Name: ______________________________________ Date: ________________________

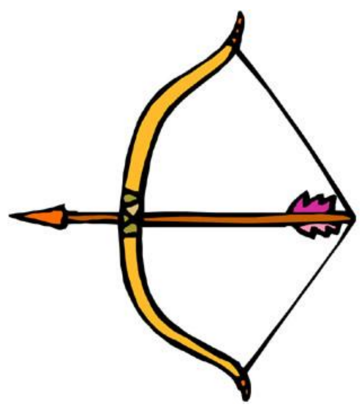
Guided Learning: Chemical Energy

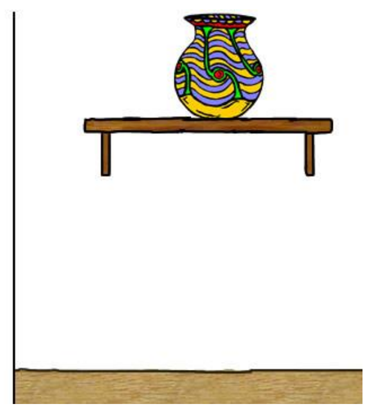
Learning goals
After completing this activity, you will be able to …

- Describe what chemical energy has in common with other forms of potential energy.
- Explain why the formation of chemical bonds causes kinetic energy to be released.
- Define and identify exothermic and endothermic reactions.
- Understand how chemists measure the enthalpy change in a chemical reaction.

Vocabulary: activation energy, catalyst, chemical energy, covalent bond, elastic potential energy, electrostatic forces, endothermic reaction, enthalpy, exothermic reaction, gravitational potential energy, ionic bond, kinetic energy, latent heat of fusion, latent heat of vaporization, metallic bonding, polar, potential energy

Warm-up questions
Potential energy is stored energy that depends on shape or position. For example, a stretched wooden bow has the potential to shoot an arrow. The bow has **elastic potential energy**. A vase placed on a high shelf has the potential to topple over and crash to the floor. The vase has **gravitational potential energy**.





1. In general, what do you have to do to give an object potential energy? (Hint: Think about the bow and vase examples.)

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2. Suppose you moved two magnets a short distance apart. Do these magnets have potential energy? Explain. ____

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3. What would happen if you released the two magnets? ______________________________

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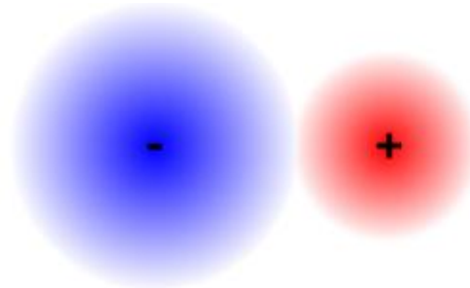
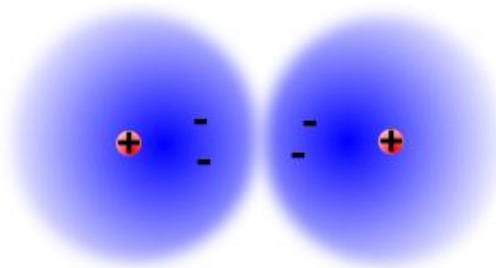
Chemical bonds

Just as magnets are attracted together or repulsed by magnetic forces, atoms are pulled together by **electrostatic forces**. Positive charges are attracted to negative charges, while two positive charges and two negative charges will repel.

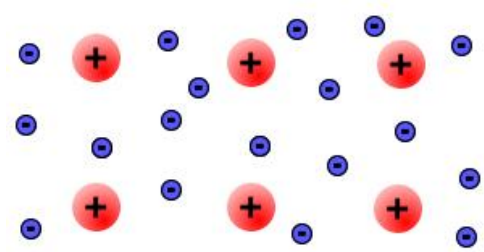
When atoms form chemical bonds, they are held together by electrostatic forces. There are three main types of chemical bonds. In an **ionic bond**, a negatively charged ion (anion) is attracted to a positively charged ion (cation). In a **covalent bond**, electrons are pulled into the space between two positively charged nuclei. Each nucleus is attracted to the electrons in the center. In **metallic bonding**, a free-flowing "sea" of electrons provides the glue that holds nuclei together.



