# Chapter 7

## Gases, Liquids, and Solids

# Gases, Liquids, and Solids

→ **CO 7.1**

Ice, water, and mist are simultaneously present in this winter scene in Yellowstone National Park.

# Hydro-power from potential energy

Betty Weiser/Photo Researchers

← Fig. 7.1

**The water in the lake behind the dam has potential energy as a result of its position.**

# Kinetic Energy and Collisions

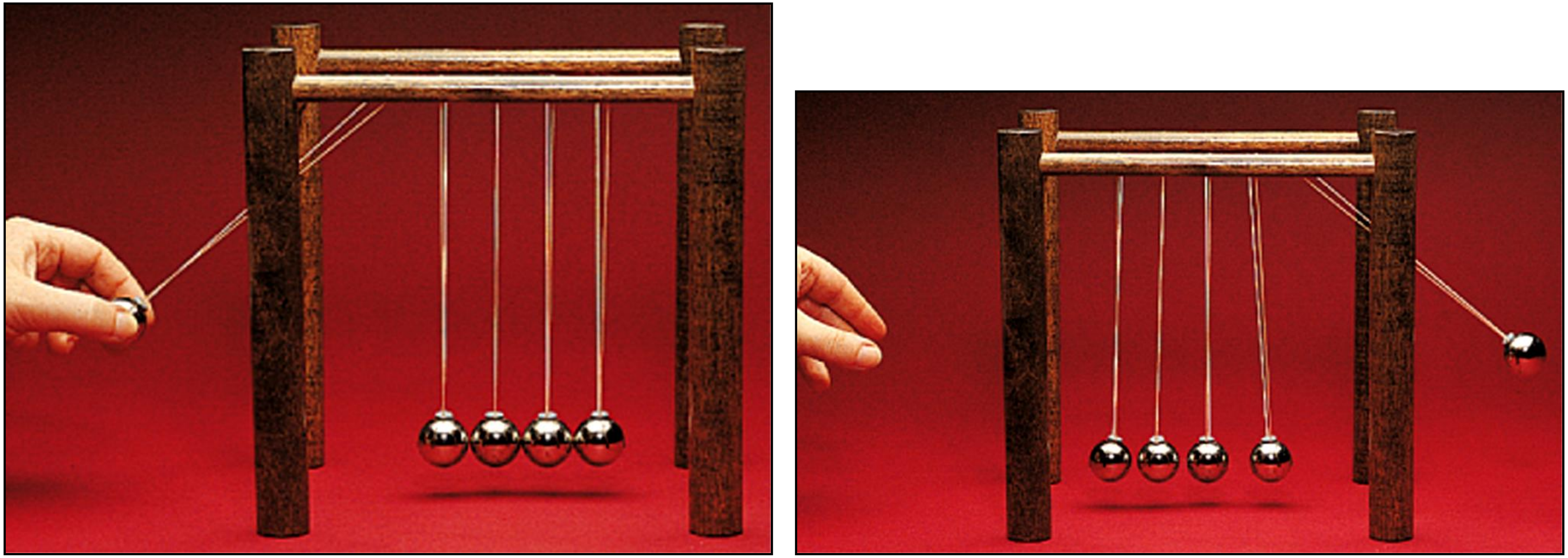


**Fig. 7.2**

Upon release, the steel ball on the left transmits its kinetic energy through a series of elastic collisions to the ball on the right.

# Properties of Gases, Liquids, and Solids

| Property | Solid State | Liquid State | Gaseous State |
|---|---|---|---|
| volume and shape | definite volume and definite shape | definite volume and indefinite shape; takes the shape of its container to the extent that it is filled | indefinite volume and indefinite shape; takes the volume and shape of the container that it completely fills |
| density | high | high, but usually lower than corresponding solid | low |
| compressibility | small | small, but usually greater than corresponding solid | large |
| thermal expansion | very small: about 0.01% per °C | small: about 0.10% per °C | moderate: about 0.30% per °C |

Table 7.1

# Arrangement of Particles in three States

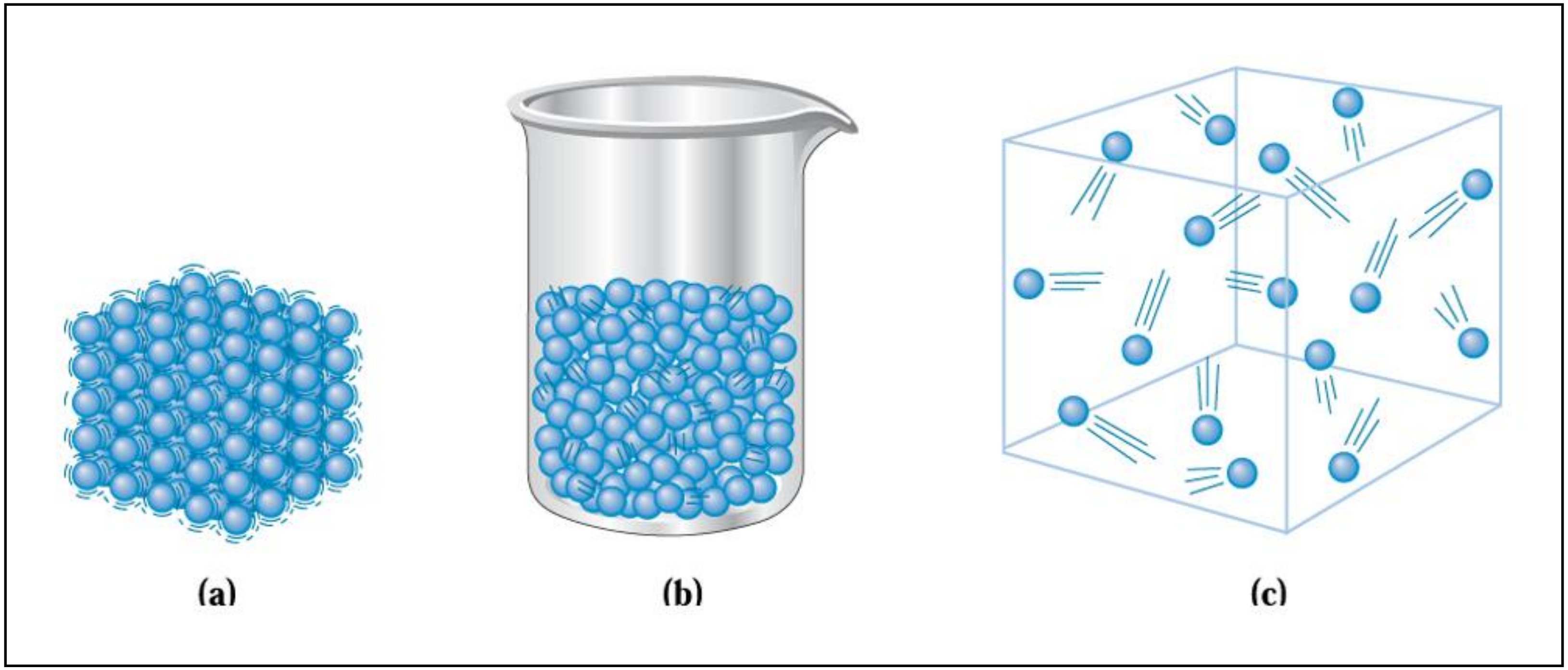



Fig. 7.3

**(a) In a solid, the particles are close together. (b) In a liquid, the particles slide freely over one another. (c) In a gas, the particles are in random motion.**

