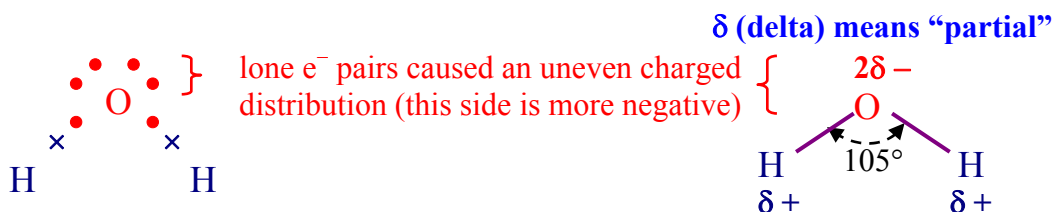
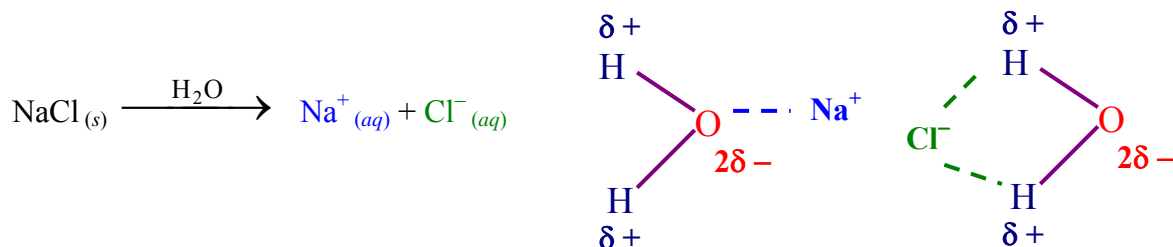


Chapter 4: Reactions in Aqueous Solutions**4.1: General Properties of Aqueous Solutions****Solute:** - the matter that is being dissolved.**Solvent:** - the matter that is doing the dissolving.**Structure of Water:**

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

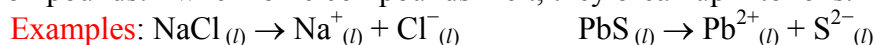
**Reason that Water is a Common Solvent:**

- Polar Molecule:** - dissolves many ionic compounds due to its ability to attract cations and anions (electrolytes).
- *Note:* **Not** all ionic compounds have high solubility (ability to dissolve) in water.

Hydration: - when ionic compound dissolves in water as water molecules surround the dissociated ions.**Example 1:** Sodium Chloride (Soluble Ionic Compounds)**Example 2:** Sulfuric Acid (Strong Acid)**Electrolytes:** - ions which have dissolved completely in water and allow the conduction of electricity.

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

- All ionic compounds containing Na^+ , NH_4^+ , NO_3^- , CH_3COO^- , ClO_3^- , or ClO_4^- .
(except RbClO_4 , CsClO_4 , AgCH_3COO , $\text{Hg}_2(\text{CH}_3\text{COO})_2$)
- Strong Acids: $\text{HClO}_4_{(aq)}$, $\text{HI}_{(aq)}$, $\text{HBr}_{(aq)}$, $\text{HCl}_{(aq)}$, $\text{H}_2\text{SO}_4_{(aq)}$, and $\text{HNO}_3_{(aq)}$
- Strong Bases: $\text{NH}_4\text{OH}_{(aq)}$, $\text{LiOH}_{(aq)}$, $\text{NaOH}_{(aq)}$, $\text{KOH}_{(aq)}$, $\text{Ca}(\text{OH})_2_{(aq)}$, $\text{Sr}(\text{OH})_2_{(aq)}$, $\text{Ba}(\text{OH})_2_{(aq)}$
- Molten Ionic Compounds: - when ionic compounds melt, they break up into ions.



b. **Weak Electrolytes**: - ionic compounds that **dissociate partially** into their ions and **conduct electricity poorly**. (Check Solubility Table in the next section.)

- Some Ionic Compounds: $\text{AgCl}_{(s)}$, $\text{PbCl}_{2(s)}$, $\text{Hg}_2\text{Cl}_{2(s)}$, $\text{HgCl}_{2(s)}$, and $\text{CuCl}_{(s)}$
- Weak Acids: $\text{HF}_{(aq)}$, $\text{HCH}_3\text{COO}_{(aq)}$, $\text{H}_2\text{SO}_{3(aq)}$, and other acids.
- Weak Bases: $\text{Mg}(\text{OH})_{2(aq)}$, $\text{Al}(\text{OH})_{3(aq)}$, $\text{NH}_3_{(aq)}$, and other bases.

Reversible Reactions: - reactions that can proceed forward and in reverse.

- when the rate of forward reaction is equalled to the rate of reverse reaction, we say that the process is at **equilibrium**. (\rightleftharpoons)
- all weak electrolyte dissociations are classified as reversible reactions.

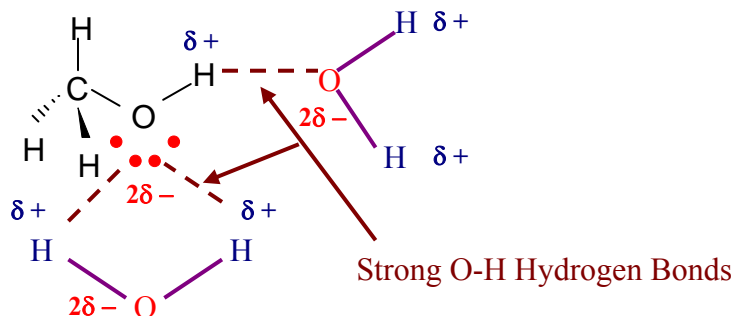
Example 3: Silver (I) Chloride (Slightly Soluble Ionic Compound) $\text{AgCl}_{(s)} \rightleftharpoons \text{Ag}^+_{(aq)} + \text{Cl}^-_{(aq)}$

Example 4: Hydrofluoric Acid (Weak Acid) $\text{HF}_{(aq)} \rightleftharpoons \text{H}^+_{(aq)} + \text{F}^-_{(aq)}$

Example 5: Magnesium Hydroxide (Weak Base) $\text{Mg}(\text{OH})_{2(s)} \rightleftharpoons \text{Mg}^{2+}_{(aq)} + 2 \text{OH}^-_{(aq)}$

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

Example 6: $\text{CH}_3\text{OH}_{(l)}$ (Methanol)



“Like-Dissolves-Like”: - polar solvents tend to dissolve polar solutes; non-polar solvents tend to dissolve non-polar solutes.

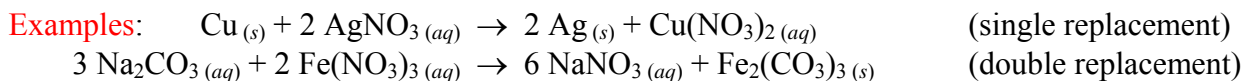
Non-Electrolytes: - soluble molecular compounds and some ionic compounds that **do not dissociate** in the solvent. **Non-electrolytes do not conduct any electricity at all**.

Examples: $\text{C}_{12}\text{H}_{22}\text{O}_{11(aq)}$ and $\text{C}_6\text{H}_{12}\text{O}_6(aq)$

(see Animation at http://www.mhhe.com/physsci/chemistry/animations/chang_7e_esp/clm2s3_4.swf)

4.2: Precipitation Reactions

Precipitation Reactions: - chemical reactions that involve a formation of an insoluble product (**precipitate**).



Metathesis Reactions: - also called double replacement reactions.

- some double replacement reactions are precipitation reactions.
- others are acid-base neutralization reactions.
- if there is no precipitate or new molecular compound form, there is no reaction. (see section on ionic equations).

Solubility: - the amount of solute that can dissolve in a given amount of solvent at a specific temperature.

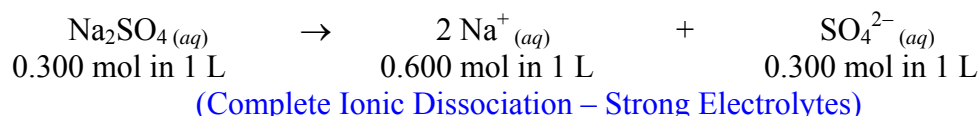
Soluble: - when a fair amount of solute can dissolve in a solvent.

Insoluble: - when a solid does not seem to be soluble by the naked eye.

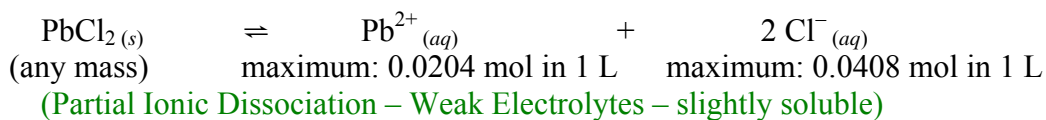
Slightly Soluble: - in reality, insoluble substances dissolve a tiny amount in water.

Examples:

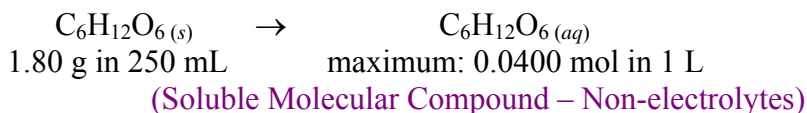
a. Dissociation of a **Soluble Ionic Compound** and Ions Concentrations.



b. Dissociation of a **Slightly Soluble Ionic Compound** and Ions Concentrations.



c. Dissolving a **Molecular Compound** and Concentrations.



General Rules for Salts (Ionic Compounds) in Water:

- All Na^+ , NH_4^+ , NO_3^- , ClO_3^- , ClO_4^- and CH_3COO^- salts are soluble (except RbClO_4 , CsClO_4 , AgCH_3COO , $\text{Hg}_2(\text{CH}_3\text{COO})_2$).
- Most F^- are soluble (except with Li^+ , Mg^{2+} , Ca^{2+} , Sr^{2+} , Ba^{2+} and Fe^{2+} , Hg_2^{2+} and Pb^{2+}).
- Most Cl^- , Br^- , and I^- salts are soluble (except with Cu^+ , Ag^+ , Hg_2^{2+} , Hg^{2+} , and Pb^{2+}).
- Most SO_4^{2-} are soluble (except with Ca^{2+} , Sr^{2+} , Ba^{2+} , Hg_2^{2+} , Pb^{2+} and Ag^+).
- Only H^+ , NH_4^+ , Na^+ , K^+ cations with PO_4^{3-} , SO_3^{2-} and CO_3^{2-} are soluble (exception Li_2CO_3 is soluble).
- Only H^+ , NH_4^+ , Li^+ , Na^+ , K^+ , Ni^{2+} , Zn^{2+} cations with IO_3^- and OOC^- are soluble (exceptions: $\text{Co}(\text{IO}_3)_2$ and $\text{Fe}_2(\text{OOC})_3$ are soluble).
- Only H^+ , NH_4^+ , Li^+ , Na^+ , K^+ , Mg^{2+} , Ca^{2+} cations with S^{2-} salts are soluble.
- Only H^+ , NH_4^+ , Li^+ , Na^+ , K^+ , Ca^{2+} , Sr^{2+} , Ba^{2+} cations with OH^- salts are soluble (these are strong bases).

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 (25°C)

Ion	H_3O^+ (H^+), Na^+ , NH_4^+ , NO_3^- , ClO_3^- , ClO_4^- , CH_3COO^-	F^-	Cl^- Br^- I^-	SO_4^{2-}	CO_3^{2-} PO_4^{3-} SO_3^{2-}	IO_3^- $\text{OOC}^-\text{COO}^{2-}$	S^{2-}	OH^-
Solubility greater than or equal to 0.1 mol/L (very soluble)	Most	most	most	most	NH_4^+ H^+ Na^+ K^+	NH_4^+ H^+ Li^+ Na^+ K^+ Ni^{2+} Zn^{2+}	NH_4^+ H^+ Li^+ Na^+ K^+ Mg^{2+} Ca^{2+}	NH_4^+ H^+ Li^+ Na^+ K^+ Ca^{2+} Sr^{2+} Ba^{2+}
Solubility less than 0.1 mol/L (slightly soluble)	RbClO_4 CsClO_4 AgCH_3COO $\text{Hg}_2(\text{CH}_3\text{COO})_2$	Li^+ Mg^{2+} Ca^{2+} Sr^{2+} Ba^{2+} Fe^{2+} Hg_2^{2+} Pb^{2+}	Cu^+ Ag^+ Hg_2^{2+} Hg^{2+} Pb^{2+}	Ca^{2+} Sr^{2+} Ba^{2+} Hg_2^{2+} Pb^{2+} Ag^+	most Exception: Li_2CO_3 is soluble	most Exceptions: $\text{Co}(\text{IO}_3)_2$ $\text{Fe}_2(\text{OOC}^-\text{COO})_3$ are soluble	most	most

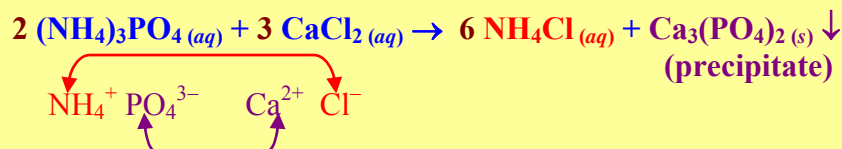
Molecular Equation: - a chemical equation where compounds are written in their chemical formulas.

Complete Ionic Equation: - a chemical equation where all compounds that are soluble are written in the ionic components (slightly soluble compounds are not separated into ions).

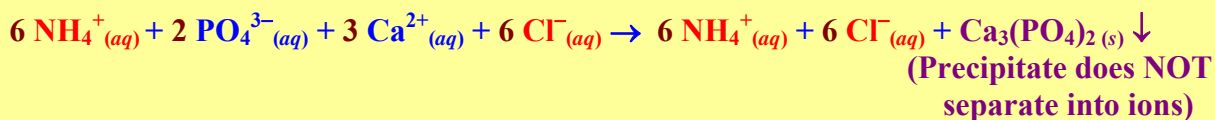
Net Ionic Equation: - an ionic equation that only shows the ions responsible in forming the precipitate. **Spectator Ions** (ions that do not form the precipitate) are omitted.

Example 1: Predict all products form when an ammonium phosphate solution reacts with a calcium chloride solution. Explain the reaction in a form of a balanced

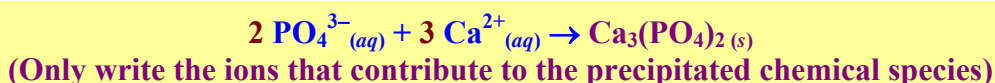
a. Molecular Equation



b. Complete Ionic Equation

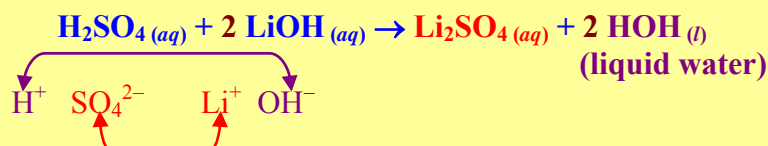


c. Net Ionic Equation

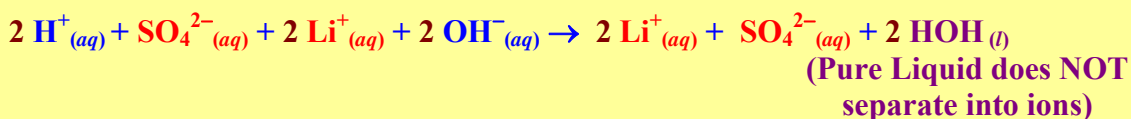


Example 2: Predict all products form when sulfuric acid reacts with a lithium hydroxide solution. Explain the reaction in a form of a balanced

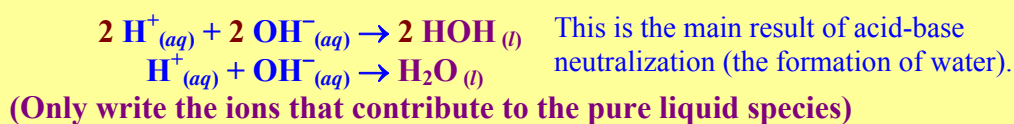
a. Molecular Equation



b. Complete Ionic Equation

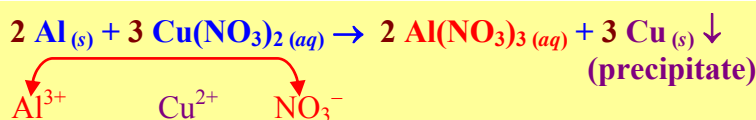


c. Net Ionic Equation

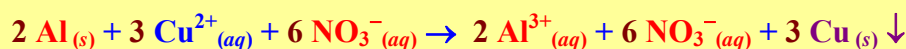


Example 3: Predict all products form when solid aluminium reacts with a copper (II) nitrate solution. Explain the reaction in a form of a balanced

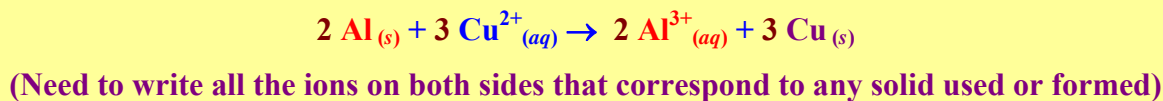
a. Molecular Equation



b. Complete Ionic Equation

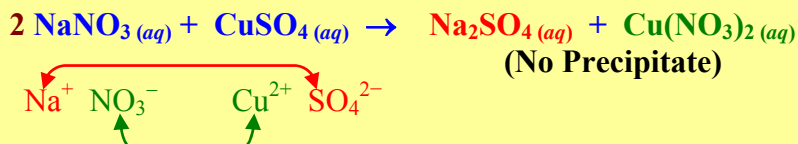


c. Net Ionic Equation

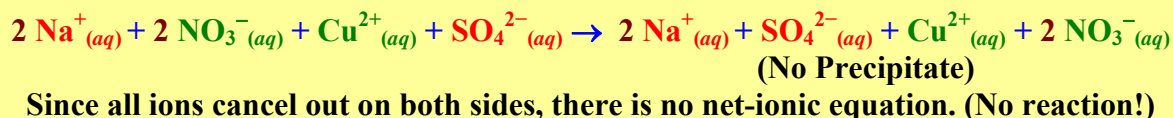


Example 4: Predict all products form when a sodium nitrate solution is mixed with a copper (II) sulfate solution. Explain the reaction in a form of a balanced

a. Molecular Equation



b. Complete Ionic Equation



Assignment**4.1 pg. 124–125 #1, 2, 3, 5, 6, 8 to 13****4.2 pg. 125 #16 to 24****4.3: Acids-Base Reactions****Physical and Chemical Properties of Acid and Base**

Acids	Bases
Taste Sour (Citric Acids).	Taste Bitter.
Burning Sensation (Stomach Acid).	Feels Slippery (Detergent, Degreaser).
Corrosive with Metals (reacts to give off $H_2(g)$).	Alkaline in Nature (NaOH, Baking Soda).
Red litmus remains Red; Blue litmus turns Red.	Red litmus turns Blue; Blue litmus remains Blue.
Bromothymol Blue turns Yellow	Bromothymol Blue turns Blue.
Phenolphthalein turns Colourless.	Phenolphthalein turns Pink.
pH < 7	pH > 7



Example 1: Five unlabeled aqueous solutions were tested and the results are shown. If these substances are $C_2H_5OH(l)$, $H_2SO_4(aq)$, $HF(aq)$, $NaNO_3(aq)$, and $Ba(OH)_2(aq)$, identify the unknown solutions.

