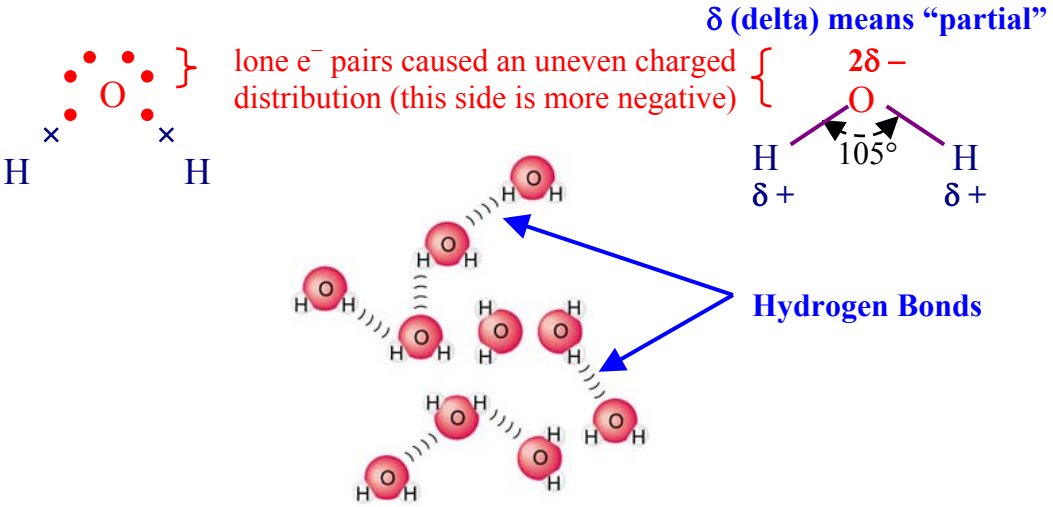
Unit 2: MATTER AS SOLUTIONS AND GASES

Chapter 17: Water as Aqueous Systems

17.1: Liquid Water and its Properties

Structure of Water:

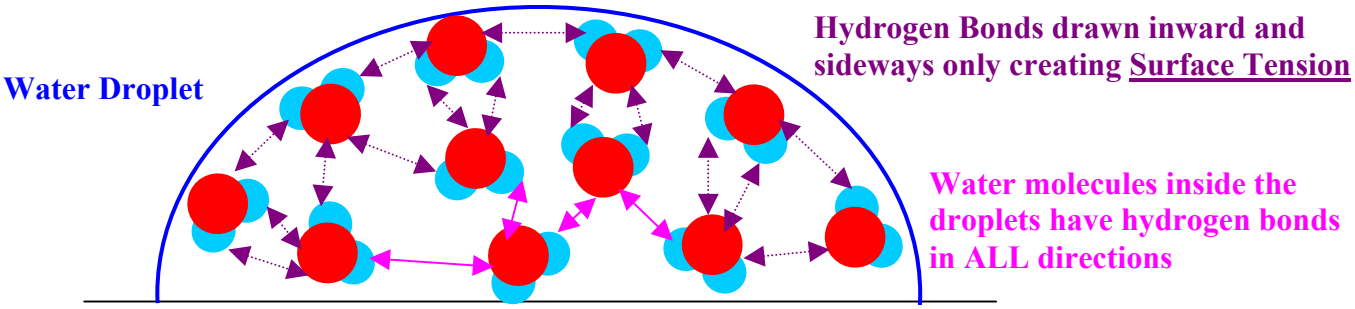
1. **V-Shaped**: - the two O–H bonds form 105° from each other, which leads to its polarity.
2. **Polar Molecule**: - unequal charged distribution due to the electron pairs around the oxygen atom.
3. **Strong O–H Hydrogen Bond**: - a type of hydrogen bond that is fairly strong compared to other types of intermolecular bonds (bonds between molecules).



Some Properties of Water:

1. **Strong Surface Tension**

Surface Tension: - the inward force of a liquid to minimize its surface area.
- water's hydrogen bond cannot attract with molecules in the air. Therefore, the higher net force is pushed inward, leaving a spherical surface with a high surface tension.



Surfactant: - a surface-active agent that decreases the surface tension (Example: detergent)

2. High Specific Heat Capacity

Specific Heat Capacity: - the amount of heat it takes to raise the temperature of 1 g of substance by 1°C.
- water has a high specific heat capacity, making it a good storage of solar energy

Specific Heat Capacity of Some Common Substances

Substance	Specific Heat Capacity	Substance	Specific Heat Capacity
Water	4.19 J/(g • °C)	Copper	0.385 J/(g • °C)
Ice	2.00 J/(g • °C)	Iron	0.449 J/(g • °C)
Ammonia	2.06 J/(g • °C)	Oxygen	0.918 J/(g • °C)
Ethanol	2.44 J/(g • °C)	Nitrogen	1.040 J/(g • °C)

17.2: Water Vapour and Ice

Other Properties of Water:

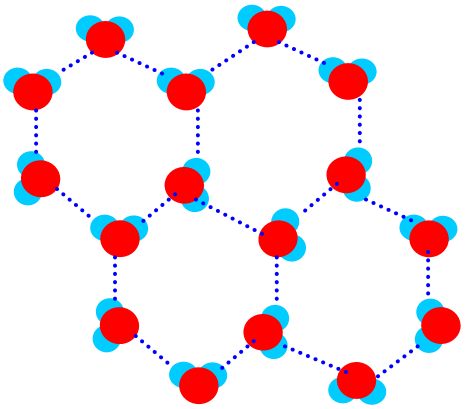
3. High Boiling Point and Heat of Vaporization

Boiling Point: - the temperature where the liquid form of a substance evaporates into its gaseous form (or a gaseous form of a substance condenses into its liquid form).
- water has a high boiling point (100°C) due to the strong hydrogen bonds. It takes more energy to break these intermolecular bonds.

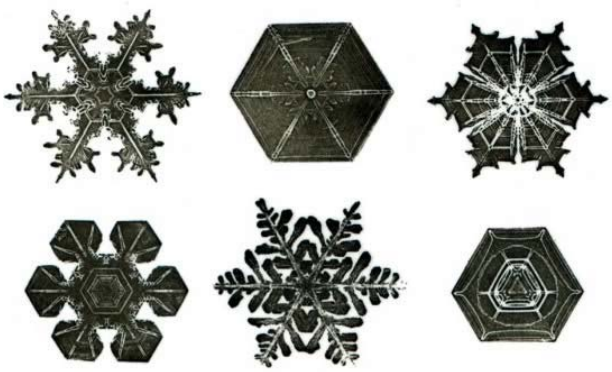
Heat of Vaporization: - the amount of heat to turn 1 mole or 1 gram of liquid substance into its gaseous form.
- water has a molar heat of vaporization of 40.65 kJ/mol or 2.256 kJ/g.
- because it takes an enormous amount of energy to evaporate water, temperatures are regulated on Earth. (Tropical regions are cooled as water absorbs heat. Polar regions are warmed as water vapour condenses to liquid giving off heat).

4. Low Density in Ice compared to Liquid Water

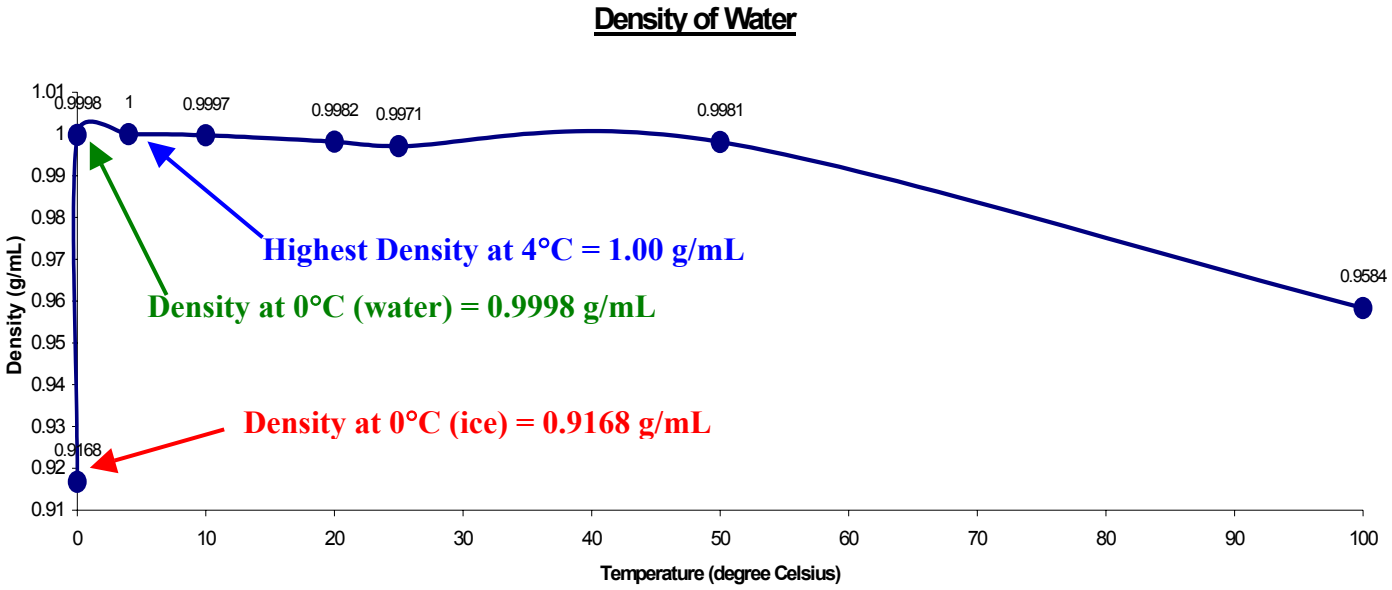
Hexagonal Arrangement of Ice Crystals: - due to the hydrogen bonds in water, it forms a honeycomb shape and expands in volume when it crystalizes into ice.



Due to the hydrogen bonds in water, it forms a *honeycomb shape* and expands in volume when it crystallizes into ice



Even though no two snowflakes are alike, all of them have a basic hexagonal shape as dictated by the bent shape of water molecule and its hydrogen bonds



Assignment
17.1 pg. 478 #1 to 4
17.2 pg. 481 #5 and 6

17.3: Aqueous Solutions

Solute: - the matter that is being dissolved.

Solvent: - the matter that is doing the dissolving.

Solvation: - the process where a solute is dissolving into a solvent.

Aqueous Solution: - solution where water is the solvent (*aq*).

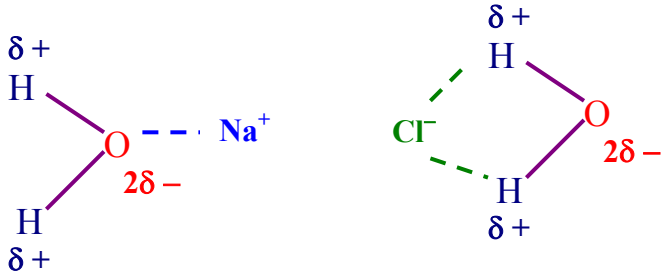
Reason that Water is a Common Solvent:

1. **Polar Molecule**: - dissolves many ionic compounds due to its ability to attract cations and anions (electrolytes).
   - *Note*: Some ionic compounds have low solubility (ability to dissolve) in water.

Hydration: - when ionic compound dissolves in water.

