Recap

- ❑ Vapor pressure vs. temperature
- ❑ Boiling point
- ❑ Unusual behavior of water
- ❑ P-T phase diagram

Phase diagram

Phase Diagrams: CO_2

A phase diagram is a map showing phases at different <u>pressures</u> and <u>temperatures</u>.

The region marked "**solid**" is the solid phase of ice. Other regions are marked "**liquid**" and "**vapor**."

This shows that carbon dioxide at 10°C and 2 atm will be a gas. At 10 atm and 10°C, CO_2 will condense to a liquid.

(10°C, 10 atm)

Solid Liquid

73

Liquid

5.1

Solid
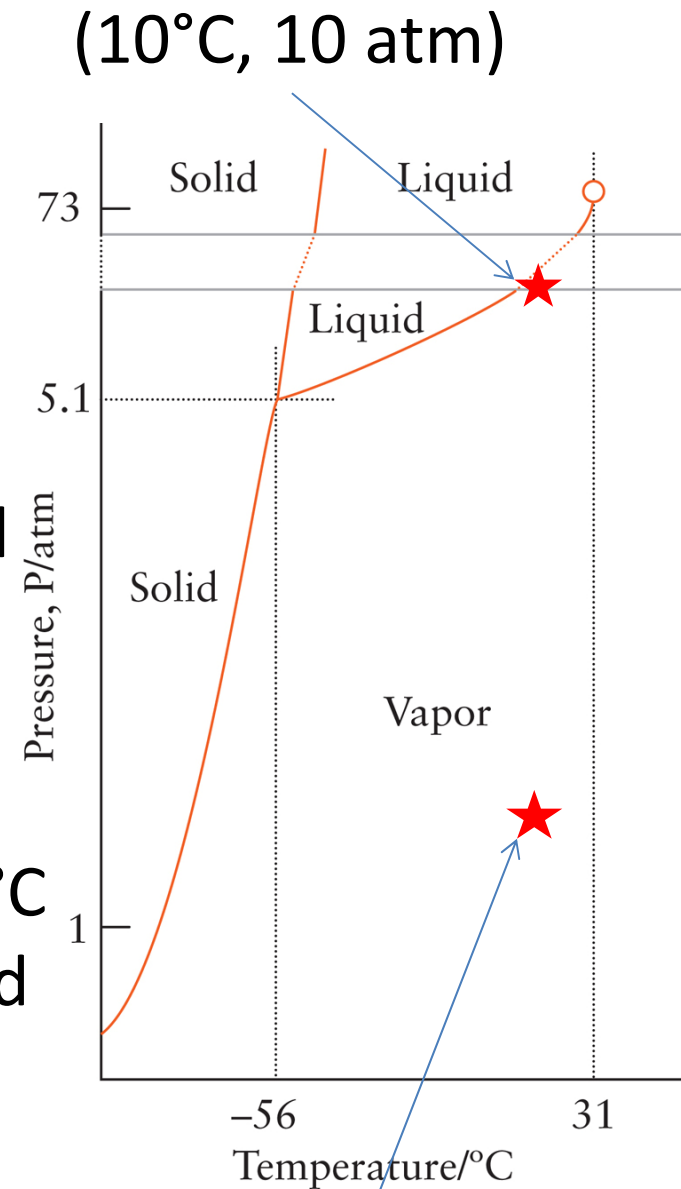
Pressure, P/atm

Vapor

1

−56   31

Temperature/°C

Figure 5B.3
Atkins, *Chemical Principles: The Quest for Insight*, 7e
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(10°C, 2 atm)   38

Phase Diagrams: Water

Point "C" is the <u>normal freezing point</u> (0.0°C, 760 Torr).

A triple point is the temperature and pressure at which water exists as a **solid**, **liquid**, and **vapor**.

The **slope** of the solid–liquid boundary is the *density* and for liquid water it is more dense than its solid. Therefore, **notice** that as we increase the pressure of ice, it will eventually be converted to liquid.
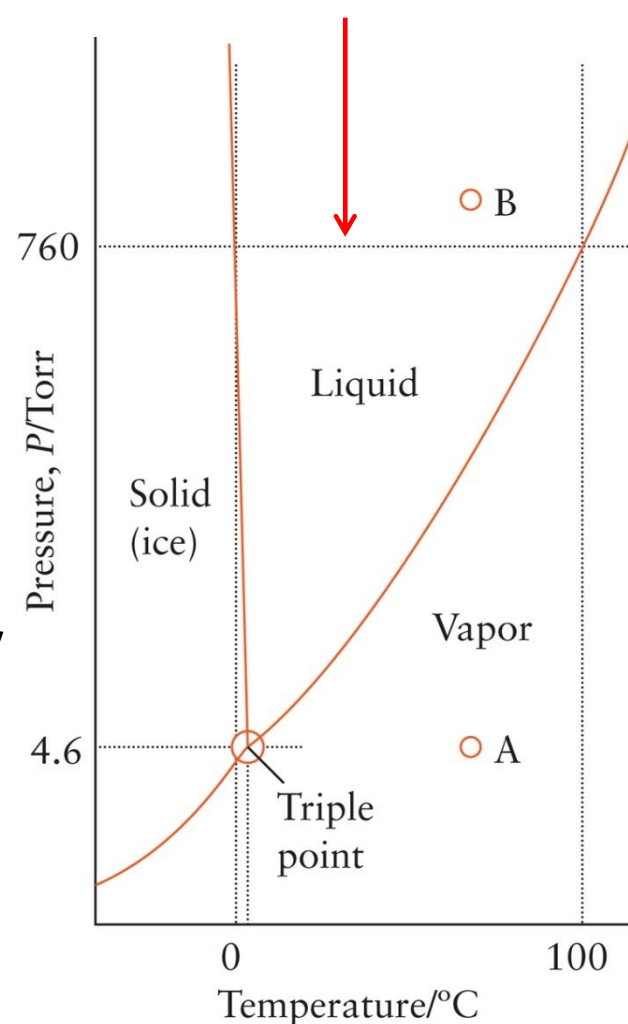

Normal boiling point.

Figure 5B.2
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From the phase diagram for carbon dioxide, predict which is more dense, the solid or the liquid phase. Explain your conclusion.
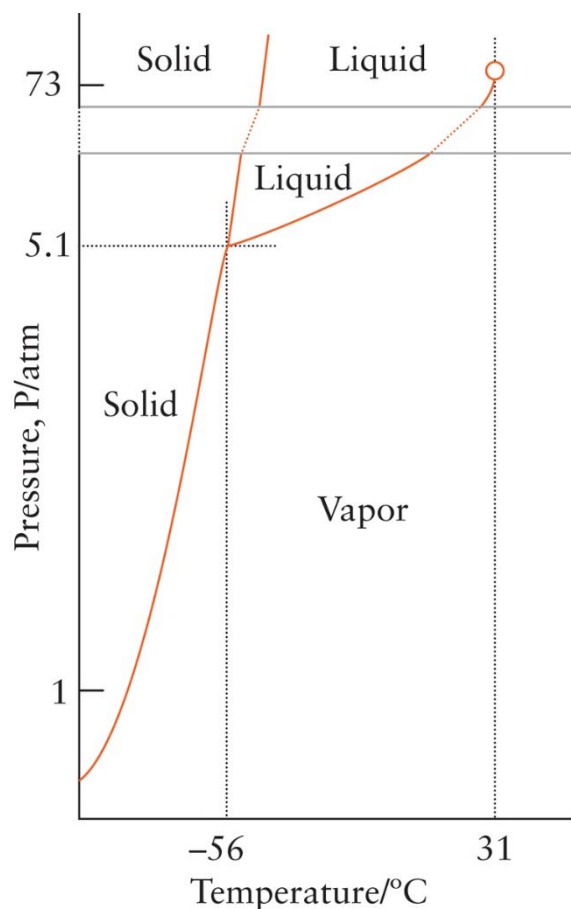

CO₂

Figure 5B.3
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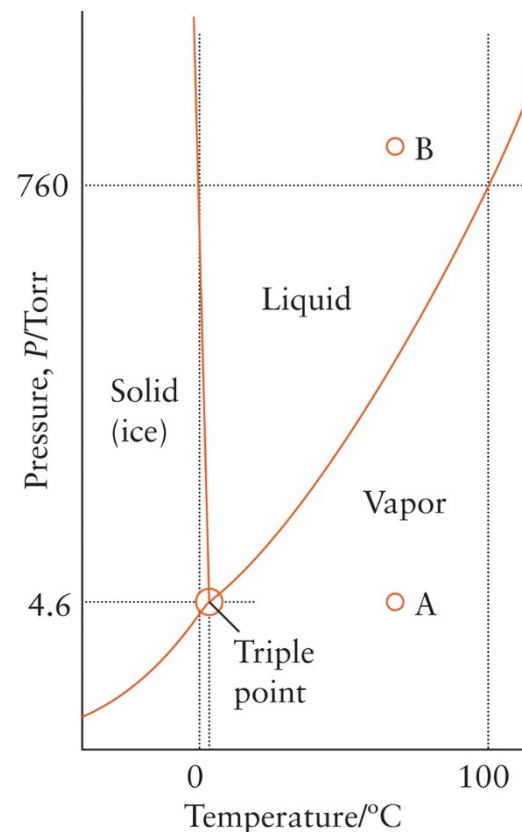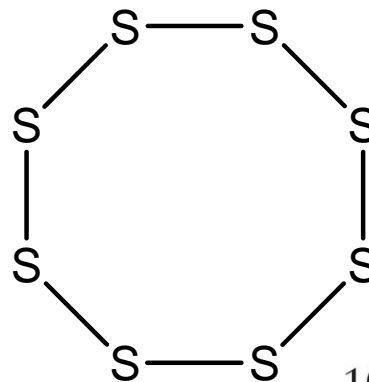

H₂O

Figure 5B.2
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The solid, because it is the stable phase at higher pressures.

The **slope** of the solid–liquid boundary (triple point) reveals that the density of liquid water is more dense than its solid.

Phase Diagrams: Sulfur, S_8

Sulfur has two solid phases rhombic and monoclinic, for the two ways its crownlike S_8 molecules stack.

Sulfur has three triple points, and four different phases.

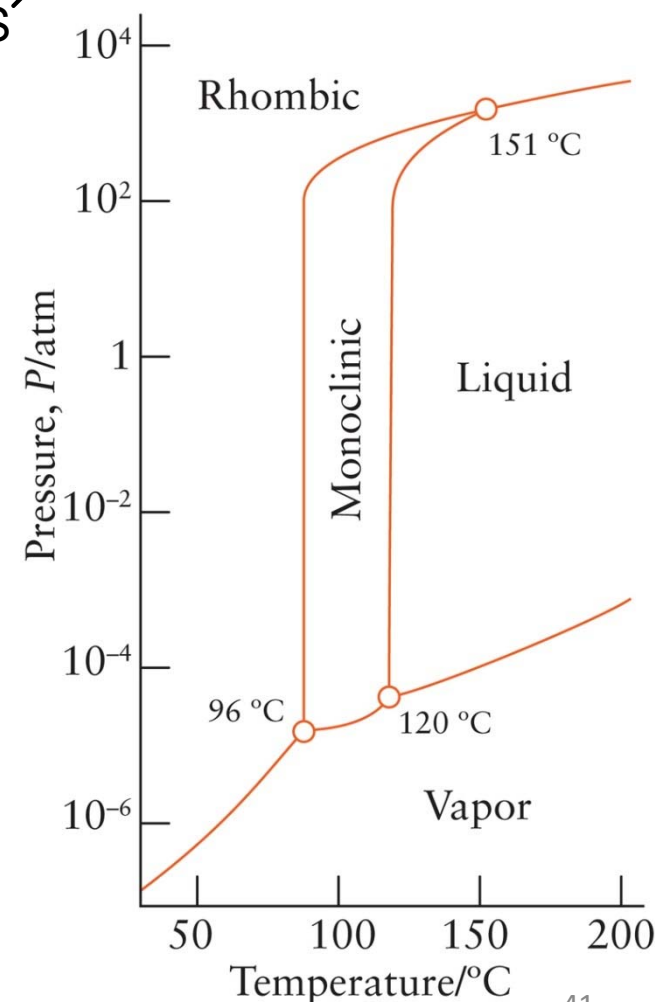Many substances have several solid phases. A "quadruple point" has never been observed.

