### Properties of States of Matter

<table>
<thead>
<tr>
<th>State</th>
<th>Density</th>
<th>Volume, Shape</th>
<th>Molecular motion</th>
</tr>
</thead>
<tbody>
<tr>
<td>Solid</td>
<td>High</td>
<td>Well-defined volume and shape</td>
<td>Vibrate and rotate around fixed positions</td>
</tr>
<tr>
<td>Liquid</td>
<td>High</td>
<td>Well-defined volume, changes shape easily</td>
<td>Vibrate and rotate around positions that change from time to time</td>
</tr>
<tr>
<td>Gas</td>
<td>Low</td>
<td>Fills the entire volume of any shape</td>
<td>Random motion of molecules throughout the entire available volume</td>
</tr>
<tr>
<td>Plasma</td>
<td>Low</td>
<td>Takes different shapes</td>
<td>Random motion of ions and electrons</td>
</tr>
</tbody>
</table>

### Common States of Matter

A phase is a homogeneous part of the system separated by a well-defined boundary (interface).

- **Liquid phase**
- **Air-water interface**
- **Gas phase**
- **Plasma**

Multiphase (not a homogeneous system)

### 11.2 – Intermolecular Forces

**Intramolecular forces** hold atoms together in a molecule.

**Intermolecular forces** are forces between molecules.

Intermolecular forces are generally weaker than intramolecular forces:
- 930 kJ to break all O-H bonds in 1 mole of water (intra)
- 41 kJ to vaporize 1 mole of water (inter)

Evidence of large intermolecular forces in a solid or liquid state:
- High boiling point of a liquid
- High melting point of a solid
- Large energy needed to vaporize a liquid ($\Delta H_{vap}$)
- Large energy needed to melt a solid ($\Delta H_{fus}$)

Movies generated by Prof. Douglas Tobias (click to view; require QuickTime)
**Ion-Dipole Forces**
Ion-dipole forces attract an ion (anion or cation) and a polar molecule (dipole) to each other.

Ion dipole forces are electrostatic in nature (opposing charges attract each other per Coulomb law).

The larger is the separation between the ion and the dipole, the weaker is the force.

**Dipole-Dipole Forces**
Dipole-dipole forces are attractive forces between polar molecules (dipoles). They are also electrostatic (or Coulombic) in nature.

In a liquid or solid state, polar molecules try to orient so as to maximize the dipole-dipole attractive forces.

---

**Hydrogen Bond**
The hydrogen bond is an unusually strong dipole-dipole interaction between the hydrogen atom in a polar N-H, O-H, or F-H bond and an electronegative O, N, or F atom.

Hydrogen bond is especially important in liquid water and ice, where each molecule can accept up to 2 hydrogen bonds and donate up to two hydrogen bonds.

1 H-bond accepted
2 H-bonds donated (typical for liquid water)

2 H-bonds accepted
2 H-bonds donated (typical for ice)

**Signatures of Hydrogen Bonding**

- Decreasing molar mass
- Decreasing boiling point
- Anomalous increase in boiling point due to strong hydrogen bonding

Hydrogen bonds keep DNA strands paired.
Hydrogen Bond

They also affect how proteins fold

Polarization Forces

Attractive forces that arise as a result of dipoles induced in otherwise non-polar atoms or molecules by other charges

unpolarized molecule...

... polarizes into a dipole when placed next to an ion and is attracted to the ion through "ion – induced dipole" forces

... polarizes into a dipole when placed next to another dipole and is attracted to it through "dipole – induced dipole" forces

Polarizability

Polarizability is the extent to which the electron distribution in the atom or molecule can be distorted by an external field.

\[
\text{Induced Dipole} = \text{Polarizability} \times \text{Electric Field}
\]

Polarizability (C×m²×V⁻¹) increases with:
- greater number of electrons (molecular size)
- more diffuse electron cloud

Helium (2 electrons) – rigid and virtually unpolarizable atom
Argon (18 electrons) – more polarizable atom

Compare binding energies:

\[
\begin{align*}
\text{NH}_4^+ \text{-- He complex} & : D_a = 0.10 \text{ kJ/mol} \\
\text{NH}_4^+ \text{-- Ar complex} & : D_a = 0.34 \text{ kJ/mol}
\end{align*}
\]

Dispersion (London) Forces

Attractive forces that arise as a result of instantaneous (temporary) dipoles induced in atoms or molecules by random fluctuations in the motion of the electrons

Dispersion Forces and Polarizability

Dispersion forces usually increase as polarizability (and molecular size) increases.