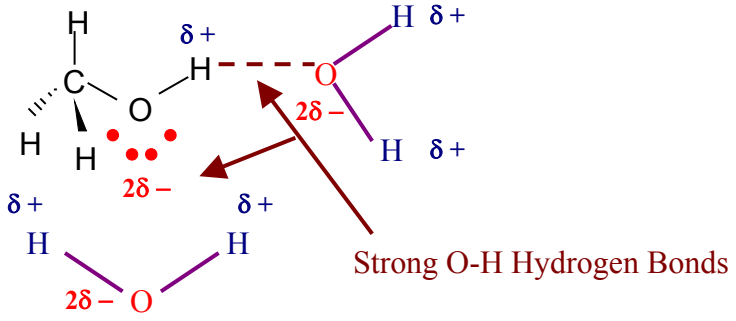
Example 1: Sodium Chloride (Salt)

$$NaCl_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + Cl^-_{(aq)}$$



2. **Strong O–H Hydrogen Bond**: - dissolves many molecular compounds that have O–H hydrogen bonds.

Example 2: CH_3OH (Methanol)



"**Like-Dissolves-Like**": - polar solvents tend to dissolve polar solutes; non-polar solvents tend to dissolve non-polar solutes.

Solubility Table: - a chart that shows the ability of various ion combinations to dissolve in water.

Solubility of Some Common Ionic Compounds in Water at 298.15 K

Ion	**Group 1** **NH_4^+** **H_3O^+ (H^+)**	**ClO_3^-** **NO_3^-** **ClO_4^-**	**CH_3COO^-**	**Cl^-** **Br^-** **I^-**	**SO_4^{2-}**	**S^{2-}**	**OH^-**	**PO_4^{3-}** **SO_3^{2-}** **CO_3^{2-}**
Solubility greater than or equal to 0.1 mol/L (very soluble)	all	all	most	most	most	Group 1 Group 2 NH_4^+	Group 1 NH_4^+ Sr^{2+} Ba^{2+} Tl^+	Group 1 NH_4^+
Solubility less than 0.1 mol/L (slightly soluble)	none	none	Ag^+ Hg^+	Ag^+ Pb^{2+} Hg^+ Cu^+ Tl^+	Ca^{2+} Sr^{2+} Ba^{2+} Ra^{2+} Pb^{2+} Ag^+	most	most	most

1. **Strong Electrolytes**: - ionic compounds that dissociates completely into their ions and conduct electricity very effectively.

Example 3:
a. All ionic compounds containing NH_4^+, NO_3^-, ClO_3^-, or ClO_4^-.
b. Strong Acids: $HClO_{4\,(aq)}$, $HI_{(aq)}$, $HBr_{(aq)}$, $HCl_{(aq)}$, $H_2SO_{4\,(aq)}$ and $HNO_{3\,(aq)}$
c. Strong Bases: $NH_4OH_{\,(aq)}$, $Sr(OH)_{2\,(aq)}$, $Ba(OH)_{2\,(aq)}$, $TlOH_{(aq)}$, and Group 1 with OH^- ($LiOH_{(aq)}$, $NaOH_{(aq)}$, $KOH_{(aq)}$, $RbOH_{(aq)}$, $CsOH_{(aq)}$, and $FrOH_{(aq)}$)

2. **Weak Electrolytes**: - ionic compounds that **dissociate partially** into their ions and **conduct electricity poorly**.

Example 4: a. $AgCl_{(s)}$, $PbCl_{2(s)}$, $HgCl_{(s)}$, $CuCl_{(s)}$, and $TlCl_{(s)}$
b. Weak Acids: $HF_{(aq)}$, $HCH_3COO_{(aq)}$, $H_2SO_{3\,(aq)}$, and other weak acids.
c. Weak Bases: $Ca(OH)_{2\,(aq)}$, $Al(OH)_{3\,(aq)}$, $NH_{3\,(aq)}$, and other weak bases.

3. **Non-Electrolytes**: - ionic compounds that **do not dissociate** (generally molecular compounds) in the solvent and **do not conduct electricity at all**.

Example 5: $C_{12}H_{22}O_{11\,(aq)}$ and $C_6H_{12}O_{6\,(aq)}$

Hydrates: - sometimes called water of hydration or water crystallization.
- when water particles are locked into the crystal structure of an ionic compound.
- naming contains the name of the ionic compound with a prefix followed by the word "*hydrate*".

Prefixes for Hydrates

1 – mono	4 – tetra	7 – hepta	10 – deca
2 – di	5 – penta	8 – octa	
3 – tri	6 – hexa	9 – nona	

Example 6: Write the equation of the formation of iron (II) sulfate heptahydrate and determine the percent mass of water in this hydrate.

iron (II) sulfate + water → iron (II) sulfate heptahydrate

$\mathbf{FeSO_{4\,(s)}} + \mathbf{7\,H_2O_{(l)}} \rightarrow \mathbf{FeSO_4 \bullet 7\,H_2O}$

$M = 18.02$ g/mol

$M = 55.85 + 32.06 + 4(16.00) + 7(18.02)$

$M = 277.91$ g/mol

$$\%\ \text{water by mass} = \frac{7(18.02\ \text{g/}\cancel{\text{mol}})}{277.91\ \text{g/}\cancel{\text{mol}}} \times 100\%$$

% H_2O = 45.4%

Chapter 18: Solutions

18.1: Properties of Solution

Solubility: - the amount of solute that can dissolve in a given amount of solvent at a specific temperature.
- usually measures in g / 100 g of solvent.

1. **Unsaturated Solution**: - when a solvent can still dissolve more solute at a particle temperature.
2. **Saturated Solution**: - when a solvent cannot dissolve anymore solute.
   - the rate of solvation is the same as the rate of crystallization.
3. **Supersaturated Solution**: - when a saturated solution is heated and more solute is dissolved, which is then followed by a slow cooling without any recrystallization.
   - the density and the viscosity of a supersaturated solution is higher than a saturated solution. (Example: Sugar Syrup – Supersaturated Sugar Water)
4. **Miscible**: - when two liquids can dissolve each other.
   (**Example**: Vinegar-Water Solution – Both have polar bonds.)
5. **Partially Miscible**: - when two liquids are slightly soluble in each other.
   (**Example**: Anaesthetic Diethyl ether-Water Solution – weak polar versus strong polar bonds.)
6. **Immiscible**: - when two liquids cannot dissolve in each other. They form a mechanical (heterogeneous) mixture.
   (**Example**: Italian Salad Dressing – Vinegar and Oil: polar versus non-polar bonds.)

Factors Affecting Solubility

1. **Temperature**: - as temperature increases, solubility of solid solutes generally increases.
   - as temperature increases, solubility of gaseous solutes decreases in aqueous solution.

Temperature ↑ **Solid Solute Solubility ↑** **Gas Solute Solubility ↓**

2. **Partial Pressure**: - as pressure increases, solubility of gas solutes generally increases. (More force is pushed down to force gas particles to dissolve in a denser liquid solvent).

Pressure ↑ **Gas Solute Solubility ↑**

3. **Agitation (Stirring)**: - as agitation increases, solubility of solid solutes generally increases.

Agitation ↑ **Solid Solute Solubility ↑**

4. **Particle Size**: - as the particle size of the solute increases, the rate of solubility decreases.

Particle Size ↑ **Rate of Solubility ↓**

Example 1: A maximum 6.70 g of carbon dioxide gas is dissolved in 2000 g of water at 0°C under normal atmosphere pressure (101.3 kPa). Calculate the solubility of carbon dioxide gas in g / 100 g of H_2O.

$$\text{Solubility} = \frac{\textbf{Mass of Solute}}{\textbf{100 g of Solvent}}$$

$$\text{Solubility} = \frac{6.70\text{ g } CO_2}{2000\text{ g water}} = \frac{x\text{ g } CO_2}{100\text{ g water}} \qquad \text{(cross-multiply)}$$

$$2000x = (6.70)(100)$$

$$x = \frac{(6.70)(100)}{2000}$$

$$x = 0.335\text{ g } CO_2$$

Solubility = 0.335 g CO_2 / 100 g of water

Example 2: The solubility of sucrose (cane sugar) is 230.9 g / 100 g of H_2O. Determine the amount of solvent needed to dissolve 3.00 kg of can sugar.

$$\text{Solubility} = \frac{230.9\text{ g sucrose}}{100\text{ g } H_2O} = \frac{3000\text{ g sucrose}}{x\text{ g } H_2O}$$

3.00 kg = 3000 g

(cross-multiply)

$$230.9x = (100)(3000)$$

$$x = \frac{(100)(3000)}{230.9}$$

$$x = 1299.263751\text{ g water}$$

x = 1.30 kg of H_2O

Assignment

17.3 pg. 488 #8 to 10 and 12

18.1 pg. 507 #3, 5, 6 and 7

18.2: Concentration of Solution

Molarity (M): - moles of solute per volume of solution in Litres (mol/L).
- commonly referred to as **molar concentration (C)**.

Molar Concentration:

$$\text{Molarity (mol/L)} = \frac{\text{moles (mol)}}{\text{Volume (L)}} \qquad C = \frac{n}{V} \qquad 1\text{ mol/L} = 1\text{ M}$$

C = Molar Concentration n = moles V = Volume

Example 1: 3.46 g of copper (II) nitrate is dissolved in 250.0 mL of water. Calculate the molarity of the solution formed.

$m = 3.46\text{ g } Cu(NO_3)_2$

$n = \dfrac{3.46\text{ g}}{187.564\text{ g/mol}}$

$n = 0.0184470367\text{ mol } Cu(NO_3)_2$

$V = 250.0\text{ mL} = 0.2500\text{ L}$

$C = ?$

$C = \dfrac{n}{V} = \dfrac{0.0184470367\text{ mol}}{0.2500\text{ L}}$

$C = 0.0738$ mol/L or 73.8 mmol/L

Example 2: Determine the mass of sodium dichromate needed for 500.0 mL of 0.0300 M.

$V = 500.0\text{ mL} = 0.5000\text{ L}$

$C = 0.0300\text{ mol/L}$

$M = 261.98\text{ g/mol } Na_2Cr_2O_7$

$n = ?$ $m = ?$

$C = \dfrac{n}{V}$

$n = CV = (0.0300\text{ mol/}\cancel{\text{L}})(0.5000\ \cancel{\text{L}})$

$n = 0.0150$ mol

$n = \dfrac{m}{M}$

$m = nM = (0.0150\ \cancel{\text{mol}})(261.98\text{ g/}\cancel{\text{mol}})$

$m = 3.93$ g

Example 3: The Palmense Mineral Water from the city of Fermo in Italy has a sodium hydrogen carbonate concentration of 5.207 mmol/L. What volume of this mineral water will contain 4.00 mg of sodium hydrogen carbonate?

$m = 4.00\text{ mg } NaHCO_3$

$n = \dfrac{4.00\text{ mg}}{84.0079\text{ g/mol}}$

$n = 0.0476145696\text{ mmol } NaHCO_3$

$C = 5.207\text{ mmol/L}$

$V = ?$

$C = \dfrac{n}{V}$ $V = \dfrac{n}{C} = \dfrac{0.0476145696\ \cancel{\text{mmol}}}{5.207\ \cancel{\text{mmol}}/\text{L}}$

$V = 0.00914$ L or 9.14 mL

Dissociation: - when ionic compounds completely dissolve in water (100% soluble), the ionic bonds are severed and the ions "swim" freely in the new aqueous environment.