Figure 5B.4
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Phase Diagrams: Waters Multiply Solid Phases

Water, like sulfur, has different phases.

The phase diagram of water drawn logarithmically, in order to show the different solid phases of water.

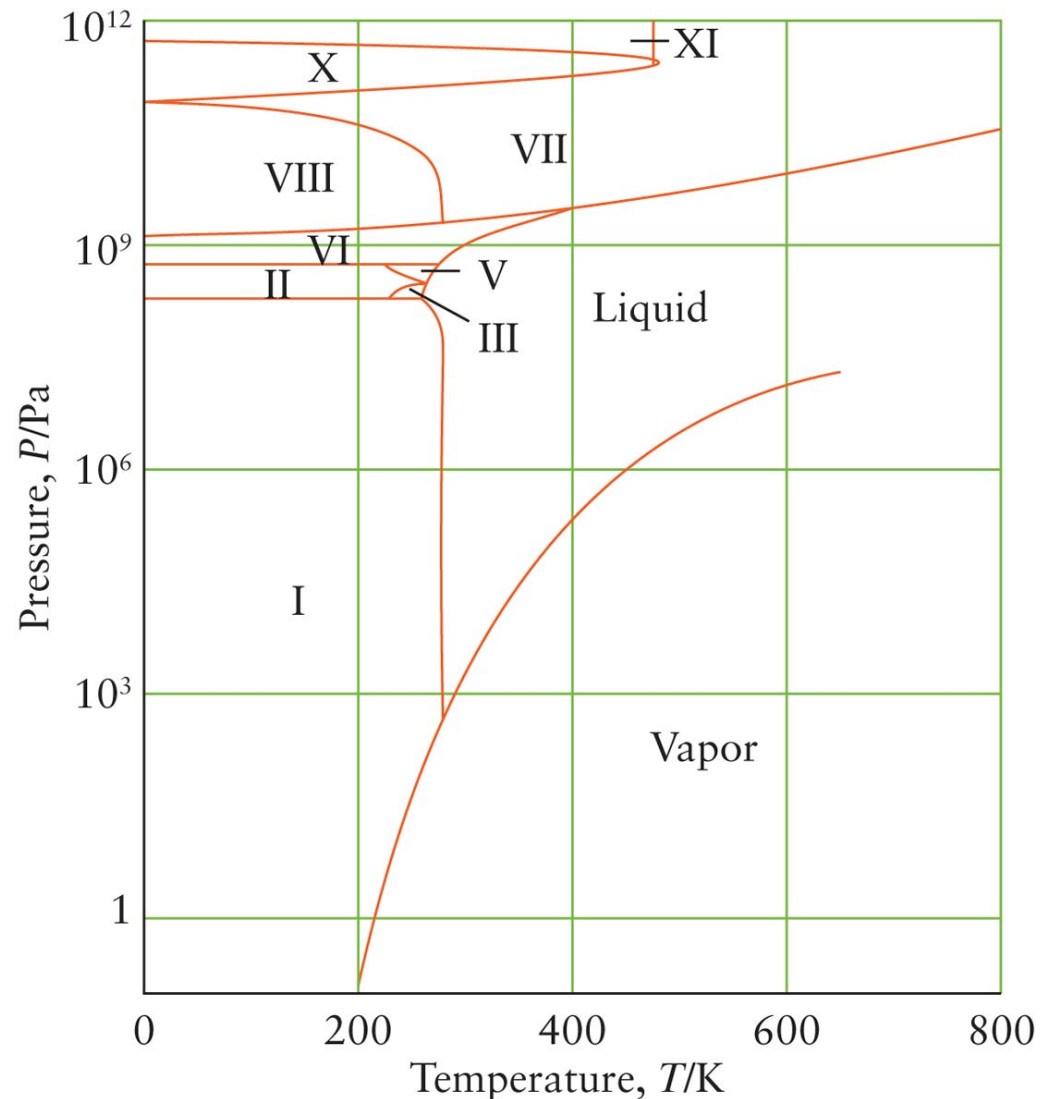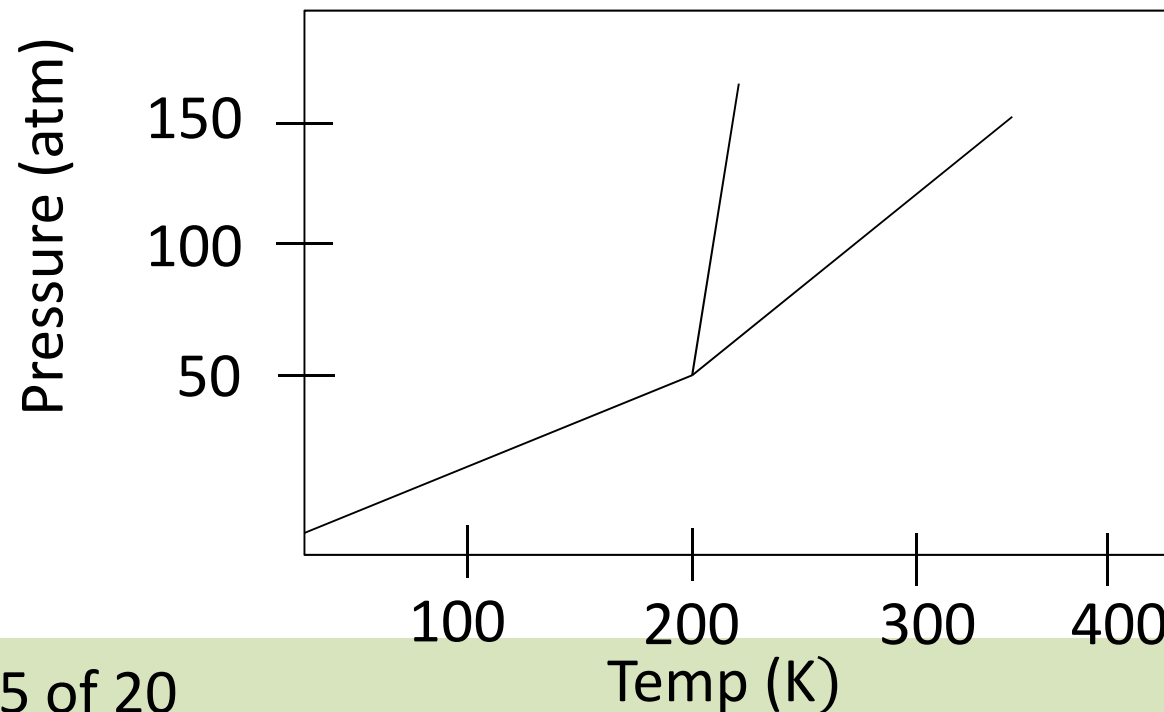Ice-VIII is stable only above 20,000 atm and 100°C.

Figure 5B.5
Atkins, *Chemical Principles: The Quest for Insight*, 7e
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The phase diagram for a pure substance is given below. The solid sublimes:

A. at 400 K and 200 atm.
B. at 300 K and 100 atm.
C. at 300 K and 75 atm.
D. if warmed at any pressure below 50 atm.
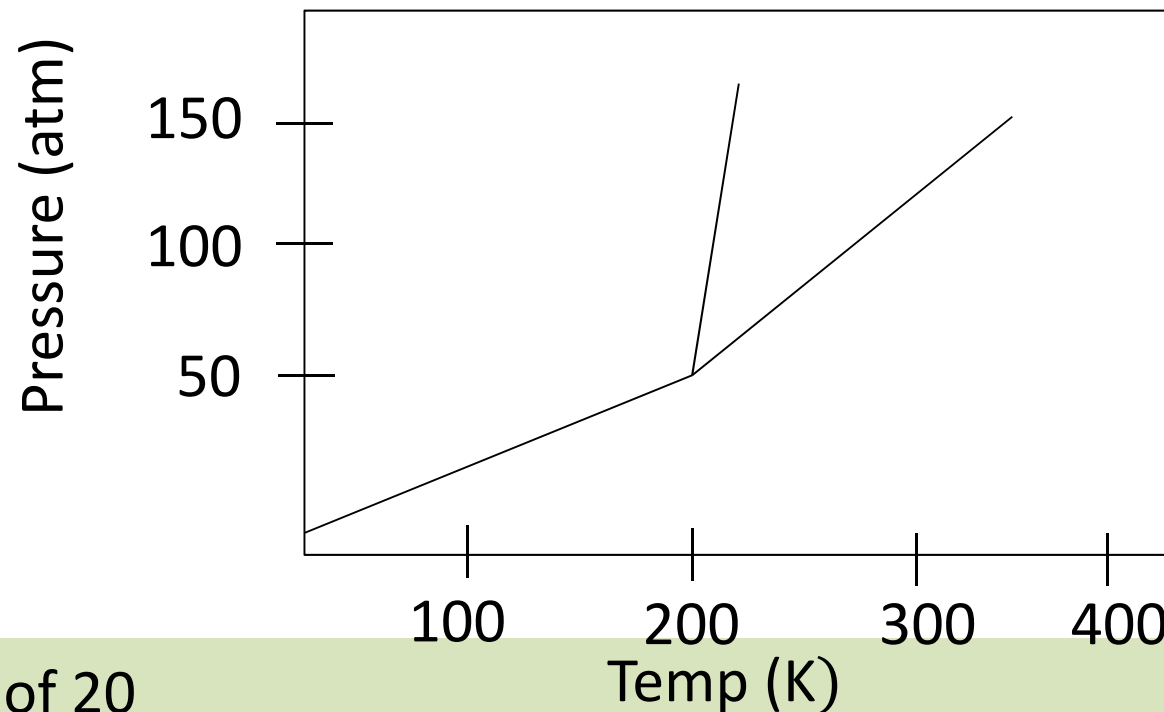
The phase diagram for a pure substance is given below. The solid sublimes:

A. at 400 K and 200 atm.
B. at 300 K and 100 atm.
C. at 300 K and 75 atm.
D. if warmed at any pressure below 50 atm.

Self-test 5B.1B From the phase diagram for sulfur, predict which phase is more dense, liquid sulfur or monoclinic sulfur. Explain your conclusion.

