Ch. 11
States of matter

States of Matter

Solid

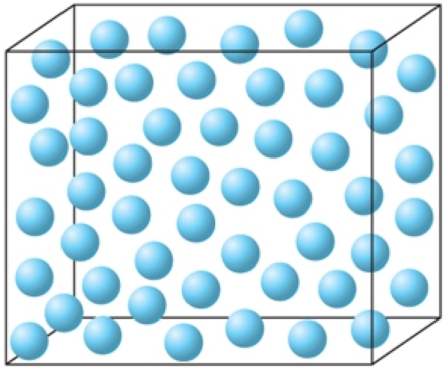
- Definite volume
- Definite shape

Liquid

- Definite volume
- Indefinite shape (conforms to container)

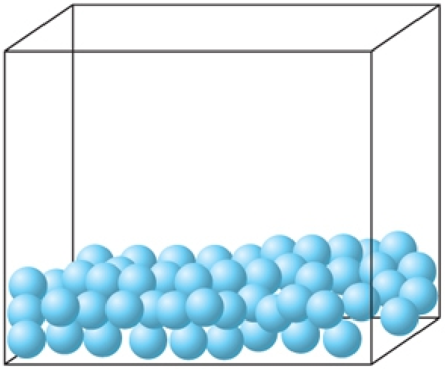
Gas

- Indefinite volume (fills any container)
- Indefinite shape (conforms to container)



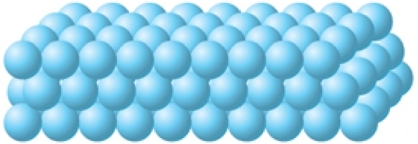
Gas

Molecules far apart and disordered
Negligible interactions among molecules



Liquid

Molecules close together and disordered
Moderately strong interactions among molecules



Solid

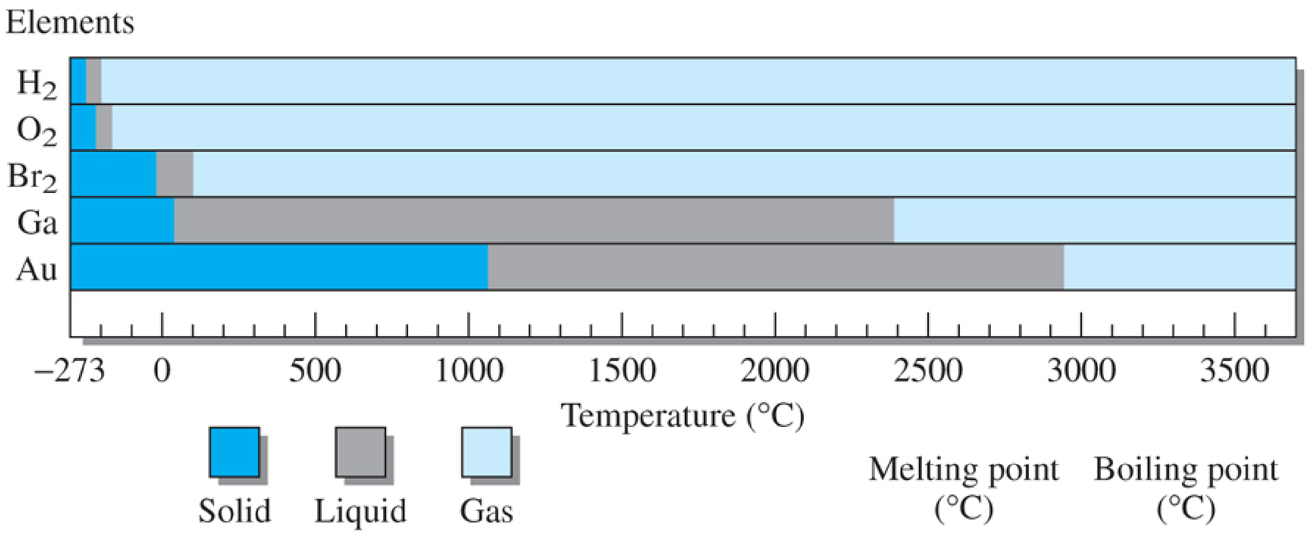
Molecules close together and ordered
Strong interactions among molecules

Solids and Liquids

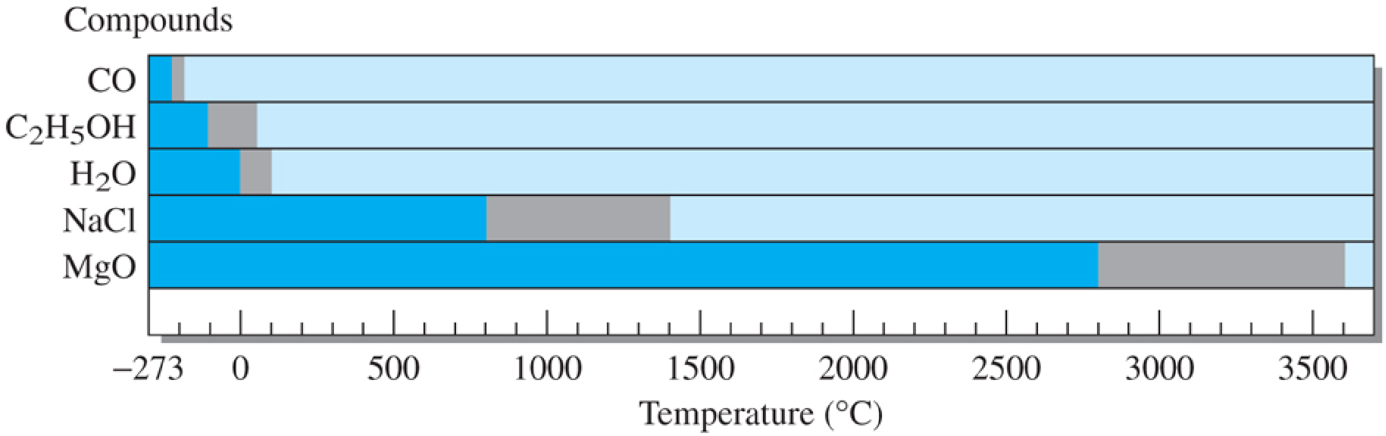
- High density – solid usually slightly more dense than the corresponding liquid
- Small compressibility
- Small thermal expansion

Gases

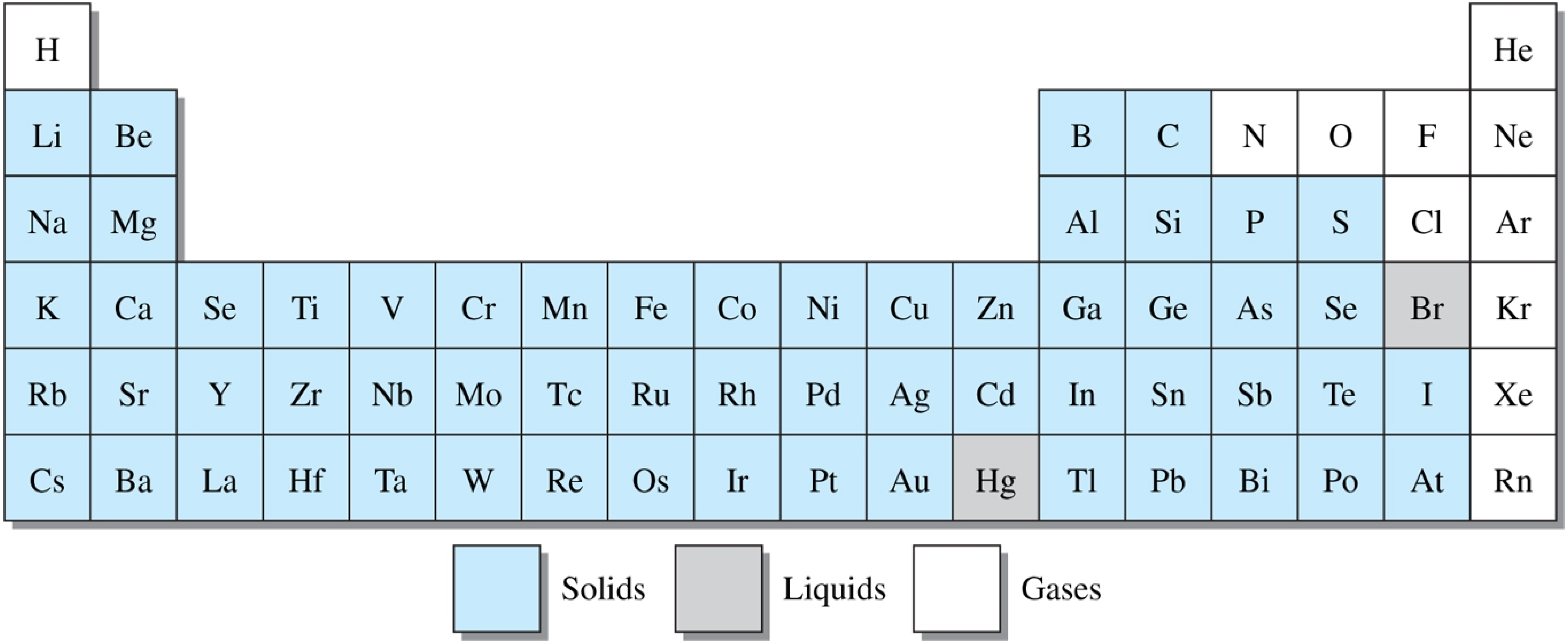
- Low density
- Large compressibility
- Large thermal expansion



	Melting point (°C)	Boiling point (°C)	Liquid range (°C)
H_2	–259	–252	7
O_2	–218	–183	35
Br_2	–7	59	66
Ga	30	2403	2373
Au	1063	2966	1903



	Melting point (°C)	Boiling point (°C)	Liquid range (°C)
CO	–199	–192	7
C_2H_5OH	–115	78	193
H_2O	0	100	100
NaCl	801	1413	612
MgO	2800	3600	800



H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Se	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn

🟦 Solids ▨ Liquids ☐ Gases

Kinetic Molecular Theory

1)Matter is composed of particles that have definite sizes
2)The particles are in constant motion (have kinetic energy
3)Particles interact through attraction and repulsion, so have potential energy
4)The kinetic energy of particles increases with the temperature
5)Particles transfer energy to one another through elastic collisions

Changes of State

When matter changes from one phase to another:

- Solid ↔ liquid
  - melting or fusion (s → l)
  - freezing (l → s)

- Liquid ↔ gas
  - evaporation (l → g)
  - condensation (g → l)

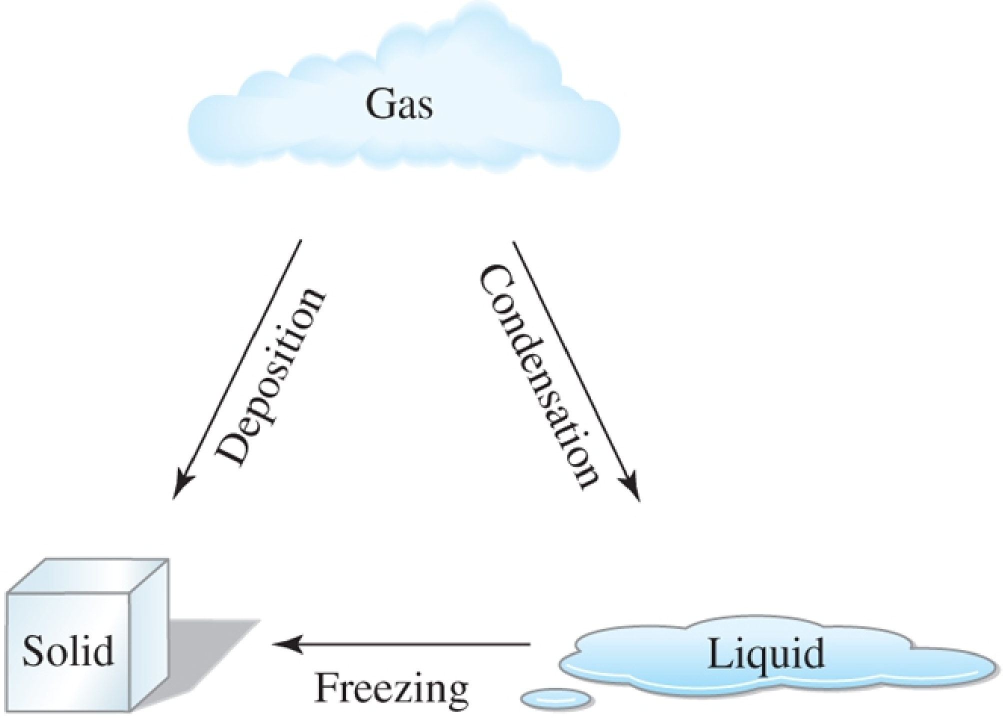
- Solid ↔ gas
  - sublimation (s → g)
  - deposition (g → s)

