SCH4U-C

11

Exothermic and Endothermic Reactions

November 28, 2014

Introduction

In the first two lessons of this unit, you studied chemical reactions that involved the transfer of electrons. Galvanic cells use the energy associated with electron transfers for many types of batteries, such as those used to start cars, power flashlights, and run watches and personal sound equipment.

You are now going to learn about the heat energy that is released or absorbed during physical, nuclear, but primarily chemical, reactions. It is this heat energy that is used to heat and generate electricity for homes and industries, and to power automobiles, airplanes, and other forms of transportation.

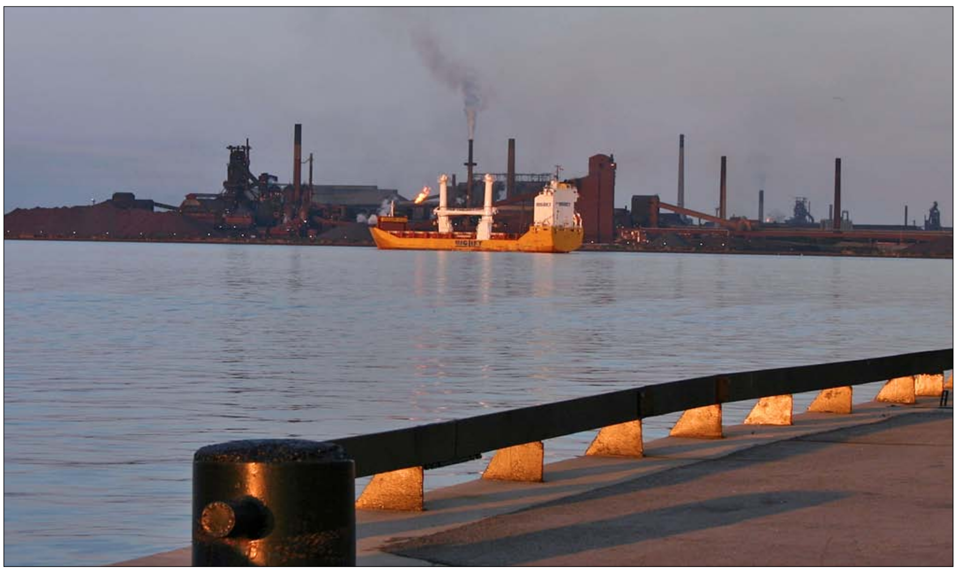

Figure 11.1: A lake freighter uses heat energy to deliver coal to Stelco Hamilton's plant, which uses coal combustion to smelt steel.

Source: http://commons.wikimedia.org/wiki/File:Yellow_Boat_-_Hamilton_Harbour_-_Hamilton_Flickr_Meet.jpgPlanning Your Study

You may find this time grid helpful in planning when and how you will work through this lesson.

Suggested Timing for This Lesson (Hours)	
Energy Sources	½
Referencing Your Information Sources	¼
Energy in Chemical Reactions	½
Quantity of Heat/GUESSS Problem-solving Format/Significant Digits	1
Calorimetry	½
Molar Enthalpy and Other Calculations	1
Geothermal Energy	¼
Key Questions	½

What You Will Learn

After completing this lesson, you will be able to

- describe examples of conventional and alternative energy sources and explain advantages and disadvantages of each
- explain the difference between exothermic and endothermic reactions
- calculate enthalpy changes for a variety of reactions using both experimental data and written problems

Energy Sources

There are three conventional energy sources used in the province of Ontario to produce heat and generate electricity: hydroelectric power, fossil fuels, and nuclear power. Hydroelectric power harnesses the energy of falling water. Niagara Falls accounts for the bulk of hydroelectricity, but there are also some smaller waterfalls and dams (which create falling water artificially) that also contribute to this source. Fossil fuels can be burned to release their energy. Most homes in urban areas are heated by the combustion of natural gas, but oil and propane are used in areas where natural gas is not readily available. Electricity may be generated by burning fossil fuels. Often, coal is used in these generating plants, but newer ones that use natural gas are being built. Nuclear power is used to generate the biggest share of Ontario's electricity.

These three conventional means of generating electricity are very similar. All three involve a turbine, a device that uses the motion of falling water or expanding steam to turn a generator. The generator spins a coil of wire within a strong magnetic field. The magnetic field induces the flow of electrons in the wire, creating electric current.

Hydroelectric Power

Figure 11.2: Sir Adam Beck Generating Complex near Niagara Falls, Ontario—the largest source of hydroelectric power in Ontario.

Source: http://commons.wikimedia.org/wiki/File:Adam_Beck_Complex.jpg

Natural or non-dammed hydroelectric facilities, like Niagara Falls, have few environmental problems. The construction of dams, on the other hand, results in the flooding of a large amount of land, which displaces communities and destroys or permanently alters wildlife habitats. The larger the project is, the greater the impact on the environment. The construction of dams along the St. Lawrence River, as part of the St. Lawrence Seaway, resulted in the relocation of a number of small villages. Some of the historical buildings from these villages were moved and preserved in Upper Canada Village, a tourist attraction near Cornwall, in Eastern Ontario. Dams also block the migration of fish to the spawning grounds upstream.

Populations of some game fish have been reduced, because of dams. One solution to the problem has been the building of fish ladders (a gently sloping series of tiny waterfalls that resemble a set of stairs) to help fish bypass the dam. These are only successful in small, low dams.

Fossil Fuels

Figure 11.3: The coal power-generating station near Thunder Bay, Ontario.

Source: http://commons.wikimedia.org/wiki/File:ThunderBayCoalGenStn.jpg

Fossil-fuel combustion heats water into steam or directly heats a home. Coal often contains impurities, which result in polluted air and acid rain. In Ontario, many coal-fired generators are being replaced by ones that use natural gas (methane), which burns more cleanly. However, both produce carbon dioxide, which is a major greenhouse gas linked to global warming and climate change. Newer generators and high-efficiency furnaces are much better for the environment than older generators and furnaces. They produce less carbon dioxide because they burn less fuel to provide the same amount of heat. Increasing the amount of insulation in the walls and ceiling of a home, as well as repairing its weatherstripping and replacing its windows, also reduces the amount of fuel burned to achieve the same level of heating. It helps the environment and, in the long run, saves money.

Nuclear Power

Nuclear power requires a nuclear, not a chemical, reaction. In chemical reactions, the relatively weak electron bonds within molecules are broken, releasing energy. In nuclear reactions, the strong forces binding protons and neutrons in the nucleus of the atom are broken to release energy. The resulting heat is used to turn water into steam, changing its state from liquid to gas. This exothermic reaction releases heat into the surrounding cool water, causing it to boil. Steam particles move more quickly and expand, because they have gained energy from the heat of the nuclear reaction. The forceful expansion of the steam turns a turbine to create electric current.