The positive slope indicates the monoclinic is more dense. Higher pressures indicate formation of a solid.
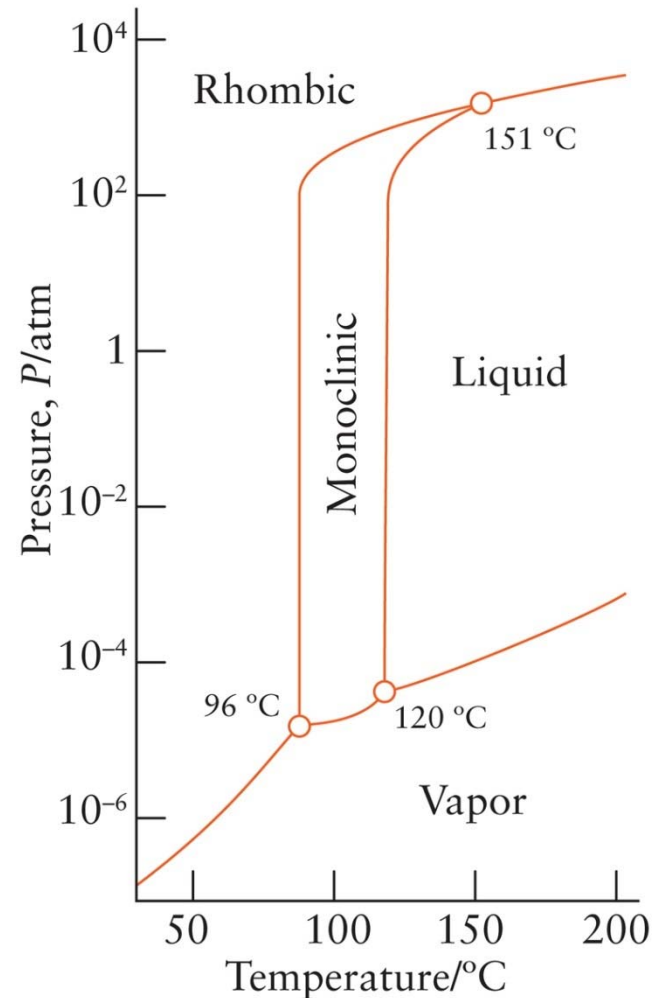
Example 5B.1 Use the phase diagram to describe the physical states and phase changes of water as the pressure on it is increased from 5 Torr (point A) to 800 Torr (point B) at 70°C.

Point A is at 5 Torr and 70°C which is a vapor. Increasing the pressure passes the liquid-vapor phase boundary, at which point liquid begins to form. At 800 Torr, point B, the phase is in the liquid region.
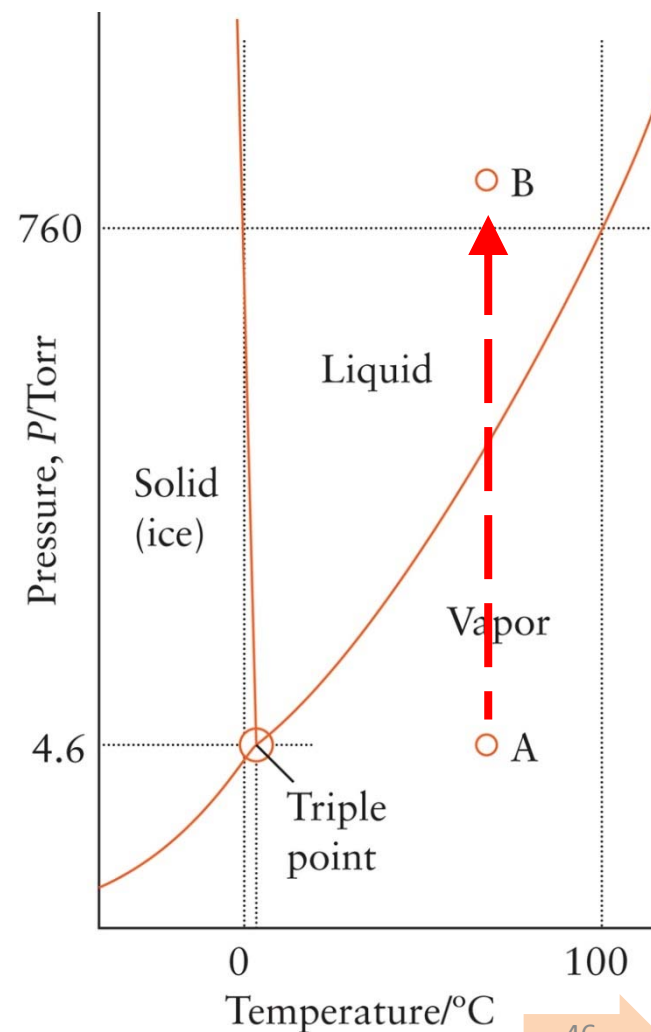

Figure 5B.2
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W. H. Freeman & Company, © 2016 by P. W. Atkins, L. L. Jones, and L. E. Laverman

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Critical Properties
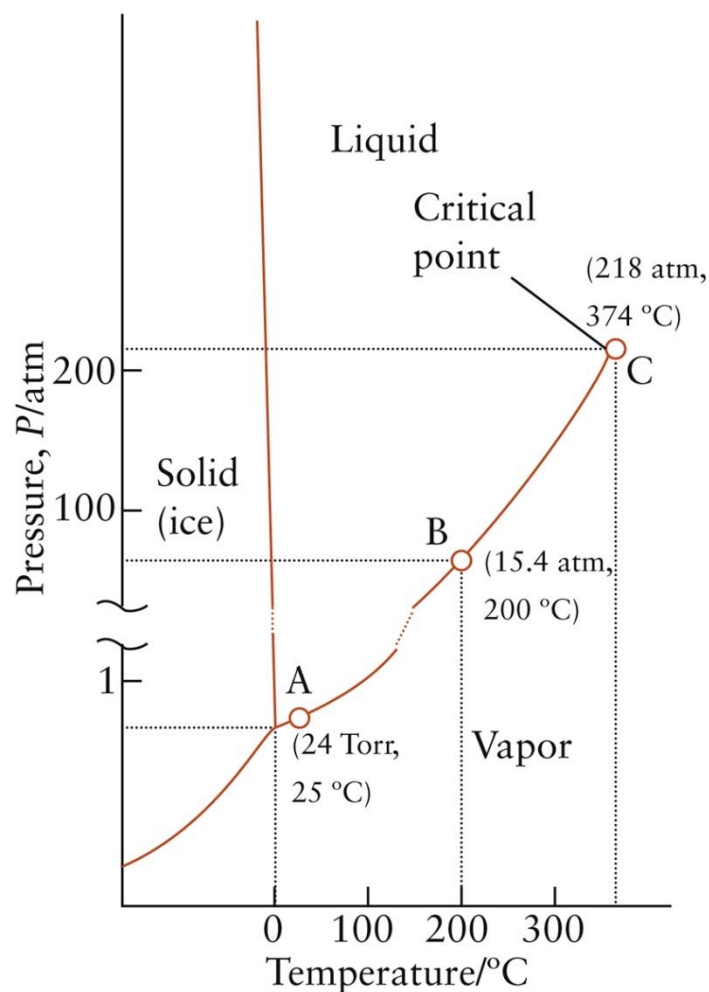
Figure 5B.8
Atkins, *Chemical Principles: The Quest for Insight*, 7e
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There is an end in the liquid–vapor phase boundary called the critical point.

The **density of the vapor** is so great that it is **equal** to the **density of the liquid**.

The **surface boundary disappears** into a single, **uniform phase**.

Here, the critical pressure and critical temperature mark the end of either liquid or vapor, and is now a **supercritical fluid**, a very dense fluid.

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Critical Point: vapor **density** = liquid **density**

Figure 5B.9
Atkins, *Chemical Principles: The Quest for Insight*, 7e
Courtesy Professor Walter Leitner, RWTH Aachen University, and Dr. Nils Theyssen, Max-Planck-Institut für Kohlenforschung, Mülheim/Ruhr, Germany.

Increasing the temperature of a liquid in a sealed container to a point where…

the liquid density **decreases** and the *vapor density* **increases**…

a single phase exists called the critical *pressure* and *temperature.*

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supercritical fluid region

Supercritical region
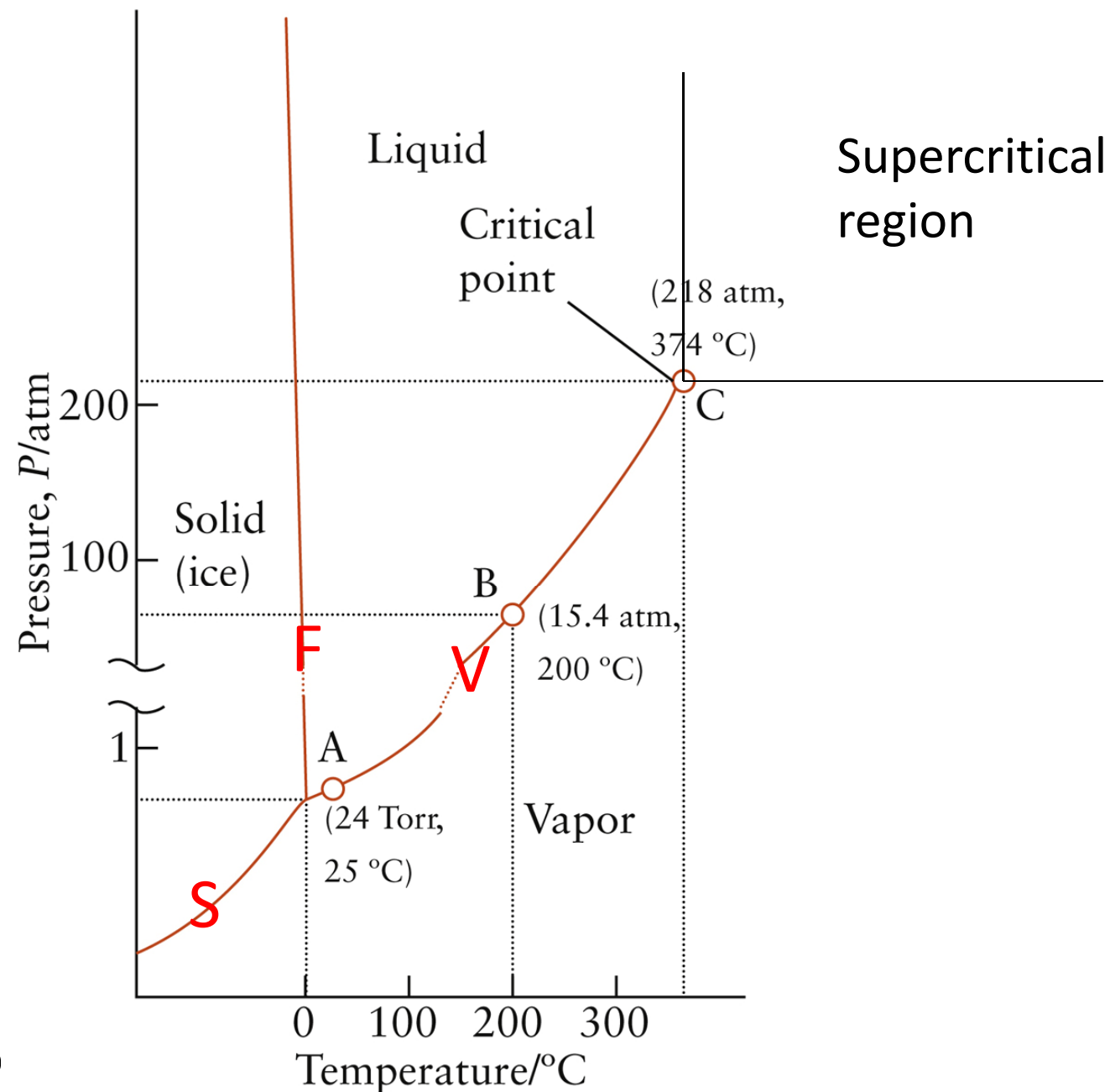
Liquid

Critical point

(218 atm, 374 °C)

C

Pressure, P/atm

200 —

Solid (ice)

100 —

B

(15.4 atm, 200 °C)

F

V

1 —

A

(24 Torr, 25 °C)

Vapor

S

F = ΔH fusion
V = ΔH vapor
S = ΔH sublimati

0   100   200   300

Temperature/°C

Self-test 5B.3B Identify trends in the effect of hydrogen bonding on the critical temperature.

For dipole molecules like H_2O and NH_3, a higher critical temperature and pressure indicate it takes **more energy** to force these types of molecules into a supercritical phase.

