{2+}	+	$2OH^-$
I			0		0
C	$-x$		$+x$		$+2x$
E			$+x$		$+2x$

   Substituting the equilibrium values into the expression:
   $(x)(2x)^2 = 5.6 \times 10^{-12}$
   $4x^3 = 5.6 \times 10^{-12}$
   $x = 1.12 \times 10^{-4}$
   $[Mg^{2+}] = 1.12 \times 10^{-4}$ M
   $[OH^-] = 2.24 \times 10^{-4}$ M
   Using $[H^+][OH^-] = 1.0 \times 10^{-14}$
   $[H^+]\ 2.24 \times 10^{-4}$ M $= 1.0 \times 10^{-14}$
   $[H^+] = 4.46 \times 10^{-11}$ M
   Using pH = $-\log [H^+]$
   pH = $-\log [4.46 \times 10^{-11}] = 10.35$

15. The chemical equilibrium equation for the dissolving of MX_2 is

$MX_2 \rightleftharpoons M^{2+} + 2X^-$

The K_{sp} expression is as follows:
$K_{sp} = [M^{2+}][X^-]^2$

	MX_2	$\rightleftharpoons$	M^{2+}	+	$2X^-$
I			0		0
C	$-x$		$+x$		$+2x$
E			$+x$		$+2x$

Substituting the equilibrium values into the expression:
$K_{sp} = (x)(2x)^2$
$K_{sp} = 4x^3$

17. The chemical equation for the dissolving of $PbBr_2$ is

$PbBr_2(s) \rightleftharpoons Pb^{2+}(aq) + 2Br^-(aq)$

The K_{sp} expression is as follows:
$K_{sp} = [Pb^{2+}][Br^-]^2$

	$PbBr_2$	$\rightleftharpoons$	Pb^{2+}	+	$2Br^-$
I			0		0
C	$-x$		$+1.33 \times 10^{-5}$		$+2.66 \times 10^{-5}$
E			$+1.33 \times 10^{-5}$		$+2.66 \times 10^{-5}$

a. From the table $[Br^-] = 2.66 \times 10^{-5}$ M.
b. Substituting the equilibrium values into the expression:
$K_{sp} = (1.33 \times 10^{-5})(2.66 \times 10^{-5})^2$
$K_{sp} = 9.41 \times 10^{-15}$

CHAPTER 14

Oxidation and Reduction

1. OXIDATION-REDUCTION REACTIONS

Test Yourself

1. By rule 3, hydrogen and oxygen are assigned oxidation numbers of +1 and −2, respectively. According to rule 4, the sum of the oxidation numbers on all atoms must equal the charge on the species, so we have the simple algebraic equation

   $$x + 3(+1) + 4(-2) = 0$$

   where x is the oxidation number of the P atom and the 0 represents the charge on the species. Evaluating for x,

   $$x + (-5) = 0 = +5$$

   Thus the oxidation number on the P atom is +5.

Test Yourself

1. Both reactants are the elemental forms of their atoms, so the C and O atoms have oxidation numbers of 0. In the product, the O atom has an oxidation number of −2, while the C atom has an oxidation number of +4.

   $$\begin{array}{lllll} & C + & O_2 & \rightarrow C & O_2 \\ \text{ox\#} & 0 & 0 & +4 & -2 \end{array}$$

   Carbon is increasing its oxidation number from 0 to +4, so it is being oxidized; oxygen is decreasing its oxidation number from 0 to −2, so it is being reduced.

   Because oxidation numbers are changing, this is a redox reaction. The total number of electrons being lost by carbon (four) is gained by oxygen (four, two gained by each O atom).

Exercises

1. Both reactants are the elemental forms of their atoms, so the K and Br atoms have oxidation numbers of 0. In the product, the K ion has an oxidation number of +1, while the Br ion has an oxidation number of −1.

   $$\begin{array}{lllll} & 2K + & Br_2 & \rightarrow 2K & Br \\ \text{ox\#} & 0 & 0 & +1 & -1 \end{array}$$

   Potassium is increasing its oxidation number from 0 to +1, so it is being oxidized; bromine is decreasing its oxidation number from 0 to −1, so it is being reduced.

   Because oxidation numbers are changing, this is a redox reaction.

3. Oxidation numbers assigned to atoms in the reactants - Mg = 0, O = 0

   Oxidation numbers assigned to atoms in the products - MgO: Mg = +2, O = −2

   $$\begin{array}{lllll} & 2Mg + & O_2 & \rightarrow 2Mg & O \\ \text{ox\#} & 0 & 0 & +2 & -2 \end{array}$$

   Magnesium is increasing its oxidation number from 0 to +2, so it is being oxidized; oxygen is decreasing its oxidation number from 0 to −2, so it is being reduced.

   The total number of electrons lost by magnesium atom (two) is gained by the oxygen atom (two, one gained by each O atom).

5. Oxidation numbers assigned to atoms in the reactants - Li = 0; O = 0

Oxidation numbers assigned to atoms in the products - Li_2O_2: Li = +2, O = −1

$$\begin{array}{lcccc} & 2Li + & O_2 \rightarrow & Li_2 & O_2 \\ \text{ox\#} & 0 & 0 & +2 & -1 \end{array}$$

Lithium is increasing its oxidation number from 0 to +2, so it is being oxidized; oxygen is decreasing its oxidation number from 0 to −1, so it is being reduced.

7. increase in oxidation number; loss of electrons

9. a. By rule 1, phosphorous is assigned an oxidation number 0, as it exists in its elemental form.

b. By rule 3, oxygen is assigned an oxidation number −2.

According to rule 4, the sum of the oxidation numbers on all atoms must equal the charge on the species, so we have the simple algebraic equation

$$x + 3(\ -2\ ) = 0$$

where x is the oxidation number of sulfur, and 0 is the net charge on SO_3. Solving for x,

$$x + (\ -6\ ) = 0$$

$$x = \ +6$$

Thus the oxidation number of sulfur is +6.

c. By rule 3, oxygen is assigned an oxidation number −2.

According to rule 4, the sum of the oxidation numbers on all atoms must equal the charge on the species, so we have the simple algebraic equation

$$x + 3(\ -2\ ) = \ -2$$

where x is the oxidation number of sulfur, and −2 is the net charge on SO_3. Solving for x,

$$x + (\ -6\ ) = \ -2 = \ +4$$

Thus the oxidation number of sulfur is +4.

d. $Ca_3(PO_3)_2$ dissociates to form Ca^{2+} and PO_3^{3-}ions.

By rule 2, calcium is assigned an oxidation number +2, which is equal to its charge.

By rule 3, oxygen is assigned an oxidation number −2.

According to rule 4, the sum of the oxidation numbers on all atoms must equal the charge on the species, so we have the simple algebraic equation

$$x + 3(\ -2\ ) = \ -3$$

where x is the oxidation number of phosphorous, and −3 is the net charge on PO_3. Solving for x,

$$x - 6 = \ -3$$

$$x = \ +3$$

Thus the oxidation number on the P atom is +3.

11. a. By rule 3, oxygen is assigned an oxidation number −2.

According to rule 4, the sum of the oxidation numbers on all atoms must equal the charge on the species, so we have the simple algebraic equation

$$x + (\ -2\ ) = 0$$

where x is the oxidation number of nitrogen, and 0 is the net charge on NO. Solving for x,

$$x = \ +2$$

Thus the oxidation number on the N atom is +2.

b. By rule 3, oxygen is assigned an oxidation number −2. According to rule 4, the sum of the oxidation numbers on all atoms must equal the charge on the species, so we have the simple algebraic equation

$$x + 2(\ -2\ ) = 0$$

where x is the oxidation number of nitrogen, and 0 is the net charge on NO_2. Solving for x,

$$x - 4 = 0$$

$$x = \ +4$$

Thus the oxidation number on the N atom is +4.

c. By rule 2, the Cl atom is assigned an oxidation number −1, which is equal to its charge.

According to rule 4, the sum of the oxidation numbers on all atoms must equal the charge on the species, so we have the simple algebraic equation

$$x + 2(\ -1\ ) = 0$$

where x is the oxidation number of chromium, and 0 is the net charge on $CrCl_2$. Solving for x,

$$x - 2 = 0$$

$$x = \ +2$$

Thus the oxidation number on the Cr atom is +2.

d. By rule 2, the Cl atom is assigned an oxidation number −1, which is equal to its charge.

According to rule 4, the sum of the oxidation numbers on all atoms must equal the charge on the species, so we have the simple algebraic equation

$$x + (\ -3\ ) = 0$$

where x is the oxidation number of chromium, and 0 is the net charge on $CrCl_3$. Solving for x,

$$x = \ +3$$

Thus the oxidation number on the Cr atom is +3.

13. a. By rule 3, the oxygen and the hydrogen atoms are assigned oxidation numbers of −2 and +1, respectively.

According to rule 4, the sum of the oxidation numbers on all atoms must equal the charge on the species, so we have the simple algebraic equation

$$x + (+2) + (-2) = 0$$

where x is the oxidation number of carbon, and 0 is the net charge on CH_2O. Solving for x,

$$x = 0$$

Thus the oxidation number on the C atom is 0.

b. By rule 3, the hydrogen atom is assigned an oxidation number +1.

According to rule 4, the sum of the oxidation numbers on all atoms must equal the charge on the species, so we have the simple algebraic equation

$$x + (+3) = 0$$

where x is the oxidation number of nitrogen, and 0 is the net charge on NH_3. Solving for x,

$$x = -3$$

Thus the oxidation number on the N atom is −3.

c. Rb_2SO_4 dissociates to form Rb^+ and SO_4^{2-} ions.

By rule 2, Rb is assigned an oxidation number +1, which is equal to its charge.

By rule 3, the oxygen atom is assigned an oxidation number −2.

According to rule 4, the sum of the oxidation numbers on all atoms must equal the charge on the species, so we have the simple algebraic equation

$$x + (-8) = -2$$

where x is the oxidation number of sulfur, and −2 is the net charge on SO_4^{2-}. Solving for x,

$$x = +6$$

Thus the oxidation number on the S atom is +6.

d. $Zn(C_2H_3O_2)_2$ dissociates to form Zn^{2+} and $C_2H_3O_2^{-}$ ions.

By rule 2, Zn is assigned an oxidation number +2, which is equal to its charge.

By rule 3, the oxygen and hydrogen atoms are assigned oxidation numbers of −2 and +1, respectively.

According to rule 4, the sum of the oxidation numbers on all atoms must equal the charge on the species, so we have the simple algebraic equation

$$2x + (+3) + (-4) = -1$$

where x is the oxidation number of carbon, and −1 is the net charge on $C_2H_3O_2^{-}$. Solving for x,

$$x = 0$$

Thus the oxidation number on the C atom is 0.

15. Oxidation number assigned to atoms in the reactants - NO: N = +2, O = −2; Cl_2: Cl = 0

Oxidation number assigned to atoms in the products - NOCl: N = +3, O = −2, Cl = −1

$$\begin{array}{lcccccccc} & 2\text{N} & \text{O} & + & \text{Cl}_2 & \rightarrow & 2\text{N} & \text{O} & \text{Cl} \\ \text{ox\#} & +2 & -2 & & 0 & & +3 & -2 & -1 \end{array}$$

Nitrogen is increasing its oxidation number from +2 to +3, so it is being oxidized; chlorine is decreasing its oxidation number from 0 to −1, so it is being reduced.

17. Oxidation number assigned to atoms in the reactants - KrF_2: Kr = +2, F = −1; H_2O: H = +1, O = −2

Oxidation number assigned to atoms in the products - Kr: Kr = 0; HF: H = +1, F = −1; O_2 = 0

$$\begin{array}{lcccccccccccc} & 2\text{Kr} & \text{F}_2 & + & 2\text{H}_2 & \text{O} & \rightarrow & \text{Kr} & + & 4\text{H} & \text{F} & + & \text{O}_2 \\ \text{ox\#} & +2 & -1 & & +1 & -2 & & 0 & & +1 & -1 & & 0 \end{array}$$

Oxygen is increasing its oxidation number from −2 to 0, so it is being oxidized; krypton is decreasing its oxidation number from +2 to 0, so it is being reduced.

19. Oxidation number assigned to atoms in the reactants - Rb: Rb = 0; $MgCl_2$: Mg = +2, Cl = −1

Oxidation number assigned to atoms in the products - RbCl: Rb = +1, Cl = −1; Mg: Mg = 0

$$\begin{array}{lcccccccc} & 2\text{Rb} & + & \text{Mg} & \text{Cl}_2 & \rightarrow & 2\text{Rb} & \text{Cl} & + & \text{Mg} \\ \text{ox\#} & 0 & & +2 & -1 & & +1 & -1 & & 0 \end{array}$$

Rubidium is increasing its oxidation number from 0 to +1, so it is being oxidized; magnesium is decreasing its oxidation number from +2 to 0, so it is being reduced.

