Chem With Com

Intermolecular Forces, States of Matter, and Energy

- Chapter 1.3 (Physical and Chemical Properties)
- Chapter 10.1 (Intermolecular Forces)
- Chapter 10.2 (Properties of Liquids)
- Chapter 10.3 (Phase Transitions) – heating and cooling curves only
- Chapter 10.4 (Phase Diagrams)
- Chapter 9.1 (Energy Basics)

Adapted From: Paul Flowers, Edward J. Neth, William R. Robinson, PhD, Klaus Theopold, Richard Langley; Chemistry: Atoms First 2e; OpenStax; Feb 14, 2019; Houston Texas; https://openstax.org/books/chemistry-atoms-first-2e/pages/1-introduction

Intermolecular Forces

- Chapter 1.3 (Physical and Chemical Properties)
- Chapter 10.1 (Intermolecular Forces)
- Chapter 10.2 (Properties of Liquids)

Adapted From: Paul Flowers, Edward J. Neth, William R. Robinson, PhD, Klaus Theopold, Richard Langley; Chemistry: Atoms First 2e; OpenStax; Feb 14, 2019; Houston Texas; https://openstax.org/books/chemistry-atoms-first-2e/pages/1-introduction

Materials experiencing physical changes have different physical properties

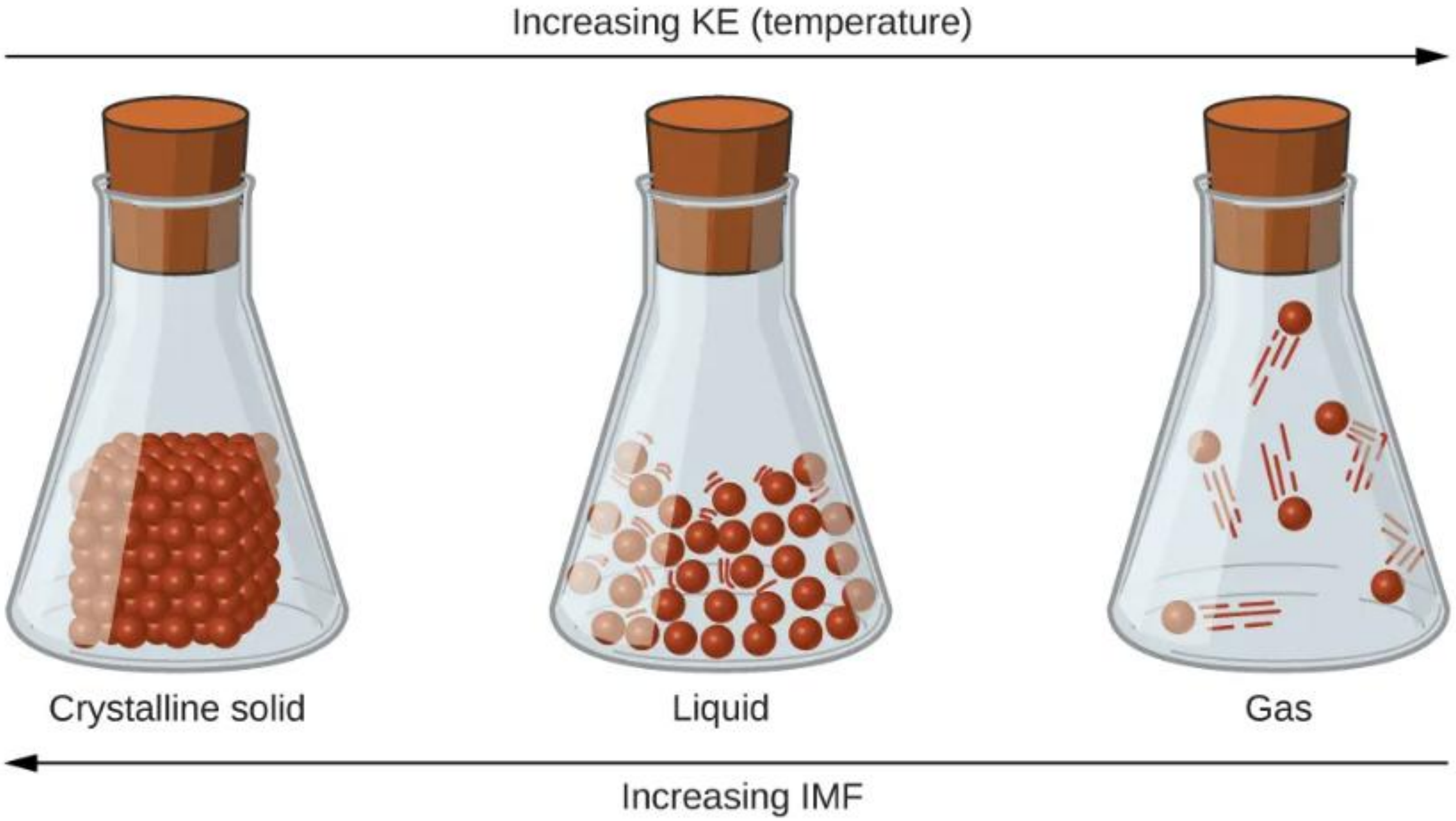


Figure 10.2 Transitions between solid, liquid, and gaseous states of a substance occur when conditions of temperature or pressure favor the associated changes in intermolecular forces. (Note: The space between particles in the gas phase is much greater than shown.)

Intermolecular Forces govern different physical states

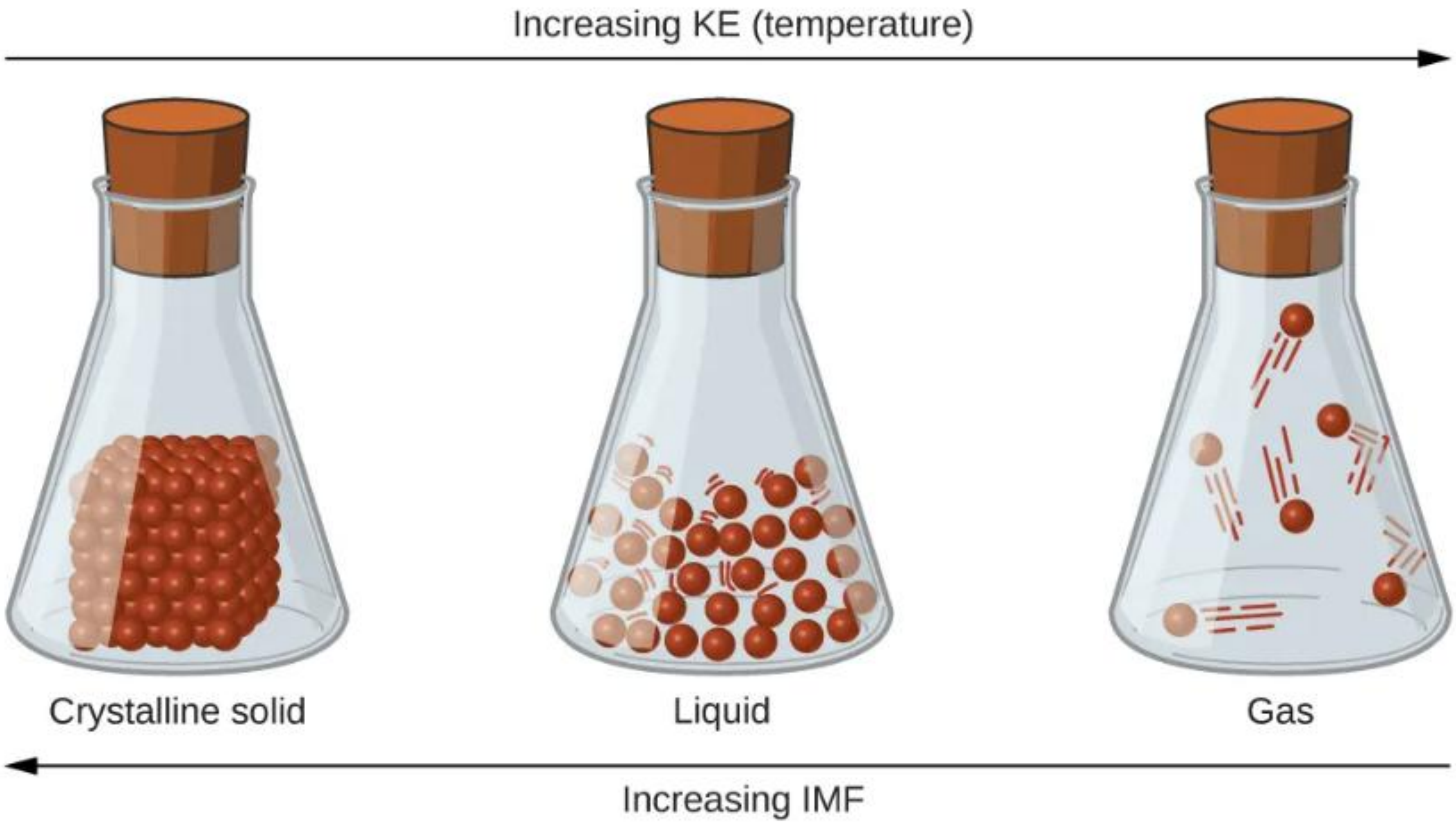


Figure 10.2 Transitions between solid, liquid, and gaseous states of a substance occur when conditions of temperature or pressure favor the associated changes in intermolecular forces. (Note: The space between particles in the gas phase is much greater than shown.)

Intermolecular Forces are weaker than covalent or ionic bonds

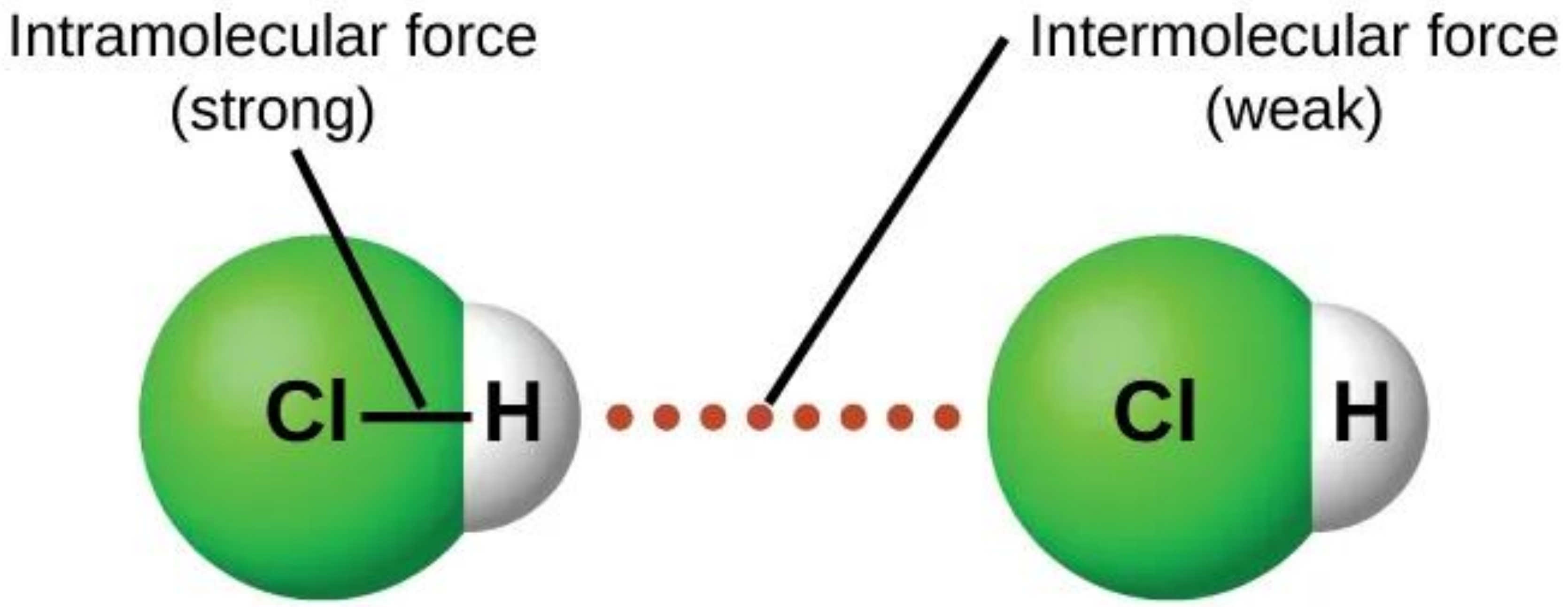
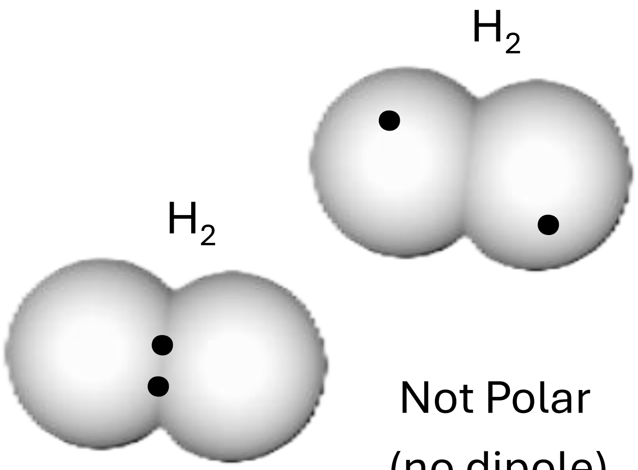