Chemical Dissociation Equation

$$A_xB_{y\,(s)} \rightarrow xA^{y+}_{(aq)} + yB^{x-}_{(aq)}$$

$[A^{y+}]$ = Molar Concentration of Cation A^{y+} **$[B^{x-}]$ = Molar Concentration of Anion B^{x-}**

Example 4: Write the chemical dissociation equation for the following ionic compounds when they dissolve in water.

a. $NaCl_{(s)}$

$$NaCl_{(s)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)}$$

b. $Cu(NO_3)_{2\,(s)}$

$$Cu(NO_3)_{2\,(s)} \rightarrow Cu^{2+}_{(aq)} + 2\ NO_3^-{}_{(aq)}$$

Example 5: Calculate the molar concentration for each ion when the following ionic compounds dissolve in water.

a. 0.150 M of $NaOH_{(aq)}$

$$NaOH_{(s)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

$[NaOH] = 0.150$ M

$$[Na^+] = 0.150\ \text{mol/L}\ NaOH_{(s)} \times \frac{1\ \text{mol}\ Na^+}{1\ \text{mol}\ NaOH_{(s)}}$$

$[Na^+] = 0.150$ mol/L

$$[OH^-] = 0.150\ \text{mol/L}\ NaOH_{(s)} \times \frac{1\ \text{mol}\ OH^-}{1\ \text{mol}\ NaOH_{(s)}}$$

$[OH^-] = 0.150$ mol/L

b. 3.45 g of potassium phosphate in 500 mL of water.

$$K_3PO_{4\,(s)} \rightarrow 3\,K^+_{(aq)} + PO_4^{3-}{}_{(aq)}$$

$m = 3.45$ g K_3PO_4

$$n = \frac{3.45\ \text{g}}{212.274\ \text{g/mol}}$$

$n = 0.0162525792$ mol K_3PO_4

$V = 500$ mL $= 0.500$ L

$C = ?$

$$C = \frac{n}{V} = \frac{0.0162525792\ \text{mol}}{0.500\ \text{L}}$$

$[K_3PO_4] = 0.0325$ mol/L

$$[K^+] = 0.0325\ \text{mol/L}\ K_3PO_{4\,(s)} \times \frac{3\ \text{mol}\ K^+}{1\ \text{mol}\ K_3PO_{4(s)}}$$

$[K^+] = 0.0975$ mol/L

$$[PO_4^{3-}] = 0.0325\ \text{mol/L}\ K_3PO_{4\,(s)} \times \frac{1\ \text{mol}\ PO_4^{3-}}{1\ \text{mol}\ K_3PO_{4(s)}}$$

$[PO_4^{3-}] = 0.0325$ mol/L

Dilution: - the process where additional solvent is added to lower the concentration of an original solution.

Dilution

$$C_1V_1 = C_2V_2$$

C_1 = Concentration of Original Solution **V_1 = Volume of Original Solution**
C_2 = Concentration of Diluted Solution **V_2 = Total Volume of Diluted Solution**

Example 6: Concentrated hydrochloric acid comes in 17.4 M. What is the volume of concentrated $HCl_{(aq)}$ needed to obtain 250 mL of 1.50 M of $HCl_{(aq)}$?

$C_1 = 17.4$ mol/L
$V_1 = ?$
$C_2 = 1.50$ mol/L
$V_2 = 250$ mL

$$C_1V_1 = C_2V_2$$

$$V_1 = \frac{C_2V_2}{C_1} = \frac{(1.50\ \text{mol/L})(250\ \text{mL})}{(17.4\ \text{mol/L})}$$

$V_1 = 21.6$ mL

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Example 7: Determine the volume of water needed to dilute 30.0 mL of 0.500 M $CuSO_{4\,(aq)}$ to 0.0750 M.

$C_1 = 0.500$ mol/L
$V_1 = 30.0$ mL
$C_2 = 0.0750$ mol/L
$V_2 = ?$

$C_1V_1 = C_2V_2$

$$V_2 = \frac{C_1V_1}{C_2} = \frac{(0.500\ \cancel{\text{mol/L}})(30.0\ \text{mL})}{(0.0750\ \cancel{\text{mol/L}})}$$

$V_2 = 200$ mL

$V_{water} = V_2 - 30.0\ \text{mL} = ?$

$V_{water} = V_2 - 30.0$ mL
$V_{water} = 200$ mL $- 30.0$ mL

$V_{water} = 170$ mL

Percent Solutions: - concentration of solution expressed in **% volume/volume** or **% mass/volume**.

Percent Solutions

$$\text{\% by Volume (\% (v/v))} = \frac{\text{Volume of Solute}}{\text{Total Volume of Solution}} \times 100\%$$

(Same Unit for Both Volumes)

$$\text{\% by Mass (\% (m/v))} = \frac{\text{Mass of Solute (g)}}{\text{Total Volume of Solution (mL)}} \times 100\%$$

Example 8: Calculate the percent solutions of the following.

a. 1.32 g of salt in 325 mL of water

$$\%\,(m/v) = \frac{\text{Mass of Solute (g)}}{\text{Total Volume of Solution (mL)}} \times 100\%$$

$$\%(m/v) = \frac{1.32\ \text{g salt}}{325\ \text{mL water}} \times 100\%$$

0.406% salt (m/v)

b. 45.0 mL of vinegar with 250 mL of water

$$\%\,(v/v) = \frac{\text{Volume of Solute}}{\text{Total Volume of Solution}} \times 100\%$$

$$\%\,(v/v) = \frac{45.0\ \text{mL vinegar}}{(250\ \text{mL}\ H_2O + 45.0\ \text{mL solute})} \times 100\%$$

15.3% vinegar (v/v)

Example 9: What is the mass of solute required to make a 3.75 L of a 22.5% sugar (m/v) solution?

$$\%\,(m/v) = \frac{\text{Mass of Solute (g)}}{\text{Total Volume of Solution (mL)}} \times 100\%$$

$$22.5\%\,(m/v) = \frac{m_{solute}}{3750\ \text{mL water}} \times 100\%$$

3.75 L = 3750 mL

$$0.225 = \frac{m_{solute}}{3750\ \text{mL water}}$$

$m_{solute} = (0.225)(3750) = 843.75$ g

$m_{solute} = 844$ g of sugar

Assignment

18.2 pg. 511 #8 to 11; pg. 513 #12, 13; pg. 514 #14, 15; pg. 515 #16 to 23
Ch 17 Review: pg. 496 – 497 #19, 20, 22, 23, 24, 29, 31 to 35, 38, 40 and 41
Ch 18 Review: pg. 528 #40 to 47, 49 to 54

Chapter 10: States of Matter

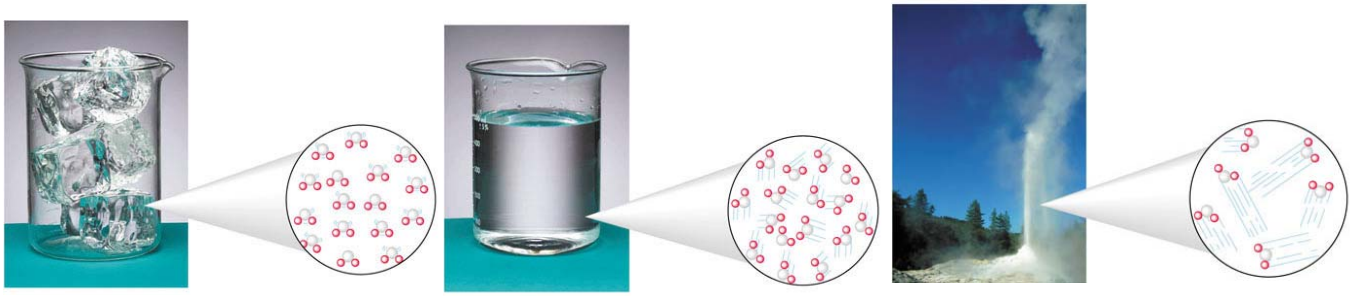
10.1: The Nature of Gases

Solids: - the state of matter where it has a definite volume with a constant shape.

Liquid: - the state of matter where it has a definite volume but an indefinite shape.

Gas: - the state of matter where it has an indefinite volume and shape (compressible).

Kinetic Molecular Theory and States of Matter

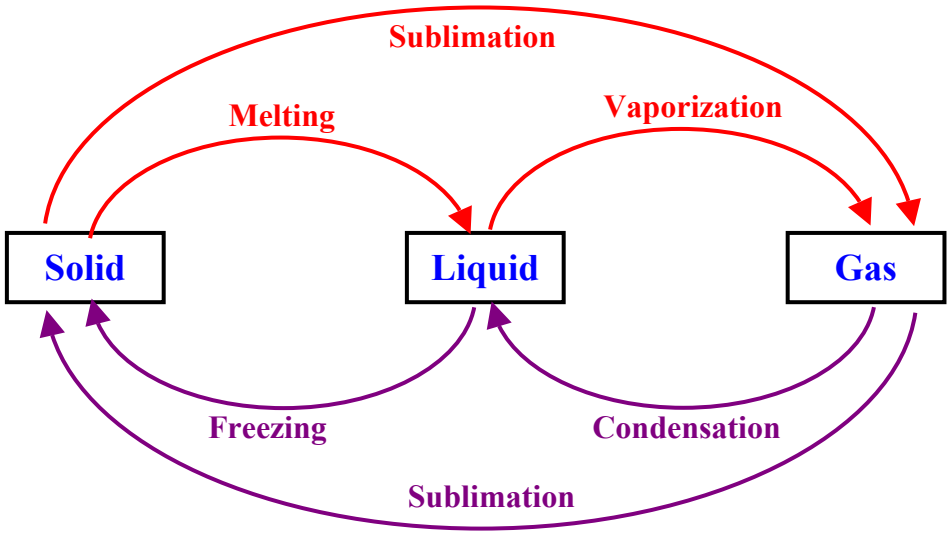


Solids have particles are in fixed positions.

Liquids have particles that can "rolled" past one another.

Gases have particles that have truly random motion and have very weak interactions between molecules.

Phase Change

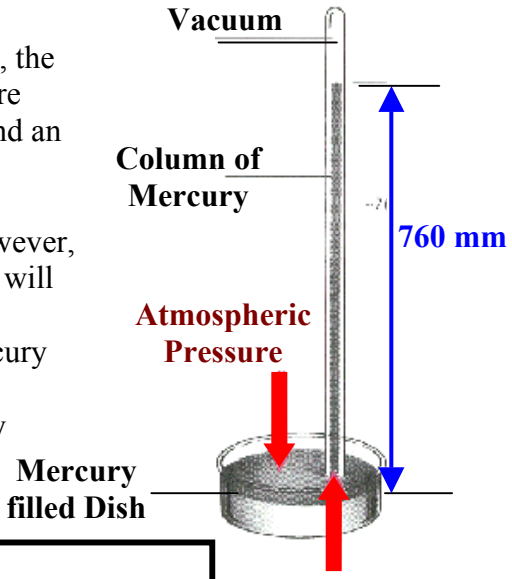


Sublimation: - the phase change from solid to gas directly or vice versa.

Example: Dry ice (Solid Carbon Dioxide) sublimates from solid to gas directly, skipping the liquid phase.

Pressure: - the amount of force per unit of area [1 N/m^2 = 1 Pascal (Pa)], measured in **kiloPascal (kPa), mm Hg**, **torr**, **atmospheric pressure (atm)**.
- in a pressurized container, a **pressure can be felt as the particles pushed on the inside wall** of the container **as well as when they collide with each other elastically** ***(no lost in energy during collision)*.**

Barometer: - an instrument (invented by Evangelisto Torricelli, hence, the old unit of pressure *torr* was named after him) to measure atmospheric pressure using a dish filled with mercury and an inverted glass tube filled with mercury.
- when the glass tube is placed into the dish filled with mercury, the level of mercury in the tube will drop. However, the atmospheric pressure of the surface of mercury dish will keep the mercury column at a certain level.
- the higher the atmospheric pressure, the higher the mercury is in the column.
- at normal atmospheric pressure at sea level, the mercury column is at 760 mm.