Endothermic process – an event in which energy is absorbed by the system

- melting, evaporation, sublimation

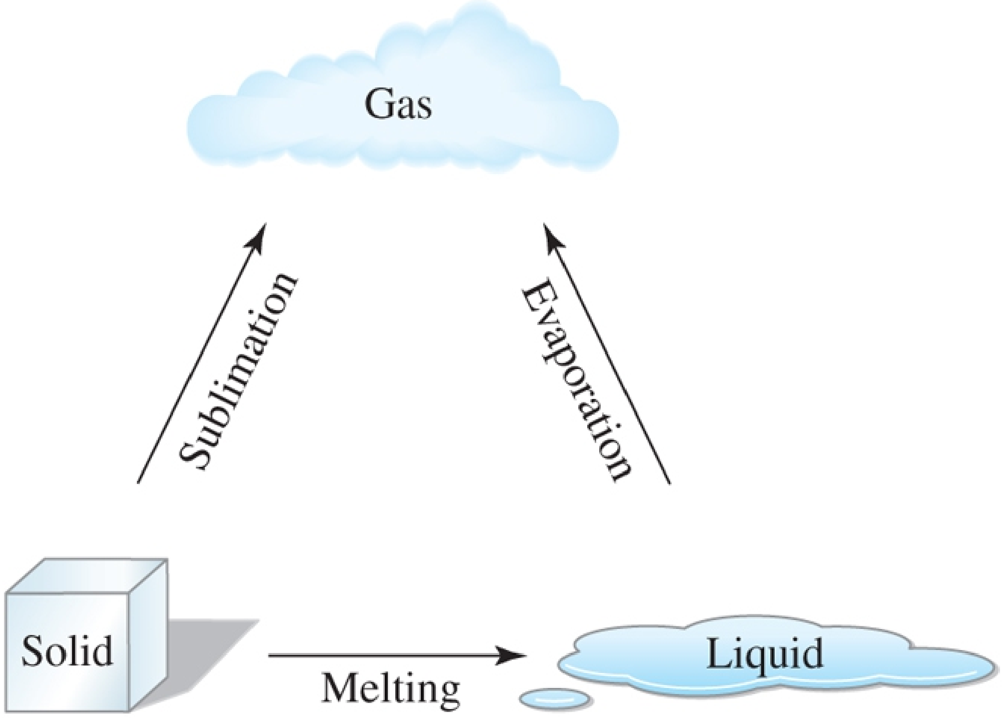
Exothermic process – an event in which energy is released by the system

- freezing, condensation, deposition



Exothermic changes of state

(a)



Endothermic changes of state

(b)

Energy – The capacity to do work

- kinetic
- potential
- chemical
- nuclear
- thermal

Energy is measured in:

- Joules (J)
- calories (cal): 1 cal = 4.18 J
- Calories, or food calories (Cal):
  1 Cal = 1000 cal = 1 kcal

Specific Heat – relates the amount of energy that is needed to raise 1 g of a substance (in a given state) by 1 °C

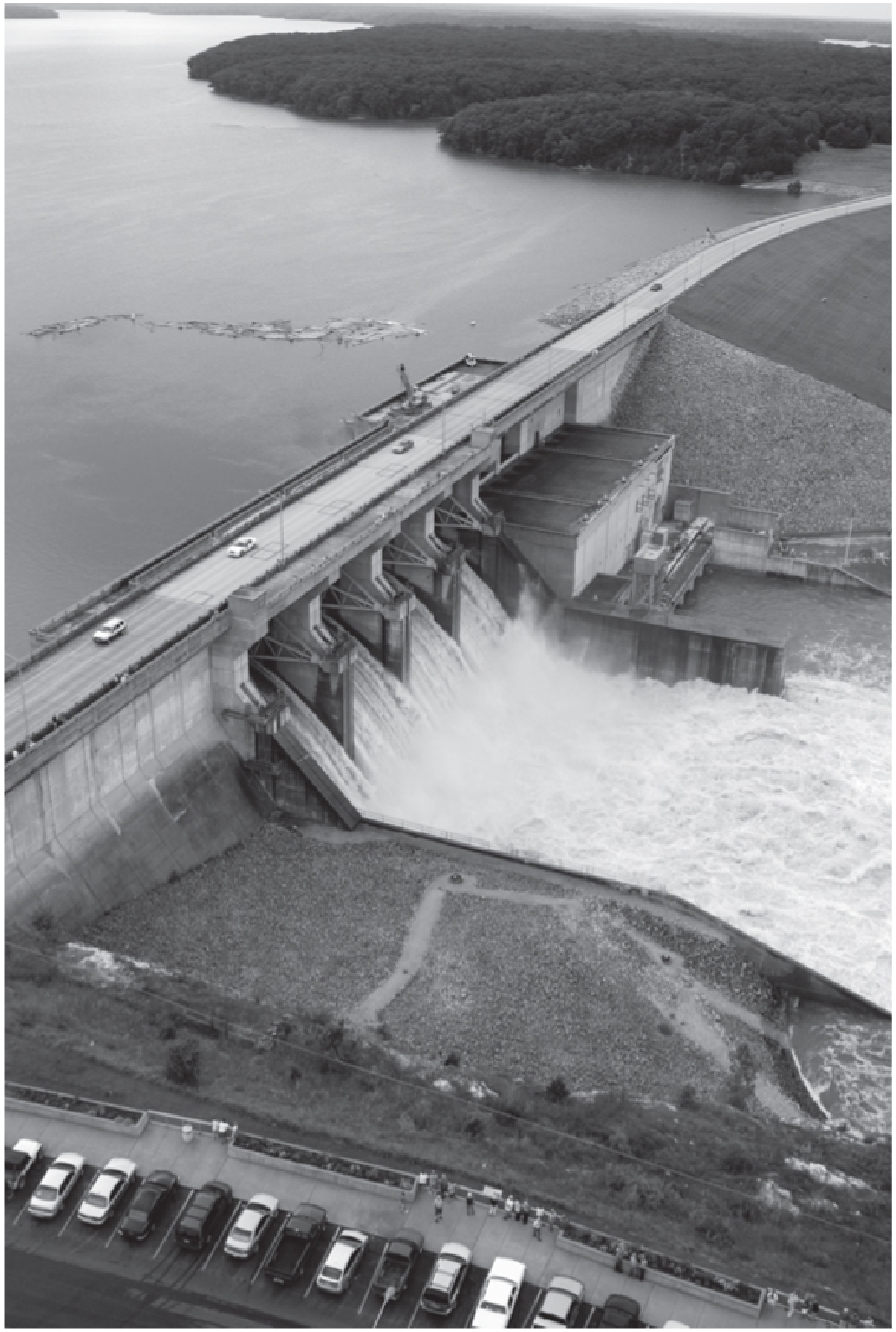
- Larger specific heat – more energy needed to raise the temperature of a sample of the matter

- Each phase of matter has a different specific heat

- The units for specific heat are typically:
  - J/(g °C)
  - or J/(mol °C)

TABLE 11.2 Specific Heats of Selected Pure Substances

Substance	Physical State	Specific Heat [J/(g · °C)]
Aluminum	solid	0.908
Copper	solid	0.382
Ethyl alcohol	liquid	2.42
Gold	solid	0.13
Iron	solid	0.444
Nitrogen	gas	1.0
Oxygen	gas	0.92
Silver	solid	0.24
Sodium chloride	solid	0.88
Water (ice)	solid	2.09
Water	liquid	4.18
Water (steam)	gas	2.03
Zinc	solid	0.388



For a sample of matter, the thermal energy change is related to the temperature change by:

$$q = m\ C\ \Delta T$$

where:
q = thermal energy change
m = mass of the sample
C = specific heat
ΔT = final – initial temperatures

If $q > 0$, then the object absorbs energy: endothermic.

If $q < 0$, then the object releases energy: exothermic.

- If two objects are in contact, thermal energy is transferred from the warm object to the cold object.
- The objects will eventually reach thermal equilibrium and be at the same temperature
- Ideally, the total energy lost by the warm object exactly equals the total energy gained by the cool object.

Specific Heat vs Heat Capacity

Specific Heat

- Relates thermal energy change to temperature change per gram/mole of matter
- Different values for different phases/substances

Heat Capacity

- Relates thermal energy change to temperature change for a specific object
- Depends on specific heat of the substance and the amount of substance present
- Heat capacity = specific heat * mass

Changes of State

When matter changes from one phase to another:

- Energy is gained or lost

- The temperature remains constant until all of the matter is converted to one phase

- The energy is stored in kinetic energy of the atoms or in intermolecular forces

To calculate the thermal energy needed for phase changes, another thermodynamic constant is needed.

This value depends on the material and on the phase change:

- Heat of fusion = ΔH_{fus}
  - Heat of freezing = $-\Delta H_{fus}$
- Heat of vaporization = ΔH_{vap}
  - Heat of condensation = $-\Delta H_{vap}$
- Heat of sublimation = ΔH_{sub}
  - Heat of deposition = $-\Delta H_{sub}$

The units for these values are either J/g or J/mol.

For ΔH in J/g, the energy is related by the equation:

$$q = m\,\Delta H$$

For ΔH in J/mol, the energy is related by the equation:

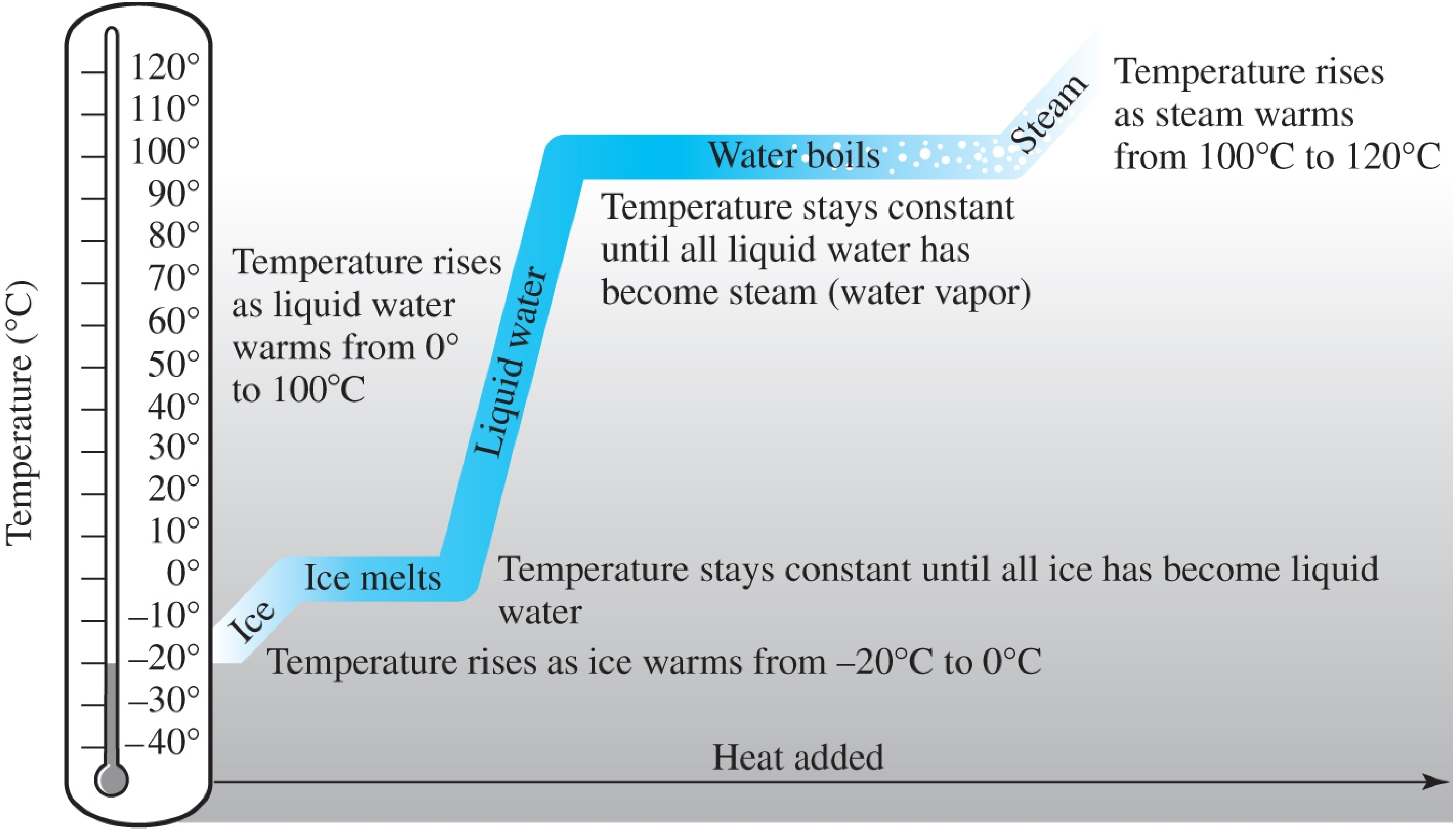
$$q = n\,\Delta H \quad \text{where } n = \text{moles}$$