Nuclear power plants do not create any greenhouse gases. Canadian nuclear power plants use a very safe technology, with minimal risk of leaks and meltdowns. However, the water used to produce steam for the turbines is taken from nearby bodies of water. Once the steam has been condensed back into water and recycled, the now-warm water is piped into the lake. This water is not radioactive, but the higher temperature can disrupt habitats along the shores. As well, the solid nuclear waste produced at Canadian nuclear power plants remains radioactive for many years. Currently, it is being stored at various power facilities but, at present, there is no long-term storage solution for this waste.

Figure 11.4: The Bruce Nuclear Generating Station near Kincardine, Ontario.

Source: http://commons.wikimedia.org/wiki/File:Bruce-Nuclear-Szmurlo.jpg

The amount of energy released in a nuclear power plant is a result of the type of reaction taking place, which harnesses energy at the subatomic level. Physical changes release only a small amount of energy: from 1 to 100 kilojoules per mole (kJ/mol) of substance changed. This energy is known as latent heat. The latent heat of fusion is the energy required to melt one gram or one mole of a chemical. The latent heat of vaporization is the heat needed to turn one gram or one mole of a liquid into gas. This is not enough to be useful in heating homes or generating electricity. Chemical reactions, like combustion, release between 100 and 10 000 kJ/mol of substance reacted. Nuclear reactions release incredible amounts of energy, in comparison.

Their range is between 10^{10} and 10^{12} kJ/mol. A small amount of fuel yields a huge amount of energy.

Both nuclear and fossil fuels require processing and refining costs, but because far less nuclear fuel is needed to produce the same amount of energy as fossil fuels, its transportation costs are much lower. Ontario stores its high-level nuclear waste in water pools for a few years, and then it is placed in dry storage inside concrete canisters or structures. It can be stored in these canisters or structures for over 50 years. People have a tendency to overestimate the amount of nuclear waste lying around. All of the spent nuclear fuel that Canada has ever produced could fit inside six hockey rinks if the fuel pellets (which are each about the size of a small marshmallow) were stacked to the top.

Now watch three videos that will give you an overview of conventional energy sources used in Ontario:

- How It Works: Nuclear Power
- How It Works: Hydroelectric Power
- How It Works: Fossil-Fuel Power

Alternative Energy Sources

The conventional sources just discussed are not without their problems. Fuel supply can be an issue, in some instances, and there are always some environmental concerns that need to be addressed. As a result, a number of alternative energy sources are being researched, developed, and constructed throughout the province.

Figure 11.5: Wind turbines along Highway 10 near Shelburne, Ontario.

Source: http://commons.wikimedia.org/wiki/File:Ontario_windfarm_on_Hwy_10_-b.jpg

Ontario's wind turbines are being built either individually, as seen near the shores of Lake Ontario, or in large groups known as wind farms, such as the one north of Shelburne. Solar energy is used to heat hot-water tanks, passively heat some homes, or actively generate electricity through solar cells. Your calculator and some types of garden lights use this technology. Homes are also being heated and cooled using geothermal energy, the energy of the earth. In the past, a large amount of land was needed to set up a geothermal system, but technological developments have changed that. As a result, a number of subdivisions across Ontario are giving buyers the option to include geothermal heating and cooling in their new homes.

Alternative energy sources have some advantages over conventional sources. For instance, there is no fuel cost associated with wind or solar power, and no by-products that could pollute the air, soil, or water. Solar energy does not require a turbine or generator. The energy of the sun causes the flow of electrons in solar panels. The windmill itself is the turbine in wind power, turning the generator. There are a number of proposed sites for wind turbines and wind farms across the province indicating that, sooner or later, Ontarians will benefit from this affordable, clean energy source.

There are also some disadvantages associated with alternative energy. Solar panels are expensive and not extremely efficient, so they are often limited to small, specialized applications until they become cheaper and can produce more power. There are a number of concerns about wind turbines. Some people who live near them are bothered by the constant hum. Others feel that the large turbines are responsible for killing or disrupting bird populations. Research is currently underway to limit the negative impacts of wind turbines on people and animals.

Referencing Your Information Sources

On the whole, there are more advantages than disadvantages to alternative energy sources, as you will see when you look at many recently published books, articles, and websites. Better, safer, and more affordable alternative energy technologies are becoming available every day. As mentioned earlier, heating and cooling your home via geothermal energy is a technology that is making its way into the mainstream with few, if any, drawbacks. You will learn a lot more about this throughout the lesson and during your own research. As you are doing research for some of the Support and Key Questions in this course, keep in mind that many energy-related Internet sites are produced by groups with vested interests in specific forms of energy, and so the information on these sites can be biased.

In this course, when you are asked to provide references or a "bibliography" (the sources of your information), you may choose the style of citation, as long as you use the style consistently, and all of the key information is present. In your citation, you should include the name of the person or organization who wrote the material, the year/date of publication, and the full title. If you are referencing printed works (such as books, magazines, and newspapers), provide the page numbers used and the publisher's information. A popular format for providing this information is the style described in the Publication Manual of the American Psychological Association. Here are some examples of how this manual would format various sources of information at the end of a research assignment, in a list titled "references."

Format for a Book

Suzuki, D. (2006). *David Suzuki: The Autobiography* (28-29). Vancouver: Greystone Books.

Note: When citing books in this course, you are encouraged to mention specific page numbers (optional), as shown above, although this is not part of the APA style.

Format for a Newspaper

Schultz, S. (2005, December 28). Calls made to strengthen state energy policies. *The Country Today*, pp. 1A, 2A.

Format for a Magazine

Leclair, L., & Maclagan, M. (2006, January). Credit due: Ontarians reconnect with their ambitions. *TVOntario Magazine,* 12-13.

Video Clip from a Website

Norton, R. (2006, November 4). How to train a cat to operate a light switch (Video file). Video posted to http://www.youtube.com/watch?v=Vja83KLQXZs

Web page

Lynds, B. T. (1995, November 21). About temperature. Retrieved on December 2, 2009 from http://eo.ucar.edu/skymath/tmp2.html

Note: If there is a page like http://www.somewebsite.org/thepageyouwant.html, and thepageyouwant.html doesn't have the publishing information you're looking for, try shortening or "moving up" the URL to http://www.somewebsite.org/ to find out who authored the page and when it was last updated.

Interviews

Interview with Bill Symons, furnace repair technician, Barrie.

Support Questions

Be sure to try the Support Questions on your own before looking at the suggested answers provided.

25. Which changes of state release heat into their surroundings, and which changes of state absorb heat? Explain your answer. Suggested answer

26. This question is based on the three videos about nuclear, hydroelectric, and fossil-fuel power plants that you watched earlier in this lesson. Evaluate these three technologies in terms of their impact on the environment by assigning them a rank from 1 (best for the environment) to 3 (worst for the environment). Justify your ranking with evidence from the videos and your own research, including at least one non-video reference. Suggested answer

Energy in Chemical Reactions

Chemical reactions involve the change of reactants into products. During a chemical reaction, the creation of one or more products takes place, as well as a change in energy. This change in energy is known as enthalpy change and is symbolized by ΔH. The "Δ" symbol, delta, means "the change in," while the "H" is used to represent stored heat energy. Sometimes, the products created have more chemical stored energy than the reactants, and sometimes this happens the other way around. Where does the extra energy come from and where does it end up? And why do reactants and products have different amounts of energy?

