14.1 Formation and Combustion Reactions

Section Preview/ Specific Expectations

In this section, you will

- **write** balanced chemical equations for the complete and incomplete combustion of hydrocarbons
- **perform** an experiment to produce and burn a hydrocarbon
- **recognize** the importance of hydrocarbons as fuels and as precursors for the production of petrochemicals
- **identify** the risks and benefits of the uses of hydrocarbons, for society and the environment
- **communicate** your understanding of the following terms: *complete combustion*, *incomplete combustion*

One of our most common uses of hydrocarbons is as fuel. (See Figure 14.1.) The combustion of fossil fuels gives us the energy we need to travel and to keep warm in cold climates. Fossil fuel combustion is also an important source of energy in the construction and manufacturing industries. As well, many power plants burn natural gas when generating electricity. Even goldsmiths use hydrocarbons, such as butane, as a heat source when crafting gold jewellery. At home, we often burn fossil fuels, such as natural gas, to cook our food.

How do we get energy from these compounds? In this section, you will learn how complete and incomplete combustion can be expressed as chemical equations. Combustion in the presence of oxygen is a *chemical* property of all hydrocarbons. (In Chapter 13, you learned about some *physical* properties of hydrocarbons, such as boiling point and solubility.)

Figure 14.1 How are fossil fuels being used in these photographs?

Chemistry Bulletin

Science | Technology | Society | Environment

Lamp Oil and the Petroleum Age

Abraham Gesner was born in 1797 near Cornwallis, Nova Scotia. Although Gesner became a medical doctor, he was much more interested in fossils. Gesner was fascinated by hydrocarbon substances, such as coal, asphaltum (asphalt), and bitumen. These substances were formed long ago from fossilized plants, algae, fish, and animals.

When Gesner was a young man, the main light sources available were fire, candles, and whale oil lamps. Gesner had made several trips to Trinidad. He began to experiment with asphaltum, a semisolid hydrocarbon from Trinidad's famous "pitch lake." In 1846, while giving a lecture in Prince Edward Island, he startled his audience by lighting a lamp that was filled with a fuel he had distilled from asphaltum. Gesner's lamp fuel gave more light and produced less smoke than any other lamp fuel the audience had ever seen used.

Gesner needed a more easily obtainable raw material to make his new lamp fuel. He tried a solid, black, coal-like bitumen from Albert County, New Brunswick. This substance, called albertite, worked better than any other substance that Gesner had tested.

Making Kerosene

One residue from Gesner's distillation process was a type of wax. Therefore, he called his lamp fuel *kerosolain*, from the Greek word for "wax oil." He soon shortened the name to *kerosene*. To produce kerosene, Gesner heated chunks of albertite in a retort (a distilling vessel with a long downward-bending neck). As the albertite was heated, it gave off vapours. The vapours passed into the neck of the retort, condensed into liquids, and trickled down into a holding tank. Once Gesner had finished the first distillation, he let the tank's contents stand for several hours. This allowed water and solid to settle to the bottom. Then he drew off the oil that remained on top.

Gesner distilled this oil again, and then treated it with sulfuric acid and calcium oxide. Finally he distilled the oil once more.

By 1853, Gesner had perfected his process. In New York, he helped to start the North American Kerosene Gas Light Company. Gesner distinguished between three grades of kerosene: grades A, B, and C. Grade C, he said, was the best lamp oil. Grades A and B could also be burned in lamps, but they were dangerous because they could cause explosions and fires.

Although Gesner never knew, his grades A and B kerosene became even more useful than the purer grade C. These grades were later produced from crude oil, or petroleum, and given a new name: gasoline!

Gesner laid the groundwork for the entire petroleum industry. All the basics of later petroleum refining can be found in his technology.

Making Connections

1. In the early nineteenth century, whales were hunted extensively for their oil, which was used mainly as lamp fuel. When kerosene became widely available, the demand for whale oil decreased. Find out what effect this had on whalers and whales.

2. How do you think the introduction of kerosene as a lamp oil changed people's lives at the time? What conclusions can you draw about the possible impact of technology?