For nonpolar molecules, London forces dominate, and show that it takes **less energy** to put the molecules into a supercritical phase.

TABLE 5B.1	Critical Temperatures and Pressures of Selected Substances	
Substance	Critical temperature/ °C	Critical pressure P_c/atm
He	−268 (5.2 K)	2.3
Ne	−229	27
Ar	−123	48
Kr	−64	54
Xe	17	58
H_2	−240	13
O_2	−118	50
H_2O	374	218
N_2	−147	34
NH_3	132	111
CO_2	31	73
CH_4	−83	46
C_6H_6	289	49

Multiple-component system
(Binary)

Vapor-Pressure Lowering

The French scientist Francois-Marie Raoult, discovered that, at any temperature, the vapor pressure of a solvent is proportional to its mole fraction in a solution. Now known as Raoult's law,

$$P = \chi_{solvent} P_{pure}$$

P is the vapor pressure of the solvent, $\chi_{solvent}$ is the mole fraction of the solvent, and P_{pure} is the vapor pressure of the pure solvent.

Example 5C.1 Calculate the vapor pressure of water at 20°C in a solution prepared by dissolving 10 g of the nonelectrolyte sucrose, $C_{12}H_{22}O_{11}$, in 100. g of water.

Using $P = \chi_{solvent}P_{pure}$ to find P we need to find $\chi_{solvent}$

$$\chi_{solvent} = \frac{n_{solvent}}{n_{solute} + n_{solvent}}$$

P_{pure} is 17.54 Torr (Table 10.3)

342.3 g·mol^{-1} $C_{12}H_{22}O_{11}$ this is our solute, it is the lesser of the two, 18.015 g·mol^{-1} H_2O this is our solvent, it is the greater of the two.

½ 1%
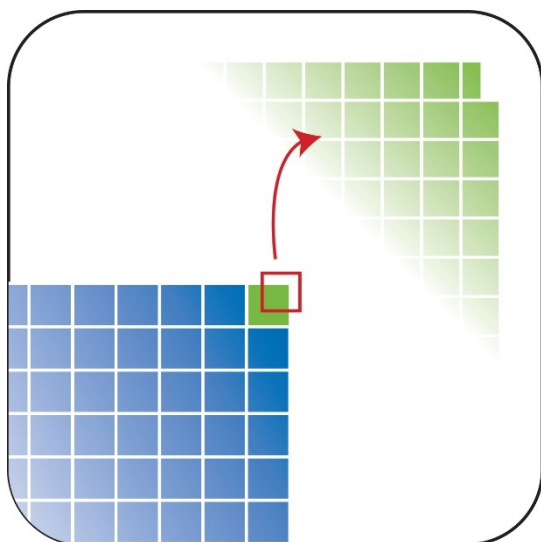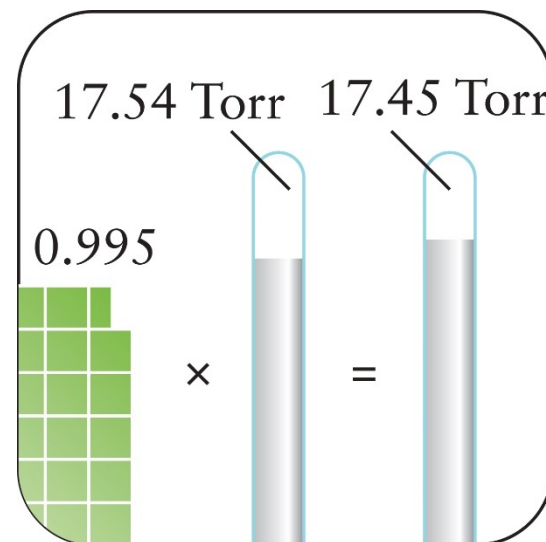
17.54 Torr   17.45 Torr

0.995

×   =

Example 5C.1 Calculate the vapor pressure of water at 20 °C in a solution prepared by dissolving 10. g of the nonelectrolyte sucrose, $C_{12}H_{22}O_{11}$, in 100 g of water.

$$n_{surose} = \frac{1 \text{ mol } C_{12}H_{22}O_{11}}{342.3 \text{ g}} \times \frac{10.00 \text{ g}}{1} = 0.02921 \text{ mol } C_{12}H_{22}O_{11}$$

$$n_{water} = \frac{1 \text{ mol } H_2O}{18.015 \text{ g}} \times \frac{100.0 \text{ g}}{1} = 5.551 \text{ mol } H_2O$$

$$\chi_{solvent} = \frac{5.551 \text{ mol } H_2O}{0.02921 \text{ mol } C_{12}H_{22}O_{11} + 5.551 \text{ mol } H_2O} = 0.995$$

$P = \chi_{solvent}P_{pure}$, $(0.995) \times 17.54$ Torr = 17.45 Torr   (not much of a change)

Introducing an ideal binary mixture of volatile liquids A and B.

Liquid A, benzene, C_6H_6, and B toluene ($C_6H_5CH_3$), when mixed are nearly ideal because these two compounds have similar molecular structures.

We can treat ideal mixtures using Raoult's law: $P = \chi_{solvent}P_{pure}$

$P_A = \chi_{A,liquid}P_{A,pure}$ and $P_B = \chi_{B,liquid}P_{B,pure}$

χ_A is the mole fraction and P_A is the vapor pressure of pure sample.

According to Dalton's law (Section 4.8), the total pressure of the vapor, P, is the sum or **total** of these two partial pressures:

$P = P_A + P_B =$ or P (above) $= \chi_{A,liquid}P_{A,pure} + \chi_{B,liquid}P_{B,pure}$
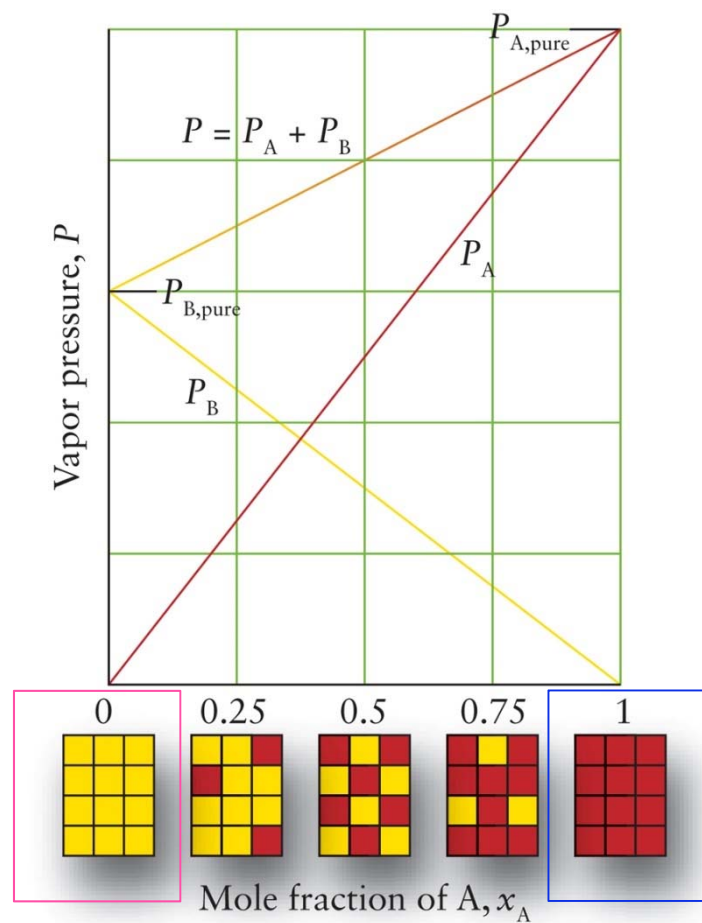
Figure 5C.2
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Pure solute

Pure solvent

Example 5C.2 What is the vapor pressure of each component at 25°C and the total vapor pressure of a mixture in which one-third of the molecules are benzene (so $x_{Benzene,liquid} = 1/3$ and $x_{Toluene,liquid} = 2/3$)? The vapor pressures of pure benzene and pure toluene at 25°C are 94.6 and 29.1 Torr, respectively.

$P_{Benzene} = 1/3 \, (94.6 \text{ Torr}) = 31.5 \text{ Torr}$
$P_{Toluene} = 2/3 \, (29.1 \text{ Torr}) = 19.4 \text{ Torr}$

Applying Dalton's law we can find the total pressure:

$P = P_{Benzene} + P_{Toluene} = 31.5 \text{ Torr} + 19.4 \text{ Torr} = 50.9 \text{ Torr}$

Raoult's Law Is Another Example of a Limiting Law

A hypothetical solution that obeys Raoult's law exactly at all concentrations is called an ***ideal solution***, and if it does not then it is a ***nonideal solution*** (a limiting law, $PV = nRT$ does not work outside of STP or during phase changes).

Real solutions are approximately ***ideal*** at concentrations below about **0.1 M for nonelectrolyte** (sugar) solutions and **0.01 M for electrolyte** solutions (ionic).

Binary Liquid Mixtures: Finding Quantitative Mole Ratios of Vapor

We were just asked to find the total pressure. What is the **composition** of that vapor? We can use $P_A = \chi_{A,\textbf{vapor}} \, P_T$ (Dalton's law)

We have: $P_A = {}_{solvent}P_{pure,A}$ (Raoult's) and $P_T = P_A + P_B$ (Dalton's law)

$$\chi_{A,\textbf{vapor}} = \frac{P_A}{P_T} = \frac{P_A}{P_A + P_B} \quad \text{(Note: this is vapor, not a liquid as before)}$$

Substitute for each sample $P_A = \chi_{A,liquid} \, P_{A,pure}$ (Raoult's)

$$\chi_{A,\textbf{vapor}} = \frac{\chi_{A,\textbf{liquid}} \, P_{A,pure}}{\chi_{A,\textbf{liquid}} \, P_{A,pure} + \chi_{B,\textbf{liquid}} \, P_{B,pure}}$$

Composition of vapor in terms of liquid.

We now have an expression that relates **composition of vapor** (**above the liquid**) in terms of **mole fractions of liquids**.

Example 5C.3 Find the mole fraction of benzene at 25°C in the **vapor** of a solution of benzene in toluene in which one-third of the molecules in the liquid are benzene (so $\chi_{Benzene,liquid} = 1/3$ and $\chi_{toluene,liquid} = 2/3$). The vapor pressures of benzene and toluene at 25°C are 94.6 and 29.1 Torr, respectively.

We're asked to find $\chi_{A,vapor}$ and $\chi_{B,vapor}$.
We know ratio of liquids $\chi_{A,liquid} = 0.333$, and $\chi_{B,liquid} = 0.667$

$$\chi_{A,vapor} = \frac{\chi_{A,liquid}\, P_{A,pure}}{\chi_{A,liquid}\, P_{A,pure} + \chi_{B,liquid}\, P_{B,pure}}$$

$$\chi_{benzene,vapor} = \frac{0.333 \times 94.6 \text{ Torr}}{0.333 \times 94.6 \text{ Torr} + 0.667 \times 29.1 \text{ Torr}} = 0.619$$

We note that for any mole fraction the ratio = 1, so $\chi_A = 1 - \chi_B$.
We can apply $\chi_A = 1 - \chi_{Benzene}$, 1- 0.619 = 0.381 or just work it out.
The mole fraction of Ben. in vapor is nearly twice that in the liquid.

The normal boiling point is at 1 atm. Heating pure benzene it would boil at 80.1°C, and toluene at 110.6°C.

A temperature–composition diagram for benzene and toluene.

Vapor

The **vapor pressure** as well as the **boiling point** of the mixture will be <u>intermediate</u> between the two pure liquids.