The *system* of a chemical reaction consists of the actual chemicals, which are both the reactants and the products. The *surroundings* are the air and everything else around the system. Energy, in the form of heat, is easily exchanged between the system and its surroundings.

When a chemical reaction takes place, bonds must be broken in the reactant molecules. Sometimes, individual atoms or ions result, while at other times, groups of atoms or groups of ions (polyatomic ions) remain. Whatever the end result, the bond-breaking process requires energy, and the source of this energy is the surroundings. New bonds form when the atoms or ions rearrange themselves to form the products. The bond-forming process releases energy, which ends up in the surroundings.

A chemical reaction is a combination of a bond-breaking (energy absorption) and a bond-forming (energy release) process. The change in enthalpy is the net energy released or absorbed due to bonds breaking and then reforming into new chemicals.

An *exothermic* reaction is one where the energy of the reactants is greater than the energy of the products. The bond-breaking process absorbs less energy from the surroundings than the bond-making process releases. As a result, the system has a net loss of energy and the surroundings have a net gain. Energy or heat has **ex**ited the system, explaining the term **ex**othermic. During an exothermic reaction, the surroundings will increase in temperature.

An *endothermic* reaction is one where the energy of the products is greater than the energy of the reactants. The bond-breaking process absorbs more energy from the surroundings than the bond-making process releases. As a result, the system has a net gain of energy and the surroundings have a net loss and so cool down. Energy or heat has become stored within the system, explaining the term **en**dothermic.

A burning candle is an example of an exothermic reaction. When you hold your hand above the flame, your hand becomes part of the surroundings and is gaining heat from the system. When you bend a "glow stick" to make a safe light at a concert, you are mixing two chemicals together to create an exothermic reaction. This time, the energy is released mostly as light, with a little heat.

Figure 11.6: A ball made of a dozen glow sticks.

Source: http://commons.wikimedia.org/wiki/File:Glowsticks_ball.jpg

If you have been injured while playing sports, you may have encountered an endothermic reaction: the chemical ice packs that are used to treat injuries and reduce swelling contain two chemicals that react when mixed together. The cold that you feel is heat being absorbed by the system from its surroundings, in this case, your injured knee or ankle.

Enthalpy change or ΔH means energy changes in the chemical system, not in the surroundings. The value of ΔH is negative for exothermic reactions because of the decrease or loss in energy as reactants become products. The value of ΔH is positive for endothermic reactions because of the increase or gain in energy as reactants become products. This information can be communicated using three formats:

1. Thermochemical equation with the ΔH incorporated as a product
2. Thermochemical equation with ΔH as a separate term
3. Potential energy diagram

A thermochemical equation is balanced and shows reactants, products, and energy. If the equation does not show the energy term, then it is simply a chemical equation. When the energy term is included, then the equation is called a *thermo*chemical equation.

The synthesis of ammonia from nitrogen gas and hydrogen gas is exothermic. You can represent this reaction in the three formats mentioned above:

$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)} + 92\text{ kJ}$

or

$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)} \qquad \Delta H = -92\text{ kJ}$

The first thermochemical equation shows the 92 kJ as part of the balanced equation, as a product of the reaction. The second form of the thermochemical equation shows the enthalpy change to the right of the equation. The negative sign in the ΔH notation reminds us that heat was lost by the system and released into the surroundings.

or

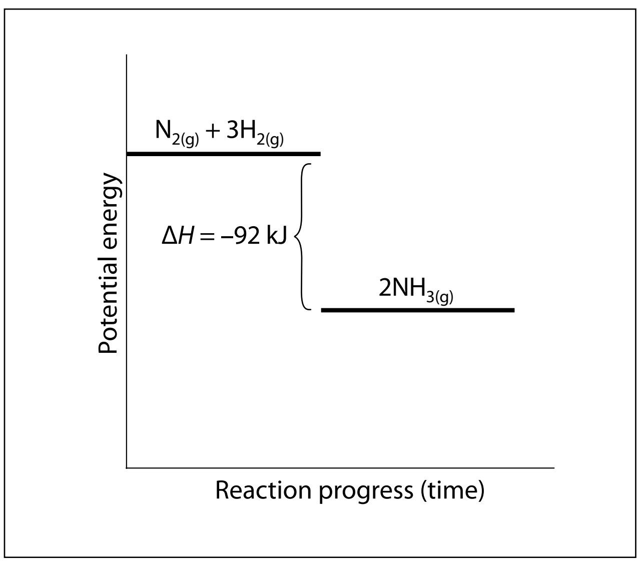

Figure 11.7: Potential energy diagram for an exothermic reaction

A potential energy diagram shows the energy of reactants and products graphically. The difference between the two energies is ΔH.

The synthesis of hydrogen iodide from hydrogen gas and iodine vapour is endothermic. You can represent this reaction as:

$H_{2(g)} + I_{2(g)} + 270\ kJ \rightarrow 2HI_{(g)}$

or

$H_{2(g)} + I_{2(g)} \rightarrow 2HI_{(g)} \qquad \Delta H = +270\ kJ$

or

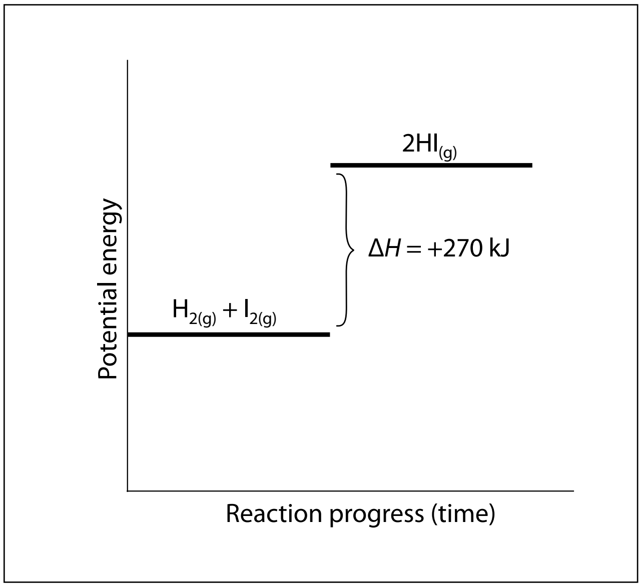

Figure 11.8: Potential energy diagram for an endothermic reaction

Figure 11.8 is a simple potential energy diagram (an enthalpy diagram) showing how the enthalpy (stored energy of the chemical system) has increased by 270 kJ. Here, the chemicals have absorbed that energy as they became the products, HI. This represents an endothermic reaction.

Support Questions

27. Fully explain the difference between exothermic and endothermic reactions, in terms of the system and the surroundings.