Figure 10.5 *Intra*molecular forces keep a molecule intact. *Inter*molecular forces hold multiple molecules together and determine many of a substance's properties.

Intermolecular Forces

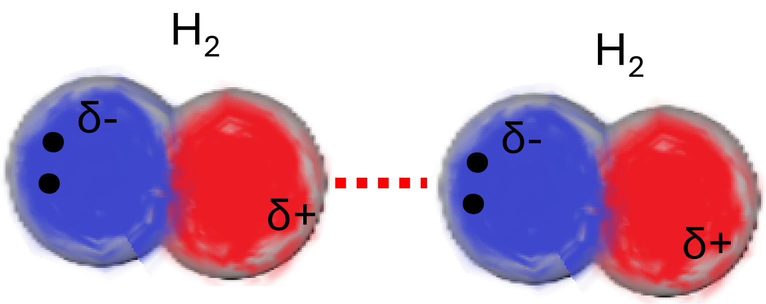
Weakest

Dispersion Forces



Not Polar

(no dipole)



Temporarily Polar

(temporary instantaneous dipole)

Dipole-Dipole Attractions

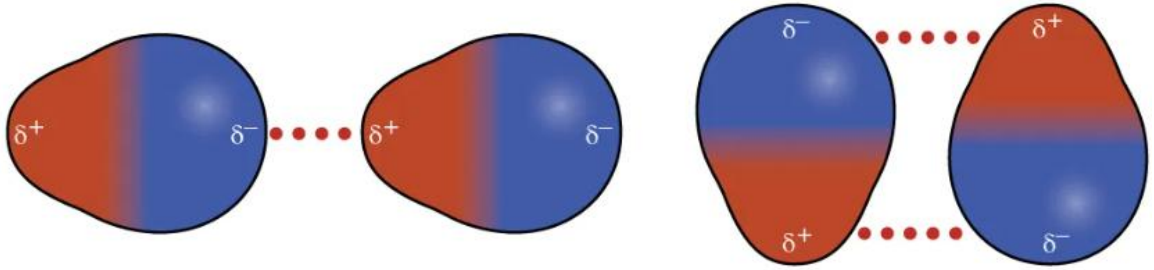


Figure 10.9 This image shows two arrangements of polar molecules, such as HCl, that allow an attraction between the partial negative end of one molecule and the partial positive end of another.

Hydrogen Bonds

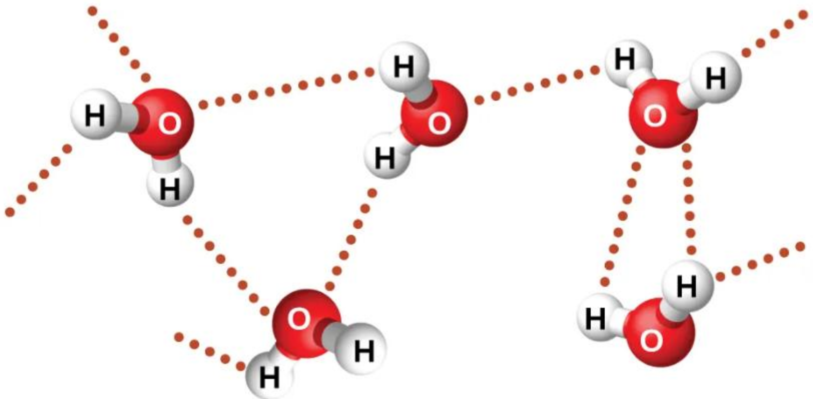


Figure 10.10 Water molecules participate in multiple hydrogen-bonding interactions with nearby water molecules.

Strongest

Dipole-Dipole Attractions (polar-polar)

Occur between two or more polar molecules

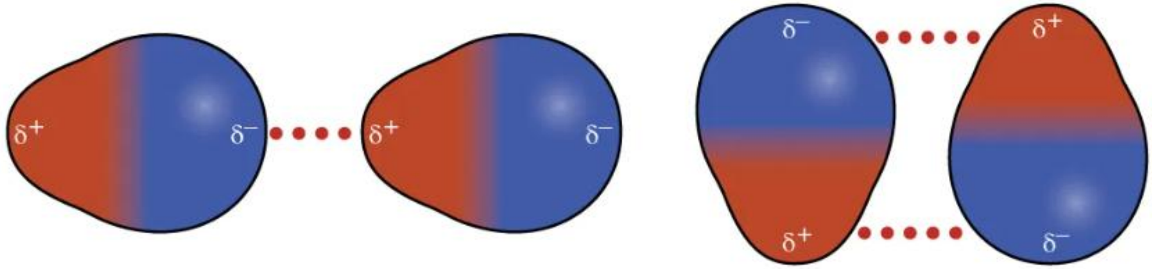


Figure 10.9 This image shows two arrangements of polar molecules, such as HCl, that allow an attraction between the partial negative end of one molecule and the partial positive end of another.

Hydrogen Bonds

Something can hydrogen bond if a hydrogen is directly covalently bonded to a nitrogen, oxygen, or fluorine (the three most electronegative atoms on the periodic table)

H-N

H-O

H-F

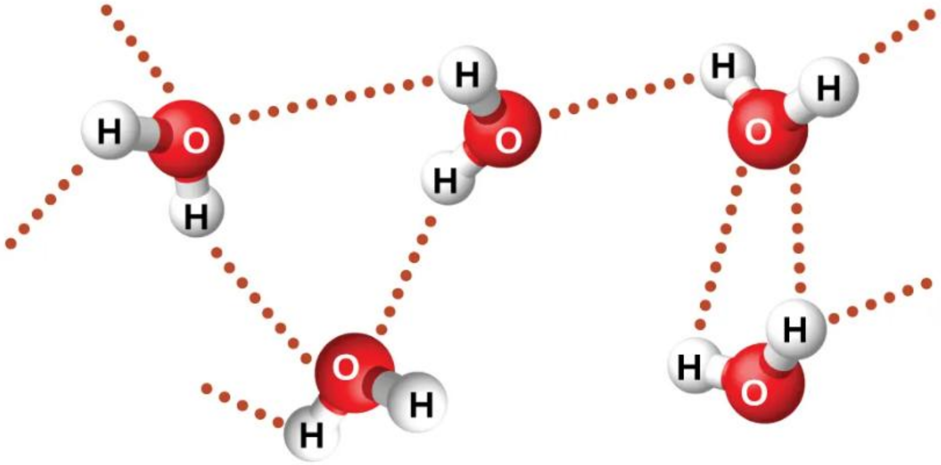
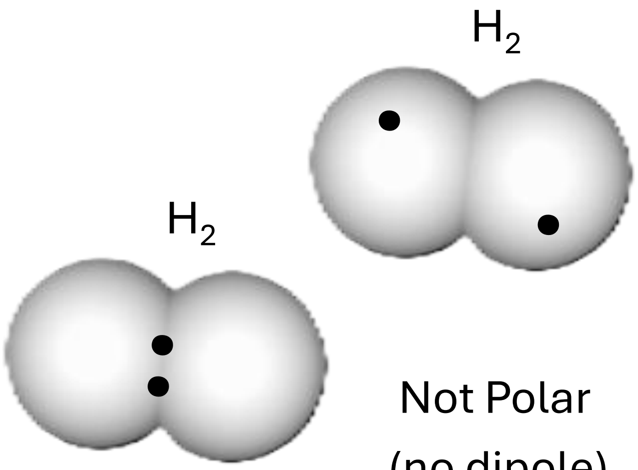


Figure 10.10 Water molecules participate in multiple hydrogen-bonding interactions with nearby water molecules.

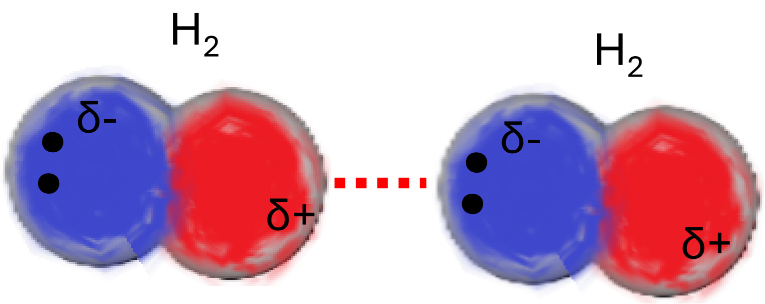
Dispersion Forces

Temporarily dipole moment (temporarily polar) that occurs instantaneously in usually nonpolar molecules or atoms



Not Polar

(no dipole)



Temporarily Polar

(temporary instantaneous dipole)

