Is freezing exothermic change

The processes on the right side of Figure \(\PageIndex{1}\) – freezing, condensation, and disposal, which are reversed by fusion, sublimation, and evaporation — are exotic. Thus heat pumps that use refrigerants are mainly air conditioners running in reverse. Heat from the environment is used to evaporate the refrigerant, which is then condensed into a liquid in coils in a house to provide heat. The energy changes that occur during phase changes can be quantified using a heat or cooling curve. Figure \(\PageIndex{3}\) shows a heat curve, a temperature plot versus heating time, for a 75 g sample of water. The sample is basically ice at 1 atm and −23°C; as heat is added, the temperature of the ice increases linearly with time. The slope of the line depends on both the mass of the ice and the specific heat (Cs) of ice, which is the number of joules required to increase the temperature of 1 g of ice by 1 ° C. As the temperature of the ice increases, the water molecules in the ice crystal absorb more and more energy and vibrate more powerfully. At the melting point, they have enough kinetic energy to overcome attractive forces and move with respect to each other. As more heat is added, the temperature in the system does not increase further, but remains constant at 0°C until all the ice has melted. When all the ice is converted into liquid water, the temperature of the water again begins to increase. Now, however, the temperature is increasing more slowly than before because the specific heat capacity of water is greater than the ice. When the temperature of the water reaches 100 ° C, the water begins to boil. Here, too, the temperature remains constant at 100 °C until all the water is converted into steam. At this point, the temperature again begins to rise, but faster than seen in the other phases because the heat capacity of steam is less than for ice or water. Figure \(\PageIndex{3}\): A water heat curve. This temperature plot shows what happens to a 75g ice sample initially at 1 ATM and −23 °C, as the heat is added at constant speed: A-B: heating solid ice; B-C: melting ice; C–D: warming of liquid water; D–E: evaporating water; E-F: hot steam. Thus, the temperature of a system does not change during a phase change. In this example, as long as even a small amount of ice is present, the temperature of the system remains at 0 ° C during the melting process, and as long as even a small amount of liquid water is present, the temperature of the system remains at 100 ° C during the cooking process. The speed at which the heat is added does not affect the temperature of the ice/water/steam mixture because the extra heat is used solely to overcome the attractive forces that hold the more condensed phase together. Many chefs believe that the food will cook faster if the heat is higher so that the water boils faster. Instead, the pot of water will boil to dryness earlier, but the temperature of the water does not depend on how powerfully it boils. The temperature of a sample does not change during a phase change. If heat is added at constant speed, as in Figure \(\PageIndex{3}\), the length of the horizontal lines, representing the time when the temperature does not change, is directly proportional to the size of the enthalpies associated with the phase changes. In figure \(\PageIndex{3}\) the horizontal line at 100 °C is much longer than the line at 0 °C because the entalyen for evaporation of water is several times greater than the enthalpy of fusion. An overheated liquid is a sample of a liquid at the temperature and pressure at which there should be a gas. Overheated fluids are not stable; the liquid will eventually boil, sometimes violently. The phenomenon of superheating causes bumping when a liquid heats up in the lab. For example, when a test tube containing water is heated over a Bunsen burner, part of the liquid can easily become too hot. When the overheated liquid is converted into a gas, it can push or bump the rest of the liquid out of the test tube. Placing a battered rod or a small piece of ceramic (a boiling chip) in the test tube allows steam bubbles to form on the surface of the object so that the liquid boils instead of overheating. Superheating is why a liquid heated in a uniform cup in a microwave oven cannot boil until the cup is moved, when the movement of the cup allows bubbles to form. The cooling curve, a temperature plot versus cooling time, in Figure \(\PageIndex{4}\) plots temperature versus time as a 75 g steam sample, originally at 1 atm and 200 °C, is cooled. Although we might expect the cooling curve to be the mirror image of the heat curve in Figure \(\PageIndex{3}\), the cooling curve is not an identical reflection. As the heat is removed from the steam, the temperature drops until it reaches 100°C. At this temperature, the steam begins to condense into liquid water. No further temperature change occurs until all the steam is converted into the liquid; then the temperature again decreases as the water has cooled. We can expect to reach another plateau at 0°C, where the water is converted into ice; in reality, however, this does not always happen. Instead, the temperature often falls below freezing for a while, as shown by the small dip in the cooling basket below 0 °C. This region corresponds to an unstable form of liquid, a super-cooled liquid. If the liquid is allowed to stand, if the cooling continues, or if a small crystal of the solid phase is added (a seed crystal), the super-cooled liquid will convert into a solid, sometimes quite sudden. As the water freezes, the temperature increases slightly until the heat developed during the freezing process and then keeps constant at the melting point as the rest of the water freezes. Then the temperature of the ice