Suggested answer

28. Communicate or show enthalpy changes in various formats (graphs and equations), and in three different ways, for each of the following reactions:

a) One mole of propane undergoes complete combustion, releasing 2220 kJ of energy.

b) Two moles of ammonium chloride react with barium hydroxide, producing water, ammonia gas, and barium chloride in solution, absorbing 50 kJ of energy from the surroundings.

Suggested answer

Quantity of Heat

In this part of the lesson, you will learn how to calculate the quantity of heat lost or gained during reactions. The mathematical equation that is used in these calculations is:

$Q = mc\Delta T$

The quantity of heat energy, Q, will always be expressed as a positive quantity. The symbols in the equation have the following meanings and units associated with them.

Table 11.1: Symbols used in solving specific heat-capacity problems

Symbol	Meaning	Unit
Q	Quantity of heat lost or gained	Measured in joules (J) or kilojoules (kJ), the metric unit of energy
m	Mass of substance that is gaining or losing heat	Measured in grams (g); 1 kg = 1000 g
c	Specific heat capacity of substance; each substance has its own unique value	J/kg•°C
ΔT	Change in temperature: if a substance loses heat its temperature decreases, if it gains heat its temperature increases; $\Delta T = T_f - T_i$ (T_f is the final temperature and T_i is the initial temperature)	°C

A high value of specific heat capacity, as seen for water at 4.18, means that it takes a relatively large quantity of heat (4.18 J) to raise the temperature of 1 gram of water by only 1°C (from 25° to 26°, or 71° to 72° for example). Examples of specific heat capacity values are provided in the following table. You don't have to memorize them, but you may have to refer to this table when you answer Support Questions and Key Questions.

Table 11.2 indicates that the heat capacity of water is approximately 10 times that of iron. For the same masses, water would rise in temp by 1°, while iron would rise by 10° if they absorbed the same amount of heat.

Table 11.2: Specific heat capacities

Substance	J/g°C
Water/Solutions	4.18
Aluminum	0.900
Copper	0.385
Iron	0.450
Lead	0.130
Methanol	2.92
Silver	0.240
Zinc	0.390

Notice that the units of heat capacity are joules per gram degrees Celsius. This means that heat capacity is a measure of how many joules of energy are needed to heat one gram of the substance by one degree Celsius.

How do the specific heat capacity values compare for solids and liquids? The table shows that solids have lower specific heat capacities than liquids. Although this is a limited list, a more complete list would show exactly the same pattern. The pot on the stove that is used to boil water will heat up more quickly than the water inside, and will cool down more quickly when the heat is turned off. That is why when the sun shines on sand and water for the same amount of time, the sand becomes very hot, but the water remains cool. Communities who live near large bodies of water are cooler in the summer, but warmer in the winter, because of the high specific heat capacity of water. The density of water also explains many of its unique properties.

The volume of water is often measured in laboratories using a marked glass measuring container called a graduated cylinder. The density of water is 1.00 grams per 1.00 millilitres (1.00 g/mL). The fact that 1.00 mL of water has a mass of 1.00 g provides you with an easy way of measuring the mass of water in a container without having to subtract the weight of the container. For instance, 100 mL of water has a mass of 100 g. This is much faster than using a balance, but not as accurate. This method only works for water and "dilute solutions," which contain mostly water. Why? It's because other liquids have different densities.

The specific heat capacity for water is used as a very close estimate of the specific capacity value for dilute solutions, which are mostly water.

GUESSS Problem-solving Format

This part of the lesson includes the first problem solving that you have done in this course. In previous courses, you might have been taught a problem-solving process or format known by a number of acronyms: GRASP, GUESS, GRASS, or GUESSS.

A problem-solving format is very useful, because it gives you a strategy to follow when solving problems, but it also reveals your thinking process to anyone looking at your solution. The model solutions given in these lessons will show you how to organize your own solutions. You should practise the proper format when answering the Support Questions. The proper format will also be expected in the Key Questions and in the Final Test and if you use it, you can get some partial credit for your answers, even if your final statement is wrong. In SCH4U-C, the proper format is the GUESSS process, which includes the following steps:

Given: Read through the problem and list any measurements, along with their units. If you are not sure how to approach the problem, seeing this list might help. Drawing a diagram and labelling what is known can be very helpful in understanding what is happening in a problem. You use your visual learning with this technique.

Unknown: What are you trying to find? Sometimes, you will need to find more than one missing value to get your final solution.

Equation: Write down any formulas (equations) that should apply. For example if, in reading the problem, you recognized that heat was being absorbed or released, then the formula for Q would come to mind. This is where the list of "givens" helps. You might have to rearrange the equation.

Substitute: Substitute the given value **with units** into the equation.

Solve: Solve the equation and record the answer using the proper units and significant figures.

Statement: Write a sentence that gives meaning to the final numerical answer.

Significant Digits

Significant digits, also called significant figures, are necessary because not all measurements that are made in the laboratory are equally precise. If you are using these measurements in a calculation, you don't want your results to seem more precise than they really are. If the electronic balance you are using measures to the nearest 1 g, then an object with a 10 g must be recorded as 10 g. But if the balance measures to the nearest one hundredth of a gram (that is, 0.01 g), then the mass is recorded as 10.00 g, to show the increased precision. Use the following guidelines to help you to determine the number of significant figures in a measurement.

Guidelines for determining significant digits:

1. Every "non-zero" digit in a recorded measurement is significant. For example, 24.6 m, 0.742 m, and 724 m all have three significant figures.
2. "Sandwiched" zeros are significant. For example, the measurements 7003 m, 40.79 m, and 1.503 m all have four significant figures.
3. Zeros to the right of the decimal point, before all "non-zeros," are not significant. For example, 0.0000099 only has two significant figures.
4. "Un-sandwiched" zeros to the right of a "non-zero" are not significant. For example, 101 000 has only three significant figures. One of the zeros is sandwiched (see guideline 2).
5. Zeros are significant if they are at the end of the number and a decimal point is present. In these cases, the zeros at the end of the number and the zeros to the right of the decimal are all significant. For example, the measurements 1241.20 m, 210.100 m, and 5600.00 m all have six significant digits.

Mathematical Operations with Significant Digits

Correctly using significant digits becomes crucial when you are dealing with more than one measured (observed) number during experiments. For example, imagine a child's swimming pool that holds about 200 L of water. If you add another 0.471 L, can you scientifically say that it now holds precisely 200.471 L? No, you cannot, because the first number was not measured as precisely as the second number. Scientists have developed the following rules for measured quantities and their mathematical operations, such as addition, subtraction, multiplication, and division.

Addition and subtraction: Calculations involving addition and/or subtraction of measured quantities must have no more significant digits to the right of the decimal point than the fewest possessed by any measured quantity to the right of the decimal point.

For example: 12.21 m + 324.0 m + 6.25 m = 342.46 m, but the answer must be rounded to 342.5 m. The weakest number in this list, in terms of decimal places, has only 1 decimal, so that limits the answer to having only one decimal place. The calculator answer is rounded to one decimal then. As always, make sure that you include the correct SI units, such as metres (m), litres (L), and grams (g).

