

Fuels and Enthalpy

Skills to Develop

- Perform enthalpy calculations and describe the process and results

What is Fuel?

One of the most important applications of chemistry is the study of **fuels**. Fuel basically means a chemical that can provide energy. We use fuel to do nearly everything. Food provides fuel to our bodies. Gasoline provides fuel for our cars. Many different chemicals, from coal and gas to uranium, provide fuel for the power plants that make electric power. When we use a fuel, it becomes a reactant in some type of reaction. Most often, fuels are burned in air to provide heat, and the heat is converted to work. For this reason, we often want to know how much heat we can get out of a chemical reaction.

Enthalpy

When our systems includes chemicals that participate in a reaction, the system may gain or lose heat because of the reaction, and the system may also do work, or work may be done on the system. How would the system do work? Suppose you burn some propane (C_3H_8). Write a balanced equation for this reaction, and convince yourself that there are more moles of gas present in the products than the reactants. Also, you know that burning propane raises the temperature. If you burn some propane, the system will increase in volume, because there is more gas, and it is hotter. (If you burn the propane in a very strong container, then the volume will stay the same and the pressure will increase.) When the volume of the system increases in an open container, it will push on the surroundings, like the atmosphere, and do pV work. We might be able to use this work (in an internal combustion engine like in a car, for instance) but only if we control the volume. If we do the reaction open to the atmosphere, at constant pressure (atmospheric pressure), then we won't be able to use the work done by the reaction.

In chemistry, because reactions are often done at atmospheric pressure, we often want to know how much heat is available from a reaction that occurs at constant pressure. To make this convenient, we can define a new quantity, related to internal energy, which is called **enthalpy**, abbreviated H :

$$H = E + pV \quad (1)$$

Here's why this definition is useful. The work done by a reaction at constant pressure is

$$w = -p\Delta V \quad (2)$$

where p is the pressure, and ΔV is the change in volume of the reaction system, and the sign of w is negative because the system is doing work. The change in enthalpy for a reaction is

$$\Delta H = \Delta(E + pV) = \Delta E + p\Delta V \text{ (at constant } p) \quad (3)$$

Since $\Delta E = q + w$,

$$\Delta H = (q + w) + p\Delta V = (q + w) - w = q \quad (4)$$

Thus, the enthalpy change in a reaction tells us exactly how much heat the reaction can provide if it runs at constant pressure. We don't need to worry about calculating the work done by the reaction pushing back the atmosphere because it is already removed from the definition of enthalpy. Enthalpy is a state function because E , p and V are all state functions. Enthalpy doesn't have a molecular meaning like internal energy, but usually pV is small, so enthalpy is similar to internal energy.

When ΔH is positive, then heat has entered the system, and the process is called **endothermic**. Endothermic reactions feel cold to the touch because they pull heat from your hand into the reaction system. Evaporation is endothermic, which is why sweat helps us cool off. When ΔH is negative, heat leaves the system, and the process is called **exothermic**. Exothermic processes feel hot. This is why flames will burn you.

Outside Links

- [Khan Academy: Enthalpy](#) (15 min)
- [CrashCourse Chemistry: Enthalpy](#) (11 min)

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