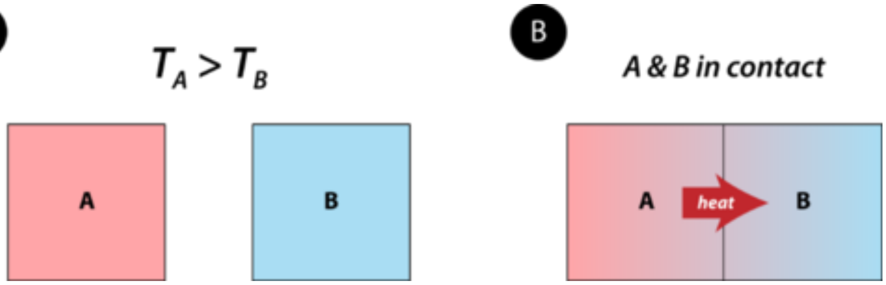
Heat Flow

Heat and Work

Energy is the capacity for doing work or supplying heat. When you fill your car with gasoline, you are providing it with potential energy. **Chemical potential energy** *is the energy stored in the chemical bonds of a substance*. The various chemicals in gasoline contain a large amount of chemical potential energy that is released when the gasoline is burned in a controlled way in the engine of the car. The release of that energy does two things. Some of the potential energy is transformed into work, which is used to move the car. At the same time, some of the potential energy is converted to heat, making the car's engine very hot. The energy changes of a system occur as either heat or work, or some combination of both.

A dragster is able to accelerate because of the chemical potential energy of its fuel. The burning of the fuel also produces large amounts of heat.

Heat *is energy that is transferred from one object or substance to another because of a difference in temperature between them*. Heat always flows from an object at a higher temperature to an object at a lower temperature (**Figure** below). The flow of heat will continue until the two objects are at the same temperature.



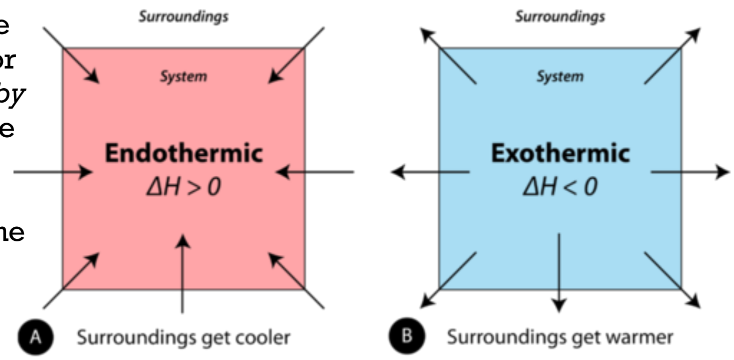
(A) Object A starts with a higher temperature than object B. No heat flows when the objects are isolated from each other. (B) When brought into contact, heat flows from A to B until the temperatures of the two objects are the same.

Thermochemistry *is the study of energy changes that occur during chemical reactions and during changes of state*. When chemical reactions occur, some chemical bonds are broken, while new chemical bonds form. As a result of the rearrangement of atoms, the total chemical potential energy of the system either increases or decreases.

Exothermic and Endothermic Processes

When physical or chemical changes occur, they are generally accompanied by a transfer of energy. The **law of conservation of energy** *states that in any physical or chemical process, energy is neither created nor destroyed*. In other words, the entire energy in the universe is conserved. In order to better understand the energy changes taking place during a reaction, we need to define two parts of the universe, called the system and the surroundings. The **system** *is the specific portion of matter in a given space that is being studied during an experiment or an observation*. The **surroundings** *is everything in the universe that is not part of the system*. In practical terms for a laboratory chemist, the system is generally the reaction being investigated, while the surroundings include the immediate vicinity within the room. During most processes, energy is exchanged between the system and the surroundings. If the system loses a certain amount of energy, that same amount of energy is gained by the surroundings. If the system gains a certain amount of energy, that energy is supplied by the surroundings.

In the study of thermochemical processes, things are viewed from the point of view of the system. A chemical reaction or physical change is **endothermic** *if heat is absorbed by the system from the surroundings*. In the course of an endothermic process, the system gains heat from the surroundings, so the temperature of the surroundings decreases. The quantity of heat for a process is represented by the letter q. The sign of q for an endothermic process is positive because the system is gaining heat. A chemical reaction or physical change is **exothermic** *if heat is released by the system into the surroundings*. Because the surroundings are gaining heat from the system, the temperature of the surroundings increases. The sign of q for an exothermic process is negative because the system is losing heat.



Units of Heat

Heat flow is measured in one of two common units: the calorie and the joule. The joule (J), introduced in the chapter *Measurements*, is the SI unit of energy. The calorie is familiar because it is commonly used when referring to the amount of energy contained within food. A calorie (cal) is the quantity of heat required to raise the temperature of 1 gram of water by 1°C. For example, raising the temperature of 100 g of water from 20°C to 22°C would require 100 × 2 = 200 cal.

Calories contained within food are actually kilocalories (kcal). In other words, if a certain snack contains 85 food calories, it actually contains 85 kcal or 85,000 cal. In order to make the distinction, the dietary calorie is written with a capital C.

1 kilocalorie = 1 Calorie = 1000 calories

To say that the snack "contains" 85 calories means that 85 kcal of energy are released when that snack is processed by your body.

Heat changes in chemical reactions are typically measured in joules rather than calories. The conversion between a joule and a calorie is shown below.

1 J = 0.2390 cal or 1 cal = 4.184 J

We can calculate the amount of heat released in kilojoules when a 400. calorie hamburger is digested.