Multiplication and division: Calculations involving only multiplication and/or division of measured quantities must have the same number of significant digits as the fewest possessed by any measured quantity in the calculation.

Multiplication examples:

12.21 m × 44.00 m = 537.24 m^2. The answer must be rounded to four significant figures, which results in 537.2 m^2.

12.21 m × 44 m = 537.24 m^2. Of all the numbers being multiplied, the weakest (in terms of significant figures) has only two significant figures (44). This requires that the answer have two significant figures: 540 m^2. (**Note:** The final zero in 540 is not a significant digit. See guideline 4 on page 16.)

If any number used in the calculation was a counted number, such as one of the whole number coefficients in a balanced equation, that number is ignored for purposes of significant figures. Counted numbers are considered to have no uncertainty in them.

Division example:

0.112 m ÷ 0.041 m = 2.7317073170 m, but the answer must be rounded to two significant figures, which results in 2.7. (**Note:** The zero to the right of the decimal in 0.041 is not a significant digit. See guideline 3 on page 16)

Example

During an experiment, 125 mL of water is heated from 12.5°C to 37.8°C. What quantity of heat did the water gain?

Solution

Use the GUESSS process and the correct number of significant digits.

Given:

m = 125 g (125 mL = 125 g for water)

ΔT = 25.3°C (37.8°C – 12.5°C)

c = 4.18 J/g°C (from Table 11.2)

Unknown:

Q = ?

Equation:

$Q = mc\Delta T$

Substitute:

= 125 g × 4.18 J/g°C × 25.3°C (Note that the units, g and °C, cancel out, leaving the heat unit, J.)

Solve:

= 13 219.25 J (The three numbers used in the calculation all have three significant figures, so this answer is rounded at the third figure, the first 2.)

= 1.32×10^4 J or 13.2 kJ (Writing this answer as 13 200 J would also be three significant figures.)

Statement:

When 125 mL of water warms by 25.3°C, the water gains 13 200 J of heat energy.

The answer could be expressed in scientific notation or converted to kJ since the quantity of heat was fairly large in value.

Example

What mass of aluminum loses 16 kJ of heat when cooled from 59.7°C to 11.8°C?

Solution

Use the GUESSS process and the correct number of significant digits.

Given:

Q = 16 kJ = 16 000 J (must be in J to match the unit of c, **J**/g°C)

c = 0.900 J/g°C (from Table 11.2)

ΔT = 47.9°C

Unknown:

m = ?

Equation:

$m = Q/c\Delta T$

Substitute:

= 16 000 J / (0.900 J/g°C × 47.9°C)

Solve:

= 371.1435862 g (round to two significant figures due to the number 16 000)

= 3.7×10^2 g or 370 g

Statement:

When 370 g of aluminum drops in temperature by 47.8°C, it loses 16 kJ of heat energy.

Support Questions

29. When solving the following, use the GUESSS problem-solving format and the proper number of significant figures in your answer. Get into the habit of using this procedure. Be careful to always include the proper units.

a) A kettle is used to heat 1.50 L of water from 15.4°C to 92.3°C. How much energy is required? (**Hint:** See example 1.) Suggested answer

b) What mass of copper is heated from 25.4°C to 57.9°C, using 8.00 kJ of heat? (**Hint:** See example 2.) Suggested answer

c) During an experiment, 683 J of energy is used to heat a 125 g piece of metal from 42.3°C to 54.4°C. Identify the metal. (**Hint:** The value of part c) will identify the metal.) Suggested answer

d) During another experiment, 58.0 g of methanol is heated, using 8410 J of energy. If the initial temperature was 17.2°C, what is the final temperature? (**Hint:** Find the change in temperature first.) Suggested answer

e) How much heat energy is used to heat an aluminum container with a mass of 48.5 g, filled with 150 g of water, from 25.0°C to 55.5°C? Assume that the aluminum container and water have the same initial and final temperatures. (**Hint:** Find Q for the water and Q for the aluminum container separately.) Suggested answer

Calorimetry

The process of determining heat lost or gained by experiment is called "calorimetry." This word derives from two words, "calorie" and "meter" (to measure). The calorie is a non-metric unit of heat. So calorimetry is the process of measuring heat.

A calorimeter is a device that allows someone to experimentally make measurements as some endothermic or exothermic process occurs.

Calorimetry Activities

You will now have the opportunity to see a calorimetry experiment in an online activity called Heat of Combustion. After simulating the first calorimetry activity online, you will be given the data for the final three activities.

First Activity

The first experiment is online and involves a candle. For the final three experiments that are not online, you will be given the data for three alcohol burners that burn methanol, ethanol, and propanol fuel.

Materials:

Here are the materials used in the first experiment, which is online:

- 50 mL of H_2O
- Retort stand or similar vertical support
- Steel (iron) tomato-paste can, with lid and label removed (serves as our calorimeter)
- Large tomato can, with bottom, lid, and label removed (to enclose the candle)
- Ring clamp or similar clamp to hold can
- Balance or scale
- Candle
- Matches
- Graduated cylinder
- Thermometer

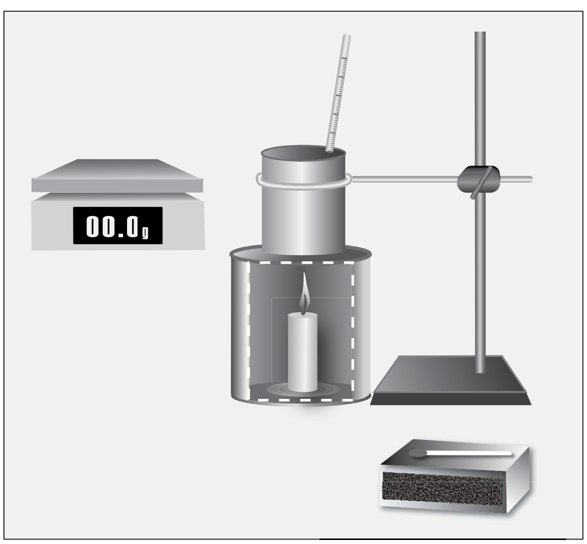

Figure 11.9: Laboratory set-up for the calorimetry activity

Procedure:

Step 1: Read through the list of eight preparation items needed for setting up the lab. The preparation for the experiment should all be done before you begin. (This includes affixing a candle to the lid of a can and measuring the mass of the candle and lid.)

Step 2: Light the candle, and immediately lower the clamp so that the tomato-paste can of water is 1 cm above the candle flame.

Step 3: Allow the candle to burn for approximately five minutes. Blow out the candle, and measure the final temperature of the water.

Step 4: Measure the mass of the candle and lid again, to get their final mass.

The steps for this procedure are listed above for your convenience, in case you cannot read them online or would like to attempt your own calorimetry trials. You are not required to do this or another activity requiring hard-to-find materials, such as a retort stand.