Examples of Dispersion Forces

Van de Waals Complex: pair of molecules weakly held together, only stable under extremely low temperatures

Br₂ – Ar, Bromine argon complex (Prof. Janda)

He – He, Helium dimer, the weakest known complex with atoms separated by 5 nm and binding energy is as low as 0.001 K!

Helium Droplet: formed when He is expanded into vacuum as a supersonic jet. Superfluid, extremely cold, used to trap other molecules
**Intermolecular Forces: Summary**

- Ionic: between two ions, \( \text{Na}^+---\text{Br}^- \)
- Ion-dipole: an ion and a polar molecule, \( \text{Na}^+---\text{H}_2\text{O} \)
- Hydrogen-bonds: O-H, F-H, N-H bound to O, F, or N atom in another molecule, \( \text{H}_2\text{O}---\text{HF} \)
- Dipole-dipole: two polar molecules, \( \text{HCl}---\text{HCl} \)
- Ion-induced dipole: an ion and a polarizable molecule, \( \text{NH}_4^+---\text{Ar} \)
- Dipole-induced dipole: a polar molecule and a polarizable molecule, \( \text{Ar}---\text{HCl} \)
- Dispersion: any two polarizable molecules, \( \text{CO}_2---\text{CO}_2 \)

*Several types of intermolecular forces can be present at the same time.*

**Sample Problem**

What type(s) of intermolecular forces exist between each of the following molecules?

- **HBr**
  - HBr is a polar molecule: dipole-dipole forces should be important. There are also dispersion forces between HBr molecules because Br has a lot of electrons.
- **CH\(_4\)**
  - CH\(_4\) is nonpolar: only dispersion forces are possible.
- **H\(_2\)O**
  - Hydrogen bonds of dipole-dipole origin keep H\(_2\)O molecules together. Some dispersion forces are present also but they are much smaller than dipole-dipole forces.

**Choose the substance with the lowest boiling point:**

- CH\(_3\)OH
- MgCl\(_2\)
- CH\(_3\)(CH\(_2\))\(_3\)CH\(_3\)
- Xe

**Sample Problem**

1. Determine the types of intermolecular forces in each substance.

- CH\(_3\)OH
  - dipole-bonding
  - dispersion

- MgCl\(_2\)
  - ionic bonding

- CH\(_3\)(CH\(_2\))\(_3\)CH\(_3\)
  - dispersion

- Xe
  - dispersion

2. Compare the strength of different intermolecular forces: ionic > H-bonding > dipole-dipole > dispersion

<table>
<thead>
<tr>
<th></th>
<th>CH(_3)OH</th>
<th>MgCl(_2)</th>
<th>CH(_3)(CH(_2))(_3)CH(_3)</th>
<th>Xe</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ionic bonding</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>H-bonding</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Dipole-dipole</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Dispersion</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

3. Compare the size of molecules bound by dispersion forces: larger size → stronger forces → higher boiling point

- CH\(_3\)OH: 351 K
- MgCl\(_2\): 1685 K
- CH\(_3\)(CH\(_2\))\(_3\)CH\(_3\): 309 K
- Xe: 166 K

**Surface Tension**

*Surface tension* is the amount of energy required to stretch or increase the surface of a liquid by a unit area.

Molecules in the bulk are pulled equally in all directions.

Molecules on the surface are pulled inside. The resulting surface tension forces water droplets adopt spherical shapes in the absence of gravity.
**Adhesion and Cohesion**

Water droplets stick differently to different substrates:
- **Glass (hydrophilic)**
- **Teflon (hydrophobic)**

**Competition between:**
- **Surface tension:** attractive forces between water molecules work to make the droplet spherical and maximize the water-water interactions.
- **Liquid-substrate forces:** work to spread the droplet on the surface to maximize the substrate-water interactions.

**Adhesion:** liquid/substrate forces are larger than liquid/liquid forces (water on a hydrophilic surface).

**Cohesion:** liquid/liquid forces are larger than liquid/substrate forces (water on a hydrophobic surface).

---

**Viscosity**

Viscosity is a measure of the fluid's resistance to flow.