2. BALANCING REDOX REACTIONS

Test Yourself

1. Add the coefficient 2 before Al to balance the number of Al atoms on both the sides:

   $2Al + O_2 \rightarrow Al_2O_3$

   We can balance the number of O atoms by adding the coefficient 3 before O_2 and 2 before Al_2O_3.

   $2Al + 3O_2 \rightarrow 2Al_2O_3$

   This imbalances the number of Al atoms. Add 4 before Al to balance the number of Al atoms.

   $4Al + 3O_2 \rightarrow 2Al_2O_3$

   This gives us four Al atoms on both sides and a total of six O atoms on both sides of the chemical equation. This redox reaction is now balanced.

Test Yourself

1. Oxygen is being oxidized, and fluorine is being reduced:

   $2O^{2-} \rightarrow O_2$ oxidation

   $F_2 \rightarrow 2F^-$ reduction

   Then we include the appropriate number of electrons on the proper side to balance the charges for each reaction:

   $2O^{2-} \rightarrow O_2 + 4e^-$

   $F_2 + 2e^- \rightarrow 2F^-$

   The first reaction involves four electrons, while the second reaction involves two electrons. The least common multiple of these two numbers is four, so to get four electrons in each reaction we need to double the second reaction:

   $2F_2 + 4e^- \rightarrow 4F^-$

   We can combine the two final reactions, noting that the electrons cancel:

   $2O^{2-} + 2F_2 + 4e^- \rightarrow O_2 + 4e^- + 4F^-$

   The overall, balanced redox reaction is

   $2O^{2-} + 2F_2 \rightarrow O_2 + 4F^-$

Test Yourself

1. The oxidation reaction is as follows:

   $Cl^{-} \rightarrow ClO_3{}^{-}$

   The Cl atom is going from a −1 to a +5 oxidation state and loses six electrons in the process. We add those six electrons to the product side:

   $Cl^{-} \rightarrow ClO_3{}^{-} + 6e^{-}$

   Now we must balance the O atoms. Because the solution is basic, we should use OH^{-} rather than H_2O:

   $3OH^{-} + Cl^{-} \rightarrow ClO_3{}^{-} + 6e^{-}$

   We have introduced H atoms as part of the reactants; we can balance them by adding H^{+} as products:

   $3OH^{-} + Cl^{-} \rightarrow ClO_3{}^{-} + 6e^{-} + 3H^{+}$

   If we check the atoms and the overall charge on both sides, we see that this reaction is balanced. However, if the reaction is occurring in a basic solution, it is unlikely that H^{+}ions will be present in quantity. The way to address this is to add additional OH^{-}ions to each side of the equation:

   $3OH^{-} + 3OH^{-} + Cl^{-} \rightarrow ClO_3{}^{-} + 6e^{-} + 3H^{+} + 3OH^{-}$

   The six OH^{-}ions on the left side can be grouped together as $6OH^{-}$. On the right side, the H^{+} and OH^{-}ions can be grouped into $3H_2O$ molecules:

   $6OH^{-} + Cl^{-} \rightarrow ClO_3{}^{-} + 6e^{-} + 3H_2O$

   This is a more appropriate form for a basic solution.
   Now we balance the reduction reaction:

   $MnO_4{}^{-} \rightarrow MnO_2$

   The Mn atom is going from +7 to +4 in oxidation number, which requires a gain of three electrons:

   $3e^{-} + MnO_4{}^{-} \rightarrow MnO_2$

   Then we balance the O atoms and then the H atoms:

   $3e^{-} + MnO_4{}^{-} \rightarrow MnO_2 + 2OH^{-}$

   $2H^{+} + 3e^{-} + MnO_4{}^{-} \rightarrow MnO_2 + 2OH^{-}$

   We add two OH^{-}ions to each side to eliminate the H^{+} ion in the reactants; the reactant species combine to make two water molecules, and the number of OH^{-}ions in the product increases to four:

   $2H_2O + 3e^{-} + MnO_4{}^{-} \rightarrow MnO_2 + 4OH^{-}$

   This reaction is balanced for a basic solution.
   Now we combine the two balanced half reactions. The oxidation reaction has six electrons, while the reduction reaction has three. The least common multiple of these two numbers is six, so we multiply the reduction reaction by 2 so that the electrons are balanced:

   $2 \times [2H_2O + 3e^{-} + MnO_4{}^{-} \rightarrow MnO_2 + 4OH^{-}]$

$4H_2O + 6e^- + 2MnO_4^- \rightarrow 2MnO_2 + 8OH^-$

Combining these two equations results in the following equation:

$6OH^- + Cl^- + 4H_2O + 6e^- + 2MnO_4^- \rightarrow ClO_3^- + 6e^- + 3H_2O + 2MnO_2 + 8OH^-$

The six electrons cancel. What remains is $Cl^- + H_2O + 2MnO_4^- \rightarrow ClO_3^- + 2MnO_2 + 2OH^-$ which is our final balanced redox reaction.

Exercises

1. a. Na atoms are balanced on both the sides.

   Add the coefficient 2 before NaF to balance the number of F atoms on both the sides:

   $Na+F_2 \rightarrow 2NaF$

   This imbalances the number of Na atoms. Add the coefficient 2 before Na to balance the number of Na atoms on both the sides.

   $2Na+F_2 \rightarrow 2NaF$

   This redox reaction is now balanced.

   b. Add the coefficient 2 before Al to balance the number of Al atoms on both the sides:

   $Al_2O_3 + H_2 \rightarrow 2Al + H_2O$

   We can balance the number of O atoms by adding the coefficient 3 before H_2O.

   $Al_2O_3 + H_2 \rightarrow 2Al + 3H_2O$

   This imbalances the number of H atoms. Add 3 before H_2 to balance the number of H atoms.

   $Al_2O_3 + 3H_2 \rightarrow 2Al + 3H_2O$

   This redox reaction is now balanced.

3. a. C atoms are balanced on both the sides.

   Add the coefficient 2 before H_2O to balance the number of H atoms on both the sides.

   $CH_4 + O_2 \rightarrow CO_2 + 2H_2O$

   Add the coefficient 2 before O_2 to balance the number of O atoms on both the sides.

   $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

   This redox reaction is now balanced.

   b. Add the coefficient 2 before PCl_3 to balance the number of P atoms on both the sides.

   $P_2O_5 + Cl_2 \rightarrow 2PCl_3 + O_2$

   Add the coefficient 3 before Cl_2 to balance the number of Cl atoms on both the sides.

   $P_2O_5 + 3Cl_2 \rightarrow 2PCl_3 + O_2$

   Add the coefficient 5 before O_2 and 2 before P_2O_5 to balance the number of O atoms on both the sides.

   $2P_2O_5 + 3Cl_2 \rightarrow 2PCl_3 + 5O_2$

   This imbalances the number of P atoms. Double the coefficient 2 before PCl_3 to balance the number of P atoms.

   $2P_2O_5 + 3Cl_2 \rightarrow 4PCl_3 + 5O_2$

   This imbalances the number of Cl atoms. Double the coefficient 3 before Cl_2 to balance the number of Cl atoms.

   $2P_2O_5 + 6Cl_2 \rightarrow 4PCl_3 + 5O_2$

   This redox reaction is now balanced.

5. a. Calcium is being oxidized, and hydrogen is being reduced:

$Ca \rightarrow Ca^{2+}$ oxidation

$2H^{+} \rightarrow H_2$ reduction

Then we include the appropriate number of electrons on the proper side to balance the charges for each reaction:

$Ca \rightarrow Ca^{2+} + 2e^{-}$

$2H^{+} + 2e^{-} \rightarrow H_2$

We can combine the two final reactions, noting that the electrons cancel:

$Ca+2H^{+} + 2e^{-} \rightarrow Ca^{2+} + 2e^{-} + H_2$

The overall, balanced redox reaction is

$Ca+2H^{+} \rightarrow Ca^{2+} + H_2$

b. Tin is being oxidized and reduced:

$Sn^{2+} \rightarrow Sn^{4+}$ oxidation

$Sn^{2+} \rightarrow Sn$ reduction

Then we include the appropriate number of electrons on the proper side to balance the charges for each reaction:

$Sn^{2+} \rightarrow Sn^{4+} + 2e^{-}$

$Sn^{2+} + 2e^{-} \rightarrow Sn$

We can combine the two final reactions, noting that the electrons cancel:

$Sn^{2+} + Sn^{2+} + 2e^{-} \rightarrow Sn^{4+} + 2e^{-} + Sn$

The overall, balanced redox reaction is

$2Sn^{2+} \rightarrow Sn^{4+} + Sn$

7. a. Sodium is being oxidized, and mercury is reduced:

$Na \rightarrow NaCl$ oxidation

$Hg_2Cl_2 \rightarrow 2Hg$ reduction

Then we include the appropriate number of electrons on the proper side to balance the charges for each reaction:

$Na \rightarrow NaCl + e^{-}$

$Hg_2Cl_2 + 2e^{-} \rightarrow 2Hg$

The first reaction involves one electron, while the second reaction involves two electrons. The least common multiple of these two numbers is two, so to get two electrons in each reaction we need to multiply the first reaction by 2:

$2 \times [Na \rightarrow NaCl + e^{-}] = 2Na \rightarrow 2NaCl + 2e^{-}$

We can combine the two final reactions, noting that the electrons cancel:

$2Na+Hg_2Cl_2 + 2e^{-} \rightarrow 2NaCl + 2e^{-} + 2Hg$

The overall, balanced redox reaction is

$2Na+Hg_2Cl_2 \rightarrow 2NaCl + 2Hg$

b. Carbon is being oxidized, and aluminum is reduced:

$C \rightarrow CO2$ oxidation

$Al_2O_3 \rightarrow 2Al$ reduction

Then we include the appropriate number of electrons on the proper side to balance the charges for each reaction:

$C \rightarrow CO_2 + 4e^{-}$

$Al_2O_3 + 6e^{-} \rightarrow 2Al$

The first reaction involves four electrons, while the second reaction involves six electrons. The least common multiple of these two numbers is twelve, so to get twelve electrons in each reaction we need to multiply the first reaction by 3 and the second reaction by 2:

$3 \times [C \rightarrow CO_2 + 4e^{-}] = 3C \rightarrow 3CO_2 + 12e^{-}$

$2 \times [Al_2O_3 + 6e^{-} \rightarrow 2Al] = 2Al_2O_3 + 12e^{-} \rightarrow 4Al$

We can combine the two final reactions, noting that the electrons cancel:

$3C+2Al_2O_3 + 12e^{-} \rightarrow 3CO_2 + 12e^{-} + 4Al$

The overall, balanced redox reaction is

$$2Al_2O_3 + 3C \rightarrow 4Al + 3CO_2$$

9. a. **Basic solution**

$$Cu + NO_3^- \rightarrow Cu^{2+} + NO_2$$

The oxidation reaction is as follows:

$$Cu \rightarrow Cu^{2+}$$

The Cu atom is going from 0 to a +2 oxidation state and loses two electrons in the process. We add those two electrons to the product side:

$$Cu \rightarrow Cu^{2+} + 2e^-$$

Balancing the reduction reaction:

$$NO_3^- \rightarrow NO_2$$

The N atom is going from +5 to +4 in oxidation number, which requires a gain of one electron:

$$1e^- + NO_3^- \rightarrow NO_2$$

We balance the O atoms and then the H atoms:

$$1e^- + NO_3^- \rightarrow NO_2 + OH^-$$

$$H^+ + 1e^- + NO_3^- \rightarrow NO_2 + OH^-$$

We add one OH^- ion to each side to eliminate the H^+ ion in the reactants; the reactant species combine to make a water molecule, and the number of OH^-ions in the product increases to two:

$$H_2O + 1e^- + NO_3^- \rightarrow NO_2 + 2OH^-$$

This is the balanced equation for the basic solution.

Now we combine the two balanced half reactions. The oxidation reaction has two electrons, while the reduction reaction has one. The least common multiple of these two numbers is two, so we multiply the reduction reaction by 2 so that the electrons are balanced:

$$2 \times [H_2O + 1e^- + NO_3^- \rightarrow NO_2 + 2OH^-]$$

$$= 2H_2O + 2e^- + 2NO_3^- \rightarrow 2NO_2 + 4OH^-$$

Combining the oxidation reaction and the multiplied reduction reaction results in the following equation:

$$2H_2O + Cu + 2e^- + 2NO_3^- \rightarrow Cu^{2+} + 2e^- + 2NO_2 + 4OH^-$$

The two electrons cancel:

$$2H_2O + Cu + 2NO_3{}^- \rightarrow Cu^{2+} + 2NO_2 + 4OH^-$$

The above equation is the final balanced redox reaction for the basic solution.