# Gases in random motion

→ Fig. 7.4

**Gas molecules can be compared to billiard balls in random motion, bouncing off one another.**

# Popping into different states

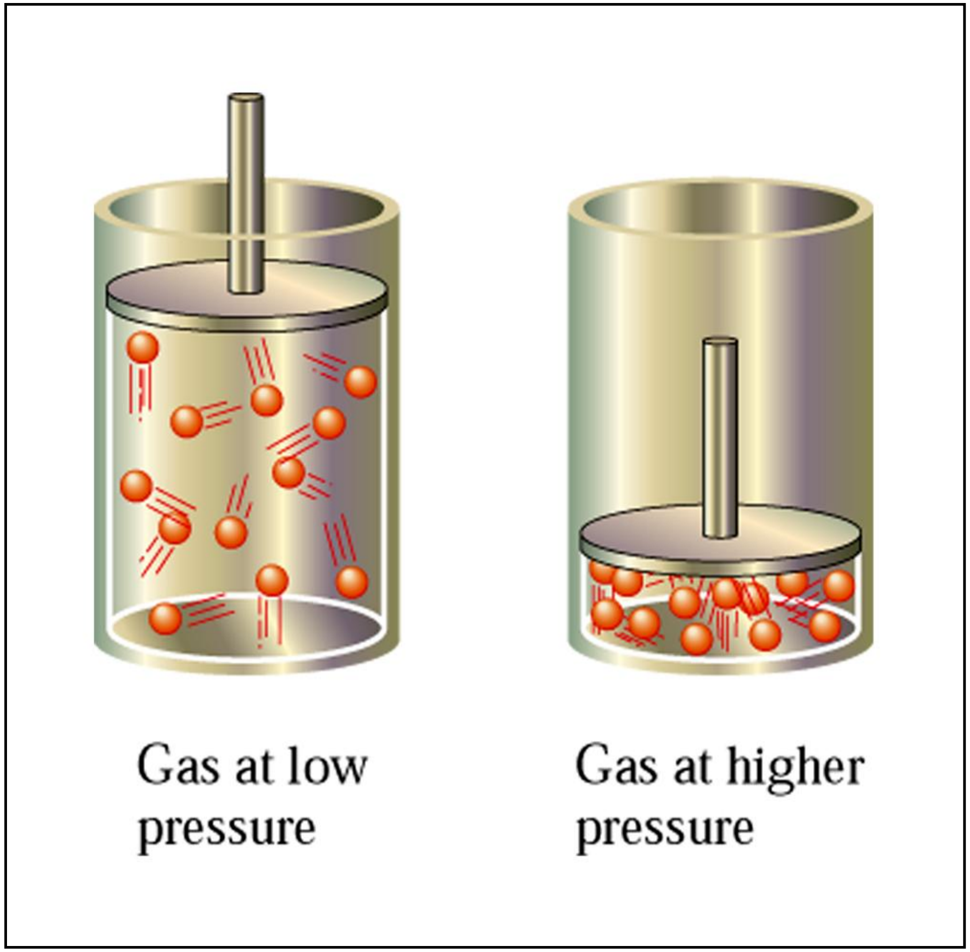
→ **CC 7.1**

Phil Degginger/Color-Pic

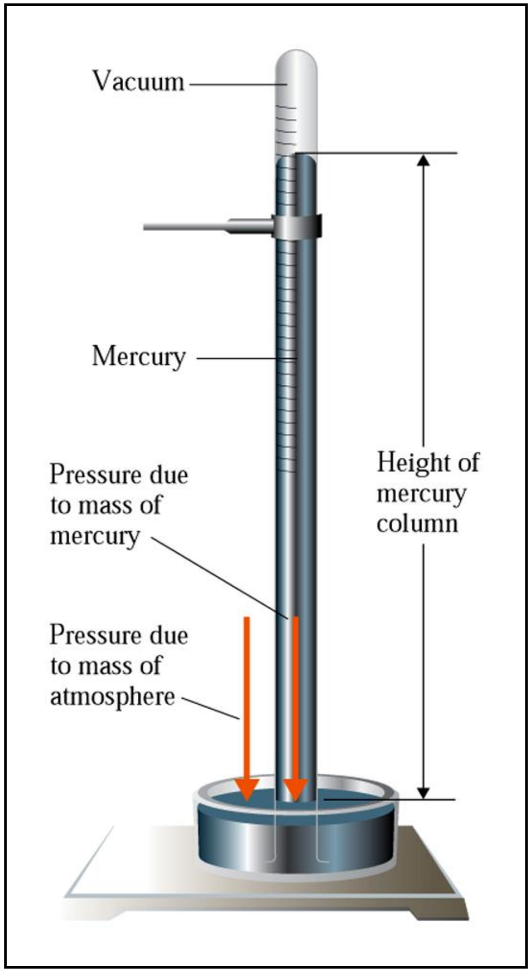
# Compressibility of gases

→ Fig. 7.5

**When a gas is compressed, the amount of empty space in the container is decreased.**



# Barometer: Measuring the Pressure of gases



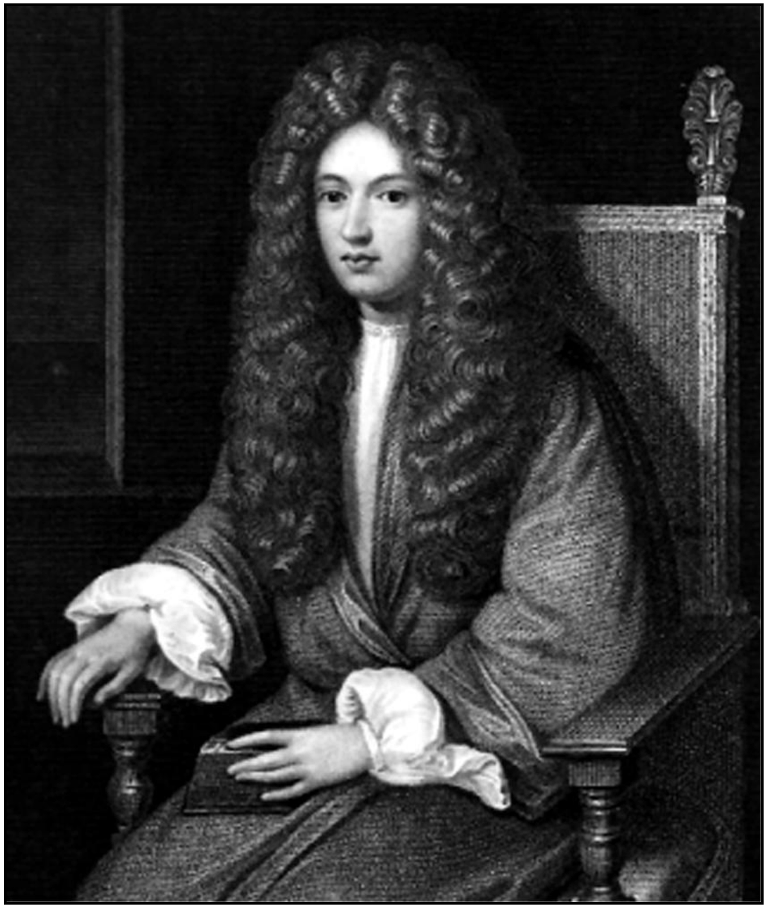

← Fig. 7.6

**The essential components of a mercury barometer are a graduated glass tube, a glass dish, and liquid mercury.**

# Boyle and the gas Laws

→ Fig. 7.7

**Robert Boyle was self-taught. Through his efforts, the true value of experimental investigation was first recognized.**



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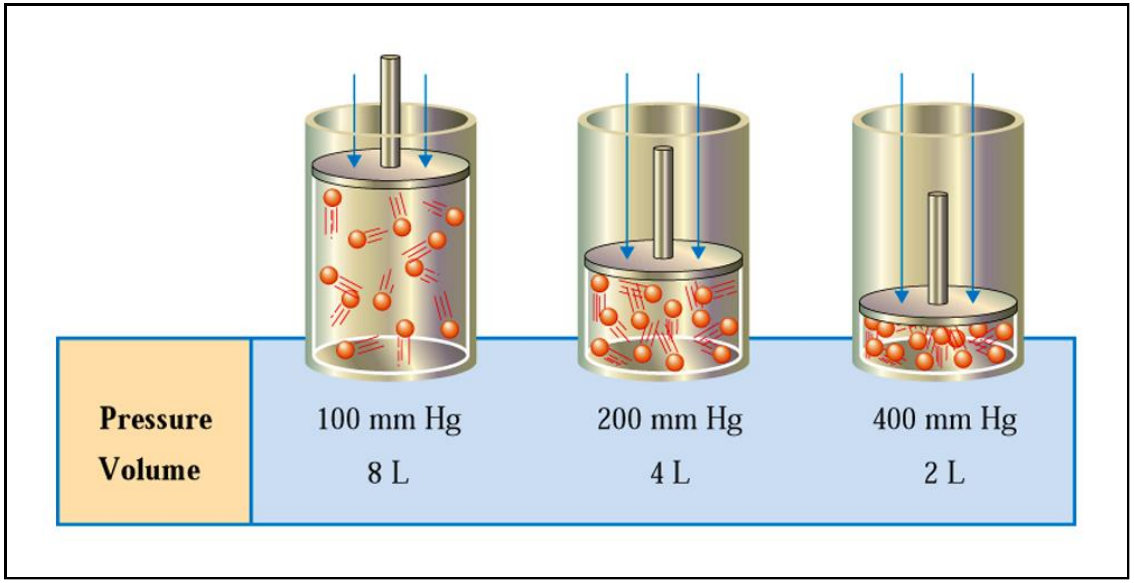
# Boyles' Law:




← Fig. 7.8
**Data illustrating the inverse proportionality associated with Boyle's law.**

# Charles Law

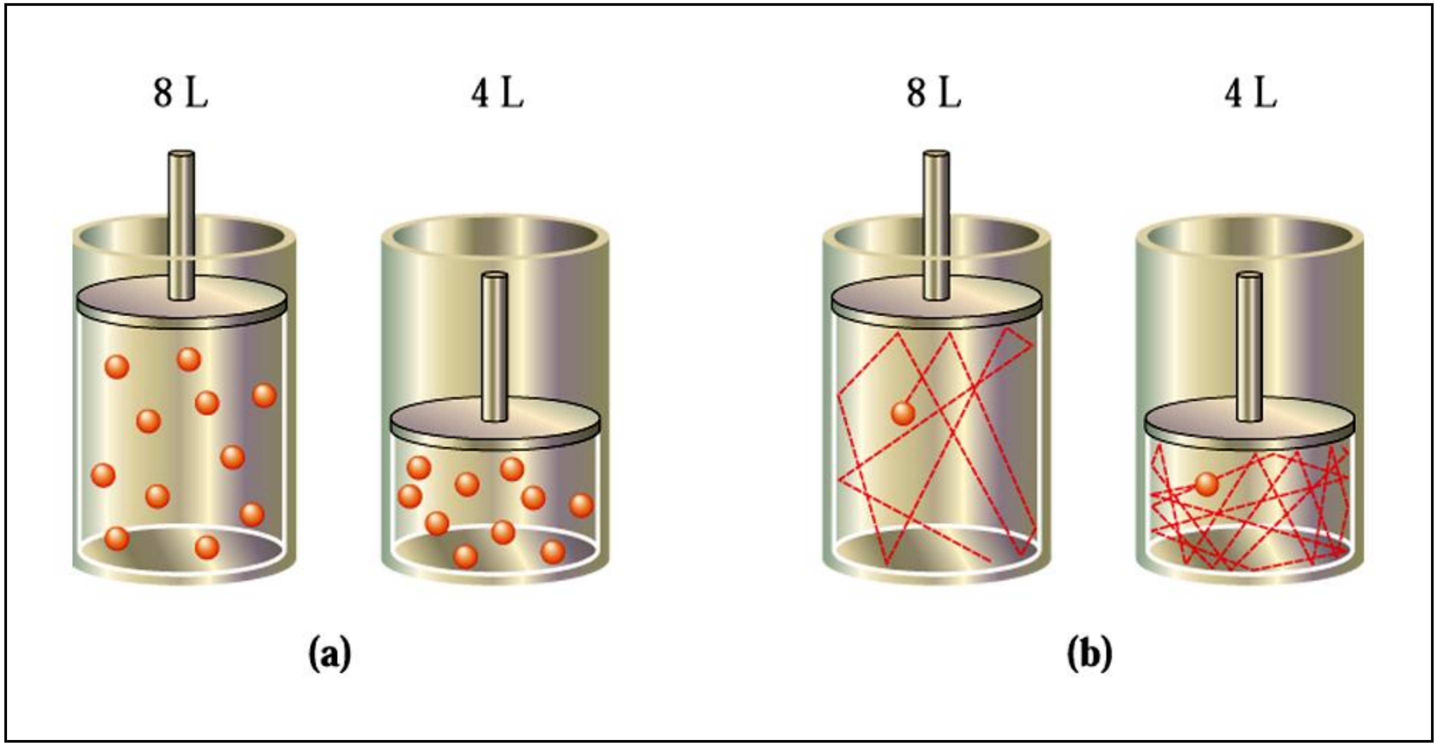
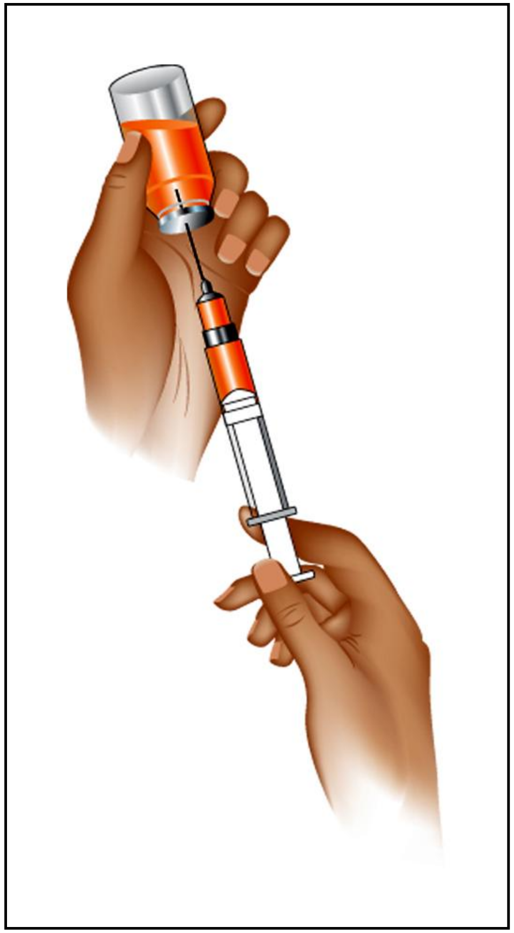



Fig. 7.9 **When the volume of a gas at constant temperature decreases by half, the average number of times a molecule hits the container walls is doubled.**

# Application of Boyle's Law

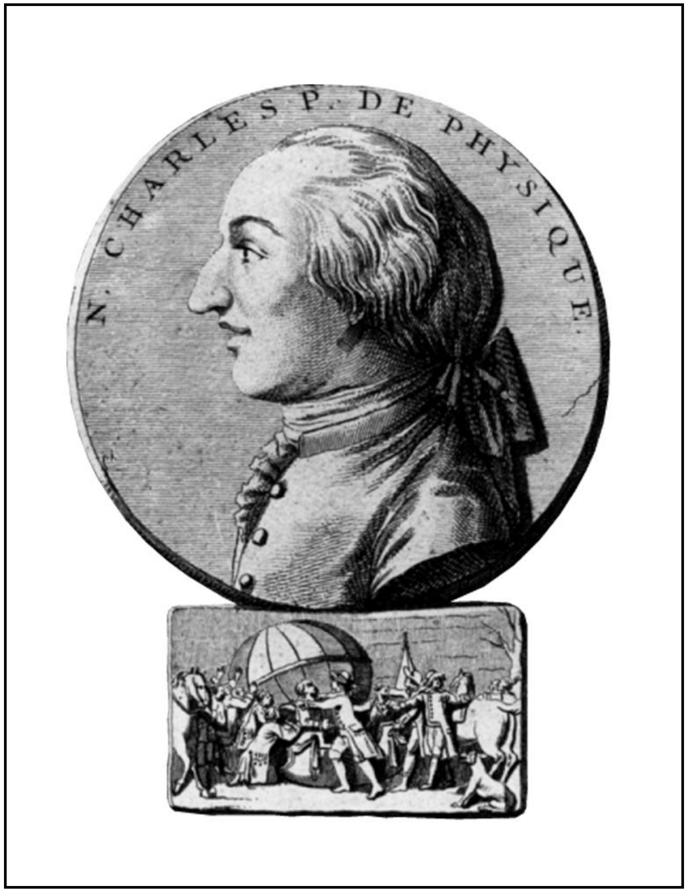


← Fig. 7.10
**Filling a syringe with a liquid is an application of Boyle's law.**

# Charles and gas laws

→ Fig. 7.11

**Jacques Charles in the process of working with hot-air balloons made the observations that led to the formulation of what is now known as Charles's law.**