Thermometer

Water out

Condenser

Solution   Water in

Heat

Distillate
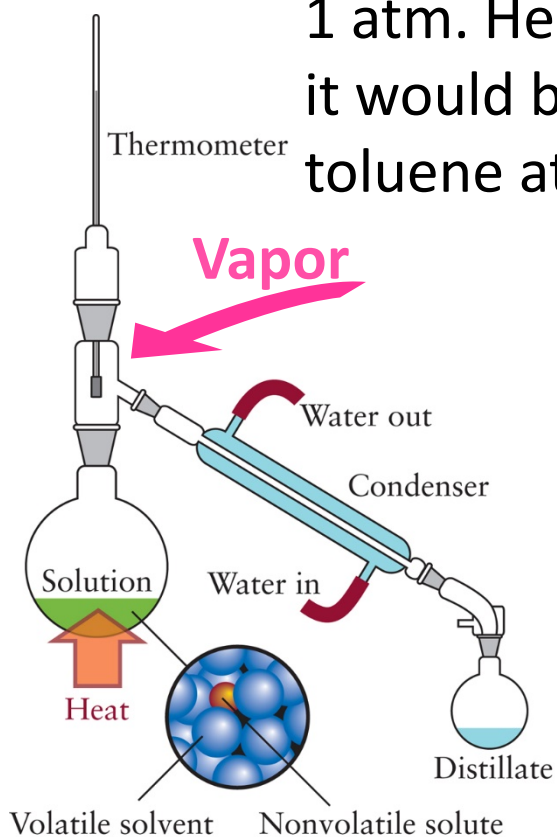
Volatile solvent   Nonvolatile solute

Figure G.6
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110.6

Mole fraction in vapor, $x_{benzene}$(g)

Temperature/°C

Boiling point

Tie line

A   B

80.1

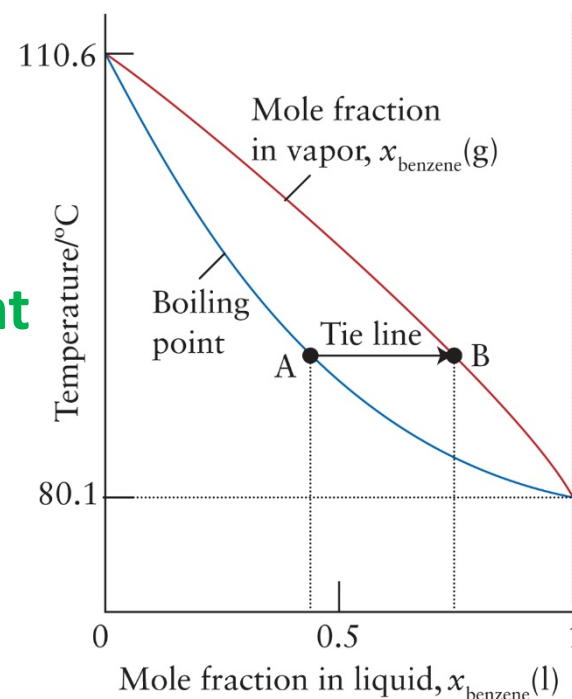0   0.5   1
Mole fraction in liquid, $x_{benzene}$(l)

Figure 5C.5
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This is called a temperature–composition diagram; the lower curve shows how the normal boiling point of the mixture varies with the composition. To find the *composition of the vapor*, we simply look along the **tie line**.

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Boiling two liquids in a distillation process: benzene and toluene.

When **liquid** benzene $x_{benzene} = 0.20$ (point A), it has a composition of **vapor** $x_{benzene} = 0.45$ (point B).

Condensing that vapor at B into a liquid is a distillate (because it was distilled) and it is richer in benzene than toluene.

Taking B distillate and **reheating** it, then condensing it, takes us to point C, where the vapor above the boiling solution now has a composition D and $x_{benzene} = 0.73$ is now more concentrated.
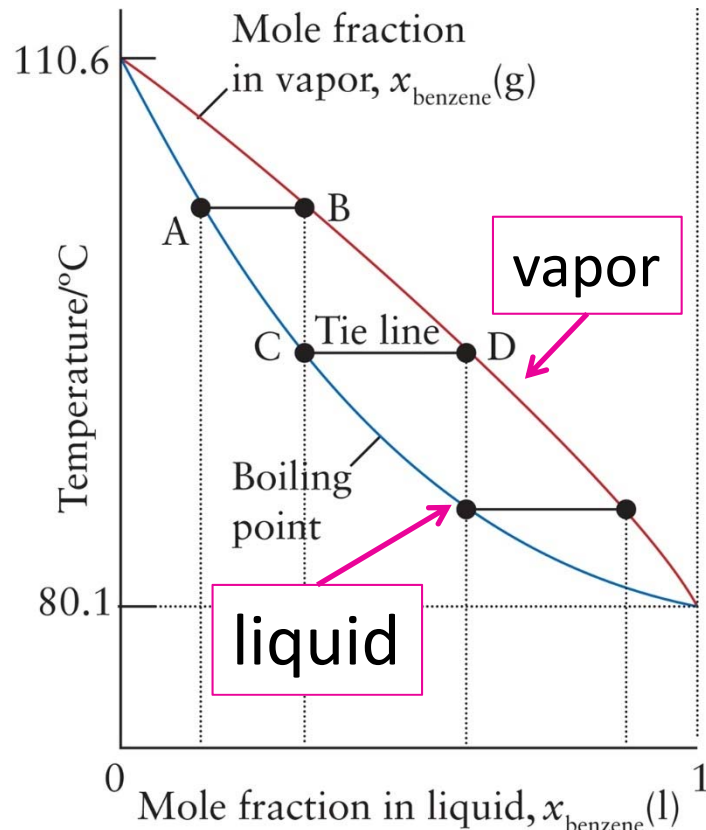

Figure 5C.6
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Distillation: During a Distillation

Mole fraction in vapor, $x_{benzene}(g)$

vapor

Tie line

Boiling point

liquid

Temperature/°C

Mole fraction in liquid, $x_{benzene}(l)$
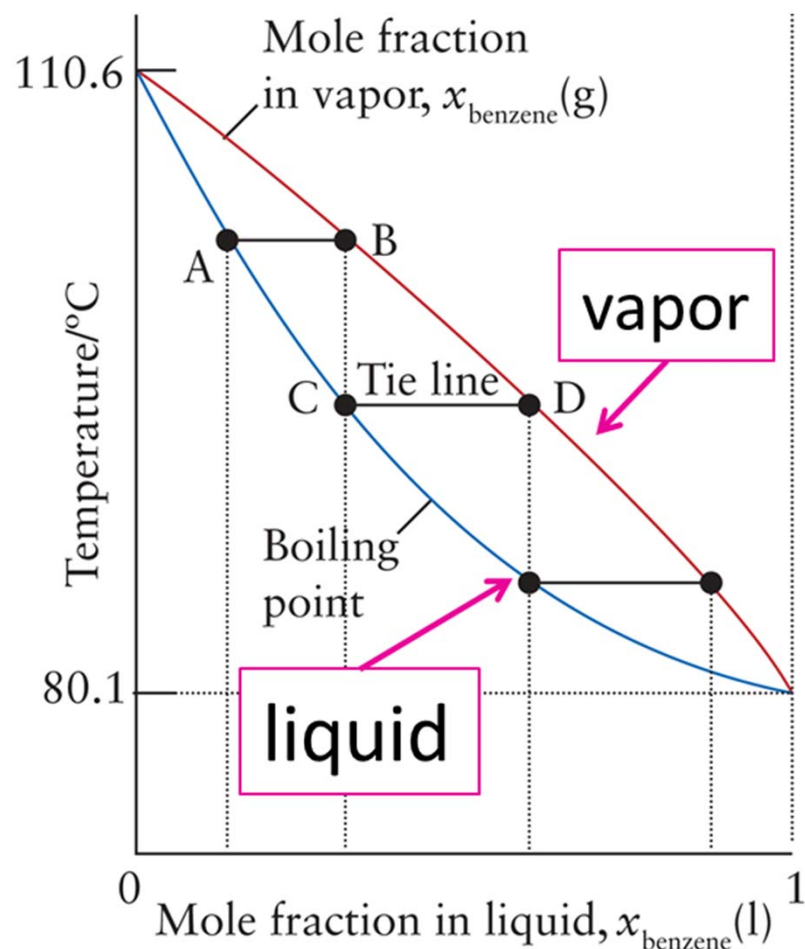
110.6

80.1

0

1

A

B

C

D

Figure 5C.6
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Note that the distillate from this **second stage** of distillation <u>is</u> <u>richer in benzene</u> than the distillate from the first stage.

If we continued these steps of boiling, condensation, and boiling again, we would eventually obtain a *very tiny amount* of <u>nearly pure</u> <u>benzene</u>. This is called **fractional distillation**.

Distillation: Fractional Distillation

Fractional distillation is a continuous redistillation, using tall columns packed with high surface area glass beads.

Higher BP vapor condenses and vaporizes over and over as it rises; the lower BP liquid drips back into the boiling mixture.

Vapor becomes richer in the component with the lower boiling point.

The final distillate is nearly pure, and the liquid in the pot is pure in the lower BP liquid.

Low-boiling-point component
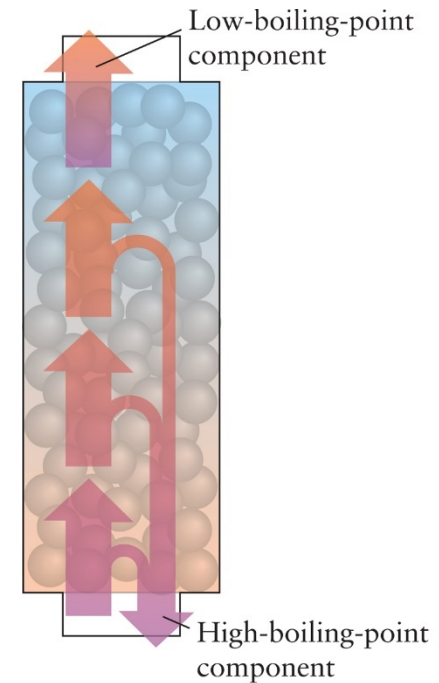
High-boiling-point component

Figure 5C.7
Atkins, *Chemical Principles: The Quest for Insight*, 7e
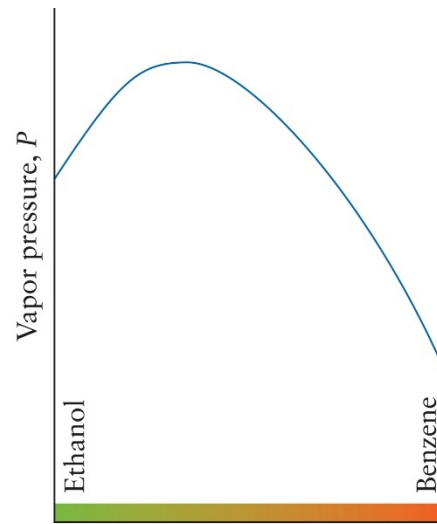W. H. Freeman & Company, © 2016 by P. W. Atkins, L. L. Jones, and L. E. Laverman

Figure 5C.8
Atkins, *Chemical Principles: The Quest for Insight*, 7e
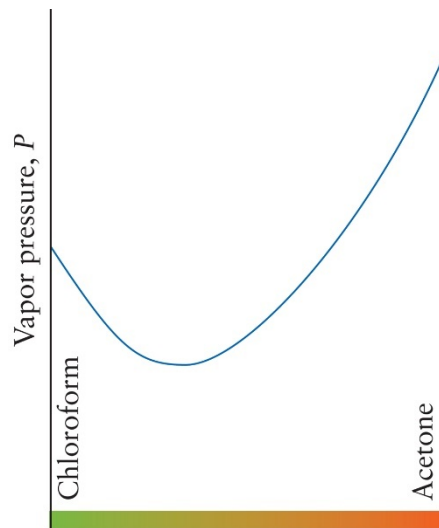© Pablo Paul/Alamy.