Complete and Incomplete Combustion

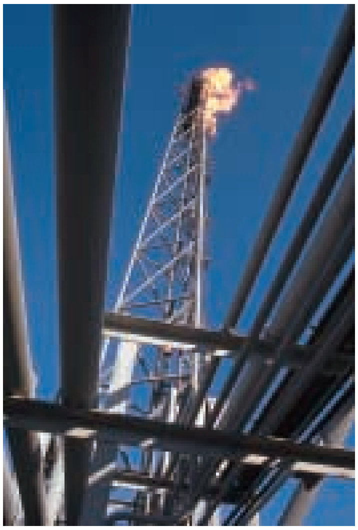

Figure 14.2 Sour gas, $H_2S_{(g)}$, is sometimes "flared off" (burned) from an oil well. This combustion produces sulfur dioxide gas, which reacts with the water in the atmosphere to produce acid rain. Oil companies are now making an effort to reduce this type of pollution.

During a typical combustion reaction, an element or a compound reacts with oxygen to produce oxides of the element (or elements) found in the compound. Figure 14.2 shows an example of a combustion reactions.

Hydrocarbon compounds will burn in the presence of air to produce oxides. This is a chemical property of all hydrocarbons. **Complete combustion** occurs if enough oxygen is present. A hydrocarbon that undergoes complete combustion produces carbon dioxide and water vapour. The following equation shows the complete combustion of propane. (See also Figure 14.3.)

$$C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)}$$

If you burn a fuel, such as propane, in a barbecue, you want complete combustion to occur. Complete combustion ensures that you are getting maximum efficiency from the barbecue. More importantly, toxic gases can result from **incomplete combustion**: combustion that occurs when not enough oxygen is present. During incomplete combustion, other products (besides carbon dioxide and water) can form. The equation below shows the incomplete combustion of propane. Note that unburned carbon, $C_{(s)}$, and carbon monoxide, $CO_{(g)}$, are produced as well as carbon dioxide and water.

$$2C_3H_{8(g)} + 7O_{2(g)} \rightarrow 2C_{(s)} + 2CO_{(g)} + 2CO_{2(g)} + 8H_2O_{(g)}$$

Figure 14.4 shows another example of incomplete combustion.
Go back to the equation for the complete combustion of propane. Notice that the mole ratio of oxygen to propane for the complete combustion (5 mol oxygen to 1 mol propane) is higher than the mole ratio for the incomplete combustion (7 mol oxygen to 2 mol propane, or 3.5 mol oxygen to 1 mol propane). These ratios show that the complete combustion of propane used up more oxygen than the incomplete combustion. In fact, the incomplete combustion probably occurred because not enough oxygen was present. You just learned that incomplete combustion produces poisonous carbon monoxide. This is why you should never operate a gas barbecue or gas heater indoors, where there is less oxygen available. This is also why you should make sure that any natural gas or oil-burning furnaces and appliances in your home are working at peak efficiency, to reduce the risk of incomplete combustion. Carbon monoxide detectors are a good safeguard. They warn you if there is dangerous carbon monoxide in your home, due to incomplete combustion.

Figure 14.4 The yellow flame of this candle indicates that incomplete combustion is occurring. Carbon, $C_{(s)}$, emits light energy in the yellow wavelength region of the visible spectrum.

Balancing Combustion Equations

Have you ever seen a construction worker using an oxyacetylene torch? (See Figure 14.6.) A brilliant white light comes from the torch as it cuts through steel. The intense heat that is associated with this flame comes from the combustion of ethyne, a very common alkyne. Ethyne is also known as *acetylene*.

Figure 14.3 Propane burning in a propane torch: A blue flame indicates that complete combustion is occurring.

How do you write the balanced equation for the complete combustion of acetylene (ethyne)? Complete hydrocarbon combustion reactions follow a general format:

hydrocarbon + oxygen → carbon dioxide + water vapour

You can use this general format for the complete combustion of any hydrocarbon, no matter how large or how small. For example, both acetylene and propane burn completely to give carbon dioxide and water vapour. Each hydrocarbon, however, produces different amounts, or mole ratios, of carbon dioxide and water.