Acidic solution

$$Cu + NO_3{}^- \rightarrow Cu^{2+} + NO_2$$

The oxidation reaction is as follows:

$$Cu \rightarrow Cu^{2+}$$

The Cu atom is going from 0 to a +2 oxidation state and loses two electrons in the process. We add those two electrons to the product side:

$$Cu \rightarrow Cu^{2+} + 2e^-$$

Balancing the reduction reaction:

$$NO_3{}^- \rightarrow NO_2$$

The N atom is going from +5 to +4 in oxidation number, which requires a gain of one electron:

$$1e^- + NO_3{}^- \rightarrow NO_2$$

We balance the O atoms and then the H atoms:

$$1e^- + NO_3{}^- \rightarrow NO_2 + OH^-$$

$$H^+ + 1e^- + NO_3{}^- \rightarrow NO_2 + OH^-$$

We add one H^+ ion to each side to eliminate the OH^- ion in the products; the product species combine to make a water molecule and the number of H^+ions in the reactant increases to two:

$$2H^+ + 1e^- + NO_3{}^- \rightarrow NO_2 + H_2O$$

This is the balanced equation for the acidic solution.

Now we combine the two balanced half reactions. The oxidation reaction has two electrons, while the reduction reaction has one. The least common multiple of these two numbers is two, so we multiply the reduction reaction by 2 so that the electrons are balanced:

$$2 \times [2H^+ + 1e^- + NO_3{}^- \rightarrow NO_2 + H_2O]$$

$$= 4H^+ + 2e^- + 2NO_3{}^- \rightarrow 2NO_2 + 2H_2O$$

Combining the oxidation reaction and the multiplied reduction reaction results in the following equation:

$$4H^+ + 2e^- + 2NO_3{}^- + Cu \rightarrow 2NO_2 + 2H_2O + Cu^{2+} + 2e^-$$

The two electrons cancel:

$$4H^+ + 2NO_3{}^- + Cu \rightarrow Cu^{2+} + 2NO_2 + 2H_2O$$

The above equation is the final balanced redox reaction for the acidic solution.

b. **Basic solution**

$$Fe + MnO_4{}^- \rightarrow Fe^{3+} + Mn$$

The oxidation reaction is as follows:

$$Fe \rightarrow Fe^{3+}$$

The Fe atom is going from 0 to a +3 oxidation state and loses three electrons in the process. We add those three electrons to the product side:

$$Fe \rightarrow Fe^{3+} + 3e^-$$

Balancing the reduction reaction:

$$MnO_4{}^- \rightarrow Mn$$

The Mn atom is going from +7 to 0 in oxidation number, which requires a gain of seven electrons:

$$7e^- + MnO_4{}^- \rightarrow Mn$$

We balance the O atoms and then the H atoms:

$$7e^- + MnO_4{}^- \rightarrow Mn + 4OH^-$$

$$4H^+ + 7e^- + MnO_4{}^- \rightarrow Mn + 4OH^-$$

We add four OH^-ions to each side to eliminate the H^+ ion in the reactants; the reactant species combine to make four water molecules, and the number of OH^-ions in the product increases to eight:

$$4H_2O + 7e^- + MnO_4{}^- \rightarrow Mn + 8OH^-$$

This is the balanced equation for the basic solution.

Now we combine the two balanced half reactions. The oxidation reaction has three electrons, while the reduction reaction has seven. The least common multiple of these two numbers is twenty one, so we multiply the oxidation reaction by 7 and the reduction reaction by 3 so that the electrons are balanced:

$$7 \times [Fe \rightarrow Fe^{3+} + 3e^-] = 7Fe \rightarrow 7Fe^{3+} + 21e^-$$

$$3 \times [4H_2O + 7e^- + MnO_4^- \rightarrow Mn + 8OH^-]$$

$$= 12H_2O + 21e^- + 3MnO_4^- \rightarrow 3Mn + 24OH^-$$

Combining these two equations results in the following equation:

$$7Fe + 12H_2O + 21e^- + 3MnO_4^-$$

$$\rightarrow 7Fe^{3+} + 21e^- + 3Mn + 24OH^-$$

The twenty one electrons cancel:

$$7Fe + 12H_2O + 3MnO_4^- \rightarrow 7Fe^{3+} + 3Mn + 24OH^-$$

The above equation is the final balanced redox reaction for the basic solution.

Acidic solution

$$Fe + MnO_4^- \rightarrow Fe^{3+} + Mn$$

The oxidation reaction is as follows:

$$Fe \rightarrow Fe^{3+}$$

The Fe atom is going from 0 to a +3 oxidation state and loses three electrons in the process. We add those three electrons to the product side:

$$Fe \rightarrow Fe^{3+} + 3e^-$$

Balancing the reduction reaction:

$$MnO_4^- \rightarrow Mn$$

The Mn atom is going from +7 to 0 in oxidation number, which requires a gain of seven electrons:

$$7e^- + MnO_4^- \rightarrow Mn$$

We balance the O atoms and then the H atoms:

$$7e^- + MnO_4^- \rightarrow Mn + 4OH^-$$

$$4H^+ + 7e^- + MnO_4^- \rightarrow Mn + 4OH^-$$

We add four H^+ ions to each side to eliminate the OH^- ion in the products; the product species combine to make four water molecules and the number of H^+ ions in the product increases to eight:

$$8H^+ + 7e^- + MnO_4^- \rightarrow Mn + 4H_2O$$

This is the balanced equation for the basic solution.

Now we combine the two balanced half reactions. The oxidation reaction has three electrons, while the reduction reaction has seven. The least common multiple of these two numbers is twenty one, so we multiply the oxidation reaction by 7 and the reduction reaction by 3 so that the electrons are balanced:

$$7 \times [Fe \rightarrow Fe^{3+} + 3e^-] = 7Fe \rightarrow 7Fe^{3+} + 21e^-$$

$$3 \times [8H^+ + 7e^- + MnO_4^- \rightarrow Mn + 4H_2O]$$

$$= 24H^+ + 21e^- + 3MnO_4^- \rightarrow 3Mn + 12H_2O$$

Combining these two equations results in the following equation:

$$7Fe + 24H^+ + 21e^- + 3MnO_4^-$$

$$\rightarrow 7Fe^{3+} + 21e^- + 3Mn + 12H_2O$$

The twenty one electrons cancel:

$$7Fe + 24H^+ + 3MnO_4^- \rightarrow 7Fe^{3+} + 3Mn + 12H_2O$$

The above equation is the final balanced redox reaction for the acidic solution.

11. a. **Basic solution**

$$ClO^- + Ti^{2+} \rightarrow Ti^{4+} + Cl^-$$

The oxidation reaction is as follows:

$$Ti^{2+} \rightarrow Ti^{4+}$$

The Ti atom is going from a +2 to a +4 oxidation state and loses two electrons in the process. We add these 2 electrons to the product side:

$$Ti^{2+} \rightarrow Ti^{4+} + 2e^-$$

Balancing the reduction reaction:

$$ClO^- \rightarrow Cl^-$$

The Cl atom is going from +1 to −1 in oxidation number, which requires a gain of two electrons:

$$ClO^- + 2e^- \rightarrow Cl^-$$

We balance the O atoms and then the H atoms:

$$ClO^- + 2e^- \rightarrow Cl^- + OH^-$$

$$H^+ + ClO^- + 2e^- \rightarrow Cl^- + OH^-$$

We add one OH^- ion to each side to eliminate the H^+ ion in the reactants; the reactant species combine to make a water molecule, and the number of OH^-ions in the product increases to two:

$$H_2O + ClO^- + 2e^- \rightarrow Cl^- + 2OH^-$$

This is the balanced equation for the basic solution.
Now we combine the two balanced half reactions.

$$H_2O + ClO^- + 2e^- + Ti^{2+} \rightarrow Cl^- + 2OH^- + Ti^{4+} + 2e^-$$

The two electrons cancel:

$$H_2O + ClO^- + Ti^{2+} \rightarrow Cl^- + 2OH^- + Ti^{4+}$$

The above equation is the final balanced redox reaction for the basic solution.

Acidic solution

$$ClO^- + Ti^{2+} \rightarrow Ti^{4+} + Cl^-$$

The oxidation reaction is as follows:

$$Ti^{2+} \rightarrow Ti^{4+}$$

The Ti atom is going from a +2 to a +4 oxidation state and loses two electrons in the process. We add these 2 electrons to the product side:

$$Ti^{2+} \rightarrow Ti^{4+} + 2e^-$$

Balancing the reduction reaction:

$$ClO^- \rightarrow Cl^-$$

The Cl atom is going from +1 to −1 in oxidation number, which requires a gain of two electrons:

$$ClO^- + 2e^- \rightarrow Cl^-$$

We balance the O atoms and then the H atoms:

$$ClO^- + 2e^- \rightarrow Cl^- + OH^-$$

$$H^+ + ClO^- + 2e^- \rightarrow Cl^- + OH^-$$

We add one H^+ ion to each side to eliminate the OH^- ion in the products; the product species combine to make a water molecule, and the number of H^+ions in the reactant increases to two:

$$2H^+ + ClO^- + 2e^- \rightarrow Cl^- + 2H_2O$$

This is the balanced equation for the acidic solution.
Now we combine the two balanced half reactions.

$$2H^+ + ClO^- + 2e^- + Ti^{2+} \rightarrow Cl^- + 2H_2O + Ti^{4+} + 2e^-$$

The two electrons cancel:

$$2H^+ + ClO^- + Ti^{2+} \rightarrow Cl^- + 2H_2O + Ti^{4+}$$

The above equation is the final balanced redox reaction for the acidic solution.

b. **Basic solution**

$$BrO_3{}^- + Ag \rightarrow Ag^+ + BrO_2$$

The oxidation reaction is as follows:

$$Ag \rightarrow Ag^+$$

The Ag atom is going from 0 to a +1 oxidation state and loses one electron in the process. We add this 1 electron to the product side:

$$Ag \rightarrow Ag^+ + 1e^-$$

Balancing the reduction reaction:

$$BrO_3{}^- \rightarrow BrO_2$$

The Br atom is going from +5 to +4 in oxidation number, which requires a gain of one electron:

$BrO_3{}^- + 1e^- \rightarrow BrO_2$

We balance the O atoms and then the H atoms:

$BrO_3{}^- + 1e^- \rightarrow BrO_2 + OH^-$

$H^+ + BrO_3{}^- + 1e^- \rightarrow BrO_2 + OH^-$

We add one OH^- ion to each side to eliminate the H^+ ion in the reactants; the reactant species combine to make a water molecule, and the number of OH^-ions in the product increases to two:

$H_2O + BrO_3{}^- + 1e^- \rightarrow BrO_2 + 2OH^-$

This is the balanced equation for the basic solution. Now we combine the two balanced half reactions.

$H_2O + BrO_3{}^- + 1e^- + Ag \rightarrow BrO_2 + 2OH^- + Ag^+ + 1e^-$

The electrons cancel:

$H_2O + BrO_3{}^- + Ag \rightarrow BrO^2 + 2OH^- + Ag^+$

The above equation is the final balanced redox reaction for the basic solution.

Acidic solution

$BrO_3{}^- + Ag \rightarrow Ag^+ + BrO_2$

The oxidation reaction is as follows:

$Ag \rightarrow Ag^+$

The Ag atom is going from 0 to a +1 oxidation state and loses one electron in the process. We add this 1 electron to the product side:

$Ag \rightarrow Ag^+ + 1e^-$

Balancing the reduction reaction:

$BrO_3{}^- \rightarrow BrO_2$

The Br atom is going from +5 to +4 in oxidation number, which requires a gain of one electron:

$BrO_3{}^- + 1e^- \rightarrow BrO_2$

We balance the O atoms and then the H atoms:

$BrO_3{}^- + 1e^- \rightarrow BrO_2 + OH^-$

$H^+ + BrO_3{}^- + 1e^- \rightarrow BrO_2 + OH^-$

We add one H^+ ion to each side to eliminate the OH^- ion in the products; the product species combine to make a water molecule, and the number of H^+ions in the reactants increases to two:

$2H^+ + BrO_3{}^- + 1e^- \rightarrow BrO_2 + H_2O$

This is the balanced equation for the acidic solution. Now we combine the two balanced half reactions.

$2H^+ + BrO_3{}^- + 1e^- + Ag \rightarrow BrO_2 + Ag^+ + H_2O + 1e^-$

The electrons cancel:

$2H^+ + BrO_3{}^- + Ag \rightarrow BrO_2 + Ag^+ + H_2O$

The above equation is the final balanced redox reaction for the acidic solution.