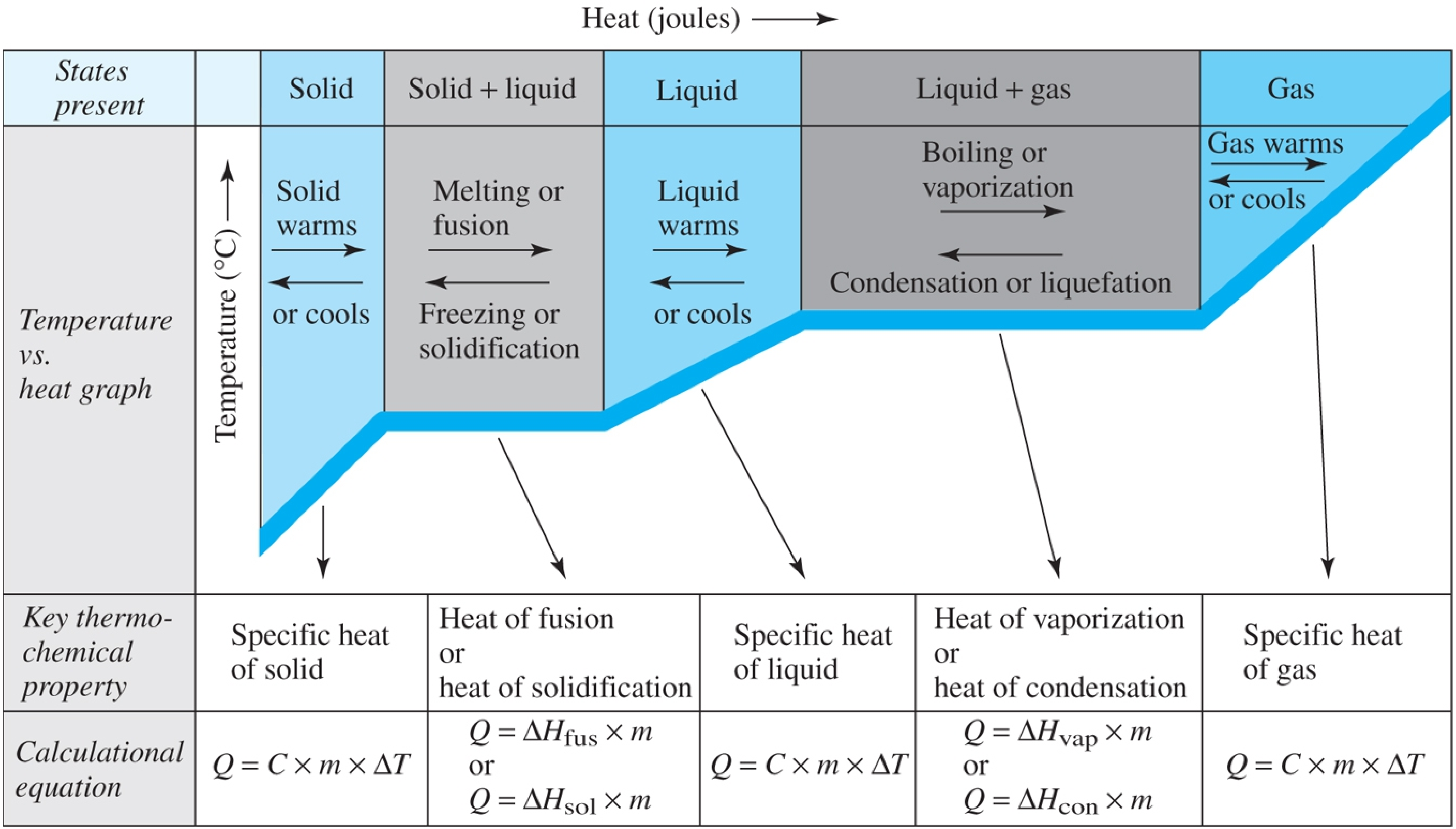
TABLE 11.3 Heats of Fusion for Various Substances at Their Melting Points

Solid	Melting Point (°C)	Heat of Fusion	
		J/g	kJ/mole
Methane	–182	59	0.94
Ethyl alcohol	–117	109	5.01
Carbon tetrachloride	–23	16.3	2.51
Water	0	334	6.01
Benzene	6	126	9.87
Aluminum	658	393	10.6
Copper	1083	205	13.0

TABLE 11.4 Heats of Vaporization for Various Substances at Their Boiling Points

		Heat of Vaporization	
Liquid	**Boiling Point (°C)**	**J/g**	**kJ/mole**
Methane	–161	648	10.4
Ammonia	–33	1380	23.4
Diethyl ether	34.6	375	27.8
Carbon tetrachloride	77	195	30.0
Ethyl alcohol	78.3	837	38.6
Benzene	80.1	394	30.8
Water	100	2260	40.7


Temperature (°C)
120°
110°
100°
90°
80°
70°
60°
50°
40°
30°
20°
10°
0°
–10°
–20°
–30°
–40°
Ice
Ice melts
Liquid water
Water boils
Steam
Temperature rises as steam warms from 100°C to 120°C
Temperature stays constant until all liquid water has become steam (water vapor)
Temperature rises as liquid water warms from 0° to 100°C
Temperature stays constant until all ice has become liquid water
Temperature rises as ice warms from –20°C to 0°C
Heat added



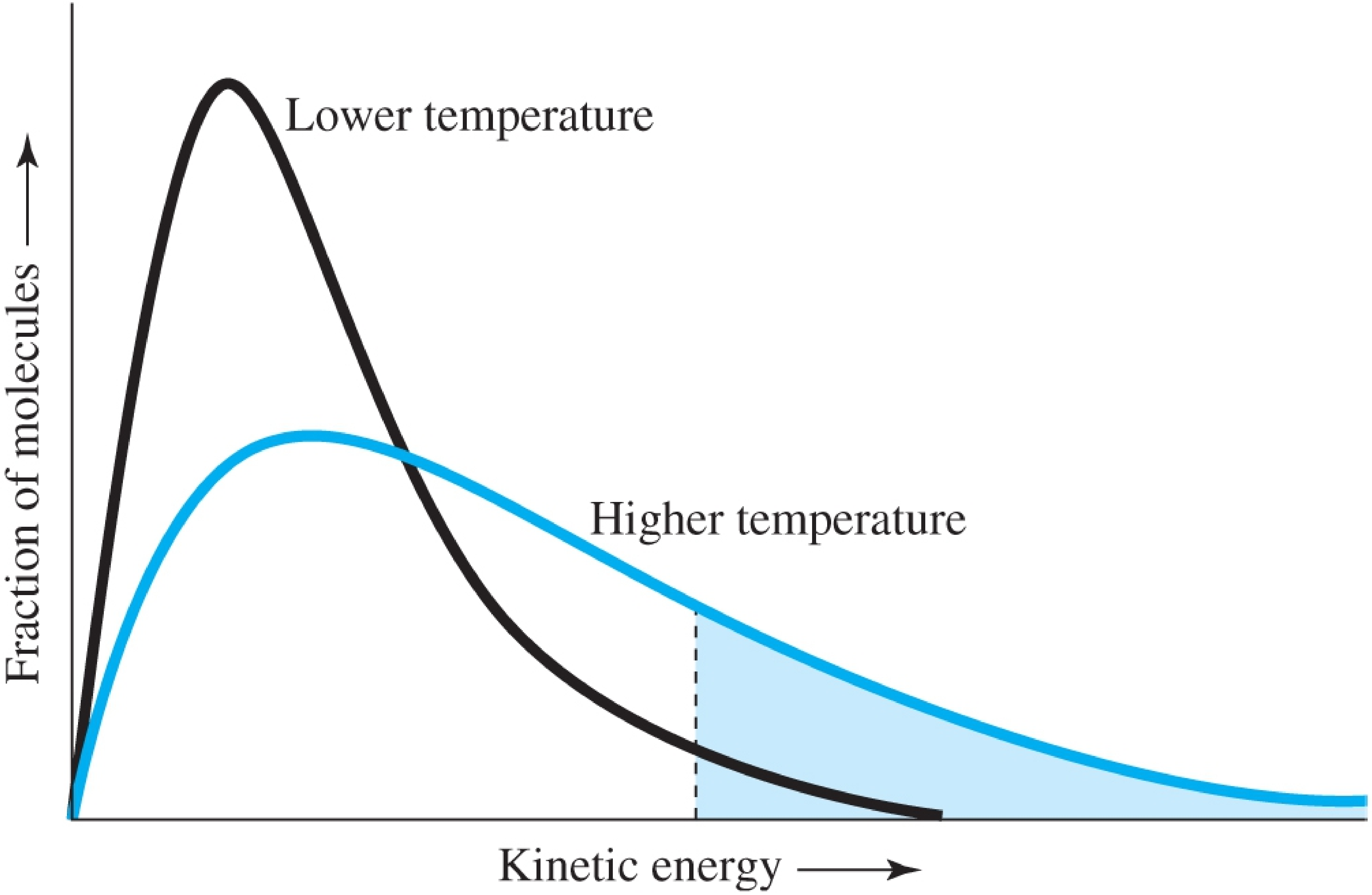
States present	Solid	Solid + liquid	Liquid	Liquid + gas	Gas
Key thermo-chemical property	Specific heat of solid	Heat of fusion or heat of solidification	Specific heat of liquid	Heat of vaporization or heat of condensation	Specific heat of gas
Calculational equation	$Q = C \times m \times \Delta T$	$Q = \Delta H_{\text{fus}} \times m$ or $Q = \Delta H_{\text{sol}} \times m$	$Q = C \times m \times \Delta T$	$Q = \Delta H_{\text{vap}} \times m$ or $Q = \Delta H_{\text{con}} \times m$	$Q = C \times m \times \Delta T$

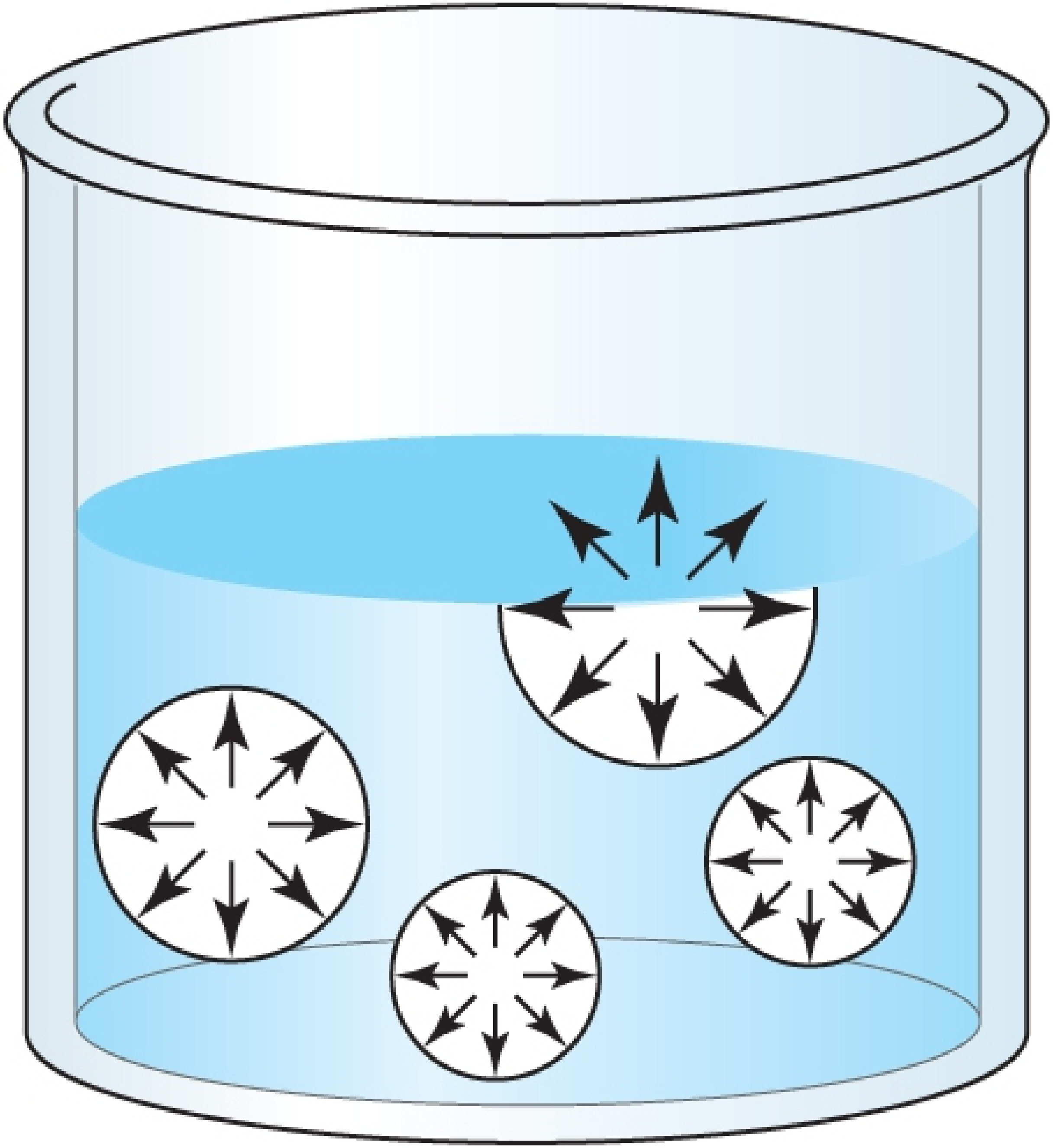
Under normal conditions, a pure substance in liquid phase will boil at a specific temperature.

