
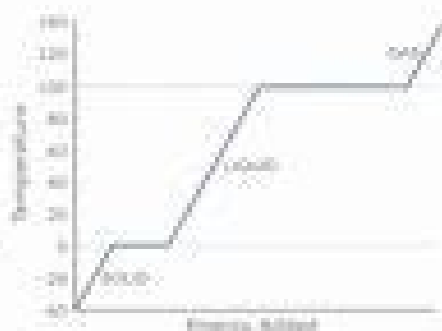


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### The Heating curve of water



This is a graph showing what phase changes if you have a piece of ice and heated it over time. When the temperature is measured it goes up and up most of the time but at certain points it stays the same. By studying these patterns we can know what water is doing at these points and what this can tell us.

- At what temperature does water undergo the first?
- Between what temperatures is the graph in the liquid or solid?
- What is happening to the substance when the graph goes flat at 0°C? The 100°C?

They exhibit essentially no surface tension and very low viscosities, however, so they can more effectively penetrate very small openings in a solid mixture and remove soluble components. These properties make supercritical fluids extremely useful solvents for a wide range of applications. For example, supercritical carbon dioxide has become a very popular solvent in the food industry, being used to decaffeinate coffee, remove fats from potato chips, and extract flavor and fragrance compounds from citrus oils. It is nontoxic, relatively inexpensive, and not considered to be a pollutant. After use, the CO<sub>2</sub> can be easily recovered by reducing the pressure and collecting the resulting gas. Figure 1 (PageIndex{7}): (a) A sealed container of liquid carbon dioxide slightly below its critical point is heated, resulting in (b) the formation of the supercritical fluid phase. Cooling the supercritical fluid lowers its temperature and pressure below the critical point, resulting in the reestablishment of separate liquid and gaseous phases (c and d). Colored floats illustrate differences in density between the liquid, gaseous, and supercritical fluid states. (credit: modification of work by "mrmrobin"/YouTube) Example (PageIndex{4}): The Critical Temperature of Carbon Dioxide If we shake a carbon dioxide fire extinguisher on a cool day (18 °C), we can hear liquid CO<sub>2</sub> sloshing around inside the cylinder. However, the same cylinder appears to contain no liquid on a hot summer day (35 °C). Explain these observations. Solution On the cool day, the temperature of the CO<sub>2</sub> is below the critical temperature of CO<sub>2</sub>, 304 K or 31 °C (Table (PageIndex{1})), so liquid CO<sub>2</sub> is present in the cylinder. On the hot day, the temperature of the CO<sub>2</sub> is greater than its critical temperature of 31 °C. Above this temperature no amount of pressure can liquefy CO<sub>2</sub> so no liquid CO<sub>2</sub> exists in the fire extinguisher. Exercise (PageIndex{4}): Ammonia can be liquefied by compression at room temperature; oxygen cannot be liquefied under these conditions. Why do the two gases exhibit different behavior? Answer The critical temperature of ammonia is 405.5 K, which is higher than room temperature. The critical temperature of oxygen is below room temperature; thus oxygen cannot be liquefied at room temperature. Coffee is the world's second most widely traded commodity, following only petroleum. Across the globe, people love coffee's aroma and taste. Many of us also depend on one component of coffee—caffeine—to help us get going in the morning or stay alert in the afternoon.

But late in the day, coffee's stimulant effect can keep you from sleeping, so you may choose to drink decaffeinated coffee in the evening. Since the early 1900s, many methods have been used to decaffeinate coffee. All have advantages and disadvantages, and all depend on the physical and chemical properties of caffeine. Because caffeine is a somewhat polar molecule, it dissolves well in water, a polar liquid. However, since many of the other 400-plus compounds that contribute to coffee's taste and aroma also dissolve in H<sub>2</sub>O, hot water decaffeination processes can also remove some of these compounds, adversely affecting the smell and taste of the decaffeinated coffee. Dichloromethane (CH<sub>2</sub>Cl<sub>2</sub>) and ethyl acetate (CH<sub>3</sub>CO<sub>2</sub>CH<sub>3</sub>) have similar polarity to caffeine, and are therefore very effective solvents for caffeine extraction, but both also remove some flavor and aroma components, and their use requires long extraction and cleanup times. Because both of these solvents are toxic, health concerns have been raised regarding the effect of residual solvent remaining in the decaffeinated coffee. Figure (PageIndex{8}): (a) Caffeine molecules have both polar and nonpolar regions, making it soluble in solvents of varying polarities. (b) The schematic shows a typical decaffeination process involving supercritical carbon dioxide. Supercritical fluid extraction using carbon dioxide is now being widely used as a more effective and environmentally friendly decaffeination method (Figure (PageIndex{8})).