Higher viscosity correlates with stronger intermolecular forces.

<table>
<thead>
<tr>
<th>Compound</th>
<th>Int. mol. Forces</th>
<th>Viscosity (N s/m²)</th>
</tr>
</thead>
<tbody>
<tr>
<td>acetone</td>
<td>weak</td>
<td>3.16 x 10⁻⁴</td>
</tr>
<tr>
<td>ethanol</td>
<td>moderate</td>
<td>1.20 x 10⁻³</td>
</tr>
<tr>
<td>water</td>
<td>strong</td>
<td>1.01 x 10⁻³</td>
</tr>
<tr>
<td>glycerol</td>
<td>very strong</td>
<td>1.49</td>
</tr>
</tbody>
</table>

**Pitch drop experiment:** Started in 1930, first drop fell 8 years later! Last (8th) drop fell in 2000! Viscosity of pitch is about 10¹¹ time higher than that of water, but pitch is still regarded as a liquid because it flows.

---

**Unique Properties of Water**

Density of ice is less than that of liquid water – ice floats on water.

- Molecules in ice and liquid are organized in tetrahedral arrangements and kept in place by hydrogen bonds.
- When ice melts some of the hydrogen bonds break and water can fill up the cavities.

---

**11.4 – Crystal Structures of Solids**

**Ordered solids**
- Particles are organized in a crystalline manner.

**Disordered solids**
- Particles are organized in a non-ordered manner as amorphous.
A crystalline solid possesses rigid and long-range order. In a crystalline solid, atoms, molecules or ions occupy specific (predictable) positions. A unit cell is the basic repeating structural unit of a crystalline solid.

At lattice points:
- Atoms
- Molecules
- Ions

Common types of Unit Cells

Cubic Cells
- Simple Cubic (SC) – example Polonium metal
- Body-Centered Cubic (BCC) – example Uranium metal
- Face-Centered Cubic (FCC) – example Gold metal

Particles in the Unit Cells
- Depending on the location of the particle in the unit cell, it may be shared by multiple unit cells
  - Corner particle
  - Edge particle
  - Face particle

Sample Problem
When silver crystallizes, it forms FCC cells. The unit cell edge length is 409 pm. Calculate the density of silver.

\[ \rho = \frac{m}{V} \]

\[ V = a^3 = (409 \text{ pm})^3 = 6.83 \times 10^{-23} \text{ cm}^3 \]

4 atoms/unit cell in a face-centered cubic cell

\[ m = 4 \text{ Ag atoms} \times \frac{107.9 \text{ g}}{\text{mole Ag}} \times \frac{1 \text{ mole Ag}}{6.022 \times 10^{23} \text{ atoms}} = 7.17 \times 10^{-22} \text{ g} \]

\[ \rho = \frac{m}{V} = \frac{7.17 \times 10^{-22} \text{ g}}{6.83 \times 10^{-23} \text{ cm}^3} = 10.5 \text{ g/cm}^3 \]
11.6 – Types of Crystals

**Atomic solid** – interconnected identical atoms

**Ionic solid** – alternating positive and negative ions

**Molecular solid** – molecules held together by intermolecular forces

---

**Ionic Crystals**

- Lattice points occupied by cations and anions
- Held together by electrostatic attraction
- Hard, brittle, high melting point
- Poor conductors of heat and electricity

![Ionic Crystals Diagram](image)

- CsCl
- ZnS
- CaF₂

---

**Atomic Crystals**

- Lattice points occupied by atoms
- Held together by covalent bonds
- Hard, high melting point
- Poor conductors of heat and electricity

- Carbon
- Diamond
- Graphite

---

**Molecular Crystals**

- Lattice points occupied by molecules (or atoms)
- Held together by intermolecular forces
- Soft, low melting points
- Poor conductors of heat and electricity

- Example: dry ice, frozen CO₂

---

**Molecular Crystals**

Crystallizing molecules becomes increasingly more difficult if the molecules become larger. Nonetheless, crystals can be made from molecules as big as large proteins

---

**Metallic Crystals**

- Lattice points occupied by metal atoms
- Metal ions "swimming" in a "sea" of mobile elections
- Soft to hard, low to high melting point
- Good conductors of heat and electricity

![Metallic Crystals Diagram](image)
Metallic Crystals

Closest vs Randomized Packing

Animation of a close-packed structure from Wikipedia

Hexagonal Close Packing

Atoms in metals tend to organize themselves so as to pack in the closest possible arrangement leaving the least empty space.