At lower temperatures, though, some of the liquid particles will still evaporate.

Temperature is a measure of the average kinetic energy of matter: at an instant in time, some particles have more than this average energy.

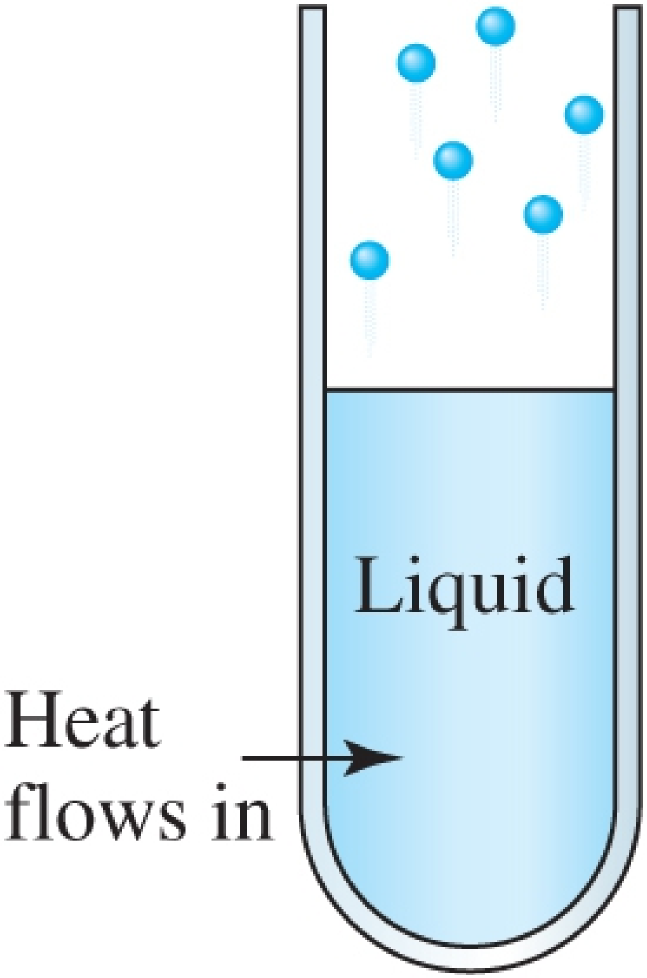
At a given temperatures, some of the liquid particles near the surface will have enough energy to escape and become gases.


Lower temperature
Higher temperature
Fraction of molecules ⟶
Kinetic energy ⟶

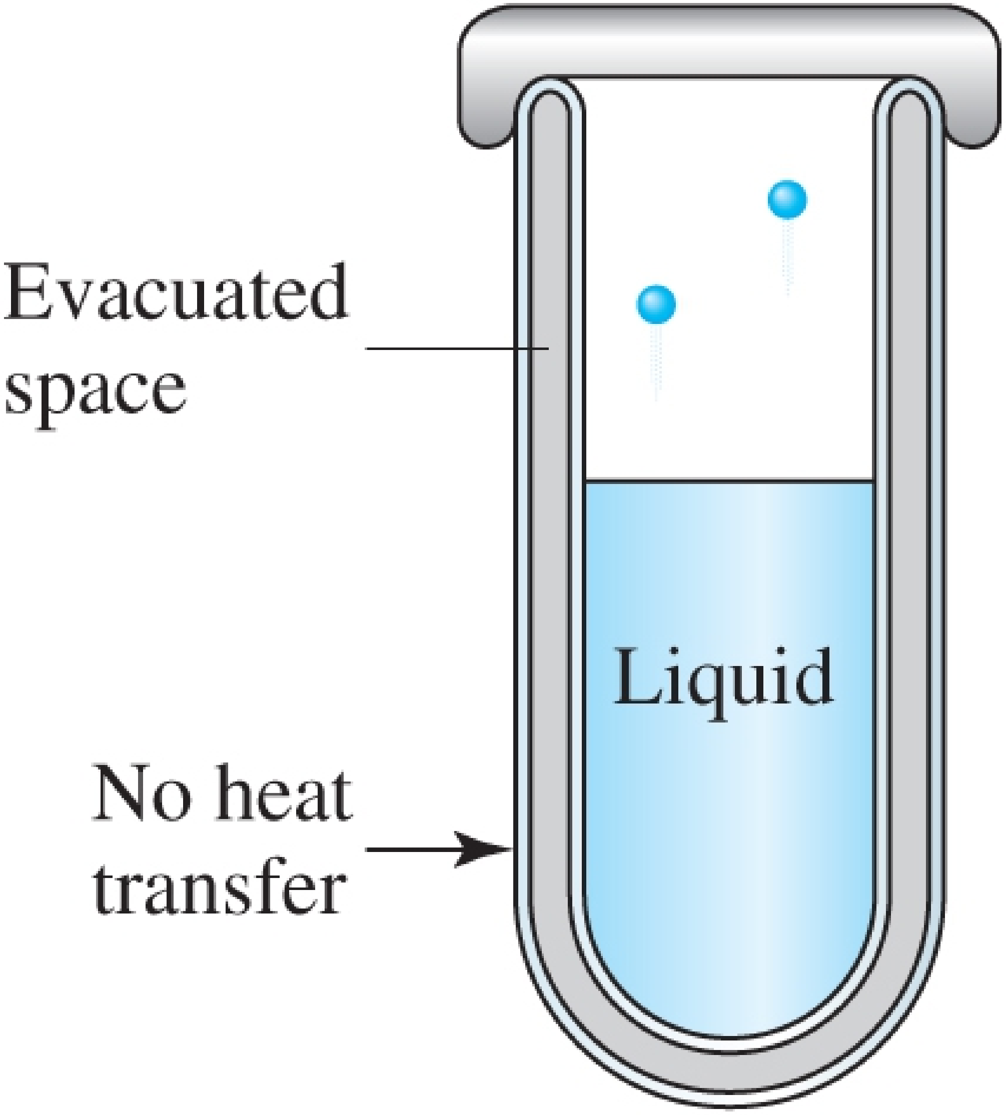


If the liquid is in an open container, eventually all of the liquid will evaporate.

If the container is sealed, however, then the gas molecules will eventually collide with and reenter the liquid phase.



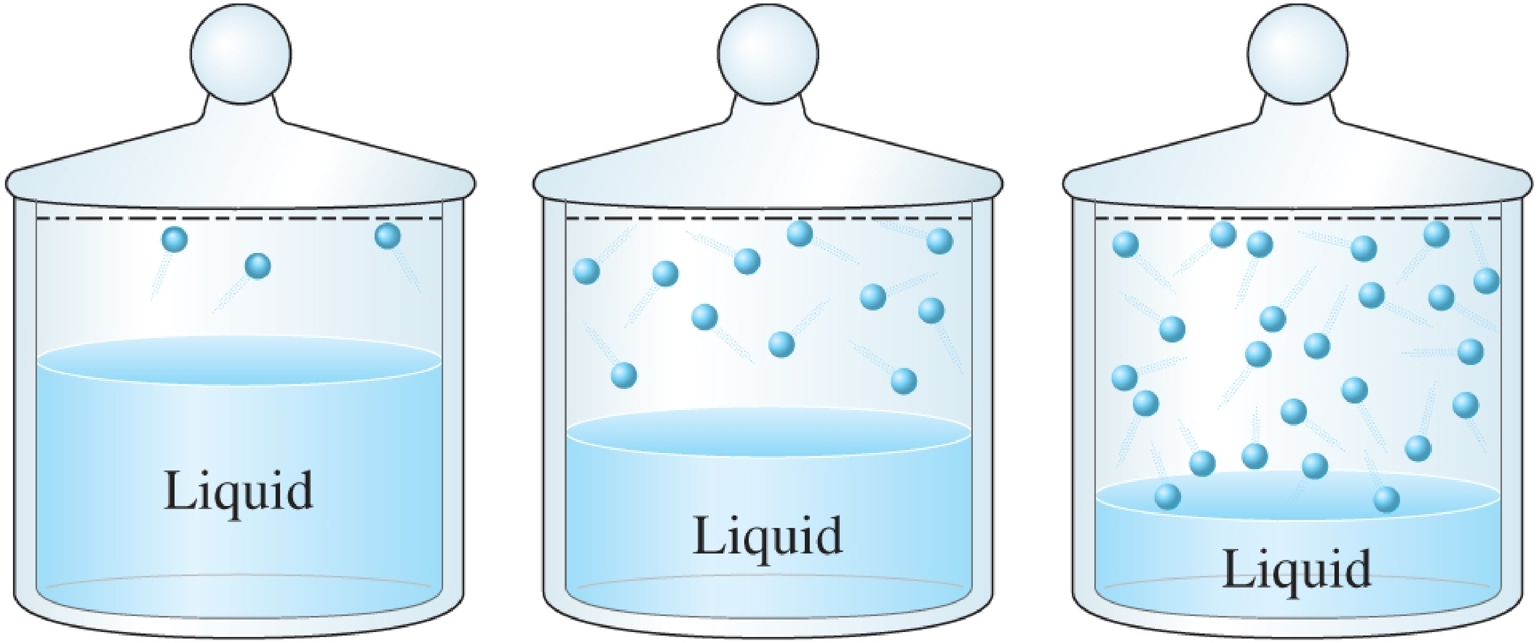
Uninsulated container



Insulated container

The evaporation and condensation processes will eventually reach equilibrium, and the concentration of gas particles becomes constant.

This concentration is referred to as 'vapor pressure', which is temperature dependent (increase with temperature).



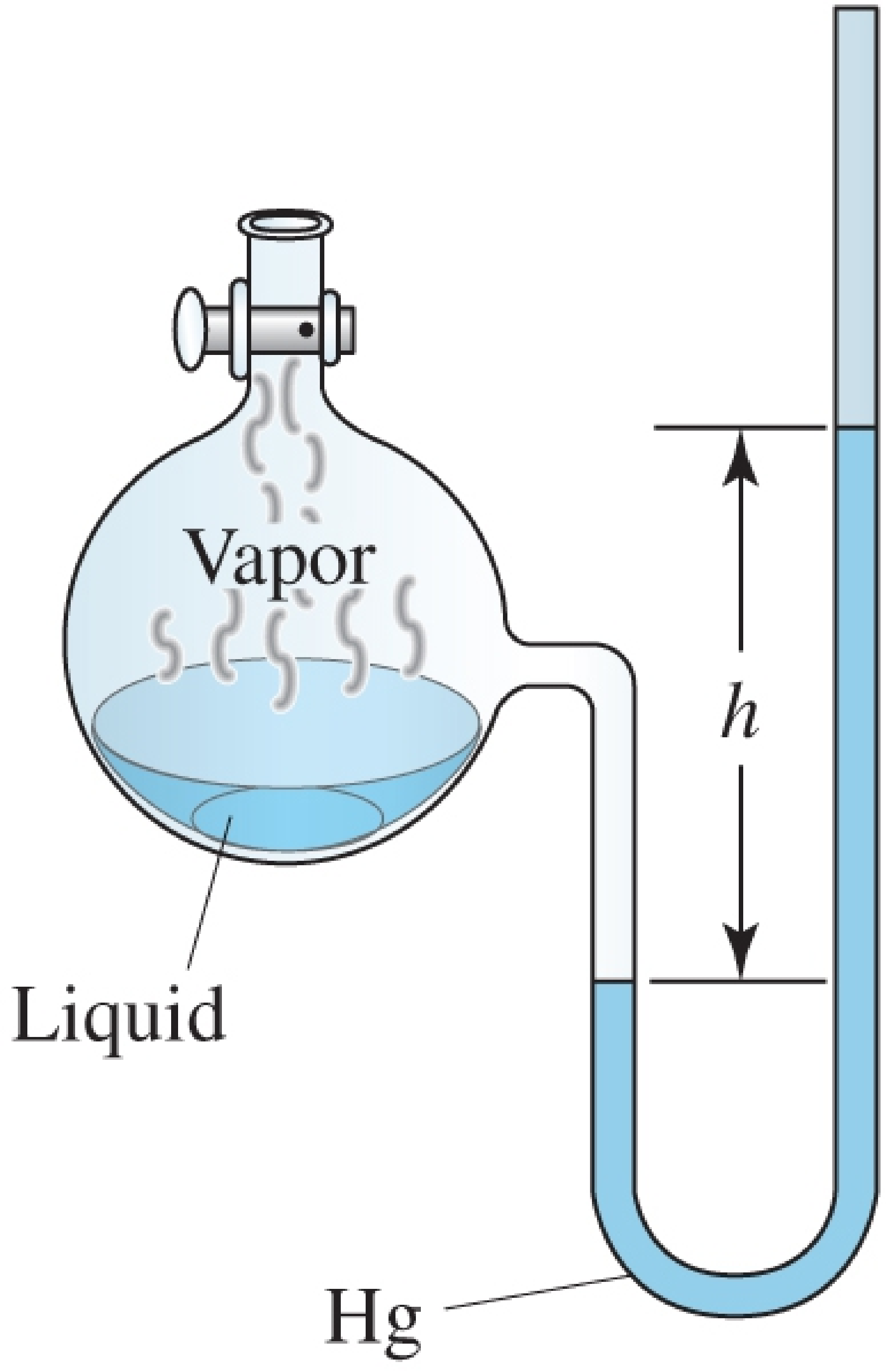
(a) Initially (b) After some time (c) At equilibrium