At temperatures above 304.2 K and pressures above 7376 kPa, CO<sub>2</sub> is a supercritical fluid, with properties of both gas and liquid. Like a gas, it penetrates deep into the coffee beans; like a liquid, it effectively dissolves certain substances. Supercritical carbon dioxide extraction of steamed coffee beans removes 97–99% of the caffeine, leaving coffee's flavor and aroma compounds intact. Because CO<sub>2</sub> is a gas under standard conditions, its removal from the extracted coffee beans is easily accomplished, as is the recovery of the caffeine from the extract. The caffeine recovered from coffee beans via this process is a valuable product that can be used subsequently as an additive to other foods or drugs. Video (PageIndex{4}): An overview of phase changes and phase diagrams. The temperatures at which phase transitions occur are determined by the relative strengths of intermolecular attractions and are, therefore, dependent on the chemical identity of the substance. The temperature and pressure conditions at which a substance exists in solid, liquid, and gaseous states are summarized in a phase diagram for that substance. Phase diagrams are combined plots of three pressure-temperature equilibrium curves: solid-liquid, liquid-gas, and solid-gas. These curves represent the relationships between phase-transition temperatures and pressures. The point of intersection of all three curves represents the substance's triple point—the temperature and pressure at which all three phases are in equilibrium. At pressures below the triple point, a substance cannot exist in the liquid state, regardless of its temperature.

The terminus of the liquid-gas curve represents the substance's critical point, the pressure and temperature above which a liquid phase cannot exist. Key Equations 
$$P=Ae^{-\frac{\Delta H_{\text{vap}}}{RT}}$$
 
$$\ln P=-\frac{\Delta H_{\text{vap}}}{RT}+\ln A$$
 
$$\ln\left(\frac{P_2}{P_1}\right)=-\frac{\Delta H_{\text{vap}}}{R}\left(\frac{1}{T_2}-\frac{1}{T_1}\right)$$
 critical point temperature and pressure above which a gas cannot be condensed into a liquid phase diagram pressure-temperature graph summarizing conditions under which the phases of a substance can exist supercritical fluid substance at a temperature and pressure higher than its critical point, exhibits properties intermediate between those of gaseous and liquid states triple point temperature and pressure at which the vapor, liquid, and solid phases of a substance are in equilibrium Contributors Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at 9.110).

Adelaide Clark, Oregon Institute of Technology Crash Course Physics: Crash Course is a division of Complexly and videos are free to stream for educational purposes. Have feedback to give about this text? Click here. Found a typo and want extra credit? Click here. Energy is required to convert a liquid into a gas (i.e. vaporization; a phase change). The energy required is called the heat (or enthalpy) of vaporization and is denoted as ΔH<sub>vap</sub> and is generally given as energy per amount of substance (often kJ mol<sup>-1</sup>). If enthalpy is given as a "per mole" quantity, we use the following equation to determine the amount of heat (q) required based on the amount of substance (n in moles) present. 
$$q = n\Delta H_{\text{vap}}$$
 Water 
$$\Delta H_{\text{vap}} = 44.01 \text{ kJ mol}^{-1}$$
 Substance ΔH° (kJ mol<sup>-1</sup>) Thermodynamic values from openStax The enthalpy of the process is determined as follows: 
$$\Delta H_{\text{vap}} = \Delta H_{\text{final}} - \Delta H_{\text{initial}}$$
 
$$= -241.82 \text{ kJ mol}^{-1} - (-285.83 \text{ kJ mol}^{-1}) = 44.01 \text{ kJ mol}^{-1}$$
 Exercise: Determine the amount of energy (q in kJ) required to transform the follow amounts of water from a liquid to a gas. Amount q (kJ) Example: 1 mol 
$$q = n\Delta H_{\text{vap}}$$
 
$$= 1 \text{ mol} \left( \frac{44.01 \text{ kJ}}{1 \text{ mol}} \right) = 44.01 \text{ kJ}$$
 Methanol 
$$\Delta H_{\text{vap}} = 38.2 \text{ kJ mol}^{-1}$$
 Exercise: Determine the amount of energy (q in kJ) required to transform the follow amounts of water from a liquid to a gas. Amount q (kJ) Heat of Fusion The heat of fusion is the amount of energy required to turn an amount of substance from a solid to a liquid (i.e. melting; a phase change) and is given in units similar to that of heat of vaporization. 
$$q = n\Delta H_{\text{fus}}$$
 Water 
$$\Delta H_{\text{fus}} = 6.01 \text{ kJ mol}^{-1}$$
 Exercise: Determine the amount of energy (q in kJ) required to transform the follow amounts of water from a solid to a liquid Amount q (kJ) Heat is required to heat a substance from one temperature to another without undergoing a phase change. This can be determined across a temperature range (ΔT) for an amount of substance (m) using the specific heat for that substance in the appropriate state (c in J g<sup>-1</sup> °C<sup>-1</sup>). 
$$q = mc\Delta T$$
 Below are the specific heats (in J g<sup>-1</sup> °C<sup>-1</sup>) for water and methanol.