Observations and Calculations

The calculations can be found online. Reviewing them online will help you to understand the final three calorimetry activities. The observations for the first activity are provided in Table 11.3:

Table 11.3: Data table for web-based calorimetry activity

Material used in activity	Measurement
Mass of tomato-paste can (g)	30.7
Mass of water and can (g)	93.6
Initial temperature of water and can (°C)	19.0
Final temperature of water and can (°C)	35.5
Initial mass of candle and lid (g)	23.8
Final mass of candle and lid (g)	22.9

Second, Third, and Fourth Calorimetry Activities

The second, third, and fourth activities are not shown online. Using the same materials, the same procedure was performed three times, each time using a different fuel: methanol, ethanol, and propanol. Each of these fuels was contained in a burner, and the mass of fuel used can be found by subtracting the mass of fuel and burner after the experiment, from the mass of the fuel and burner before the experiment.

All of the measurements and data have been recorded and provided in a data table (Table 11.4) for you. You do not need to measure and record the mass of the iron calorimeter can, the initial and final temperatures of water in the can, or the initial and final masses of the alcohol burner.

These activities are almost identical to what was done with the candle. In the candle activity, the fuel was paraffin wax. In these three activities, the fuels will be three different alcohols.

Observations

Table 11.4: Suggested results table for calorimetry activities

Measurement	Methanol (CH_3OH)	Ethanol (C_2H_5OH)	Propanol (C_3H_7OH)
Mass of iron can (g)	74.60	74.60	74.60
Mass of water and can (g)	124.60	124.60	124.60
Mass of water (g)	50.0	50.0	50.0
Initial temperature of water and can (°C)	20.0	20.0	20.0
Final temperature of water and can (°C)	61.1	71.3	77.5
Initial mass of burner (g)	35.53	37.21	36.82
Final mass of burner (g)	34.92	36.65	36.29

The above table shows the measurements you would have been able to make during the experiments. These are your quantitative observations. An observation table does not contain any calculated values. Use this data to answer the following Support Questions.

Calculating Percentage Error

In this course you will need to calculate the percentage of experimental error. It is possible to do this whenever you have observed “experimental results” that you can compare to the expected or “theoretical results.” This is done using the percentage error equation:

$$\%\ \text{error} = \left(\frac{\text{theoretical value} - \text{experimental value}}{\text{theoretical value}}\right) \times 100\%$$

Be aware that the modulus symbols “ | | ” are used to enclose an absolute value. Any negative value inside these symbols becomes positive, while positive values remain unchanged. In other words, | x | is the absolute value of x, which is the numerical value of x without regard to its sign.

Example

Calculate the percentage error when you have observed 4.1 g of NaCl produced during a reaction when, theoretically, you expected the reaction to produce 4 g of NaCl.

Solution

Use the percentage error equation and substitute the known value into it.

$$\begin{aligned}\%\ \text{error} &= \left(\frac{[4.00 - 4.10]}{4.00}\right) \times 100\% \\ &= \frac{0.10}{4.10} \times 100\% \\ &= 2.5\%\end{aligned}$$

Therefore, the percentage error is 2.5%.

Support Questions

30. Use the data in Table 11.4 to answer the following questions about the combustion of methanol, ethanol, and propanol. Review the rules for significant figures for the operations of addition, subtraction, multiplication, and division. For each different fuel used, what is the

a) ΔT?

b) heat gained by the water? (Calculate the Q value for the water.)

c) heat gained by the iron calorimeter can? (Calculate the Q value for the iron can.)

d) total heat gained?

e) total heat released by the combustion of alcohol? (**Hint:** The heat gained by the water plus the iron can came from the alcohol combustion.)

f) mass of fuel burned? (We will also need to know how many moles of alcohol were burned.)

g) quantity of heat released per gram of fuel burned?

h) quantity of heat released per mole of fuel burned?

i) percentage error (theoretical values will be provided for calculating error)?

j) Write the thermochemical equation for the combustion of one mole of the fuel. Use the experimental heat term calculated.

Suggested answer

k) You can set up a table similar to the following one to show the results of these calculations, and you should include at least one complete set of calculations, for example, the calculation of the methanol trial. This is known as a demonstration or "sample" calculation.

In other words, if you were handing this question in for credit on a Key Question or test, you would need to demonstrate how you calculated your answers for parts a) through i) for methanol, as a demonstration, but you would not be required to hand in all subsequent calculations for ethanol and propanol. This is standard practice in most journals that publish scientific studies because it saves a great deal of time and many pages of calculations. You should do all three sets of calculations here, however, to practise the skill.

Suggested table for calorimetry calculations

	Methanol (CH_3OH)	Ethanol (C_2H_5OH)	Propanol (C_3H_7OH)
a) ΔT (°C)			
b) Heat gained by the water (kJ)			
c) Heat gained by the iron calorimeter can (kJ)			
d) Total heat gained (kJ)			
e) Total heat released by the combustion of alcohol (kJ)			
f) Mass of fuel burned (g) and (mol)			
g) Quantity of heat released per gram of fuel burned (kJ/g)			
h) Quantity of heat released per mole of fuel burned (kJ/mol)			
Theoretical heat released per mole (kJ/mol)	727	1367	2020
i) Percentage error (%)			

Suggested answer

31. **a)** In each case, your calculated heat released per mole was less than the theoretical value. Using the illustration of the set-up (Figure 11.9), describe three possible sources of error.

b) Suggest two methods of minimizing the errors described above. Suggested answer

32. List four safety rules that apply to this activity. Suggested answer

33. Which of these fuels would be the best to use in remote locations, such as wilderness camping, where everything you take with you has to be carried? Explain. Suggested answer

Molar Enthalpy and Other Calculations

In order to make valid comparisons among energy released or absorbed during chemical reactions, there must be some kind of standard established. The measurements for these reactions should occur under standard conditions. Standard conditions are set at 100 kPa pressure and 25.0°C (298 K) temperature. If the reaction occurs under these standard conditions, then the symbol $\Delta H°$ is used. The symbol ΔH without the "°" indicates that the reaction took place at conditions other than standard.

As well, different masses of reactants are used in reactions, releasing or absorbing different quantities of energy. Again, to make valid comparisons between reactions, the amount of energy released or absorbed is calculated in terms of moles of the key reactant. In combustion reactions, the key reactant is the fuel. That is why you converted the mass of fuel burned to moles, in the previous set of Support Questions, and calculated the molar enthalpy for each reaction. The units of molar enthalpy are kJ/mol.

Sometimes, the molar mass of a substance is not known, or, as in the camping question you answered in Support Question 34, there may be some reason to want to compare masses of substance instead of moles. Therefore, sometimes you may be asked to calculate kJ/g instead of kJ/mol.

The calorimeter that you used in the simulation activity was designed for use with a heat source. You discovered that it was not that accurate because so much of the heat released by the reaction ended up escaping into the surrounding air, and could not be measured and accounted for. To get more accurate measurements, such as the theoretical values that you used to calculate percentage error, a different type of calorimeter is used. This is called a "bomb calorimeter." In this calorimeter, the reaction takes place inside a metal reaction chamber that is immersed in an insulated container of water. Heat from the reaction is transferred to the metal chamber and to the water, but not beyond. Heat does not escape into the air.