You have seen, written, and balanced several types of reaction equations so far in this textbook. In the following sample problem, you will learn an easy way to write and balance hydrocarbon combustion equations.

Sample Problem

Complete Combustion of Acetylene

Problem

Write the balanced equation for the complete combustion of acetylene (ethyne).

What Is Required?

You need to write the equation. Then you need to balance the atoms of the reactants and the products.

What Is Given?

You know that acetylene (ethyne) and oxygen are the reactants. Since the reaction is a complete combustion reaction, carbon dioxide and water vapour are the products.

Plan Your Strategy

Step 1 Write the equation.

Step 2 Balance the carbon atoms first.

Step 3 Balance the hydrogen atoms next.

Step 4 Balance the oxygen atoms last.

Act on Your Strategy

Step 1 Write the chemical formulas and states for the reactants and products.

ethyne	+	oxygen	→	carbon dioxide	+	water vapour
$C_2H_{2(g)}$	+	$O_{2(g)}$	→	$CO_{2(g)}$	+	$H_2O_{(g)}$

Step 2 Balance the carbon atoms first.

$$C_2H_{2(g)} + O_{2(g)} \rightarrow 2CO_{2(g)} + H_2O_{(g)}$$

(2 carbons) (2 carbons)

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Step 3 Balance the hydrogen atoms next.

$$C_2H_{2(g)} + O_{2(g)} \rightarrow 2CO_{2(g)} + H_2O_{(g)}$$

(2 hydrogens) (2 hydrogens)

Step 4 Balance the oxygen atoms last.

The product coefficients are now set. Therefore, count the total number of oxygen atoms on the product side. Then place an appropriate coefficient in front of the reactant oxygen.

$$C_2H_{2(g)} + ?O_{2(g)} \rightarrow 2CO_{2(g)} + H_2O_{(g)}$$

(4 + 1 oxygens)

You end up with an *odd* number of oxygen atoms on the product side of the equation. When this happens, use a fractional coefficient so that the reactant oxygen balances. Here you have

$$C_2H_{2(g)} + \frac{5}{2}O_{2(g)} \rightarrow 2CO_{2(g)} + H_2O_{(g)}$$

You may prefer to balance the equation with whole numbers. If so, multiply everything by a factor that is equivalent to the denominator of the fraction. Since the fractional coefficient of $O_{2(g)}$ has a 2 in the denominator, multiply *all* the coefficients by 2 to get whole number coefficients.

$$2C_2H_{2(g)} + 5O_{2(g)} \rightarrow 4CO_{2(g)} + 2H_2O_{(g)}$$

Check Your Solution

The same number of carbon atoms appear on both sides of the equation.

The same number of hydrogen atoms appear on both sides of the equation.

The same number of oxygen atoms appear on both sides of the equation.

Sample Problem

Incomplete Combustion of 2,2,4-Trimethylpentane

Problem

2,2,4-trimethylpentane is a major component of gasoline. Write one possible equation for the incomplete combustion of 2,2,4-trimethylpentane.

What Is Required?

You need to write the equation for the incomplete combustion of 2,2,4-trimethylpentane. Then you need to balance the atoms of the reactants and the products. For an incomplete combustion reaction, more than one balanced equation is possible.

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What is given?

You know that 2,2,4-trimethylpentane and oxygen are the reactants. Since the reaction is an incomplete combustion reaction, the products are unburned carbon, carbon monoxide, carbon dioxide, and water vapour.

Plan Your Strategy

Draw the structural diagram for 2,2,4-trimethylpentane to find out how many hydrogen and oxygen atoms it has. Then write the equation and balance the atoms. There are many carbon-containing products but only one hydrogen-containing product, water. Therefore, you need to balance the hydrogen atoms first. Next balance the carbon atoms, and finally the oxygen atoms.