Intermolecular Forces govern different physical states

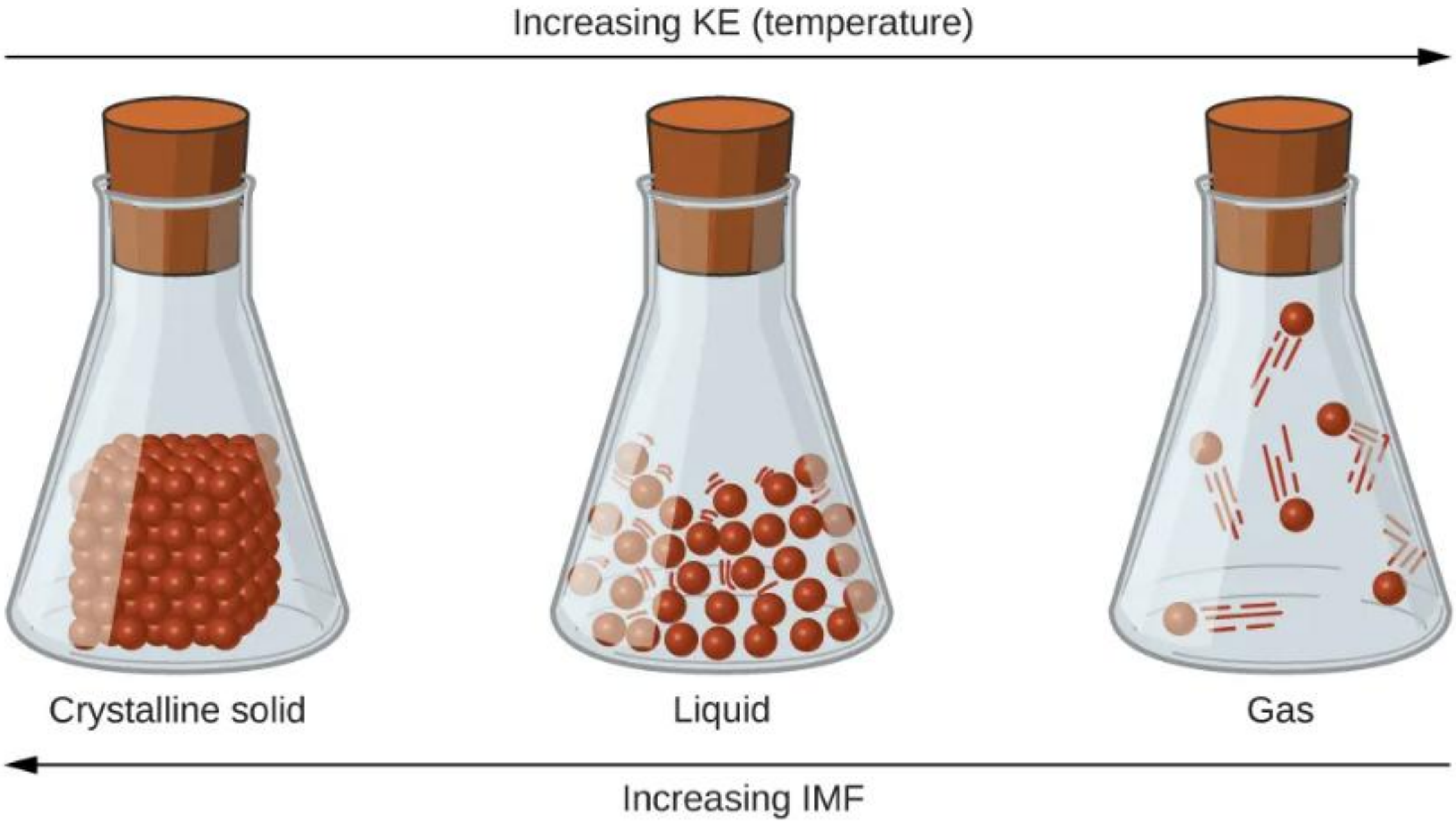
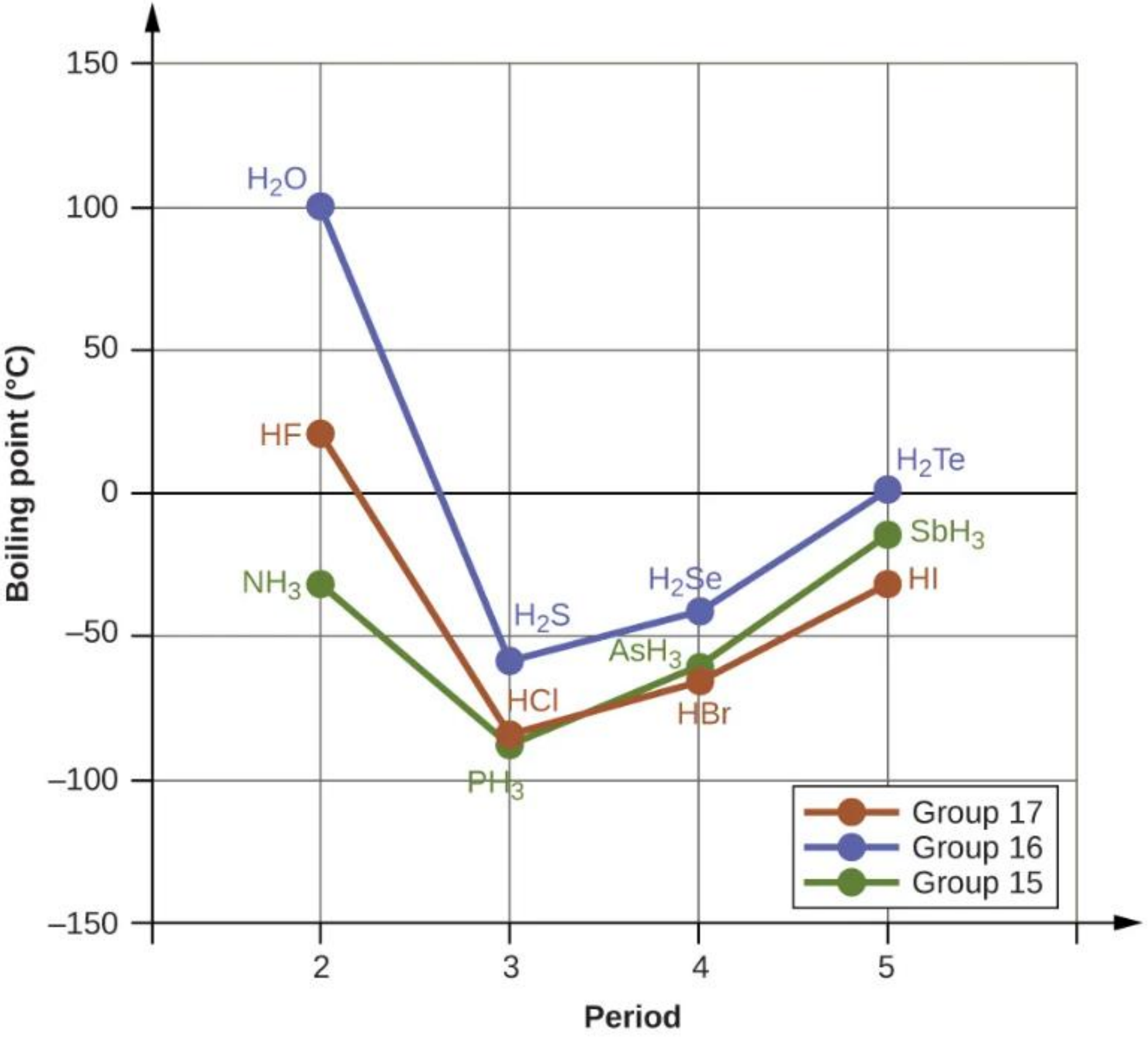


Figure 10.2 Transitions between solid, liquid, and gaseous states of a substance occur when conditions of temperature or pressure favor the associated changes in intermolecular forces. (Note: The space between particles in the gas phase is much greater than shown.)

Stronger Intermolecular Forces mean higher boiling points



Chapter 10.1

Figure 10.12 In comparison to periods 3–5, the binary hydrides of period 2 elements in groups 17, 16 and 15 (F, O and N, respectively) exhibit anomalously high boiling points due to hydrogen bonding.

Viscosity of a liquid is a measure of its resistance to flow

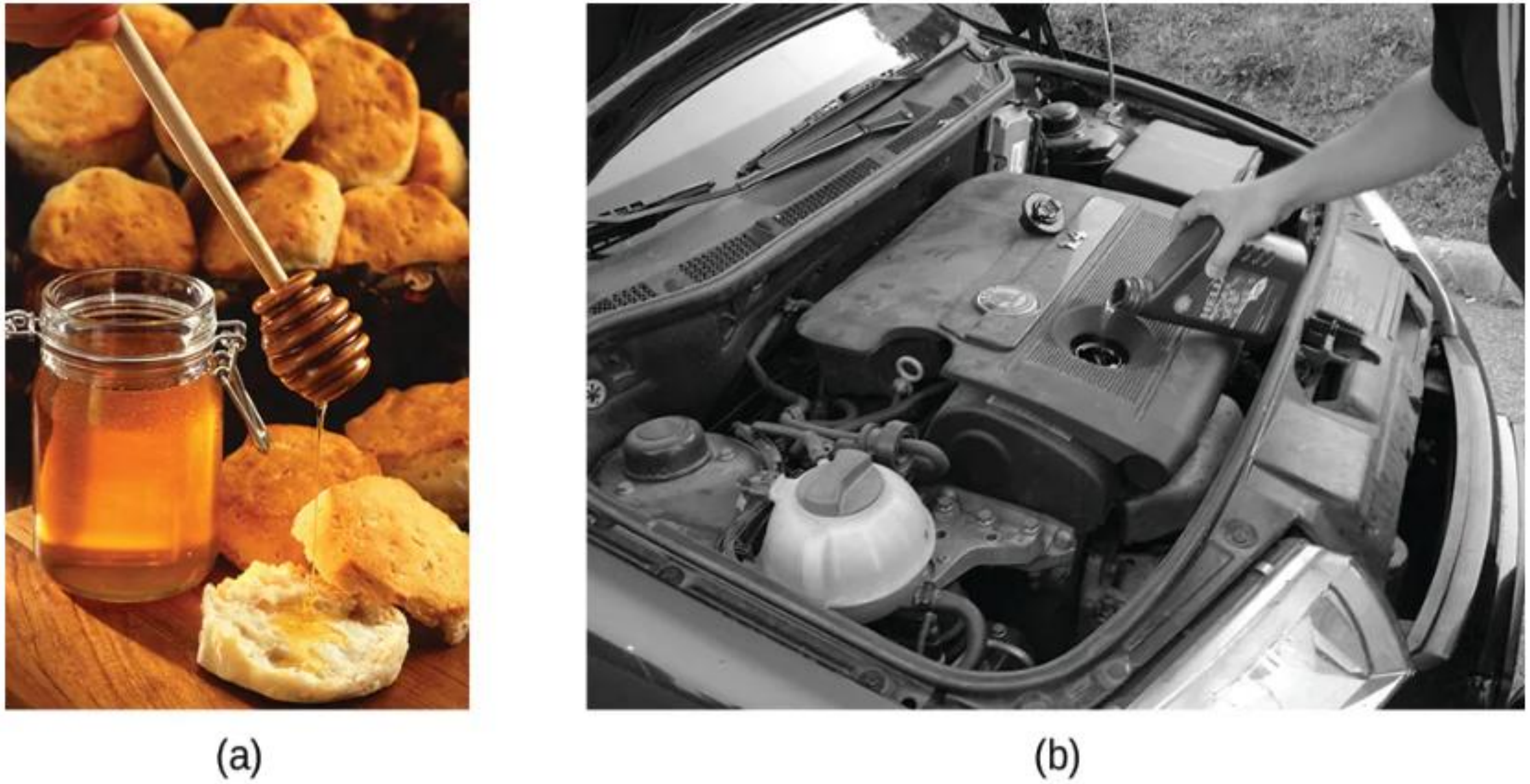


Figure 10.15 (a) Honey and (b) motor oil are examples of liquids with high viscosities; they flow slowly. (credit a: modification of work by Scott Bauer; credit b: modification of work by David Nagy)