Units of Pressure

1 standard atmosphere (atm) = 760 mm Hg = 760 torr = 101.325 kPa

Note: 1 mm Hg = 1 torr

Example 1: Convert 525. mm Hg to atm and kPa.

$$525.\ \cancel{\text{mm Hg}} \times \frac{1\ \text{atm}}{760\ \cancel{\text{mm Hg}}} = \mathbf{0.691\ atm} \qquad 525.\ \cancel{\text{mm Hg}} \times \frac{101.325\ \text{kPa}}{760\ \cancel{\text{mm Hg}}} = \mathbf{70.0\ kPa}$$

Example 2: Convert 350. kPa to atm, and mm Hg.

$$350.\ \cancel{\text{kPa}} \times \frac{1\ \text{atm}}{101.325\ \cancel{\text{kPa}}} = \mathbf{3.45\ atm} \qquad 350.\ \cancel{\text{kPa}} \times \frac{760\ \text{mm Hg}}{101.325\ \cancel{\text{kPa}}} = 2625.2\ \text{mm Hg} = \mathbf{2.62 \times 10^3\ mm\ Hg}$$

Gaseous Pressure and Temperature in Kinetic Molecular Theory

1. **The smaller the volume, the higher the pressure**. This is there is less space for the molecules to move about in the container. Hence, each molecule will exert more force on the wall of the container, causing a higher pressure.
2. **The higher the temperature, the higher the pressure**. When the temperature is high, it means molecules move faster and have a higher average kinetic energy. With more force they can exert on the wall of the container, the pressure increases.

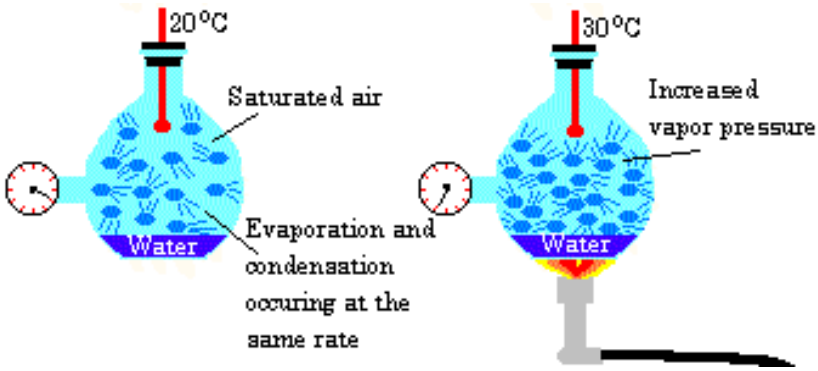
10.2: The Nature of Liquids

Vaporization: - the physical process of changing a pure substance from a liquid phase to a gaseous phase.

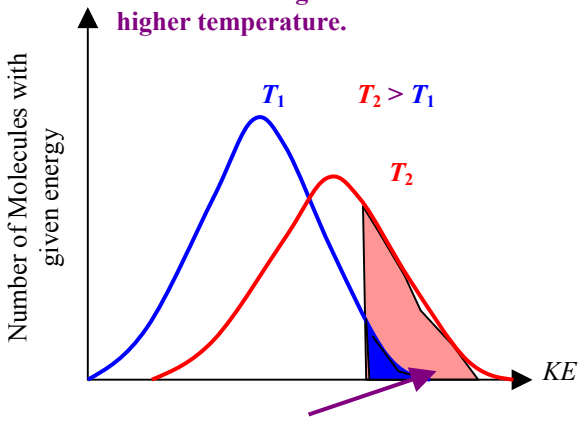
Evaporation: - vaporization below the boiling point.
- occurs because some molecules have attained sufficient energy to leave the liquid surface.
- as the temperature increases, more molecules will have the threshold energy to leave the surface of the liquid.
- when evaporation occurs, there is a cooling effect because it takes energy from the surrounding for molecules to move fast enough to escape from the liquid's surface.

Vapour Pressure: - the pressure existed above a liquid when its rate of evaporation is the same as the rate of its condensation.
- sometimes refers to as **equilibrium vapour pressure** because **equilibrium means the same rate of a two-way process**.
- in general, for a particular compound, **as temperature increases**, **the vapour pressure increases**.
- This is because the **higher the temperature**, **kinetic energy increases** and **more lquid molecules will overcome** the **intermolcular forces (force of attraction between molecules)** to become a gas. Therby, **increasing vapour presuure**.

Temperature ↑, Vapour Pressure ↑



Note that more molecules can overcome the intermolecular forces needed to be in gaseous state at a higher temperature.



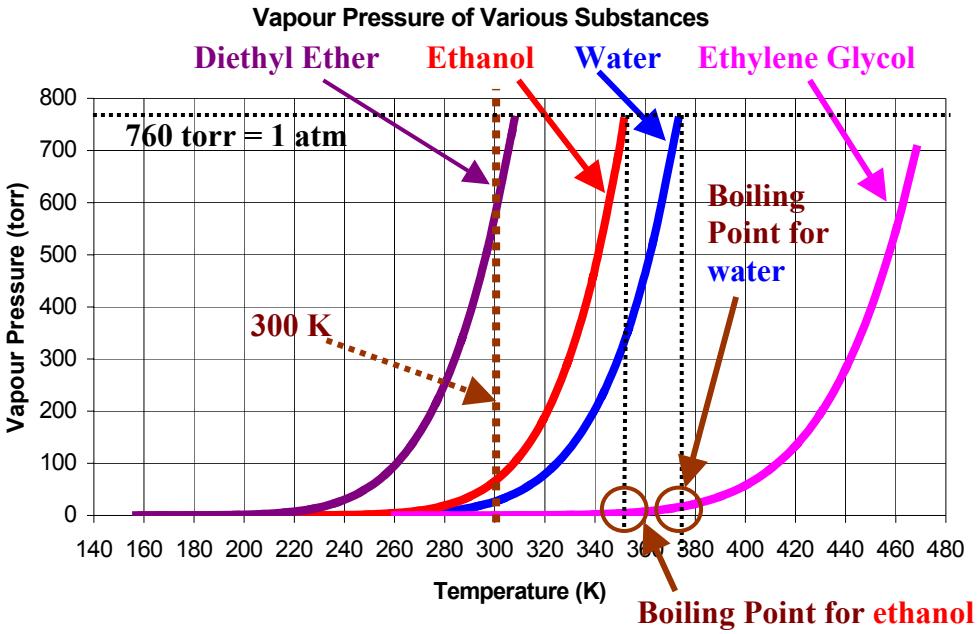
Energy required to overcome intermolecular forces.

Boiling: - vaporization at the boiling point.

Boiling Point (Bp): - the temperature at which the vapour pressure equals to the outside external pressure (like atmospheric pressure at sea level).
- **the lower the atmospheric pressure, the lower the boiling point**. This is because, there is less external air molecules pushing down on the surface of the liquid. Hence, the molecules under the liquid surface find it easier to attain that threshold energy to escape.

Normal Boiling Point: - the boiling point of a substance at 1 atm or 101.325 kPa (at sea level altitude).

We can use this graph to find the boiling points of various substances by moving across from a desired vapour pressure (which is equal to external air pressure at boiling). The temperature corresponds to that particular pressure for a certain substance is its boiling point at the air pressure.



Note: For substances, when P_{vap} = 1 atm = 760 torr, the corresponding temperature represents the boiling point.

10.4: Changes of State

Phase Diagrm: - a digaram with axes of Presuure versus Temperature to illustrate all three phases of a substance.
- it is more useful than a heating curve because boiling and melting points are pressure dependent, and the user now can see where the melting and boiling point are at any pressure. The curve of the vapour pressure versus temperature of a liquid is only a line on the phase diagram.
- it allows the user to identify the phase of a substance at a certain pressure and temperature.

Solid-Liquid Line: - a line that outlines the temperatures and the corresponding pressures where solid and liquid phases coexist.
- used to find the **melting point** of a substance at any given pressure.

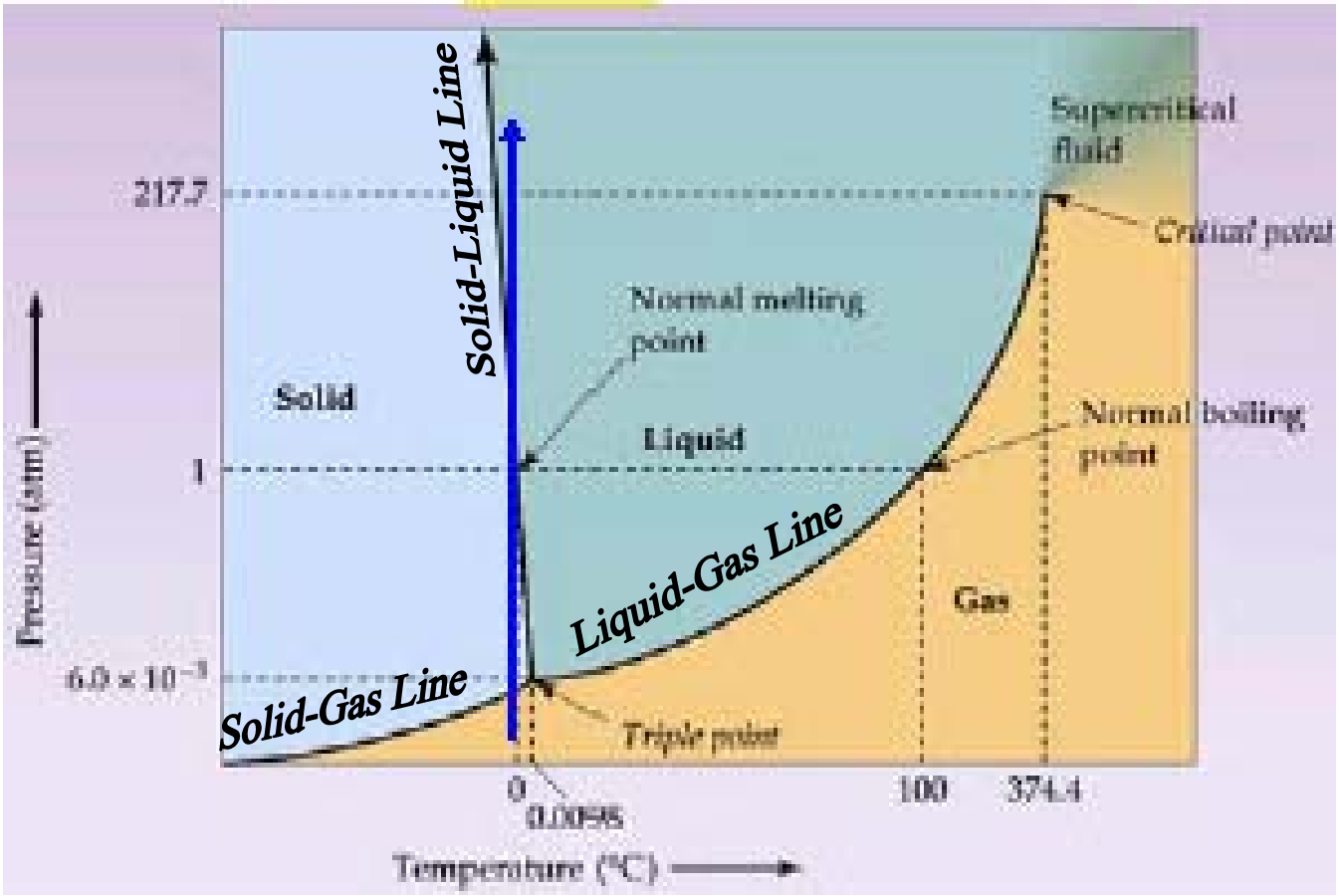
Liquid-Gas Line: - a line that outlines the temperatures and the corresponding pressures where liquid and gas phases coexist.
- used to find the **boiling point** of a substance at any given pressure.

Solid-Gas Line: - a line that outlines the temperatures and the corresponding pressures where solid and gas phases coexist.
- used to find the **sublimation point** of a substance at any given pressure.

Triple Point: - a point indicating the pressure and temperature where all solid-liquid, liquid solid, and solid-gas line meet.
- this is the **pressure and temepertaure conditions where all three phases can coexist**.

