

## Teacher Background

### Chapter 5, Lesson 9

The Teacher Background in Chapter 2 on evaporation and condensation discussed the concept that it takes energy to break bonds and that energy is released when bonds are formed. In the context of evaporation, “breaking bonds” refers to overcoming the attractions between water molecules as they change from a liquid to a gas. In the context of condensation, “making bonds” refers to the attractions of water vapor molecules bringing them together as they change from a gas to a liquid.

It was also explained that the energy “used” in breaking bonds is actually an *energy conversion* from kinetic to potential energy. The energy “released” in making bonds is an energy conversion from potential to kinetic energy.

Since evaporation is only a *bond-breaking* process, it only *uses* energy, resulting in a temperature *decrease*. Therefore, evaporation is endothermic. Since condensation is only a *bond-making* process, it only *releases* energy, resulting in a temperature *increase*. Therefore, condensation is exothermic.

But the process of dissolving involves both bond-making and bond-breaking. In the context of dissolving, “making bonds” refers to water molecules attracting and bonding to the ions or molecules of the solute. “Breaking bonds” refers to the action of the water molecules in separating the ions or molecules of the solute from one another as they go into solution.

The energy conversions in bond-making and bond-breaking in dissolving are the same as in evaporation and condensation. Energy is released as water molecules bond to the solute and energy is used as the ions or molecules of the solute are separated.

The dissolving process is exothermic if more energy is released when water molecules bond to the ions or molecules of the solute than is used to break the bonds holding the solute together.

The dissolving process is endothermic if less energy is released when water molecules bond to the ions or molecules of the solute than is used to break the bonds holding the solute together.

## Teacher Background

### Chapter 5, Lesson 9

The idea of counting molecules was discussed in the Teacher Background section in Chapter 5, lesson 1. There, the question was whether it was fair to do an evaporation test between water and alcohol using one drop of each when these drops contain a different number of molecules. It was explained that there is a way to “count” molecules using the atomic mass and the mole concept. In this way, if needed, an evaporation test could be conducted using the same number of molecules of alcohol and water.

Here, the question is whether it is fair to compare the temperature change when equal masses of two different substances are dissolved. Again, since substances are composed of different atoms, they have different masses. So comparing the same mass of two different substances means that each sample contains a different number of particles.

For example, an equal mass of calcium chloride and potassium chloride must have a different number of units of each substance. Calcium chloride is  $\text{CaCl}_2$  with a mass of  $40 + 2(35.5) = 111$ . This means that 111 grams of calcium chloride contain one mole or  $6.02 \times 10^{23}$  calcium chloride units.

Potassium chloride is  $\text{KCl}$  with a mass of  $39 + 35.5 = 74.5$ . This means that 74.5 grams of potassium chloride contains one mole or  $6.02 \times 10^{23}$  potassium chloride units. So, it takes fewer grams of potassium chloride to have the same number of units as calcium chloride. This means that if you dissolve equal masses of these substances, you are actually using more units of potassium chloride than calcium chloride.

If you wanted to use the same number of units of each, you would need to use masses that are in the same ratio as 111 grams calcium chloride / 74.5 grams potassium chloride which is very close to 1.5/1.