Most first aid kits contain a small pouch called a cold pack. Cold packs are not cold at first, but they can be activated to become cold. They are useful when someone gets injured and needs an instant ice pack to keep the injured area from swelling.

When you activate a cold pack, you set in motion a chemical reaction that causes a drop in temperature. Other chemical reactions cause temperature to increase. Chemists deal with the temperature changes that accompany chemical reactions when they study *thermochemistry*. What is thermochemistry, and what does it tell us about chemical reactions?

**All chemical reactions involve a change in energy.**

*Thermochemistry* is the study of energy changes that occur in chemical systems. All chemical reactions involve some change in energy as the atoms in reactants recombine to form new products. Depending on the reaction, the energy change may be very small or it may be very large. Each reaction is different, and each reaction can be defined by a specific energy change associated with it.

To analyze this more closely, consider a set of reactants before a reaction occurs. This set of reactants can be considered a system. Everything else is considered the *surroundings*. If the reaction is initiated between the reactants, a change in energy of the system occurs. For any given reaction, the energy of the system (E_{\text{system}}) either increases or decreases as reactants react to form products. If the energy in the system decreases, energy is released into the surroundings. If the energy in the system increases, energy is absorbed from the surroundings.

Reactions that release more energy than they require are called **exothermic reactions**. Those that absorb more energy than they release are called **endothermic reactions**.

**Heat Lost or Heat Gained**

We can calculate the amount of heat lost or gained by substances in chemical reactions. Heat is also called enthalpy (Q or ΔH) and is measured using the mass of the substance, specific heat of the substance (energy required to raise one gram of the substance one degree Celsius), and the change in initial vs. final temperature. Heat is measured in joules (SI) or calories.
Energy and Reactions
All chemical reactions involve a change in energy, as the atoms in reactants recombine to form new products. Energy change can be very large (combustion reactions) or very small, depending on the reaction. It starts with the collision of the reactants. Then there are two other factors that must be satisfied before it is determined whether the reactants can indeed react. The reactants must collide with the proper orientation and with the minimum amount of energy required to start the reaction. This is called the activation energy. Without a collision that has both the right orientation and the minimum activation energy, the reactants will simply bounce off each other and move on. This energy is needed to break the bonds that are in the reactants and is usually in the form of heat or light.

Bond Energy
Since the law of conservation says that energy before and after the reaction must be the same, we can account for the total energy by looking at the bond energies of each substance in the reaction. We can subtract the sum of the energies of the reactants from the sum of the energies of the reactants. The enthalpy ($\Delta H_{\text{rxn}}$) of the reaction can be calculated this way. The sign for this helps us determine whether the reaction is endothermic or exothermic.

Endothermic Reactions
If the products have greater energy than the reactants, the reaction is endothermic. In an endothermic reaction, the enthalpy ($\Delta H$) of the products is greater than the enthalpy of the reactants, showing that energy was absorbed (gained) and $\Delta H$ is positive. Endothermic reactions absorb heat, making them feel cold. Generally, endothermic reactions take place more slowly than exothermic reactions because their activation energies are higher. In an endothermic reaction, heat flows into the system from the surroundings. Therefore, $\Delta H$ is positive from the energy absorbed during the reaction. Examples include the instant cold packs that mix water and ammonium nitrate to cool an area of the body by absorbing heat, and the photosynthetic process of producing sugar.

Exothermic Reactions
If the products have lower energy than the reactants, the reaction is said to be exothermic. Exothermic reactions give off heat, causing the things that were formed to have less energy than they did before they reacted. In an exothermic reaction, the enthalpy of the products is less than the enthalpy of the reactants, meaning that energy was released (lost) during the reaction and $Q$ or $\Delta H$ is negative. Examples include cooking bread, burning a candle, and crystallizing liquid salts (as in sodium acetate in chemical hand warmers).
What Do You Think?

If you look on the package of just about any food item, you will find a nutrition label. This label lists the quantities of fats, proteins, carbohydrates, and various minerals and vitamins present in one serving of the food. This information lets you know about the chemical makeup of the food so that you are aware of the types of nutrients the food will supply to your body.

The nutrition label goes beyond listing the chemical content of the food, however. It also states the number of calories present in a serving of the food item. The number of calories lets you know how much energy your body will take in if you eat a serving of the food. This number is obtained by food chemicals using lab techniques in calorimetry. What is calorimetry, and how does it enable scientists to measure the energy content of foods?

Measuring Energy Changes in Chemical Reactions

Calorimetry is a topic within the larger field of chemistry. The word calorimetry comes from the Latin word calor, which means “heat,” and from the Greek word metron, which means “measure.” Basically, calorimetry is the science of measuring changes in energy that accompany a chemical reaction. Scientists use the term enthalpy change (ΔH) to talk about the energy differences between the reactants and products of a chemical reaction. When reactions occur under constant pressure conditions, all of this energy is in the form of heat. ΔH is often expressed in either joules (J) or calories (cal).

One calorie is equivalent to 4.184 joules.