Critical Point: - a point at the end of the liquid-gas line which indicates an intermediate "fluid" region where liquid can be coverted to gas instantaneously without waiting for phase change .
- the temperature at critical point is called critical temperature and the pressure at critical point is called critical pressure.
- at this region (beyond the critical pressure and temperature), the liquid will become a gas withoutout going through a change of state.

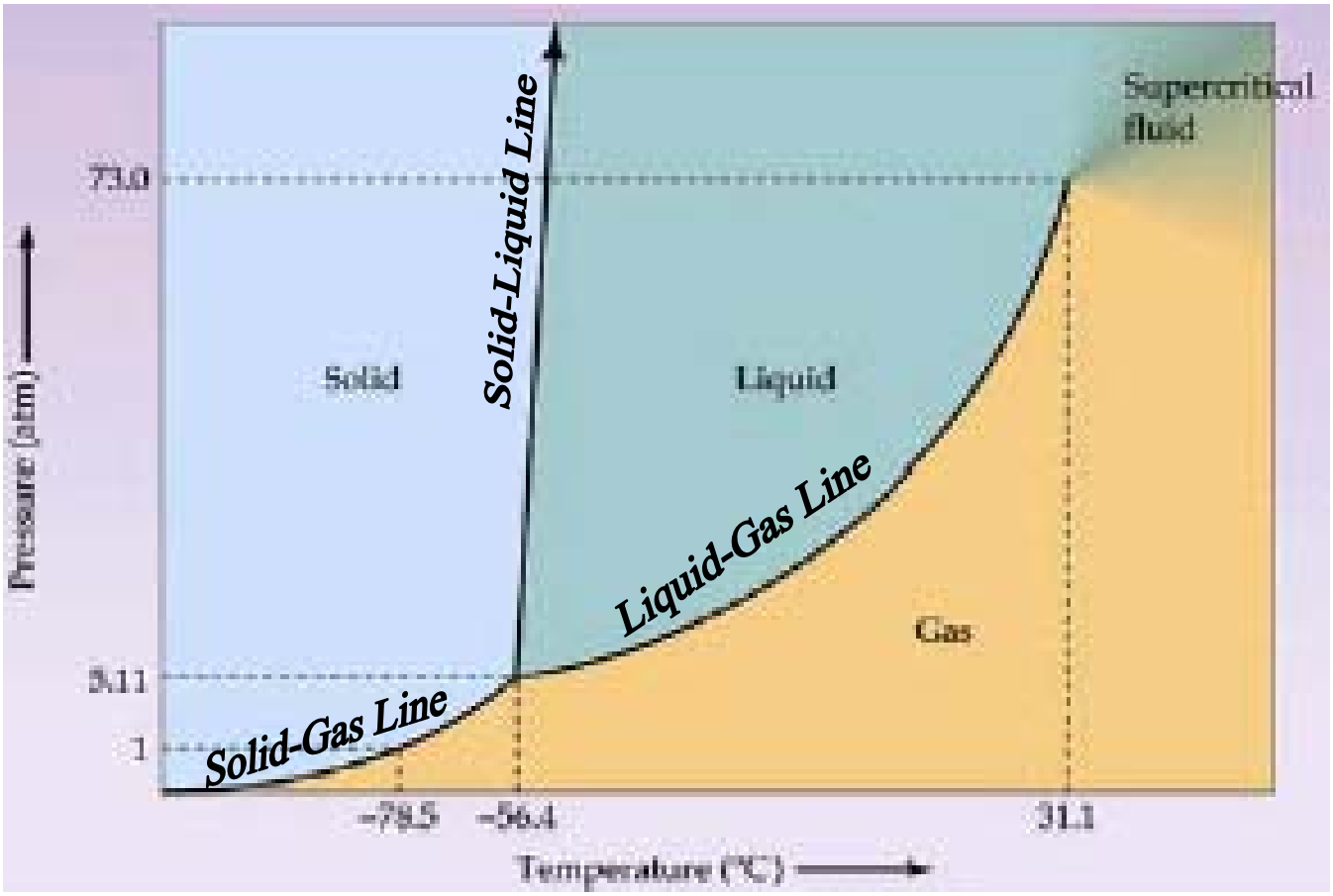
Phase Diagram of Water



From the phase diagram of water:

- at 1 atm, water has a normal melting point of 0°C and a normal boiling point at 100°C.
- there is no normal sublimation point because the solid-gas line ends well under 1 atm.
- the negative slope of the solid-liquid line indicates that as pressure increases, the melting point of water decreases.
- the positive slope of the liquid-gas line illustates that as pressure increases, the boiling point of water also increases.
- the positive slope of the solid-gas line shows that as pressure increases, the sublimation point of water increases as well.
- the triple point, where all three phases can coexist, is at 0.0098°C and 6.0×10^{-3} atm (4.56 torr).
- the critical point of water is at 374.4°C and 217.7 atm. Beyond this critical temperature and pressure, water will become gaseous instantenously.

Phase Diagram of Carbon Dioxide



From the Phase Diagram of Carbon Dioxide:

- at 1 atm, carbon dioxide has a normal sublimation point of –78.5°C.
- there is no normal melting and boiling points because carbon dioxide's solid-gas line is well above normal atmospheric condition of 1 atm.
- the positive slope of the solid-liquid line indicates that as pressure increases, the melting point of carbon dioxide increases.
- the positive slope of the liquid-gas line illustates that as pressure increases, the boiling point of carbon dioxide also increases.
- the positive slope of the solid-gas line shows that as pressure increases, the sublimation point of carbon dioxide increases as well.
- the triple point, where all three phases can coexist, is at –56.4°C and 5.11 atm (518 kPa or 3884 torr).
- the critical point of carbon dioxide is at 31.1°C and 73 atm. Beyond this critical temperature and pressure, liquid carbon dioxide will become gaseous instantaneously.

Assignment

10.1 pg. 271 #1, 2; pg. 272 #3 to 7
10.2 pg. 279 #8 to 13
10.4 pg. 286 #17 to 19
Ch 10 Review: pg. 289 #20 to 42, 46 and 47

Chapter 12: The Behaviour of Gases

12.1: The Properties of Gases

Properties of Gases:

1. **Compressibility**: - a main property of gas where the amount of volume can decrease under increase pressure.

2. **Lack of Particle Interaction**: - Unlike liquids and solids, gas particles have no attractive and repulsive forces between them as assumed in the kinetic theory.

3. **Rapid and Constant Motion**: - gas particles move in a straight line and independent of each other.
   - at collisions, gas particles are completely **elastic** (total kinetic energy remains constant before and after the collision).

Variables to Describe a Gas:

1. **Pressure (*P*)**: - the amount of force per unit of area, measures in **kiloPascal (kPa)**.
   - in a pressurized container, a pressure can be felt as the particles pushed on the inside wall of the container.

2. **Volume (*V*)**: - the amount of space the gas is occupied; measures in **Litre (L)**.

3. **Temperature (*T*)**: - the average of kinetic energy of the gas; measures in **Kelvin (K)**.

4. **Moles (*n*)**: - the amount of gas particle in a closed system; measures in **moles (mol)**.

12.2: Factors Affecting Gas Pressure

1. **Amount of Gas**: as amount of gas particles increases, there are more particles pushing on the inside wall of the container, causing an increase in pressure.

   Amount of Gas ↑ Pressure ↑

2. **Volume**: - as the volume decreases, there is less room for the gas particles to move about. This causes the particles to push "harder" onto the inside wall of the container, causing an increase in pressure.

   Volume ↑ Pressure ↓

3. **Temperature**: - as the temperature increases, gas particles move faster and push harder on the inside wall of the container, causing an increase in pressure.

   Temperature ↑ Pressure ↑

7.2: Mole-Volume Relationship

1. **Standard Temperature and Pressure (STP)**: - the amount of any gas at 0°C and 101.325 kPa (Earth's atmospheric pressure at sea level).

2. **Standard Ambient Temperature and Pressure (SATP)**: - the amount of any gas at 25°C and 100 kPa.

STP = 22.4 L/mol @ 0°C and 101.325 kPa (1 atm)
SATP = 24.8 L/mol @ 25°C and 100 kPa

Note: The amount of gas is determined by temperature, pressure and volume. The type of gas particles have no effect on these variables. (*Avogadro's Hypothesis)*

Example 1: Determine the amount of oxygen gas in a 5.00 L container under STP and SATP.

a. STP

STP = 22.4 L/mol

$n = 5.00\ \cancel{\text{L}} \times \frac{1\ \text{mol}}{22.4\ \cancel{\text{L}}}$ **$n = 0.223$ mol**

b. SATP

SATP = 24.8 L/mol

$n = 5.00\ \cancel{\text{L}} \times \frac{1\ \text{mol}}{24.8\ \cancel{\text{L}}}$ **$n = 0.202$ mol**

Example 2: Determine the volume of 3.50 g of nitrogen gas under STP and SATP.

a. STP

$n = \frac{m}{M} = \frac{3.50\ \text{g}}{28.014\ \text{g/mol}} = 0.1249375312$ mol

STP = 22.4 L/mol

$V = (0.1249375312\ \cancel{\text{mol}})(22.4\ \text{L}/\cancel{\text{mol}})$

$V = 2.80$ L

b. SATP

$n = \frac{m}{M} = \frac{3.50\ \text{g}}{28.014\ \text{g/mol}} = 0.1249375312$ mol

SATP = 24.8 L/mol

$V = (0.1249375312\ \cancel{\text{mol}})(24.8\ \text{L}/\cancel{\text{mol}})$

$V = 3.10$ L

Example 3: Calculate the molar mass of a gaseous compound containing carbon and hydrogen if its density is 0.645 g/L at SATP.

SATP = 24.8 L/mol
$D = 0.645$ g/L
M = ? (g/mol)

$M = D \times (\text{SATP})$
$M = (0.645\ \text{g}/\cancel{\text{L}})\ (24.8\ \cancel{\text{L}}/\text{mol})$
$M = 15.996$ g/mol

$M = 16.0$ g/mol

Example 4: Calculate the density in g/L of a sulfur dioxide gas at STP.

STP = 22.4 L/mol
$M = 64.06$ g/mol (SO_2)
D = ? (g/L)

$M = D \times (\text{STP})$

$D = \frac{M}{(\text{STP})} = \frac{(64.06\ \text{g}/\cancel{\text{mol}})}{(22.4\ \text{L}/\cancel{\text{mol}})} = 2.859821429$ g/L

$D = 2.86$ g/L

Assignment

12.1 pg. 328 #1 to 4
12.2 pg. 332 #5 to 9
7.2 pg. 184 #20, 21; pg. 185 #22, 23; pg. 186 #25, 26, 28; pg. 348 #31 to 33; pg. 349 #34 to 36; pg. 353 #40 and 41 and (Redo all 7.2 questions except #28 using SATP)

Answers to 7.2 questions using SATP

pg. 184 to 186
#20a. 0.0794 L b. 23.8 L c. 91.8 L
#21a. 2.71 mol b. 0.0355 mol c. 40.3 mol
#22. 88.8 g/mol #23. 3.38 g/L #25b. 1.33 L Cl_2 c. 11.6 g CH_4
#26. The volume is the same as the number of moles in any type of gas, but the mass will be different because of their different molar masses.

pg. 349 and pg. 353

#34. 1.66 L #35 4.96 L #36 85.7 L #40 $n = \frac{m}{M}$; $V = n \times 24.8$ L/mol (SATP)

#41a. 42 L b. 0.45 L c. 6.2 kL

12.3: The Gas Laws

1. **Boyle's Law**: - at a constant temperature, **pressure is inversely proportional to the volume**.