13. The charges are not properly balanced. The correct balanced equation is

$3Cr^{2+} + Cl_2 \rightarrow 2Cr^{3+} + 2Cl^-$

3. APPLICATIONS OF REDOX REACTIONS: VOLTAIC CELLS

Test Yourself

1. The given reaction is the reverse of this reaction:

   $Na^{+} + e^{-} \rightarrow Na \qquad E_{1/2} = -2.714\ V$

   Therefore, the $E_{1/2}$ of the given reaction is 2.714 V.

Test Yourself

1. The overall redox reaction is formed from these two half reactions:

   $Ni^{2+} + 2e^{-} \rightarrow Ni \qquad E_{1/2} = -0.25\ V$

   $MnO_4^{-} + 8H^{+} + 5e^{-} \rightarrow Mn^{2+} + 4H_2O \qquad E_{1/2} = 1.51\ V$

   First equation is reversed:

   $5Ni \rightarrow 5Ni^{2+} + 10e^{-} \qquad E_{1/2} = 0.25\ V$

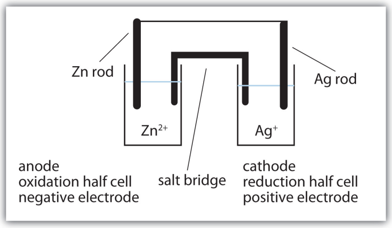
   Combining the two voltages of the half reactions:

   $E = 1.51 + 0.25 = 1.76\ V$

   Because the overall voltage is positive, the reaction is spontaneous as written.

Exercises

1.



3. The given reaction is the reverse of this reaction:

   $F_2 + 2e^{-} \rightarrow 2F^{-} \qquad E_{1/2} = 2.87\ V$

   Therefore, the $E_{1/2}$ of the given reaction is −2.87 V.

5. Voltaic cell: $Zn + 2Ag^{+} \rightarrow Zn^{2+} + 2Ag$

   The overall redox reaction of the cell is formed from these two half reactions:

   $Zn^{2+} + 2e^{-} \rightarrow Zn \qquad E_{1/2} = -0.76\ V$

   $Ag^{+} + e^{-} \rightarrow Ag \qquad E_{1/2} = 0.80V$

   Reversing the first equation:

   $Zn \rightarrow Zn^{2+} + 2e^{-} \qquad E_{1/2} = 0.76V$

   $Ag^{+} + e^{-} \rightarrow Ag \qquad E_{1/2} = 0.80V$

   Combining the two voltages of the half reactions:

   $E = 0.76 + 0.80 = 1.56V$

7. Aluminum is being oxidized, and lithium is reduced:

$Al \rightarrow Al^{3+}$ oxidation

$Li^{+} \rightarrow Li$ reduction

Then we include the appropriate number of electrons on the proper side to balance the charges for each reaction:

$Li^{+} + 1e^{-} \rightarrow Li$

$Al \rightarrow Al^{3+} + 3e^{-}$

We can combine the two final reactions, by multiplying the first equation by 3:

$Al + 3Li^{+} + 3e^{-} \rightarrow 3Li + 3e^{-} + Al^{3+}$

The overall, balanced redox reaction is

$Al + 3Li^{+} \rightarrow 3Li + Al^{3+}$

The overall redox reaction of the cell is formed from these two half reactions:

$Al + 3e^{-} \rightarrow Al^{3+}$ $\qquad E_{1/2} = -1.66V$

$Li^{+} + 1e^{-} \rightarrow Li$ $\qquad E_{1/2} = -3.045V$

Reverse the first equation:

$Al^{3+} \rightarrow Al + 3e^{-}$ $\qquad E_{1/2} = +1.66V$

Combining the two voltages of the half reactions:

$E = 1.66 + (-3.045) = -1.39V$

Because the overall voltage is negative, the reaction is nonspontaneous.

9. Silver is being oxidized, and copper is reduced:

$Ag + Cl^{-} \rightarrow AgCl$ oxidation

$Cu^{2+} \rightarrow Cu$ reduction

Then we include the appropriate number of electrons on the proper side to balance the charges for each reaction:

$Ag + Cl^{-} \rightarrow AgCl + 1e^{-}$

$Cu^{2+} + 2e^{-} \rightarrow Cu$

After multiplying the first reaction with 2 we can combine the two final reactions:

$2Ag + 2Cl^{-} + Cu^{2+} + 2e^{-} \rightarrow 2AgCl + 2e^{-} + Cu$

The overall balanced redox reaction is

$2Ag + 2Cl^{-} + Cu^{2+} \rightarrow 2AgCl + Cu$

The overall redox reaction of the cell is formed from these two half reactions:

$AgCl + e^{-} \rightarrow Ag + Cl^{-}$ $\qquad E_{1/2} = 0.222V$

$Cu^{2+} + 2e^{-} \rightarrow Cu$ $\qquad E_{1/2} = 0.34V$

Reversing the second equation:

$Cu \rightarrow Cu^{2+} + 2e^{-}$ $\qquad E_{1/2} = -0.34V$

Combining the two voltages of the half reactions:

$E = 0.222 + (-0.34) = -0.12V$

Because the overall voltage is negative, the reaction is nonspontaneous.

11. cathode reaction: $Li^{+} + e^{-} \rightarrow Li$; anode reaction: $Al \rightarrow Al^{3+} + 3e^{-}$

13. cathode reaction: $Cu^{2+} + 2e^{-} \rightarrow Cu$; anode reaction: $Ag + Cl^{-} \rightarrow AgCl + e^{-}$

15. The overall redox reaction of the cell is formed from these two half reactions:

$\mathrm{Ni^{2+} + 2e^{-} \rightarrow Ni} \qquad E_{1/2} = -0.25\mathrm{V}$

$\mathrm{2Au^{+} \rightarrow 2Au + 2e^{-}} \qquad E_{1/2} = x\mathrm{V}$

Reversing the first equation:

$\mathrm{Ni \rightarrow Ni^{2+} + 2e^{-}} \qquad E_{1/2} = 0.25\mathrm{V}$

Combining the two voltages of the half reactions:

$0.33 = x + 0.25$

$x = 0.08$

Standard reduction potential of the $\mathrm{Au^{+} + e^{-} \rightarrow Au}$ half reaction is 0.08 V.

17. oxidized: Zn; reduced: Mn

19. The equation for a silver oxide button battery is represented by

$\mathrm{Zn + Ag_2O \rightarrow ZnO + 2Ag}$

From the equation we can see that Zn is oxidized and Ag is reduced.

21. 5.92 V from the reaction of F_2 and Li

$\mathrm{F_2 + 2e^{-} \rightarrow 2F^{-}} \qquad E_{1/2} = 2.87\mathrm{V}$

$\mathrm{Li^{+} + e^{-} \rightarrow Li} \qquad E_{1/2} = -3.045\mathrm{V}$

Reversing the second reaction:

$\mathrm{Li \rightarrow Li^{+} + e^{-}} \qquad E_{1/2} = 3.045\mathrm{V}$

$E = 2.87 + 3.045 = 5.92\mathrm{V}$

4. ELECTROLYSIS

Exercises

1. an electrochemical cell in which charge is forced through and a nonspontaneous reaction occurs
3. any three of the following: Al, K, Li, Na, Cl_2, or Mg
5. $Cu^{2+} + 2e^- \rightarrow Cu \qquad E_{1/2} = 0.34 V$

   0.34 V must be applied to an electrolytic cell to electroplate copper from Cu^{2+}.

5. END-OF-CHAPTER MATERIAL

Additional Exercises

1. As oxygen is added to magnesium, it is being oxidized. In modern terms, the Mg atoms are losing electrons and being oxidized, while the electrons are going to the O atoms.
3. a. By rule 1, krypton is assigned an oxidation number 0, since it is in its elemental form.
   b. By rule 3, fluorine is assigned an oxidation number of −1.
   According to rule 4, the sum of the oxidation numbers on all atoms must equal the charge on the species, so we have the simple algebraic equation
   $$x + 4(-1) = 0$$
   where x is the oxidation number of the Kr atom and the 0 represents the charge on the species. Evaluating for x,
   $$x = +4$$
   Thus the oxidation number on the Kr atom is +4.
   c. By rule 3, fluorine is assigned an oxidation number of −1.
   According to rule 4, the sum of the oxidation numbers on all atoms must equal the charge on the species, so we have the simple algebraic equation
   $$2x + 2(-1) = 0$$
   where x is the oxidation number of the O atom and the 0 represents the charge on the species. Evaluating for x,
   $$x = +1$$
   Thus the oxidation number on the O atom is +1.
5. a. By rule 3, oxygen is assigned an oxidation number of −2.
   According to rule 4, the sum of the oxidation numbers on all atoms must equal the charge on the species, so we have the simple algebraic equation
   for N_2O_5, $2x + 5(-2) = 0$
   where x is the oxidation number of the N atom and the 0 represents the charge on the species. Evaluating for x,
   $$x = +5$$
   Thus the oxidation number on the N atom is +5.
   b. For N_2O_4, $2x + 4(-2) = 0$
   where x is the oxidation number of the N atom and the 0 represents the charge on the species. Evaluating for x,
   $$x = +4$$
   Thus the oxidation number on the N atom is +4.
   c. For NO2, $x + 2(-2) = 0$
   where x is the oxidation number of the N atom and the 0 represents the charge on the species. Evaluating for x,
   $$x = +4$$
   Thus the oxidation number on the N atom is +4.
   d. For NO, $x + 1(-2) = 0$
   where x is the oxidation number of the N atom and the 0 represents the charge on the species. Evaluating for x,
   $$x = +2$$
   Thus the oxidation number on the N atom is +2.
   e. By rule 3, hydrogen is assigned an oxidation number of +1.
   According to rule 4, the sum of the oxidation numbers on all atoms must equal the charge on the species, so we have the simple algebraic equation
   for N_2H_4, $2x + 4(+1) = 0$
   where x is the oxidation number of the N atom and the 0 represents the charge on the species. Evaluating for x,
   $$x = -2$$
   Thus the oxidation number on the N atom is −2.
   f. For NH_3, $x + 3(+1) = 0$
   where x is the oxidation number of the N atom and the 0 represents the charge on the species. Evaluating for x,
   $$x = -3$$

7. To balance the number of S atoms add the coefficient 8 before SO_2.

$$S_8 + O_2 \rightarrow 8SO_2$$

This imbalances the number of O atoms. Add the coefficient 4 before O_2 to balance the number of O atoms.

$$S_8 + 4O_2 \rightarrow 8SO_2$$

9. $Cr_2O_7{}^{2-} + Fe \rightarrow Cr^{3+} + Fe^{3+}$

The oxidation reaction is as follows:

$$Fe \rightarrow Fe^{3+}$$

The Fe atom is going from 0 to a +3 oxidation state and loses three electrons in the process. We add these electrons to the product side:

$$Fe \rightarrow Fe^{3+} + 3e^-$$

Balancing the reduction reaction:

$$Cr_2O_7{}^{2-} \rightarrow Cr^{3+}$$

The Cr atom is going from +6 to +3 in oxidation number, which requires a gain of three electrons. Since $Cr_2O_7{}^{2-}$ has two Cr atoms, the number of electrons gained is +6:

$$Cr_2O_7{}^{2-} + 6e^- \rightarrow 2Cr^{3+}$$

We balance the O atoms and then the H atoms:

$$Cr_2O_7{}^{2-} + 6e^- \rightarrow 2Cr^{3+} + 7OH^-$$

$$7H^+ + Cr_2O_7{}^{2-} + 6e^- \rightarrow 2Cr^{3+} + 7OH^-$$

We add seven H^+ ions to each side to eliminate the OH^- ion in the products; the product species combine to make seven water molecules, and the number of H^+ ions in the reactant increases to fourteen:

$$14H^+ + Cr_2O_7{}^{2-} + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$$

This is the balanced equation for the acidic solution.

Now we combine the two balanced half reactions. The oxidation reaction has three electrons, while the reduction reaction has six. The least common multiple of these two numbers is six, so we multiply the oxidation reaction by 2 so that the electrons are balanced:

$$2 \times [Fe \rightarrow Fe^{3+} + 3e^-] = 2Fe \rightarrow 2Fe^{3+} + 6e^-$$

Combining the multiplied oxidation reaction and the reduction reaction results in the following equation:

$$2Fe + 14H^+ + Cr_2O_7{}^{2-} + 6e^- \rightarrow 2Fe^{3+} + 6e^- + 2Cr^{3+} + 7H_2O$$

The six electrons cancel:

$$2Fe + 14H^+ + Cr_2O_7{}^{2-} \rightarrow 2Fe^{3+} + 2Cr^{3+} + 7H_2O$$

The above equation is the final balanced redox reaction for the acidic solution.