Hexagonal arrangement of atoms offers the closest packing, such as in the Hexagonal Close Packed (hcp) structure

Cubic Close Packing

Cubic Close Packed (ccp) structure is another way to achieve close packing of spheres

Note that ccp is an fcc structure

Sample Problem

The element copper has ccp packing with a face-centered cubic unit cell. The density of Cu is 8920 kg/m³. Calculate the volume (m³) of the unit cell of Cu.

1. Determine the number of atoms per unit cell
   8 x 1/8 corner atoms = 1 atom
   6 x 1/2 face atoms = 3 atoms
   4 atoms/unit cell

2. Determine the mass of a unit cell
   \[ m = \frac{4 \text{ atoms/unit cell} \times 63.55 \text{ g Cu}}{1 \text{ mol} \times 6.022 \times 10^{23} \text{ atoms}} = 4.221 \times 10^{-22} \text{ g} \]

3. Determine the volume by making use of the density of Cu
   \[ \rho = \frac{m}{V} \quad V = \frac{m}{\rho} = \frac{4.221 \times 10^{-22} \text{ g}}{8920 \times 10^3 \text{ g/m}^3} = 4.73 \times 10^{-29} \text{ m}^3 \]

Sample Problem

Based on the unit cell representation shown, enter the coordination number for the cation.

The coordination number is equivalent to the number of atoms of the other species by which the atom is surrounded.

Lithium is the cation. Its nearest neighbors are:

There are 4 nearest neighbors, so its coordination number is 4
### 11.7 – Amorphous Solids

Solids that have no well-defined long-range arrangement of the atoms, but are approximately ordered over short distances.

<table>
<thead>
<tr>
<th>Material</th>
<th>Molecular Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sucrose and many other sugars</td>
<td>(CH₂O)ₙ</td>
</tr>
<tr>
<td>Amber, a blend of organic resins</td>
<td>C₁₀H₁₆O</td>
</tr>
<tr>
<td>Fat, with the average formula of</td>
<td>(BigMac)₂(CocaCola)₃</td>
</tr>
</tbody>
</table>

### 11.8 – Phase Changes

#### Types of Phase Changes

<table>
<thead>
<tr>
<th>Phase Change</th>
<th>Description</th>
</tr>
</thead>
<tbody>
<tr>
<td>Gas-liquid equilibrium</td>
<td>Molecules at the liquid/gas interface are constantly exchanging roles.</td>
</tr>
</tbody>
</table>

#### Dynamic Nature of Phase Changes

- **Evaporation**: Molecules escape from the liquid phase into the gas phase.
- **Condensation**: Molecules return from the gas phase into the liquid.

#### Liquid Vapor Equilibrium

- **Evaporation**: The number of molecules that escape form the liquid interface increases with the kinetic energy of the molecules (per Boltzmann energy distribution discussed in detail in Chem 130).
- **Condensation**: The number of molecules that condense into the liquid decreases with the kinetic energy of the molecules.

Because $E_{\text{kin}} = \frac{3}{2}RT$, the number of molecules that evaporate grows with temperature.

- **Low Temperature ($T_1$)**: Fewer molecules evaporate, resulting in lower vapor pressure.
- **High Temperature ($T_2$)**: More molecules evaporate, resulting in higher vapor pressure.

#### Liquid Vapor Equilibrium

- **Evaporation**: Empty space above the liquid will be filled up with evaporated molecules (gas).
- **Condensation**: As the number of molecules in the gas increases some of them start to condense back into the liquid.
**Liquid Vapor Equilibrium**

The *equilibrium vapor pressure* is the vapor pressure measured when the number of molecules that evaporate equals the number of molecules that condense per unit time.

\[
\text{H}_2\text{O (l)} \rightleftharpoons \text{H}_2\text{O (g)}
\]

**Dynamic Equilibrium**

\[
\text{Rate of evaporation} = \text{Rate of condensation}
\]

At 298 K (room temperature) the equilibrium vapor pressure of water is 24 mmHg.