Act on Your Strategy

$$\begin{array}{ccccccccl} & & CH_3 & & & & CH_3 & & \\ & & | & & & & | & & \\ CH_3 & - & C & - & CH_2 & - & CH & - & CH_3 + O_2 \rightarrow C + CO + CO_2 + H_2O \\ & & | & & & & & & \\ & & CH_3 & & & & & & \\ & & (C_8H_{18}) & & & & & & \end{array}$$

Count the carbon and hydrogen atoms, and balance the equation. Different coefficients are possible for the carbon-containing product molecules. Follow the steps you learned in the previous sample problem to obtain the balanced equation shown below.

$$\begin{array}{ccccccccl} & & CH_3 & & & & CH_3 & & \\ & & | & & & & | & & \\ CH_3 & - & C & - & CH_2 & - & CH & - & CH_3 + \frac{15}{2}O_2 \rightarrow 4C + 2CO + 2CO_2 + 9H_2O \\ & & | & & & & & & \\ & & CH_3 & & & & & & \end{array}$$

or

$$\begin{array}{cccccccccl} & & & CH_3 & & & & CH_3 & & \\ & & & | & & & & | & & \\ 2\Big(& CH_3 & - & C & - & CH_2 & - & CH & - & CH_3\Big) + 15O_2 \rightarrow 8C + 4CO + 4CO_2 + 18H_2O \\ & & & | & & & & & & \\ & & & CH_3 & & & & & & \end{array}$$

Check Your Solution

The same number of carbon atoms appear on both sides of the equation.

The same number of hydrogen atoms appear on both sides of the equation.

The same number of oxygen atoms appear on both sides of the equation.

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FROM PAGE 583

Practice Problems

1. The following equation shows the combustion of 3-ethyl-2,5-dimethylheptane:

$$C_{11}H_{24} + 17O_2 \rightarrow 11CO_2 + 12H_2O$$

   (a) Does this equation show *complete* or *incomplete* combustion?

   (b) Draw the structural formula for 3-ethyl-2,5-dimethylheptane.

2. (a) Write a balanced equation for the complete combustion of pentane, C_5H_{12}.

   (b) Write a balanced equation for the complete combustion of octane, C_8H_{18}.

   (c) Write two possible balanced equations for the incomplete combustion of ethane, C_2H_6

3. (a) The flame of a butane lighter is usually yellow, indicating incomplete combustion of the gas. Write a balanced chemical equation to represent the incomplete combustion of butane in a butane lighter. Use the condensed structural formula for butane.

   (b) If you supplied enough oxygen, the butane would burn with a blue flame. Write a balanced chemical equation for the complete combustion of butane.

4. The paraffin wax in a candle burns with a yellow flame. If it had sufficient oxygen to burn with a blue flame, it would burn rapidly and release a lot of energy. It might even be dangerous! Write the balanced chemical equation for the complete combustion of candle wax, $C_{25}H_{52(s)}$.

5. 4-propyldecane burns to give solid carbon, water vapour, carbon monoxide, and carbon dioxide.

   (a) Draw the structural formula for 4-propyldecane.

   (b) Write two different balanced equations for the reaction described in this problem.

   (c) Name the type of combustion. Explain.

Large quantities of acetylene are produced each year by an inexpensive process that combines calcium carbide and water. In the next investigation, you will use this process to produce your own acetylene.

SKILL FOCUS

Predicting

Performing and recording

Analyzing and interpreting

Communicating results

Investigation 14-A

The Formation and Combustion of Acetylene

In this investigation, you will produce acetylene (ethyne) gas by mixing solid calcium carbide with water.

$$CaC_{2(s)} + 2H_2O_{(\ell)} \rightarrow C_2H_{2(g)} + Ca(OH)_{2(s)}$$

Then you will combine the acetylene with different quantities of air to determine the best reaction ratio for complete combustion. Make sure that you follow all the safety precautions given in this investigation and by your teacher.

Question

What is the ideal ratio of fuel to air for the complete combustion of acetylene (ethyne) gas?