11. $UO_2{}^{2+} + HN_3 \rightarrow U + N_2$

The oxidation reaction is as follows:

$HN_3 \rightarrow N_2$

The N atom is going from a −1/3 to a 0 oxidation state. Since 2HN3 contains six N atoms, it loses 2 electrons in the process. We add these electrons to the product side:

$2HN_3 \rightarrow 3N_2 + 2e^-$

We balance the H atoms:

$2HN_3 \rightarrow 3N_2 + 2e^- + 2H^+$

Balancing the reduction reaction:

$UO_2{}^{2+} \rightarrow U$

The U atom is going from +6 to 0 in oxidation number, which requires a gain of 6 electrons.

$UO_2{}^{2+} + 6e^- \rightarrow U$

We balance the O atoms and then the H atoms:

$UO_2{}^{2+} + 6e^- \rightarrow U + 2OH^-$

$2H^+ + UO_2{}^{2+} + 6e^- \rightarrow U + 2OH^-$

We add two H^+ ions to each side to eliminate the OH^- ion in the product; the product species combine to make two water molecules, and the number of H^+ ions in the reactant increases to four:

$4H^+ + UO_2{}^{2+} + 6e^- \rightarrow U + 2H_2O$

This is the balanced equation for the acidic solution.

Now we combine the two balanced half reactions. The oxidation reaction has two electrons, while the reduction reaction has six. The least common multiple of these two numbers is six, so we multiply the oxidation reaction by 3 so that the electrons are balanced:

$$\begin{aligned} &3 \times [2HN_3 \rightarrow 3E_2 + 2e^- + 2H^+] \\ &\quad = 6HN_3 \rightarrow 9N_2 + 6e^- + 6H^+ \end{aligned}$$

Combining the multiplied reduction reaction and the oxidation reaction results in the following equation:

$$\begin{aligned} &6HN_3 + 4H^+ + UO_2{}^{2+} + 6e^- \\ &\quad \rightarrow 9N_2 + 6e^- + 6H^+ + U + 2H_2O \end{aligned}$$

The six electrons cancel:

$$6HN_3 + UO_2{}^{2+} \rightarrow 9N_2 + 2H^+ + U + 2H_2O$$

The above equation is the final balanced redox reaction for the acidic solution.

13. The overall redox reaction is formed from these two half reactions:

$$Zn^{2+} + 2e^- \rightarrow Zn \qquad E_{1/2} = -0.76V$$

$$2H^+ + 2e^- \rightarrow H_2 \qquad E_{1/2} = -0.000V$$

Reverse the first reaction:

$$Zn \rightarrow Zn^{2+} + 2e^- \qquad E_{1/2} = 0.76V$$

$$E = 0.76 - 0 = 0.76V$$

The voltage of the reaction is +0.76 V, which implies a spontaneous reaction.

15. $Cr^{2+} \rightarrow Cr^{3+} + 1e^-$ Oxidation

$Cr^{2+} + 2e^- \rightarrow Cr$ Reduction

Multiplying the oxidation reaction by 2 and combining with the reduction reaction results in the equation:

$$2Cr^{2+} + Cr^{2+} + 2e^- \rightarrow 2Cr^{3+} + 2e^- + Cr$$

The electrons cancel and the final equation:

$$3Cr^{2+} \rightarrow Cr + 2Cr^{3+}$$

The overall redox reaction is formed from these two half reactions:

$$Cr^{3+} + 1e^- \rightarrow Cr^{2+} \qquad E_{1/2} = -0.41V$$

$$Cr^{2+} + 2e^- \rightarrow Cr \qquad E_{1/2} = -0.91V$$

Reverse the first equation:

$$Cr^{2+} \rightarrow Cr^{3+} + 1e^- \qquad E_{1/2} = 0.41V$$

$$E = 0.41 + (-0.91) = -0.50V$$

17. $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

19. yes because reduction occurs at the cathode

21. Area to be plated = 55.0 cm^2

Thickness of the award to be plated with gold = 3.00 μm

Converting 3.00 μm to cm:

$$3.00\ \mu m \times \frac{1\ m}{10^6\ \mu m} \times \frac{100\ cm}{1\ m} = 3.00 \times 10^4\ cm^2$$

Volume of the plate = Area × Thickness = 55.0 × 3.00 × 10^{-4} cm^3 = 1.65 × 10^{-2} cm^3

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

$$19.3\ g/cm^3 = \frac{\text{Mass}}{1.65 \times 10^{-2}\ cm^3}$$

Mass = 0.318 g

CHAPTER 15

Nuclear Chemistry

1. RADIOACTIVITY

Test Yourself

1. Polonium has an atomic number of 84, so the parent isotope is represented as $^{208}_{84}\text{Po}$. We represent the alpha particle as $^{4}_{2}\text{He}$ and use subtraction (208 – 4 = 204 and 84 – 2 = 82) to identify the daughter isotope as lead:

   $^{208}_{84}\text{Po} \rightarrow {}^{4}_{2}\text{He} + {}^{204}_{82}\text{Pb}$; daughter isotope: $^{204}_{82}\text{Pb}$

Test Yourself

1. The parent isotope is $^{133}_{43}\text{Tc}$, while one of the products is $^{0}_{-1}\text{e}$. So that the mass and atomic numbers have the same value on both sides, the mass number of the daughter isotope must be 133, and its atomic number must be 44. The element having an atomic number of 44 is ruthenium. Thus the complete nuclear equation is as follows:

   $^{133}_{43}\text{Tc} \rightarrow {}^{133}_{44}\text{Ru} + {}^{0}_{-1}\text{e} + \gamma$; daughter isotope: ruthenium-133

Exercises

1. Radioactivity is the spontaneous emission of particles and electromagnetic radiation from nuclei of unstable atoms.

3. a. $^{11}_{5}\text{B}$

   Number of protons = Atomic number = 5

   Number of neutrons = Mass number – Atomic number = 11 – 5 = 6

   b. $^{27}_{13}\text{Al}$

   Number of protons = Atomic number = 13

   Number of neutrons = Mass number – Atomic number = 27 – 13 = 14

   c. ^{56}Fe

   Iron has an atomic number of 26.

   Number of protons = Atomic number = 26

   Number of neutrons = Mass number – Atomic number = 56 – 26 = 30

   d. ^{224}Rn

   Radon has an atomic number of 86.

   Number of protons = Atomic number = 86

   Number of neutrons = Mass number – Atomic number = 224 – 86 = 138

5. An alpha particle is a collection of two protons and two neutrons and is equivalent to a helium nucleus.

7. Gamma rays are high-energy electromagnetic radiation given off in radioactive decay.

9. Plutonium has an atomic number of 94, so the parent isotope is represented as $^{244}_{94}\text{Pu}$. We represent the alpha particle as $^{4}_{2}\text{He}$ and use subtraction (244 – 4 = 240 and 94 – 2 = 92) to identify the daughter isotope as uranium:

   $^{244}_{94}\text{Pu} \rightarrow {}^{240}_{92}\text{U} + {}^{4}_{2}\text{He}$; daughter isotope: $^{240}_{92}\text{U}$

11. The parent isotope is $^{121}_{50}\text{Sn}$, while one of the products is $^{0}_{-1}\text{e}$. So that the mass and atomic numbers have the same value on both sides, the mass number of the daughter isotope must be 121, and its atomic number must be 51. The element having an atomic number of 51 is antimony. Thus the complete nuclear equation is as follows:

    $^{121}_{50}\text{Sn} \rightarrow {}^{121}_{51}\text{Sb} + {}^{0}_{-1}\text{e}$; daughter isotope: $^{121}_{51}\text{Sb}$

13. In the radioactive decay of radon-222 the gamma radiation emitted has energy of 8.2×10^{-14} J per nucleus decayed. Converting J to MeV:

    $$8.2 \times 10^{-14}\text{J} \times \frac{1\ \text{MeV}}{1.602 \times 10^{-13}\ \text{J}} = 0.51\ \text{MeV}$$

    The energy in megaelectron volts of the gamma ray emitted when radon-222 decays is 0.51 MeV.

15. Beta particles penetrate more. A thick wall of inert matter is sufficient to block both particles.

17. Nuclear fission is the breaking down of large nuclei into smaller nuclei, usually with the release of excess neutrons.

2. HALF-LIFE

Test Yourself

1. If we compare the time that has passed to the isotope's half-life, we note that 240.0 y is exactly 4 half-lives, so $n = 4$. Substituting and solving results in the following:

$$\text{amount remaining} = 0.600\text{ g} \times \left(\frac{1}{2}\right)^4$$

$$\text{amount remaining} = 0.0375\text{ g}$$

Test Yourself

1. Here we identify the initial amount as 0.600 g, $t = 100.0$ y, and $t_{1/2} = 60.0$ y. Substituting into the equation:

   amount remaining = (0.600 g) × $e^{-(0.693)(100.0\text{ y})/60.0\text{ y}}$

   Evaluating the exponent (and noting that the *s* units cancel), we get

   amount remaining = (0.600 g) × $e^{-1.155}$

   Solving, the amount remaining is 0.189 g.

Exercises

1. Only radioactive isotopes have a half-life.
3. Here we identify the initial amount as 1.00 g, final amount as 0.125 g, and $t_{1/2} = 17.0$ d. Substituting into the equation:

$$0.125\text{ g} = (1.00\text{ g}) \times e^{-(0.693)(t)/17.0\text{ d}}$$

$$-(0.693)(t)/17.0\text{ d} = \ln(0.125)$$

$$t = 51.0\text{ d}$$

5. Here we identify the initial amount as 10.0 g, final amount as 1.25 g, and $t = 75$ y. Substituting into the equation:

$$1.25\text{ g} = (10.0\text{ g}) \times \text{e}^{-(0.693)(75)\,/\,t_{1/2}}$$

$$-(0.0693)(75\text{ y})\,/\,t_{1/2} = \ln(0.125)$$

$$t_{1/2} = 25\text{ y}$$

7. Here we identify the initial amount as 0.0002 g, $t_{1/2} = 432$ y, and $t = 100.0$ y. Substituting into the equation:

$$\begin{aligned}\text{final amount} &= (0.0002\text{ g}) \times e^{-(0.693)(100.0\text{ y})/(432\text{ y})}\\ \text{final amount} &= 0.000170\text{ g}\end{aligned}$$

   When $t = 1{,}000.0$ y

   Substituting into the equation:

$$\begin{aligned}\text{final amount} &= (0.0002\text{ g}) \times e^{-(0.693)(1{,}000.0\text{ y})/(432\text{ y})}\\ \text{final amount} &= 0.0000402\text{ g}\end{aligned}$$

9. Radioactive decay is an exponential process, not a linear process.
11. Here we identify the final amount as 8.4×10^{-9} g, $t_{1/2} =$ 5,730 y, and $t = 10{,}670$ y. Substituting into the equation:

$$8.4 \times 10^{-9}\text{ g} = \text{initial amount} \times e^{-(0.693)(10{,}670\text{ y})/(5{,}730\text{ y})}$$

$$\text{initial amount} = 3.1 \times 10^{-8}\text{ g}$$

3. UNITS OF RADIOACTIVITY

Test Yourself

1. The following table shows the activity of the radon sample over multiple half-lives:

Time in Years	Activity
0	60,000 Bq
15 h	30,000 Bq
30 h	15,000 Bq
45 h	7,500 Bq
60 h	3,750 Bq

Over a period of 4 half-lives, the activity of the radon will be halved four times, at which point its activity will be 3,750 Bq. Thus it takes 4 half-lives, or 4 × 15 h = 60 h, for the activity to decrease to 3,750 Bq.

Test Yourself

1. If the initial amount is represented by 60,000 Bq and the final amount is 10,000 Bq, we have 10,000 Bq = (60,000 Bq)$e^{-0.693t/(15\ \text{y})}$

Solve for t,

$$\frac{10{,}000\ \text{Bq}}{60{,}000\ \text{Bq}} = e^{-0.693t/(15\ \text{y})}$$

$$\ln(0.167) = -0.693t/(15\ \text{y})$$

$$t = 38.8\ \text{y}$$

Exercises

1. a unit of radioactive exposure equal to 0.01 J of energy per gram of tissue
3. A becquerel is 1 decay/s, whereas a curie is 3.7×10^{10} decays/s.
5. The following table shows the activity of the radon gas over multiple half-lives:

Time in Years	Activity
0	140.0 mCi
1,500 y	70.0 mCi
3,000 y	35.0 mCi
4,500 y	17.5 mCi
6,000 y	8.75 mCi

Over a period of 4 half-lives, the activity of the radon gas will be halved four times, at which point its activity will be 8.75 mCi. Thus it takes 4 half-lives, or 4 × 1,500 y = 6,000 y, for the activity to decrease to 8.75 mCi.