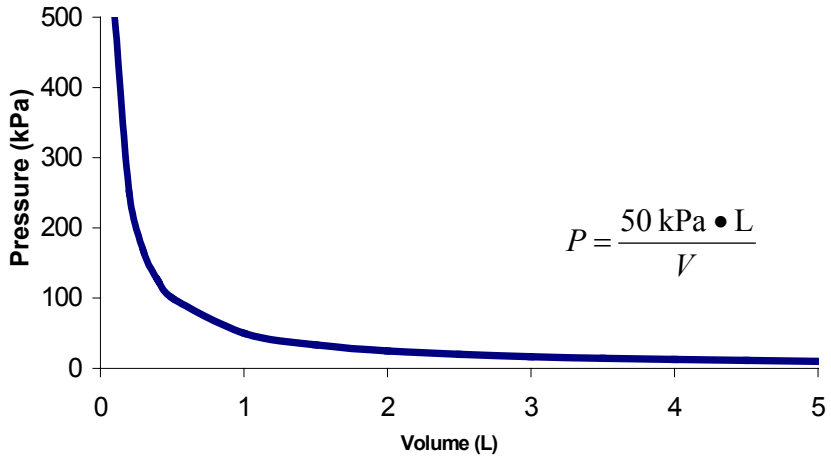
$$P \propto \frac{1}{V}$$

$$P = \frac{k}{V} \quad \text{or} \quad PV = k$$

k = constant of proportion

Volume ↑ Pressure ↓

Boyle's Law (Constant Temperature)



Boyle's Law

$$P_1V_1 = P_2V_2$$

P_1 = Pressure of Initial Condition V_1 = Volume of Initial Condition
P_2 = Pressure of Final Condition V_2 = Volume of Final Condition

Example 1: A gas cylinder changed its volume from 2.50 L to 6.25 L. If it were at 101.325 kPa initially, what would be its final pressure?

$P_1 = 101.325$ kPa
$P_2 = ?$
$V_1 = 2.50$ L
$V_2 = 6.25$ L

$$P_1V_1 = P_2V_2$$

$$\frac{P_1V_1}{V_2} = P_2$$

$$P_2 = \frac{(101.325\,\text{kPa})(2.50\,\cancel{\text{L}})}{(6.25\,\cancel{\text{L}})}$$

As Volume ↑, Pressure ↓

$P_2 = 40.5$ kPa

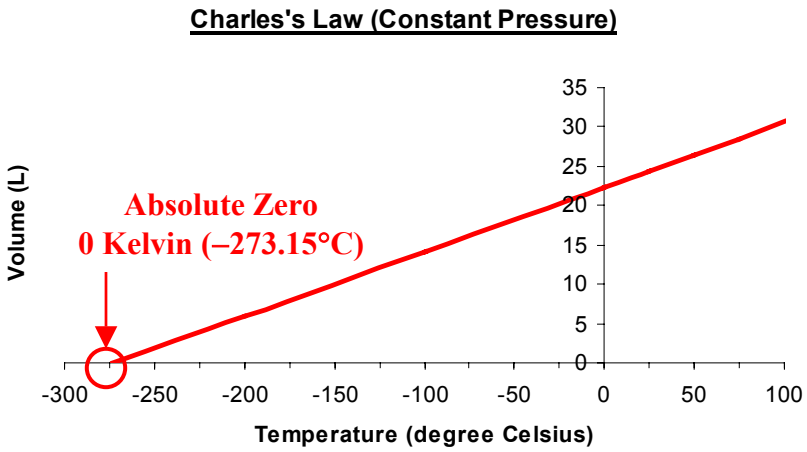
2. **Charles's Law**: - at constant pressure, **volume is directly proportional to the temperature**.

$$V \propto T$$

$$V = kT \quad \text{or} \quad \frac{V}{T} = k$$

k = constant of proportion

Temperature ↑ Volume ↑



Charles's Law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

T_1 = Temperature of Initial Condition
T_2 = Temperature of Final Condition
V_1 = Volume of Initial Condition
V_2 = Volume of Final Condition

Example 2: A balloon is has a volume of 3.25 L at 25.0°C. Determine the volume of the same balloon when the temperature is dropped to 5.00°C.

$V_1 = 3.25$ L
$T_1 = 25.0°C = 298.15$ K
$V_2 = ?$
$T_2 = 5.00°C = 278.15$ K

(Change °C to K)

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$\frac{V_1T_2}{T_1} = V_2$$

$$V_2 = \frac{(3.25\,\text{L})(278.15\,\cancel{\text{K}})}{(298.15\,\cancel{\text{K}})}$$

As Temp ↓, Volume ↓

$V_2 = 3.03$ L

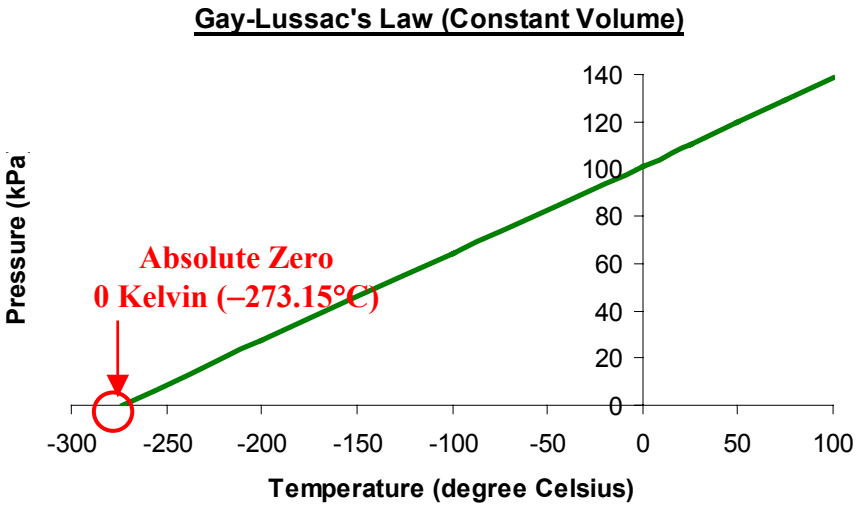
3. Gay-Lussac's Law: - at constant volume, **pressure is directly proportional to the temperature**.

$P \propto T$

$P = kT$ or $\frac{P}{T} = k$

k = constant of proportion

Temperature ↑ Pressure ↑



Gay-Lussac's Law

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

T_1 = Temperature of Initial Condition
T_2 = Temperature of Final Condition
P_1 = Pressure of Initial Condition
P_2 = Pressure of Final Condition

Example 3: A canister is has a pressure of 8.00 atm at 15.0°C. Calculate its pressure if the temperature was to increase to 100.0°C. **(Don't do this at home!)**

$P_1 = 8.00$ atm
$T_1 = 15.0°C = 288.15$ K
$P_2 = ?$
$T_2 = 100.0°C = 373.15$ K

(Change °C to K)
(P can be in atm)

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

$$\frac{P_1 T_2}{T_1} = P_2$$

$$P_2 = \frac{(8.00\,\text{atm})(373.15\,\cancel{\text{K}})}{(288.15\,\cancel{\text{K}})}$$

As Temp ↑, Pressure ↑

$P_2 = 10.4$ atm

4. Combined Gas Law: - a formula that summarizes Boyle's Charles's and Guy-Lussac's Gas Laws.
- allow the user of the formula to determine the change in conditions of the same amount of gas.

Combined Gas Law

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

P_1 = Pressure of Initial Condition
V_1 = Volume of Initial Condition
T_1 = Temperature of Initial Condition
P_2 = Pressure of Final Condition
V_2 = Volume of Final Condition
T_2 = Temperature of Final Condition

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Example 4: A high altitude weather balloon has a volume of 57.2 L at SATP. Determine its volume at its maximum height where the pressure is 37.0 kPa and the temperature is –45.0°C.

$V_1 = 57.2$ L

At SATP:
$P_1 = 100$ kPa
$T_1 = 25.0°C = 298.15$ K

$V_2 = ?$
$P_2 = 37.0$ kPa
$T_2 = -45.0°C = 228.15$ K

(Change °C to K)

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\frac{P_1V_1T_2}{T_1P_2} = V_2$$

$$V_2 = \frac{(100\ \cancel{\text{kPa}})(57.2\ \text{L})(228.15\ \cancel{\text{K}})}{(298.15\ \cancel{\text{K}})(37.0\ \cancel{\text{kPa}})} = 118.2986978\ \text{L}$$

$V_2 = 118$ L

Because P ↓ much more than T ↓, the final Volume ↑.

Assignment

12.3 pg. 335 #10, 11; pg. 337 #12, 13; pg. 338 #14; pg. 339 #15; pg. 340 #16 to 21

12.4: Ideal Gases

Ideal Gas Law: - a formula that relates pressure, volume, amount, and temperature of an **ideal gas** (gaseous volume does not account for total particles volumes) at one specific condition.

Ideal Gas Law

$$PV = nRT$$

P = Pressure (kPa) **V = Volume (L)**
n = Amount of Gas (mol) **T = Temperature (K)**
R = Gas Constant = 8.314 (L • kPa)/(K • mol)

Example 1: Determine the mass of propane if it is in a 200 L container at 15.0°C and at 32.0 atm.

$V = 200$ L
$T = 15.0°C = 288.15$ K

$$P = 32.0\ \cancel{\text{atm}} \times \frac{101.325\ \text{kPa}}{1\ \cancel{\text{atm}}}$$

$P = 3242.4$ kPa
[need to change atm to kPa because R is in (L • **kPa**)/(K • mol)]

$R = 8.314$ (L • kPa)/(K • mol)

$m = ?$
$n = ?$ (need to find n first)

$$PV = nRT$$

$$\frac{PV}{RT} = n$$

$$n = \frac{(3242.4\ \cancel{\text{kPa}})(200\ \cancel{\text{L}})}{\left(8.314\ \frac{\cancel{\text{L}} \bullet \cancel{\text{kPa}}}{\cancel{\text{K}} \bullet \text{mol}}\right)(288.15\ \cancel{\text{K}})}$$

$n = 270.670642$ mol

For propane, C_3H_8, $M = 44.11$ g/mol

$m = nM$
$m = (270.670642\ \cancel{\text{mol}})(44.11\ \text{g}/\cancel{\text{mol}})$
$m = 11939.28202$ g

$m = 1.19 \times 10^4$ g = 11.9 kg

Example 2: It is known that air is less dense at higher altitude. Calculate the density of pure oxygen gas near the top of Mount Everest where the temperature is –30.0°C and 31.0 kPa. (By comparison, pure oxygen has a density of 1.29 g/L at SATP.)

$T = -30.0°C = 243.15\ K$
$P = 31.0\ kPa$
$R = 8.314\ (L \bullet kPa)/(K \bullet mol)$
$M = 32.00\ g/mol$

$D = \frac{m}{V} = ?$ (Substitute $\frac{m}{M}$ for n. Then, solve for $\frac{m}{V}$.)

$PV = nRT$

$PV = \frac{m}{M}RT$

$\frac{PM}{RT} = \frac{m}{V}$

$$D = \frac{PM}{RT} = \frac{(31.0\ \cancel{kPa})(32.0\ g/\cancel{mol})}{\left(8.314\ \frac{L \bullet \cancel{kPa}}{\cancel{K} \bullet \cancel{mol}}\right)(243.15\ \cancel{K})} = 0.4907127905\ g/L$$

$D = 0.491\ g/L$

Real Gas: - a gas where its properties deviate from an ideal gas.
- **when we account for the fact that gaseous particles have volume, and they actually have attraction between particles (intermolecular bonds), the ideal gas law become less accurate**.
- This ***happens mainly when pressure is extremely high, volume is small and temperature is low***. At such conditions, the small volume will make the particle size an important matter. This is the same for attraction forces between particles because they are moving a lot slower.