Surface Tension – IMF cause the surface of a liquid to be as small as possible

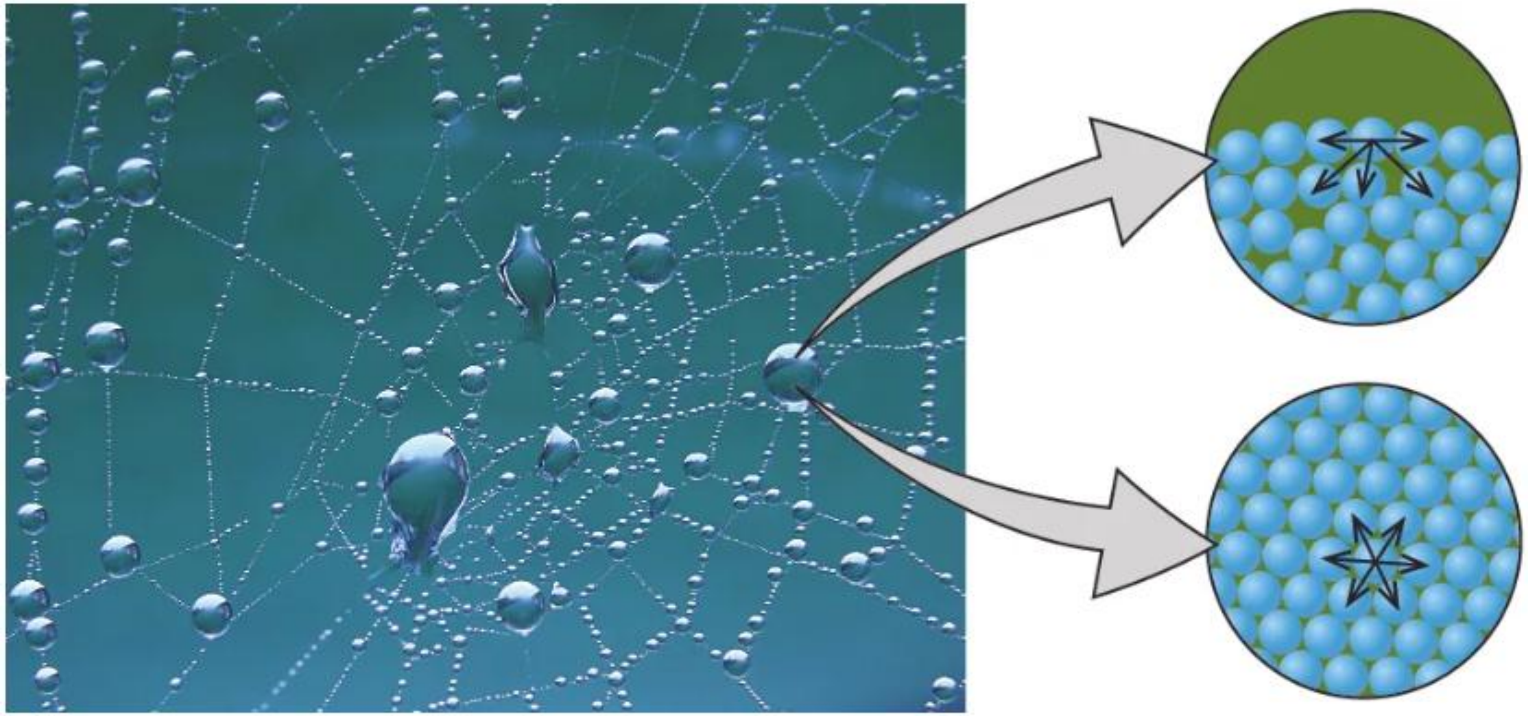


Figure 10.16 Attractive forces result in a spherical water drop that minimizes surface area; cohesive forces hold the sphere together; adhesive forces keep the drop attached to the web. (credit photo: modification of work by "OliBac"/Flickr)

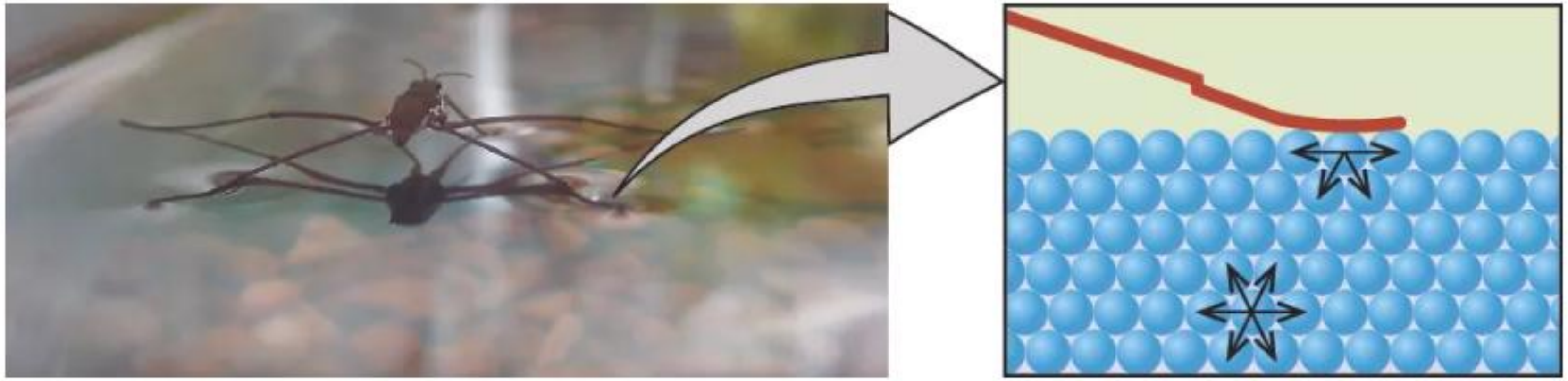


Figure 10.17 Surface tension (right) prevents this insect, a "water strider," from sinking into the water.

Meniscus is from IMF, the liquid is either attracted or repelled by the tube

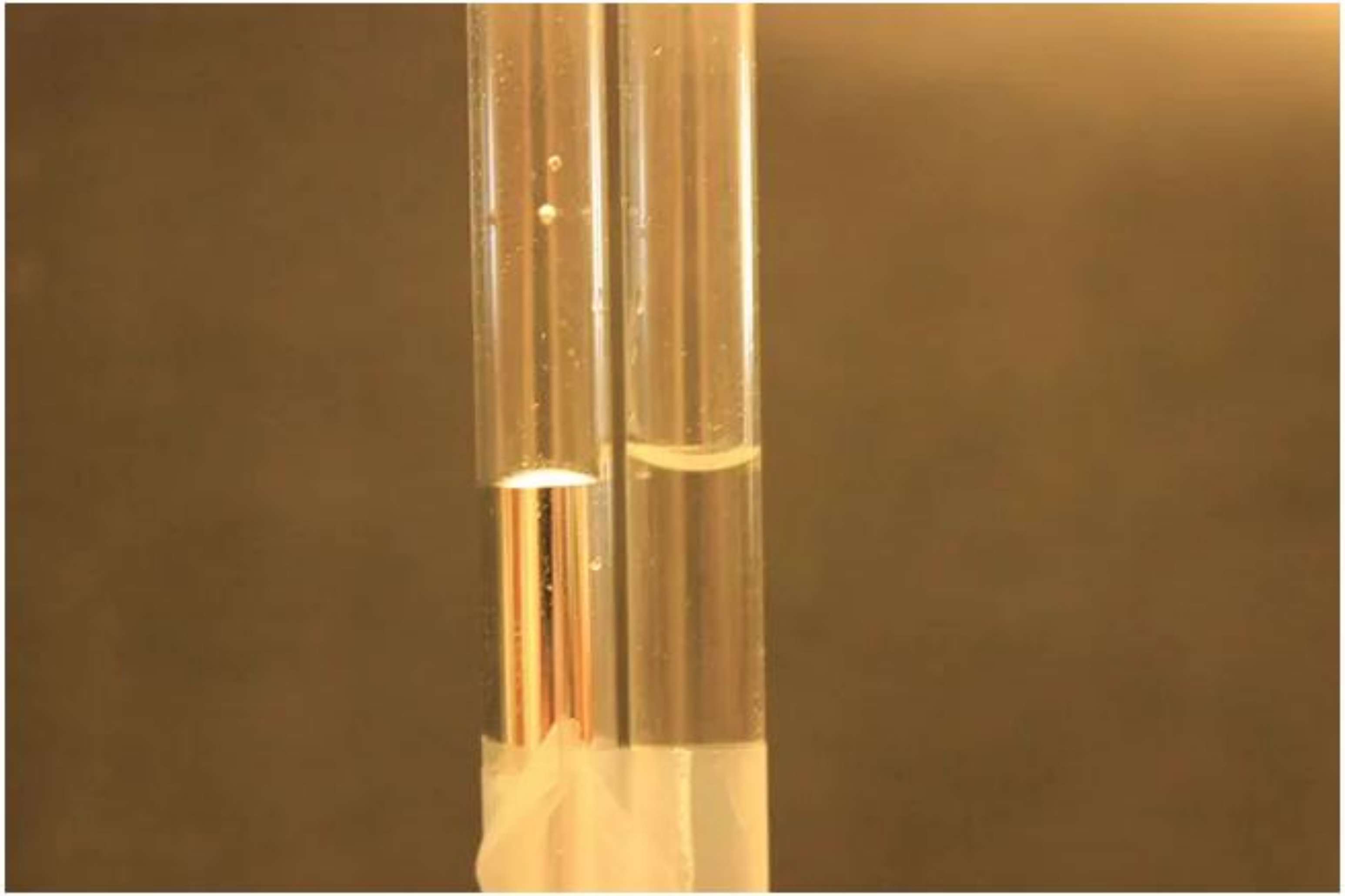


Figure 10.18 Differences in the relative strengths of cohesive and adhesive forces result in different meniscus shapes for mercury (left) and water (right) in glass tubes. (credit: Mark Ott)