Solution	Conductivity	Red Litmus	Blue Litmus
1	High	Red	Blue
2	Low	Red	Red
3	High	Blue	Blue
4	None	Red	Blue
5	High	Red	Red

Solution 1: Ionic (Conductivity High) and Neutral (No change in litmus) → **Neutral Ionic** → $NaNO_3(aq)$

Solution 2: Slightly Ionic (Conductivity Low) and Acidic (Both litmus are Red) → **Weak Acid** → $HF(aq)$

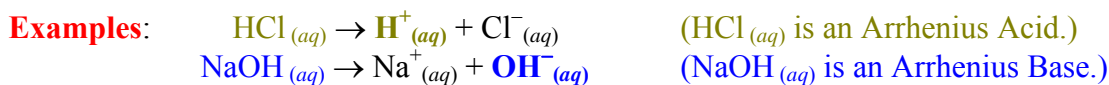
Solution 3: Ionic (Conductivity High) and Basic (Both litmus are Blue) → **Strong Base** → $Ba(OH)_2(aq)$

Solution 4: Molecular (No Conductivity) and Neutral (No change in litmus) → **Molecular** → $C_2H_5OH(l)$

Solution 5: Ionic (Conductivity High) and Acidic (Both litmus are Red) → **Strong Acid** → $H_2SO_4(aq)$

Conceptual Definition: - an explanation that attempts to describe why things are the way they are.

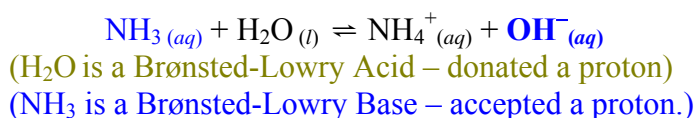
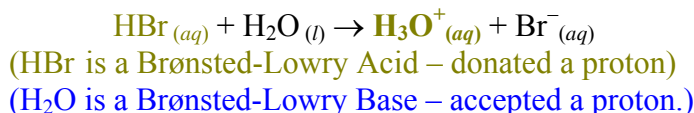
Arrhenius Concept: - **acids are H^+ (proton) producers** and **bases are OH^- producers**.



Brønsted-Lowry Model: - acids and bases **react with water** to **dissociate** where **acids are H^+ (proton) donors** and **bases are H^+ (proton) acceptors**.
- first proposed by Johannes Brønsted and Thomas Lowry.

Hydronium Ion: - an ion formed when an **acid “donated” H^+ ion combined with a H_2O molecule** to form a **H_3O^+ ion (hydronium ion)**.
- essentially has the same function as a H^+ ion, but H_3O^+ denotes that we are using the Brønsted-Lowry model.

Examples:



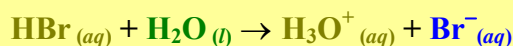
Strong Acids: - acids that **dissociate completely (100%) in water**.
- **Note: Strong Acids DO NOT MEAN that they are VERY CORROSIVE. It is the $[H_3O^+]$ that defines acidity.**

$$HA_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + A^-_{(aq)}$$

	[HA]		$[H_3O^+]$	$[A^-]$
Initial	x		0	0
Change	-x		+x	+x
Equilibrium	0		x	x

Examples: Strong Acids: $HClO_4_{(aq)}$, $HI_{(aq)}$, $HBr_{(aq)}$, $HCl_{(aq)}$, $H_2SO_4_{(aq)}$ and $HNO_3_{(aq)}$

Example 2: Write the Brønsted-Lowry dissociation reaction of $HBr_{(aq)}$.



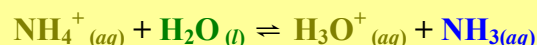
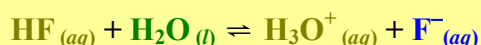
Weak Acids: - acids that **dissociate LESS than 100% in water**.
- **Note: Weak Acids DO NOT MEAN that they are NOT CORROSIVE. It is the $[H_3O^+]$ that defines acidity. At a high enough concentration, a weak acid can be corrosive.**

$$HA_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + A^-_{(aq)}$$

	[HA]		$[H_3O^+]$	$[A^-]$
Initial	x		0	0
Change	-y (where $y \ll x$)		+y	+y
Equilibrium	$(x - y) \approx x$		y	y

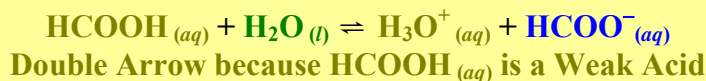
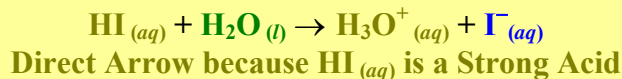
Examples: Some Weak Acids: $HOOC-COOH_{(aq)}$, $H_2SO_3_{(aq)}$, $HSO_4^-_{(aq)}$, $H_3PO_4_{(aq)}$, $HNO_2_{(aq)}$, $H_3C_6H_5O_7_{(aq)}$, $HF_{(aq)}$, $HCOOH_{(aq)}$, $C_6H_8O_6_{(aq)}$, $C_6H_5COOH_{(aq)}$, $CH_3COOH_{(aq)}$, $H_2CO_3_{(aq)}$, $H_2S_{(aq)}$, $HOCl_{(aq)}$, $HCN_{(aq)}$, $NH_4^+_{(aq)}$, and $H_3BO_3_{(aq)}$

Example 3: Write the Brønsted-Lowry dissociation reaction of $HF_{(aq)}$ and $NH_4^+_{(aq)}$.

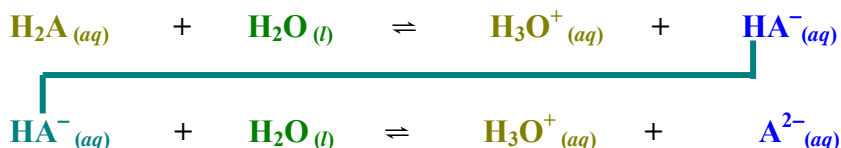


Monoprotic Acids: - acids that can donate a maximum of one proton.

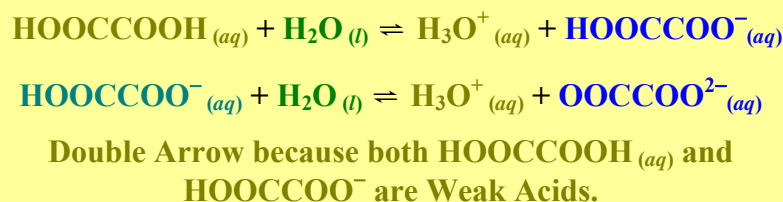
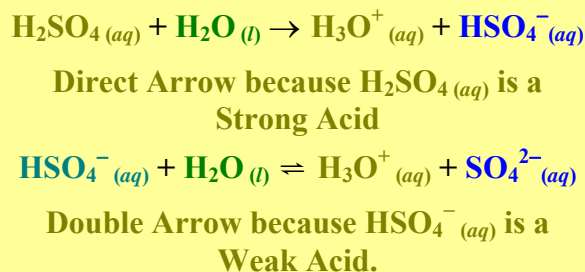
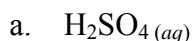
Example 4: Write the Brønsted-Lowry dissociation reaction for the following monoprotic acids.



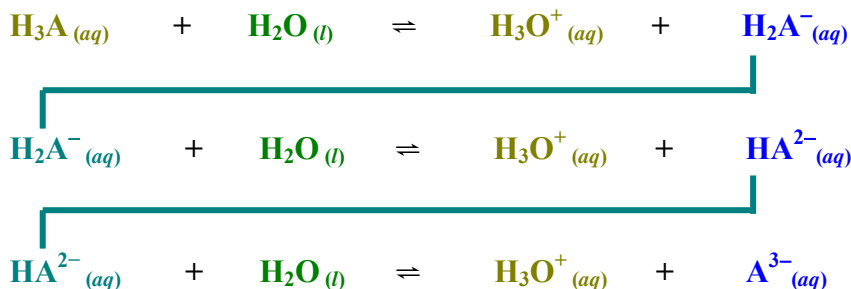
Diprotic Acids: - acids that can donate a maximum of two protons in stepwise dissociation.



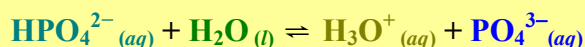
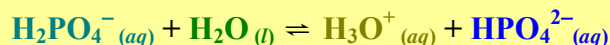
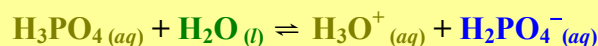
Example 5: Write the stepwise Brønsted-Lowry dissociation reaction for the following diprotic acids.



Triprotic Acids: - acids that can donate a maximum of three protons in stepwise dissociation.



Example 6: Write the stepwise Brønsted-Lowry dissociation reaction for $\text{H}_3\text{PO}_4_{(aq)}$.

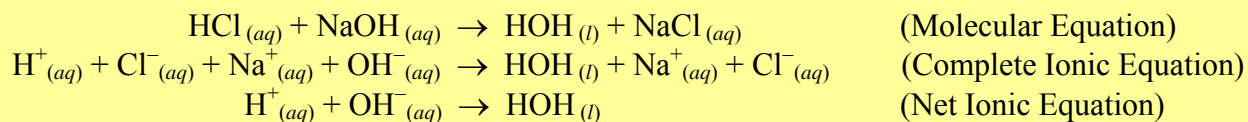


Double Arrow because $\text{H}_3\text{PO}_4_{(aq)}$, H_2PO_4^- , and $\text{HPO}_4^{2-}_{(aq)}$ are all Weak Acids.

Acid-Base Neutralization: - the reaction between acid and base to produce water and salt.

(**acid-base salt:** - cation of the base and anion of the acid)

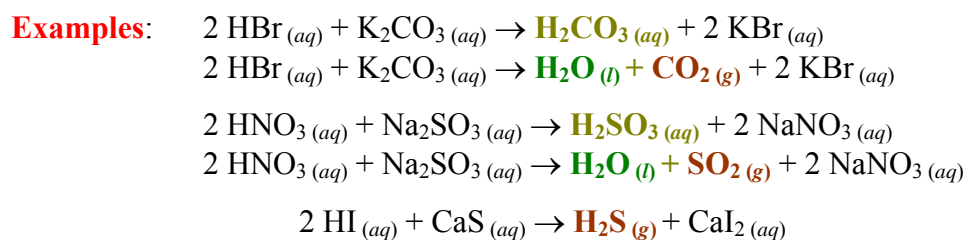
Example 7: Write the molecular, complete ionic, and net-ionic equation when $\text{NaOH}_{(aq)}$ neutralizes $\text{HCl}_{(aq)}$.



Acid-Base Neutralization that involves Gas Formation

- Some products of acid-base neutralization ($\text{H}_2\text{CO}_3_{(aq)}$ and $\text{H}_2\text{SO}_3_{(aq)}$) dissociates into gases ($\text{CO}_2_{(g)}$ and $\text{SO}_2_{(g)}$) because these gases are not very soluble in water. Hence, these products from an acid-base neutralization must be dissociated further into water and gas.

- Another product, $\text{H}_2\text{S}_{(g)}$ has a low solubility in water as well.



Assignment

4.3 pg. 126 #26 to 34

4.5: Concentration of Solutions

Molarity (M): - moles of solute per volume of solution in Litres (mol/L) or **molar (M)**.

- commonly referred to as **molar concentration (C)**.

Molar Concentration

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

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.

$$m = 3.46 \text{ g Cu(NO}_3)_2$$

$$n = \frac{3.46 \text{ g}}{187.57 \text{ g/mol}}$$

$$n = 0.0184464467 \text{ mol Cu(NO}_3)_2$$

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

$$C = ?$$

$$C = \frac{n}{V} = \frac{0.0184464467 \text{ mol}}{0.2500 \text{ L}}$$

$$\mathbf{C = 0.0738 \text{ mol/L or } 73.8 \text{ mmol/L}}$$

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

$$\begin{aligned}
 V &= 500.0 \text{ mL} = 0.5000 \text{ L} & C &= \frac{n}{V} & n &= CV = (0.0300 \text{ mol/L})(0.5000 \text{ L}) \\
 C &= 0.0300 \text{ mol/L} & & & n &= 0.0150 \text{ mol} \\
 M &= 261.98 \text{ g/mol Na}_2\text{Cr}_2\text{O}_7 & & & m &= nM = (0.0150 \text{ mol})(261.98 \text{ g/mol}) \\
 n &=? & m &=? & & m = 3.93 \text{ g}
 \end{aligned}$$

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?

$$\begin{aligned}
 m &= 4.00 \text{ mg NaHCO}_3 & C &= \frac{n}{V} & V &= \frac{n}{C} = \frac{0.0476133794 \text{ mmol}}{5.207 \text{ mmol/L}} \\
 n &= \frac{4.00 \text{ mg}}{84.01 \text{ g/mol}} & & & & V = 0.00914 \text{ L or } 9.14 \text{ mL} \\
 n &= 0.0476133794 \text{ mmol NaHCO}_3 & & & & \\
 C &= 5.207 \text{ mmol/L} & & & & \\
 V &=? & & & &
 \end{aligned}$$

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

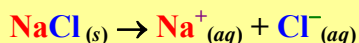


$[A^{y+}]$ = Molar Concentration of Cation A^{y+}

$[B^{x-}]$ = Molar Concentration of Anion B^{x-}

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

a. $\text{NaCl}_{(s)}$

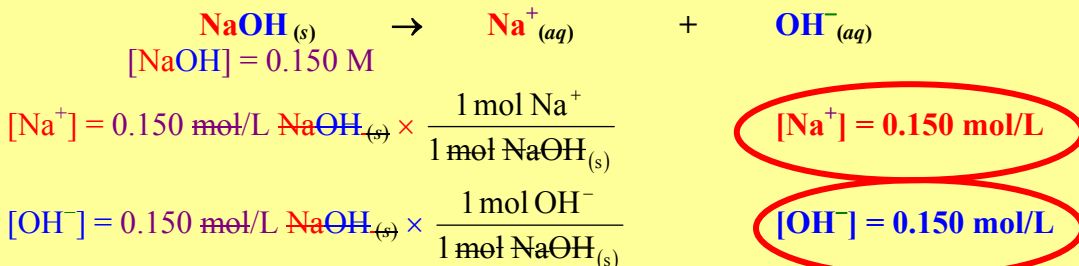


b. $\text{Cu}(\text{NO}_3)_2_{(s)}$



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

a. 0.150 M of $\text{NaOH}_{(aq)}$



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

$$\text{K}_3\text{PO}_4(s) \rightarrow 3\text{K}^+(aq) + \text{PO}_4^{3-}(aq)$$

$m = 3.45 \text{ g K}_3\text{PO}_4$
 $n = \frac{3.45 \text{ g}}{212.27 \text{ g/mol}}$
 $n = 0.0162528855 \text{ mol K}_3\text{PO}_4$
 $V = 500. \text{ mL} = 0.500 \text{ L}$
 $C = ?$

$C = \frac{n}{V} = \frac{0.0162528855 \text{ mol}}{0.500 \text{ L}}$
 $[\text{K}_3\text{PO}_4] = 0.0325 \text{ mol/L}$

$[\text{K}^+] = 0.0325 \text{ mol/L K}_3\text{PO}_4(s) \times \frac{3 \text{ mol K}^+}{1 \text{ mol K}_3\text{PO}_4(s)} = 0.0975 \text{ mol/L}$

$[\text{PO}_4^{3-}] = 0.0325 \text{ mol/L K}_3\text{PO}_4(s) \times \frac{1 \text{ mol PO}_4^{3-}}{1 \text{ mol K}_3\text{PO}_4(s)} = 0.0325 \text{ mol/L}$

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

Dilution

$$C_1V_1 = C_2V_2$$

$C_1 = \text{Concentration of Original Solution}$ $V_1 = \text{Volume of Original Solution}$
 $C_2 = \text{Concentration of Diluted Solution}$ $V_2 = \text{Total Volume of Diluted Solution}$

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)?

$C_1 = 17.4 \text{ mol/L}$
 $V_1 = ?$
 $C_2 = 1.50 \text{ mol/L}$
 $V_2 = 250. \text{ mL}$

$$C_1V_1 = C_2V_2$$

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

$V_1 = 21.6 \text{ mL}$

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

$C_1 = 0.500 \text{ mol/L}$
 $V_1 = 30.0 \text{ mL}$
 $C_2 = 0.0750 \text{ mol/L}$
 $V_2 = ?$

$$C_1V_1 = C_2V_2$$

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

$V_2 = 200. \text{ mL}$

$V_{\text{water}} = V_2 - 30.0 \text{ mL}$
 $V_{\text{water}} = 200. \text{ mL} - 30.0 \text{ mL}$
 $V_{\text{water}} = 170. \text{ mL}$

Solution Preparations from Solids (<http://media.norquest.ca/chemistry/prepstandard.wmv>)

Solution Preparations using Dilution Video (<http://www.youtube.com/watch?v=kMDC4vNEoVo>)

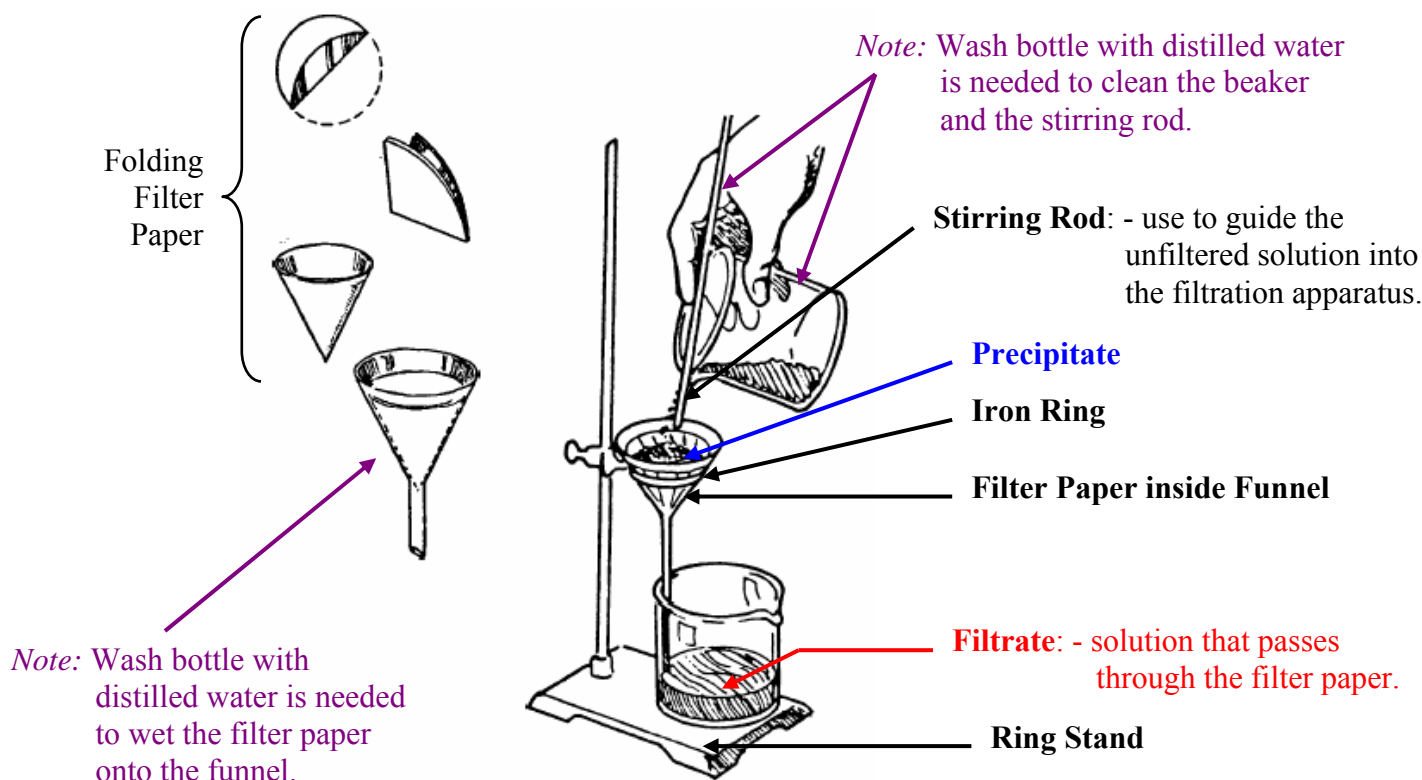
Assignment

4.5 pg. 127 #52 to 60, pg. 129 # 98, 107; pg. 127 #61, 63 to 68

4.6A: Gravimetric Analysis**Steps to Solve a Precipitation Reaction:**

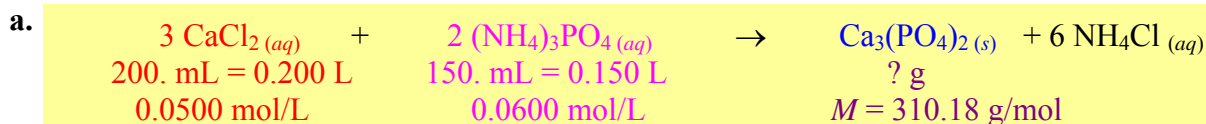
1. Write a balanced molecular equation. Identify the precipitate.
2. Put the given information underneath the proper chemicals. Identify the limiting reagent if any.
3. Using $n = CV$, convert all given information to moles.
4. Identify and use the information of the limiting reagent if necessary.
5. Determine the moles of precipitate form by using the mole ratio $\left(\frac{\text{Require Coefficient}}{\text{Given Coefficient}} \right)$.
6. Covert moles of precipitate to mass ($m = nM$).