Figure 11.10: A bomb calorimeter with bomb vessel exposed.

Source: http://commons.wikimedia.org/wiki/File:Bombenkalorimeter_mit_bombe.jpg

Example

During an experiment, 1.00 g of octane (C_8H_{18}) undergoes combustion in a lead bomb calorimeter. The mass of the lead chamber is 52.82 g and it is immersed in 1.00 L of water. What is the molar enthalpy of the combustion of octane if the temperature of the water increases by 11.4°C?

List what is known:

Mass of fuel burned (octane) = 1.00 g

Mass of lead calorimeter = 52.82 g

Mass of water in calorimeter = 1.00 L × 1.00 kg/L = 1.00 kg

= 1.00×10^3 g (notice that 3 significant figures are maintained)

ΔT for both the water and the lead container = 11.4°C

(It is logical to assume that the water was placed into the calorimeter and they came to the same initial temperature and then they both warmed to the same final temperature.)

Solution

First we determine the heat absorbed by the water in the calorimeter.

$Q_{water} = mc\Delta T$

$= 1.0 \times 10^3 \text{ g} \times 4.18 \text{ J/g°C} \times 11.4\text{°C}$

= 47652 J (round to 3 "sig-figs")

= 47.7 kJ

Now we calculate the portion of the heat that was absorbed by the lead container.

$Q_{lead} = mc\Delta T$

$= 52.82 \text{ g} \times 0.130 \text{ J/g°C} \times 11.4\text{°C}$ (the *c* value for lead was found on Table 11.2)

= 78.2794 J (round to 3 sig-figs)

= 0.0783 kJ

Q_{total} = 47.7 kJ + 0.0783 kJ = 47.8 kJ (Rules for adding tell us that 1 decimal is permitted in the answer.)

We now know that 47.8 kJ of energy was absorbed in total by the water and the lead calorimeter. This means that 47.8 kJ of heat was absorbed when 1.00 g of octane was burned.

Finally, we need to find the molar enthalpy by dividing as follows.

$$\frac{47.8\text{kJ}}{8.75 \times 10^{-3}\text{mol}} = 5460 \text{ kJ/mol}$$

Therefore, the molar enthalpy of the combustion of octane is 5460 kJ/mol.

Octane is C_8H_{18}. Molar mass is calculated, using the periodic table, as follows:

$M_{octane} = (8 \times 12.01) + (18 \times 1.01) = 114.26$ g (Note that really we are adding together all of those numbers that have 2 decimals in each one, so we keep two decimals in the answer.)

We use that molar mass to convert 1.00 g of octane to moles.

$n = m / M = 1.00 \text{ g} / 114.26 \text{ g/mol} = 0.00875 \text{ mol}$

We have determined that 1.00 g of octane releases 47.8 kJ of energy when it burns. We then determined that if 1 mole of octane were burned, it would release 5460 kJ of heat.

This was an exothermic reaction, and we learned earlier that ΔH for an exothermic reaction is expressed as a negative quantity. Let's now take what we know about octane and write a thermochemical equation for the combustion of octane.

The balanced equation, assuming complete combustion, is:

$C_8H_{18(l)} + 12.5O_{2(g)} \rightarrow 8CO_{2(g)} + 9H_2O_{(g)}$

The equation was written to involve one mole of octane because we determined the heat released when one mole of octane burns. This was determined to be 5460 kJ.

We can now write the thermochemical equation:

$C_8H_{18(l)} + 12.5O_{2(g)} \rightarrow 8CO_{2(g)} + 9H_2O_{(g)} + 5460 \text{ kJ}$

or

$C_8H_{18(l)} + 12.5O_{2(g)} \rightarrow 8CO_{2(g)} + 9H_2O_{(g)} \quad \Delta H = -5460 \text{ kJ}$

Notice that the negative sign appears when we use the ΔH notation.

A "coffee-cup" calorimeter is often used when studying enthalpy changes for reactions that take place in water. These reactions include neutralization of acids and bases, and the enthalpy change that accompanies the dissolving of a solute. Often, a Styrofoam cup is used as the reaction vessel, hence the name "coffee cup." The water gains or loses heat, but the insulated Styrofoam walls of the cup do not gain or lose heat. Heat can still escape into, or be absorbed from, the air if the cup has no lid.

Example

In an experiment, 100. mL of water is measured and poured into a coffee-cup calorimeter. Then, 3.00 g of potassium nitrate (KNO_3) is dissolved in the water, causing the temperature to drop from 20.0°C to 17.5°C. What is the molar enthalpy for dissolving potassium nitrate?

Solution

First list what we know about what occurred here:

Mass of KNO_3 that was dissolved in the water = 3.00 g

Mass of water in the calorimeter cup = 100.mL × 1.00 g/mL = 100.g (note the decimal signifying that the number 100.g has three significant figures)

$\Delta T = 2.5°C$ (A drop in temperature occurred to the water. This was an endothermic reaction that occurred with the potassium nitrate dissolving. The potassium nitrate absorbed heat from the water as it dissolved.)

We assume that the insulated cup was not involved. First determine the heat lost by the water.

Heat lost, $Q = m\, c\, \Delta T$

$= 100.\text{g} \times 4.18\ \text{J/g°C} \times 2.5°\text{C}$

$= 1.0 \times 10^3\ \text{J}$ (we are only allowed to keep 2 significant figures)

$= 1.0\ \text{kJ}$

We now know that when 3.00 g of KNO_3 dissolves in water, 1.0 kJ of heat is absorbed by the KNO_3. We would like to know how much heat would be absorbed if 1 mole of KNO_3 were dissolved in water. Change 3.00 g to moles.

$M_{KNO3} = 101.11$ g/mol

$n_{KNO3} = 3.00\ \text{g}/(101.11\ \text{g/mol}) = 0.0297$ mol

Heat absorbed per mole of KNO_3 dissolving is 1.0 kJ/0.0297 mol = 34 kJ/mol

As was done for the octane combustion example, the thermochemical equation can be written for this dissolving process. The equation for the dissolving of a solid in water is simple:

$KNO_{3(s)} \rightarrow KNO_{3(aq)}$ Notice that this shows a solid becoming aqueous and it is for 1 mole of KNO_3.

The thermochemical equation includes the heat term determined earlier for one mole.

$KNO_{3(s)} + 34\ \text{kJ} \rightarrow KNO_{3(aq)}$

or

$KNO_{3(s)} \rightarrow KNO_{3(aq)} \quad \Delta H = +34\ \text{kJ}$

Note (again) that the sign is utilized when the ΔH notation is used.

Support Questions

When solving the following problems, pay attention to the problem-solving format and significant figures in your answers.