Chapter 10.2

Capillary Action is from IMF, the liquid is attracted to the capillary tube

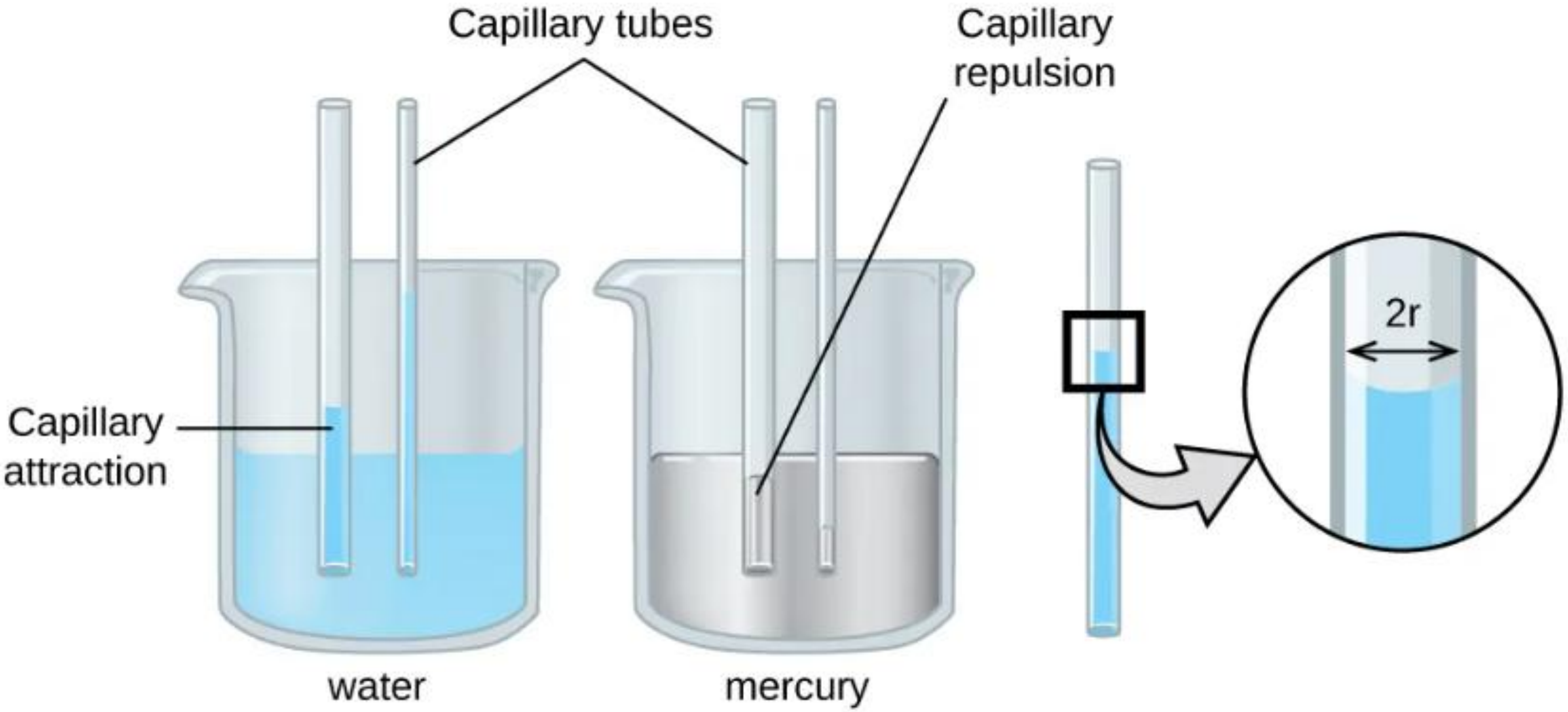


Figure 10.20 Depending upon the relative strengths of adhesive and cohesive forces, a liquid may rise (such as water) or fall (such as mercury) in a glass capillary tube. The extent of the rise (or fall) is directly proportional to the surface tension of the liquid and inversely proportional to the density of the liquid and the radius of the tube.

States of Matter and Energy

- Chapter 10.3 (Phase Transitions) – heating and cooling curves only
- Chapter 10.4 (Phase Diagrams)
- Chapter 9.1 (Energy Basics)

Adapted From: Paul Flowers, Edward J. Neth, William R. Robinson, PhD, Klaus Theopold, Richard Langley; Chemistry: Atoms First 2e; OpenStax; Feb 14, 2019; Houston Texas; https://openstax.org/books/chemistry-atoms-first-2e/pages/1-introduction

Intermolecular Forces govern different physical states

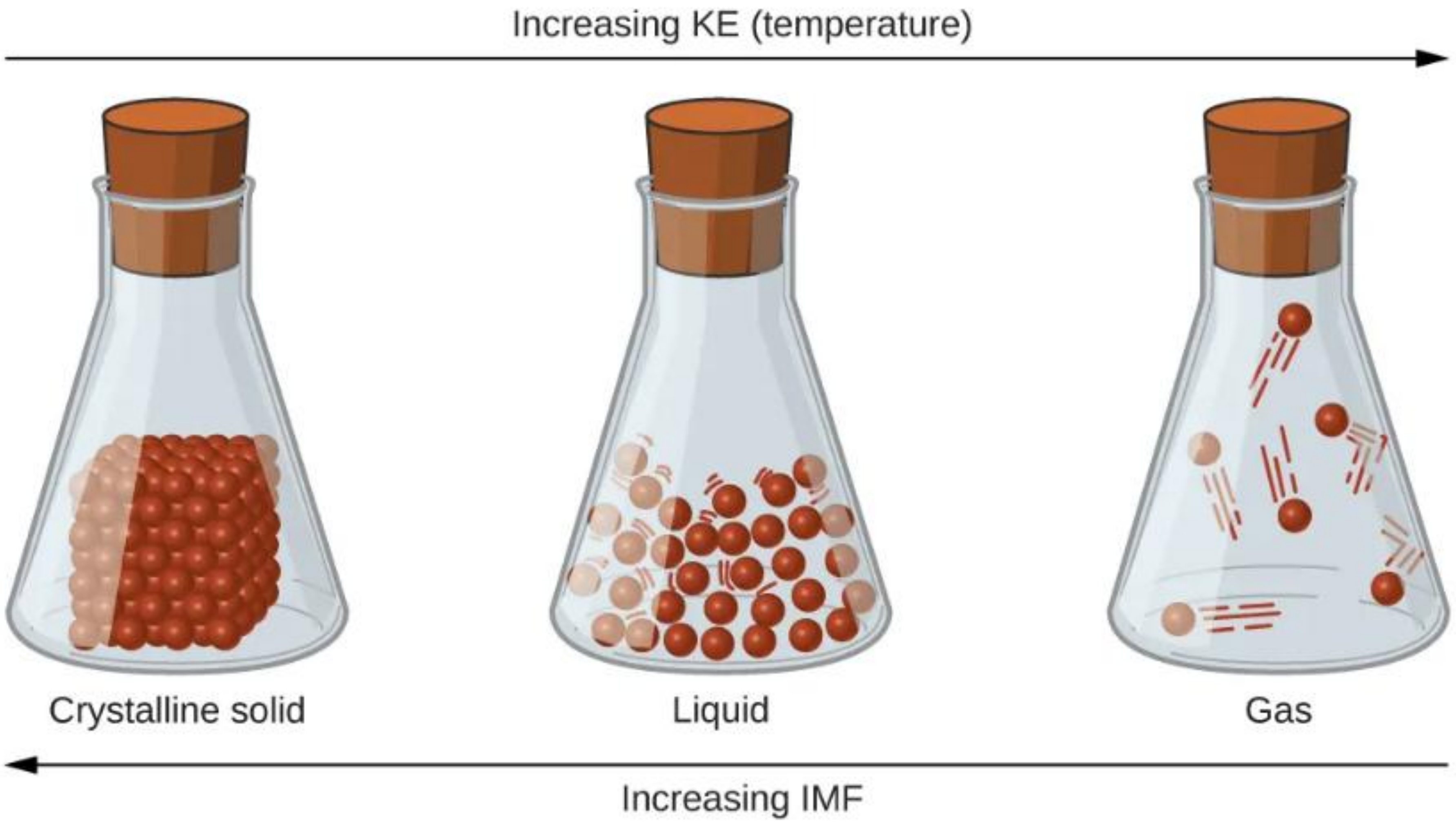
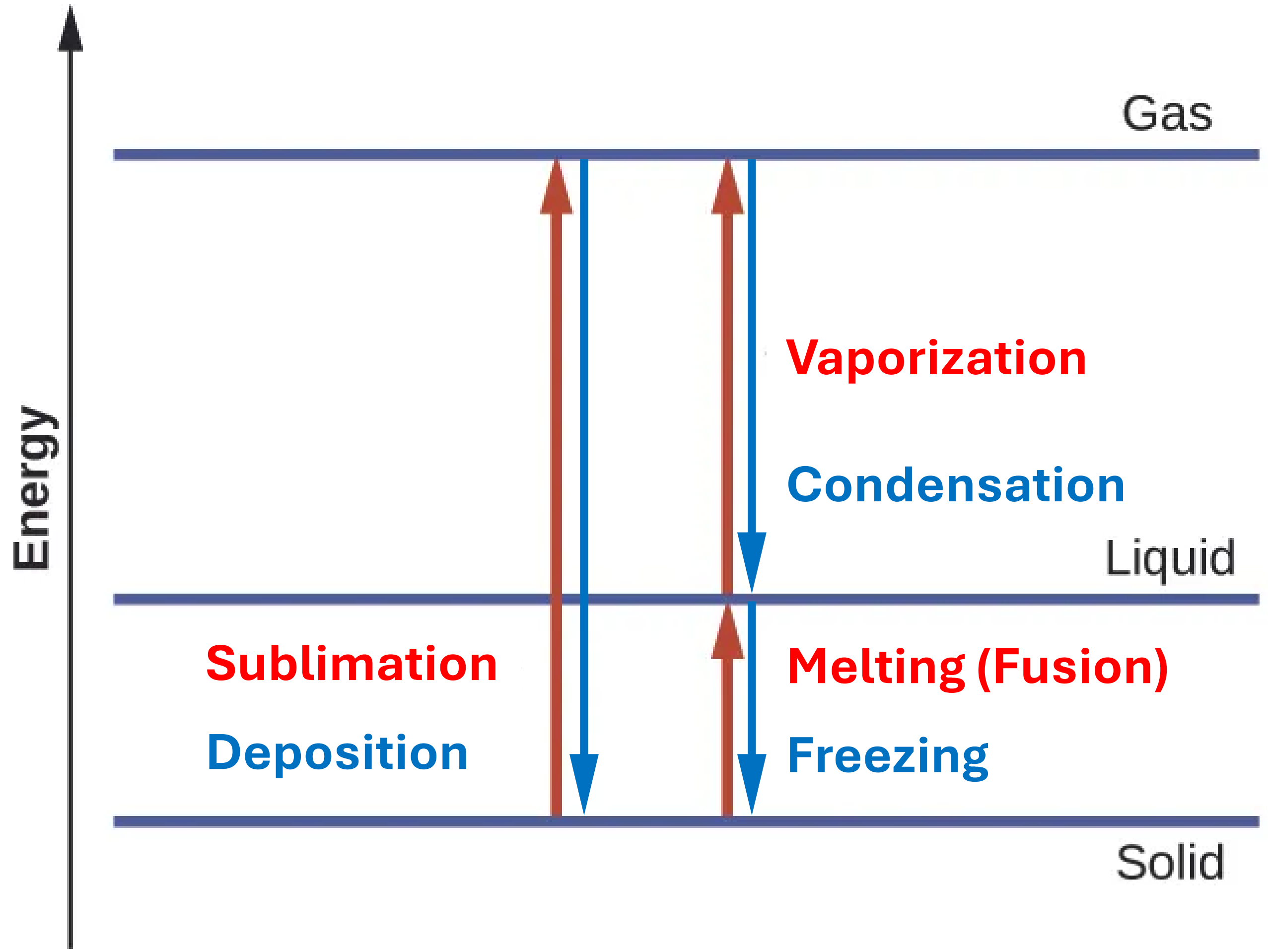


Figure 10.2 Transitions between solid, liquid, and gaseous states of a substance occur when conditions of temperature or pressure favor the associated changes in intermolecular forces. (Note: The space between particles in the gas phase is much greater than shown.)