Prediction

Air contains 20% oxygen, $O_{2(g)}$. How much air do you think is needed for the complete combustion of acetylene gas? Predict which proportion will react best:

- $\frac{1}{2}$ acetylene to $\frac{1}{2}$ air
- $\frac{1}{3}$ acetylene to $\frac{2}{3}$ air
- $\frac{1}{5}$ acetylene to $\frac{4}{5}$ air
- $\frac{1}{10}$ acetylene to $\frac{9}{10}$ air.

Safety Precautions

- Be careful of the flames from Bunsen burners. Check that there are no flammable solvents close by. If your hair is long, tie it back. Confine loose clothing.

Materials

4 test tubes (100 mL)
4 rubber stoppers
grease pencil
ruler
400 mL beaker
tweezers
matches (or Bunsen burner and splints)
1 or 2 calcium carbide chips
phenolphthalein indicator
limewater
medicine dropper
distilled water
test tube tongs

Procedure

1. Make a table to record your observations. Give your table a title.
2. Mark each test tube with a grease pencil to indicate one of the following volumes: $\frac{1}{2}$, $\frac{1}{3}$, $\frac{1}{5}$, and $\frac{1}{10}$. To find out where to mark the test tube, measure the total length of the test tube with a ruler. Then multiply the length by the appropriate fraction. Measure the fraction from the *bottom* of the test tube.
3. Fill the four test tubes completely with distilled water.
4. Invert the four test tubes in a 400 mL beaker, half full of distilled water. Make sure that the test tubes stay completely full.
5. Add three to five drops of phenolphthalein to the water in the beaker.
6. Using tweezers, drop a small chip of calcium carbide into the water. **CAUTION** Do not touch calcium carbide with your hands!
7. Capture the gas that is produced by holding the test tube marked $\frac{1}{2}$ over the calcium carbide chip. Fill the test tube to the $\frac{1}{2}$ mark with the gas. Remove the test tube from the beaker, still inverted. Let the water drain out. Air will replace the water and mix with the gas in the test tube. Insert a rubber stopper. Invert the test tube a few times to mix the acetylene gas with the air in the test tube.

8. Repeat step 7 with the other three test tubes. Fill each test tube to the volume that you marked on it: $\frac{1}{3}$, $\frac{1}{5}$, or $\frac{1}{10}$. After filling each test tube, remove it from the water and insert a rubber stopper.

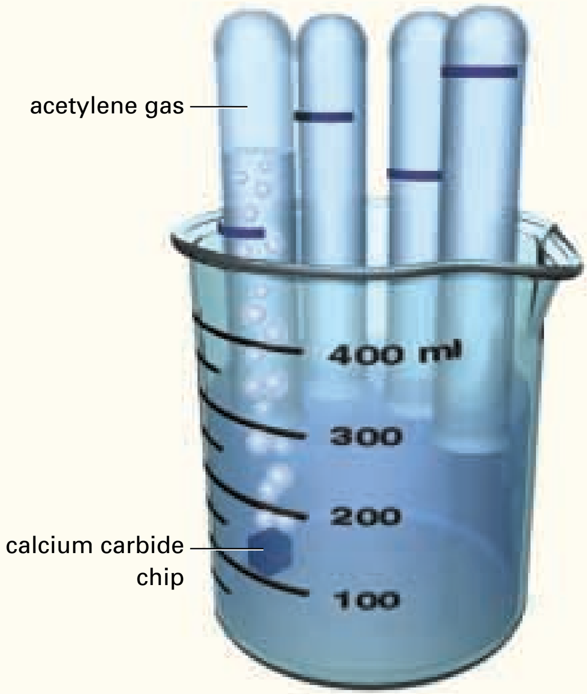

Figure 14.5

9. Invert the first test tube ($\frac{1}{2}$ full). Light a match or a splint. Using test tube tongs, hold the test tube inverted, and take out the stopper. Ignite the gas in the test tube by holding the lighted match or splint near the mouth of the test tube. **CAUTION** If you are using a Bunsen burner and wooden splints, take appropriate safety precautions. Extinguish burning splints by immersing them in water. Be aware of the Bunsen burner flame. Make sure that long hair is tied back and loose clothing is confined.