7. 1 curie = 3.7×10^{10} disintegrations/s

Converting μCi to disintegrations/s:

$$65\ \mu\text{Ci} \times \frac{1\ \text{Ci}}{10^{6}\ \mu\text{Ci}} \times 3.7 \times 10^{10}\ \frac{\text{disintegrations/s}}{1\ \text{Ci}}$$

$$= 2.41 \times 10^{6}\ \text{disintegrations/s}$$

9. If the initial amount is represented by 2.44 mCi, we have

$$\text{final amount} = (2.44\ \text{mCi})e^{-0.693(50.0\ \text{s})/(11.0\ \text{s})}$$

$$\text{final amount} = 0.104\ \text{mCi}$$

11. If the initial amount is represented by 100.0 mCi and the final amount is 10.0 mCi, we have

$$10.0\ \text{mCi} = (100\ \text{mCi})e^{-0.693(t)/(11.0\ \text{y})}$$

$$\ln(0.1) = 0.693(t)/(11.0\ \text{y})$$

$$t = 36.5\ \text{s}$$

13. by using a film badge, which is exposed by the radiation, or a Geiger counter
15. Radioactive atoms in the body, most terrestrial sources, cosmic sources, and nuclear energy sources are likely unavoidable, which is about 27% of the total exposure. If exposure to radon gas is added, the total unavoidable exposure increases to 82%.

17. Film is exposed by the radiation. The more radiation film is subjected to, the more exposed it becomes.

4. USES OF RADIOACTIVE ISOTOPES

Exercises

1. A tracer is a radioactive isotope that can be detected far from its original source to trace the path of certain chemicals. Hydrogen-3 can be used to trace the path of water underground.
3. If the initial amount of a radioactive isotope is known, then by measuring the amount of the isotope remaining, a person can calculate how old that object is since it took up the isotope.
5. If the initial amount is represented by 14.0 Bq and the final amount is 3.5 Bq, we have

$$3.5\text{ Bq} = (14.0\text{ Bq})e^{-0.693(t)/(5{,}730\text{ y})}$$

$$\ln(0.25) = -0.693(t)/(5{,}730\text{ y})$$

$$t = 11{,}500\text{ y}$$

7. increased shelf life (answers will vary)
9. The thyroid gland absorbs most of the iodine, allowing it to be imaged for diagnostic purposes or preferentially irradiated for treatment purposes.
11. gamma rays

5. NUCLEAR ENERGY

Test Yourself

1. Using the data given, we can write the following initial equation:

$${}^{1}_{0}\text{n} + {}^{238}_{92}\text{U} \rightarrow {}^{135}_{55}\text{Cs} + {}^{96}_{37}\text{Rb} + ?{}^{1}_{0}\text{n}$$

In balanced nuclear equations, the sums of the subscripts on each sides of the equation are the same, as are the sums of the superscripts. The subscripts are already balanced: 0 + 92 = 92 and 55 + 37 = 92. The superscripts on the left equal 239 (1 + 238) but equal 231 (135 + 96) on the right. We need eight more mass number units on the right. Eight neutrons should be products of the process for the mass numbers to balance. Thus the balanced nuclear equation is as follows:

$${}^{1}_{0}\text{n} + {}^{238}_{92}\text{U} \rightarrow {}^{135}_{55}\text{Cs} + {}^{96}_{37}\text{Rb} + 8{}^{1}_{0}\text{n}$$

Test Yourself

1. We start by adding the masses of all species on each side of the nuclear equation. Then we determine the difference in mass as the reaction proceeds and convert this to an equivalent amount of energy. The total mass of the reactants is as follows:

$$1.0087 + 238.0508 = 239.0595 \text{ g}$$

The total mass of the products is as follows:

$$95.9342 + 134.9060 + (8 \times 1.0087) = 238.9098 \text{ g}$$

The change in mass is determined by subtracting the mass of the reactants from the mass of the products:

$$\text{change the mass} = 239.0595 - 238.9098 = 0.1497 \text{ g}$$

This mass change must be converted into kilogram units:

$$0.1497 \text{ g} \times \frac{1 \text{ kg}}{1{,}000 \text{ g}} = 0.0001497 \text{ kg}$$

Now we can use Einstein's equation to determine the energy change of the nuclear reaction:

$$E = -(0.0001497 \text{ kg})(3.00 \times 10^{8} \text{ m/s})^{2} = -1.35 \times 10^{13} \text{ J}$$

Exercises

1. Using Einstein's equation,

$$E = -(0.00100 \text{ kg})(3.00 \times 10^{8} \text{ m/s})^{2} = -9 \times 10^{13} \text{ J}$$

3. Change in mass = mass lost = 0.1002 g

This mass change must be converted into kilogram units:

$$0.1002 \text{ g} \times \frac{1 \text{ kg}}{1{,}000 \text{ g}} = 0.0001002 \text{ kg}$$

Now we can use Einstein's equation to determine the energy change of the nuclear reaction:

$$E = (0.0001002 \text{ kg})(3.00 \times 10^{8} \text{ m/s})^{2} = 9.02 \times 10^{12} \text{ J}$$

5. Change in mass = mass lost = 0.1002 g

This mass change must be converted into kilogram units:

$$0.1002 \text{ g} \times \frac{1 \text{ kg}}{1{,}000 \text{ g}} = 0.0001002 \text{ kg}$$

Now we can use Einstein's equation to determine the energy change of the nuclear reaction:

$$E = (0.0001002 \text{ kg})(3.00 \times 10^{8} \text{ m/s})^{2} = 9.02 \times 10^{12} \text{ J}$$

In terms of the number of grams of lead-208 reacted:

$$\frac{9.02 \times 10^{12} \text{J}}{208 \text{ g}} = 4.34 \times 10^{10} \text{J/g}$$

7. We start by adding the masses of all species on each side of the nuclear equation. Then we determine the difference in mass as the reaction proceeds and convert this to an equivalent amount of energy.

   Total mass of the reactants: 241.0569 g

   Total mass of the products: 139.9106 + 89.9077 + (11 × 1.0087) = 240.914 g

   The change in mass is determined by subtracting the mass of the reactants from the mass of the products:

   change in mass = 241.0569 - 240.914 g = 0.1429 g

   This mass change must be converted into kilogram units:

   $$0.1429 \text{ g} \times \frac{1 \text{ kg}}{1{,}000 \text{ g}} = 0.0001429 \text{ kg}$$

   Now we can use Einstein's equation to determine the energy change of the nuclear reaction:

   $$E = -(0.0001429 \text{ kg})(3.00 \times 10^8 \text{ m/s})^2 = -1.28 \times 10^{13} \text{ J}$$

9. Change in mass = mass lost = 0.01888 g

   This mass change must be converted into kilogram units:

   $$0.01888 \text{ g} \times \frac{1 \text{ kg}}{1{,}000 \text{ g}} = 0.00001888 \text{ kg}$$

   Now we can use Einstein's equation to determine the energy change of the nuclear reaction:

   $$E = (0.00001888 \text{ kg})(3.00 \times 10^8 \text{ m/s})^2 = 1.70 \times 10^{12} \text{ J}$$

11. A nuclear reactor controls a nuclear reaction to produce energy in usable amounts. The energy produced generates steam, which is used to turn a turbine that generates electricity for general use.

13. a process that generates more reaction pathways for each previous reaction

6. END-OF-CHAPTER MATERIAL

Additional Exercises

1. Acids can dissolve many metals; a spilled acid can lead to contamination.
3. Convert grams of Th^{3+} to mol of Th^{3+} and then using the balanced chemical equation convert mol of Th^{3+} to mol of I^{-}:

$$0.567\ \text{g Th}^{3+} \times \frac{1\ \text{mol Th}^{3+}}{232.04\ \text{g Th}^{3+}} \times \frac{3\ \text{mol I}^{-}}{1\ \text{mol Th}^{3+}} = 0.00733\ \text{mol I}^{-}$$

$$\text{Molarity} = \frac{\#\ \text{mol I}^{-}}{\text{liters of solution}}$$

$$0.500\ \text{M} = \frac{0.00733\ \text{mol I}^{-}}{\text{liters of solution}}$$

liters of solution = 0.0147 L

Converting L to mL:

$$0.0147\ \text{L} \times \frac{1{,}000\ \text{mL}}{1\ \text{L}} = 14.7\ \text{mL}$$

5. The parent isotope is ${}^{90}_{38}Sr$, while one of the products is ${}^{0}_{-1}e$. So that the mass and atomic numbers have the same value on both sides, the mass number of the daughter isotope must be 90, and its atomic number must be 39. The element having an atomic number of 39 is yitrium. Thus the complete nuclear equation is as follows:

$${}^{90}_{38}\text{Sr} \rightarrow {}^{90}_{39}\text{Y} + {}^{0}_{-1}\text{e}$$

7.

1 U mass	238.03 u	238.03 g
14 O masses	14 × 16.00 u =	224.00 g
2 N masses	2 × 14.01 u =	28.02 g
12 H masses	12 × 1.00 u =	12.00 g
Total		502.15 g

9. 0.0117% radioactive potassium-40:

$$\frac{0.0117}{100} \times 600\ \text{mg} \times \frac{1\ \text{g}}{1{,}000\ \text{mg}} = 7.02 \times 10^{-5}\ \text{g}$$

If 1 g of potassium-40 has an activity of 2.626×10^{5} Bq, the banana has an activity of:

$$7.02 \times 10^{-5}\ \text{g} \times \frac{2.626 \times 10^{5}\ \text{Bq}}{1\ \text{g}} = 18.4\ \text{Bq}$$

11. U atoms on both the sides of the equation are balanced. To balance H atoms add the coefficient 2 before HF.

$$UF_6 + H_2O \rightarrow UO_2F_2 + 2HF$$

To balance the number of O atoms add the coefficient 2 before H_2O.

$$UF_6 + 2H_2O \rightarrow UO_2F_2 + 2HF$$

This imbalances the number of H atoms. To balance the number of H atoms double the coefficient of HF.

$$UF_6 + 2H_2O \rightarrow UO_2F_2 + 4HF$$

The number of F atoms on both the sides is balanced. Now the equation is balanced.

13. Initial amount = 100%

Final amount =100 − 99.0 = 1%

$t_{1/2} = 12.3$ y

Substituting into the equation:

$$1\% = (100\%) \times e^{-(0.693)(t)/12.3\ \text{y}}$$

$$\ln(0.01) = -(0.693)(t)/12.3\ \text{y}$$

$$t = 81.7\ \text{y}$$

15. $t_{1/2} = 1.9 \times 10^{19}$ y

$t = 1.38 \times 10^{10}$ y

Using the equation,

$$\text{final amount} = \text{initial amount} \times e^{-(0.693)(1.38 \times 10^{10}\text{y}) \big/ 1.9 \times 10^{19}\text{y}}$$

$$\text{final amount} = \text{initial amount} \times 0.9999999995$$

$$\%\ \text{bismuth decayed} = \frac{\text{initial amount} - \text{final amount}}{\text{initial amount}} \times 100$$

$$= \frac{\text{initial amount} - 0.9999999995\ \text{initial amount}}{\text{initial amount}} \times 100$$

$$= 5.033 \times 10^{-8}\%$$

17. Radioactive atoms in the body, terrestrial sources, and cosmic sources are truly involuntary, which is about 27% of the total. Radon exposure, medical sources, consumer products, and even nuclear energy sources can be avoided.

CHAPTER 16

Organic Chemistry

1. HYDROCARBONS

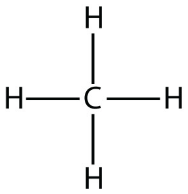
Test Yourself

1. With six C atoms, we will use the *hex-* stem, and with a C—C double bond, this is an alkene, so this molecule is a hexene. In numbering the C atoms, we use the number 3 because it is the lowest possible label. So this molecule is named 3-hexene.