Chapter 10.1

Energy can break intermolecular forces (IMF)



Vaporization

Condensation

Sublimation

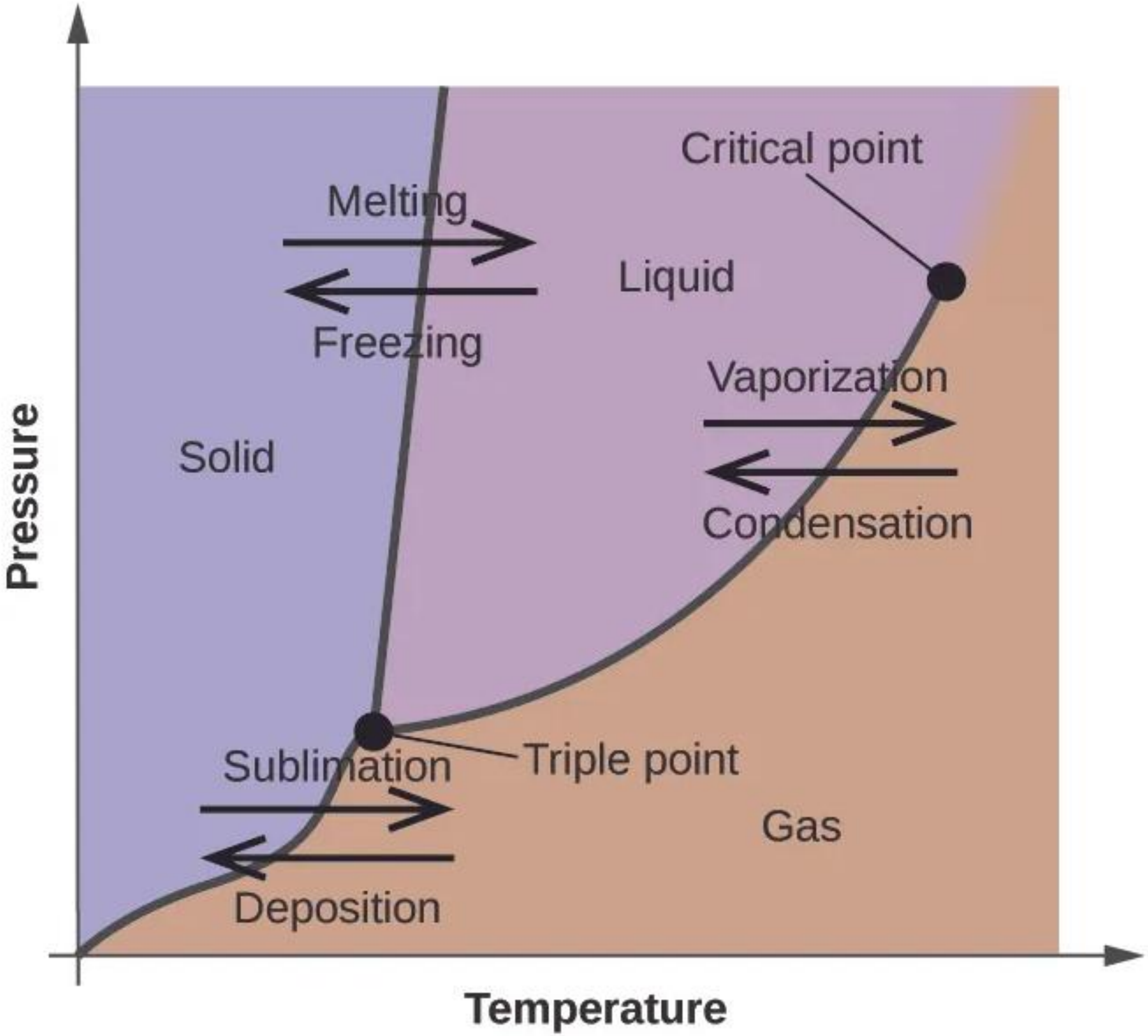
Deposition

Melting (Fusion)

Freezing

Figure 10.28 For a given substance, the sum of its enthalpy of fusion and enthalpy of vaporization is approximately equ
to its enthalpy of sublimation.

Pressure can also affect phase changes (and physical states)



Chapter 10.4

Figure 10.30 The physical state of a substance and its phase-transition temperatures are represented graphically in a phase diagram.

Phase diagrams can help identify the physical state of a substance

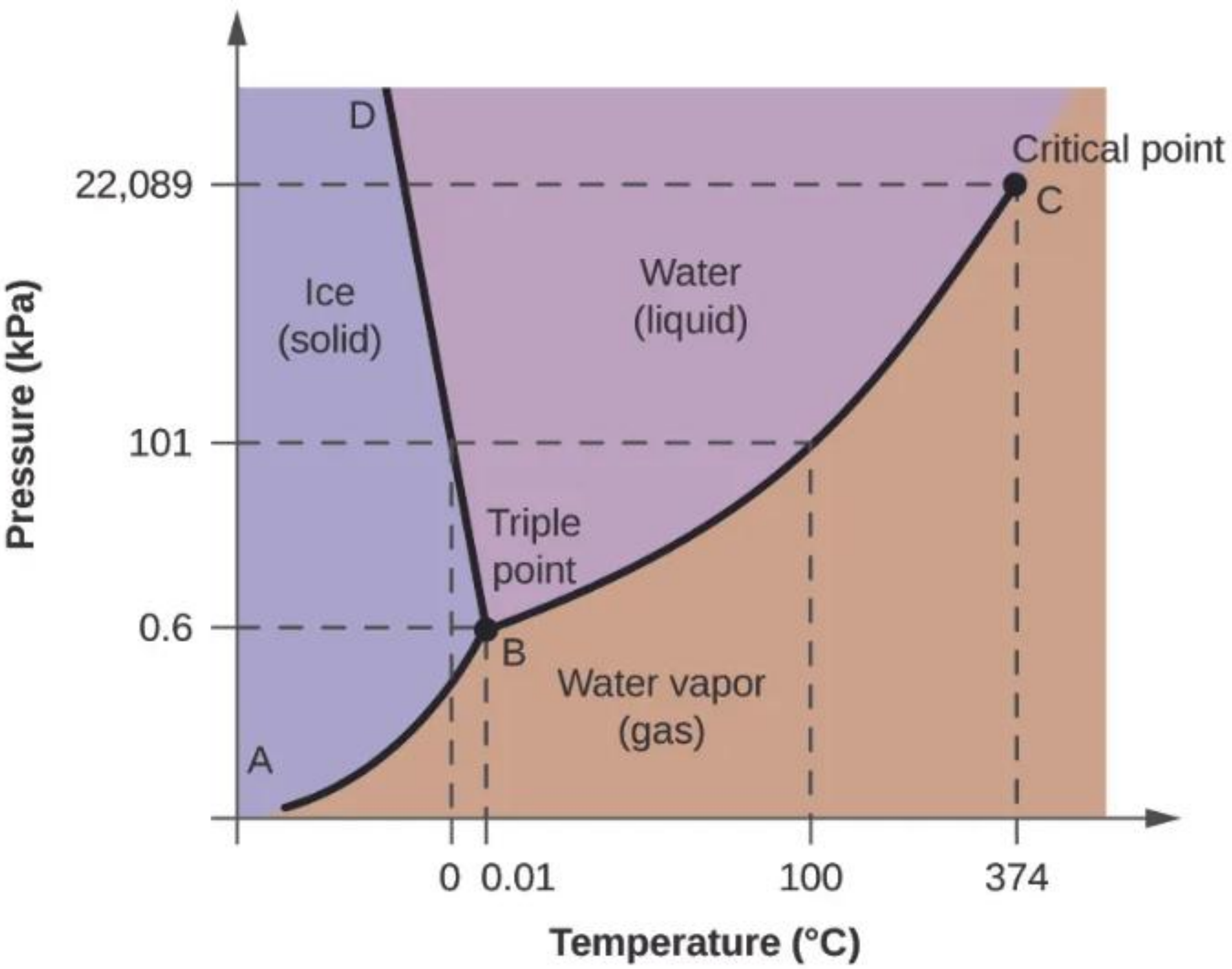


Figure 10.31 The pressure and temperature axes on this phase diagram of water are not drawn to constant scale in order to illustrate several important properties.

Check Your Learning

What phase changes can water undergo as the temperature changes if the pressure is held at 0.3 kPa? If the pressure is held at 50 kPa?

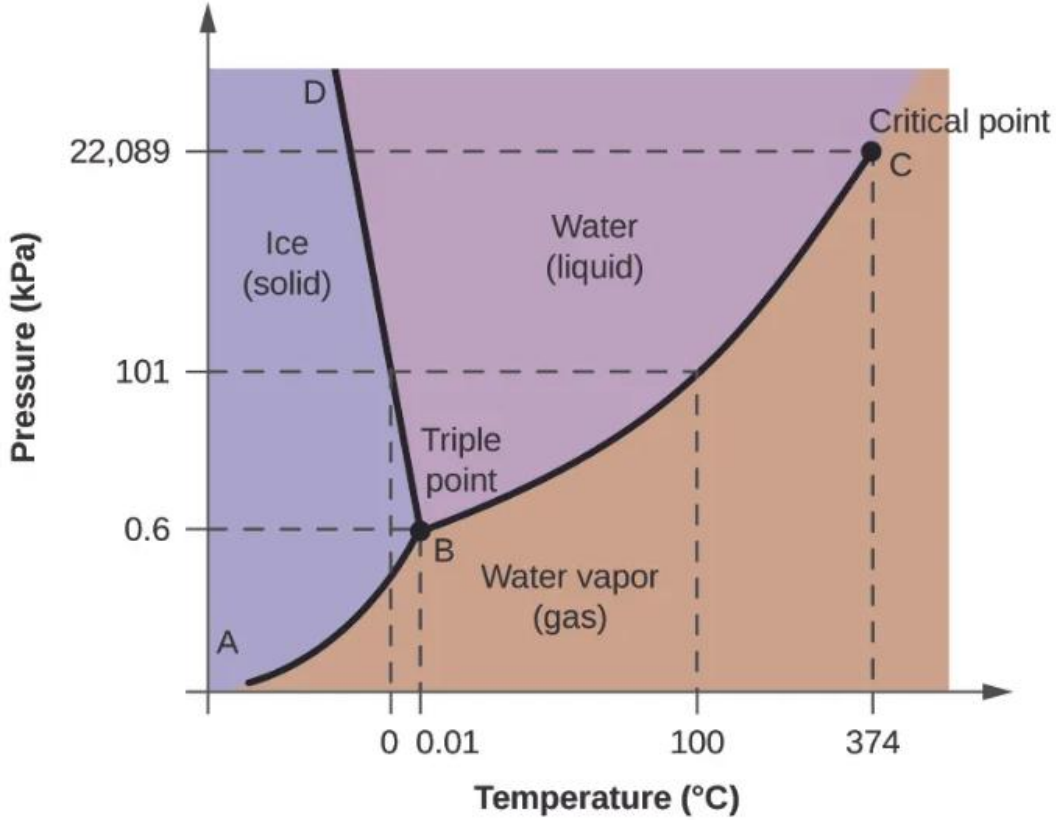


Figure 10.31 The pressure and temperature axes on this phase diagram of water are not drawn to constant scale in order to illustrate several important properties.

Chapter 10.4

Check Your Learning

Identify the phase changes that carbon dioxide will undergo as its temperature is increased from −100 °C while holding its pressure constant at 1500 kPa. At 50 kPa. At what approximate temperatures do these phase changes occur?