decreases again as more heat is removed from the system. Figure \(\PageIndex{4}\): A water cooling curve. This temperature plot shows what happens to a 75g steam sample initially at 1 ATM and 200°C, as the heat is removed at constant speed: A-B: cooling steam; B–C: condensing steam; C-D: cool liquid water to provide a super-cooled liquid; D–E: heating the liquid when it starts to freeze; E–F: freezing of liquid water; F-G: cooling ice. Supercooling effects have a huge impact on earth's climate. For example, supercooling water droplets in clouds can prevent the clouds from releasing precipitation over regions that are persistently dry as a result. Clouds consist of small drops of water, which in principle should be dense enough to fall like rain. In fact, however, the drops must accumulate to reach a certain size before they can fall to the ground. Usually, a small particle (a nucleus) is required for the drops to be aggregated; the nucleus can be a dust particle, an ice crystal or a particle of silver iodide dispersed in a cloud during sowing (a method of inducing rain). Unfortunately, the small water droplets usually remain as a super-cooled liquid down to about 10°C, instead of freezing in ice crystals that are more suitable cores for raindrop formation. One approach to producing precipitation from an existing cloud is to cool the water droplets so that they crystallize to give nuclei around which raindrops can grow. This is best done by spreading small granules of solid CO2 (dry ice) into the cloud from an aircraft. Solid CO2 is sublimated directly to the gas at pressures of 1 atm or lower, and enthalpy of sublimation is significant (25.3 kJ / mol). As CO2 sublimes, it absorbs heat from the cloud, often with the desired results. Example \(\PageIndex{1}\): Cooled If a 50.0 g ice cube at 0.0 °C is added to 500 ml of tea at 20.0 °C, what is the temperature of the tea when the ice cube has just melted? Assume that no heat is transmitted to or from the environment. The density of water (and iced tea) is 1.00 g/ml above the range of 0 °C –20 °C, the specific hotness of liquid water and ice is 4,184 J/(g•°C) and 2,062 J/(g•°C), respectively), and enthalpy of ice fusion is 6.01 kJ/mol, respectively. Given: mass, volume, initial temperature, density, specific heat and \(Δ H_{fus}\) Requested: final temperature Strategy Replace the given values in the general equation of heat obtained (by the ice) to heat lost (by tea) to achieve the final temperature of the mixture. Solution When two substances or objects at different temperatures are brought into contact, the heat will flow from warmer to cooler. The amount of heat flowing is of \[q=mC_s ΔT \] where \(q\) is hot, \(m\) is mass, \(C_s\) is the specific heat, and \(ΔT\) is the temperature change. Finally, the temperatures of the two substances will be equal to a value somewhere between their original temperatures. Calculating the temperature of iced tea after adding an ice cube is a little more complicated. The general equation of heat obtained and heat lost is still valid, but in this case we also need to take into account how much heat is required to melt the ice cube from ice at 0.0 ° C to liquid water at 0.0 ° C. The amount of heat that the ice cube gets when it melts is determined by its entalpy of fusion in kJ/mol: \[q=nΔ H_{fus} \] For our 50.0 g ice cube: \[q_ {ice} = 50.0 g\dfrac{1\: mol}{18.02\:g}↑6.01\: kJ/mol = 16.7 kJ \] Thus, when the ice cube has just melted, it has absorbed 16.7 kJ heat from tea. We can then replace this value in the first formula to determine the temperature change: \[q_{tea} = - 16,700 J =500 ml\dfrac{1.00\:g}{1\: mL}↑4.184 J/(g•°C) ΔT \] \[ΔT = -7.98 °C = T_f - T_i \] \[T_f = 12.02 °C \] This would be the temperature of the tea when the ice cube has just finished melting; However, this leaves the melted ice still at 0.0 °C. We will perhaps more practically know what the final temperature of the mixture of tea will be when melted ice has reached thermal equilibrium with tea. To find out, we can add one more step to the calculation by connecting to the general equation of heat obtained and heat lost again: \[q_{ice} = - q_{tea} \] \] \[q_{ice} = m_{ice}C_s ΔT = 50.0g↑4.184 J/184 J/184 J/184 J/184 184 J/184 J/184 J/184 J/184 J/184 J/184 (g•°C)↑(T_f - 0.0°C) = 209.2 J/°C T_f \] \[q_{te} = m_{te}C_s ΔT = 500g-4.184 J/(g•°C)(T_f - 12.02°C) = 2092 J/°C T_f - 25,150 J\]T_f[209.2 J/°C T_f= - 2092 J/°C <2> <7> + 25,150 J \] \] \[2301.2 J/°C T_f = 25,150 J \] \[T_f = 10.9 °C \] The final temperature is between the first temperatures of the tea (12.02 °C) and molten ice (0.0 °C) , so this answer makes sense. In this example, the tea loses much more heat in melting the ice than in mixing with cold water, which shows the importance of taking into account the heat of phase changes! Exercise \(\PageIndex{1}\): Death by freezing Assume that you get overrun by a blizzard while skiing and you seek refuge in a tent. You're thirsty, but you forgot to bring liquid water. You have a choice to eat a handful of snow (say 400 g) at -5.0°C immediately to quench your thirst or set up the propane oven, melt the snow and heat the water to body temperature before drinking it. You remember that the survival guide you browsed through at the hotel said something about not eating snow, but you can't remember why – after all, it's just frozen water. To understand the guide's recommendation, calculate how much heat your body needs to to bring 400 g of snow at −5.0 °C to the body's internal temperature of 37 °C. Use the data in Example \(\PageIndex{1}\) Answer 200 kJ (4.1 kJ to bring the ice from −5.0 °C to 0.0 °C, 133.6 kJ to melt the ice at 0.0 °C, and 61.9 kJ to bring the water from 0.0 °C to 37 °C), which is energy that would not have been used had you first melted the snow. Snow.