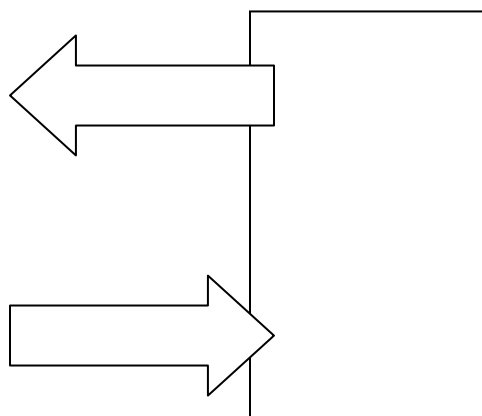
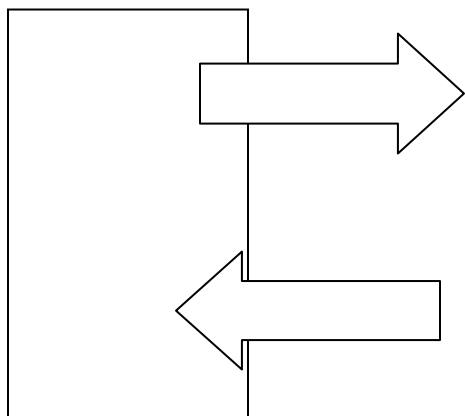


Section 5.1 practice problems

1. What is the kinetic energy, in joules, of a mole of Argon atoms moving with a speed of 650 m/s?
2. Calculate the kinetic energy of a 7.3-kg steel ball traveling at 18.0 m/s.



ΔE is a State function..what does this mean?

**First Law Practice Problem**

1. Calculate ΔE and determine whether the process is endothermic or exothermic for the following cases. (a) A system releases 113 kJ of heat to the surroundings and does 39 kJ of work on the surroundings; (b) $q = 1.62$ kJ and $w = -874$ J; (c) the system absorbs 77.5 kJ of heat while doing 63.5 kJ of work on the surroundings.

Section 5.3 – 5.4 Enthalpy and Enthalpies of reactions

Enthalpy

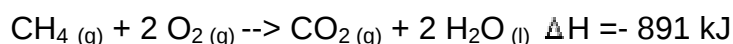
Enthalpy is the heat that is transferred between systems.

We get the very simple result that the change in enthalpy is equal to the heat

$$\Delta H = H_{\text{final}} - H_{\text{initial}} = q_p$$

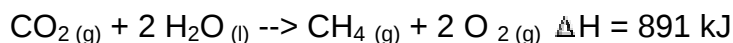
To emphasize that the pressure is constant, we have added a subscript 'p' to the heat in the equation above. Remember that this is not an exotic condition. It is a good approximation when reactions are run in containers open to the atmosphere like a laboratory desk top or even a furnace in a house.

In teaching energy concepts at the high school and college general chemistry level, you will probably spend more time talking about enthalpy than about internal energy. One important concept at this level is whether a chemical reaction produces energy when it occurs or requires that energy be supplied for it to occur. For example, when you burn natural gas in a furnace you expect that heat will be liberated. Natural gas is mostly methane, CH₄, and the combustion reaction is

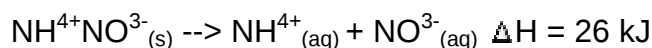


This equation states that when one mole of methane gas combines with two moles of oxygen gas to produce one mole of carbon dioxide gas and two moles of liquid water, 891 kilojoules of enthalpy will be liberated. The value given for enthalpy assumes that the reaction is carried out at 100 kPa (1 bar) and 298 K. The minus sign indicates that the system will *lose* 891 kJ of enthalpy and the surroundings (your house) will *gain* 891 kJ of enthalpy. Since the reaction takes place at constant pressure (in air), the change of enthalpy will be entirely in the form of heat. Reactions, such as the combustion of methane, that give off heat are called [exothermic](#).