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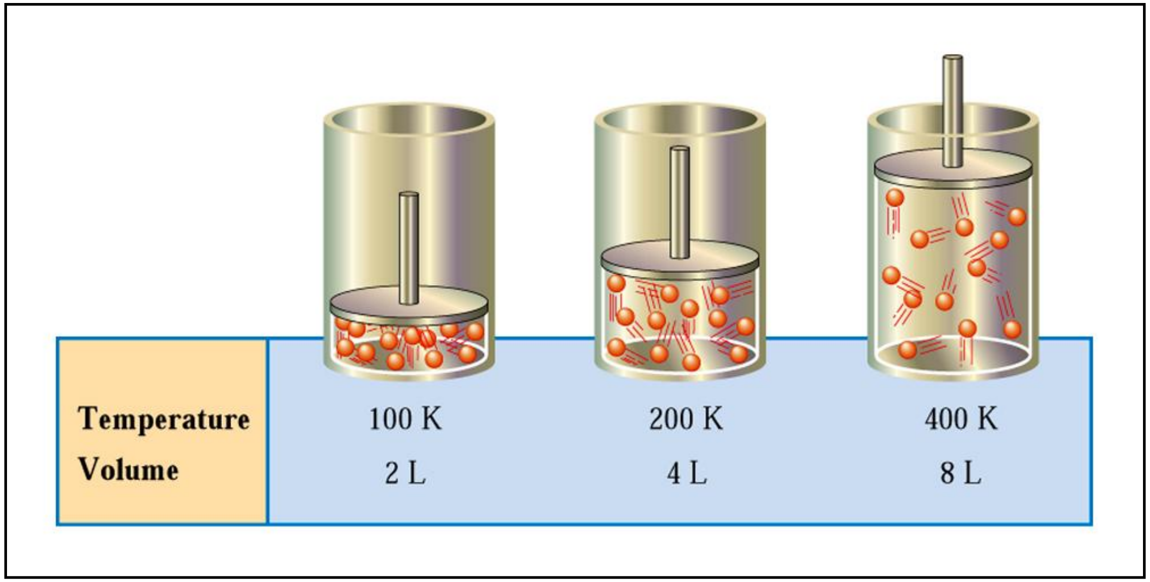
# Charles Law




← Fig. 7.12

**Data illustrating the direct proportionality associated with Charles's law.**

# Dalton and the Partial Pressure

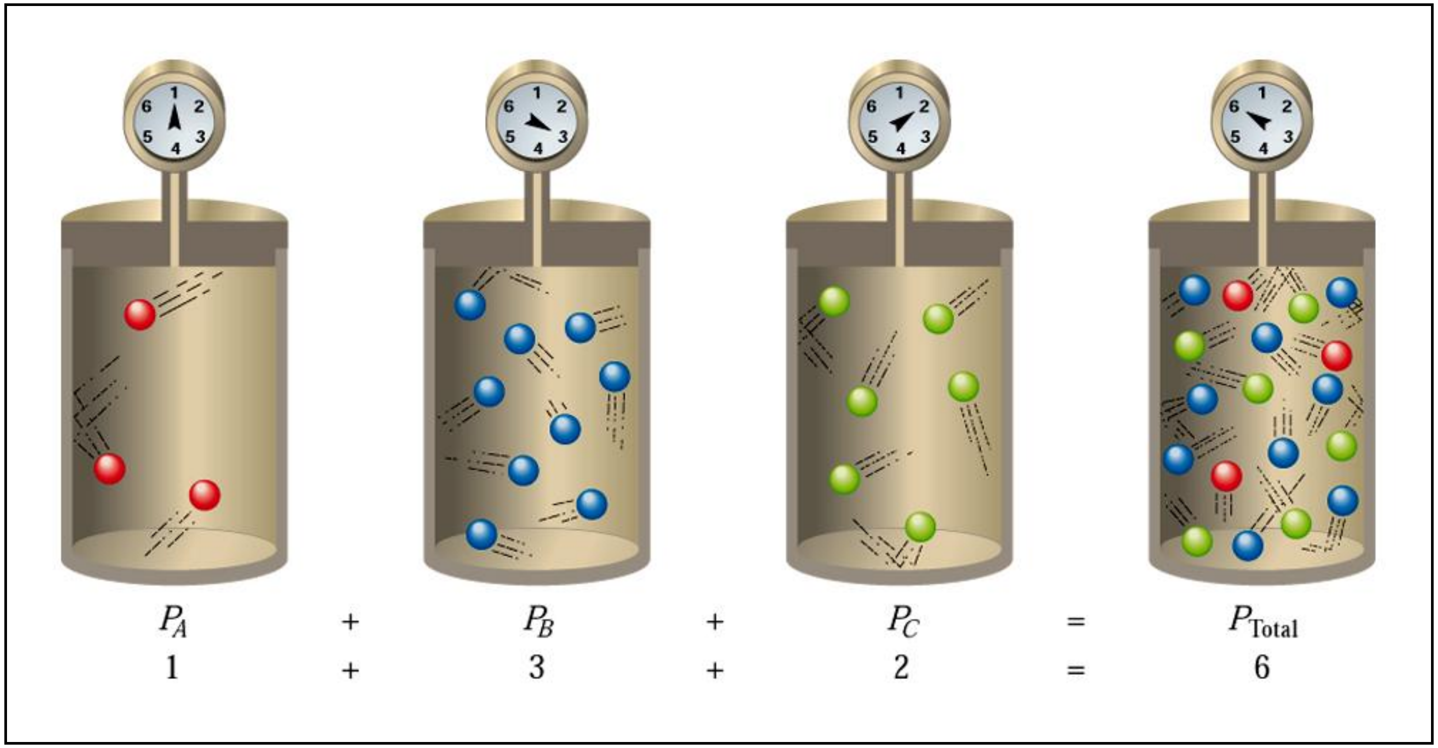
→ Fig. 7.13

**John Dalton had an interest in the study of weather.**

# Dalton's law of partial pressures

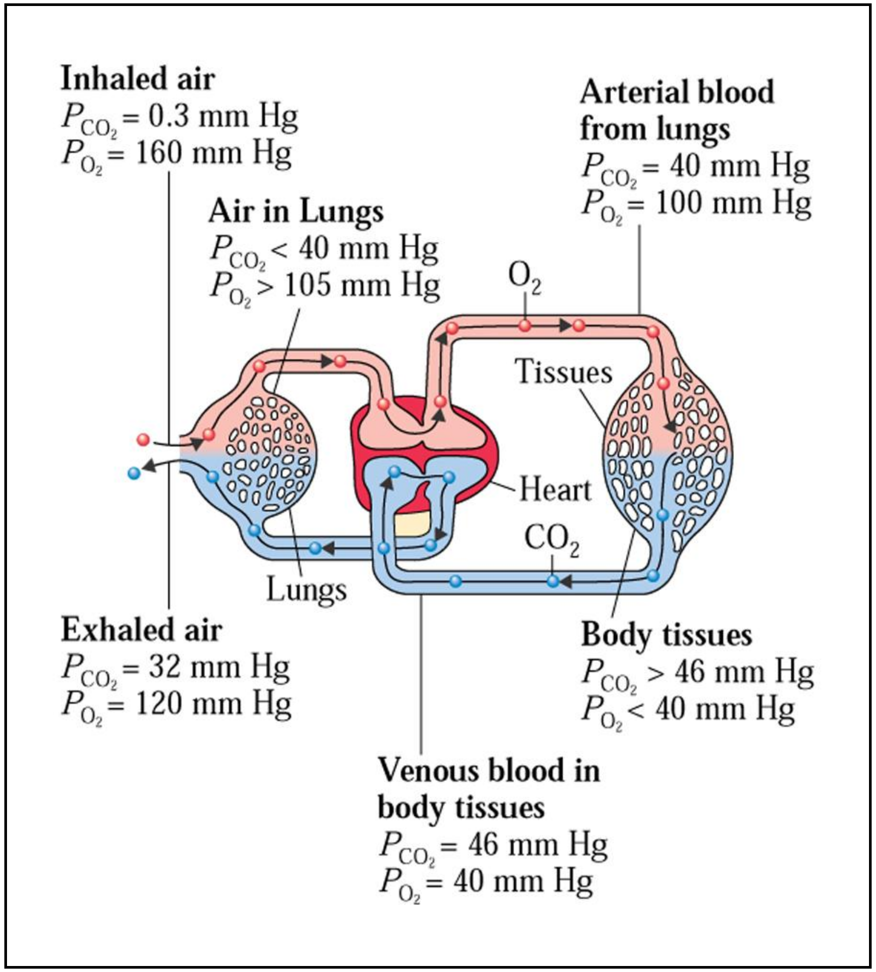
Fig. 7.14

**A set of four containers can be used to illustrate Dalton's law of partial pressures. The pressure in the fourth container equals the sum of the first three.**