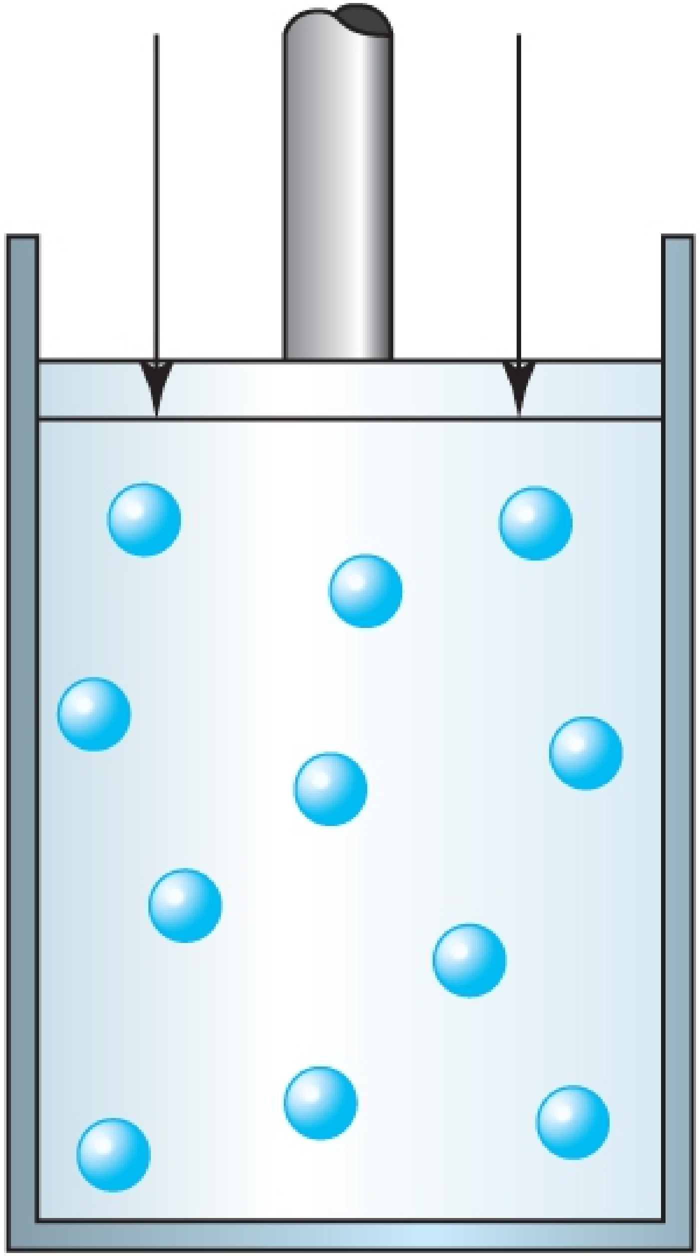
Constant liquid level

At higher temperatures, more particles have enough energy to form gases.

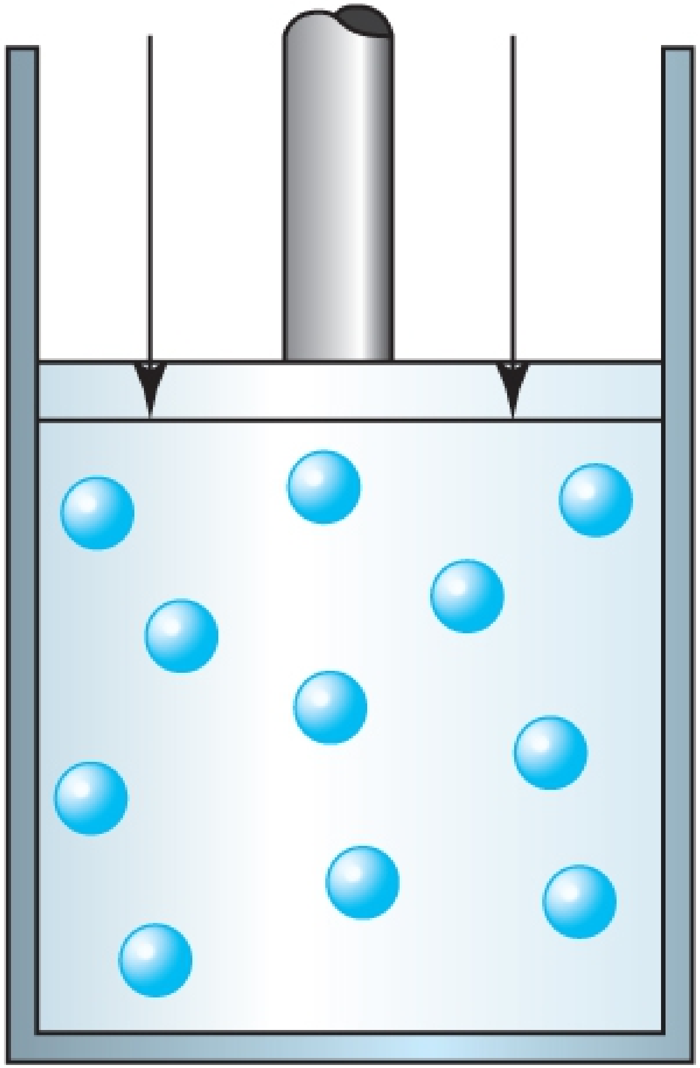
These gas particles partially contribute to the total pressure of all of the gases present (atmospheric pressure under normal conditions).

So as the temperature is increased, the vapor pressure exerted by the evaporated liquid increases.


Vapor
Liquid
h
Hg



Gas at low pressure



Gas at higher pressure

The boiling point occurs when this vapor pressure is equal to the total atmospheric pressure.

At higher altitudes, the atmospheric pressure is lower, so the boiling point of a liquid is lower than at sea level.

Water, for example, boils at 100 °C at sea level, but in Denver, Colorado, water boils at 95 °C.

TABLE 11.5 Vapor Pressure of Water at Various Temperatures

Temperature (°C)	Vapor Pressure (mm Hg)*	Temperature (°C)	Vapor Pressure (mm Hg)*
0	4.6	60	149.4
10	9.2	70	233.7
20	17.5	80	355.1
30	31.8	90	525.8
40	55.3	100	760.0
50	92.5		

**The units used to specify vapor pressure in this table will be discussed in detail in Section 12.2.*

TABLE 11.6 Variation of the Boiling Point of Water with Elevation

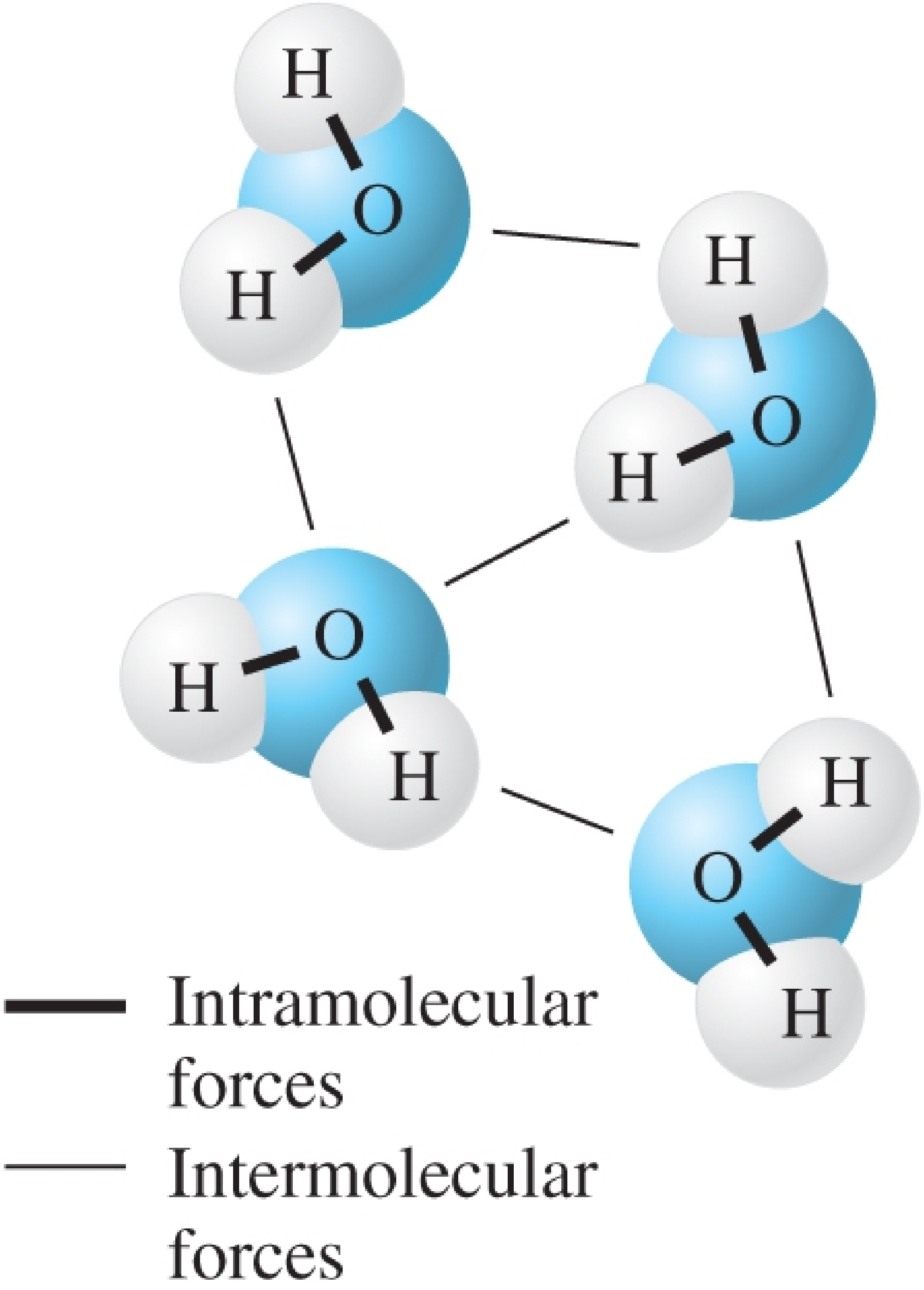
Location	Elevation (ft above sea level)	Boiling Point of Water (°C)
San Francisco, CA	0	100.0
Salt Lake City, UT	4,390	95.6
Denver, CO	5,280	95.0
La Paz, Bolivia	12,795	91.4
Mount Everest	28,028	76.5

TABLE 11.7 Boiling Points of Water in a Pressure Cooker

Pressure above Atmospheric		
lb/in.^{2}	mm Hg	Boiling Point of Water (°C)
5	259	108
10	517	116
15	776	121

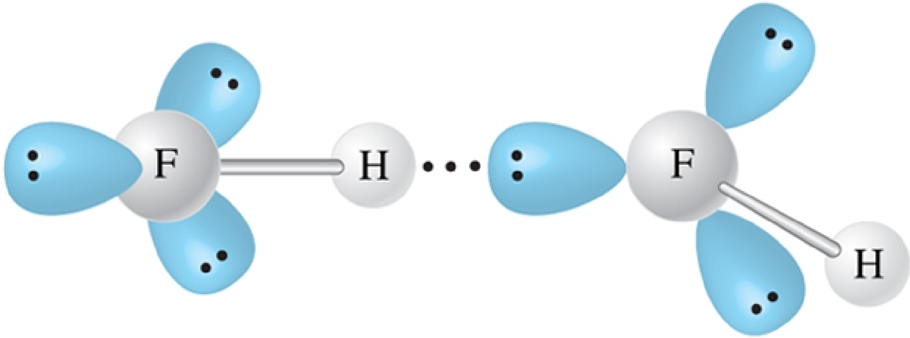
Intermolecular Forces

- Attractive forces between two molecules
- Also used to refer to attraction between atoms or other small particles
- Electrostatic based – weak attraction between opposite charges
- Much weaker than covalent or ionic bonds