Filtration: - a separation process to isolate the precipitate formed.

Filtration Set-up

Example 1: 200. mL of 0.0500 M of calcium chloride is reacted with 150. mL of 0.0600 M of ammonium phosphate.

- Determine the mass of the precipitate formed in this reaction.
- If the experimental mass of the precipitate is 1.28 g, calculate the % error. How can you interpret this result?
- Calculate the concentration of all spectator ions in the final solution. (Assume volumes are additive.)



Since there is enough information to determine the moles of two reactants, we need to determine which one is the limiting reagent.

$$\textcircled{1} n_{\text{CaCl}_2} = CV = (0.0500 \text{ mol/L})(0.200 \text{ L}) = 0.01 \text{ mol}$$

$$\textcircled{2} n_{(\text{NH}_4)_3\text{PO}_4} = CV = (0.0600 \text{ mol/L})(0.150 \text{ L}) = 0.009 \text{ mol}$$

Let's assume $(\text{NH}_4)_3\text{PO}_4$ is the limiting reagent. Calculate the mol CaCl_2 actually needed.

$$\textcircled{3} n_{\text{CaCl}_2} = 0.009 \text{ mol } (\text{NH}_4)_3\text{PO}_4 \times \frac{3 \text{ mol CaCl}_2}{2 \text{ mol } (\text{NH}_4)_3\text{PO}_4} = 0.0135 \text{ mol CaCl}_2 \text{ needed}$$

But we don't have 0.0135 mol of CaCl_2 , we only have 0.01 mol of CaCl_2 . Therefore, CaCl_2 is the limiting reagent. (Note: the limiting reagent is NOT always the chemical with the smaller number of moles. You have to always compare like we did above.)

Now, we calculate the moles of $\text{Ca}_3(\text{PO}_4)_2$ formed by using moles of limiting reagent CaCl_2 .

$$\textcircled{4} n_{\text{Ca}_3(\text{PO}_4)_2} = 0.01 \text{ mol CaCl}_2 \times \frac{1 \text{ mol Ca}_3(\text{PO}_4)_2}{3 \text{ mol CaCl}_2} = 0.003333... \text{ mol Ca}_3(\text{PO}_4)_2$$

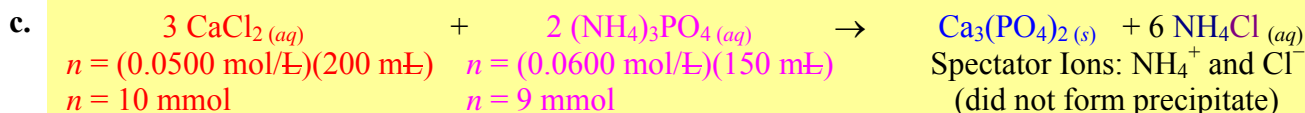
$$\textcircled{5} m_{\text{Ca}_3(\text{PO}_4)_2} = nM = (0.003333... \text{ mol Ca}_3(\text{PO}_4)_2)(310.18 \text{ g/mol})$$

$$m_{\text{Ca}_3(\text{PO}_4)_2} = 1.03 \text{ g}$$

b.

$$\% \text{ error} = \frac{|\text{Experimental} - \text{Theoretical}|}{\text{Theoretical}} \times 100\% = \frac{|1.28 \text{ g} - 1.03 \text{ g}|}{1.03 \text{ g}} \times 100\% = 24.3\%$$

This is a significant error. Since the experimental is much higher than the theoretical, we can say that there were a lot of impurities in the precipitate (from the excess ammonium phosphate).



$$[\text{Cl}^-] = \frac{2(10 \text{ mmol})}{(200. \text{ mL} + 150. \text{ mL})}$$

$$[\text{NH}_4^+] = \frac{3(9 \text{ mmol})}{(200. \text{ mL} + 150. \text{ mL})}$$

$$\text{Total Volume} \\ 200. \text{ mL} + 150. \text{ mL}$$

$$[\text{Cl}^-] = 0.0571 \text{ mol/L}$$

$$[\text{NH}_4^+] = 0.0771 \text{ mol/L}$$

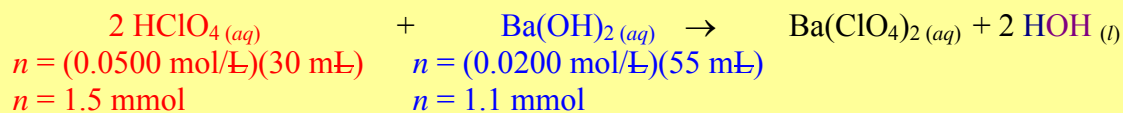
4.6B: Acid-Base Titration**Steps to Solve a Neutralization Reaction:**

1. Write a balanced molecular equation.
2. Put the given information underneath the proper chemicals.
3. Using $n = CV$, convert the given information to moles.
4. Determine the moles of the required chemical by using the mole ratio $\left(\frac{\text{Require Coefficient}}{\text{Given Coefficient}}\right)$.
5. Convert moles of the required chemical to concentration or volume $\left(C = \frac{n}{V} \text{ or } V = \frac{n}{C}\right)$.

Steps to Solve a Neutralization Reaction involving Limiting Reagent:

1. Write a balanced molecular equation.
2. Put the given information underneath the proper chemicals.
3. Convert all information to moles. Identify the limiting and excess reagent.
4. Determine the surplus number of moles of the excess reagent.
5. Use the total volume of both solutions; calculate the final concentration of the excess reagent.
6. Write the dissociation equation of the excess reagent.
7. Determine the concentration of the H^+ or OH^- ion.

Example 7: 30.0 mL of 0.0500 M of perchloric acid is mixed with 55.0 mL of 0.0200 M of barium hydroxide. Determine the final concentration of H^+ or OH^- ion present.



Let's assume Ba(OH)_2 is the limiting reagent. Calculate the mol HClO_4 actually needed.

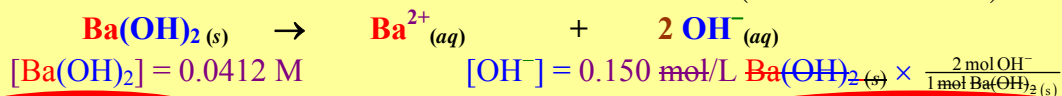
$$\textcircled{1} n_{\text{HClO}_4} = 1.1 \text{ mmol Ba(OH)}_2 \times \frac{2 \text{ mmol HClO}_4}{1 \text{ mmol Ba(OH)}_2} = 2.2 \text{ mmol HClO}_4 \text{ needed}$$

But we don't have 2.2 mmol of HClO_4 , we only have 1.5 mmol of HClO_4 . Therefore, HClO_4 is the limiting reagent. (Note: the limiting reagent is NOT always the chemical with the smaller number of moles. You have to always compare like we did above.)

$$n_{\text{Ba(OH)}_2 \text{ reacted}} = 1.5 \text{ mmol HClO}_4 \times \frac{1 \text{ mol Ba(OH)}_2}{2 \text{ mol HClO}_4} = 0.75 \text{ mmol Ba(OH)}_2$$

$$n_{\text{Ba(OH)}_2 \text{ left over}} = 1.1 \text{ mmol Ba(OH)}_2 \text{ available} - 0.75 \text{ mmol Ba(OH)}_2 \text{ reacted}$$

$$n_{\text{Ba(OH)}_2 \text{ left over}} = 0.35 \text{ mmol Ba(OH)}_2 \text{ left over} \quad [\text{Ba(OH)}_2] = \frac{0.35 \text{ mmol}}{(30.0 \text{ mL} + 55.0 \text{ mL})} = 0.00412 \text{ M}$$



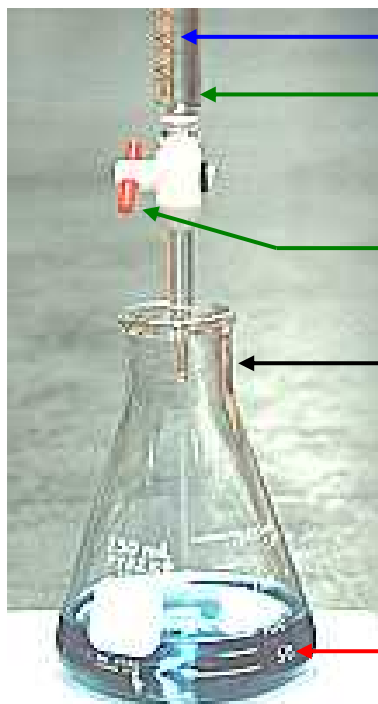
$$[\text{H}^+] = 0 \text{ mol/L}$$

$$[\text{OH}^-] = 0.00824 \text{ mol/L}$$

(H^+ is a Limiting Reagent – all used up to form H_2O)

Titration: - a volumetric analysis that involves measuring the volume of known concentration solution to measure a given volume of an unknown concentration solution.

Titration Set-up



Titrant: - the solution of known concentration.

Buret: - a precise apparatus to deliver the titrant.
- the volume of the titrant added is read by subtracting the final volume and the initial volume.

Buret Valve: - can be adjusted to let one drop out at a time.

Erlenmeyer Flask: - a container commonly uses to hold the analyte. (Narrow mouth prevents splash and spillage.)

Analyte: - the solution of an unknown concentration.
- the exact volume is usually delivered by a pipet.

Acid-Base Titration: - volumetric analysis that assist in determining the unknown concentration in an acid and base neutralization.

Equivalent Point (Stoichiometric Point): - a point where the number of moles of H^+ is equivalent to the number of moles of OH^- . ($n_{H^+} = n_{OH^-}$)

Endpoint: - a point where the indicator actually changes colour to indicate neutralization is completed.

Indicator: - a chemical that changes colour due to the pH of the solution.

Common Acid-Base Indicators:

- Bromothymol Blue – Green at pH = 7
- Phenol Red – Light Orange at pH = 7
- Phenolphthalein – Light Pink at pH = 9

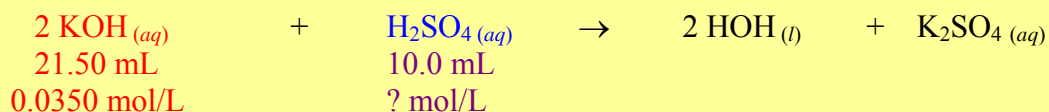
Titration Video (<http://www.youtube.com/watch?v=YDzzMcrdyB4>)

Example 8: Use the following observation table to determine the concentration of sulfuric acid.

10.0 mL of H ₂ SO ₄ (aq) titrated by 0.0350 mol/L of KOH (aq)				
	Trial 1	Trial 2	Trial 3	Trial 4
Initial Volume	0.32 mL	24.19 mL	3.48 mL	24.97 mL
Final Volume	24.19 mL	45.71 mL	24.97 mL	46.47 mL
Volume of KOH added	23.87 mL	21.52 mL	21.49 mL	21.50 mL
Bromothymol Blue Colour	Blue	Green	Green	Green

First, we have to complete the table by subtracting the final and the initial volumes. Since the titration is completed when the indicator turns green, we only average the result of the last 3 trials.

$$\text{Average Volume of KOH added} = \frac{21.52 \text{ mL} + 21.49 \text{ mL} + 21.50 \text{ mL}}{3} = 21.50 \text{ mL}$$



$$\textcircled{1} n_{\text{KOH}} = CV = (0.0350 \text{ mol/L}) (21.50 \text{ mL}) = 0.7525 \text{ mmol}$$

$$\textcircled{2} n_{\text{H}_2\text{SO}_4} = 0.7525 \text{ mmol KOH} \times \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol KOH}} = 0.37625 \text{ mmol H}_2\text{SO}_4$$

$$\textcircled{3} C_{\text{H}_2\text{SO}_4} = \frac{n}{V} = \frac{0.37625 \text{ mmol}}{10.0 \text{ mL}} = 0.037625 \text{ mol/L}$$

$$[\text{H}_2\text{SO}_4] = 0.0376 \text{ mol/L}$$

Assignment

4.6A pg. 127 #70 to 74

4.6B pg. 128 #77 to 80

4.6 pg. 128–129 #91 to 96

Chapter 5: Gases

5.1: Substance That Exist as Gases

Elements that are Gases in Room Temperature: - H_2 , N_2 , O_2 , O_3 , F_2 , Cl_2 and all Noble Gases.

Compounds that are Gases in Room Temperature:

- some non-metal oxides like NO , NO_2 , N_2O , SO_2 , CO and CO_2
- other hydrogen-non-metal compounds like HF , HCl , HBr , HI , NH_3 , H_2S , HCN , CH_4 , C_3H_8 and other light hydrocarbons

Note: Most are colourless except F_2 (pale yellow), Cl_2 (yellowish green), and NO_2 (dark brown)

Properties of Gases:

1. Gases take the shape and volume of the container. They have very low boiling points.
2. Gases can be compressed much more easily than liquids and solids.
3. Gases have much lower densities (g/L) than liquids and solids (g/mL).
4. Two or more different gases will mix readily and thoroughly when placed in the same container to form a gaseous mixture.

5.2: Pressure of a Gas

Pressure: - the amount of force per unit of area [$1 \text{ N/m}^2 = 1 \text{ Pascal (Pa)}$], measured in **kiloPascal (kPa)**, **mm Hg**, **torr**, **atmospheric pressure (atm)**.

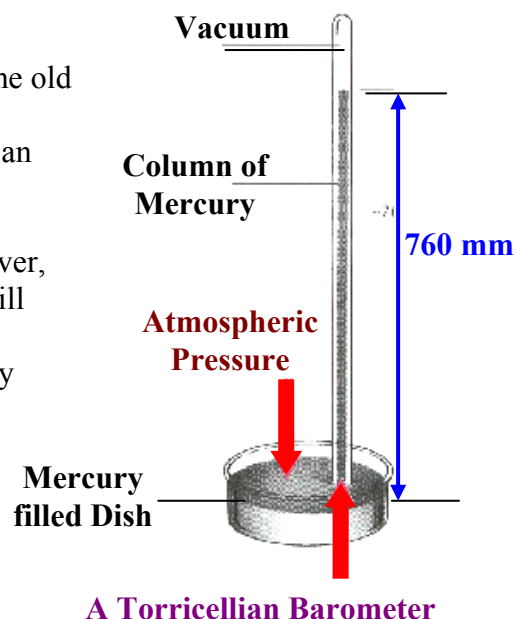
- in a pressurized container, a pressure can be felt as the particles pushed on the inside wall of the container.

Barometer: - an instrument (invented by Evangelista 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.

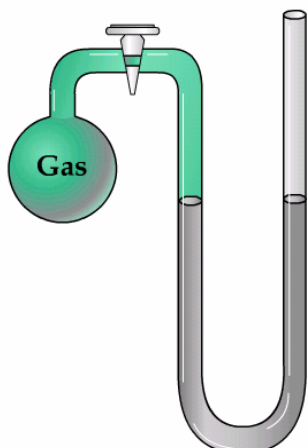
- when the glass tube is placed into the dish filled with mercury, the level of mercury in the tube will drop. However, the atmospheric pressure of the surface of mercury dish will keep the mercury column at a certain level.
- the higher the atmospheric pressure, the higher the mercury is in the column.
- at normal atmospheric pressure at sea level, the mercury column is at 760 mm.

Manometer: - a device consists of a tube connected to a container of gas to measure the pressure in the gas container.

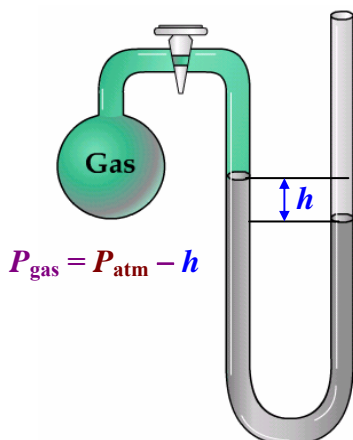
- there are two kinds of manometer: an **open-ended manometer** and a **close-ended manometer**.



1. **Open-Ended Manometer:** - the tube is open ended to allow atmospheric pressure to come in.