Assignment
12.4 pg. 342 #22, 23; pg. 343 #24, 25; pg. 346 #26 to 30

12.5: Gas Molecules: Mixtures and Movements

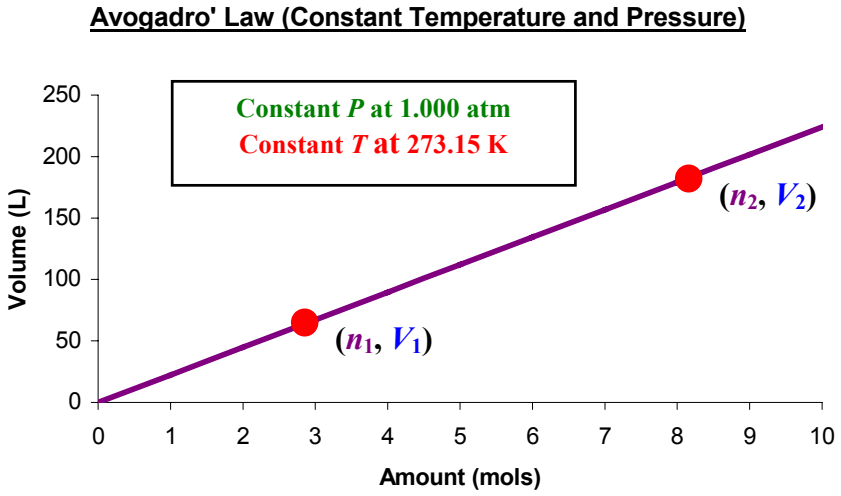
Avogadro's ~~Hypothesis~~ LAW: - at constant pressure and temperature, **volume is directly proportional to amount of moles of gas present**.

$V \propto n$

$V = kn$ or $\frac{V}{n} = k$

k = constant of proportion

Moles ↑ Volume ↑



Avogadro's Law

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

n_1 = Amount of Moles at Initial Condition
n_2 = Amount of Moles at Final Condition
V_1 = Volume at Initial Condition
V_2 = Volume at Final Condition

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Example 1: An excess amount of solid phosphorus is reacted with 9.00 mol of chlorine gas at 223.2 L to produce 3.00 mol of phosphorus trichloride gas under constant temperature and pressure. Assuming a complete reaction, what is the volume of phosphorus trichloride produced?

$n_1 = 9.00$ mol (Cl_2)
$V_1 = 223.2$ L

$n_2 = 6.00$ mol (PCl_3)
$V_2 = ?$

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

$$\frac{n_2 V_1}{n_1} = V_2$$

As Moles ↓, Volume ↓

$$V_2 = \frac{(6.00 \cancel{\text{mol}})(223.2\ \text{L})}{(9.00 \cancel{\text{mol}})}$$

$V_2 = 149.\ \text{L}$

Partial Pressure: - the pressure exerted by one component of a gas mixture.

Dalton's Law of Partial Pressures: - the total pressure in a gas mixture is equal to the sum of the pressures exerted by individual gaseous components.

Dalton's Law of Partial Pressures

$$P_{Total} = P_1 + P_2 + P_3 + \ldots$$

Example 2: Write the equations for partial pressures for air, which is composed of N_2, O_2, CO_2, Ar, and other trace gases (78.084%, 20.947%, 0.033%, 0.934%, 0.002% by volume respectively). Determine the partial pressures of these gases for 1 atm of air in kPa.

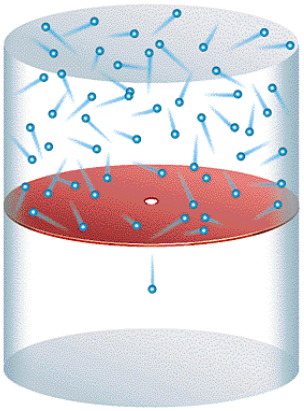
$$P_{total} = P_{N_2} + P_{O_2} + P_{CO_2} + P_{Ar} + P_{trace\ gases} = 101.325\ \text{kPa}$$

$P_{N_2} = (0.780\ 804)(101.325\ \text{kPa})$ $\quad P_{N_2} = 79.115\ \text{kPa}$
$P_{O_2} = (0.209\ 47)(101.325\ \text{kPa})$ $\quad P_{O_2} = 21.225\ \text{kPa}$
$P_{CO_2} = (0.000\ 33)(101.325\ \text{kPa})$ $\quad P_{CO_2} = 0.033\ \text{kPa}$
$P_{Ar} = (0.009\ 34)(101.325\ \text{kPa})$ $\quad P_{Ar} = 0.946\ \text{kPa}$
$P_{trace\ gases} = (0.00\ 002)(101.325\ \text{kPa})$ $\quad P_{trace\ gases} = 0.002\ \text{kPa}$

Example 3: The total pressure of the gas collected over a closed ended column partially with water is 135.20 kPa. If the vapour pressure of the water at that temperature is 3.15 kPa, what is the actual pressure of a gas?

$$P_{Total} = P_{water\ vapour} + P_{actual\ gas}$$

$P_{actual\ gas} = P_{Total} - P_{water\ vapour} = 135.20\ \text{kPa} - 3.15\ \text{kPa}$ $\quad P_{actual\ gas} = 132.05\ \text{kPa}$



Effusion: - when gas passes through from one container to the next through a small hole.

Rate of Effusion: - the speed of which a gas pass through a small hole between two containers.
- it is the same as the root mean square velocity, u_{rms}. This is because the small hole on the separting wall of the two containers allow one gas particle to pass through at a time. Therefore, the rate of effusion = u_{rms}
- **rate of effusion is generally measures in mL/min**

Graham's Law of Effusion: - the ratio between the effusion rate of two gases.

Graham's Law of Effusion

$$\frac{\text{Rate of Effusion for Gas 1}}{\text{Rate of Effusion for Gas 2}} = \frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$

M = Molar Mass in g/mol or kg/mol

M ↓ (g/mol), Effusion Rate ↑ (mL/min)

Example 4: Calculate the ratio of effusion rates between $NH_{3\,(g)}$ and $HCl_{\,(g)}$ under the same temperature.

$M_{HCl} = 36.46$ g/mol
$M_{NH_3} = 17.04$ g/mol
$\frac{r_{NH_3}}{r_{HCl}} = ?$

$$\frac{r_{NH_3}}{r_{HCl}} = \sqrt{\frac{M_{HCl}}{M_{NH_3}}} = \sqrt{\frac{36.46\ \cancel{\text{g/mol}}}{17.04\ \cancel{\text{g/mol}}}}$$

$\frac{r_{NH_3}}{r_{HCl}} = 1.463$

NH_3 effuse about 1.5 times faster than HCl.

Example 5: Determine the effusion rate F_2 through a porous barrier if the effusion rate of N_2 through the same barrier is 274. mL/min.

$r_{N_2} = 274$ mL/min
$M_{F_2} = 38.00$ g/mol
$M_{N_2} = 28.02$ g/mol
$r_{F_2} = ?$

$$\frac{r_{F_2}}{r_{N_2}} = \frac{\sqrt{M_{N_2}}}{\sqrt{M_{F_2}}} \qquad r_{F_2} = \frac{(r_{N_2})\sqrt{M_{N_2}}}{\sqrt{M_{F_2}}} = \frac{(274\text{ mL/min})\sqrt{28.02\ \cancel{\text{g/mol}}}}{\sqrt{38.00\ \cancel{\text{g/mol}}}}$$

Since F_2 is bigger than N_2, r_{F_2} is less than r_{N_2}

$r_{F_2} = 235.$ mL/min

Assignment

12.5 pg. 348 #31 to 33; pg. 349 # 34 to 36; pg. 351 #37, 38; pg. 351 #37, 38; pg. 353 #39 to 44
Ch 12 Review: pg. 356 #45 to 68

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Chapter 18: Solutions (Part 2)

18.1: Properties of Solution (Part 2)

Henry's Law: - states that the solubility of gas in a solution is directly proportional to the pressure above the solution.

$$S \propto P$$

$$S = kP \quad \text{or} \quad \frac{S}{P} = k$$

k = constant of proportion

Pressure ↑ Solubility of Gas ↑

Henry's Law

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$

S_1 = Solubility at Initial Condition
P_1 = Pressure of Initial Condition
S_2 = Solubility at Final Condition
P_2 = Pressure of Final Condition

Example 1: Some health food vendors claim that drinking oxygenated beverage will increase physical performance. Suppose an oxygenated drink has a oxygen solubility of 3.17 g/L is under 7.50 atm of pressure. If the partial pressure of oxygen in the atmosphere 0.20947 atm, what will be the solubility of oxygen in the beverage if it was left open for a long time?

$P_1 = 7.50$ atm
$S_1 = 3.17$ g/L
$P_2 = 0.20947$
$S_2 = ?$

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$

$$\frac{S_1 P_2}{P_1} = S_2$$

As Pressure ↓, Solubility ↓

$$S_2 = \frac{(3.17\ \text{g/L})(0.20947\ \cancel{\text{atm}})}{(7.50\ \cancel{\text{atm}})}$$

$P_2 = 0.0885$ g/L

18.3 and 18.4: Colligative Properties of Solutions and their Calculations

Molality: - moles of solute per kilogram of solvent (m = mol/kg).

$$\text{Molality} = \frac{n_{solute}}{m_{solvent}}$$

Molality = m (mol/kg) n = moles of Solute $m_{solvent}$ = Mass of Solvent in kg

Mole Fraction (χ): - the ratio of the moles of a solute and the moles of the entire solution (solute and solvent).

Mole Fraction

$$\chi_A = \frac{n_A}{n_{Total}}$$

n_A = moles of component A n_{Total} = Total Moles of Solution = $n_A + n_B + n_C + \ldots$

Example 1: 40.0 g of 1-propanol with a density of 0.803 g/mL is dissolved in 150 mL of water. Describe the composition of the solution by

a. molarity b. molality c. mass percent d. mole fraction

a. Molarity

$$n = \frac{40.0\text{ g}}{60.0962\text{ g/mol}} = 0.6655994888\text{ mol } C_3H_7OH$$

$$V_{propanol} = \frac{40.0\text{ g}}{0.803\text{ g/mL}} = 49.8132005\text{ mL}$$

$V_{Total} = 150\text{ mL} + 49.8132005\text{ mL}$

$V_{Total} = 199.8132005\text{ mL} = 0.1998132005\text{ L}$

$$C_{propanol} = [C_3H_7OH] = \frac{0.6655994888\text{ mol}}{0.1998132005\text{ L}}$$

$[C_3H_7OH] = 3.33$ mol/L

b. Molality

$$n = \frac{40.0\text{ g}}{60.0962\text{ g/mol}} = 0.6655994888\text{ mol } C_3H_7OH$$

$m_{solvent} = 150\text{ g} = 0.150\text{ kg}$

(Assume 1 g ≈ 1 mL H_2O)

$$\textbf{Molality} = \frac{n_{solute}}{m_{solvent}}$$

$$\text{Molality} = \frac{0.6655994888\text{ mol}}{0.150\text{ kg}}$$

Molality = 4.44 *m*

c. Mass Percent

$m_{propanol} = 40.0\text{ g}$

$m_{solution} = m_{propanol} + m_{water} = 40.0\text{ g} + 150\text{ g}$

$m_{solution} = 190\text{ g}$

$$\textbf{Mass Percent} = \frac{m_{solute}}{m_{solution}} \times 100\%$$

$$\text{Mass Percent} = \frac{40.0\text{ g}}{190\text{ g}} \times 100\%$$

Mass Percent = 21.1%

d. Mole Fraction

$$n = \frac{40.0\text{ g}}{60.0962\text{ g/mol}} = 0.6655994888\text{ mol } C_3H_7OH$$

$$n_{water} = \frac{150\text{ g}}{18.0158\text{ g/mol}} = 8.326024934\text{ mol } H_2O$$