**T-dependence of Vapor Pressure**

The higher the temperature, the more kinetic energy the molecules have, the more molecules are in the gas phase.

For water: \( P_{\text{H}_2\text{O}} \text{(mmHg)} = 7.14 \times 10^3 \times e^{\left(\frac{-51328}{T}\right)} \)

**Boiling Point**

The *boiling point* is the temperature at which the (equilibrium) vapor pressure of a liquid is equal to the external pressure.

The *normal boiling point* is the temperature at which a liquid boils when the external pressure is 1 atm.

<table>
<thead>
<tr>
<th>Substance (A)</th>
<th>Boiling Point* (°C)</th>
<th>( \Delta H_{\text{vap}} ) (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Argon (Ar)</td>
<td>−186</td>
<td>6.3</td>
</tr>
<tr>
<td>Benzene (C_6H_6)</td>
<td>80.1</td>
<td>31.0</td>
</tr>
<tr>
<td>Ethanol (C_2H_5OH)</td>
<td>78.3</td>
<td>39.3</td>
</tr>
<tr>
<td>Diethyl ether (C_4H_9O)</td>
<td>34.6</td>
<td>26.0</td>
</tr>
<tr>
<td>Mercury (Hg)</td>
<td>357</td>
<td>59.0</td>
</tr>
<tr>
<td>Methane (CH_4)</td>
<td>−164</td>
<td>9.2</td>
</tr>
<tr>
<td>Water (H_2O)</td>
<td>100</td>
<td>40.79</td>
</tr>
</tbody>
</table>

**Relative Humidity**

\[ \text{RH} = \frac{\text{Actual Partial Pressure of Water}}{\text{Equilibrium Pressure of Water}} \]

RH = 50% means corresponds to partial pressure of 12 mmHg at room temperature.

**Heat of Vaporization**

A common way to present the vapor pressure dependence on temperature (T) is:

\[
\ln P = \frac{-\Delta H_{\text{vap}}}{RT} + C
\]

C : constant

\( \Delta H_{\text{vap}} \) : Molar heat of vaporization

\( R = \text{gas constant (8.314 J/K\text{mol})} \)

The heat of vaporization depends on the intermolecular forces between molecules. If the forces are strong, \( \Delta H_{\text{vap}} \) is high. If the forces are weak, \( \Delta H_{\text{vap}} \) is low.

**Sample Problem**

The vapor pressure of benzene is 40.1 mmHg at 7.6 °C. What is the vapor pressure at 60.6 °C? The molar heat of vaporization is 31.0 kJ/mol.

HINT: Two sets of equilibrium vapor pressure and temperature, \( P_i \) & \( T_i \) and \( P_j \) & \( T_j \), are related through the Clausius-Clapeyron equation as follows:

\[
\ln \frac{P_i}{P_j} = \frac{\Delta H_{\text{vap}}}{R \left( \frac{T_j}{T_i} \right) - 1}
\]

1. Determine the conditions at the two T-P points:

\( T_1 = 7.6 \text{ °C} = 280.75 \text{ K} \)
\( T_2 = 60.6 \text{ °C} = 333.75 \text{ K} \)
\( P_1 = 40.1 \text{ mmHg} \)

\( P_2 = ? \)
Sample Problem

2. Determine $P_2$ using the modified Clausius-Clapeyron equation:

$$\ln \frac{40.1}{P_2} = \frac{31.0 \times 10^3 \text{ J/mol}}{8.314 \text{ J/K mol}} \left( \frac{280.75 \text{ K}}{333.75 \text{ K}} \right)$$

$$= 2.109$$

$$\frac{40.1}{P_2} = e^{2.109} = 1.121$$

$$P_2 = 330 \text{ mmHg}$$

Heat of Fusion

Melting requires energy. **Molar heat of fusion** ($\Delta H_{fus}$) is the energy required to melt 1 mole of a solid substance at its freezing point.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Melting Point ($^\circ$C)</th>
<th>$\Delta H_{fus}$ (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Argon (Ar)</td>
<td>−190</td>
<td>1.3</td>
</tr>
<tr>
<td>Benzene (C6H6)</td>
<td>5.5</td>
<td>10.9</td>
</tr>
<tr>
<td>Ethanol (C2H5OH)</td>
<td>−117.3</td>
<td>7.61</td>
</tr>
<tr>
<td>Diethyl ether (C4H10O)</td>
<td>−116.2</td>
<td>6.90</td>
</tr>
<tr>
<td>Mercury (Hg)</td>
<td>−39</td>
<td>23.4</td>
</tr>
<tr>
<td>Methane (CH4)</td>
<td>−183</td>
<td>0.84</td>
</tr>
<tr>
<td>Water (H2O)</td>
<td>0</td>
<td>6.01</td>
</tr>
</tbody>
</table>