Azeotropes

Figure 5C.9a
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Figure 5C.9b
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Most liquids do not follow Raoult's law.

Ethanol and benzene have a positive, **endothermic** enthalpy of mixing; their vapor pressure is higher than predicted and produces a **positive deviation**.

Acetone and chloroform have a negative, **exothermic** enthalpy of mixing; their vapor pressure is lower than predicted and produces a **negative deviation**.

Deviations from Raoult's law can make it impossible to separate liquids by distillation.

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Azeotropes

Figure 5C.9a
Atkins, *Chemical Principles: The Quest for Insight*, 7e
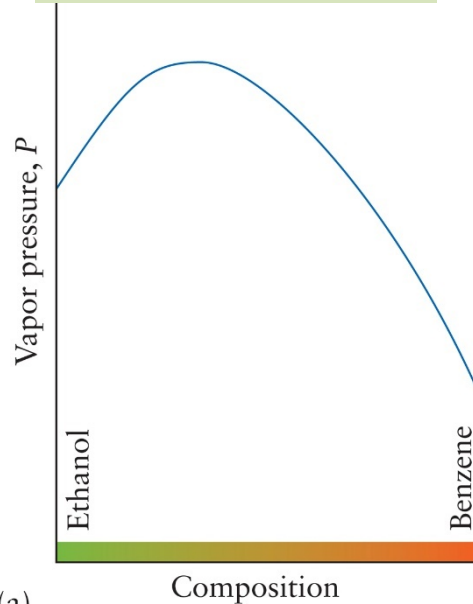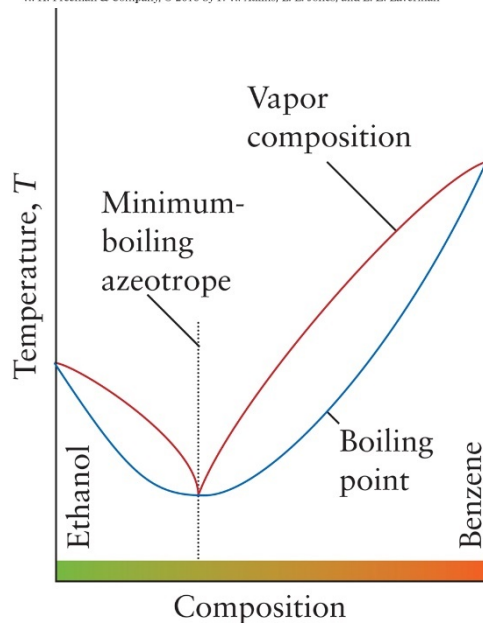W. H. Freeman & Company, © 2016 by P. W. Atkins, L. L. Jones, and L. E. Laverman

Figure 5C.10
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A positive deviation from Raoult's law means that the *attractive forces* between solute and solvent are **lower** than the pure component.

As a result, the boiling point of the mixture is lower than that predicted by Raoult's law and boils more easily.

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Solubility

In this next section we explore **solubility**, the molecular nature of these effects and see how to treat them quantitatively, specifically:

- limits of solubility
- like-dissolves-like rule
- pressure and gas solubility
- temperature and solubility
- thermodynamics of dissolving
- colloids

Gases dissolve into liquids

Pressure and Gas Solubility: Henry's Law

Oxygen, vital to all aquatic organisms, is **nonpolar** yet is slightly soluble in water (polar).

Pressure of a gas is a determining solubility factor.

First observed in 1801 by the English chemist William Henry, it is now known as **Henry's law.**

The **solubility** of a **gas** *is directly proportional* to its **partial pressure, P**

$$s = k_H P$$

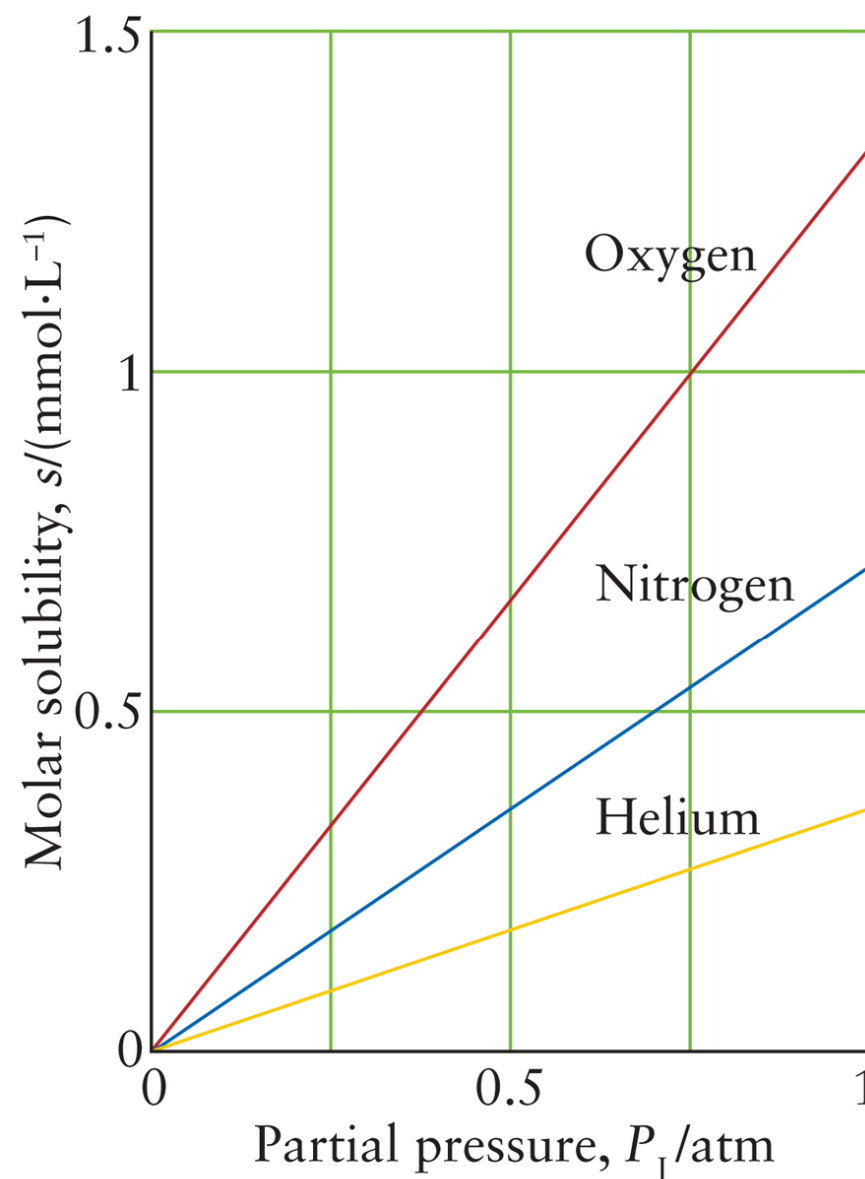k_H is called Henry's constant.

Figure 5D.8
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Example 5D.1 Verify that the concentration of oxygen in lake water is normally adequate to sustain aquatic life, which requires a concentration of at least 0.13 mmol·L^{-1}. The partial pressure of oxygen is 0.21 atm at sea level.

From $s = k_H P$

$k_H = 1.3 \times 10^{-3}$ mol·L^{-1}·atm^{-1}

$s = (1.3 \times 10^{-3}$ mol·L^{-1}·atm$^{-1}) \times 0.21$ atm

$= 2.7 \times 10^{-4}$ mol·L^{-1} which is much greater than 0.13 mmol·L^{-1}

Solids dissolve into liquids

The Enthalpy of Solution

Measuring heat released or absorbed when a substance dissolves is called molar enthalpy of solution, ΔH_{sol}.

TABLE 5D.3 Limiting Enthalpies of Solution, $\Delta H_{sol}/(kJ \cdot mol^{-1})$, at 25 °C*

	Anion							
Cation	fluoride	chloride	bromide	iodide	hydroxide	carbonate	sulfate	nitrate
lithium	+4.9	−37.0	−48.8	−63.3	−23.6	−18.2	−2.7	−29.8
sodium	+1.9	+3.9	−0.6	−7.5	−44.5	−26.7	+20.4	−2.4
potassium	−17.7	+17.2	+19.9	+20.3	−57.1	−30.9	+34.9	−23.8
ammonium	−1.2	+14.8	+16.0	+13.7	—	—	+25.7	+6.6
silver	−22.5	+65.5	+84.4	+112.2	—	+41.8	+22.6	+17.8
magnesium	−12.6	−160.0	−185.6	−213.2	+2.3	−25.3	−90.9	−91.2
calcium	+11.5	−81.3	−103.1	−119.7	−16.7	−13.1	−19.2	−18.0

Lithium chloride dissolves **exothermically** (heat releasing).
Others, such as ammonium nitrate, dissolve **endothermically** (heat absorbing).

The Enthalpy of Solution

The first step sublimes solid ions to gas ions.
Highly endothermic, this is the lattice enthalpy, ΔH_L, for NaCl:

$NaCl(s) + 787$ kJ·mol^{-1} → $Na^+(g) + Cl^-(g)$

In the second step, gaseous ions plunge into water forming the final solution. This is the enthalpy of hydration, ΔH_{hyd}, and for NaCl:

$Na^+(g) + Cl^-(g)$ → $Na^+(aq) + Cl^-(aq)$ - 784 kJ·mol^{-1}

Combining these steps : $\Delta H_{sol} = \Delta H_L + \Delta H_{hyd}$
$\Delta H_{sol} = \{787 + (-784)\}$ kJ·mol^{-1} = + 3 kJ·mol^{-1} (endothermic)
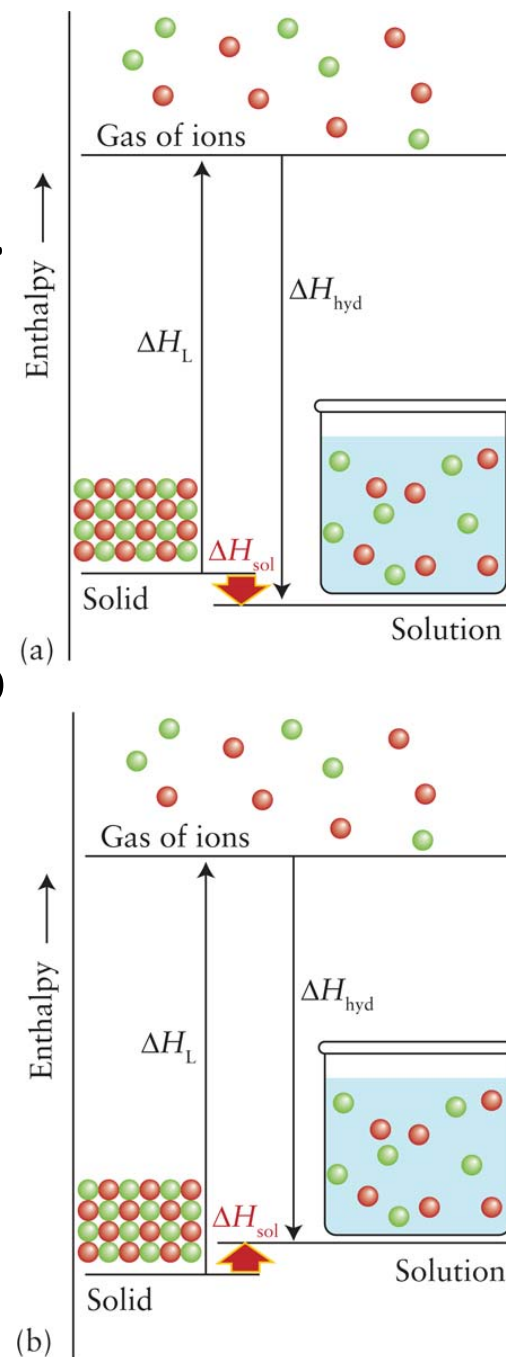

Figure 5D.10
Atkins, *Chemical Principles: The Quest for Insight,*
W. H. Freeman & Company, © 2016 by P. W. Atkin

The Enthalpy of Solution: Factors

$$\Delta H_{sol} = \Delta H_L + \Delta H_{hyd}$$
$$(+) \qquad (\text{mostly } -)$$

High charge and **small ionic radius** contribute to *high* lattice enthalpy (ΔH_L).