Substance c (J g<sup>-1</sup> °C<sup>-1</sup>) See more enthalpies and specific heats here. Water Exercise: Determine the amount of heat required (in kJ) to heat the following amounts of water from 25 °C to 50 °C.

Amount q (kJ) Note: Which specific heat will you use for this problem? What state of matter is water in for the specified temperature range? 
$$\Delta T = 18.02 - \text{math}\{g\} = 1 - \text{math}\{mol\}$$
 
$$\left( \frac{18.02 - \text{math}\{g\}}{\text{math}\{mol\}} \right) \left( \text{math}\{g\} \right) \left( 50 - \text{math}\{C\} - \text{math}\{C\} \right) = 100.0 - \text{math}\{C\}$$
 
$$25 - \text{math}\{C\} \left( \text{math}\{C\} \right) \left( \frac{\text{math}\{kJ\}}{10^3 - \text{math}\{J\}} \right) = 1.88 - \text{math}\{kJ - mol^{-1}\}$$
 Note: The amount of substance (given in moles) had to be converted to mass (in g) since the specific heat was given in g. Methanol Repeat the process for ethanol (CH<sub>3</sub>OH) across a temperature range of 25 °C to 50 °C.

Amount q (kJ) Note: Which specific heat will you use for this problem? What state of matter is water in for the specified temperature range? You have to know the melting and boiling points of methanol. Additional Practice: Repeat the exercise for water and methanol across the following temperature range: Combining it All We can combine each individual concept into an overall heating curve problem that involves one or more phase change as well as the heating of a substance in a particular phase. Water Determine the amount of heat (in kJ mol<sup>-1</sup>) required to heat 500.0 g water from -50 °C to 150 °C. 
$$\Delta H_{\text{vap}} = 44.01 \text{ kJ mol}^{-1}$$
 
$$\Delta H_{\text{fus}} = 6.01 \text{ kJ mol}^{-1}$$
 
$$c_{\text{ice}} = 2.09 \text{ J g}^{-1} \text{ }^{\circ}\text{C}^{-1}$$
 
$$c_{\text{water}} = 4.18 \text{ J g}^{-1} \text{ }^{\circ}\text{C}^{-1}$$
 
$$c_{\text{steam}} = 2.01 \text{ J g}^{-1} \text{ }^{\circ}\text{C}^{-1}$$
 Fill in the phase labels for this problem. Heating the solid 
$$T_{\text{initial}} = -50.0 \text{ }^{\circ}\text{C}$$
 
$$T_{\text{melting}} = 0.0 \text{ }^{\circ}\text{C}$$
 Heat the solid up to its melting point. The normal melting point of water is 0.0 °C. 
$$q_1 = mc\Delta T$$
 Melting the Solid 
$$T_{\text{melting}} = 0.0 \text{ }^{\circ}\text{C}$$
 
$$T_{\text{melting}} = 0.0 \text{ }^{\circ}\text{C}$$
 A phase transition occurs at the melting point of water. 
$$q_2 = n\Delta H_{\text{fus}}$$
 Heating the liquid 
$$T_{\text{boiling}} = 100.0 \text{ }^{\circ}\text{C}$$
 
$$T_{\text{boiling}} = 100.0 \text{ }^{\circ}\text{C}$$
 Heat the liquid up to its boiling point. The normal boiling point of water is 100.0 °C. 
$$q_3 = mc\Delta T$$
 Vaporizing the Liquid 
$$T_{\text{boiling}} = 100.0 \text{ }^{\circ}\text{C}$$
 
$$T_{\text{boiling}} = 100.0 \text{ }^{\circ}\text{C}$$
 A phase transition occurs at the boiling point of water. 
$$q_4 = n\Delta H_{\text{vap}}$$
 Heating the gas 
$$T_{\text{final}} = 150.0 \text{ }^{\circ}\text{C}$$
 
$$T_{\text{final}} = 150.0 \text{ }^{\circ}\text{C}$$
 Heat the gas up from the normal point boiling point of water to the final temperature of 150 °C. 
$$q_5 = mc\Delta T$$
 Combine all the heats 
$$q_{\text{total}} = q_1 + q_2 + q_3 + q_4 + q_5$$
 Methanol Repeat the exercise for methanol. Note that you will need to look up the melting and boiling points first! Additional Practice Handout: Acetone More practice problems