In an ionic bond, a positive ion is attracted to a negative ion.



In a covalent bond, two nuclei are attracted to electrons whose average position is between the atoms.



In metallic bonding, positive nuclei are attracted to freely-moving electrons (and vice-versa).

1. What do all chemical bonds have in common?

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2. Suppose you pull two bonded atoms a short distance apart, and then release them. What do you think would happen?

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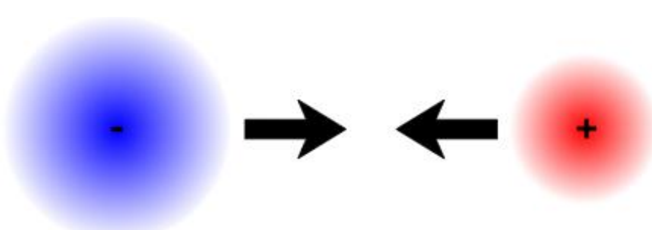
Chemical energy

If you pull two magnets apart, they gain potential energy because they are attracted to one another. Releasing the magnets causes potential energy to be converted into **kinetic energy**, or energy of motion as the magnets are pulled together.



Separated ions have chemical energy. This is converted to kinetic energy as the ions are pulled together.

If two bonded atoms are pulled apart, they gain a form of potential energy called **chemical energy** because they are attracted to one another. If the atoms are released and allowed to bond again, the chemical energy is converted to kinetic energy as the atoms accelerate toward one another.

1. What is the origin of chemical energy? Explain in your own words. ____________________

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2. How is chemical energy similar to elastic potential energy, gravitational potential energy, and magnetic potential energy? ____

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3. If a single pair of atoms were released and attracted to one another, they would accelerate towards each other and gain kinetic energy. What do you think would happen if this occurred to billions of atoms at once? How would the kinetic energy of all of these atoms be expressed?

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Exothermic and endothermic reactions

In most situations, the atoms involved in chemical bonds do not float freely. In a chemical reaction, some chemical bonds are broken and new bonds are formed. For example, consider the reaction of hydrogen (H) and oxygen (O) to form water (H_2O):

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

For this reaction to proceed, energy is required to break the H–H bonds in hydrogen molecules and the O–O bonds in oxygen molecules. The energy required to break these bonds is provided by collisions between molecules—only the most violent collisions at just the right angle will be powerful enough to separate bonded atoms. Energy is then released as the H–O–H bonds form in the water molecules.

In any chemical reaction, energy is absorbed to break existing chemical bonds. Energy is then released as new bonds form. If more energy is released than absorbed in the reaction, the product molecules will have more kinetic energy than the reactant molecules had, and the temperature of the solution will increase. This type of reaction, in which energy is released, is called an **exothermic reaction**. If more energy is absorbed than released, the reaction is an **endothermic reaction**.

The burning of wood is exothermic.

Photosynthesis is endothermic.

Exothermic reactions release energy and cause substances to heat up. The following are examples of exothermic reactions:

- The synthesis of hydrogen and oxygen to produce water
- The combustion (burning) of methane in oxygen
- The rusting of iron to produce iron oxide

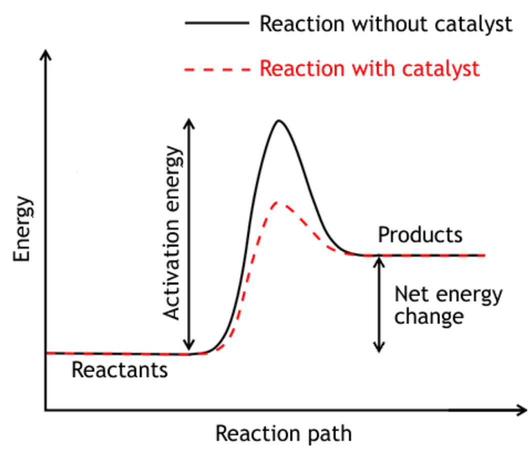
Endothermic reactions absorb energy and cause substances to cool. Examples of endothermic reactions include the following:

- The reaction of water and ammonium chloride in an ice pack
- Photosynthesis

Phase changes also involve the breaking and forming of chemical bonds, and can be classified as exothermic or endothermic. Exothermic phase changes release heat and include condensation and freezing. Endothermic phase changes absorb heat and include melting, boiling, and sublimation.