However, often *these can be the same properties* that relate to *low* enthalpy of hydration (ΔH_{hyd}).

Therefore it is *very difficult* to make reliable predictions and instead rationalize what is observed.

The Gibbs Free Energy of Solution

$$\Delta H_{sol} = \Delta H_{L} + \Delta H_{hyd}$$

A <u>negative enthalpy</u>, $-\Delta H_{sol}$, is an **energy releasing** process.

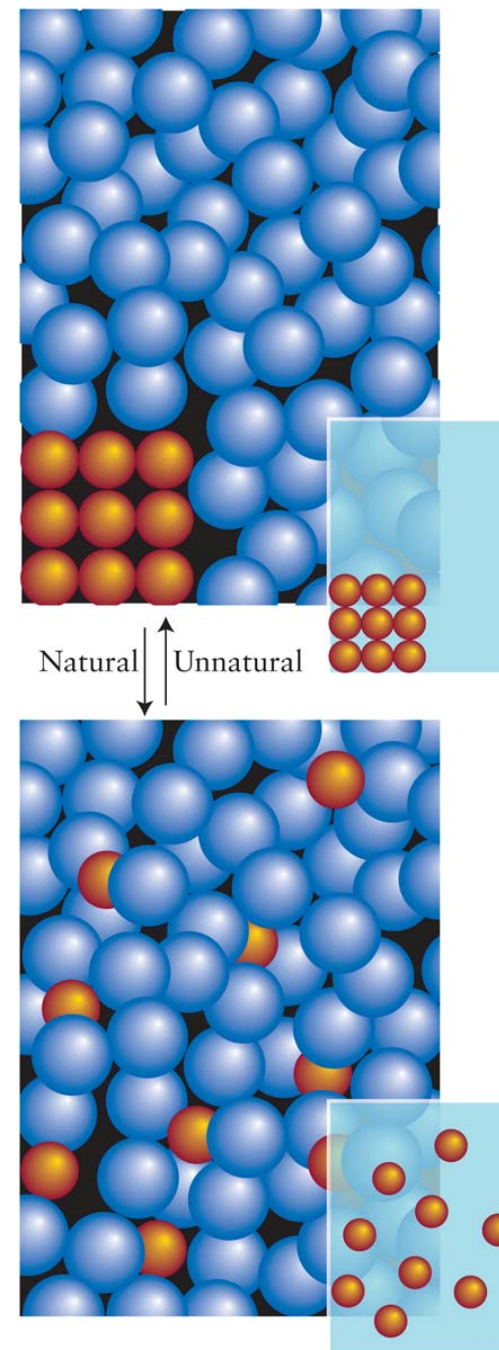An increase in ΔS

Natural | Unnatural

Figure 5D.11
Atkins, *Chemical Principles: The Quest for Insight*, 7
W. H. Freeman & Company, © 2016 by P. W. Atkins

The Gibbs Free Energy of Solution: **An Increase in ΔS (total)**

A <u>spontaneous change</u> (at constant temperature and pressure) is the Gibbs free energy, $\Delta G = \Delta H - T\Delta S$.

Dissolving a solid increases the disorder of a system, so we expect ΔS to increase or be positive.

Solubility: Limits

Forming a solution (solute and solvent)

"Free" water molecules pull crystals of glucose into the solution.

Both stirring and heating increase the solubility process.
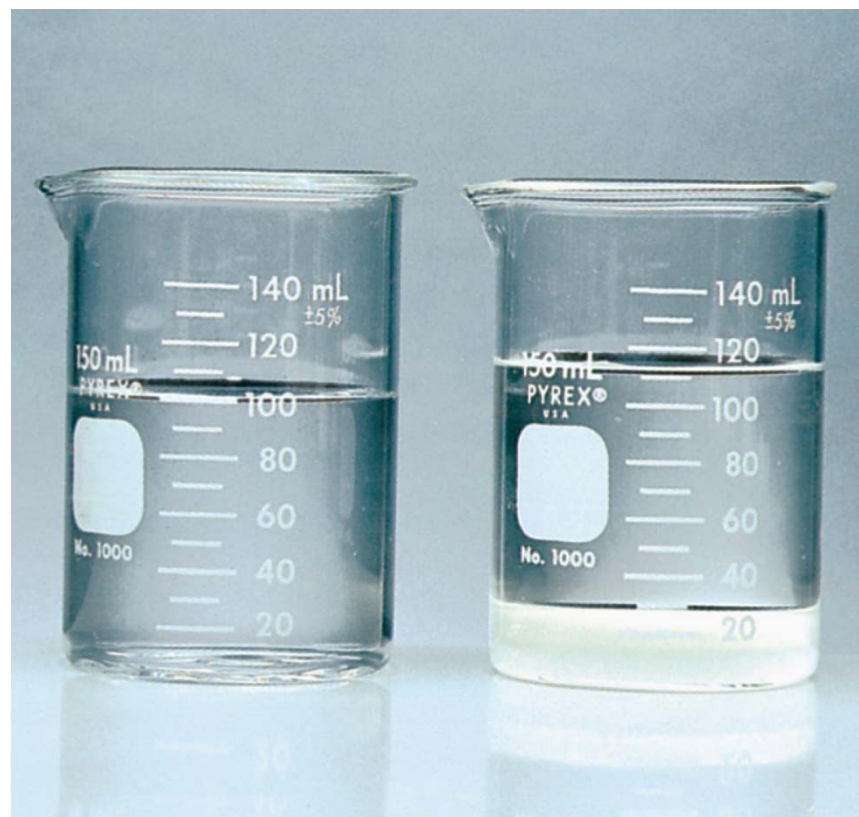

Figure 5D.2
Atkins, *Chemical Principles: The Quest for Insight*, 7e
W. H. Freeman photo by Ken Karp.
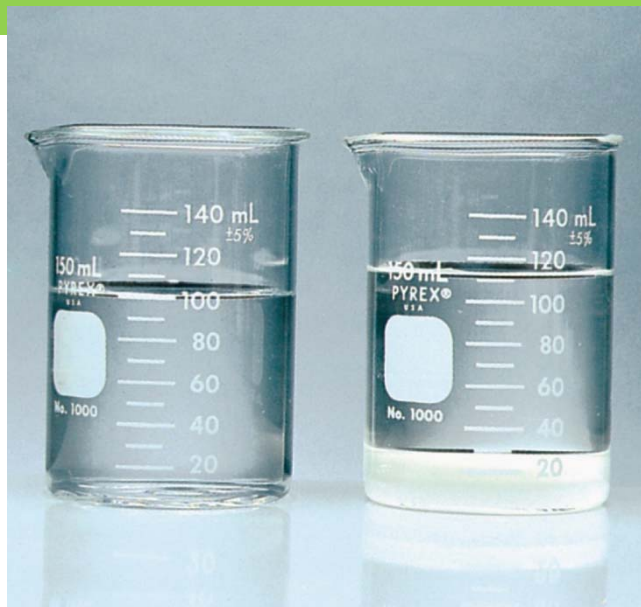
Solubility: Limits

Figure 5D.2
Atkins, *Chemical Principles: The Quest for Insight*, 7e
W. H. Freeman photo by Ken Karp.

Glucose is **unsaturated** at 20 g of glucose per 100 mL of water.

Beyond this point is a **saturated** solution. Excess glucose molecules will fall out of solution.
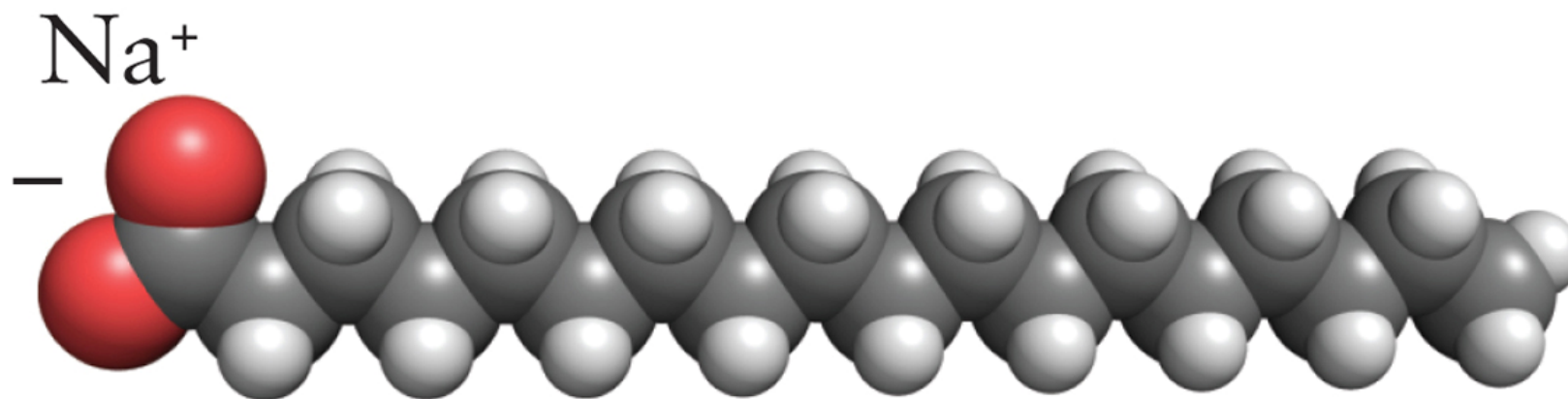
Solubility: **Like Dissolves Like**

A good guide is: **like dissolves like.**

Water (polar) can dissolve other polar and ionic compounds.

Hexane and tetrachloroethene (nonpolar), $Cl_2C=CCl_2$, dry–cleaning solvents, interact with other nonpolar compounds through London forces.

The cleaning action of soaps and detergents relies on the like-dissolves-like rule.

Na^+

$-$

1   Sodium stearate, $NaCH_3(CH_2)_{16}CO_2$

Atkins, *Chemical Principles: The Quest for Insight*, 7e
W. H. Freeman & Company, © 2016 by P. W. Atkins, L. L. Jones, and L. E. Laverman

Soaps are long-chain carboxylic acids, including sodium stearate.

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Solubility: Like Dissolves Like

Soap has polar **head** carboxylates ($-CO_2^{2-}$) attached to a nonpolar hydrocarbon **tail**.

The **hydrophobic** nonpolar hydrocarbon tails sink into blobs of grease.