34. A 4.806 g sample of salad oil is burned in an iron calorimeter having a mass of 30.30 g. The calorimeter is immersed in 2.00 L of water. The initial temperature of the set-up was 58.2°C and the final temperature was 68.3°C. Calculate the enthalpy of the combustion of salad oil in kJ/g. Suggested answer

35. Support Question 30 showed the theoretical heat released per mole of each of the three alcohols burned in the activity. For ethanol and propanol, write the thermochemical equation for the complete combustion of the alcohol, two ways. Use the theoretical heat term in the equation. Suggested answer

36. During an experiment, 10.0 g of sodium hydroxide is dissolved in 200.mL of water, raising its temperature from 20.5°C to 33.8°C. Calculate the molar enthalpy for the dissolving of NaOH in water. Write the thermochemical equation that describes this process. Suggested answer

Geothermal Energy

At the start of this lesson, you learned about some conventional and alternative forms of producing heat energy and electrical energy. One of the alternative methods mentioned was geothermal energy, which is an up-and-coming technology that is being used to heat and cool homes for less money and with less impact on the environment.

Geothermal energy systems make use of the ground's natural ability to store and release heat. Two metres below the ground's surface, the temperature is a constant 17°C all year round. A series of pipes is buried at this depth and filled with water. In the past, only homes in rural areas with large lots could use geothermal energy because the pipes ran horizontally. Since then, technology has advanced so that the pipes can now be placed vertically, going much deeper and requiring less space. A number of new subdivisions are being constructed in Ontario where buyers have the option of installing a geothermal heating and cooling system. The water in the pipes and in the ground that is in contact with the pipes serves as a heat source in the winter and a heat sink in the summer. In other words, heat is pumped into the ground in summer and out of the ground in winter.

In winter, the water passes through a heat pump, allowing heat from the water to be transferred into the home. In summer, the process is reversed: the heat pump cools the home by taking heat from the house and transferring it into the external water.

A geothermal system is very expensive to install: it costs approximately $20 000, while a standard furnace and air conditioner only costs about $3500. Without a geothermal system, the cost of natural gas to heat the home and operate the hot-water tank is around $150 per month. Operating the heat pump would cost about $100 per month but, without having to pay for electricity for an air conditioner, would save $400 per year. A geothermal system costs more initially to install, but is cheaper to operate. Maintenance costs are approximately the same for both conventional and geothermal systems.

The federal government provides grants of $7500 to encourage people to switch to geothermal energy. A geothermal system can be added to a home without altering any of the existing ductwork. People who install geothermal systems in their homes often use other strategies to reduce the costs of heating their homes. Some of these strategies include: designing the home so that it receives passive solar heat in the winter (for example, by putting large windows on the south side to let the heat in), using better wall insulation and better-insulated windows, using curtains or window quilts in the winter, and using dark stone floors in front of windows that receive sunlight. There are many other strategies that help to keep homes warmer and more ecologically friendly, for very little money.

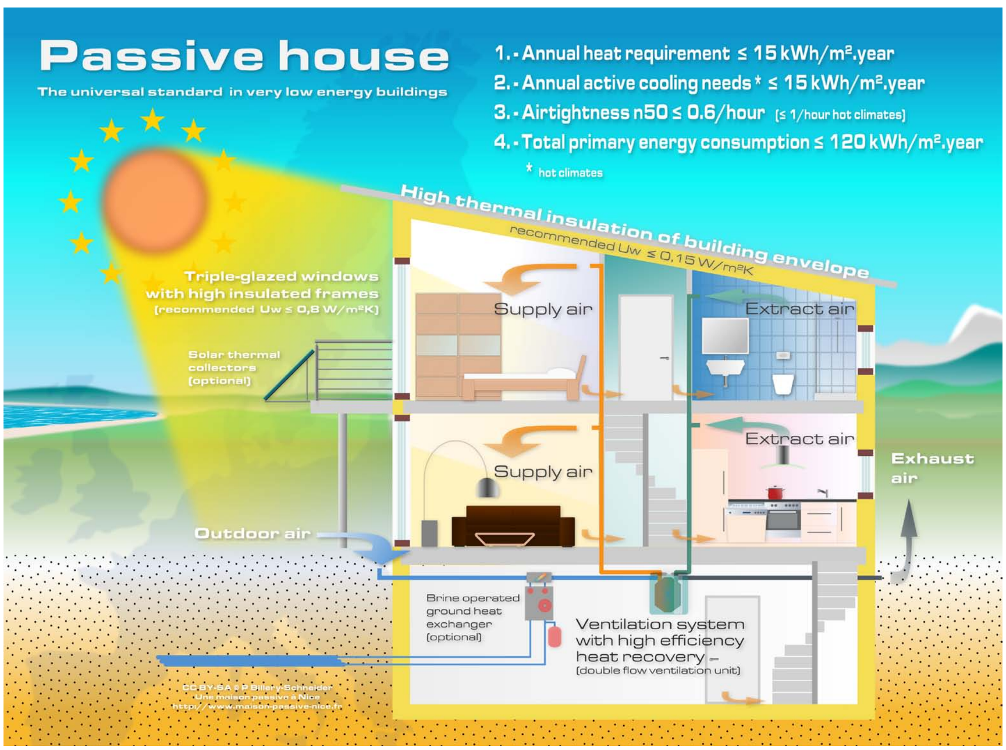

Figure 11.11: A passive house that uses the low angle of the sun in winter for heating and has a saltwater-based ground heat exchanger for geothermal heating and cooling.

Source: http://commons.wikimedia.org/wiki/File:Passive-house_scheme_HQ.png

Key Questions

Now work on your Key Questions in the online submission tool. You may continue to work at this task over several sessions, but be sure to save your work each time. When you have answered all the unit's Key Questions, submit your work to the ILC.

42. A geothermal system is initially more expensive to set up than a conventional heating and cooling system, but the annual operating costs are much cheaper.

a) How many years would it take for the geothermal system to pay for itself? Assume that you are looking to install a new system, and that you can obtain a government grant to do so. Use the costs that were mentioned on the previous page. Show all of your calculations. **(5 marks)**

Installation Costs

- geothermal system: $20 000
- conventional system: $3500
- government grant for geothermal installation: $7000

Operating Costs

- geothermal: $100/month
- conventional: $150/month, plus $400/yr for air conditioning

b) From an environmental perspective, describe two other advantages of switching to a geothermal system. **(2 marks)**

43. A 2.56 g sample of anthracene, $C_{14}H_{10}$, was burned to heat an aluminum calorimeter (mass = 948 g). The calorimeter contained 1.50 L of water with an initial temperature of 20.5°C and a final temperature of 34.3°C.

a) Calculate the molar heat of combustion of anthracene. **(6 marks)**

b) Write the thermochemical equation, two ways, for the complete combustion of anthracene. **(2 marks)**

c) If the actual value for $\Delta H = -7150$ kJ/mol, what is the percentage error? **(1 mark)**

44. Write thermochemical equations, two ways, for each of the following situations. **(4 marks total)**

a) When 1 mole of NH_4Cl is dissolved in water, 25.0 kJ is lost by the water.

b) When acetic acid is neutralized by sodium hydroxide, 76 kJ (per mole of acetic acid) of heat is released.

Now go on to Lesson 12. Send your answers to the Key Questions to the ILC when you have completed Unit 3 (Lessons 9 to 12).