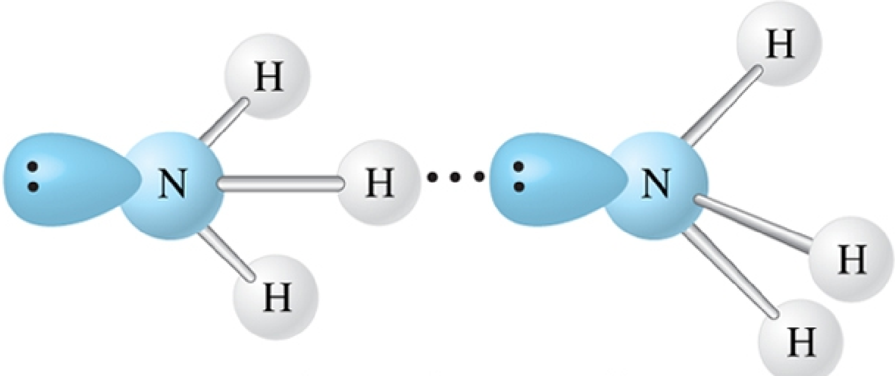

H
O
H
H
O
H
H
O
H
H
O
H
Intramolecular forces
Intermolecular forces

- Dipole-dipole interaction – attraction between two polar molecules (+ on one dipole attracted to – on the other dipole, like a weak ionic bond)

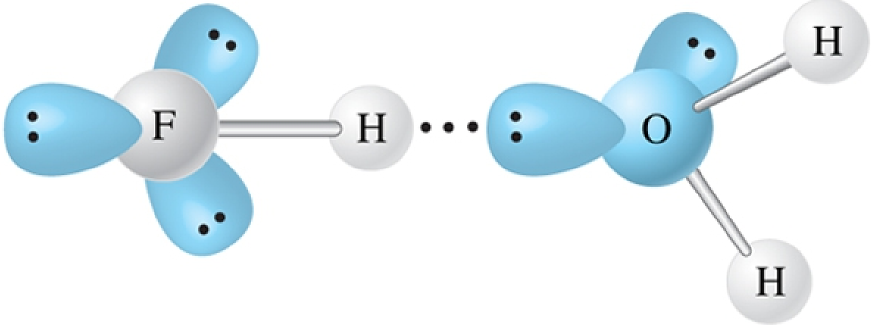
- Hydrogen 'bond' – Very strong dipole-dipole interaction.
  - Occurs between a hydrogen bonded to a small, very electronegative atom (N, O or F) and unshared electrons on a small, very electronegative atom (N,O or F)