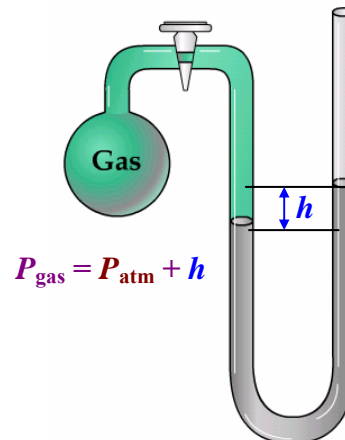


Gas Pressure =
Atmospheric Pressure



$$P_{\text{gas}} = P_{\text{atm}} - h$$

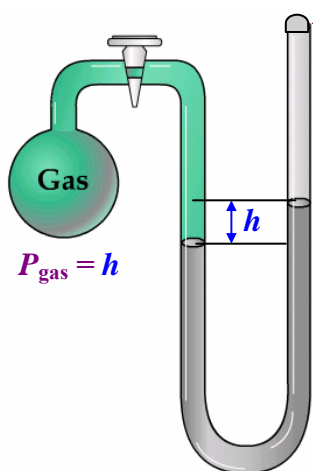
Gas Pressure <
Atmospheric Pressure



$$P_{\text{gas}} = P_{\text{atm}} + h$$

Gas Pressure >
Atmospheric Pressure

2. **Close-Ended Manometer:** - the tube is a vacuum that is closed ended. The difference in height of the U-tube is the pressure of the gas container.



$$P_{\text{gas}} = h$$

Close-Ended
Vacuum Tube

(Check out animation at

<http://www.chm.davidson.edu/ChemistryApplets/GasLaws/Pressure.html>
under Part 1: Reading a Manometer)

Units of Pressure

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

Note: 1mm Hg = 1 torr

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

$$525 \text{ mm Hg} = \mathbf{525 \text{ torr}} \quad (1 \text{ mm Hg} = 1 \text{ torr})$$

$$525 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm-Hg}} = \mathbf{0.691 \text{ atm}}$$

$$525 \text{ mm Hg} \times \frac{101.325 \text{ kPa}}{760 \text{ mm-Hg}} = \mathbf{70.0 \text{ kPa}}$$

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

$$350. \text{ kPa} \times \frac{1 \text{ atm}}{101.325 \text{ kPa}} = \mathbf{3.45 \text{ atm}}$$

$$350. \text{ kPa} \times \frac{760 \text{ torr}}{101.325 \text{ kPa}} = 2625.2 \text{ torr} = \mathbf{2.63 \times 10^3 \text{ torr}}$$

$$2.63 \times 10^3 \text{ torr} = \mathbf{2.63 \times 10^3 \text{ mm Hg}}$$

(1 mm Hg = 1 torr)

Assignment

5.1 & 5.2 pg. 163 #2, 13 and 14

5.3: The Gas Laws**Variables to Describe a Gas:**

1. **Pressure (P)**: - the amount of force per unit of area, measures in **kiloPascal (kPa)** or **standard atmosphere (atm)** or **mmHg or torr**.
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)**.

Laws that Relate Gas Variables:

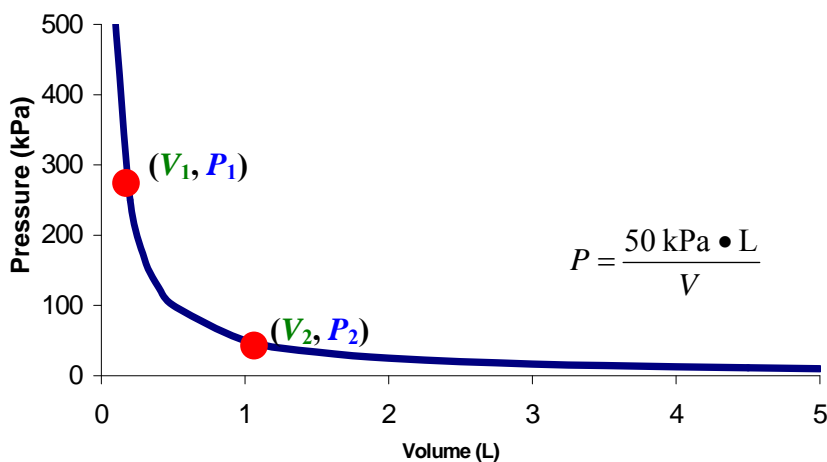
1. **Boyle's Law**: - at a constant temperature and moles, **pressure is inversely proportional to the volume**.

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

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

$k = \text{constant of proportion}$

Volume ↑ Pressure ↓

Boyle's Law (Constant Temperature)**Boyle's Law**

$$P_1V_1 = P_2V_2$$

$P_1 = \text{Pressure at Initial Condition}$ $V_1 = \text{Volume at Initial Condition}$

$P_2 = \text{Pressure at Final Condition}$ $V_2 = \text{Volume at Final Condition}$

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?

$$P_1 = 101.325 \text{ kPa}$$

$$P_2 = ?$$

$$V_1 = 2.50 \text{ L}$$

$$V_2 = 6.25 \text{ L}$$

$$P_1V_1 = P_2V_2$$

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

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

$$P_2 = 40.5 \text{ kPa}$$

As Volume ↑, Pressure ↓

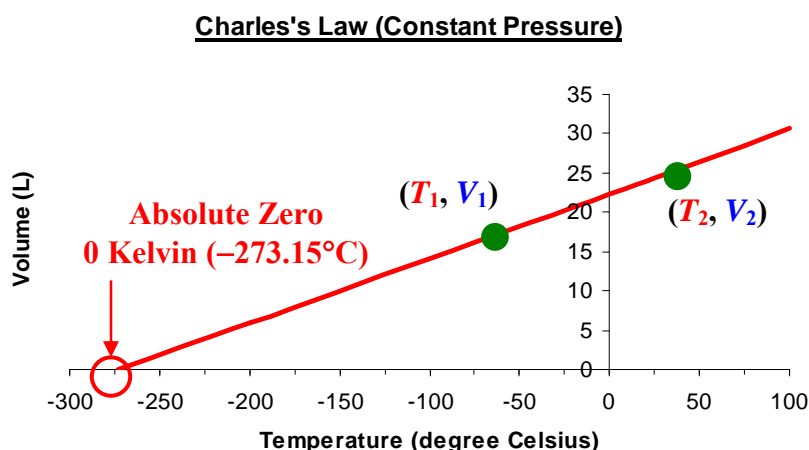
2. **Charles's Law:** - at constant pressure and moles, **volume is directly proportional to the temperature.**

$$V \propto T$$

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

$k = \text{constant of proportion}$

Temperature ↑ Volume ↑



Charles's Law

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

$T_1 = \text{Temperature at Initial Condition (in K)}$ $V_1 = \text{Volume at Initial Condition}$
 $T_2 = \text{Temperature at Final Condition (in K)}$ $V_2 = \text{Volume at Final Condition}$

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.

$V_1 = 3.25 \text{ L}$
 $T_1 = 25.0^\circ\text{C} = 298.15 \text{ K}$
 $V_2 = ?$
 $T_2 = 5.00^\circ\text{C} = 278.15 \text{ K}$
 (Change °C to K)

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

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

As Temp ↓, Volume ↓

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

$V_2 = 3.03 \text{ L}$

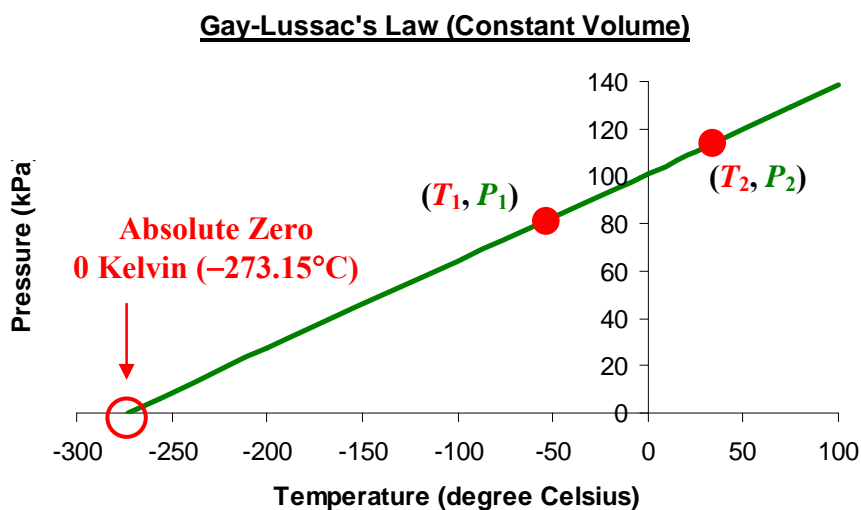
3. **Gay-Lussac's Law:** - at constant moles and volume, **pressure is directly proportional to the temperature.**

$$P \propto T$$

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

$k = \text{constant of proportion}$

Temperature ↑ Pressure ↑



Gay-Lussac's Law

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

T_1 = Temperature at Initial Condition (in K)

T_2 = Temperature at Final Condition (in K)

P_1 = Pressure at Initial Condition

P_2 = Pressure at 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!**)

$$P_1 = 8.00 \text{ atm}$$

$$T_1 = 15.0^\circ\text{C} = 288.15 \text{ K}$$

$$P_2 = ?$$

$$T_2 = 100.0^\circ\text{C} = 373.15 \text{ K}$$

(Change °C to K)

(P can be in atm)

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

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

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

As Temp ↑, Pressure ↑

$$P_2 = 10.4 \text{ atm}$$

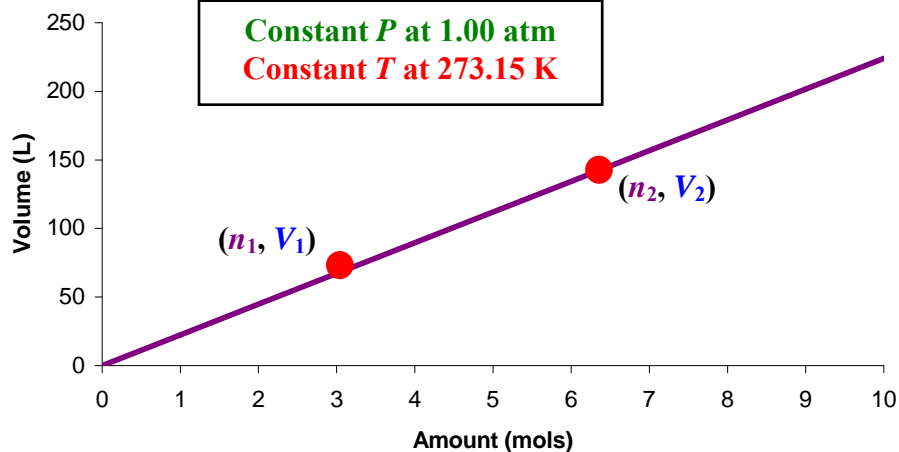
4. **Avogadro's Law:** - at constant pressure and temperature, **volume is directly proportional to amount of moles of gas present.**

$$V \propto n$$

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

k = constant of proportion

Moles ↑ Volume ↑

Avogadro' Law (Constant Temperature and Pressure)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

Example 4: An excess amount of solid phosphorus is reacted with 9.00 mol of chlorine gas at 223.2 L to produce phosphorus trichloride gas under constant temperature and pressure.

- Write a balanced equation for this reaction.
- Assuming a complete reaction, what is the volume of phosphorus trichloride produced?

a.	$\text{P}_4(\text{s}) + 6 \text{Cl}_2(\text{g}) \rightarrow 4 \text{PCl}_3(\text{g})$		
b.	$n_1 = 9.00 \text{ mol } (\text{Cl}_2)$	$\frac{V_1}{n_1} = \frac{V_2}{n_2}$	
	$V_1 = 223.2 \text{ L}$	$\frac{n_2 V_1}{n_1} = V_2$	
	$n_2 = 9.00 \text{ mol } \text{Cl}_2 \times \frac{4 \text{ mol } \text{PCl}_3}{6 \text{ mol } \text{Cl}_2}$		As Moles ↓, Volume ↓
	$n_2 = 6.00 \text{ mol } (\text{PCl}_3)$	$V_2 = \frac{(6.00 \text{ mol})(223.2 \text{ L})}{(9.00 \text{ mol})}$	$V_2 = 149 \text{ L}$
	$V_2 = ?$		

Assignment
5.3 pg. 164–165 #15 to 26

5.4: The Ideal Gas Law and Combined Gas Law

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
 $PV = nRT$

$P = \text{Pressure (kPa or atm)}$	$V = \text{Volume (L)}$
$n = \text{Amount of Gas (mol)}$	$T = \text{Temperature (K)}$

$R = \text{Gas Constant} = 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{K} \cdot \text{mol}}$ or $8.314 \frac{\text{J}}{\text{K} \cdot \text{mol}}$ or $8.314 \frac{\text{L} \cdot \text{kPa}}{\text{K} \cdot \text{mol}}$

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

$V = 200. \text{ L}$ $T = 15.0^\circ\text{C} = 288.15 \text{ K}$ $P = 32.0 \text{ atm}$ $R = 0.0821 \text{ (L} \cdot \text{atm)/(K} \cdot \text{mol)}$ $m = ?$ $n = ? \text{ (need to find } n \text{ first)}$	$PV = nRT$ $\frac{PV}{RT} = n$ $n = \frac{(32.0 \text{ atm})(200. \text{ L})}{(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{K} \cdot \text{mol}})(288.15 \text{ K})}$ $n = 270.5317195 \text{ mol}$	For propane, C_3H_8 , $M = 44.11 \text{ g/mol}$ $m = nM$ $m = (270.5317195 \text{ mol})(44.11 \text{ g/mol})$ $m = 11933.15415 \text{ g}$ $m = 1.19 \times 10^4 \text{ g} = 11.9 \text{ kg}$
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Standard Mole-Volume Relationships:

- Standard Temperature and Pressure (STP):** - the amount of any gas at 0°C and 101.325 kPa (Earth's atmospheric pressure at sea level).
- Standard Ambient Temperature and Pressure (SATP):** - the amount of any gas at 25°C and 100 kPa.

STP = 22.4 L/mol @ 0.00°C and 101.325 kPa (1 atm)

SATP = 24.8 L/mol @ 25.00°C and 100.0 kPa

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

Example 2: Determine the amount of oxygen gas in a 5.00 L container under STP and SATP.

a. STP

STP = 22.4 L/mol

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

$$n = 0.223 \text{ mol}$$

b. SATP

SATP = 24.8 L/mol

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

$$n = 0.202 \text{ mol}$$

Example 3: Determine the volume of 3.50 g of nitrogen gas under STP and SATP.

a. STP

$$n = \frac{m}{M} = \frac{3.50 \text{ g}}{28.02 \text{ g/mol}} = 0.124910778 \text{ mol}$$

STP = 22.4 L/mol

$$V = (0.124910778 \text{ mol})(22.4 \text{ L/mol})$$

$$V = 2.80 \text{ L}$$

b. SATP

$$n = \frac{m}{M} = \frac{3.50 \text{ g}}{28.02 \text{ g/mol}} = 0.124910778 \text{ mol}$$

SATP = 24.8 L/mol

$$V = (0.124910778 \text{ mol})(24.8 \text{ L/mol})$$

$$V = 3.10 \text{ L}$$

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

Combined Gas Law (since $R = \frac{PV}{nT}$ from the Ideal Gas Law)

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$$

or

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \quad (\text{for } n_1 = n_2)$$

P_1 = Pressure at Initial Condition

V_1 = Volume at Initial Condition

T_1 = Temperature at Initial Condition (in K)

n_1 = Amount of Moles at Initial Condition

P_2 = Pressure at Final Condition

V_2 = Volume at Final Condition

T_2 = Temperature at Final Condition (in K)

n_2 = Amount of Moles at Final Condition

Example 4: A high altitude weather balloon has a volume of 57.2 L at 700. mm Hg and at 25.0°C.
Determine its volume at its maximum height when the pressure is 0.365 atm and the temperature is at -45.0°C.

$n_1 = n_2$ (gas amounts did not change)

$V_1 = 57.2 \text{ L}$

$P_1 = 700. \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}}$

$P_1 = 0.9210526316 \text{ atm}$

$T_1 = 25.0^\circ\text{C} = 298.15 \text{ K}$

$V_2 = ?$

$P_2 = 0.365 \text{ atm}$

$T_2 = -45.0^\circ\text{C} = 228.15 \text{ K}$

(P can be in atm or mmHg but change °C to K)

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

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

$$V_2 = \frac{(0.9210526316 \text{ atm})(57.2 \text{ L})(228.15 \text{ K})}{(298.15 \text{ K})(0.365 \text{ atm})}$$

$$V_2 = 110. \text{ L}$$

Because $P \downarrow$ much more than $T \downarrow$, the final Volume \uparrow .

Example 5: A 0.852 mol in a 1.00 L balloon of xenon sample is heated from 23.0°C to 148.0°C. The pressure the gas has also changed from 752 mm Hg to 2854 mm Hg. What is the change in volume the sample of xenon gas is experienced if there was a leak and the amount of xenon reduced to 0.750 mol?

$$\begin{aligned}
 n_1 &= 0.852 \text{ mol} & n_2 &= 0.750 \text{ mol} \\
 T_1 &= 23.0^\circ\text{C} = 296.15 \text{ K} \\
 T_2 &= 148.0^\circ\text{C} = 421.15 \text{ K} \\
 P_1 &= 752 \text{ mm Hg} \\
 P_2 &= 2854 \text{ mm Hg} \\
 R &= 0.08206 \text{ (L} \cdot \text{atm) / (K} \cdot \text{mol)} \\
 V_1 &= 1.00 \text{ L} & V_2 &= ? \\
 \Delta V &= V_2 - V_1 = ?
 \end{aligned}$$

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2} \quad \rightarrow \quad \frac{P_1 V_1 n_2 T_2}{n_1 T_1 P_2} = V_2$$

$$V_2 = \frac{(752 \text{ mmHg})(1.00 \text{ L})(0.750 \text{ mol})(421.15 \text{ K})}{(0.852 \text{ mol})(296.15 \text{ K})(2854 \text{ mmHg})}$$

$$V_2 = 0.330 \text{ L}$$

$$\Delta V = V_2 - V_1 = 0.330 \text{ L} - 1.000 \text{ L}$$

$$\Delta V = -0.67 \text{ L or a decrease of } 0.67 \text{ L}$$

Example 6: An expandable container is filled with a gas mixture. If the temperature (in Kelvin) of the container is doubled and the pressure is decreased by one-third, how would the new volume compared to the original volume?

$T_2 = 2T_1$ (temperature doubled)
 $P_2 = \frac{2}{3} P_1$ (P is decreased by a third means new pressure is $1 - \frac{1}{3} = \frac{2}{3}$ of original pressure)

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

$$V_2 = \frac{P_1 V_1 (2T_1)}{T_1 (\frac{2}{3} P_1)} = \frac{P_1 V_1 (2T_1)}{T_1 (\frac{2}{3} P_1)} = \frac{V_1 (2)}{(\frac{2}{3})}$$

$$V_2 = 3V_1$$

The new volume will be three times the original volume.

Density and Molar Mass Determinations of a Gaseous Substance

Using the Ideal Gas Law formula, we can find the **Density (D)** or the **Molar Mass (M)** of any gas.

$$PV = nRT$$

$$PV = \left(\frac{m}{M} \right) RT \quad \text{(Substitute } \frac{m}{M} \text{ for } n)$$

Solve for $\frac{m}{V}$ for **Density ($D = \frac{m}{V}$)**

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

Density (g/L)

$$D = \frac{PM}{RT}$$

Solve for **Molar Mass (M)**

$$M = \frac{mRT}{PV}$$

$$M = \left(\frac{m}{V} \right) \frac{RT}{P}$$

$$M = \frac{DRT}{P}$$

Note: Do NOT memorize these formulas but learn the derivations!