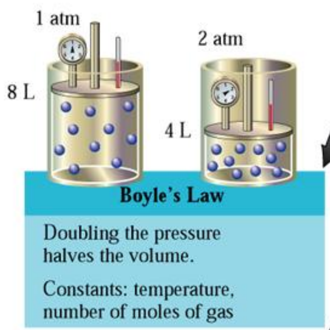
# Blood gases and their solubilty

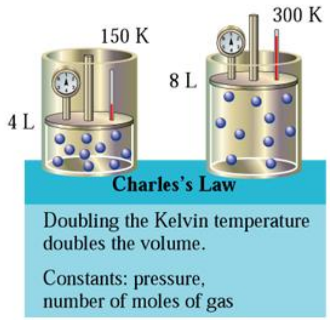
→ CC 7.2



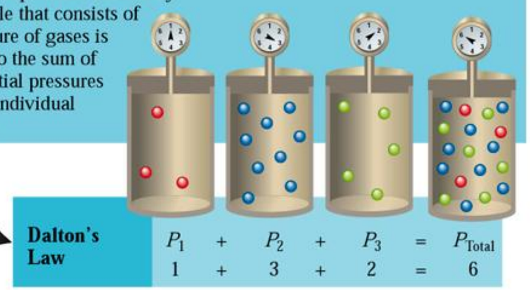
# Summary of gas laws

→ **CAG 7.1**



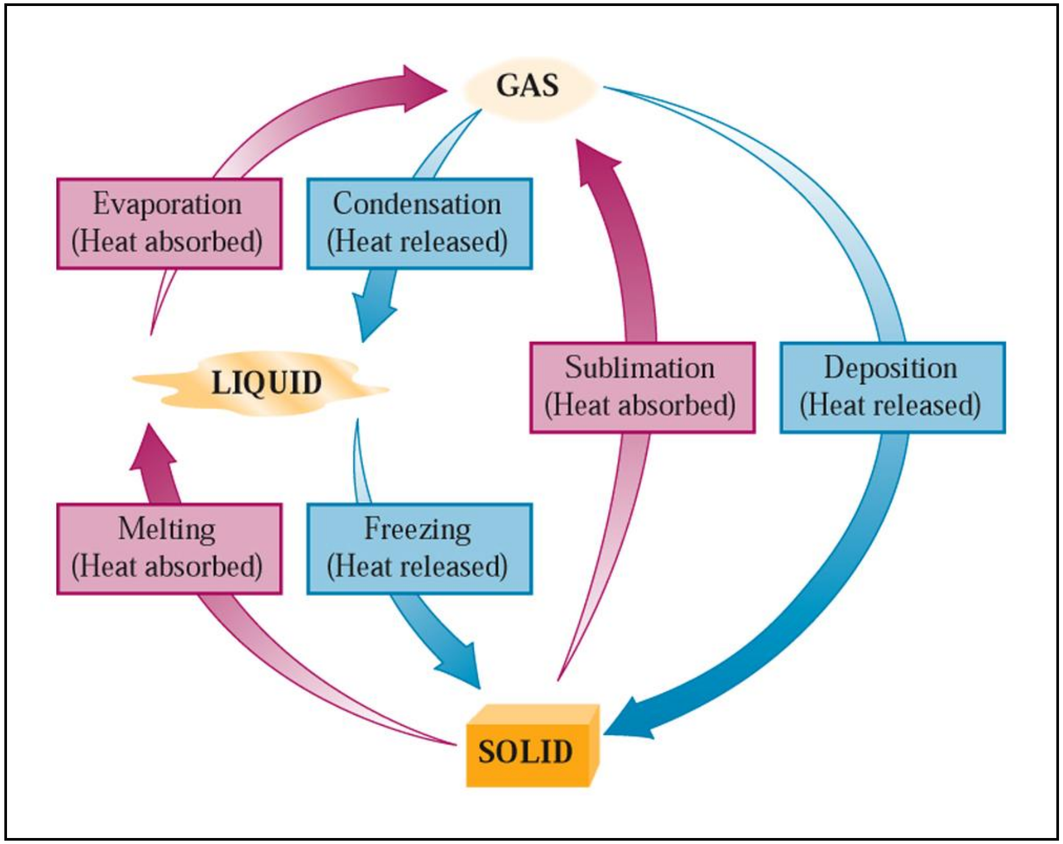

| GAS LAW | SYNOPSIS | CONSTANTS | VARIABLES |
|---|---|---|---|
| **Boyle's Law** $P_1V_1 = P_2V_2$ | At constant temperature, the volume of a fixed amount of gas is *inversely proportional* to the pressure applied to it. | temperature, number of moles of gas | pressure, volume |
| **Charles's Law** $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ | At constant pressure, the volume of a fixed amount of gas is *directly proportional* to its Kelvin temperature. | pressure, number of moles of gas | volume, temperature |
| **Combined Gas Law** $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ | The product of the pressure and the volume of a fixed amount of gas is *directly proportional* to its Kelvin temperature. | number of moles of gas | pressure, temperature, volume |
| **Ideal Gas Law** $PV = nRT$ | Relates volume, pressure, temperature, and molar amount of a gas under one set of conditions. If three of the four variables are known, the fourth can be calculated from the equation. | $R = 0.0821 \frac{L \cdot atm}{mole \cdot K}$ | pressure, volume, temperature, number of moles |
| **Dalton's Law** $P_{Total} = P_1 + P_2 + P_3$ | The total pressure exerted by a sample that consists of a mixture of gases is equal to the sum of the partial pressures of the individual gases. | | |