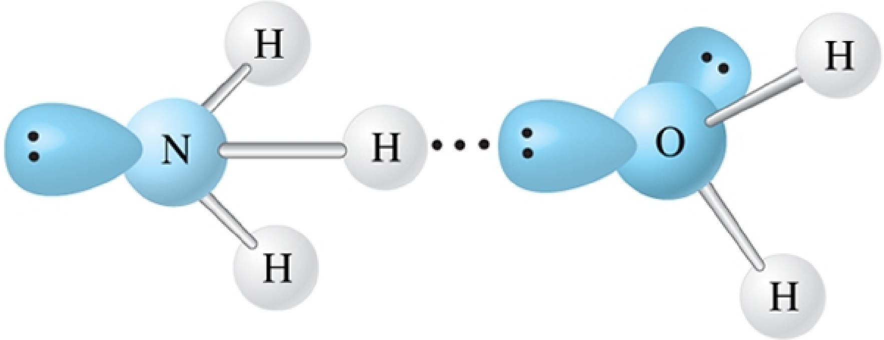
Hydrogen fluoride–hydrogen fluoride



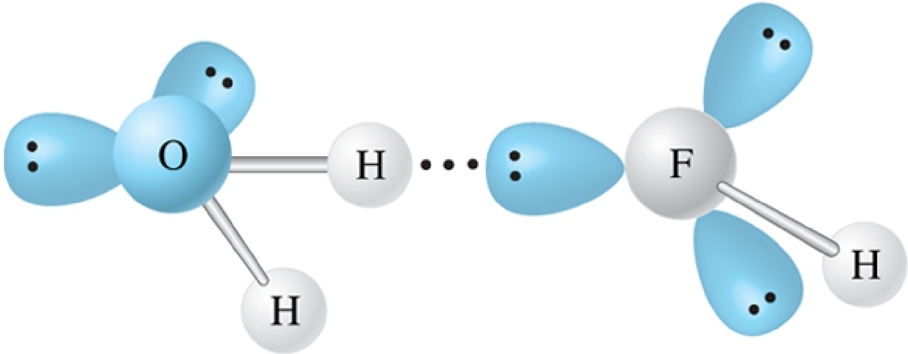
Ammonia–ammonia



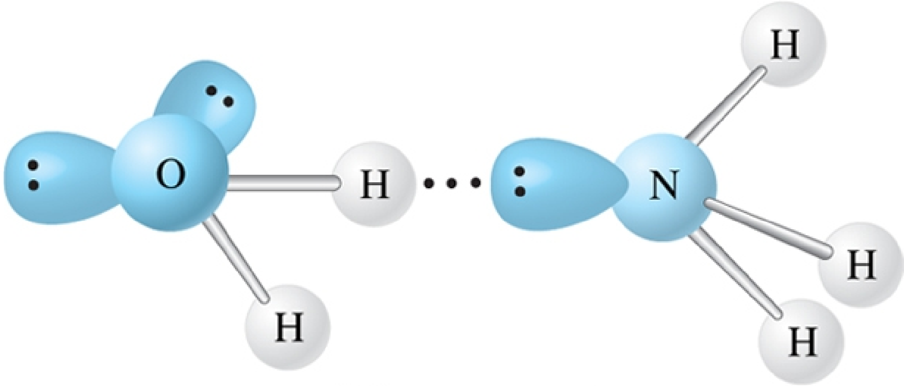
Hydrogen fluoride–water



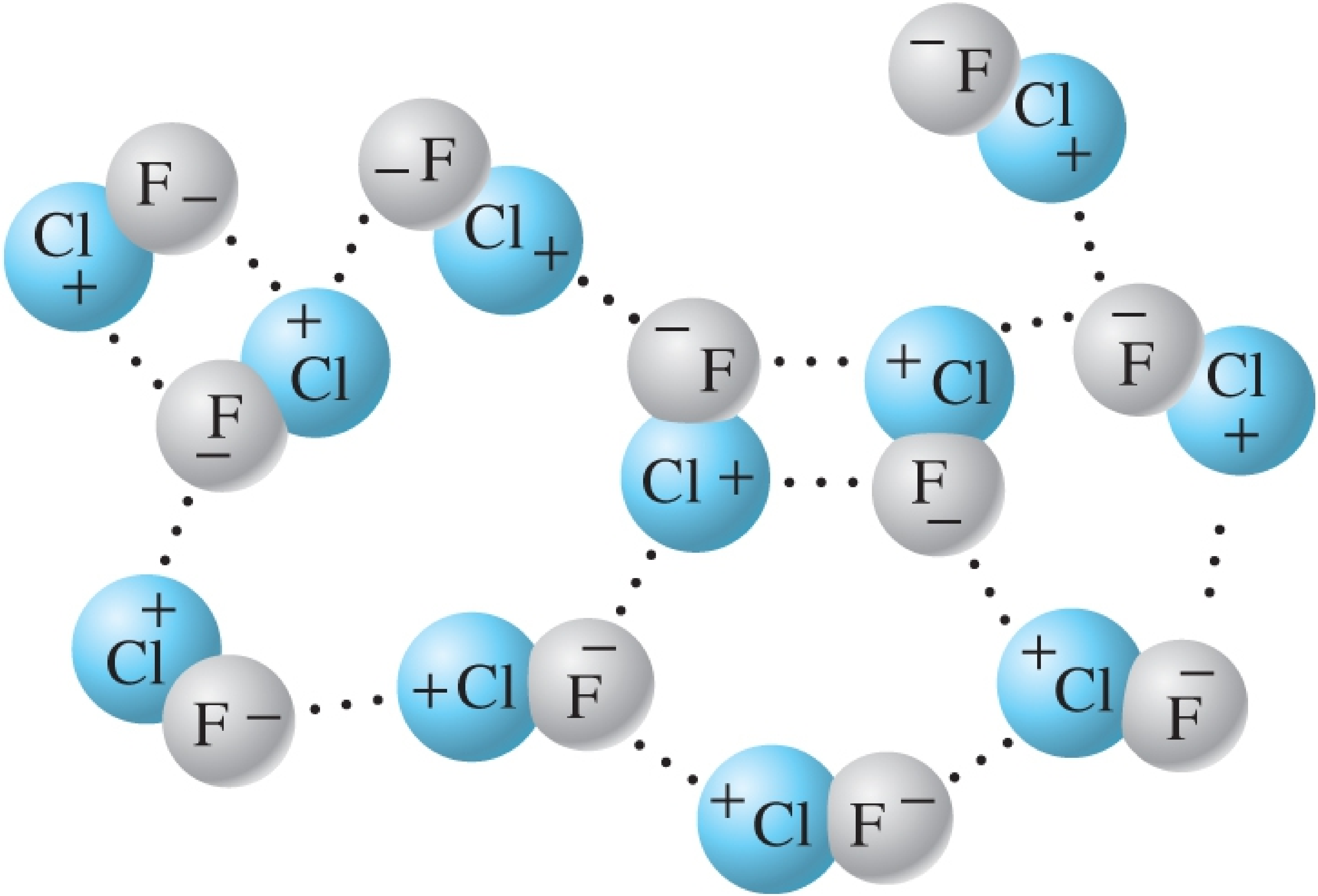
Ammonia–water

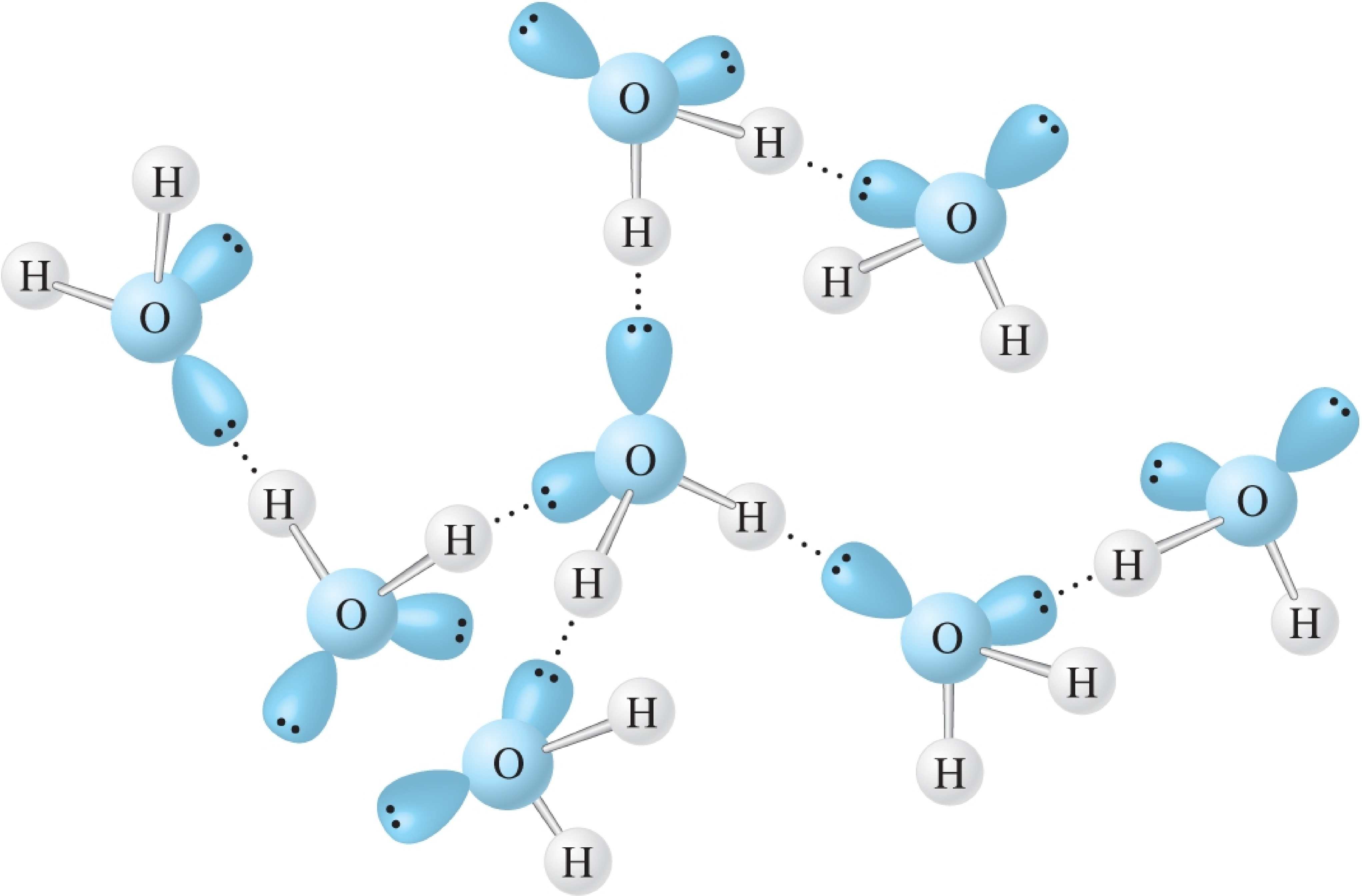


Water–hydrogen fluoride

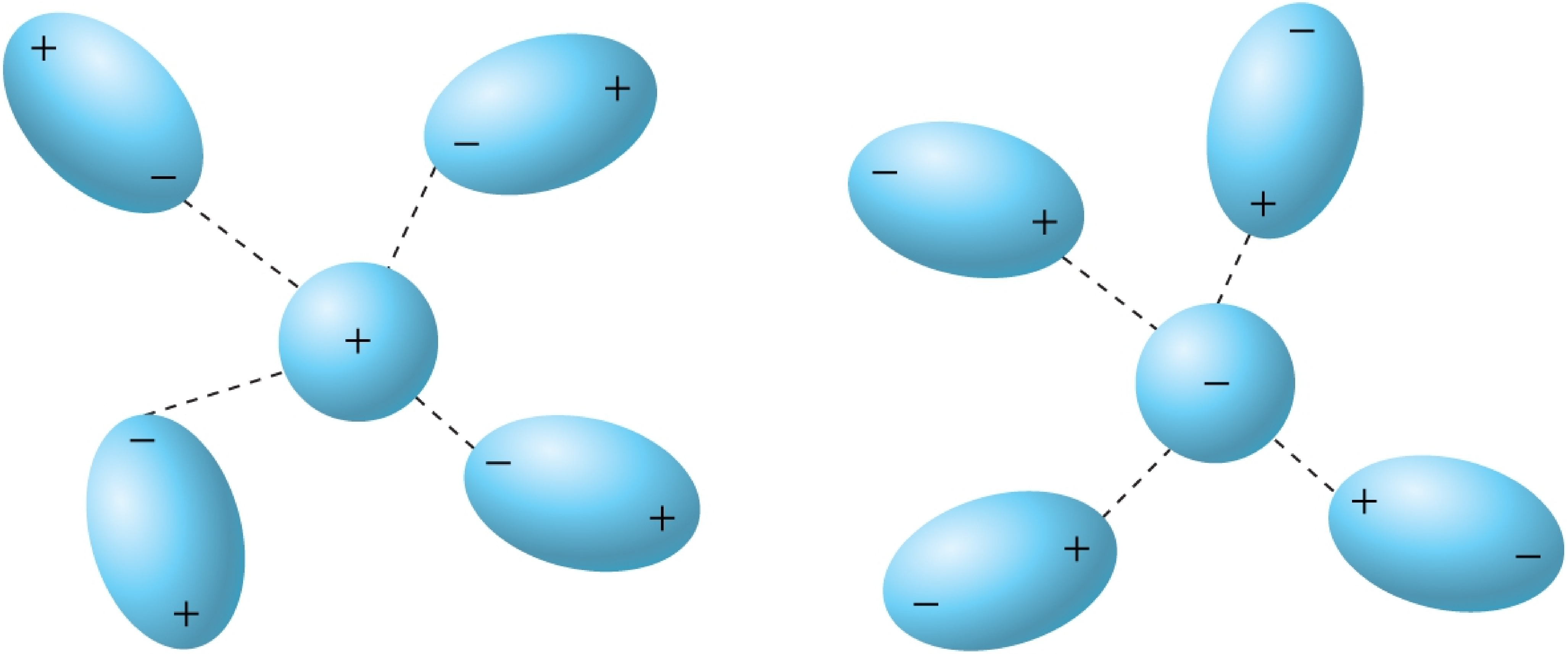


Water–ammonia


Cl
F
+
−


O
H
O
H
H
O
H
H
O
H
H
H
O
O
H
H
O
H
O
H
H
H
O
H
H

- Ion-dipole interaction
  - Attraction between an ion and the oppositely charged end of a polar bond
  - When a solvent dissolves an ion, typically several polar molecules surround the ion to bring it into solution



(a) (b)

- Dipole-induced dipole interaction – As a dipole approaches another atom or molecule, it induces a weak temporary dipole in the other atom/molecule, which results in a weak attraction.

- Induced dipole-induced dipole – As two non-polar atoms/molecules near each other, they induce weak temporary dipoles, resulting in a very weak attraction.

- Attraction due to induced dipoles are also called:
  - London forces
  - Weak forces
  - Van der Waals forces

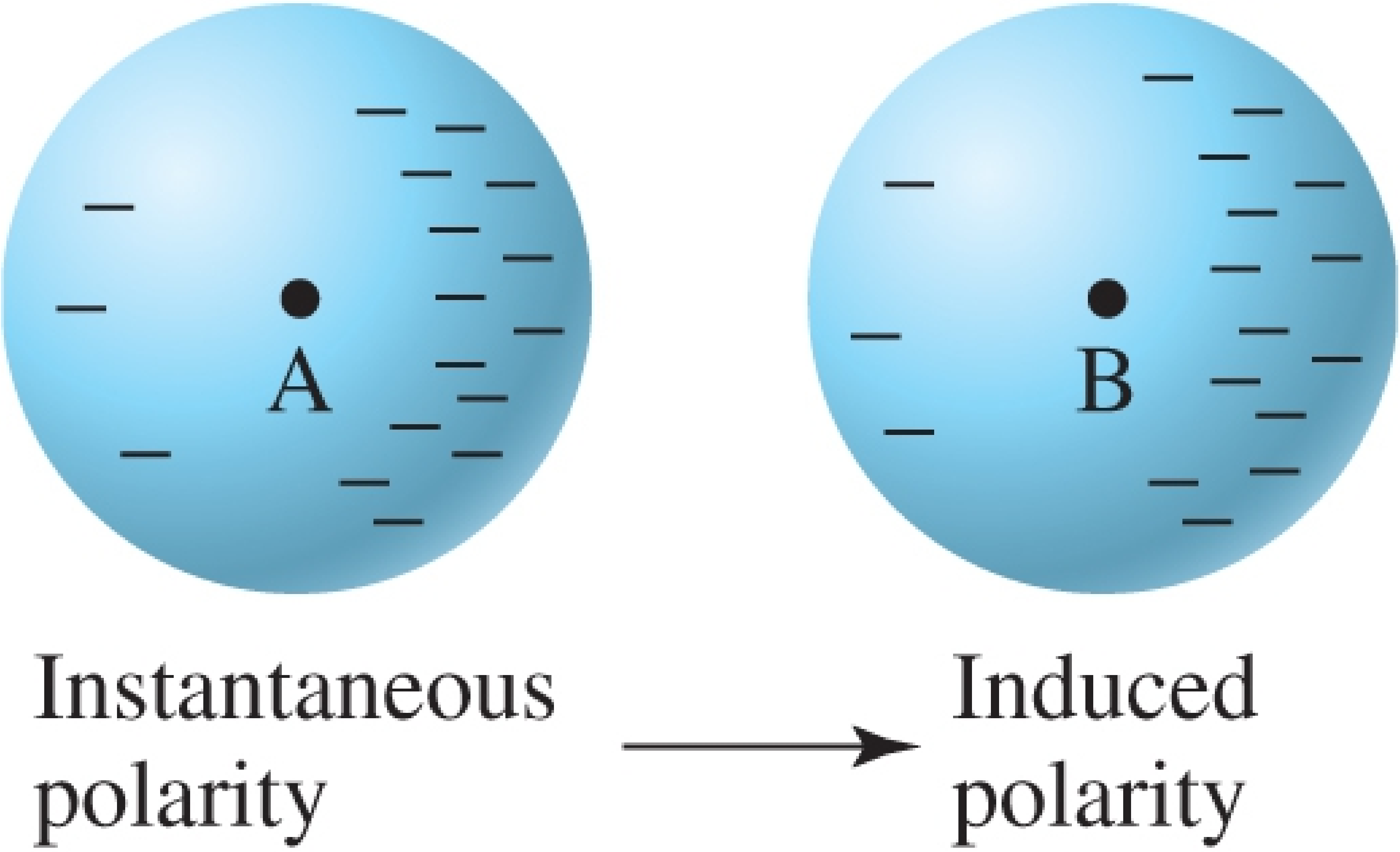

A
B
Instantaneous polarity
Induced polarity

TABLE 11.8 Boiling Point Trends for Related Series of Nonpolar Molecules

Substance	Molecular Mass (amu)	Boiling Point (°C)
Noble Gases		
He	4.0	–269
Ne	20.2	–246
Ar	39.9	–186
Kr	83.8	–153
Xe	131.3	–107
Rn	222	–62
Group VIIA Elements		
F_2	39.0	–187
Cl_2	70.9	–35
Br_2	159.8	+59
I_2	253.8	+184

Bond Strength (strongest to weakest)

- Covalent bonds
- Ionic bonds

- Ion-dipole attraction (~1/10 of ionic bond strength)
- dipole-dipole attraction
  - Hydrogen 'bonding' – strongest dipole/dipole
- London or Weak forces

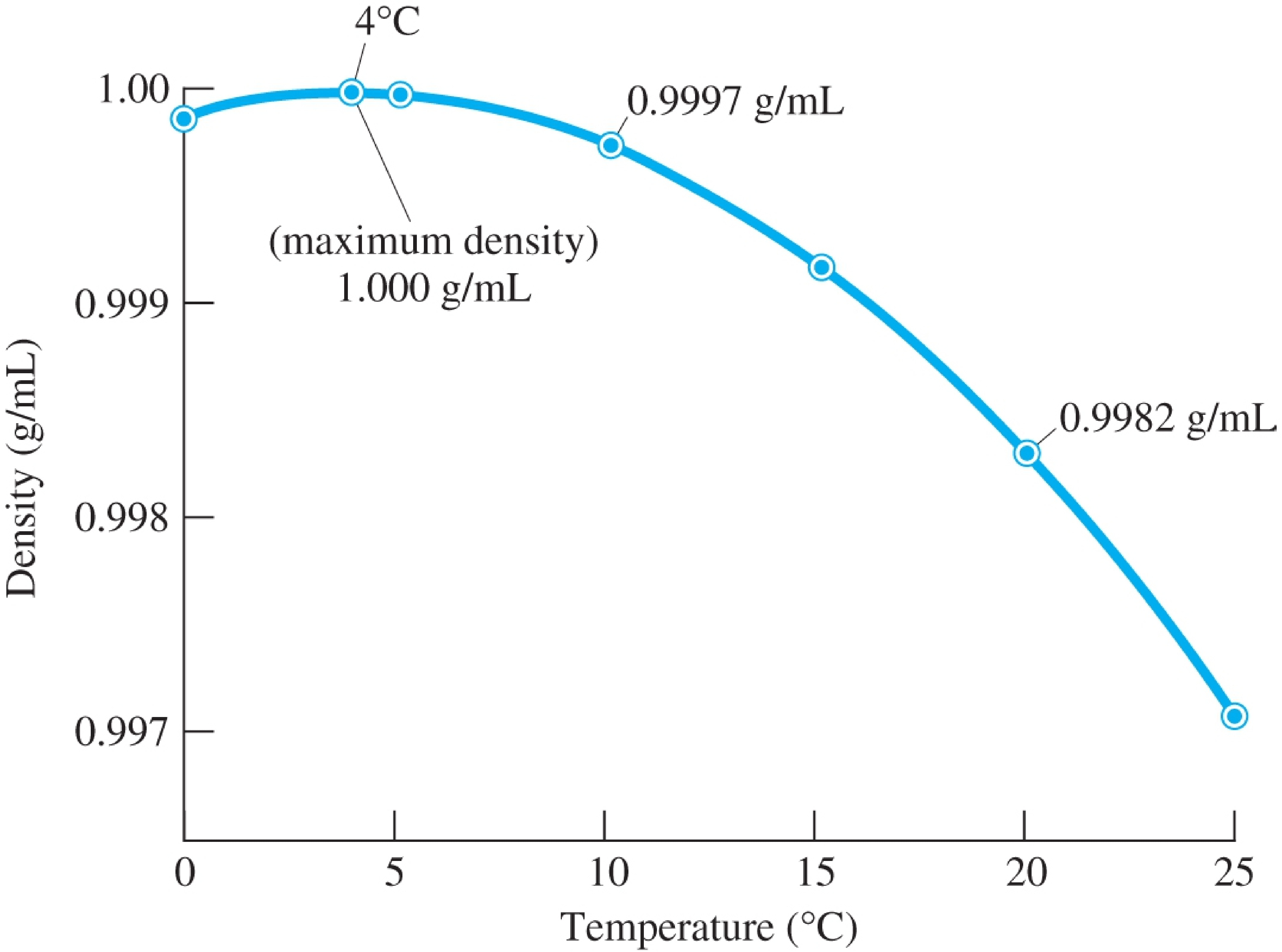
Intermolecular forces account for many of the distinct properties of water:

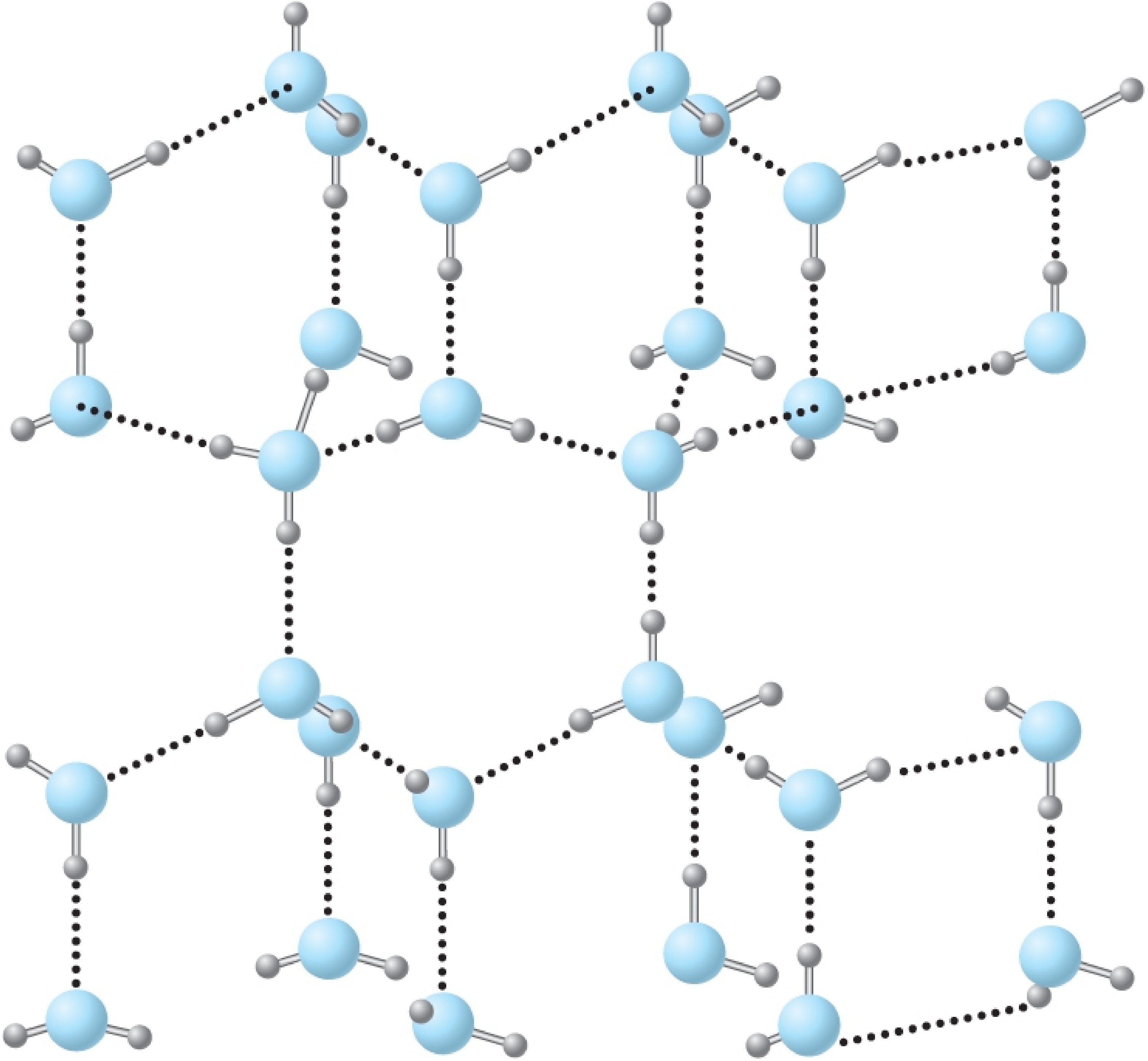
- Very high boiling point for such a small molecule, due to hydrogen bonding.

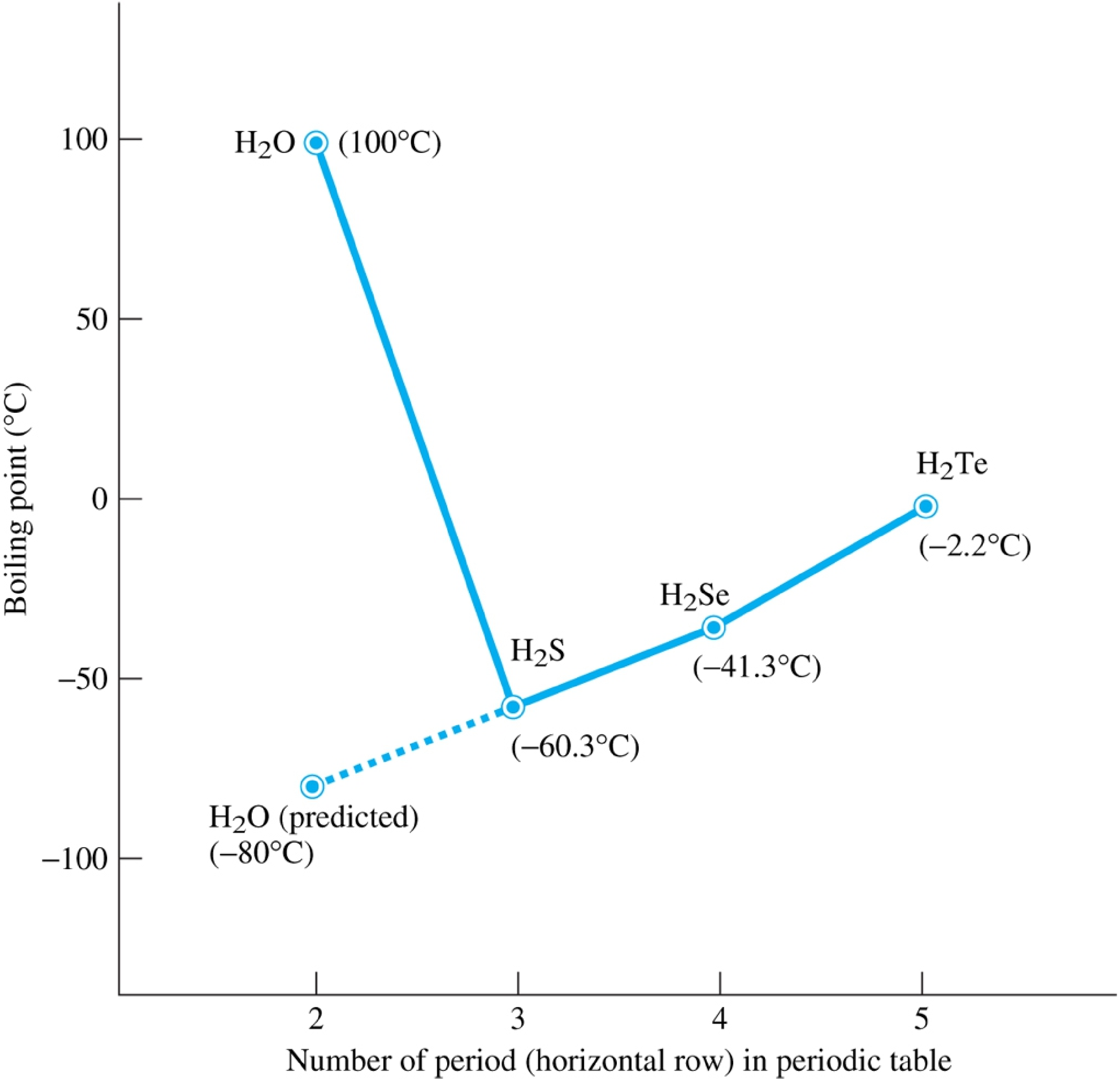
- Very low vapor pressure for such a small molecule.

- Excellent solvent – dissolves many polar molecules as well as many ionic compounds

- Solid is lower density than liquid – in solid form the molecules form a lower density crystal to maximize the attractive hydrogen bonding

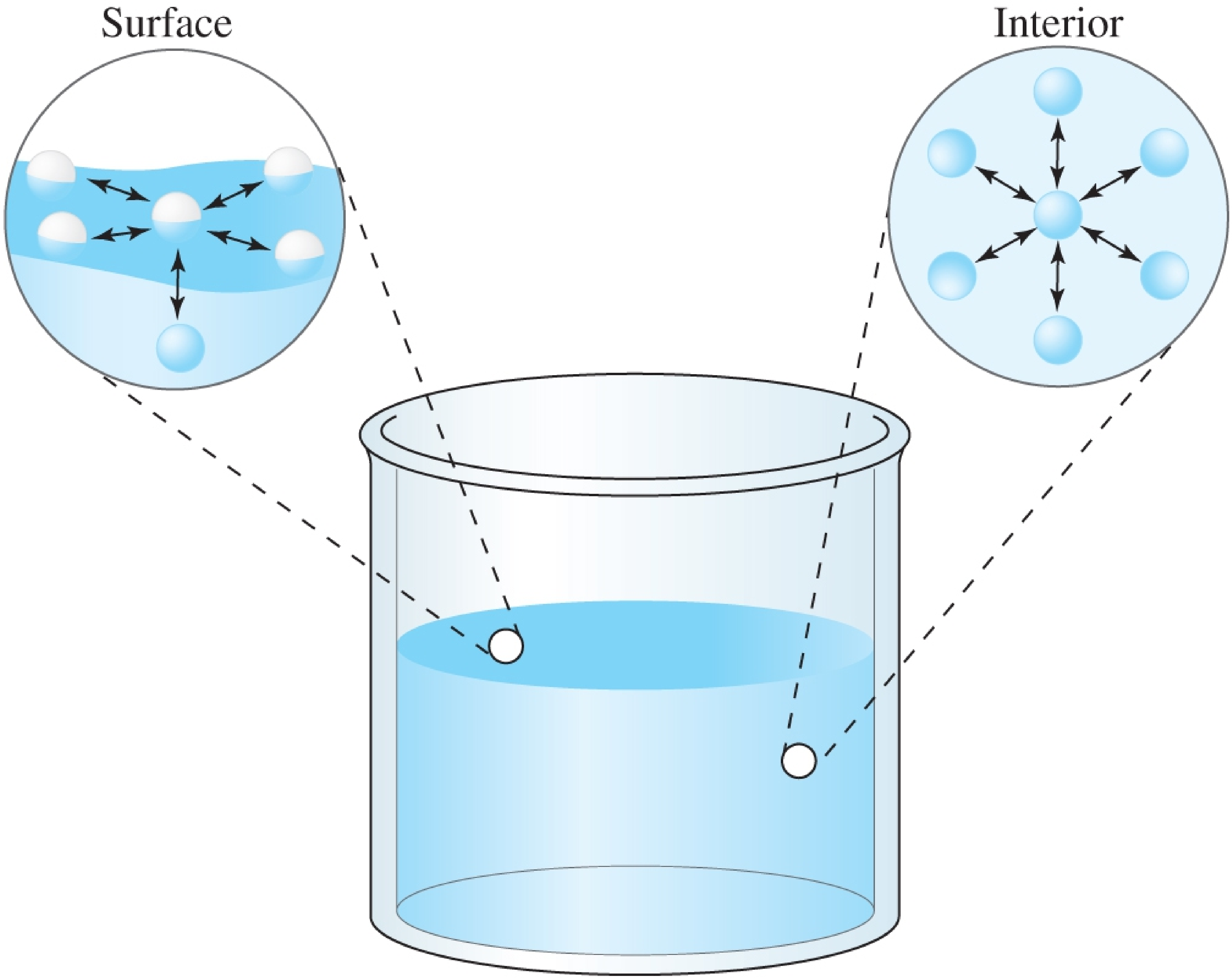

4°C
1.00
0.999
0.998
0.997
(maximum density)
1.000 g/mL
0.9997 g/mL
0.9982 g/mL
Density (g/mL)
0
5
10
15
20
25
Temperature (°C)




100
50
0
−50
−100
Boiling point (°C)
H_2O (100°C)
H_2S (−60.3°C)
H_2Se (−41.3°C)
H_2Te (−2.2°C)
H_2O (predicted) (−80°C)
2
3
4
5
Number of period (horizontal row) in periodic table

Surface Tension

- Indicates how attracted a substance is to itself.
- Due to intermolecular forces.
- Liquids with large surface tensions have strong intermolecular forces.
- Surface tension explains the meniscus of water in a glass tube.
- A liquid forms 'beads' or droplets to minimize surface area and maximize intermolecular attraction
- How well a liquid 'beads up' on a solid surface depends on:
  - The liquid surface tension
  - How attracted the liquid is to the surface


Surface
Interior