Example 7: Calculate the molar mass of a gaseous compound containing carbon and hydrogen if its density is 0.645 g/L at 100.0 kPa and at 25.0°C.

$$\begin{aligned}
 D &= 0.645 \text{ g/L} \\
 P &= 100.0 \text{ kPa} \\
 T &= 25.0^\circ\text{C} = 298.15 \text{ K} \\
 R &= 8.314 \text{ (L} \cdot \text{kPa) / (K} \cdot \text{mol)} \\
 &\text{(We use this } R \text{ because we are} \\
 &\text{given } P \text{ in kPa)} \\
 M &= ?
 \end{aligned}
 \quad
 \begin{aligned}
 M &= \frac{DRT}{P} = \frac{(0.645 \text{ g/L})(8.314 \frac{\text{L} \cdot \text{kPa}}{\text{K} \cdot \text{mol}})(298.15 \text{ K})}{(100.0 \text{ kPa})} \\
 M &= 15.9883832 \text{ g/mol} \\
 M &= 16.0 \text{ g/mol}
 \end{aligned}$$

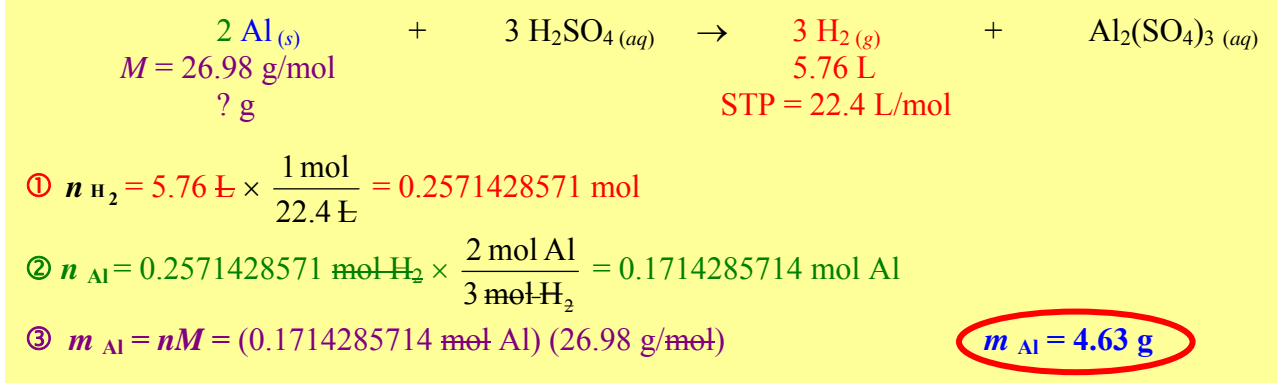
Example 8: Calculate the density in g/L of a sulfur dioxide gas at 724 torr and at -8.47°C.

$$\begin{aligned}
 P &= 724 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} \\
 P &= 0.9526315789 \text{ atm} \\
 T &= -8.47^\circ\text{C} = 264.68 \text{ K} \\
 R &= 0.0821 \text{ (L} \cdot \text{atm) / (K} \cdot \text{mol)} \\
 &\text{For sulfur dioxide, SO}_2, \\
 M &= 64.07 \text{ g/mol} \\
 D &= ?
 \end{aligned}
 \quad
 \begin{aligned}
 D &= \frac{PM}{RT} = \frac{(0.9526315789 \text{ atm})(64.07 \text{ g/mol})}{(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{K} \cdot \text{mol}})(264.68 \text{ K})} \\
 D &= 2.808765065 \text{ g/L} \\
 D &= 2.81 \text{ g/L}
 \end{aligned}$$

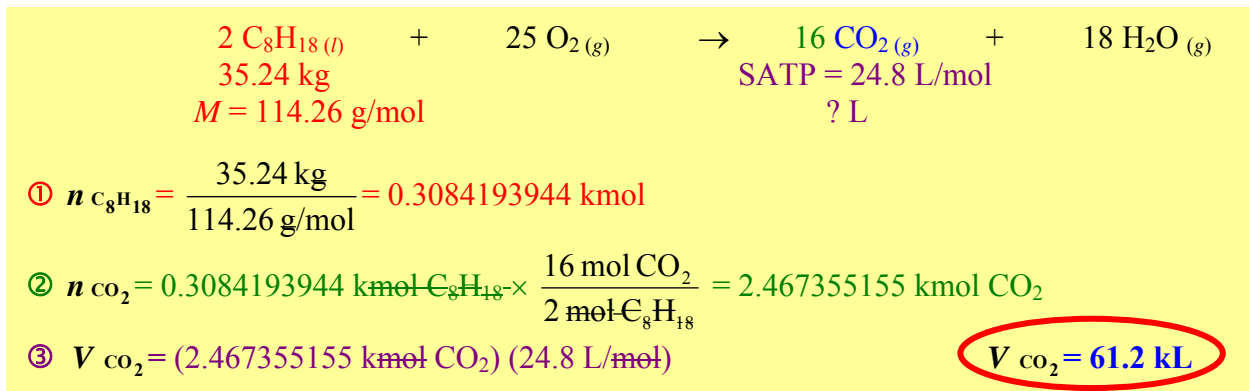
Gaseous Stoichiometry Procedure (Ideal Gas, STP, or SATP)

- Predict the products and balance the chemical equation.
- Put all the information given under the appropriate chemicals.
- Find the moles of the given chemical: $\left(n = \frac{m}{M}\right)$ or $(n = CV)$ or $\left(n = \frac{PV}{RT}\right)$ or
(at STP: $n = \text{Volume} \times \frac{1 \text{ mol}}{22.4 \text{ L}}$; at SATP: $n = \text{Volume} \times \frac{1 \text{ mol}}{24.8 \text{ L}}$).
- Check for limiting reagent if necessary (if you are given enough information to find the moles of two chemicals). Use the Limiting Reagent for further calculation. Ignore the Excess Reagent.
- Find the mole of the required chemical using mole ratio:
 $\left(\text{mol of require} = \text{mol of given} \times \frac{\text{require coefficient}}{\text{given coefficient}}\right)$
- Convert mole of the required chemical to its mass, concentration or volume equivalence: $(m = nM)$ or
 $\left(C = \frac{n}{V}\right)$ or $(PV = nRT)$ or (at STP: $V = n \times 22.4 \text{ L/mol}$; at SATP: $V = n \times 24.8 \text{ L/mol}$).

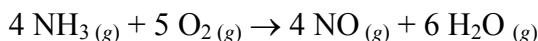
Example 9: A piece aluminium metal is placed in an excess amount of sulfuric acid until all the metal is reacted. Calculate the mass of the aluminium used if 5.76 L of hydrogen gas is evolved at STP.



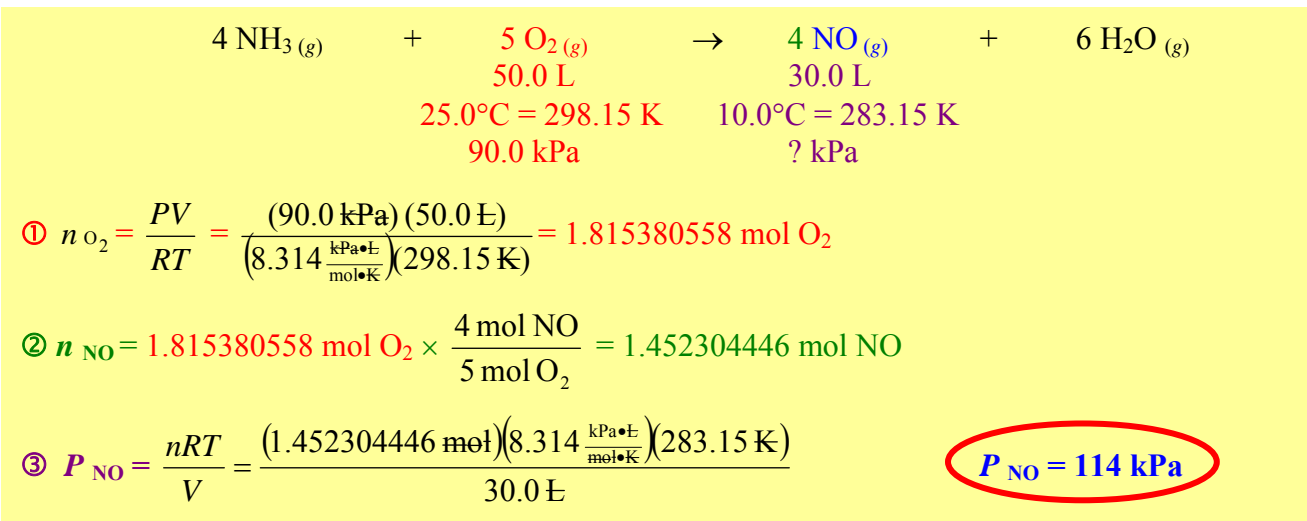
Example 10: 35.24 kg of liquid octane (C_8H_{18}) is burned under excess oxygen. Determine the volume of carbon dioxide gas produced at SATP.



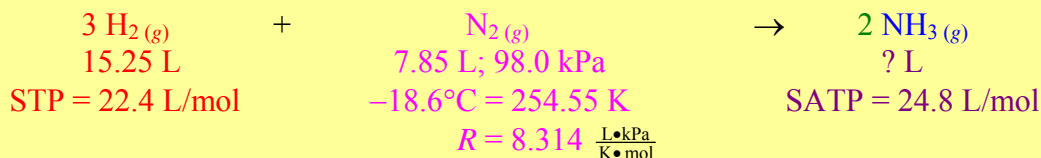
Example 11: Ammonia is reacted with oxygen to form nitrogen monoxide and water vapour.



If 50.0 L of oxygen at 90.0 kPa at 25.0°C were allowed to react with excess ammonia, what would be the pressure of nitrogen monoxide in a collector vessel measuring 30.0 L at a temperature of 10.0°C?



Example 12: If 15.25 L of hydrogen at STP is reacted with 7.85 L of nitrogen at 98.0 kPa and at -18.6°C , determine the volume of ammonia formed at SATP.



Since there is enough information to determine the moles of two reactants, we need to determine which one is the limiting reagent.

$$\textcircled{1} n_{\text{H}_2} = 15.25 \text{ L} \times \frac{1 \text{ mol}}{22.4 \text{ L}} = 0.6808035714 \text{ mol H}_2$$

$$\textcircled{2} n_{\text{N}_2} = \frac{PV}{RT} = \frac{(98.0 \text{ kPa})(7.85 \text{ L})}{(8.314 \frac{\text{kPa}\cdot\text{L}}{\text{mol}\cdot\text{K}})(254.55 \text{ K})} = 0.3635068598 \text{ mol N}_2$$

Let's assume N_2 is the limiting reagent. Calculate the mol H_2 actually needed.

$$\textcircled{3} n_{\text{H}_2} = 0.3635068598 \text{ mol N}_2 \times \frac{3 \text{ mol H}_2}{1 \text{ mol N}_2} = 1.090520579 \text{ mol H}_2 \text{ needed}$$

But we don't have 1.090520579 mol of H_2 , we only have 0.6808035714 mol of H_2 . Therefore, H_2 is the limiting reagent. (Note: the limiting reagent is NOT always the chemical with the smaller number of moles. You have to always compare like we did above.)

Now, we calculate the moles of NH_3 formed by using moles of limiting reagent, H_2 .

$$\textcircled{4} n_{\text{NH}_3} = 0.6808035714 \text{ mol H}_2 \times \frac{2 \text{ mol NH}_3}{3 \text{ mol H}_2} = 0.4538690476 \text{ mol NH}_3$$

Finally, we determine the volume of NH_3 produced.

$$\textcircled{5} V_{\text{NH}_3} = (0.4538690476 \text{ mol NH}_3)(24.8 \text{ L/mol}) = 11.25595 \dots \text{ L}$$

$$V_{\text{NH}_3} = 11.3 \text{ L}$$

Assignment

5.4 pg. 165–166 #29, 30 to 54 (do even, optional odd for extra practice); pg. 167 #89

5.5: Dalton's Law of Partial Pressures

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

$$P_{\text{Total}} = P_1 + P_2 + P_3 + \dots \quad \left(\text{Using Ideal Gas Law } P = \frac{nRT}{V} \right)$$

$$\frac{n_{\text{Total}}RT}{V} = \frac{n_1RT}{V} + \frac{n_2RT}{V} + \frac{n_3RT}{V} + \dots \quad \left(R, T, \text{ and } V \text{ are common to all components in gas mixture} \right)$$

$$n_{\text{Total}} \left(\frac{RT}{V} \right) = \left(\frac{RT}{V} \right) (n_1 + n_2 + n_3 + \dots) \quad \left(\text{Take out Common Factor } \frac{RT}{V} \text{ and cancel both sides} \right)$$

$$n_{\text{Total}} = n_1 + n_2 + n_3 + \dots \quad \left(\text{Mole Components of Gas Mixtures} \right)$$

Example 1: Write the equations for partial pressures and mole components for air, which is composed of N₂, O₂, Ar, CO₂, and other trace gases (<http://www.mistupid.com/chemistry/aircomp.htm>).

$$P_{\text{total}} = P_{\text{N}_2} + P_{\text{O}_2} + P_{\text{Ar}} + P_{\text{CO}_2} + P_{\text{trace gases}} \quad n_{\text{total}} = n_{\text{N}_2} + n_{\text{O}_2} + n_{\text{Ar}} + n_{\text{CO}_2} + n_{\text{trace gases}}$$

Example 2: Write the equations for partial pressures for air, which is composed of N₂, O₂, CO₂, 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.

$$P_{\text{total}} = P_{\text{N}_2} + P_{\text{O}_2} + P_{\text{CO}_2} + P_{\text{Ar}} + P_{\text{trace gases}} = 101.325 \text{ kPa}$$

$$P_{\text{N}_2} = (0.78084)(101.325 \text{ kPa})$$

$$P_{\text{O}_2} = (0.20947)(101.325 \text{ kPa})$$

$$P_{\text{CO}_2} = (0.00033)(101.325 \text{ kPa})$$

$$P_{\text{Ar}} = (0.00934)(101.325 \text{ kPa})$$

$$P_{\text{trace gases}} = (0.00002)(101.325 \text{ kPa})$$

$$P_{\text{N}_2} = 79.115 \text{ kPa}$$

$$P_{\text{O}_2} = 21.225 \text{ kPa}$$

$$P_{\text{CO}_2} = 0.033 \text{ kPa}$$

$$P_{\text{Ar}} = 0.946 \text{ kPa}$$

$$P_{\text{trace gases}} = 0.002 \text{ kPa}$$

Mole Fraction: - a **unit-less ratio** of the mole of any one component to the mole of the entire gas mixture.
- since pressure is directly proportional to mole ($PV = nRT$), mole fraction is also the ratio between the partial pressure of one gaseous component to the total pressure of the entire gas mixture.

$$\text{Mole Fraction} = \frac{\text{Mole of a Gas Component}}{\text{Total Moles in Gas Mixture}} = \frac{\text{Partial Pressure of a Gas Component}}{\text{Total Pressure in Gas Mixture}}$$

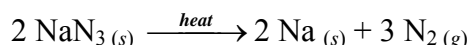
$$\chi_1 = \frac{n_1}{n_{\text{Total}}} = \frac{P_1}{P_{\text{Total}}}$$

χ = Mole Fraction (χ is pronounced *chi* – like *chi-tea*, not *tai-chi*) has **NO UNITS**.

Example 3: Air is composed of 0.20947 oxygen gas. (We would spontaneously combust if air has over 23% oxygen.) Determine the amount, in moles, and the pressure of oxygen if a compressed air tank at 4.50 atm has a total volume of 6.50 L and it is at 18.0°C.

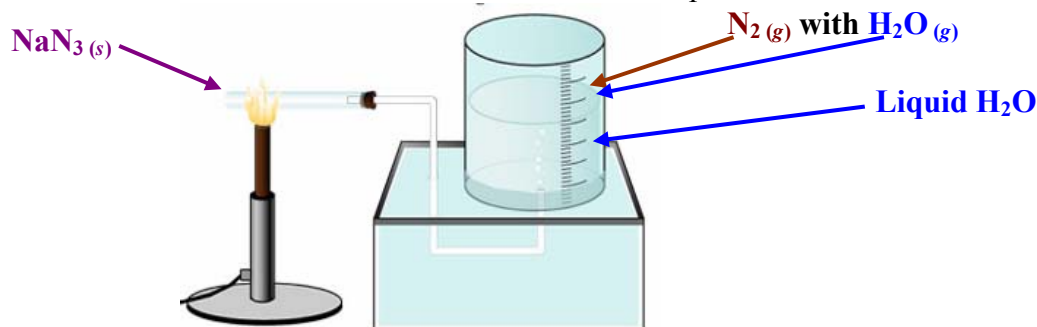
$$\begin{aligned}
 T &= 18.0^\circ\text{C} = 291.15 \text{ K} \\
 \chi &= 0.20947 \\
 P_{\text{Total}} &= 4.50 \text{ atm} \\
 V &= 6.50 \text{ L} \\
 n_{\text{oxygen}} &=? \\
 P_{\text{oxygen}} &=?
 \end{aligned}
 \quad
 \begin{aligned}
 \chi_{\text{oxygen}} &= \frac{P_{\text{O}_2}}{P_{\text{Total}}} \\
 P_{\text{oxygen}} &= \chi_{\text{oxygen}} P_{\text{Total}} = (0.20947)(4.50 \text{ atm}) = 0.942615 \text{ atm} \\
 P_{\text{oxygen}} V &= n_{\text{oxygen}} RT \quad n_{\text{O}_2} = \frac{P_{\text{O}_2} V}{RT} = \frac{(0.942615 \text{ atm})(6.50 \text{ L})}{(0.0821 \frac{\text{atm}\cdot\text{L}}{\text{mol}\cdot\text{K}})(291.15 \text{ K})}
 \end{aligned}
 \quad
 \begin{aligned}
 P_{\text{oxygen}} &= 0.943 \text{ atm} \\
 n_{\text{O}_2} &= 0.256 \text{ mol}
 \end{aligned}$$

Example 4: Sodium azide, NaN_3 , can be electronically ignited to produce nitrogen used in automobile airbags. The reaction is as followed.



An experiment is set up where the nitrogen produced is bubbled through water for collection and has become saturated with water vapour. Suppose 800 mL of nitrogen gas is collected at 20.0°C and has a total pressure of 105 kPa. (Vapour Pressure of water at 20.0°C is 17.54 torr – the amount of pressure due to evaporation of water into water vapour at a given temperature: a table can be found on pg. 152 of the textbook.)

- Determine the partial pressure of the nitrogen gas in this experiment.
- Calculate the mass of sodium azide needed to react to produce this amount of nitrogen.



$$\begin{aligned}
 P_{\text{water vapour}} &= 17.54 \text{ torr} \times \frac{101.325 \text{ kPa}}{760 \text{ torr}} \\
 P_{\text{water vapour}} &= 2.338474342 \text{ kPa} \\
 P_{\text{nitrogen}} &=? \\
 V_{\text{Total}} &= 800 \text{ mL} = 0.800 \text{ L} \\
 P_{\text{Total}} &= 105 \text{ kPa} \\
 T &= 20.0^\circ\text{C} = 293.15 \text{ K} \\
 m_{\text{NaN}_3} &=?
 \end{aligned}
 \quad
 \begin{aligned}
 \text{a. } P_{\text{Total}} &= P_{\text{water vapour}} + P_{\text{nitrogen}} \quad P_{\text{nitrogen}} = P_{\text{Total}} - P_{\text{water vapour}} \\
 P_{\text{nitrogen}} &= 105 \text{ kPa} - 2.338474342 \text{ kPa} = 102.6615257 \text{ kPa} \\
 P_{\text{nitrogen}} &= 103 \text{ kPa} \\
 \text{b. Solving for } n_{\text{N}_2} &\text{ allows us to use Stoichiometry to find } m \text{ of } \text{NaN}_3 \\
 P_{\text{nitrogen}} V &= n_{\text{nitrogen}} RT \\
 n_{\text{nitrogen}} &= \frac{P_{\text{N}_2} V}{RT} = \frac{(102.6615257 \text{ kPa})(0.800 \text{ L})}{(8.314 \frac{\text{L}\cdot\text{kPa}}{\text{K}\cdot\text{mol}})(293.15 \text{ K})} = 0.0336975078 \text{ mol} \\
 2 \text{NaN}_3 (s) &\xrightarrow{\text{heat}} 2 \text{Na} (s) + 3 \text{N}_2 (g) \\
 65.02 \text{ g/mol; ? g} &\hspace{15em} 0.0336975078 \text{ mol} \\
 \textcircled{1} n_{\text{NaN}_3} &= 0.0336975078 \text{ mol N}_2 \times \frac{2 \text{ mol NaN}_3}{3 \text{ mol N}_2} = 0.0224759646 \text{ mol NaN}_3 \\
 \textcircled{2} m_{\text{NaN}_3} &= nM = (0.0224759646 \text{ mol NaN}_3)(65.02 \text{ g/mol}) \\
 m_{\text{NaN}_3} &= 1.46 \text{ g}
 \end{aligned}$$

5.6: The Kinetic Molecular Theory of Gas

In order to explain the behaviour of an ideal gas as we have related with pressure, volume, moles and temperature, a model is needed to give us a picture of how gas particles behave in the molecular level. This model is called the **Kinetic Molecular Theory (KMT) of Gas**.