Sublimation (Solid $\rightarrow$ Gas)

Sublimation: process in which molecules go directly from the solid phase into the gas phase

$$\text{CO}_2 \text{ (s)} \rightarrow \text{CO}_2 \text{ (g)}$$

**Molar heat of sublimation** ($\Delta H_{sub}$) is the energy required to sublime 1 mole of a solid.

$$\Delta H_{sub} = \Delta H_{fus} + \Delta H_{vap}$$

( of Hess’s Law)

Sample Problems

1. Calculate the temperature of a solution resulting from dissolving 20 g of ice at 0 °C in a beaker containing 100 mL of water at room temperature (25 °C). Assume that the specific heat capacity of water is $C_{water} = 4.184 \text{ J/(g \times K)}$.

**Note:** specific heat capacity is defined as the energy needed to raise the temperature of 1 g of substance by 1 K.

2. Calculate the amount of energy required to heat 1 kg of water from room temperature, 25 °C, to 150 °C. Assume that the specific heat of water (energy needed to heat 1 g of water by 1 K) is $C_{water} = 4.184 \text{ J/(g \times K)}$, and specific heat of water vapor (energy needed to heat 1 g of water vapor by 1 K) is $C_{vapor} = 1.99 \text{ J/(g \times K)}$.

Solution to problem 1: next page
Solution to problem 2: discussion sections
11.9 – Phase Diagrams

A phase diagram graphically represents the boundaries between conditions at which a substance exists as a solid, liquid, or gas.

Phase Diagram of Water

- Triple point: all three phases co-exist
- At 1 atm, 0°C: H₂O (s) → H₂O (l) (melting)
- At 1 atm, 100°C: H₂O (l) → H₂O (g) (evaporation)

Phase Diagram of Carbon Dioxide

- At 1 atm: CO₂ (s) → CO₂ (g) (sublimation)

Critical Conditions

The critical temperature ($T_c$) is the temperature above which the gas cannot liquefy, no matter how great the applied pressure is.

The critical pressure ($P_c$) is the minimum pressure that must be applied to bring about liquefaction at the critical temperature.

<table>
<thead>
<tr>
<th>Substance</th>
<th>$T_c$ (K)</th>
<th>$P_c$ (atm)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Arsenic (As)</td>
<td>711</td>
<td>14.8</td>
</tr>
<tr>
<td>Argon (Ar)</td>
<td>151</td>
<td>1.11</td>
</tr>
<tr>
<td>Carbon dioxide (CO₂)</td>
<td>304.3</td>
<td>7.39</td>
</tr>
<tr>
<td>Carbon tetrachloride (CCl₄)</td>
<td>190</td>
<td>6.6</td>
</tr>
<tr>
<td>Silver (Ag)</td>
<td>962</td>
<td>341</td>
</tr>
<tr>
<td>Tin (Sn)</td>
<td>1316</td>
<td>25.4</td>
</tr>
<tr>
<td>Hydrogen fluoride (HF)</td>
<td>17.2</td>
<td>100.5</td>
</tr>
<tr>
<td>Nitrine (N₂O)</td>
<td>342</td>
<td>33.5</td>
</tr>
<tr>
<td>Nitric oxide (NO)</td>
<td>176.5</td>
<td>3.6</td>
</tr>
<tr>
<td>Water (H₂O)</td>
<td>374.4</td>
<td>1.8</td>
</tr>
</tbody>
</table>

More Complicated Phase Diagrams

There are different types of ice in addition to the "normal" hexagonal ice. In fact, water remains the most studied substance to this day; there are scientific conferences devoted exclusively to phase transitions in ice and water!

Sample Problem

How many triple points are there?

What are the co-existing phases at the triple points?