Some reactions require heat. Reactions for which ΔH is positive are called [endothermic](#). You can make any exothermic reaction into an endothermic one by reversing the reaction



Note that whether a reaction requires heat (endothermic) or liberates heat (exothermic) says absolutely nothing about whether it is spontaneous or not. Some reactions that occur spontaneously are exothermic. The combustion of methane at 1 bar and 298 K is one such reaction. However, many heats of solution, e.g.



are endothermic under these conditions, as are some reaction that involve the making and breaking of chemical bonds. One example of this is the photosynthesis reaction

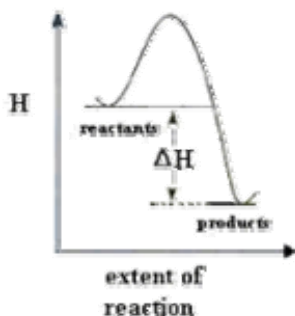


As the temperature is increased, more and more endothermic reactions become spontaneous as we supply the energy, in the form of thermal motion, necessary to drive them.

Enthalpy is an intrinsic property; it depends upon the amount of chemicals you have. If you burn 1/3 mole of methane, your combustion reaction would be



and 297 kJ of heat would be liberated. One gram of methane ($1.0 \text{ g} / 16 \text{ g/mole} = 0.062 \text{ moles}$) would generate 55 kilojoules of heat ($\Delta H = -55 \text{ kJ}$), provided there is sufficient oxygen to burn it completely.



Sometimes, the exothermicity or endothermicity of a reaction is depicted graphically, as shown on the chart to the right. In this chart, the enthalpy is depicted along the y-axis and the reaction proceeds from left to right along the x-axis. For the reaction to occur, chemical bonds first must be broken and, since this requires energy, the reaction must first surmount an energy 'hill' to a higher enthalpy state before the products can form. The graph to the right depicts an endothermic reaction, in which the enthalpy of the products is lower than that of the reactants. For a discussion of rates of reactions and other examples of energy diagrams, including the appearance of exothermic energy diagrams, click [here](#).

It is often possible to make a qualitative estimate of whether a chemical process is endothermic or exothermic. For example, melting ice (or changing any solid into its liquid) takes energy. If you want to melt a solid isobarically, you need to put energy into it in the form of heat. This is because the solid phase of a particular chemical will have much stronger intermolecular interactions than will a liquid phase of the same chemical. Similarly, boiling water takes energy because to transform a liquid into a gas, you have to break the intermolecular interactions holding the liquid together. Both of these processes are endothermic. Perhaps not so obvious is that when water vapor condenses to a liquid under constant pressure conditions, heat is given off. Remember that condensation is the reverse reaction of vaporization. If vaporization (boiling) is exothermic (requires heat), condensation must be endothermic (gives off heat).

The enthalpy of a chemical reaction can also be estimated by considering the bonds that need to be broken to separate the atoms in the molecular reactants and the bonds that need to be formed from these atoms to create the product. You can even get a semi-quantitative measure of the enthalpy of reaction by using [tabulated bond enthalpies](#). These bond energies are defined as the enthalpy of breaking an [average bond](#) which is always positive (exothermic). Combining enthalpies is possible because enthalpy is a state functions which is independent of path. A reaction can be thought of as many steps separating all bonds and then forming the new bonds. This distorted picture of the actual reaction path can be used because enthalpy depends only on the initial and final state and not the pathway.

Reaction enthalpies are estimated by summing enthalpies of all of the bonds broken. Then enthalpies of all of the bonds formed are subtracted from the total as [illustrated here](#).

$$\sum \text{Enthalpies of bonds broken} - \sum \text{Enthalpies of bonds formed} = \text{Enthalpy of reaction}$$