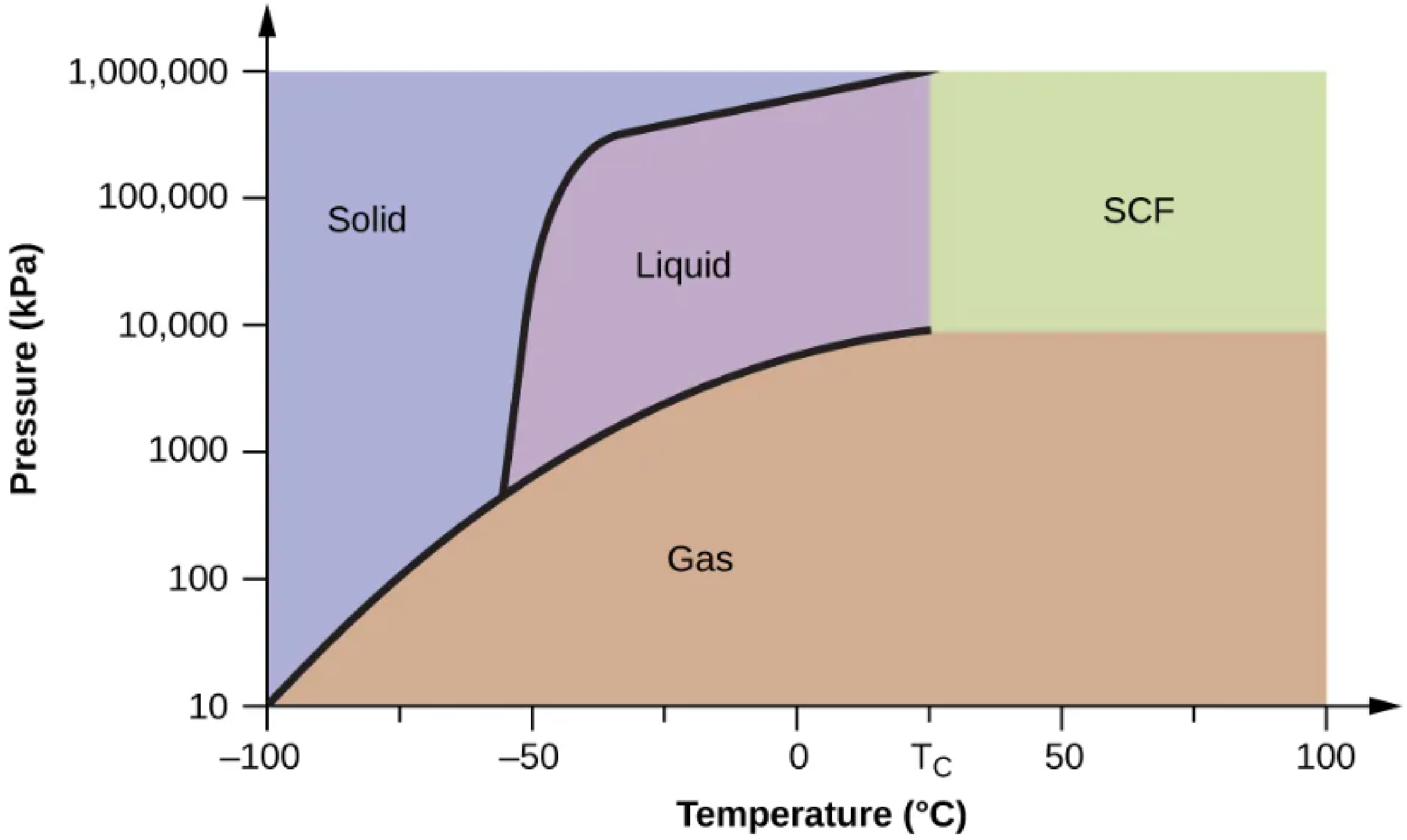


Figure 10.34 A phase diagram for carbon dioxide is shown. The pressure axis is plotted on a logarithmic scale to accommodate the large range of values.

Heating and Cooling Curves

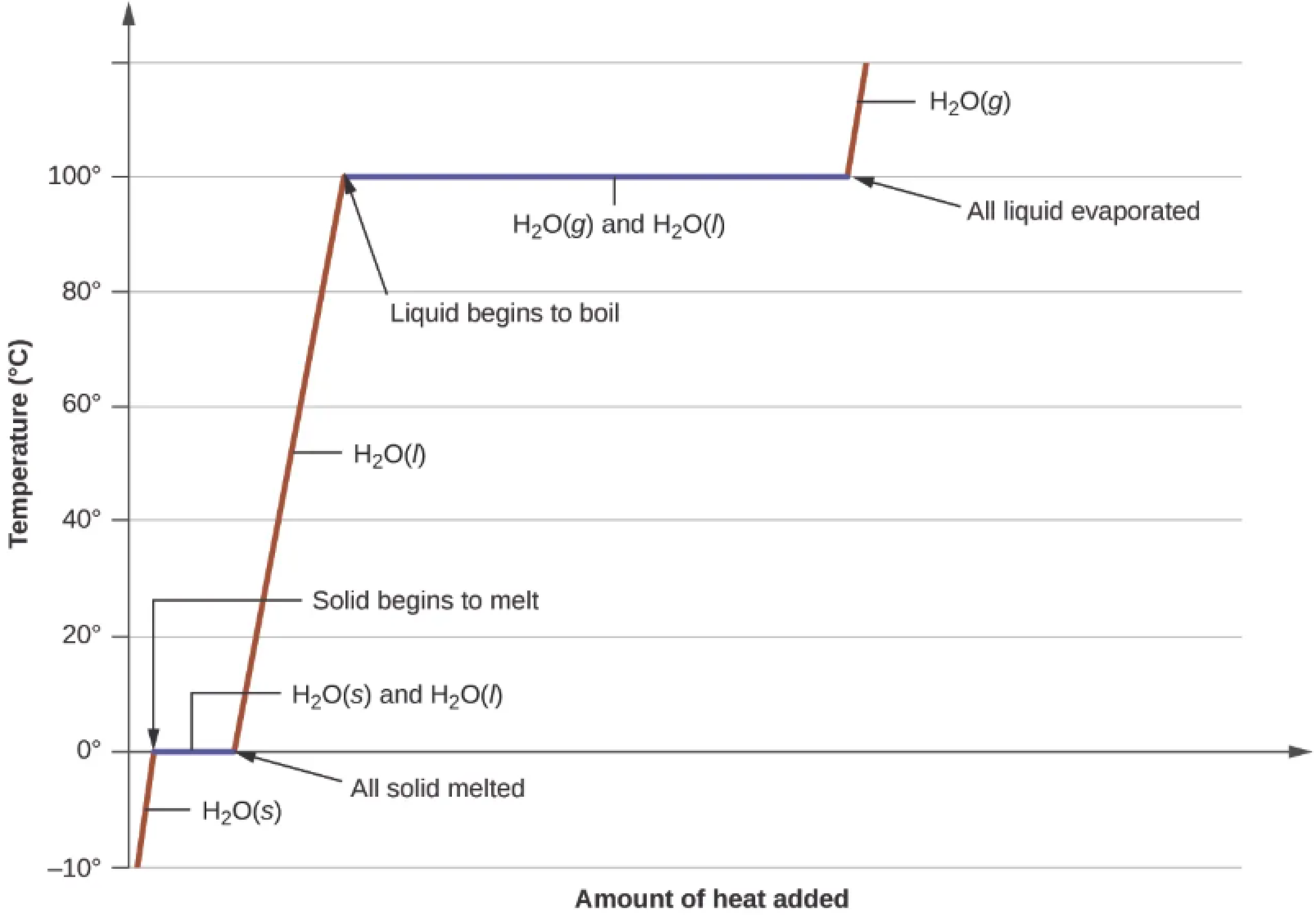


Figure 10.29 A typical heating curve for a substance depicts changes in temperature that result as the substance absorbs increasing amounts of heat. Plateaus in the curve (regions of constant temperature) are exhibited when the substance undergoes phase transitions.

Energy Key Terms

energy
capacity to supply heat or do work

heat (*q*)
transfer of thermal energy between two bodies

temperature
intensive property of matter that is a quantitative measure of “hotness” and “coldness”

intensive property
property of a substance that is independent of the amount of the substance

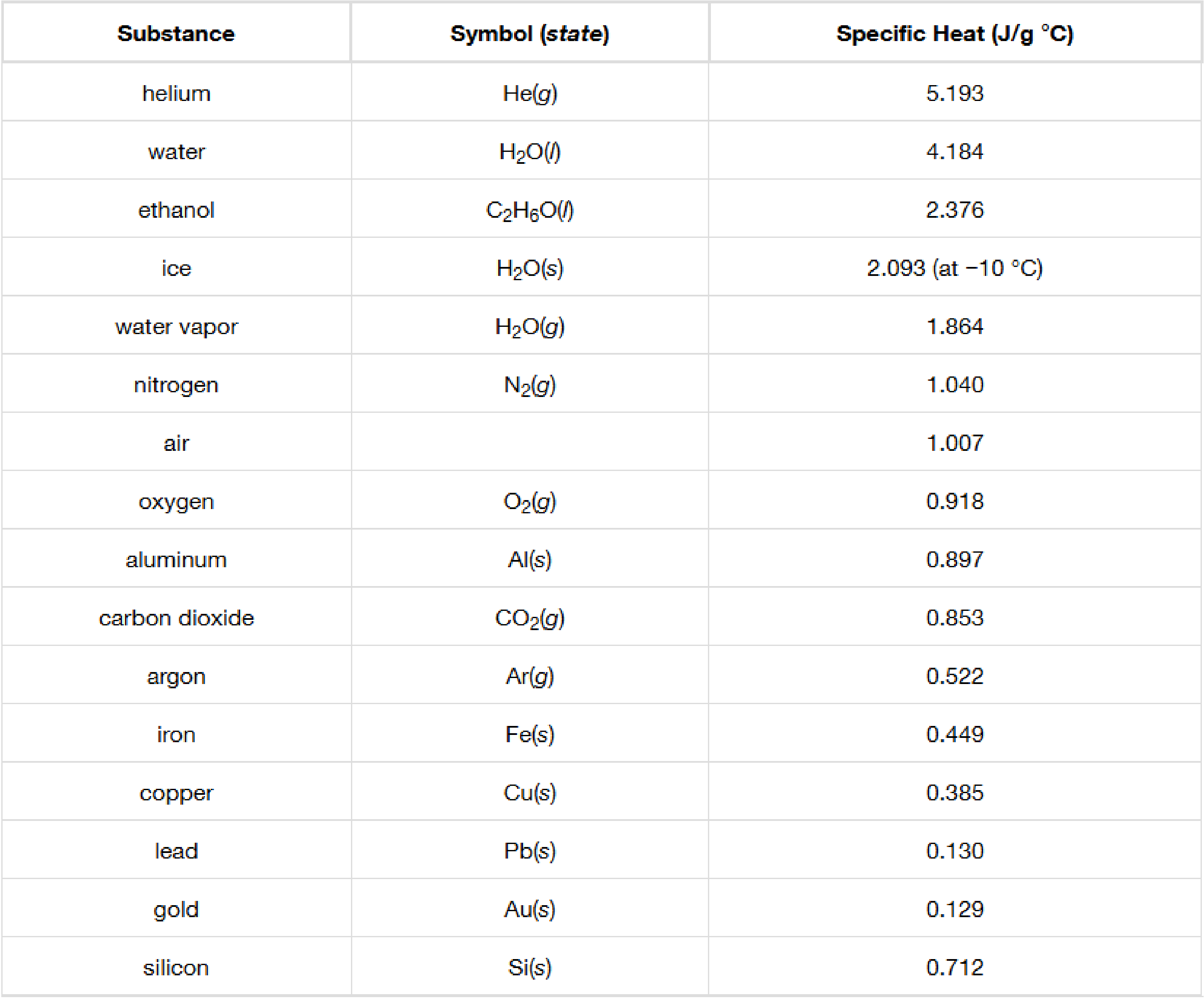
joule (J)
SI unit of energy; amount of energy used when a force of 1 newton moves an object 1 meter, 1 J = 1 kg m^2/s^2 and 4.184 J = 1 cal

specific heat capacity (*c*)
intensive property of a substance that represents the quantity of heat required to raise the temperature of 1 gram of the substance by 1 degree Celsius (or 1 kelvin)

Different substances have different heat capacities

$$q = (\text{specific heat}) \times (\text{mass of substance}) \times (\text{temperature change})$$
$$q = c \times m \times \Delta T = c \times m \times (T_{\text{final}} - T_{\text{initial}})$$

Specific Heats of Common Substances at 25 °C and 1 bar



Substance	Symbol (*state*)	Specific Heat (J/g °C)
helium	$He(g)$	5.193
water	$H_2O(l)$	4.184
ethanol	$C_2H_6O(l)$	2.376
ice	$H_2O(s)$	2.093 (at −10 °C)
water vapor	$H_2O(g)$	1.864
nitrogen	$N_2(g)$	1.040
air		1.007
oxygen	$O_2(g)$	0.918
aluminum	$Al(s)$	0.897
carbon dioxide	$CO_2(g)$	0.853
argon	$Ar(g)$	0.522
iron	$Fe(s)$	0.449
copper	$Cu(s)$	0.385
lead	$Pb(s)$	0.130
gold	$Au(s)$	0.129
silicon	$Si(s)$	0.712

Table 9.1

Exothermic processes release heat while endothermic reactions absorb heat.