400. Cal=400. kcal×4.184 kJ1 kcal=1.67×103 kJ

Heat Capacity and Specific Heat

If a swimming pool and a bucket, both full of water at the same temperature, were subjected to the same input of heat energy, the bucket of water would certainly rise in temperature more quickly than the swimming pool. **Heat capacity** *is the amount of heat required to raise the temperature of an object by 1°C*. The heat capacity of an object depends both on its mass and its chemical composition. Because of its much larger mass, the swimming pool of water has a larger heat capacity than the bucket of water.

Specific Heats of Some Common Substances

Substance	Specific Heat (J/g•°C)
Water (*l*)	4.18
Water (*s*)	2.06
Water (*g*)	1.87
Ethanol (*l*)	2.44
Aluminum (*s*)	0.897
Copper (*s*)	0.385
Iron (*s*)	0.449

Different substances respond to heat in different ways. If a metal chair sits in the bright sun on a hot day, it may become quite hot to the touch. An equal mass of water in the same sun will not become nearly as hot. Water is very resistant to changes in temperature, while metals in general are not. The **specific heat** *of a substance is the amount of energy required to raise the temperature of 1 gram of the substance by 1°C*. The table below (**Table** below) lists the specific heats of some common substances. The symbol for specific heat is cp, with the p subscript referring to the fact that specific heats are measured at constant pressure. The units for specific heat can either be joules per gram per degree (J/g•°C) or calories per gram per degree (cal/g•°C). This text will use J/g•°C for specific heat. Note that the specific heat of a substance depends not only on its identity but also its state. For example, ice, liquid water, and steam all have different specific heat values.

Notice that water has a very high specific heat compared to most other substances. Water is commonly used as a coolant for machinery because it is able to absorb large quantities of heat. Coastal climates are much more moderate than inland climates because of the presence of the ocean. Water in lakes or oceans absorbs heat from the air on hot days and releases it back into the air on cool days.

This power plant in West Virginia, like many others, is located next to a large lake so that the water from the lake can be used as a coolant. Cool water from the lake is pumped into the plant, while warmer water is pumped out of the plant and back into the lake.

Specific Heat Calculations

The specific heat of a substance can be used to calculate the temperature change that a given substance will undergo when it is either heated or cooled. The equation that relates heat (q) to specific heat (c_p), mass (m), and temperature change (ΔT) is shown below.

$$q = m \times c_p \times \Delta T$$

The heat that is either absorbed or released is measured in joules. The mass is measured in grams. The change in temperature is given by $\Delta T = T_f - T_i$, where T_f is the final temperature and T_i is the initial temperature.

Sample Problem 17.1: Calculating Specific Heat

A 15.0 g piece of cadmium metal absorbs 134 J of heat as its temperature is increased from 24.0°C to 62.7°C. Calculate the specific heat of cadmium.

Step 1: List the known quantities and plan the problem.

Known

- heat = q = 134 J
- mass = m = 15.0 g
- ΔT = 62.7°C – 24.0°C = 38.7°C

Unknown

- c_p of cadmium = ? J/g•°C

The specific heat equation can be rearranged to solve for the specific heat.

Step 2: Solve.

$$c_p = q/(m \times \Delta T) = (134\text{ J})/(15.0\text{ g} \times 38.7{\circ}\text{C}) = 0.231\text{ J/g}\cdot{\circ}\text{C}$$

Step 3: Think about your result.

The specific heat of cadmium, a metal, is fairly close to the specific heats of other metals in the table above. The result has three significant figures.

Practice Problems

1. How much heat is required to raise the temperature of 13.7 g of aluminum from 25.2°C to 61.9°C?

2. A 274 g sample of air is heated with 2250 J of heat, and its temperature rises by 8.11°C. What is the specific heat of air at these conditions?

Since most specific heats are known, they can be used to determine the final temperature attained by a substance when it is either heated or cooled. Suppose that a 60.0 g sample of water at 23.52°C was cooled by the removal of 813 J of heat. The change in temperature can be calculated using the specific heat equation.

$$\Delta T = q/(c_p \times m) = (-813\ J)/(4.18\ J/g\cdot{}^\circ C \times 60.0\ g) = -3.24{}^\circ C$$

Since the water was being cooled, heat is removed from the system. Therefore, q is negative, and the temperature decreases. The final temperature is:

$$T_f = 23.52°C - 3.24°C = 20.28°C$$

Reviewing Concepts

1. What is one potential use for substances that have a large amount of chemical potential energy? What happens to that energy?

2. Describe what happens when two objects that have different temperatures come into contact with one another.

3. Describe the difference between an endothermic and an exothermic reaction.

4. Two different reactions are performed in two identical test tubes. In reaction A, the test tube becomes very warm as the reaction occurs. In reaction B, the test tube becomes cold. Which reaction is endothermic and which is exothermic? Explain.

5. What is the sign of q for an endothermic process? For an exothermic process?

6. Classify the following as endothermic or exothermic processes.
   a. boiling water
   b. running a race
   c. burning paper
   d. water freezing

Problems

7. 98.3 J of heat is supplied to 12.28 g of a substance, and its temperature **rises by** 5.42°C. What is the specific heat of the substance?

8. 755 J of heat is supplied to 34.0 g of water, and an 755 J of heat is supplied to 34.0 g of iron. If both samples are originally at 20.0°C, calculate the final temperature of the water and the iron. Comment on the difference in your answers, and explain why water is used as a coolant in a car radiator.

9. A quantity of ethanol is cooled from 47.9°C to 12.3°C and releases 3.12 kJ of heat. What is the mass of the ethanol sample?

10. How much heat is absorbed as 7.56 g of ice is heated from −30.0°C to its normal melting point (100°C)?