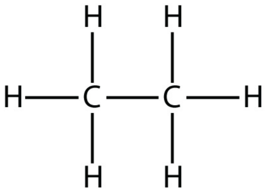
Exercises

1. an organic compound composed of only carbon and hydrogen; aliphatic hydrocarbons and aromatic hydrocarbons
3. a. aliphatic; alkane
   b. aromatic
   c. aliphatic; alkene
5. a. aliphatic; alkane
   b. aliphatic; alkene
   c. aromatic

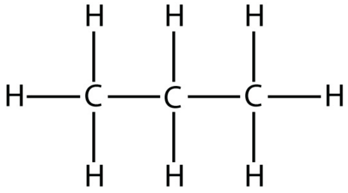
7. The four smallest alkanes are methane, ethane, propane and butane. Their structures are:



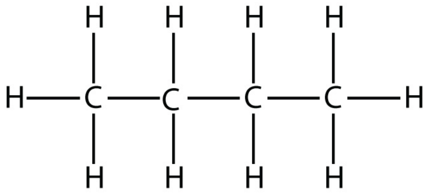
Methane



Ethane

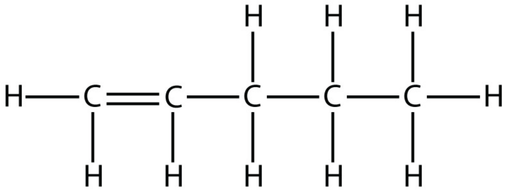


Propane

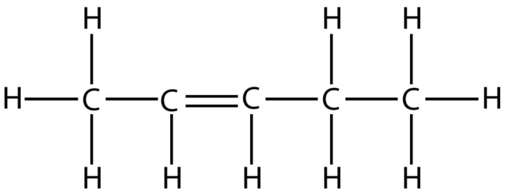


Butane

9. Aromatic means that the molecule has a benzene ring.
11. The 1 is not necessary. The name of the compound is simply propene.
13. Pentene has two isomers:



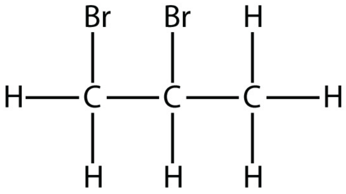
1-Pentene



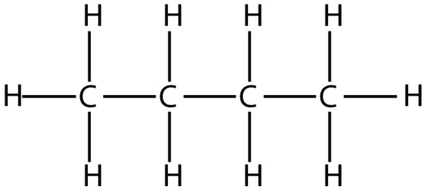
2-Pentene

15. $CH_4 + Br_2 \rightarrow CH_3Br + HBr$

17. This is an addition reaction. Bromine reacts with the C—C double bond of propene and inserts itself onto each C atom.



19. This is a hydrogenation reaction. Hydrogen is added to the double bond of 1-butene to form butane.



21. $C_7H_{16} + 11O_2 \rightarrow 7CO_2 + 8H_2O$

2. BRANCHED HYDROCARBONS

Test Yourself

1. The longest continuous carbon chain has five C atoms, so this molecule will be named as a pentane. There is a one-carbon substituent on the main chain, which is a methyl group. To give the substituent the lowest numbering, we number the chain from the *right* side and see that the substituent is on the second C atom. So this hydrocarbon is 2-methylpentane.

Test Yourself

1. The longest chain has eight C atoms, so we name this molecule as an octane. We find two three-carbon substituents on the fourth C atom and one three-carbon substituent on the fifth C atom. So this molecule is named 4,4,5-tripropyloctane.

Test Yourself

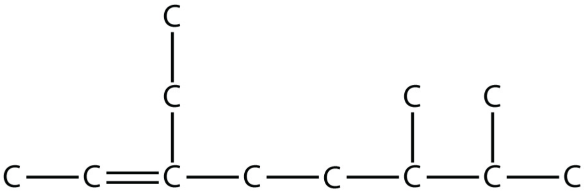
1. The longest chain that contains the C—C double bond has five C atoms, so this is a pentene molecule. The double bond starts at the second C atom, so this is a 2-pentene. Finally, there are three methyl groups on the chain; we number the chain from the left side so that the double bond gets the lowest number. So the name of this molecule is 2,3,4-trimethyl-2-pentene.

Test Yourself

1. First, we draw the eight-carbon backbone with a double bond at the second carbon:

$$\mathrm{C{-}C{=}C{-}C{-}C{-}C{-}C{-}C}$$

According to the name, there is an ethyl group attached to the third C atom and two one-carbon methyl groups attached to the sixth and seventh C atoms in the chain. We finish the carbon backbone by putting the two methyl groups and an ethyl group on the octene main chain:



$$\begin{array}{ccccccccccccccc} & & & & \mathrm{C} & & & & & & & & & & \\ & & & & | & & & & & & & & & & \\ & & & & \mathrm{C} & & & & & & \mathrm{C} & & \mathrm{C} & & \\ & & & & | & & & & & & | & & | & & \\ \mathrm{C} & - & \mathrm{C} & = & \mathrm{C} & - & \mathrm{C} & - & \mathrm{C} & - & \mathrm{C} & - & \mathrm{C} & - & \mathrm{C} \end{array}$$

Exercises

1. A branched hydrocarbon does not have all of its C atoms in a single row.
3. The longest chain that contains the C—C double bond has six C atoms, so this is a hexene molecule. The double bond starts at the second C atom, so this is a 2-hexene. Finally, there is one methyl group on the chain; we number the chain from the left side so that the double bond gets the lowest number. So the name of this molecule is 3-methyl-2-hexene.
5. The longest chain that contains the C—C double bond has five C atoms, so this is a pentene molecule. The double bond starts at the first C atom, so this is a 1-pentene. Finally, there are two methyl groups on the chain. Because the double bond should get the lowest possible number, we number the chain from the right side. So the name of this molecule is 4,4-dimethyl-1-pentene.
7. The longest chain that contains the C—C double bond has five C atoms, so this is a pentene molecule. The double bond starts at the second C atom, so this is a 2-pentene. Finally, there are two methyl groups on the chain; to give them the lowest possible number, we number the chain from the left side. So the name of this molecule is 2,4-dimethyl-2-pentene.

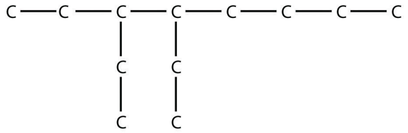
9. The longest chain has eight C atoms, so we name this molecule as an octane. We find two two-carbon substituents on the third and fourth C atoms. So this molecule is named 3,4-diethyloctane.

11. The benzene ring has two substituents- chlorine and bromine. We assign the C atom bearing Br as position 1, and the ring is numbered in a circle to give the other substituent the lowest possible number. Then, Cl is on the fourth C atom. So the name of this molecule is 1-bromo-4-chlorobenzene.

13. a. First we draw the eight-carbon backbone that represents the octane chain:

C—C—C—C—C—C—C—C

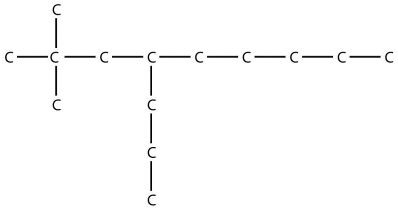
According to the name, there are two ethyl groups attached to the third and fourth C atoms in the chain. We finish the carbon backbone by putting the two ethyl groups on the octane main chain:



b. First we draw the nine-carbon backbone that represents the nonane chain :

C—C—C—C—C—C—C—C—C

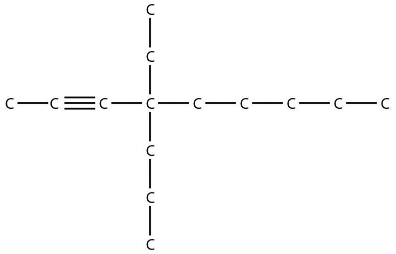
According to the name, there are two methyl groups attached to the second C atom and one propyl group attached to the fourth C atom in the chain. We finish the carbon backbone by putting the two methyl groups and a propyl group on the nonane main chain:



15. a. First we draw the eight-carbon backbone with the triple bond at the second C atom:

C—C≡C—C—C—C—C—C

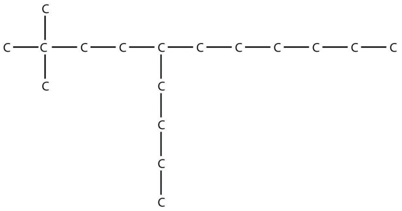
According to the name, there is an ethyl group and a propyl group attached to the fourth C atom in the chain. We finish the carbon backbone by putting the ethyl and the propyl groups on the octyne main chain:



b. First we draw the ten-carbon backbone which represents the decane chain:

C—C—C—C—C—C—C—C—C—C

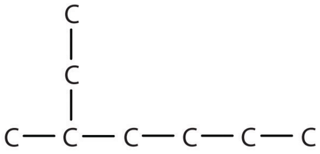
According to the name, there are two methyl groups attached to the second C atom and a butyl group attached to the fifth C atom in the chain. We finish the carbon backbone by putting the methyl groups and the butyl group on the decane main chain:



17. First we draw the six-carbon backbone that represents hexane main chain.

C—C—C—C—C—C

According to the name, there is an ethyl group attached to the second C atom. We finish the carbon backbone by putting the ethyl group on the hexane main chain:



3-methylhptane

In this structure we notice that the longest chain is the seven-carbon chain with a methyl substituent at the third C atom; and not the six-carbon chain. So, the name of this molecule is 3-methylheptane.

3. ALKYL HALIDES AND ALCOHOLS

Test Yourself

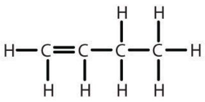
1. The longest carbon chain has three C atoms, so the molecule is propane. There are three bromine substituents- two are located on the first C atom and one is located on the second C atom. The correct name for this molecule is 1,1,2-tribromopropane.

Test Yourself

1. The longest carbon chain containing the OH group has two C atoms, so the parent hydrocarbon is ethane and the alcohol is ethanol. There are three Cl atoms on the second C atom, so the formal name for this alcohol is 2,2,2-trichloroethanol. If naming the alcohol group as a substituent, it would be 2,2,2-trichloro-1-hydroxyethane.

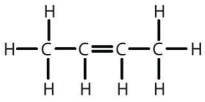
Test Yourself

1. Under these conditions, an HCl molecule will be eliminated, and an alkene will be formed. The H atom may be lost from either of the adjacent carbon atoms. If H is lost from the first C atom, 1-butene is formed; and if it is lost from the second C atom, 2-butene is formed:



1-butene

and



2-butene

Exercises

1. a group of atoms with a certain reactivity; halogen atoms and alcohol groups (answers will vary).
3. The longest carbon chain has four C atoms, so the molecule is butane. There is one bromine substituent located on the second C atom. The correct name for this molecule is 2-bromobutane.
5. The longest carbon chain has seven C atoms, so the molecule is heptane. There is one chlorine substituent located on the second C atom; one fluorine substituent on the third C atom; and a methyl group on the fourth C atom. The correct name for this molecule is 2-chloro-3-fluoro-4-methylheptane.
7. The longest carbon chain containing the OH group has three C atoms, so the parent hydrocarbon is propane. Because the OH group is on the second C atom, it is 2-propanol. There is one methyl substituent as well on the second C atom. So the formal name for this alcohol is 2-methyl-2-propanol. If naming the alcohol group as a substituent, it would be 2-hydroxy-2-methylpropane.
9. The longest carbon chain containing the OH group has eight C atoms, so the parent hydrocarbon is octane. Because the OH group is on the fourth C atom, it is 4-octanol. There are no other substituents on the chain, so the formal name of the molecule is just 4-octanol. If naming the alcohol group as a substituent, it would be 4-hydroxyoctane.
11. Under these conditions, an HOH (otherwise known as H_2O) molecule will be eliminated, and an alkene will be formed. It does not matter which adjacent carbon loses the H atom; in either case the product will be 2-pentene:

$$C-C=C-C-C$$

13. Under these conditions, an HBr molecule will be eliminated, and an alkene will be formed. The H atom may be lost from either of the adjacent carbon atoms. If H is lost from the second C atom, 2-hexene is formed; and if it is lost from the fourth C atom, 3-hexene is formed:

$$C-C-C-C=C-C$$ and

2-hexene

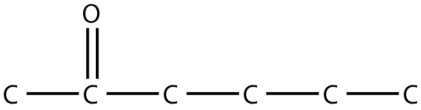
$$C-C-C=C-C-C$$

3-hexene

4. OTHER OXYGEN-CONTAINING FUNCTIONAL GROUPS

Test Yourself

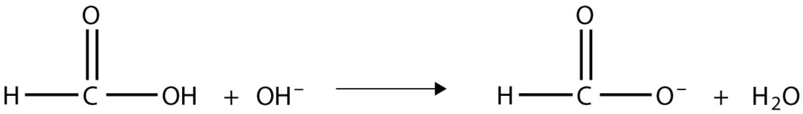
1. This molecule has a one-carbon methyl group and a four-carbon butyl group attached to the carbon of the carbonyl group. So, the molecule has six C atoms in the chain with the second C atom doubly bonded to an oxygen atom. Its structure is as follows:



$$C-\overset{O}{\overset{\Vert}{C}}-C-C-C-C$$

Test Yourself

1. The OH⁻ ion removes the H atom that is part of the carboxyl group:

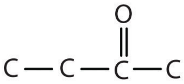


$$H-\overset{O}{\overset{\Vert}{C}}-OH + OH^- \longrightarrow H-\overset{O}{\overset{\Vert}{C}}-O^- + H_2O$$

The anion is the methanoate ion, which is commonly called the formate ion.