The Kinetic Molecular Theory of Gas

- All gaseous particles are so small that their volumes are essentially zero especially compared to the amount of space between them.
- All gaseous particles are constantly moving (hence the word “kinetic”). Pressure exists because of the collision of these gaseous particles against the wall of the container.
- There are no attracting or repelling forces between the particles (again due to the large distances between them).
- Temperature, express in Kelvin, is the average kinetic energy of the gas particles. (This is also the definition of temperature.)

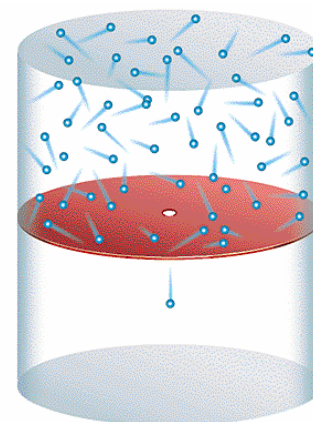
Effusion: - when gas passes through from one container to the next through a small hole.

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 separating 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 effusion rate of a gas is proportional to the square root of its molar mass.

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

$M \downarrow$ (g/mol), Effusion Rate \uparrow (mL/min)

Example 1: Calculate the ratio of effusion rates between NH_3 (g) and HCl (g) under the same temperature.

$$M_{\text{HCl}} = 36.46 \text{ g/mol}$$

$$M_{\text{NH}_3} = 17.04 \text{ g/mol}$$

$$\frac{r_{\text{NH}_3}}{r_{\text{HCl}}} = ?$$

$$\frac{r_{\text{NH}_3}}{r_{\text{HCl}}} = \sqrt{\frac{M_{\text{HCl}}}{M_{\text{NH}_3}}} = \sqrt{\frac{36.46 \text{ g/mol}}{17.04 \text{ g/mol}}}$$

$$\frac{r_{\text{NH}_3}}{r_{\text{HCl}}} = 1.463$$

NH_3 effuse about 1.5 times faster than HCl .

Example 2: 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.

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

$$M_{F_2} = 38.00 \text{ g/mol}$$

$$M_{N_2} = 28.02 \text{ g/mol}$$

$$r_{F_2} = ? \quad \text{Since } F_2 \text{ is bigger than } N_2, r_{F_2} \text{ is less than } r_{N_2} \quad r_{F_2} = 235 \text{ mL/min}$$

Example 3: Complete the table below and find the molar mass of unknown gas X.

(http://www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/flashfiles/gaslaw/effusion_macro.html)

Gas	Molar Mass (g/mol)	Volume	Time	Effusion Rate (mL/min)
H ₂		100 mL	5 sec	
O ₂		100 mL	20 sec	
Xe		100 mL	40.6 sec	
Kr		100 mL	32.4 sec	
Unknown X	?	100 mL	23.5 sec	

Gas	Molar Mass (g/mol)	Volume	Time	Effusion Rate (mL/min)
H ₂	2.02	100 mL	5.00 sec	$\frac{100 \text{ mL}}{(5.00/60) \text{ min}} = 1200$
O ₂	32.00	100 mL	20.0 sec	$\frac{100 \text{ mL}}{(20.0/60) \text{ min}} = 300.0$
Xe	131.29	100 mL	40.6 sec	$\frac{100 \text{ mL}}{(40.6/60) \text{ min}} = 147.8$
Kr	83.80	100 mL	32.4 sec	$\frac{100 \text{ mL}}{(32.4/60) \text{ min}} = 185.2$
Unknown X	?	100 mL	23.5 sec	$\frac{100 \text{ mL}}{(23.5/60) \text{ min}} = 255.3$

We can use any gas to set up a ratio with r_X . We choose O₂ because its rate is an even number.

$$r_{O_2} = 300 \text{ mL/min} \quad r_X = \frac{\sqrt{M_X}}{\sqrt{M_{O_2}}} \rightarrow \frac{(r_{O_2})\sqrt{M_{O_2}}}{(r_X)} = \sqrt{M_X} \rightarrow \left(\frac{(r_{O_2})\sqrt{M_{O_2}}}{(r_X)}\right)^2 = M_X$$

$$M_{O_2} = 32.00 \text{ g/mol}$$

$$r_X = 255.3 \text{ mL/min}$$

$$M_X = ? \quad M_X = \frac{(r_{O_2})^2 (M_{O_2})}{(r_X)^2} = \frac{(300 \text{ mL/min})^2 (32.00 \text{ g/mol})}{(255.3 \text{ mL/min})^2} \quad M_X = 44.2 \text{ g/mol}$$

Since r_X is less than r_{O_2} , X is bigger than O₂.

Diffusion: - the natural tendency of a gas from an area of high concentration travels to an area where it has a lower concentration without any barriers (we sometimes called it “*moving down the concentration gradient*”) until an equilibrium state is reached.

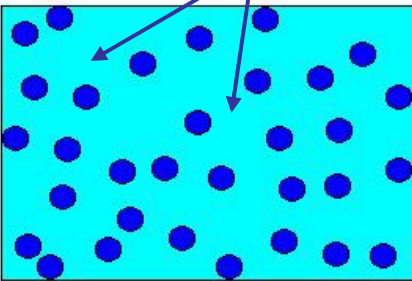
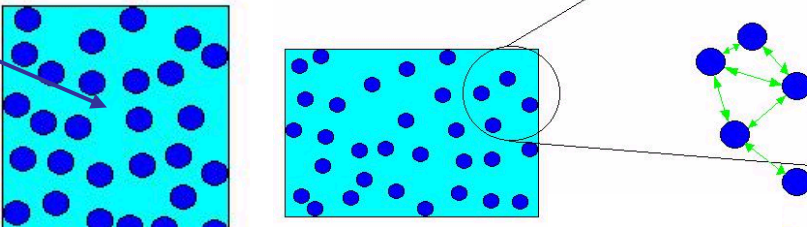
- **in general, the Rate of Diffusion is SLOWER than the Rate of Effusion.** This is because even without the barrier, there are air particles the gas has to “fight” through to reach the area of low concentration. This is also true where both areas have roughly equal pressures.
- **the only time where the Rate of Diffusion is Greater than the Rate of Effusion is when the gas is traveled FROM an area of High Pressure to an area of very Low Pressure (near vacuum)** – like a breach on the hull of a spacecraft or a plane.

(See Diffusion Animation at <http://www.biologycorner.com/resources/diffusion-animated.gif>)

5.7: Deviation from Ideal Behaviour

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 temperature is low and they are moving a lot slower.

Ideal Gas Low Pressure High Temperature	Real Gas High Pressure Low Temperature
 <p>At low pressure, there are a lot of spaces between molecules. Therefore, $V_{\text{actual}} \approx V_{\text{container}}$</p>	 <p>At high pressure, there is a lot less space between molecules and $V_{\text{actual}} < V_{\text{container}}$</p> <p>Intermolecular Forces or Electron Repulsions can significantly affect the collision rate against the wall of the container. This causes pressure to behave differently than ideal gas.</p>

Assignment

**5.5 pg. 166 #55, 56, 58, 60 to 64; pg. 167 #82 to 84, 87a, 88
5.6 & 5.7 pg. 166 #65 and Effusion Worksheet & pg. 167 #76**

Effusion Worksheet

1. Calculate the ratio of effusion rates between hydrogen gas and oxygen gas under the same temperature.
2. Calculate the ratio of effusion rates between helium gas and neon gas under the same temperature.
3. The rate of effusion of an unknown gas was measured and found to be 24.0 mL/min. Under identical conditions, the rate of effusion of propane gas was 28.83 mL/min. Determine the molar mass if the unknown gas.
4. The effusion rate of an unknown gas is measured and found to be 31.50 mL/min. Under the same conditions, the effusion rate of nitrogen gas is found to be 39.48 mL/min. If the unknown gas is a compound consists of nitrogen and oxygen, what is its chemical formula?
5. It took 4.5 minutes for 1.0 L helium to effuse through a porous barrier. How long will it take for 2.0 L of krypton gas to effuse under identical conditions?

Answers:

1. $\frac{r_{\text{H}_2}}{r_{\text{O}_2}} = 3.98$ times 2. $\frac{r_{\text{He}}}{r_{\text{Ne}}} = 2.25$ times 3. 63.7 g/mol 4. N₂O 5. 41 minutes

Chapter 12: Properties of Liquids and Phase Changes**12.1: Kinetic Molecular Theory of Liquids and Solids**

Phase: - is a homogenous part of the system in contact with other parts of the system but separated from them by a well-defined boundary.

Kinetic Molecular Theory of Liquids and Solids:

Liquid: - a phase of matter where it has a definite volume but takes up the shape of the container.

- the molecules in this phase have much less space between themselves compared to gas phase molecules. Hence, they are only slightly compressible, and resulting in a much higher density compared to densities of gases.
- this phase can still be considered as fluid. Like gas, the molecules can slide past one another freely. Therefore, they can be “pour” out of the container.

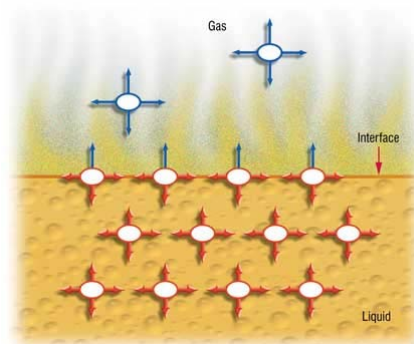
Solid: - a phase of matter where it has a definite volume and definite shape.

- the molecules in this phase have much less space between themselves compared to liquid phase molecules (with the exception of water). Hence, they are not compressible, and resulting in a very higher density.
- solid is not a fluid. Unlike gas and liquid, the molecules can only vibrate about their fixed positions. Therefore, they cannot be “pour” out of the container. (Even though one can “pour” sand or sugar out of a container, it is not a fluid because the molecules are locked in crystalline form and cannot slide past each other. They can be “pour” out because their crystal sizes are small.)

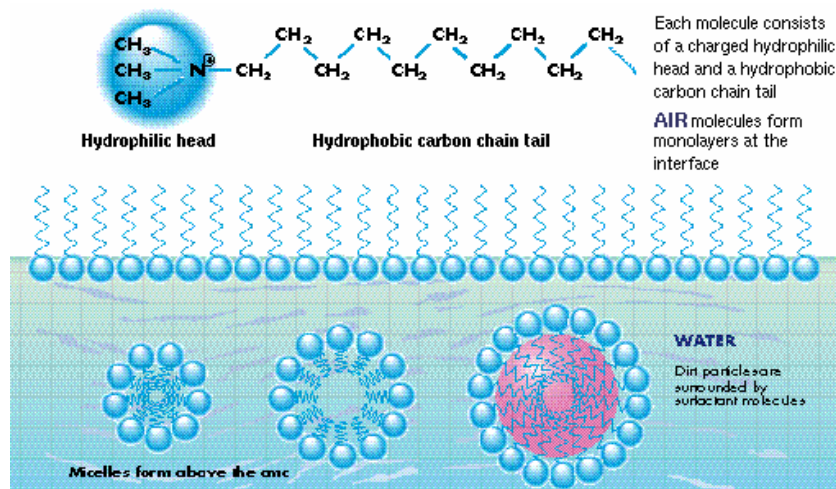
12.3: Properties of Liquids

Surface Tension: - the inward force of a liquid to minimize its surface area.

- intermolecular forces from polar molecules cannot attract with molecules in the air, which are nonpolar. Therefore, the higher net force is pushed inward, leaving a stretched surface called surface tension.



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



Hydrophobic: - non-polar nature
- “water fearing”

Hydrophilic: - polar nature
- “water loving”

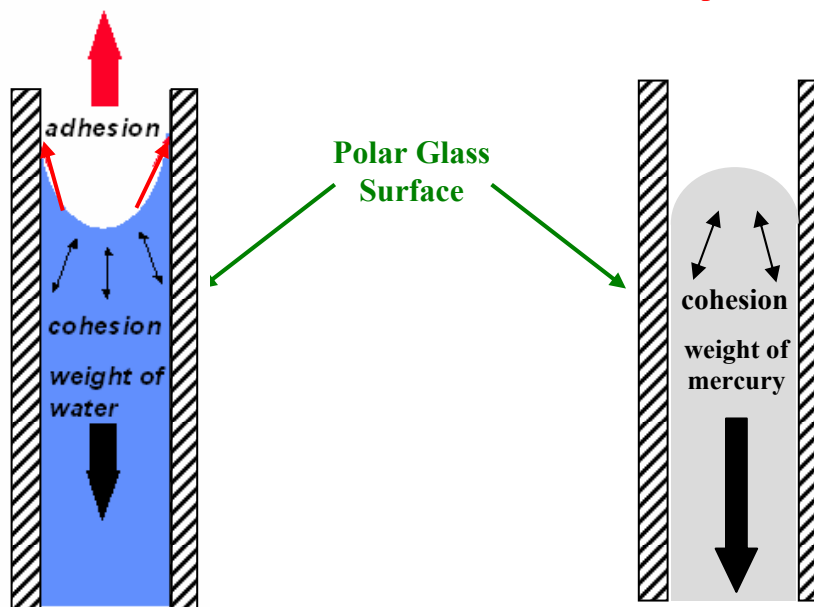
Viscosity: - the measure of a liquid resistance to flow.

- **polar liquids tend to have high viscosity** because they have both adhesive force (with the container's inner surface) and cohesive forces (with other molecules within the liquid). These attractions slows down the flow rate. Thus, viscosity increases.

Capillary Action: - when liquid suddenly rises in a narrow tube.

- there are two forces that causes capillary action to occur. They are **cohesive forces** and **adhesive forces**.

- Cohesive Forces:** - forces between molecules (intermolecular forces like London Dispersion Force, Dipole interactions, and Hydrogen Bonding).
- Adhesive Forces:** - forces between the liquid and the inner surface of the container.
 - in the case of capillary action, it would be the inside of the narrow tube.
 - **polar liquid molecules with inner surface of the container that are also polar will create a large adhesive force** (example: water and glass).
 - **non-polar liquid molecules with polar inner surface of the container will create a small adhesive force** (example: mercury and glass).



Both adhesion and cohesion forces are both strong.

However, **adhesion** > **cohesion**.

Therefore, water meniscus is **concave**.

Cohesion forces are weak (Hg is non-polar).

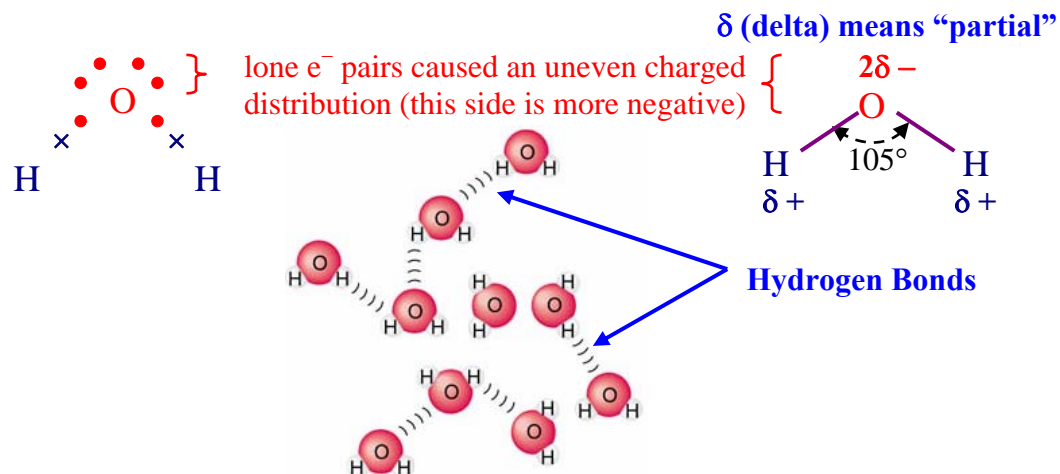
There is no adhesion.

However, **cohesion** > **adhesion**.

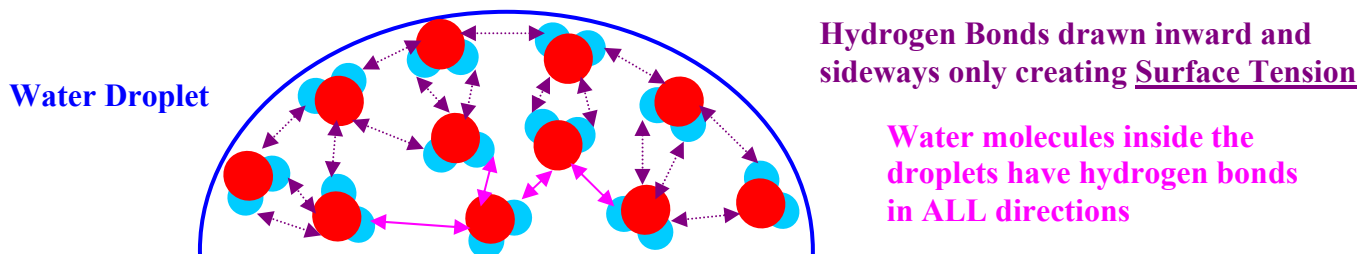
Therefore, mercury meniscus is **convex**.

Structural Features of Water:

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

**Special Properties of Water:**

- Strong Surface Tension** - 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.



(Left) Fresh dew drops on a leaf. Note the spherical drop of the water. (Centre) Surface tension of water can support a much denser paper clip. (Right) A water strider is supported by the surface tension of water.

- 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 \cdot $^\circ\text{C}$)	Copper	0.385 J/(g \cdot $^\circ\text{C}$)
Ice	2.00 J/(g \cdot $^\circ\text{C}$)	Iron	0.449 J/(g \cdot $^\circ\text{C}$)
Ammonia	2.06 J/(g \cdot $^\circ\text{C}$)	Oxygen	0.918 J/(g \cdot $^\circ\text{C}$)
Ethanol	2.44 J/(g \cdot $^\circ\text{C}$)	Nitrogen	1.040 J/(g \cdot $^\circ\text{C}$)

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).