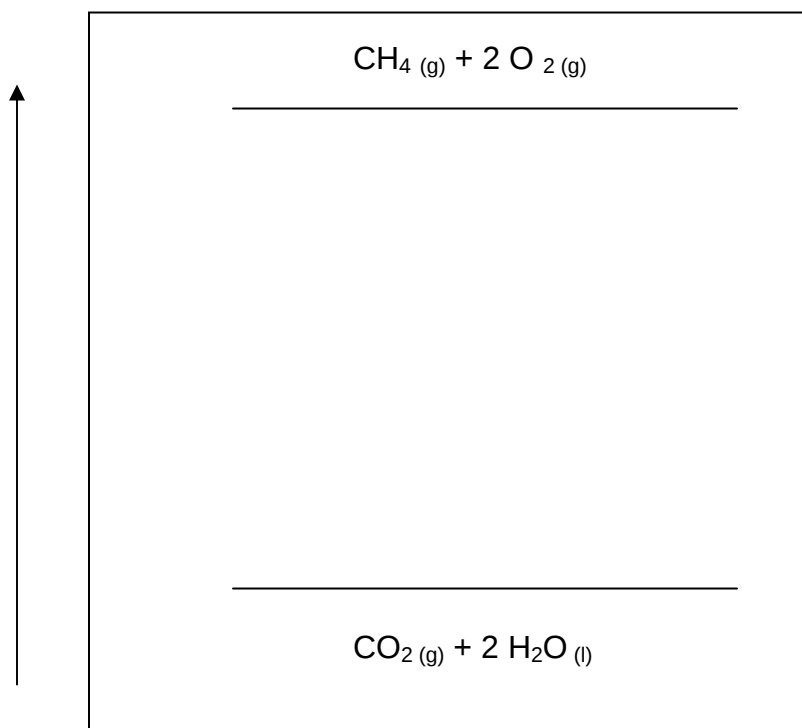
or

$$\Delta H = H_{\text{products}} - H_{\text{reactants}}$$

Reaction enthalpies calculated in this manner are only approximate because different molecules will have slightly different energies for even what appear to be similar bonds. Furthermore, keep in mind that enthalpies are both temperature and pressure dependent. Although the dependency on pressure is usually weak, the dependence on temperature can be very substantial.

We also use Enthalpy Diagrams

For Example: $\text{CO}_2 (\text{g}) + 2 \text{H}_2\text{O} (\text{l}) \rightarrow \text{CH}_4 (\text{g}) + 2 \text{O}_2 (\text{g}) \quad \Delta H = +891 \text{ kJ}$



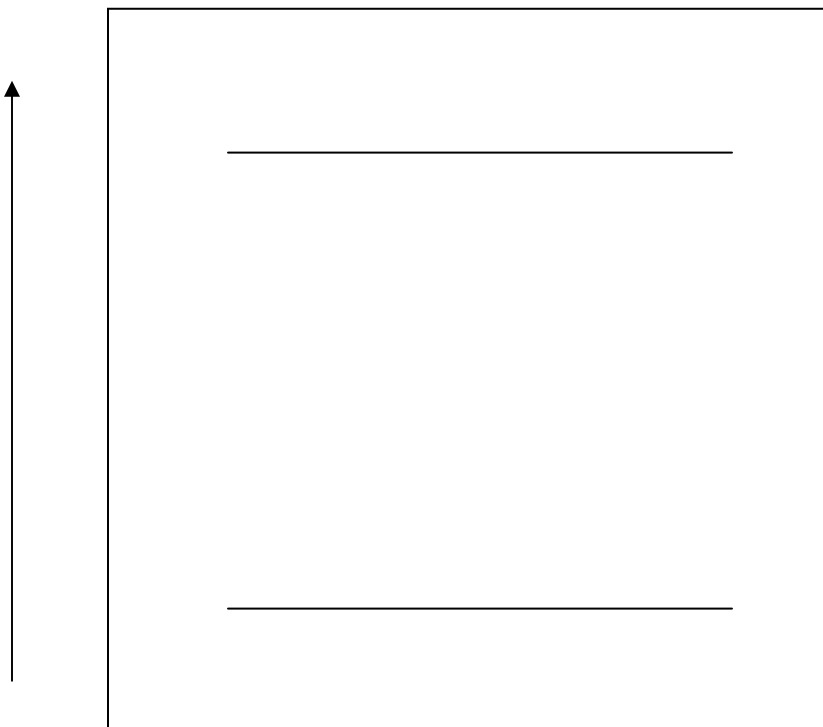
Section 5.3- 5.4 Enthalpy and Enthalpies of reaction practice problem

Hydrogen peroxide can decompose to water and oxygen by the following reaction.



Calculate the value of q when 5.00 g of $\text{H}_2\text{O}_2 (\text{l})$ decomposes at constant pressure

Draw an Enthalpy diagram for the above reaction



Is this reaction exothermic or endothermic? _____

Write the reverse reaction and ΔH below.