Because energy is required to break chemical bonds, many chemical reactions do not occur spontaneously. For example, gasoline must be heated to about 250 °C before it begins to burn. The energy required to initiate a chemical reaction is the **activation energy** of the reaction. A **catalyst** can greatly reduce the activation energy needed for a reaction to take place.

The course of a reaction can be plotted on an energy profile (right). If the reactants have a greater energy than the products, the reaction is exothermic. If the reactants have less energy than the products, the reaction is endothermic.



1. How can you tell from observation if a chemical reaction is exothermic or endothermic?

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2. How can you tell from a reaction path diagram if a chemical reaction is exothermic or endothermic? ____

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Which type of reaction is shown in the reaction path above? ________________________

3. Why do many reactions require a high temperature to get started? (Hint: Describe what is happening at the level of individual molecules.)

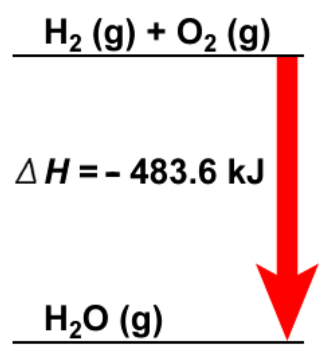
4. **Challenge:** Why do many exothermic reactions only need a spark to begin but then can proceed spontaneously after they start? Describe what is going on at a molecular level.

Enthalpy

Chemists use the term **enthalpy** (*H*) to describe the total energy in a system. This includes the chemical energy, heat, gravitational potential energy, nuclear energy, and so on. While it is difficult to measure the exact value of the enthalpy of a given system, it is usually straightforward to measure the *change* in enthalpy (ΔH).

As long as a reaction proceeds at a constant pressure (such as at atmospheric pressure), the change in enthalpy is equal to the total change in heat (q_P): $\Delta H = q_P$. The subscsript "P" indicates the pressure is constant. A negative value of ΔH means that heat is released into the surrounding environment (exothermic reaction), while a positive value of ΔH indicates heat is absorbed (endothermic reaction). Enthalpy is measured in joules (J).

For example, suppose two moles (4 g) of hydrogen react with excess oxygen to produce water. This reaction produces -483,600 J of heat. This is summarized by the equation below and the enthalpy diagram at right.



$2H_2\ (g) + O_2\ (g) \rightarrow 2H_2O\ (g) \qquad \Delta H = -483.6\ \text{kJ}$

It is necessary to indicate the state of each substance (solid, liquid, or gas) because the enthalpy of a reaction may change depending on the state of the reactants and products.

Enthalpy is used to describe the energy that is absorbed and emitted during phase changes as well. The **latent heat of fusion** is the heat absorbed by a substance during melting (or released while freezing) and the **latent heat of vaporization** is the heat absorbed by a substance during boiling (or released while condensing).

The latent heat of a substance may depend on its molecular structure. For example, water molecules are **polar**, meaning they have a slight positive charge on one side and a slight negative charge on the other. This results in an attractive force between the positive side of one molecule and the negative side of another. Because of this, water molecules are "sticky" and are relatively difficult to separate by melting or boiling. Water has a much higher latent heat of fusion and latent heat of vaporization than similar molecules such as carbon dioxide or methane.

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1. Why is it easier to calculate the change in enthalpy for a reaction than it is to calculate the enthalpy of a substance? ____

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2. Melting a kilogram of ice requires 333 kJ of heat energy. How much heat energy is released by the freezing of one kilogram of water? ____________________________________

3. How does the molecular structure of water affect its latent heat of fusion and latent heat of vaporization? ____

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