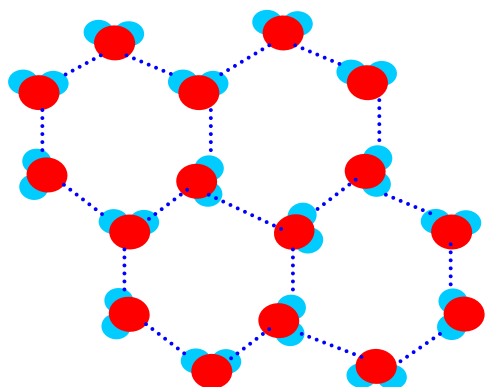
- water has a high boiling point (100°C) due to the strong hydrogen bonds. It takes more energy to break these intermolecular bonds.

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

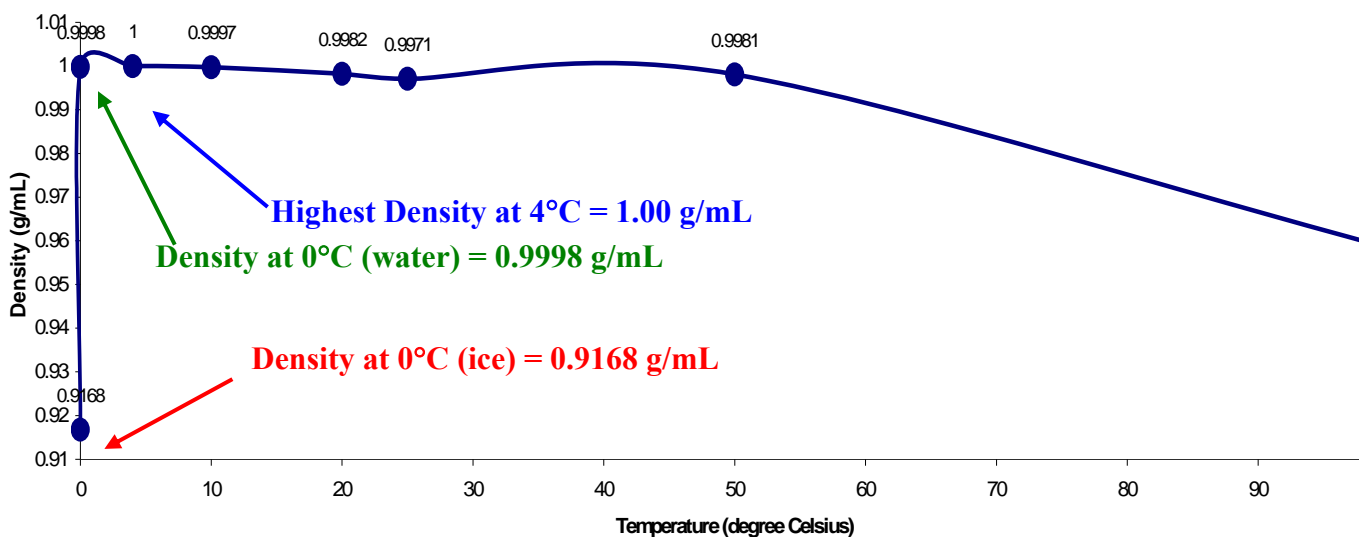
Hexagonal Arrangement of Ice Crystals: - due to the hydrogen bonds in water, it forms a honeycomb shape and expands in volume when it crystallizes into ice.



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

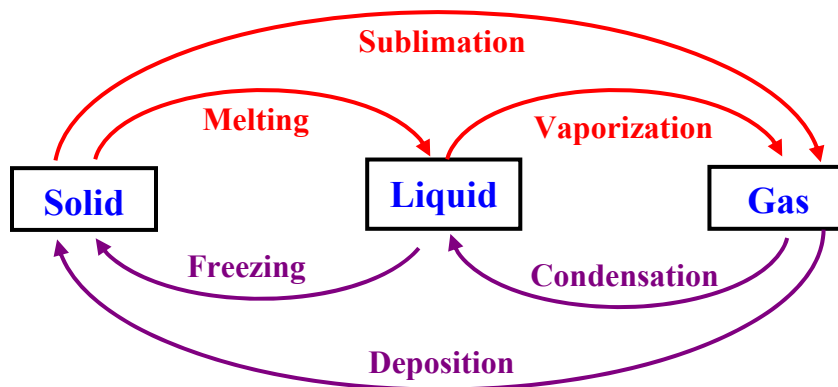
Even though no two snowflakes are alike, all of them have a basic hexagonal shape as dictated by the bent shape of water molecule and its hydrogen bonds.

Density of Water



Assignment
12.3 pg. 419 #21, 22, 24, 27 to 30

12.6 & 12.7: Phase Changes

Phase Changes

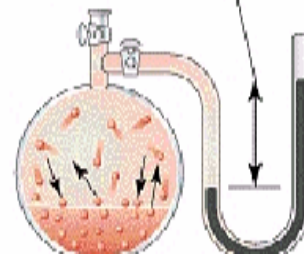
Sublimation: - when a solid becomes a gas directly without undergoing a liquid phase in between.

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

Deposition: - the phase change from gas to solid directly without undergoing a liquid phase in between.

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.

$P_{\text{gas}} = \text{Equilibrium vapor pressure}$

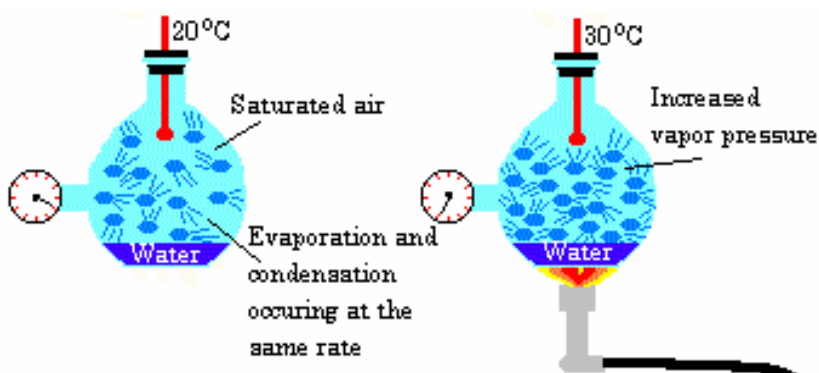
Vapour Pressure

$$P_{\text{atmosphere}} = P_{\text{vapour}} + P_{\text{Hg column}}$$

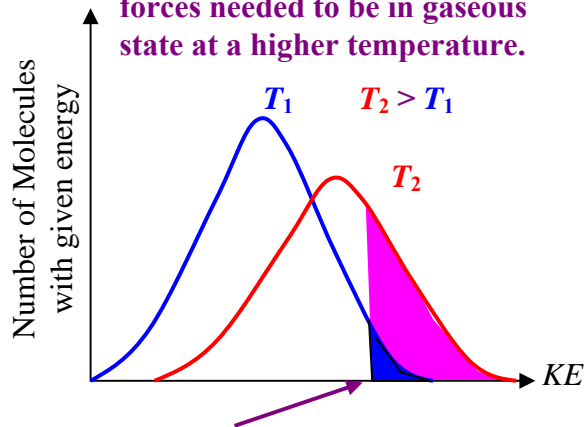
$$P_{\text{vapour}} = P_{\text{atmosphere}} - P_{\text{Hg column}}$$

- 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 liquid molecules will overcome the intermolecular forces** to become a gas. Thereby, **increasing vapour pressure.**

Temperature ↑, Vapour Pressure ↑



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

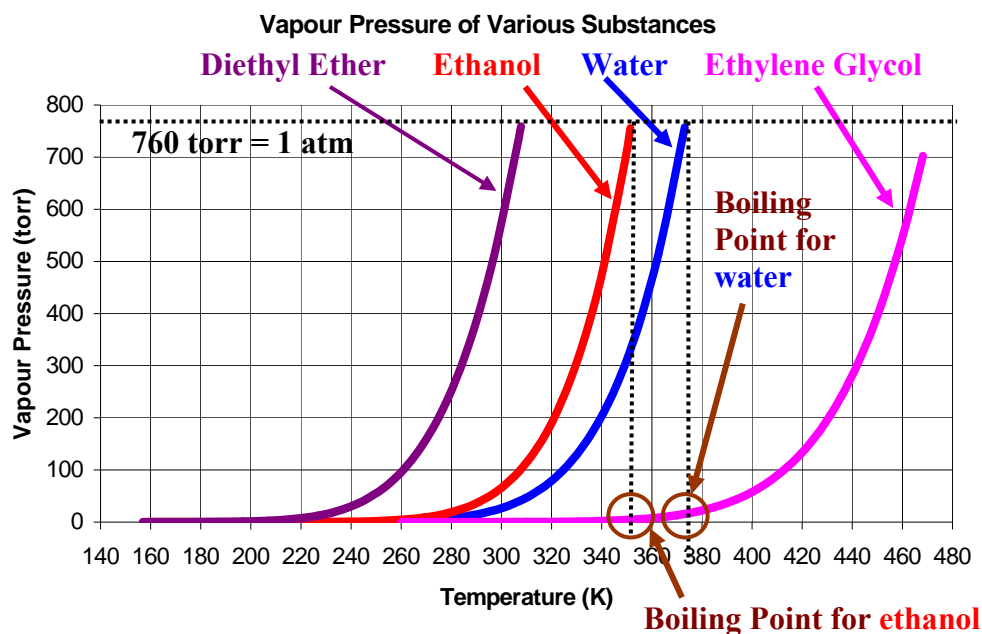


Minimum energy needed to overcome any intermolecular forces.

Volatile: - when liquids has a high vapour pressure.

- when comparing compounds at the same temperature, the compound that has a lower boiling point is considered more volatile because it will have a higher vapour pressure.
- when **intermolecular forces are weak**, the **compound is more volatile**. This is simply because it takes **less energy for the liquid to evaporate**.

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 $P_{\text{vap}} = 1 \text{ atm} = 760 \text{ torr}$, the corresponding temperature represents the **boiling point**.

Normal Melting Point: - the temperature where **vapour pressures of liquid is equaled to that of a solid** under normal atmospheric condition of 1 atm.

- when temperature is at the normal melting point, both liquid and solid coexist.
- when $T < \text{Melting Point}$, $P_{\text{vap}} \text{ Solid} < P_{\text{vap}} \text{ Liquid}$. Hence, any liquid's vapour pressure will be used by the solid to equilibrate its own vapour pressure. Thus, liquid will slowly become a solid below its melting point.
- when $T > \text{Melting Point}$, $P_{\text{vap}} \text{ Liquid} < P_{\text{vap}} \text{ Solid}$. As such, any solid's vapour pressure will be used by the liquid to equilibrate its own vapour pressure. Therefore, solid will slowly become a liquid above its melting point.

Normal Boiling Point: - the temperature where **vapour pressures of liquid is equaled to that of a gas** under normal atmospheric condition of 1 atm.

- when temperature is at the normal boiling point, both liquid and gas can coexist.
- when $T < \text{Boiling Point}$, $P_{\text{vap}} \text{ Liquid} < P_{\text{vap}} \text{ Gas}$. Hence, any vapour pressure of a gas will be used by the liquid to equilibrate its own vapour pressure. Thus, the gas will slowly become a solid below its boiling point.
- when $T > \text{Boiling Point}$, $P_{\text{vap}} \text{ Gas} < P_{\text{vap}} \text{ Liquid}$. As such, any liquid's vapour pressure will be used by the gas to equilibrate its own vapour pressure. Therefore, liquid will slowly become a gas above its boiling point.

Sublimation Point: - the temperature where **vapour pressures of solid is equaled to that of a gas.**

Normal Sublimation Point: - the temperature where **vapour pressures of solid is equaled to that of a gas** under standard atmospheric pressure of 1 atm.

Supercooled Liquid: - when a liquid is cooled too rapidly, its temperature is below the melting point.
- happens because molecules need time rearrange themselves for recrystallization to become a solid.

Superheated Liquid: - when a liquid is heated too rapidly, its temperature is above the boiling point.
- happens because molecules need time to completely break apart their intermolecular bonds to become a gas.
- the **bubbles of a superheated liquid tend to be large that burst violently** (known as **bumping**) as it's vapour pressure of the liquid is greater than that of the atmospheric pressure.

Phase Diagram: - a diagram with axes of Pressure versus Temperature to illustrate all three phases of a pure 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.

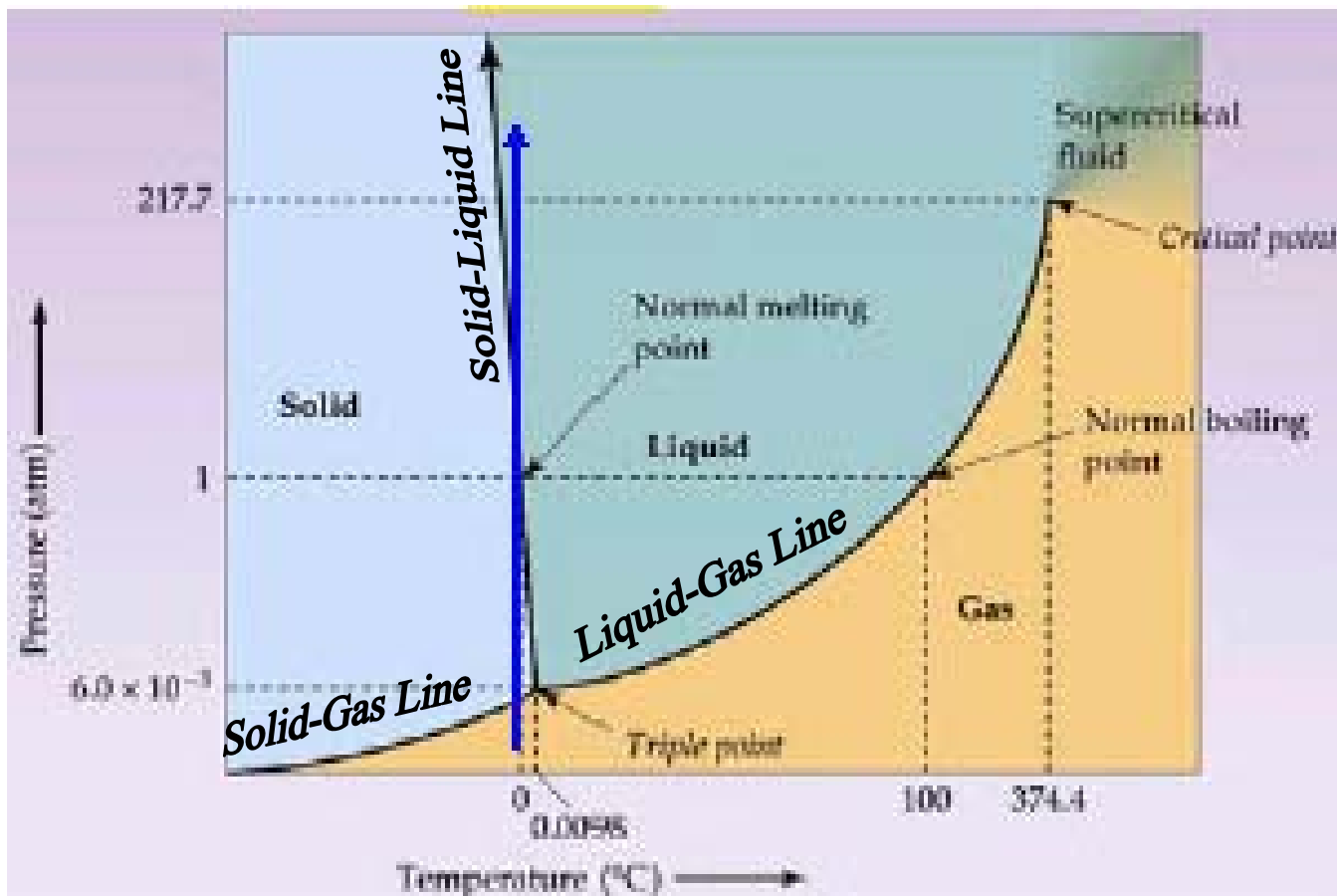
Solid-Liquid Line: - a line that outlines the temperatures and the corresponding pressures where solid and liquid phases coexist in **dynamic equilibrium** (a state where the rate of forward process is exactly balanced by the reverse process).
- used to find the **melting point** of a substance at any given pressure.

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

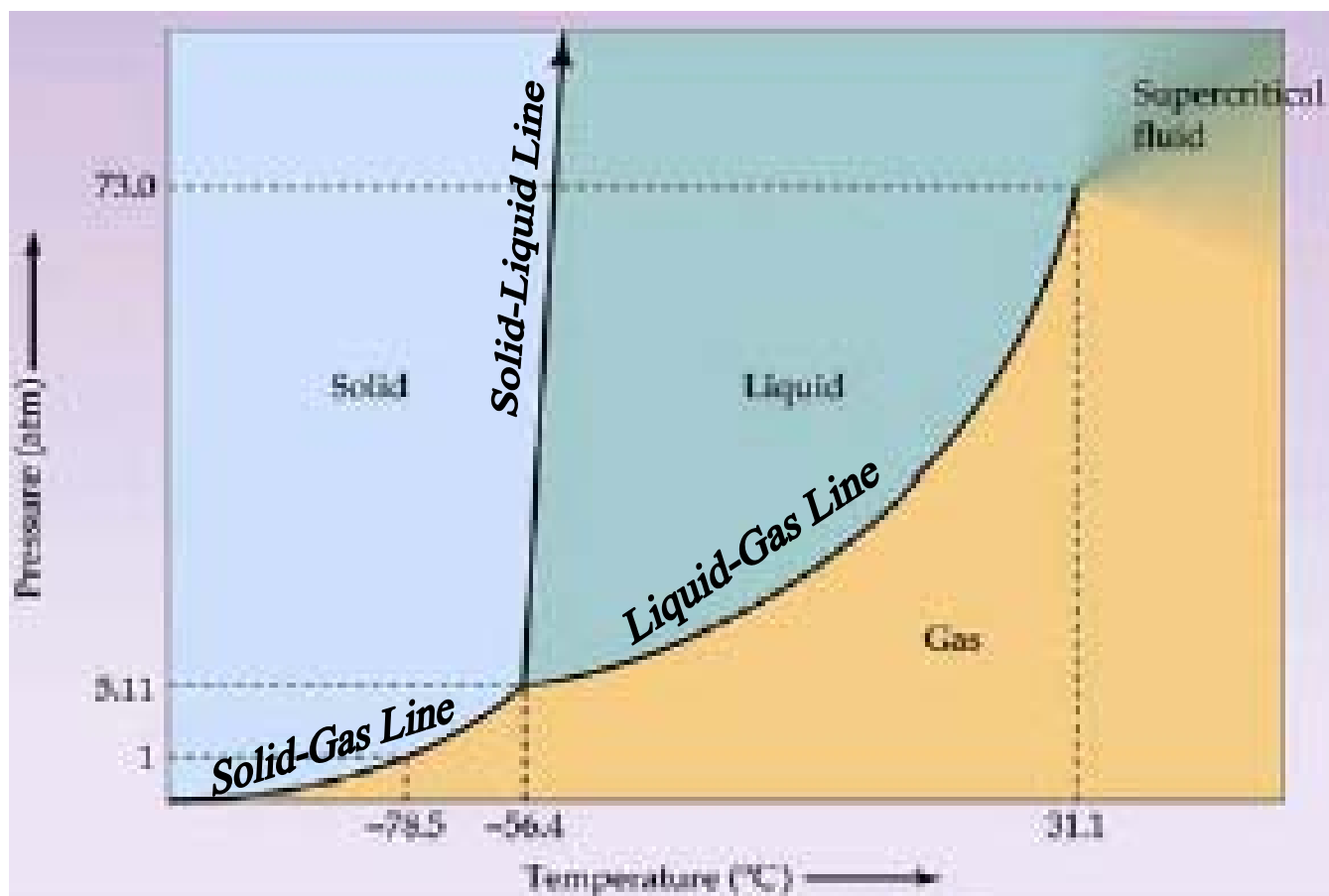
Solid-Gas Line: - a line that outlines the temperatures and the corresponding pressures where solid and gas phases coexist in dynamic equilibrium.
- used to find the **sublimation point** of a substance at any given pressure.