10. *Immediately* after the reaction in the test tube occurs, use the medicine dropper to add about 1 mL of limewater to the test tube. Stopper the test tube and shake it. **CAUTION** The mouth of the test tube may be hot.

11. Record your observations of the gas when it was ignited. Record your observations of what happened when you added the limewater. Describe any residue left on the test tube.

12. Repeat steps 9 to 11 with the other test tubes in order: $\frac{1}{3}$, $\frac{1}{5}$, and then $\frac{1}{10}$ full.

13. Dispose of all chemical materials as instructed by your teacher.

Analysis

1. What happened to the phenolphthalein indicator during the production of the gas? Explain your observation. **Note**: phenolphthalein is an acid/base indicator.

2. What products may have formed during the combustion of the gas? Support your answers with experimental evidence. **Note**: Limewater reacts with carbon dioxide to produce a milky white solid.

Conclusions

3. **(a)** Write a balanced chemical equation for the incomplete combustion of acetylene gas.

   (b) Write a balanced chemical equation for the complete combustion of acetylene gas.

4. The air that we breathe is approximately 20% oxygen. Think about the reaction you just wrote for the complete combustion of acetylene. Which ratio in this investigation ($\frac{1}{2}$, $\frac{1}{3}$, $\frac{1}{5}$, or $\frac{1}{10}$) allowed the closest amount of oxygen needed for complete combustion? Support your answer with calculations. Do the observations you made support your answer? Explain.

Applications

5. An automobile engine requires a carburetor or fuel injector to mix the fuel with air. The fuel and air must be mixed in a particular ratio to achieve maximum efficiency in the combustion of the fuel. What might happen if the fuel and air mixture is too rich (if there is too much fuel)?

6. What does the limewater test indicate? Write the balanced chemical equation for the limewater test. **Hint**: Limewater is a dilute solution of calcium hydroxide. A carbonate forms.

In this section, you were introduced to the complete and incomplete combustion of hydrocarbons. You learned that the complete combustion of a hydrocarbon produces water and carbon dioxide. You also learned that the incomplete combustion of a hydrocarbon produces additional products, such as unburned carbon and dangerous carbon monoxide. In the investigation, you had the chance to make and combust a hydrocarbon.

In the next section, you will learn about an important factor of combustion reactions: energy. The combustion of hydrocarbons produces a large amount of energy. This is why they are so useful as fuels. How can you include energy as part of a combustion or other equation? How can you calculate the energy released by fossil fuels? You will learn the answers to these questions in the rest of this chapter.

Section Review

1. K/U What is the difference between incomplete and complete hydrocarbon combustion reactions?
2. K/U Explain why you would usually write (g) for the state of the product water in these combustion reactions. When might you identify the state of water as liquid?
3. **(a)** K/U Write the balanced equation for the complete combustion of heptane, C_7H_{16}.
   (b) K/U Write a balanced equation for the incomplete combustion of 1-pentene, C_5H_{10}.
4. K/U Natural gas is mainly methane gas. If you have a natural gas furnace, stove, or water heater in your home, you must ensure that these appliances are always running at peak efficiency. In other words, the methane gas should undergo complete combustion so that carbon monoxide is not produced. Write a balanced chemical equation to show the complete combustion of methane gas. Write a second balanced equation to show the incomplete combustion of methane gas.
5. C All hydrocarbons have the chemical property of combustion in the presence of oxygen. How do the complete combustion reactions of methane, ethane, and propane differ? How are they similar? **Hint:** Compare the balanced equations for the complete combustion of each.
6. I The complete combustion of ethane is given by the following unbalanced equation:

$$C_2H_{6(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$$

   (a) Balance this equation.
   (b) If one mole of ethane is combusted, how many grams of water vapour are produced?
   (c) Assume that air contains 20% oxygen gas. What volume of air at STP is needed for the complete combustion of one mole of ethane?