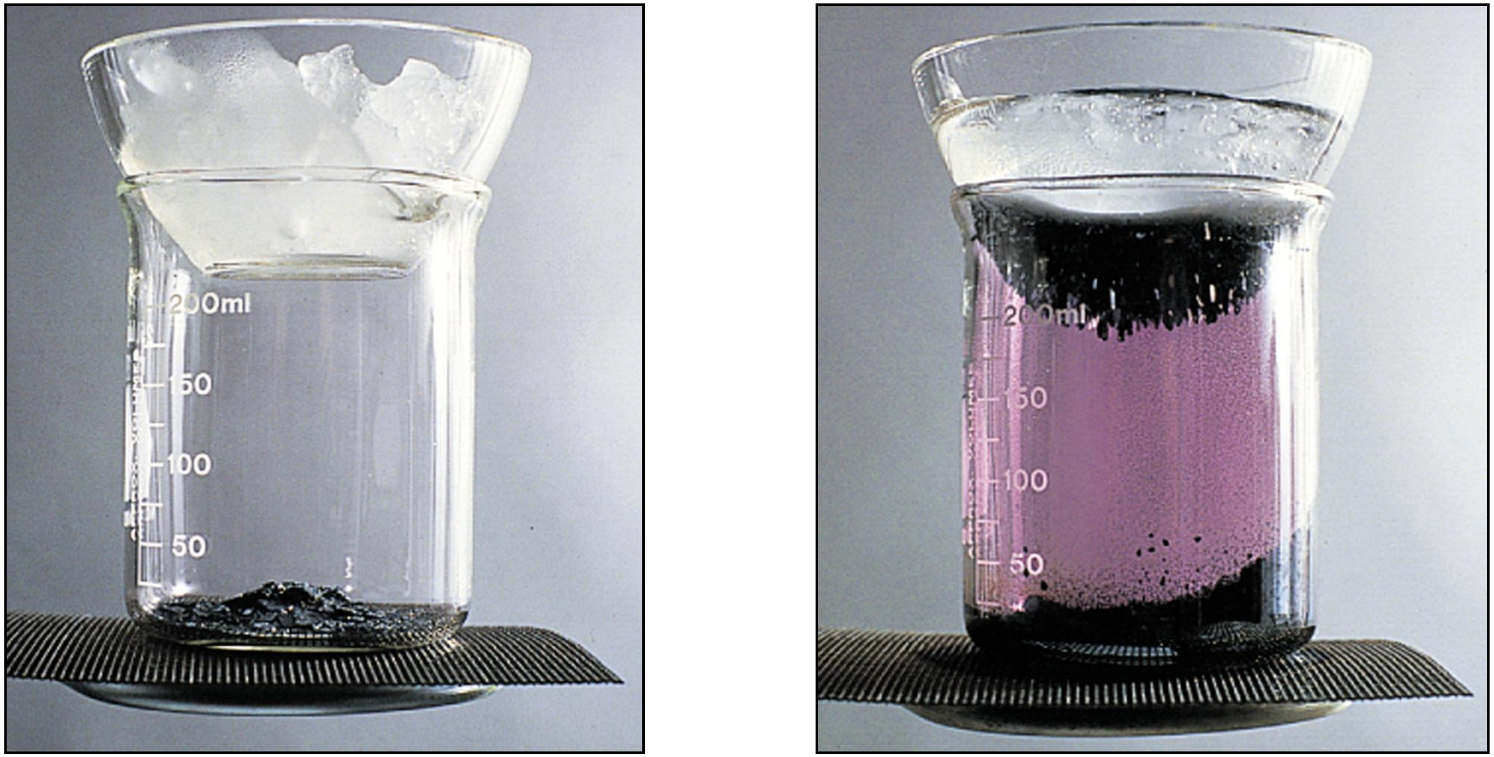
# Changes of states

→ Fig. 7.15

**There are six changes of state possible for substances.**



# Sublimation of Iodine




(a) The beaker contains iodine crystals.

(b) Iodine has an appreciable vapor pressure even below its melting point.

Fig. 7.15

# Vapor pressure and equilibrium

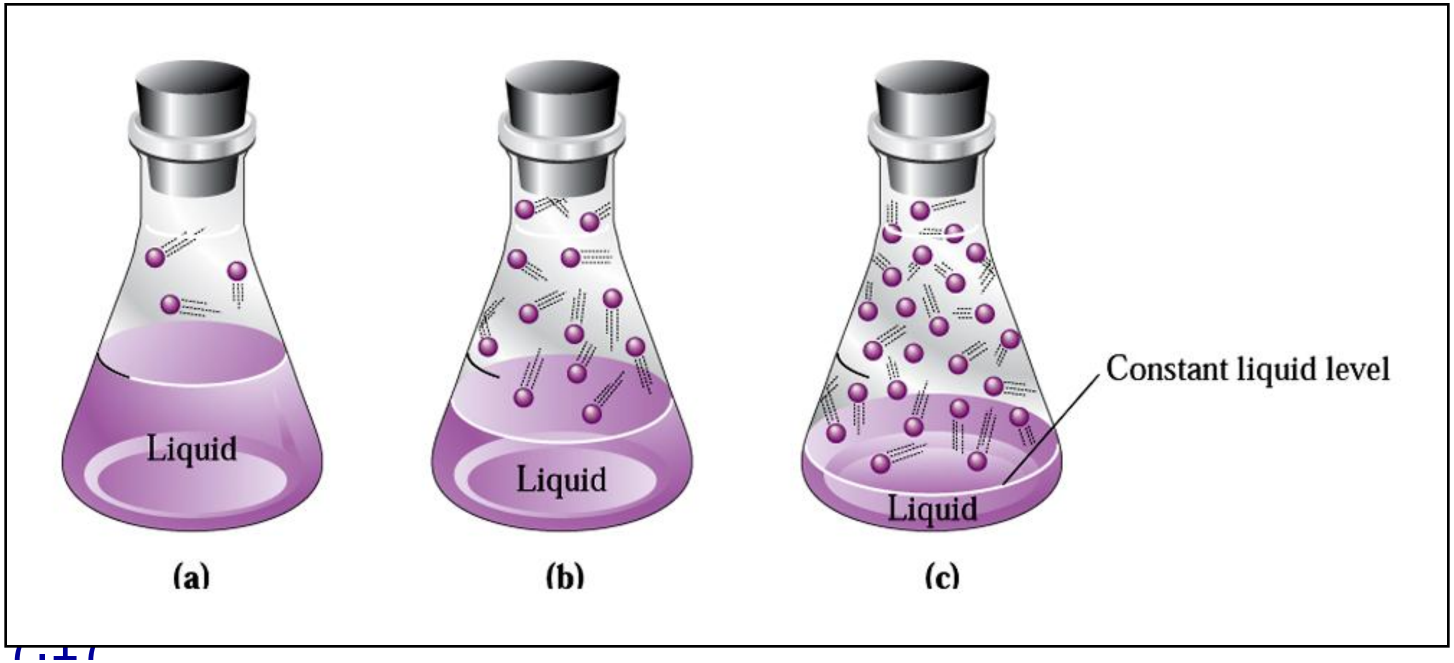



Fig. 7.17

**(a) the liquid level drops for a time, (b) then becomes constant. At that point a state of equilibrium has been reached in which (c) the rate of evaporation equals the rate of condensations.**

# Boiling point of liquids

→ Fig. 7.18

**Bubbles of vapor form within a liquid when the temperature of the liquid reaches the liquid's boiling point.**

# How you cook? Low or high pressure

Brian Bailey/Network Aspen

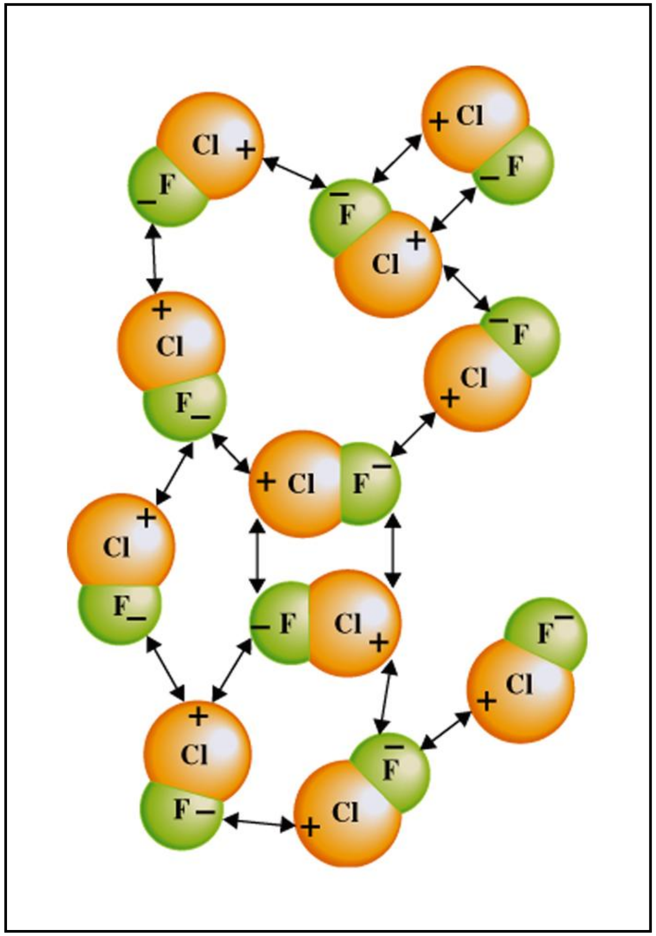
← Fig. 7.19

**The converse of the pressure cooker "phenomenon" is that food cooks more slowly at reduced pressure.**

# Intermolecular forces: Dipole-dipole

→ Fig. 7.20

**There are many dipole-dipole interactions possible between randomly arranged ClF molecules.**



# Vapor pressure and the temperature

| Temperature (°C) | Vapor Pressure (mm Hg) | Temperature (°C) | Vapor Pressure (mm Hg) |
|---|---|---|---|
| 0 | 4.6 | 50 | 92.5 |
| 10 | 9.2 | 60 | 149.4 |
| 20 | 17.5 | 70 | 233.7 |
| 25[a] | 23.8 | 80 | 355.1 |
| 30 | 31.8 | 90 | 525.8 |
| 37[b] | 37.1 | 100 | 760.0 |
| 40 | 55.3 | | |