The **hydrophilic** head encases the grease into a **micelle**, letting water wash the grease away.
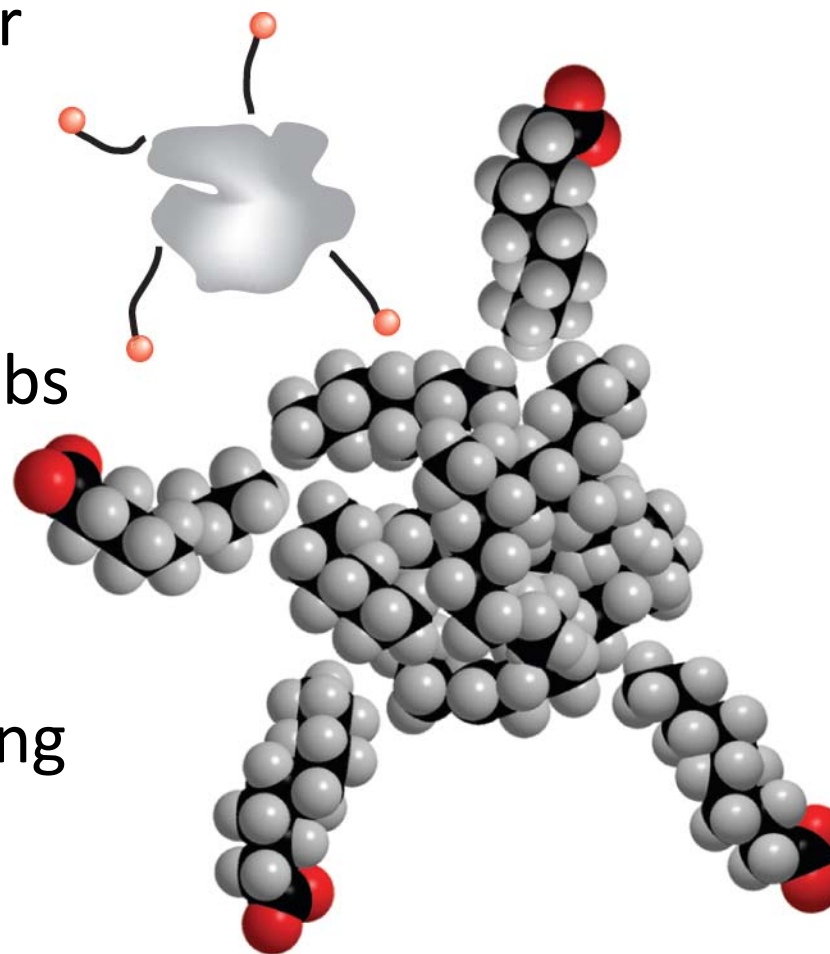
Figure 5D.7
Atkins, *Chemical Principles: The Quest for Insight*, 7e
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Temperature and Solubility

Solids dissolve more quickly (are more soluble) at higher temperatures than at low ones, but which does not necessarily hold for solubility.
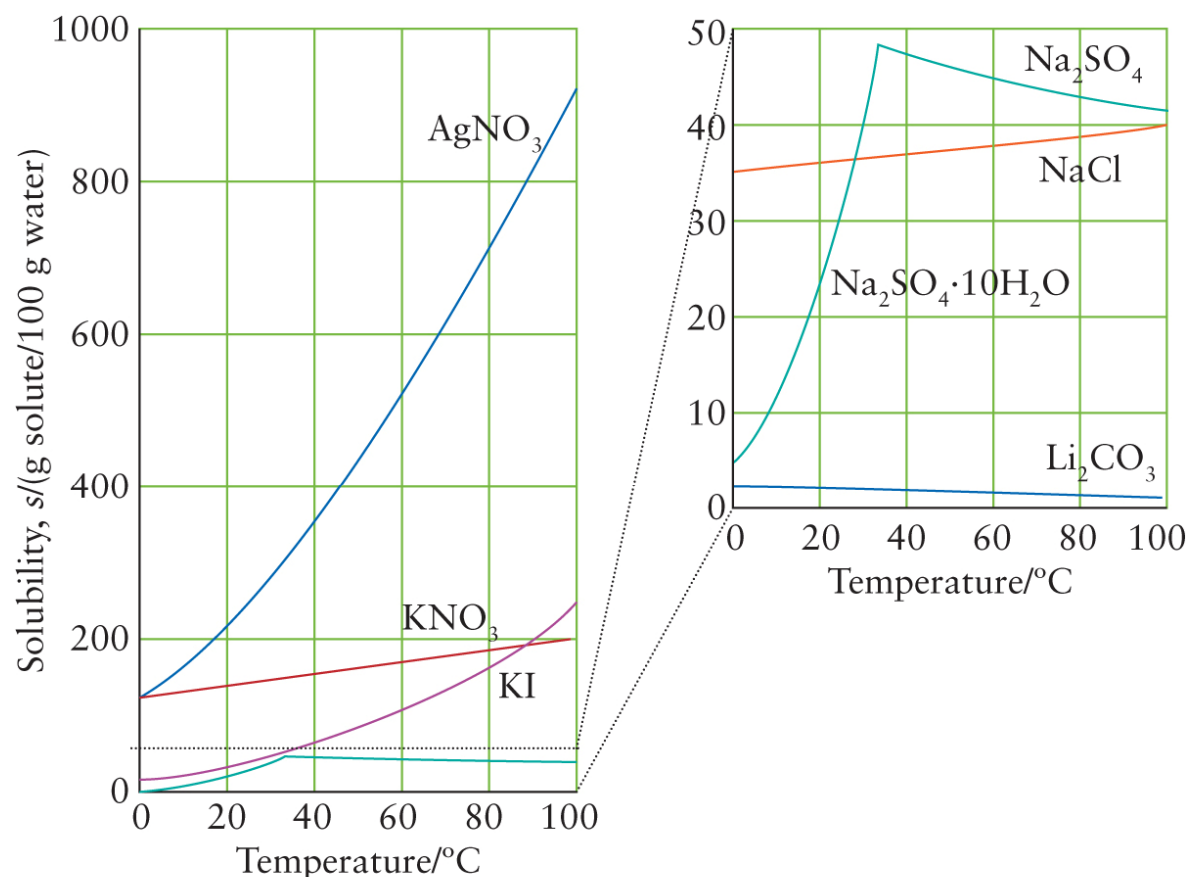
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Properties of dilute solutions

Colligative Properties

Properties that **depend on the numbers** of solute and solvent molecules and <u>not on chemical identity</u> are called **colligative properties**.

Four colligative properties of major importance are the:
1. lowering of the vapor pressure,
2. raising boiling points,
3. lowering of freezing points, and
4. osmosis.

<u>Extensive properties</u> depend on the quantity of matter (n atoms), volume, and mass whereas <u>intensive properties</u> are independent of the quantity of matter, density, and temperature. Colligative is a new third category.

Molality

Colligative properties are measured using either **mole fraction** or **molality** (next slide).

Both indicate the relative numbers of solute and solvent molecules (not mass or volume).

Mole fraction, χ, is covered in Focus 3, under Dalton's Law.

$$\chi_A = \frac{n_A}{n_A + n_B + \cdots}, \text{ note that: } \chi_A + \chi_B = 1$$

Molality

The molality of a solute is the amount of solute species (in moles) divided by the mass of the solvent (in kilograms):

$$\text{molality} = \frac{\text{amount of solute (mol)}}{\text{mass of solvent (kg)}}$$

Molality is *independent of temperature* (not true for molarity).

Units of *molality* $mol \cdot kg^{-1}$ and denoted as "***m.***" For example, a 1.0 *m* $NiSO_4$(aq) would read as 1.0 **molal** and it is exactly 1 mol $NiSO_4$ dissolved in 1 kg of water.

Calculating molality of a solution: What is the molality of NaCl in a solution prepared by dissolving 10.5 g of sodium chloride in 250 g of water?

Convert the mass of sodium chloride into moles of NaCl, then the mass of water in kilograms.

From molality $= \dfrac{n_{solute} \text{ (mol)}}{m_{solvent} \text{ (kg)}}$

$n_{solute} \text{ (mol)} = \dfrac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} \times \dfrac{10.5 \text{ g}}{1}$

$m_{solvent} \text{ (kg)} = \dfrac{1 \text{ kg}}{1000 \text{ g}} \times \dfrac{250. \text{ g water}}{1} = 0.250 \text{ kg water}$

What is the molality of NaCl in a solution prepared by dissolving 10.5 g of sodium chloride in 250 g of water?

$$\text{molality of a NaCl} = \frac{\dfrac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} \times \dfrac{10.5 \text{ g}}{1}}{0.250 \text{ kg water}} = 0.719 \text{ mol·kg}^{-1} \text{ or } 0.719 \text{ } m$$

Or dimensional analysis:

$$\frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} \times \frac{10.5 \text{ g}}{1} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1}{250. \text{ g water}} = 0.719 \text{ } m$$

Common and useful equations relating to **colligative properties**:
Toolbox 5E.1 (note the use of moles)

Mole fraction: $\chi_{solute} = \dfrac{n_{solute}}{n_{solute}+n_{solvent}}$ or $\chi_{solvent} = \dfrac{n_{solvent}}{n_{solute}+n_{solvent}}$

$1 = \chi_{solute} + \chi_{solvent}$

Calculating the mass of solvent from mole fraction (Toolbox 5E.1):

$m_{solvent}$ (g) $= n_{solvent}$ (mol) $M_{solvent}$ (g·mol^{-1}) $= \{(1-\chi_{solute})(mol)\}\, M_{solvent}$ (kg·mol^{-1})

(on the next slide we see why $(1-\chi_{solute})$ is equal to moles)

Calculating the molality, given the molarity (Example 5E.3):

m_{solute}(g) $= n_{solute}$ (mol) M_{solute} (g·mol^{-1}) $=$ molarity(mol·L^{-1})×1L× M_{solute}(g·mol^{-1})

$m_{solvent}$ (g) $= m_{solution}$ (g) $- m_{solute}$ (g)

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Self-test 5E.2A What is the molality of benzene, C_6H_6, dissolved in toluene, $C_6H_5CH_3$, in a solution for which the mole fraction of benzene is 0.150?

$$\text{molality} = \frac{n_{solute} \text{ (mol)}}{m_{solvent} \text{ (kg)}}$$

First we find the mass, kg, of toluene

benzene, C_6H_6 is the solute, and toluene is the solvent.

78.11 g·mol^{-1} C_6H_6
92.13 g·mol^{-1} $C_6H_5CH_3$
and since this is the solvent, 0.09213 kg·mol^{-1}

Self-test 5E.2A What is the molality of benzene, C_6H_6, dissolved in toluene, $C_6H_5CH_3$, in a solution for which the mole fraction of benzene is 0.150?

Calculating the mass of solvent from mole fraction:

$$m_{solvent} \text{ (g)} = n_{solvent} \text{ (mol) } M_{solvent} \text{ (g·mol}^{-1}\text{)} =$$
$$\{(1-\chi_{solute})\text{(mol)}\} M_{solvent} \text{ (kg·mol}^{-1}\text{)}$$

$$\{(1-0.150) \text{ mol}\} \times 0.09213 \text{ kg·mol}^{-1} = 0.0783 \text{ kg toluene}$$

$$m = \frac{n_{solute}}{\text{kg }_{solvent}}, \quad \frac{0.150 \text{ mol benzene}}{0.0783 \text{ kg toluene}} = 1.92 \text{ mol·kg}^{-1}$$

Find the molality of sucrose, $C_{12}H_{22}O_{11}$, in 1.06 M $C_{12}H_{22}O_{11}$(aq), which is known to have density 1.14 g.mL^{-1}.

$$\text{molality} = \frac{n_{solute} \text{ (mol)}}{m_{solvent} \text{ (kg)}}$$

342.3 g·mol^{-1} $C_{12}H_{22}O_{11}$

Calculating the molality, given the molarity:

m_{solute} (g) = n_{solute} (mol) M_{solute} (g·mol^{-1}) =
        molarity (mol·L^{-1})× 1L × M_{solute} (g·mol^{-1})

$m_{solvent}$ (g) = $m_{solution}$ (g) − m_{solute} (g)