A food chemist can determine the enthalpy change for a reaction experimentally using a device called a calorimeter. A calorimeter is an insulated container that the chemist can seal after placing reactants inside. The chemist then inserts a thermometer into the calorimeter so that it extends down into the reaction chamber. This allows scientists to monitor the temperature change that occurs when they initiate a reaction between the reactants inside the reaction chamber.
Heat and temperature are not the same thing. Heat is a form of energy; it is measured in calories or joules. Temperature is a measure of the average kinetic energy of the particles in a substance; it is measured in kelvins or degrees. As you will see in the next section, temperature changes measured during a chemical reaction inside a calorimeter cannot be used directly to measure the enthalpy change ($\Delta H$) of the reaction. Instead, the temperature change must be inserted into a mathematical equation along with other factors to calculate the enthalpy change of the reaction.

**Determining Changes in Enthalpy**

So far, we have talked about how a calorimeter can measure changes in temperature. How do scientists measure changes in the enthalpy of a reaction? Before we answer this question, let’s review some terms. First, if pressure remains constant during a reaction, change in enthalpy ($\Delta H$) of a chemical system is equivalent to the heat energy in the system ($q$):

$$\Delta H_{\text{system}} = q$$

Second, $\Delta H$ of a chemical system is equal to the difference between the sum of the enthalpies of products and reactants:

$$\Delta H_{\text{system}} = \sum H_{\text{products}} - \sum H_{\text{reactants}} = q$$

Now that we have reviewed the meanings of these terms, we can turn to another equation that relates $q$ to other factors that scientists can measure in a lab:

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

In this equation, $m$ is the mass of the system (measured in grams), $c$ is the specific heat of the system (measured in J/g°C), and $\Delta T$ (measured in °C) is the change in temperature of the system. This equation allows a scientist to measure the temperature change of a chemical reaction in a calorimeter, and then use that temperature change to calculate the change in enthalpy for that reaction.

The scientist must also measure the system’s mass ($m$) and know the system’s specific heat ($c$). For a chemical reaction, the system is composed of the substances undergoing reaction and the solvent. For reactions taking place in water, the mass is the sum of the water and other reactants. Only the specific heat of water is used because it is present in such high amounts compared to the other substances.

We haven’t yet used specific heat in any calculations, so let’s take a minute to define this term and learn how to use it. The specific heat ($c$) of a substance is defined as the amount of heat required to raise the temperature of one gram of the substance by one degree Celsius. For a particular substance, specific heat is a constant in the same way that boiling point and melting point are constants.
The table below provides specific heat values for several common substances. Note that specific heat is measured in units of joules per gram per degree Celsius (J/g·°C).

<table>
<thead>
<tr>
<th>Substance</th>
<th>Water</th>
<th>Wood</th>
<th>Glass</th>
<th>Copper</th>
</tr>
</thead>
<tbody>
<tr>
<td>Specific Heat (J/g°C)</td>
<td>4.184</td>
<td>1.76</td>
<td>0.84</td>
<td>0.385</td>
</tr>
</tbody>
</table>

**What Do You Think?**

Specific heat is a property of a substance. Each substance has a unique specific heat. To think about this, consider what it feels like to stand on a sandy beach on a hot summer day. When it is very hot, the sand becomes scorching—so hot that it hurts to walk barefoot across the beach to the water.

Yet, when you step into the water, it feels cool. You might wonder why the sand is so hot and the water is so cool when both are exposed to the same conditions of air temperature and sunlight. How can the concept of specific heat solve this puzzle?

*(Here is a hint: The specific heats of water and sand are not the same. In the table above, you can use the specific heat of glass—which is made from sand—to approximate the specific heat of sand. How do the two specific heat values differ? Why does this explain your different experiences with hot sand and cool water?)*

**Measuring the Enthalpy Change for a Reaction**

The procedure for determining ΔH, or change in enthalpy, of a chemical system begins with laboratory work and then proceeds to calculations as outlined below:

1. Measure the masses of reactants: The masses of the reactants will be important later on during the calculation phase of the procedure.
2. Measure the temperature of reactants: Reactants should be allowed to achieve the same initial temperature. Measure and record this value as the initial temperature, Tinitial.
3. Combine reactants in the calorimeter and initiate reaction: Some reactants react spontaneously when combined. Others need to be ignited to begin a reaction.
4. Monitor the temperature until the reaction is complete: The temperature will rise if the reaction is exothermic; the temperature will fall if the reaction is endothermic. Record the temperature every 30 seconds until it levels off. This indicates the reaction has come to a stop. The temperature reading corresponding to this time point can be taken as the ending temperature of the reaction, Tfinal.
5. Calculate ΔT: Subtract Tinitial from Tfinal to determine ΔT: \( \Delta T = T_{final} - T_{initial} \).
6. Calculate q: Use the following equation to calculate the enthalpy change for the reaction (q).
Sum the measured masses of the reactants to obtain the mass of the reaction solution (m). Use
ΔT (calculated above). The value for specific heat (c) can be obtained from reference sources.
For aqueous solutions, use the specific heat of water (4.184 J/g·°C).

\[ q = m \times c \times \Delta T \]

7. Adjust the sign to indicate either an exothermic or endothermic reaction: If the reaction is
exothermic, include a negative sign before the final result. If the reaction is endothermic, include
a positive sign.