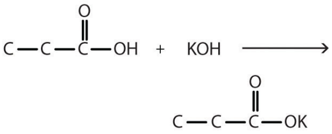
Exercises

1. They both have a carbonyl group, but an aldehyde has the carbonyl group at the end of a carbon chain, and a ketone has the carbonyl group in the middle.
3. a. The longest carbon chain containing the carbonyl group has three C atoms, so the parent hydrocarbon is propane. Since one bond of the carbonyl group is made to an H atom, it is an aldehyde. So, the name of the molecule is propanal.
   b. The longest carbon chain containing the carbonyl group has four C atoms, so the parent hydrocarbon is butane. Since the carbonyl group is in the middle of the carbon chain, it is a ketone. Because the carbonyl group is on the second C atom, the name of the molecule is 2-butanone.
5. a. The longest carbon chain containing the carbonyl group has four C atoms, so the parent hydrocarbon is butane. Since the molecule has a combination of a carbonyl functional group and an OH group, it is a carboxylic acid. There is a methyl group substituted on the third C atom. So, the name of the molecule is 3-methylbutanoic acid.
   b. In this molecule the H atom of a carboxylic acid is replaced by an alkyl group, so it is an ester. The carboxyl group has three C atoms, so the parent carboxylic acid is propanoic acid. The alkyl group is a two-carbon group, so it is ethyl group. The name of this molecule is ethyl propanoate.
7. In this molecule, the functional group is an O atom that is bonded to two organic groups; so it is ether. The O atom is bonded to two alkyl groups; one alkyl group is a two-carbon ethyl group and the other alkyl group is a three-carbon propyl group. So, the name of this molecule is ethyl propyl ether.
9. In 2-butanone, the carbonyl group is attached to a two-carbon ethyl group and a one-carbon methyl group. Hence, the alternate name for 2-butanone is ethyl methyl ketone.



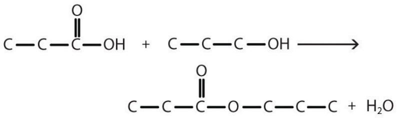
$$C-C-\overset{O}{\overset{\Vert}{C}}-C$$

11. The OH⁻ ion from KOH removes the H atom that is part of the carboxyl group:



$$C-C-\overset{O}{\overset{\Vert}{C}}-OH + KOH \longrightarrow C-C-\overset{O}{\overset{\Vert}{C}}-OK + H_2O$$

13. This molecule has an aromatic ring (benzene ring); a carboxylic acid group (a combination of a carbonyl group and an OH group); and an ester group (a carboxylic acid group whose H atom is replaced by an alkyl group or an aromatic ring).
15. The propanoic acid reacts with propanol to give an ester - propyl propanoate; a water molecule is removed in this reaction:



$$C-C-\overset{O}{\overset{\Vert}{C}}-OH + C-C-C-OH \longrightarrow C-C-\overset{O}{\overset{\Vert}{C}}-O-C-C-C + H_2O$$

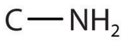
5. OTHER FUNCTIONAL GROUPS

Test Yourself

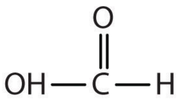
1. This amine has three ethyl groups. Hence, this amine is triethylamine.

Test Yourself

1. The structures of methylamine and formic acid are as follows:

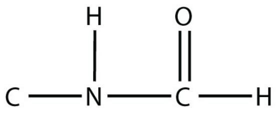


methyl amine



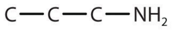
formic acid

When they come together to make an amide, an H_2O molecule is lost, and the N of the amine group bonds to the C of the carboxyl group. The resulting molecule is as follows:

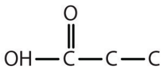


Exercises

1. CH_3NH_2; methylamine
3. a. In this molecule the nitrogen atom is bonded to only one C atom (one R group). So, it is a primary amine.
   b. In this molecule the nitrogen atom is bonded to three C atoms (three R groups). So, it is a tertiary amine.
   c. In this molecule the nitrogen atom is bonded to two C atoms (two R groups). So, it is a secondary amine.
5. a. $C_3H_3CO_2HSHNH_2 + HCl \rightarrow C_3H_3CO_2HSHNH_3Cl$
   b. $(C_6H_5)(C_2H_5)(CH_3)N + HCl \rightarrow (C_6H_5)(C_2H_5)(CH_3)NHCl$
   c. $(C_2H_5)(CH_3)NH + HCl \rightarrow (C_2H_5)(CH_3)NH_2Cl$
7. a. This amine has a methyl group and an ethyl group. Listing the names in alphabetical order, this amine is ethylmethylamine.
   b. This amine has only a phenyl group attached to the N atom, so this amine is phenylamine.
9. An amide functional group is a combination of an amine group and a carbonyl group. In this peptide chain, there are two amide bonds.
11. The structures of propylamine and propanoic acid are as follows:

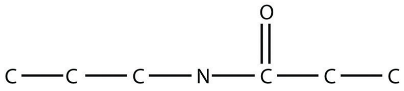


propyl amine



propanoic acid

When they come together to make an amide, an H_2O molecule is lost, and the N of the amine group bonds to the C of the carboxyl group. The resulting molecule is as follows:

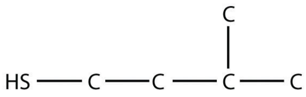


13. a. In this molecule, the S atom of thiol group is bonded to C atom of cyclohexane, so the name of this molecule is cyclohexanethiol.
    b. In this molecule, the thiol group is bonded to a four-carbon chain, so the name of this molecule is butanethiol.

15. First we draw a four-carbon chain for the parent hydrocarbon butane:

$$\mathrm{C{-}C{-}C{-}C}$$

Now we place a methyl group on the third C atom of the chain and the thiol group on the first C atom of the chain. Thus, the structure of the molecule is:

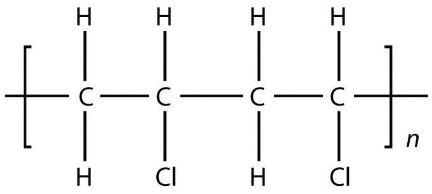


$$\mathrm{HS{-}C{-}C{-}\overset{\overset{\displaystyle C}{|}}{C}{-}C}$$

6. POLYMERS

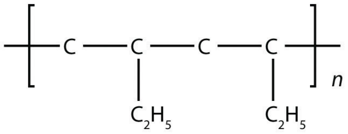
Test Yourself

1. In the case of this monomer, the double bond opens up and joins to other monomers, just as with ethylene. The polymer that is made has this structure:

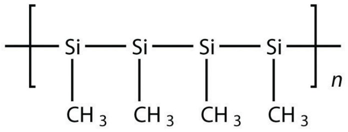


Exercises

1. A polymer is many monomers bonded together.
3. The monomer is butene. The double bond in butene opens up and joins other monomers to form the polymer polybutene:



5. In an addition polymer, no small molecule is given off as a product, whereas in a condensation polymer, small parts of each monomer come off as a small molecule.
7. solubility in H2O and other solvents, melting point, flammability, color, hardness, transparency, film thickness, wetability, surface friction, moldability, and particle size (answers will vary)
9. In the case of this monomer, the double bond opens up and joins to other monomers. The polymer that is made has a silicon backbone which has this structure:

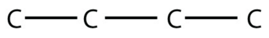


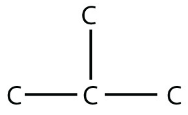
11. Starch is composed of many glucose monomer units.
13. Proteins are polymers of amino acids, which act as the monomers.

7. END-OF-CHAPTER MATERIAL

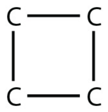
Additional Exercises

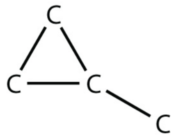
1. cyclopropane, cyclobutane, and cyclopentane
3. The carbon backbone of all noncyclic alkanes with only four C atoms may be either a straight long chain or a branched one:



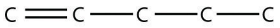


5.



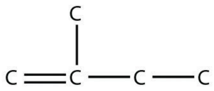


7.

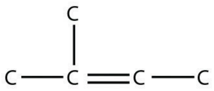


1-pentene

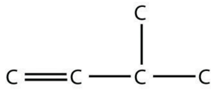
2-pentene



2-methyl-1-butene



2-methyl-1-butene



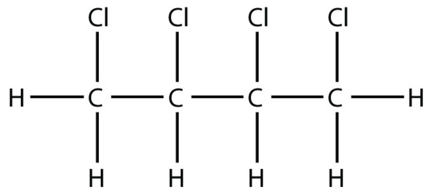
3-methyl-1-butene

9.

Both molecular formulas are C_4H_6.

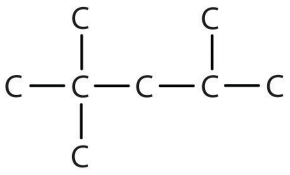
11. $C_2H_6 + 6Cl_2 \rightarrow C_2Cl_6 + 6HCl$

13. When butadiene reacts with chlorine, the double bonds are broken and Cl atoms are added to the C atoms initially involved in the double bond. The product is:

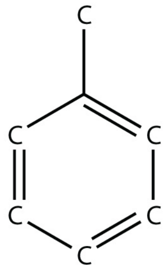


15. two

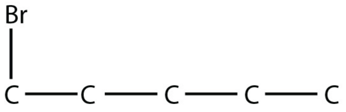
17. First we draw a five-carbon backbone representing pentane. Then we place two methyl groups on the second C atom and one on the fourth C atom. The structure of this molecule is:

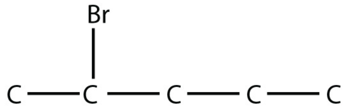


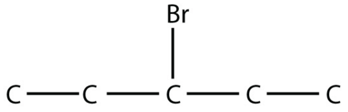
19. The structure of the parent compound toluene is:



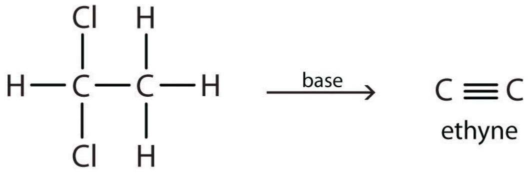
21. The straight chain isomers of bromopentane have a five-carbon backbone. The isomers are 1-bromopentane, 2-bromopentane and 3-bromopentane:



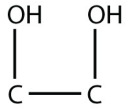




23. The elimination of one molecule of HCl results in the formation of a double bond. A successive elimination of another molecule of HCl results in the formation of a triple bond. Hence, the final product of the *double* elimination of HCl from 1,1-dichloroethane is ethyne.

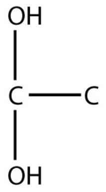


25. The elimination of 1,2-dihydroxyethane or 1,1-dihydroxyethane would give ethyne.



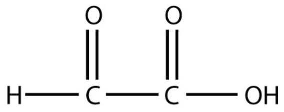
$\mathrm{HO{-}C{-}C{-}OH}$

or



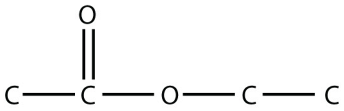
$\mathrm{C{-}C(OH)_2}$

27.



$\mathrm{H{-}C({=}O){-}C({=}O){-}OH}$

29. Ethyl acetate is an ester obtained from the combination of acetic acid (i.e. ethanoic acid) and ethyl alcohol. So the carboxyl group of the ester has two C atoms and the alkyl group also has two C atoms. The structure of ethyl acetate is:



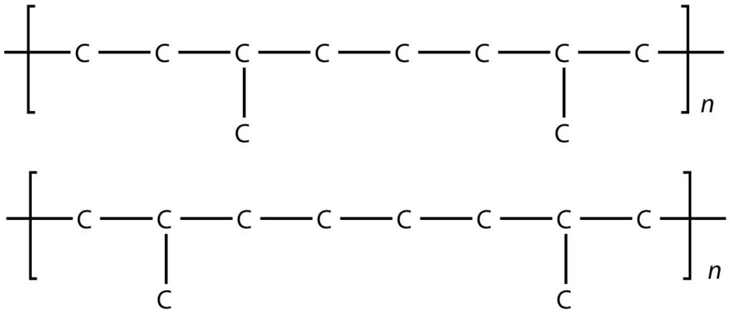
$\mathrm{C{-}C({=}O){-}O{-}C{-}C}$

31. The ether functional group has an O atom attached to two organic groups. In diethyl ether, the O atom is attached to two ethyl groups, so the structure of this molecule is:

$\mathrm{C{-}C{-}O{-}C{-}C}$

33. $(CH_3)_3N + HCl \rightarrow (CH_3)_3NHCl$

35. (answers will vary)



37.