[a]Room temperature
[b]Body temperature

Table 7.2

# Pressure cooker: how does it work?

| Pressure Cooker Setting (additional pressure beyond atmospheric, $lb/in.^2$) | Internal Pressure in Cooker (atm) | Boiling Point of Water (°C) |
|---|---|---|
| 5 | 1.34 | 108 |
| 10 | 1.68 | 116 |
| 15 | 2.02 | 121 |

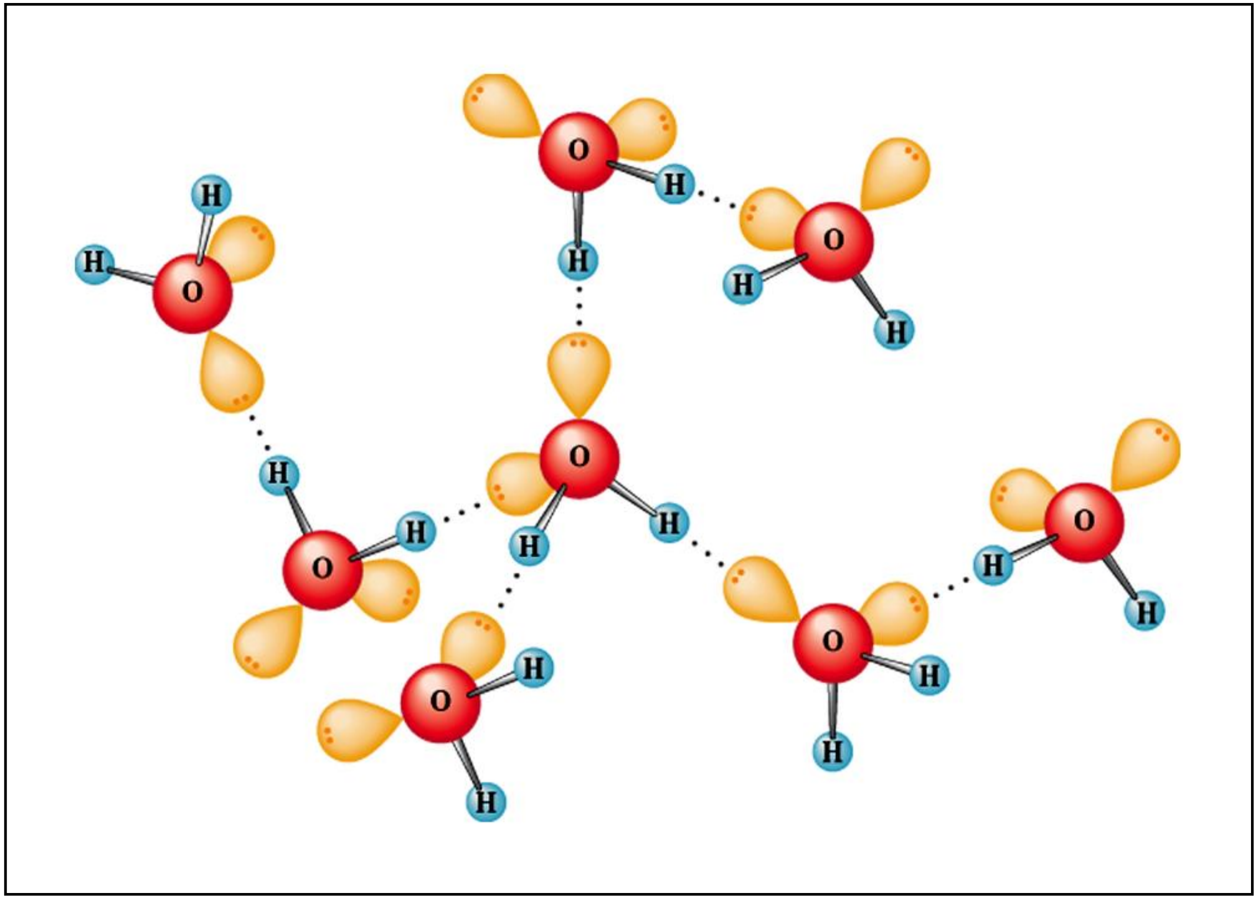
Table 7.4

# Boiling point and Eelvation

| Location | Feet Above Sea Level | $P_{atm}$ (mm Hg) | Boiling Point (°C) |
|---|---|---|---|
| top of Mt. Everest, Tibet | 29,028 | 240 | 70 |
| top of Mt. McKinley, Alaska | 20,320 | 340 | 79 |
| Leadville, Colorado | 10,150 | 430 | 89 |
| Salt Lake City, Utah | 4,390 | 650 | 96 |
| Madison, Wisconsin | 900 | 730 | 99 |
| New York City, New York | 10 | 760 | 100 |
| Death Valley, California | −282 | 770 | 100.4 |