$n_{Total} = n_{propanol} + n_{water}$

$n_{Total} = 0.6655994888\text{ mol} + 8.326024934\text{ mol}$

$n_{Total} = 8.991624422\text{ mol}$

$$\chi_{propanol} = \frac{n_{propanol}}{n_{Total}} = \frac{0.6655994888\text{ mol}}{8.991624422\text{ mol}}$$

$\chi_{propanol} = 0.0740$

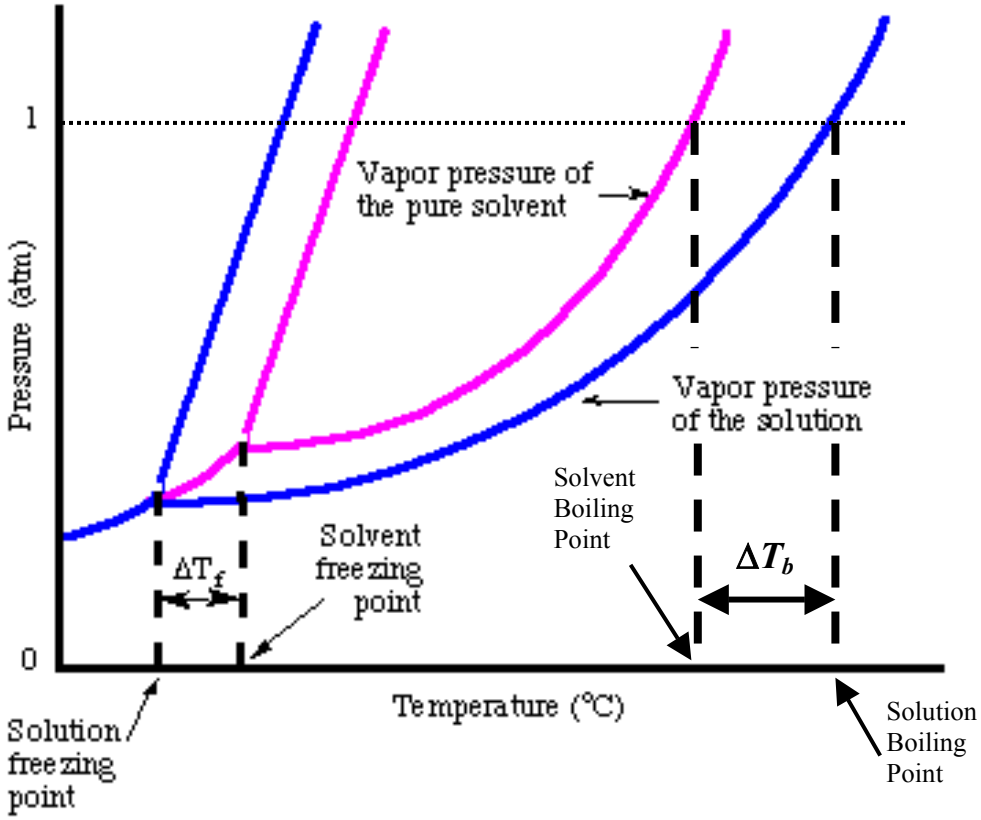
Colligative Properties: - changes in physically properties of a pure substance as it is mixed with a solute.
- these include freezing-point depression, boiling-point elevation and osmotic pressure.

Boiling-Point Elevation: - the presence of solute in a solution raises the boiling point of a pure solvent due to the extra intermolecular forces between them (more energy is needed to boil the solution as vapour pressure is lowered).
- the amount of temperature elevation to boil is directly proportional to the molality of the solute in the solution.
- can be used to **determine molar mass of a solute**.

Molal Boiling-Point Elevation Constant (K_b): - a constant relating the change in boiling point temperature and the molality of the solute in the solution.

Freezing-Point Depression: - the presence of solute decreases the vapour pressure of the solvent. Since ice has a higher vapour pressure than that of water in the solution, the freezing point has effectively been lowered. (In order for ice to form, vapour pressure of the ice has to be lowered than that of water).
- the amount of temperature depression to freeze is directly proportional to the molality of the solute in the solution.
- used most commonly to **determine molar mass of a solute**.

Molal Freezing-Point Elevation Constant (K_f): - a constant relating the change in freezing point temperature and the molality of the solute in the solution.



Boiling Point Elevation and Freezing Point Depression of Non-Electrolytic Solutions

$\Delta T_b = K_b \times \text{Molality}_{\text{solute}}$ $\Delta T_f = K_f \times \text{Molality}_{\text{solute}}$

ΔT_b = Change in Boiling Point Elevation (°C)
ΔT_f = Change in Freezing Point Depression (°C)
K_b = Molal Boiling-Point Constant (°C • kg/mol)
K_f = Molal Freezing-Point Constant (°C • kg/mol)
$\text{Molality}_{\text{solute}}$ = Molality of Solute (mol/kg of solvent)

Example 2: Antifreeze, ethylene glycol (CH_2OHCH_2OH), is commonly used to prevent water from freezing in the engine in cold temperature as well as overheating. The maximum temperature a radiater can reach is 120°C. Given the densities for ethylene glycol and water are 1.11 g/mL and 1.00 g/mL respectively, and K_b and K_f for water are 0.51°C • kg/mol and 1.86 °C • kg/mol,

a. determine the volume of antifreeze needed to add to 10.0 L of water to sustain the maximum radiater temperature.

b. what is the freezing point of the solution?

a.

$\Delta T_b = 120°C - 100°C = 20°C$

$K_b = 0.51°C \bullet kg/mol$

$m_{water} = 10.0 \times 10^3 \text{ mL} \times 1.00 \text{ g/mL}$

$m_{water} = 10.0 \times 10^3 \text{ g} = 10.0 \text{ kg}$

$D_{antifreeze} = 1.11 \text{ g/mL}$

$M_{antifreeze} = 62.0694 \text{ g/mol}$

$n_{antifreeze}$ = ? $m_{antifreeze}$ = ?

$V_{antifreeze}$ = ?

$$\Delta T_b = K_b \times \text{Molality}_{antifreeze} = K_b \times \frac{n_{antifreeze}}{\text{kg of water}}$$

$$n_{antifreeze} = \frac{\Delta T_b (\text{kg of water})}{K_b} = \frac{(20^\circ\cancel{C})(10.0\ \cancel{kg})}{(0.51^\circ\cancel{C} \bullet \cancel{kg}/mol)} = 392.1568627 \text{ mol}$$

$$m_{antifreeze} = (392.1568627\ \cancel{mol})(62.0694\ g/\cancel{mol}) = 24340.94117 \text{ g}$$

$$V_{antifreeze} = \frac{m_{antifreeze}}{D_{antifreeze}} = \frac{24340.94117\ g}{1.11\ g/mL} = 21928.77583 \text{ mL}$$

$V_{antifreeze}$ = 21.9 L

b.

$K_f = 1.86°C \bullet kg/mol$

$m_{water} = 10.0 \text{ kg}$

$n_{antifreeze} = 392.1568627 \text{ mol}$

ΔT_f = ?

$$\Delta T_f = K_f \times \text{Molality}_{antifreeze} = K_f \times \frac{n_{antifreeze}}{\text{kg of water}}$$

$$\Delta T_f = (1.86°C \bullet \cancel{kg}/\cancel{mol}) \left(\frac{392.1568627\ \cancel{mol}}{10.0\ \cancel{kg}} \right) = 72.94117646°C$$

New Freezing Point = 0°C – 72.94117646°C

$T_{f\,(soln)}$ = –72.9°C

Molar Mass Determination of a Solute using Freezing Point Depression

- one common way to determine the molar mass of a solute is dissolving it in water (if the solute is polar). After manipulating the freezing point depression formula, we can calculate molar mass of this solute.

$$\Delta T_f = K_f \times \text{Molality}_{solute}$$

$$\Delta T_f = K_f \times \frac{n_{solute}}{\text{kg of solvent}}$$

$$\frac{(\Delta T_f)(\text{kg of solvent})}{K_f} = n_{solute} \longrightarrow \frac{(\Delta T_f)(\text{kg of solvent})}{K_f} = \frac{m_{solute}}{Molar\ Mass_{solute}}$$

Molar Mass Determination from Freezing Point Depression of Non-Electrolytic Solutions

$$Molar\ Mass_{solute} = \frac{(K_f)(m_{solute} \text{ in grams})}{(\Delta T_f)(\text{kg of solvent})}$$

Example 3: A 20.0 g of newly synthesized enzyme can lowered the freezing point of 100.0 g CCl_4 by 4.70°C. Given that K_f for CCl_4 is 30.0°C • kg/mol, calculate the molar mass of this new enzyme.

$\Delta T_f = 4.70°C$
$K_f = 30.0°C \bullet kg/mol$
$m_{CCl_4} = 100.0\ g = 0.1000\ kg$
$m_{enzyme} = 20.0\ g$

$\boldsymbol{n_{enzyme} = ?}$
$\boldsymbol{M_{enzyme} = ?}$

$$\Delta T_f = K_f \times \text{Molality}_{enzyme} = K_f \times \frac{n_{enzyme}}{\text{kg of } CCl_4}$$

$$n_{enzyme} = \frac{\Delta T_f (\text{kg of } CCl_4)}{K_f} = \frac{(4.70^\circ\cancel{C})(0.1000\ \cancel{kg})}{(30.0^\circ\cancel{C} \bullet \cancel{kg}/mol)} = 0.0156666667\ mol$$

$$M_{enzyme} = \frac{m_{enzyme}}{n_{enzyme}} = \frac{20.0\ g}{0.0156666667\ mol} = 1276.595745\ g/mol$$

$$\boldsymbol{M_{enzyme} = 1.30 \times 10^3\ g/mol}$$

van't Hoff Factor (*i*): - is a factor that relates the number of moles of all ions in a solution per mole of solutes dissolved.
- because non-electrolytic solutions do not form ions, ***i* = 1 for all non-ionic solutions**.

Colligative Properties of Electrolytic Solutions

$$\textbf{van't Hoff Factor } (i) = \frac{n_{ions}}{n_{solute}}$$

Boiling Point Elevation: $\Delta T_b = iK_b \times \text{Molality}_{solute}$
Freezing Point Depression: $\Delta T_f = iK_f \times \text{Molality}_{solute}$

Example 4: Determine the freezing point of a saturated salt solution (35.7 g / 100 g H_2O) if the K_f for pure water is 1.86 °C • kg/mol.

$K_f = 1.86°C \bullet kg/mol$
$m_{water} = 100\ g = 0.100\ kg$
$n_{NaCl} = \dfrac{35.7\ g}{58.443\ g/mol}$
$n_{NaCl} = 0.610851599\ mol$
$i = 2$
(1 mol Na^+ and 1 mol Cl^-)
$\boldsymbol{\Delta T_f = ?}$

$$\Delta T_f = iK_f \times \text{Molality}_{NaCl} = iK_f \times \frac{n_{NaCl}}{\text{kg of water}}$$

$$\Delta T_f = (2)(1.86°C \bullet \cancel{kg}/\cancel{mol})\left(\frac{0.610851599\ \cancel{mol}}{0.100\ \cancel{kg}}\right) = 22.72367948°C$$

New Freezing Point = 0°C – 22.72367948°C

$$\boldsymbol{T_{f\,(soln)} = -22.7°C}$$

Assignment

18.1 (Part 2) pg. 507 #1, 2; pg. 508 #4
18.3 pg. 519 #24 to 27
18.4 pg. 521 #28, 29; pg. 522 #30, 31; pg. 524 #32, 33; pg. 525 #34 to 39
Ch 18 Review: pg. 528 #48, 55 to 60

Assignment

Unit 2 Review: pg. 291 #1 to 9; pg. 356–357 #69, 71, 73; pg. 359 #1, 2, 5 to 12
pg. 496–497 #47, 49, 51, 53, 56, 59; pg. 499 #2 to 7
pg. 529 #62 to 66, 68 to 77; pg. 531 #1 to 9, 14 to 17