Let’s look at a specific example to see how the steps above work. Suppose you have 100 g 1M
HCl and 100 g 1M NaOH. The initial temperature of both solutions is 22.5°C. You combine the
solutions in a calorimeter and observe that the temperature increases as an acid-base reaction
occurs:

\[ \text{HCl(aq)} + \text{NaOH(aq)} \rightarrow \text{H}_2\text{O(l)} + \text{NaCl(aq)} \]
The final temperature of the solution is 29.2°C. What is the enthalpy change of this reaction,
assuming the specific heat of the reaction solution is 4.184 J/g·°C?
Solution: To determine q, calculate m and ΔT, and then multiply these values by c.

\[ m = 100.0 \text{ g HCl(aq)} + 100.0 \text{ g NaOH(aq)} = 200.0 \text{ g} \]
\[ \Delta T = T_{\text{final}} - T_{\text{initial}} = 29.2^\circ\text{C} - 22.5^\circ\text{C} = 6.7^\circ\text{C} \]
\[ q = m \times c \times \Delta T = (200.0 \text{ g})(4.184 \text{ J/g·°C})(6.7^\circ\text{C}) = 5.6 \times 103 \text{ J} = 5.6 \text{ kJ} \]

Because temperature increased during the reaction, this reaction must be exothermic. Therefore,
place a negative sign in front of the ΔH value to give an enthalpy change of −5.6 kJ for this
reaction.
A chemist ran three calorimetry experiments. The table below shows data collected during the experiments. Complete the missing parts of the table.

<table>
<thead>
<tr>
<th>Experiment Number</th>
<th>1</th>
<th>2</th>
<th>3</th>
</tr>
</thead>
<tbody>
<tr>
<td>Specific Heat of a System (J/g°C)</td>
<td>4.184</td>
<td>3.921</td>
<td>2.553</td>
</tr>
<tr>
<td>Mass of System (g)</td>
<td>10.5</td>
<td>5.9</td>
<td>7.4</td>
</tr>
<tr>
<td>Temperature of Reactants (°C)</td>
<td>22.1</td>
<td>20.0</td>
<td>25.3</td>
</tr>
<tr>
<td>Temperature of Products (°C)</td>
<td>34.8</td>
<td>11.3</td>
<td>20.7</td>
</tr>
<tr>
<td>Type of Reaction</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Δ H (kJ)</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Heating or cooling a substance can cause it to undergo a phase change.**
As heat is added to a substance, the kinetic energy increases. As kinetic energy increases, motion increases, eventually breaking intermolecular attractions and causing a state of matter change.

When a substance is going through a phase change, the temperature is held constant. For instance, the melting point of salt (sodium chloride) is 801 degrees Celsius. We can take a piece of solid NaCl and heat it to its melting point. The temperature will stay the same the entire time as long as there is still solid. This is because the energy is being used to break the intermolecular attractions instead of raising the temperature. Once these intermolecular attractions are broken and solid NaCl has turned into molten NaCl, the temperature can increase again.

**Heating Curves**
A heating curve can be plotted out as temperature vs. time. During a heating curve, kinetic energy and particle motion increase as temperature increases. This shows a slope when temperature is rising and a plateau when the substance is in a phase change indicating that the temperature is constant. The phase changes that these heating curves display are melting points and boiling or vaporization points.
Your child can calculate enthalpy and draw a heating curve during a common physical process (melting ice into liquid water followed by heating the water to produce steam).

- Take the initial weight of a small pot.
- Fill a small pot two-thirds full with ice and add water to cover the ice.
- Take the final weight of the small pot. Subtract the weights to get the mass of the ice water. Record this mass.
- Place a thermometer into the pot and allow the thermometer to sit below the water line. Do not let the sensor/thermometer rest on the bottom or side of the pot.
- Carefully swirl the pot until the temperature stabilizes at or below zero degrees Celsius. Record the temperature at 30-second intervals.
- After two minutes (record the time), place the pot on the stove top and add turn on the heat. Warm the ice water and continue to record the temperature at 30-second intervals. **NOTE: Once the stove top is set, leave it at this setting and make no adjustments to the heat.**
- Record the time at which all the ice has melted. Also record the time at which the water begins to boil.
- Continue to heat the water for at least three minutes after the water begins to boil, recording the temperature at 30-second intervals. After three minutes, turn off the stove top and record the time and the final temperature.

Your child can now construct a graph (time on the x-axis and temperature on the y-axis) of a heating curve using the data they just collected.

**Guided questions to ask your students:**

1. Was this experiment an endothermic or exothermic reaction?
2. Explain, in terms of bond energy, what is happening in this reaction.
3. What do the sloped portions of the graph represent?
4. What do the plateaus on the graph represent?
5. Explain, in terms of intermolecular forces, what is happening during a phase change.
6. Calculate the heat of the reaction from 2 degrees Celsius to 60 degrees Celsius (specific heat of water is 4.184 J/g°C and use the mass recorded of the ice water at the beginning of the experiment).