Table 7.3

# Strongest intermolecular force: The hydrogen bonding



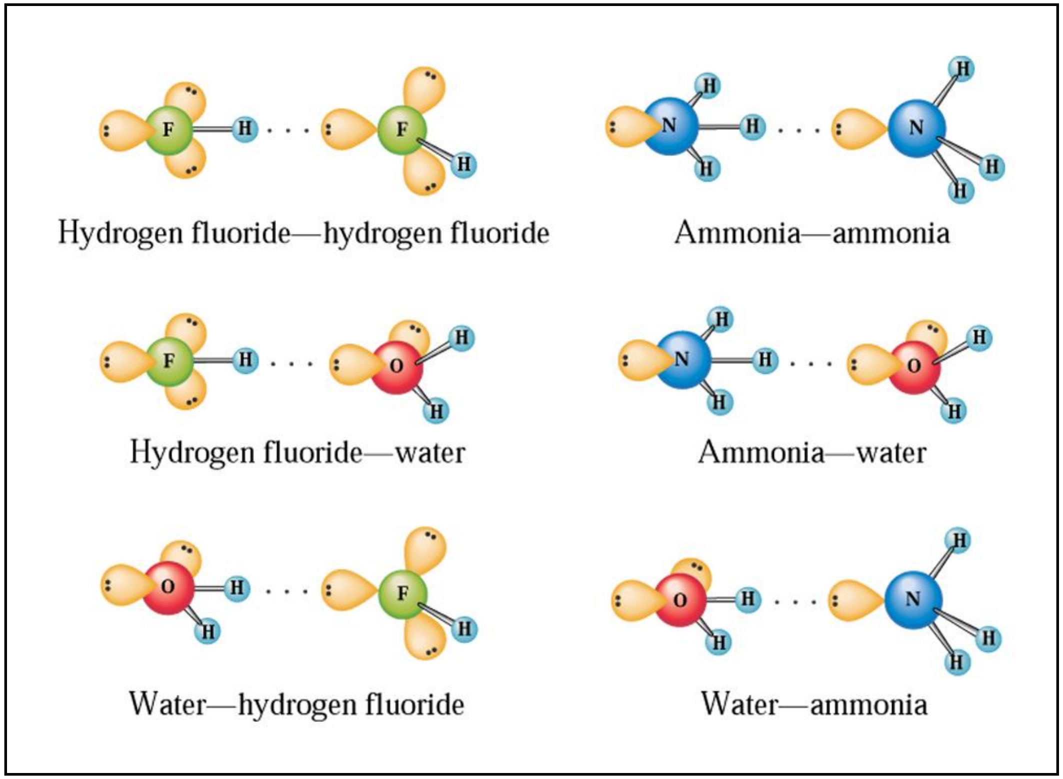

← Fig. 7.21

**Depiction of hydrogen bonding among the water molecules.**

# Three types of hydrogen bonding

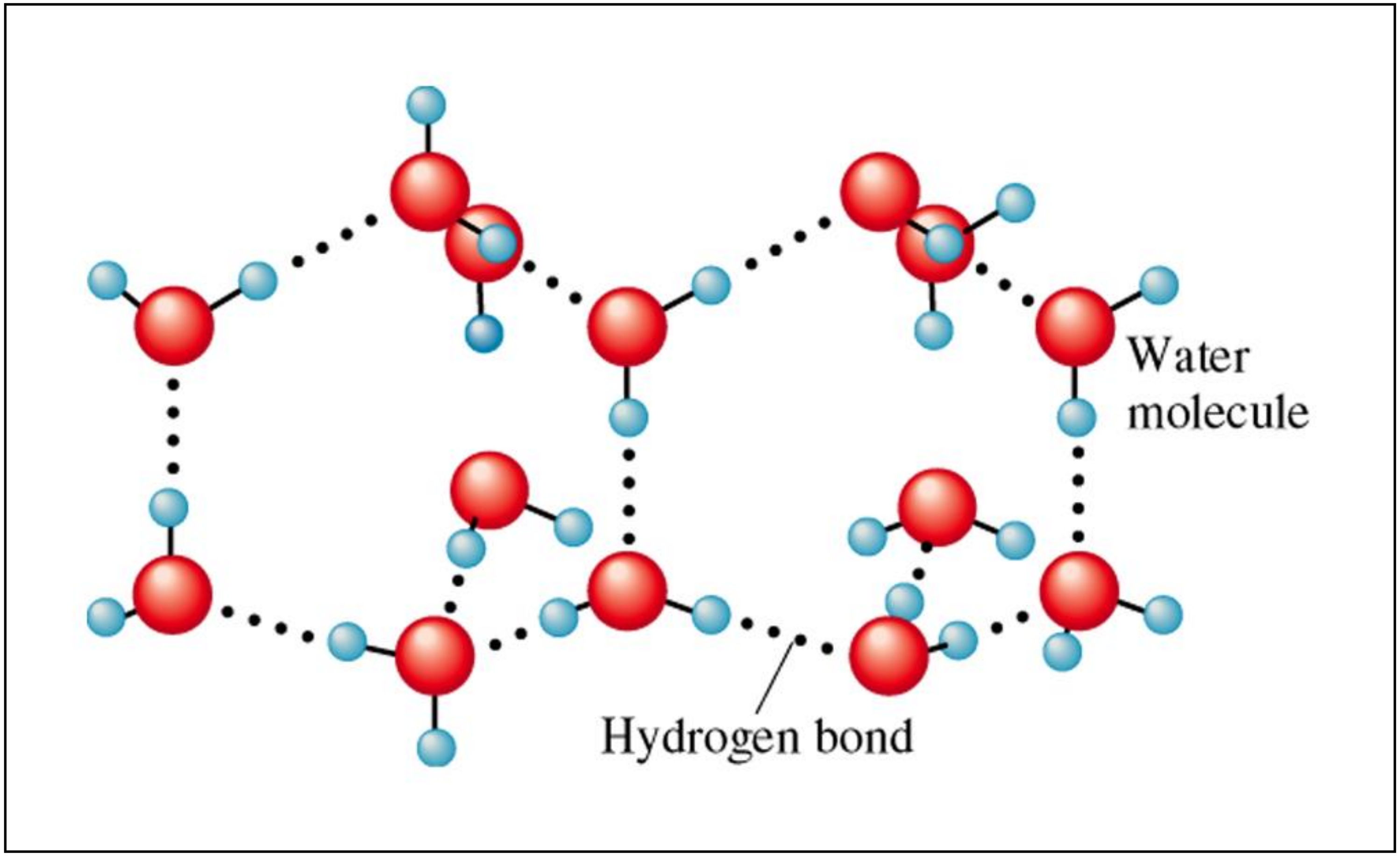
→ Fig. 7.22

**Diagrams of hydrogen bonding involving selected simple molecules.**

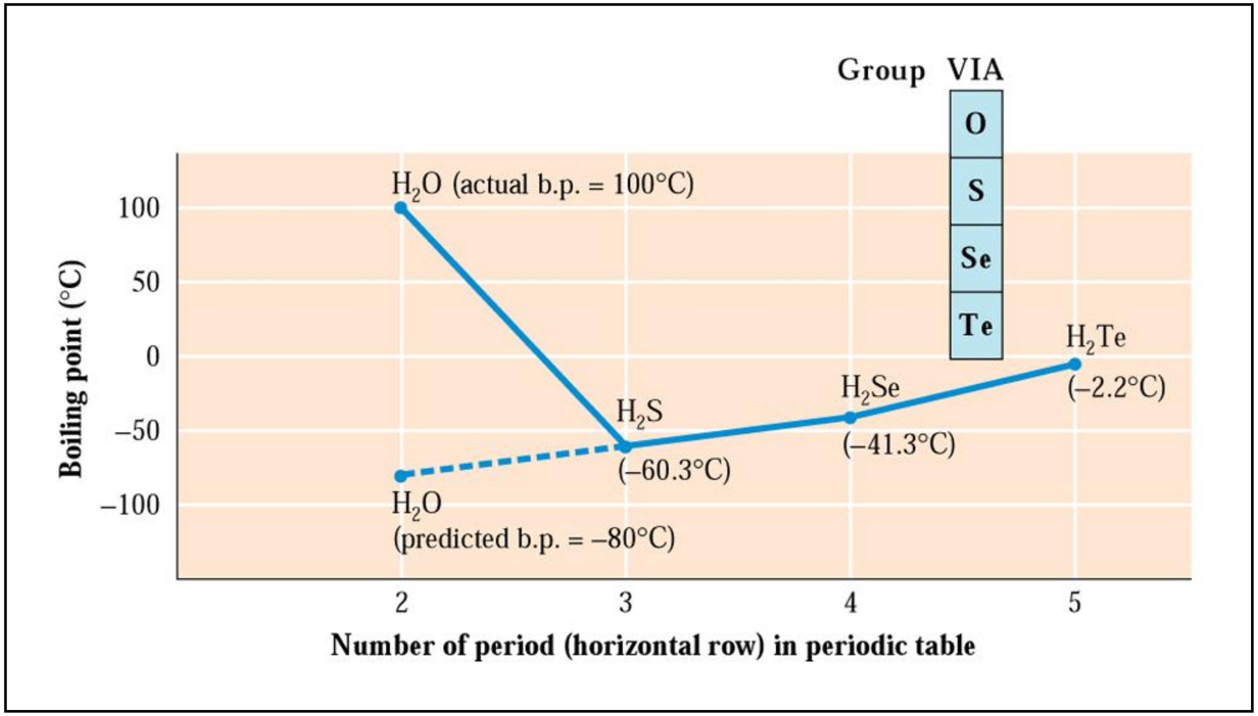


# Hydrogen bonding in ice

→ CC 7.3



# How does hydrogen bonding affect properties



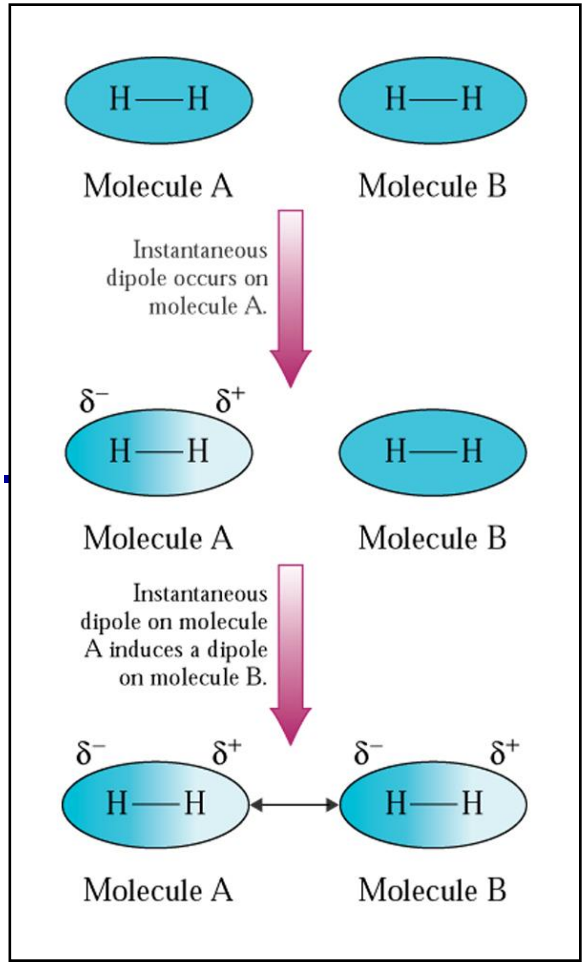

← Fig. 7.23

**If there were no hydrogen bonding between water molecules, the boiling point of water would be approximately -80C.**

# London Dispersion Forces:

→ Fig. 7.24

**Nonpolar molecules can develop instantaneous dipoles and induced dipoles.**



# Summary of intermolecular forces

→ **CAG 7.2**