Figure 9.7 (a) An oxyacetylene torch produces heat by the combustion of acetylene in oxygen. The energy released by this exothermic reaction heats and then melts the metal being cut. The sparks are tiny bits of the molten metal flying away. (b) A cold pack uses an endothermic process to create the sensation of cold. (credit a: modification of work by "Skatebiker"/Wikimedia commons)

Heating and Cooling Curves

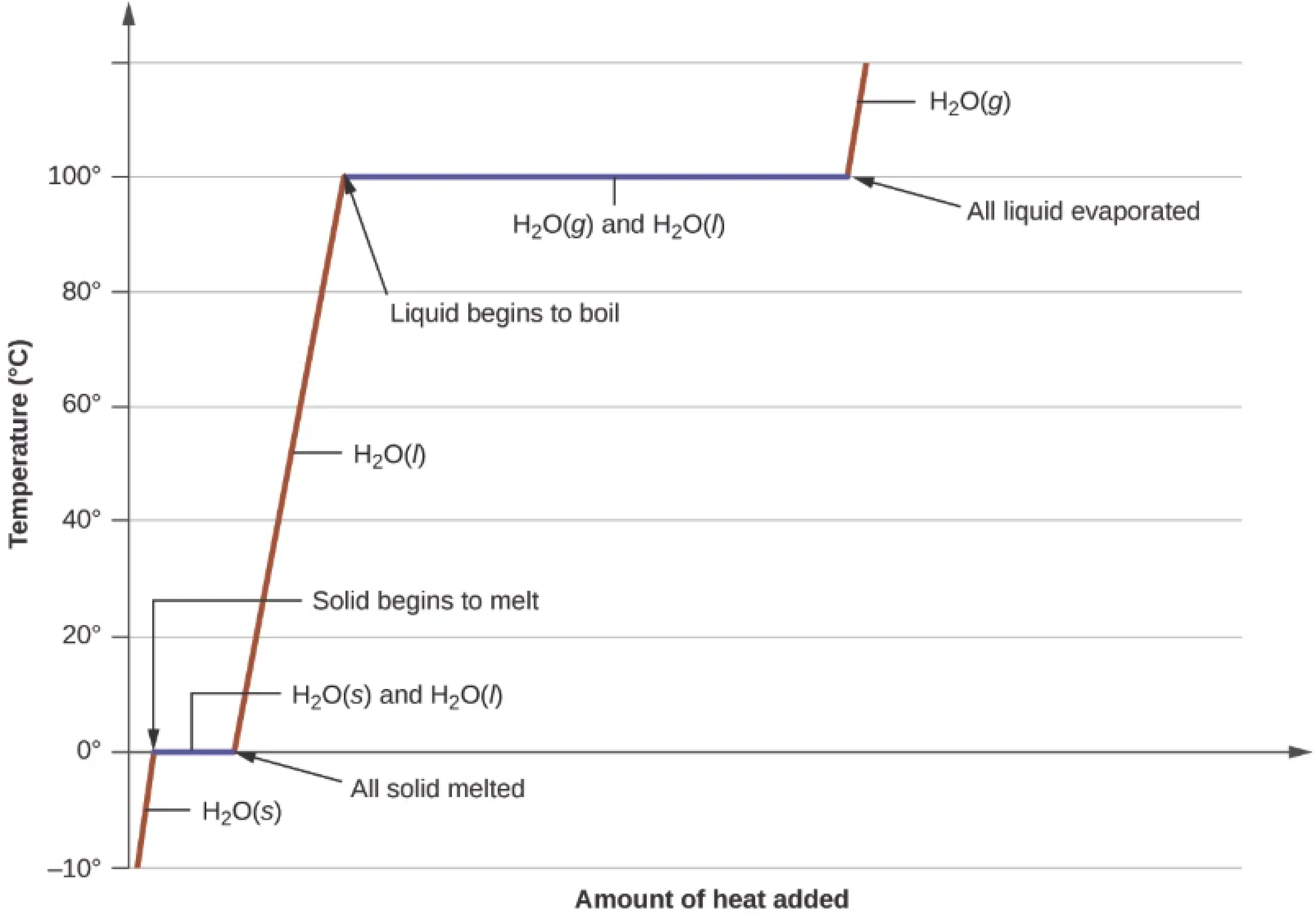
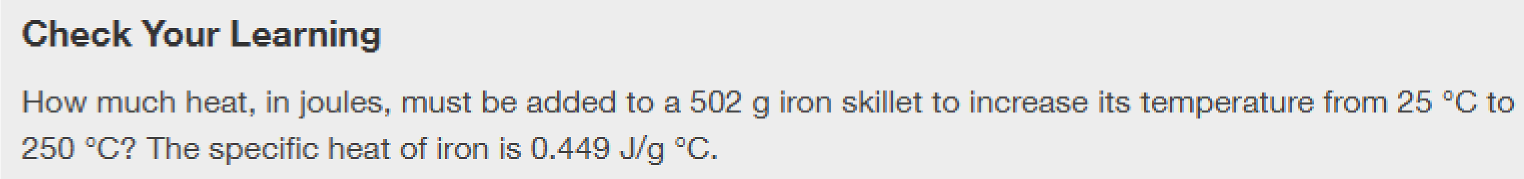


Figure 10.29 A typical heating curve for a substance depicts changes in temperature that result as the substance absorbs increasing amounts of heat. Plateaus in the curve (regions of constant temperature) are exhibited when the substance undergoes phase transitions.

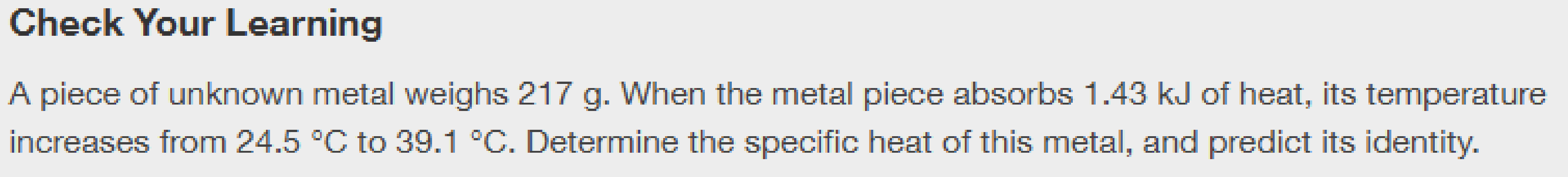


Check Your Learning

How much heat, in joules, must be added to a 502 g iron skillet to increase its temperature from 25 °C to 250 °C? The specific heat of iron is 0.449 J/g °C.

$$q = \text{(specific heat)} \times \text{(mass of substance)} \times \text{(temperature change)}$$

$$q = c \times m \times \Delta T = c \times m \times (T_{\text{final}} - T_{\text{initial}})$$

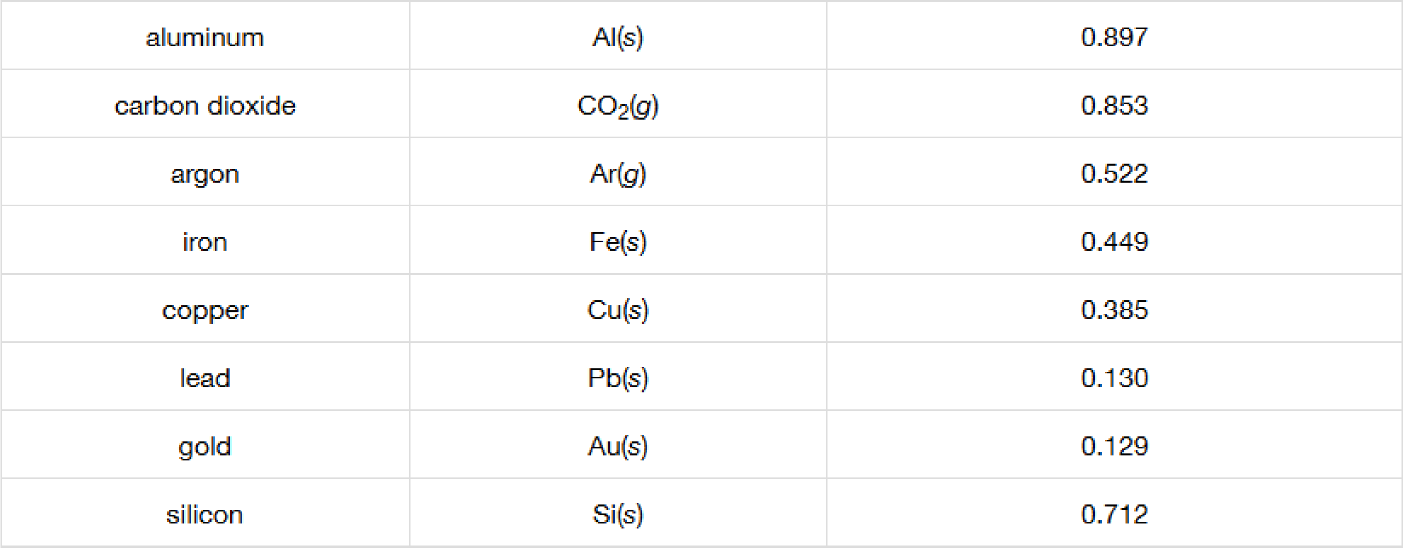


Check Your Learning

A piece of unknown metal weighs 217 g. When the metal piece absorbs 1.43 kJ of heat, its temperature increases from 24.5 °C to 39.1 °C. Determine the specific heat of this metal, and predict its identity.

$$q = (\text{specific heat}) \times (\text{mass of substance}) \times (\text{temperature change})$$

$$q = c \times m \times \Delta T = c \times m \times (T_{\text{final}} - T_{\text{initial}})$$



aluminum	Al(*s*)	0.897
carbon dioxide	CO_2(*g*)	0.853
argon	Ar(*g*)	0.522
iron	Fe(*s*)	0.449
copper	Cu(*s*)	0.385
lead	Pb(*s*)	0.130
gold	Au(*s*)	0.129
silicon	Si(*s*)	0.712

Table 9.1

Chapter 9.1