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 temperature conditions where all three phases can coexist in dynamic equilibrium.**

Critical Point: - a point at the end of the liquid-gas line which indicates an intermediate "fluid" region where liquid can be converted 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 without going through a change of state.

Phase Diagram of WaterNotes 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 illustrates 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 instantaneously.

Phase Diagram of Carbon DioxideNotes 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 illustrates 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

12.6 pg. 420–421 #55, 60, 67, 68, 73, 75 to 77

12.7 pg. 420 #65; pg. 421–422 #86, 87, 92, 97

Chapter 13: Physical Properties of Solutions**13.1: Types of Solutions**

Solution: - a homogeneous mixture of two or more substances.

Solute: - the substance that is being dissolved or the smaller component of the solution.

Solvent: - the substance that is doing the dissolving or the larger component of the solution.

Different Types of Solutions in Various Phases:

<u>Solute</u>	<u>Phases</u>		<u>Examples</u>
	<u>Solvent</u>	<u>Solution</u>	
Solid	Solid	Solid (Alloys)	Steel (97% Fe; 3% C), Common Brass (63% Cu; 37% Zn)
Solid	Liquid	Liquid	NaCl _(aq) , Sugar Drinks
Liquid	Liquid	Liquid	Alcoholic Cocktails (ethanol in water)
Gas	Liquid	Liquid	Carbonated Beverages (Soda Water – CO ₂ in water)
Gas	Solid	Solid	H ₂ gas in palladium
Gas	Gas	Gas	Air (78% N ₂ ; 21% O ₂ ; 1% CO ₂ , Ar and trace gases)

Crystallization: - the process in which solid solute comes out of the solution and form well-developed crystals over a long period of time.
 - happens when a supersaturated solution (see below) is disturbed either by adding a solid matter (seeding the crystal) or scratching the side of the container.

Precipitation: - the process in which dissolved solute comes out of the solution and form small particles over a relative short period of time.



(Left) Lead iodide precipitate from the reaction between lead chlorate and potassium iodide

(Right) Home-grown alum crystals

**General Descriptions of Solutions at Various Levels of Concentration:**

- Unsaturated Solution:** - when a solvent can still dissolve more solute at a particle temperature.
- Saturated Solution:** - when a solvent cannot dissolve anymore solute.
 - the rate of solvation is the same as the rate of crystallization.
- 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)

13.3: Concentration Units**Percent by Mass:** - the ratio of mass of a solute (g) and mass of a solution (g) expressed in percent.**Percent by Mass**

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

 m_{solute} = Mass of Solute (g) m_{solution} = Total Mass of Solution (g)**Percent by Solutions:** - concentration of solution expressed in % volume/volume or % mass/volume.**Percent by Solutions**

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

(Same Unit for Both Volumes)

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

Example 1: Calculate the percent by solutions of the following.

a. 1.32 g of salt in 325 mL of water

b. 45.0 mL of pure vinegar with 250. mL of water

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

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

0.406% salt (m/v)

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

$$\% \text{ (v/v)} = \frac{45.0 \text{ mL vinegar}}{(250. \text{ mL H}_2\text{O} + 45.0 \text{ mL solute})} \times 100\%$$

15.3% vinegar (v/v)**Example 2:** What is the mass of solute required to make a 3.75 L of a 22.5% (m/v) sugar solution?

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

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

← **3.75 L = 3750 mL**

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

 $m_{\text{solute}} = 844 \text{ g of sugar}$ **Molarity (Concentration):** - moles of solute per Litre of solution (M = mol/L).**Molarity (Concentration)**

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

C = Concentration (M = mol/L)**n = moles of Solute****V = Total Volume of Solution**

Molality: - moles of solute per kilogram of solvent ($m = \text{mol/kg}$).

Molality

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

Molality = m (mol/kg)

n = moles of Solute

m_{solvent} = Mass of Solvent in kg

Example 3: 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. percent by mass

a. Molarity

$$n = \frac{40.0 \text{ g}}{60.11 \text{ g/mol}} = 0.6654466811 \text{ mol C}_3\text{H}_7\text{OH}$$

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

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

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

$$[\text{C}_3\text{H}_7\text{OH}] = \frac{0.6654466811 \text{ mol}}{0.1998132005 \text{ L}} \quad [\text{C}_3\text{H}_7\text{OH}] = 3.33 \text{ M}$$

b. Molality

$$n = \frac{40.0 \text{ g}}{60.11 \text{ g/mol}} = 0.6654466811 \text{ mol C}_3\text{H}_7\text{OH}$$

$$m_{\text{solvent}} = 150 \text{ g} = 0.150 \text{ kg} \quad (1 \text{ g} \approx 1 \text{ mL H}_2\text{O})$$

$$\text{Molality} = \frac{n_{\text{solute}}}{m_{\text{solvent}}} = \frac{0.6654466811 \text{ mol}}{0.150 \text{ kg}}$$

$$\text{Molality} = 4.44 \text{ m}$$

c. Percent by Mass

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

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

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

$$\text{Percent by Mass} = \frac{m_{\text{solute}}}{m_{\text{solution}}} \times 100\% = \frac{40.0 \text{ g}}{190 \text{ g}} \times 100\%$$

$$\text{Percent by Mass} = 21.1\%$$

Example 4: Alcoholic beverages have labels indicating alcoholic content ($\text{C}_2\text{H}_5\text{OH}$) using (% v/v).

Suppose a Canadian whiskey has a 35.0% (v/v) alcohol content, and the density of pure ethanol is 0.789 g/cm³, determine the composition of the whiskey by

a. molarity

b. molality

c. percent by mass

a. Molarity (Assume 100. mL of whiskey)

$$35.0\% \text{ (v/v)} = \frac{35.0 \text{ mL C}_2\text{H}_5\text{OH}}{100 \text{ mL Total}}$$

$$m_{\text{ethanol}} = 0.789 \text{ g/cm}^3 \times 35.0 \text{ mL} = 27.615 \text{ g}$$

$$n = \frac{27.615 \text{ g}}{46.08 \text{ g/mol}} = 0.5992838542 \text{ mol C}_2\text{H}_5\text{OH}$$

$$V_{\text{Total}} = 100. \text{ mL} = 0.100 \text{ L}$$

$$[\text{C}_2\text{H}_5\text{OH}] = \frac{0.5992838542 \text{ mol}}{0.100 \text{ L}}$$

$$[\text{C}_2\text{H}_5\text{OH}] = 5.99 \text{ mol/L}$$

b. Molality (Assume 100. mL of whiskey)

$$n = 0.5992838542 \text{ mol C}_2\text{H}_5\text{OH}$$

$$V_{\text{solvent}} = V_{\text{Total}} - V_{\text{ethanol}}$$

$$V_{\text{solvent}} = 100. \text{ mL} - 35.0 \text{ mL} = 65. \text{ mL}$$

$$m_{\text{solvent}} = 65. \text{ g} = 0.065 \text{ kg} \quad (1 \text{ g} \approx 1 \text{ mL H}_2\text{O})$$

$$\text{Molality} = \frac{n_{\text{solute}}}{m_{\text{solvent}}} = \frac{0.5992838542 \text{ mol}}{0.065 \text{ kg}}$$

$$\text{Molality} = 9.22 \text{ m}$$

c. Percent by Mass

$$m_{\text{ethanol}} = 27.615 \text{ g}$$

$$m_{\text{solution}} = m_{\text{ethanol}} + m_{\text{water}}$$

$$m_{\text{solution}} = 27.615 \text{ g} + 65. \text{ g} = 92.615 \text{ g}$$

$$\text{Percent by Mass} = \frac{m_{\text{solute}}}{m_{\text{solution}}} \times 100\% = \frac{27.615 \text{ g}}{92.615 \text{ g}} \times 100\%$$

$$\text{Percent by Mass} = 29.8\%$$

Parts per Million (ppm): - the amount of solute (usually ion) in milligram per Litre of solution.
- use for measuring very small amount of solutes.

Parts per Billion (ppb): - the amount of solute (usually ion) in microgram per Litre of solution.
- use for measuring trace amount of solutes.

<u>Parts per Million</u>	<u>Parts per Billion</u>
$\text{ppm} = \frac{m_{\text{solute}} \text{ (mg)}}{V_{\text{solution}} \text{ (L)}}$	$\text{ppb} = \frac{m_{\text{solute}} \text{ (\mu g)}}{V_{\text{solution}} \text{ (L)}}$

Example 5: 4.25×10^{-2} g of sodium sulfate is dissolved in 500. mL of water. Calculate the concentration of sodium ion in ppm.

This is a Stoichiometry problem. We have to figure out the mass of Na^+ in mg.

$$\text{Na}_2\text{SO}_4(s) \rightarrow 2 \text{Na}^+(aq) + \text{SO}_4^{2-}(aq)$$

$$n_{\text{Na}_2\text{SO}_4} = \frac{4.25 \times 10^{-2} \text{ g}}{142.05 \text{ g/mol}} = 2.991904259 \times 10^{-4} \text{ mol Na}_2\text{SO}_4$$

$$n_{\text{Na}^+} = 2.9919 \times 10^{-4} \text{ mol Na}_2\text{SO}_4 \times \frac{2 \text{ mol Na}^+}{1 \text{ mol Na}_2\text{SO}_4} = 5.9838 \times 10^{-4} \text{ mol Na}^+$$

$$m_{\text{Na}^+} = 5.983808518 \times 10^{-4} \text{ mol Na}^+ \times 22.99 \text{ g/mol}$$

$$m_{\text{Na}^+} = 0.0137567758 \text{ g Na}^+ = 13.7567758 \text{ mg Na}^+$$

$$\text{ppm} = \frac{m_{\text{solute}} \text{ (mg)}}{V_{\text{solution}} \text{ (L)}}$$

$$\text{ppm} = \frac{13.7567758 \text{ mg}}{0.500 \text{ L}}$$

27.5 ppm

Assignment

13.1 pg. 448 #1, 2, 6

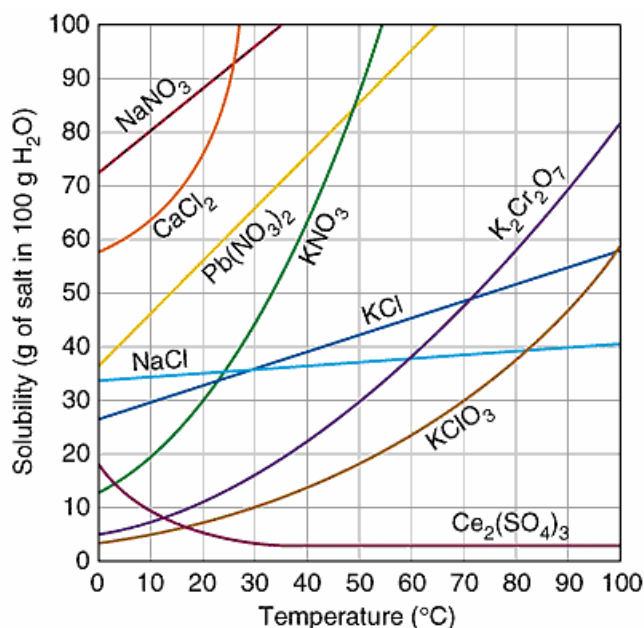
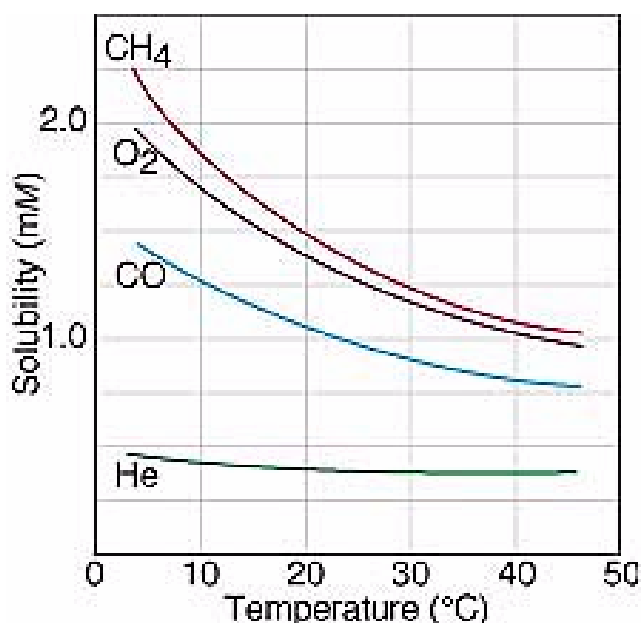
13.3 pg. 448 #13 to 22; pg. 452 #90

13.4: The Effect of Temperature on Solubility

Solubility versus Temperature:

- as **temperature increases**, solubility of **some solid solutes generally increases**. (exceptions are sodium sulfate and cerium sulfate).
- when solutes have $\Delta H_{\text{sol}} > 0$, the increase in temperature supplies the energy needed to facilitate the dissolving process. Thereby, solubility increases. (Examples: KNO_3 and NaNO_3)
- when solutes have $\Delta H_{\text{sol}} < 0$, the increase in temperature hinders the energy needed to be released during the dissolving process. Thereby, solubility decreases. (Examples: Na_2SO_4 and $\text{Ce}_2(\text{SO}_4)_3$)
- as **temperature increases**, **the solubility of gas solutes decreases**. (The increase temperature causes gas solute particles to move faster, breaking the intermolecular bonds they established with the molecules of the liquid solvent.)

Temperature ↑ Solid Solute Solubility ↑ Gas Solute Solubility ↓

Solubility of some Solid SolutesSolubility of some Gaseous Solutes

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.

General Descriptions of Solutions with Various Solubilities:

- Miscible:** - when two liquids can dissolve each other completely with infinite solubility.
(Example: Vinegar-Water Solution – Both have polar bonds.)
- Partially Miscible:** - when two liquids are slightly soluble in each other.
(Example: Anaesthetic Diethyl ether-Water Solution – weak polar versus strong polar bonds.)
- Immiscible:** - when two liquids cannot dissolve in each other, which result in a mechanical or heterogeneous mixture.
(Example: Italian Salad Dressing – Vinegar and Oil: polar versus non-polar bonds.)

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 H₂O.

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

$$\text{Solubility} = \frac{6.70 \text{ g CO}_2}{2000 \text{ g water}} = \frac{x \text{ g CO}_2}{100 \text{ g water}} \quad (\text{cross-multiply})$$

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

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

$$x = 0.335 \text{ g CO}_2$$

$$\text{Solubility} = 0.335 \text{ g CO}_2 / 100 \text{ g of water}$$

Example 2: The solubility of sucrose (cane sugar) is 230.9 g / 100 g of H₂O. Determine the amount of solvent needed to dissolve 3.00 kg of cane sugar.

$$\text{Solubility} = \frac{230.9 \text{ g sucrose}}{100 \text{ g H}_2\text{O}} \times \frac{3000 \text{ g sucrose}}{x \text{ g H}_2\text{O}} \quad \leftarrow 3.00 \text{ kg} = 3000 \text{ g} \text{ (cross-multiply)}$$

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

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

$$x = 1299.263751 \text{ g water} \quad \quad \quad x = 1.30 \text{ kg of H}_2\text{O}$$

13.5: The Effect of Pressure on Solubility of Gases

Solubility versus Pressure:

- as **pressure increases**, **solubility of gas solutes generally increases**. (More force is exerted down to force gas particles to dissolve in a denser liquid solvent).
- pressure has very little effect on solubility of liquid and solid solutes.

Pressure ↑

Gas Solute Solubility ↑

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

Henry's Law Constant (*k*): - relates concentration of a gas in solution with the pressure above the solution.
- dependant on temperature and the identity of the gas solute.

Henry's Law

$$C = kP$$

C = Concentration of Gas in Solution (mol/L)

P = Pressure above the solution (atm)

k = Henry's Law Constant [mol / (L • atm)]

Example 1: Some health food vendors claim that drinking oxygenated beverage will increase physical performance. Suppose an oxygenated drink is under 7.50 atm and the Henry's law constant for oxygen is 0.001322 mol / (L • atm) at 25°C.

- Calculate the concentration of oxygen in this beverage.
- If the partial pressure of oxygen in the atmosphere 0.20947 atm, what will be the concentration of oxygen in the beverage if it was left open for a long time.

<p>a. $k = 0.001322 \text{ mol} / (\text{L} \cdot \text{atm})$ $P = 7.50 \text{ atm}$ $C = ?$ $C = kP = (0.01322 \text{ mol} / (\text{L} \cdot \text{atm})) (7.50 \text{ atm})$ $C = 0.0992 \text{ mol/L}$</p>	<p>b. $k = 0.001322 \text{ mol} / (\text{L} \cdot \text{atm})$ $P = 0.20947 \text{ atm}$ $C = ?$ $C = kP = (0.01322 \text{ mol} / (\text{L} \cdot \text{atm})) (0.20947 \text{ atm})$ $C = 0.00277 \text{ mol/L}$</p>
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Solubility versus Structure Effect:

- in general, **polar solutes dissolve in polar solvents whereas non-polar solutes dissolve in non-polar solvents** - "**Like Dissolves Like**".
- **bigger molecules are harder to dissolve** (agitation might alleviate this problem).

Assignment

13.4 pg. 449 #24 to 26

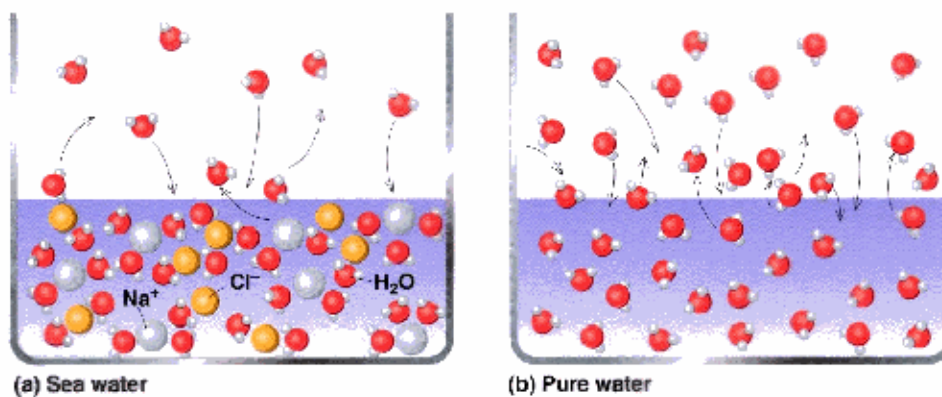
13.5 pg. 449 #27 to 36

13.6: Colligative Properties of Nonelectrolyte Solutions

Vapour Pressure of Solution: - in general, vapour pressure of solution tends to decrease as more solutes is dissolved into the solvent.



- this is because the presence of solutes molecules create more intermolecular bonds that hinder the vaporization of the solvent.



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.

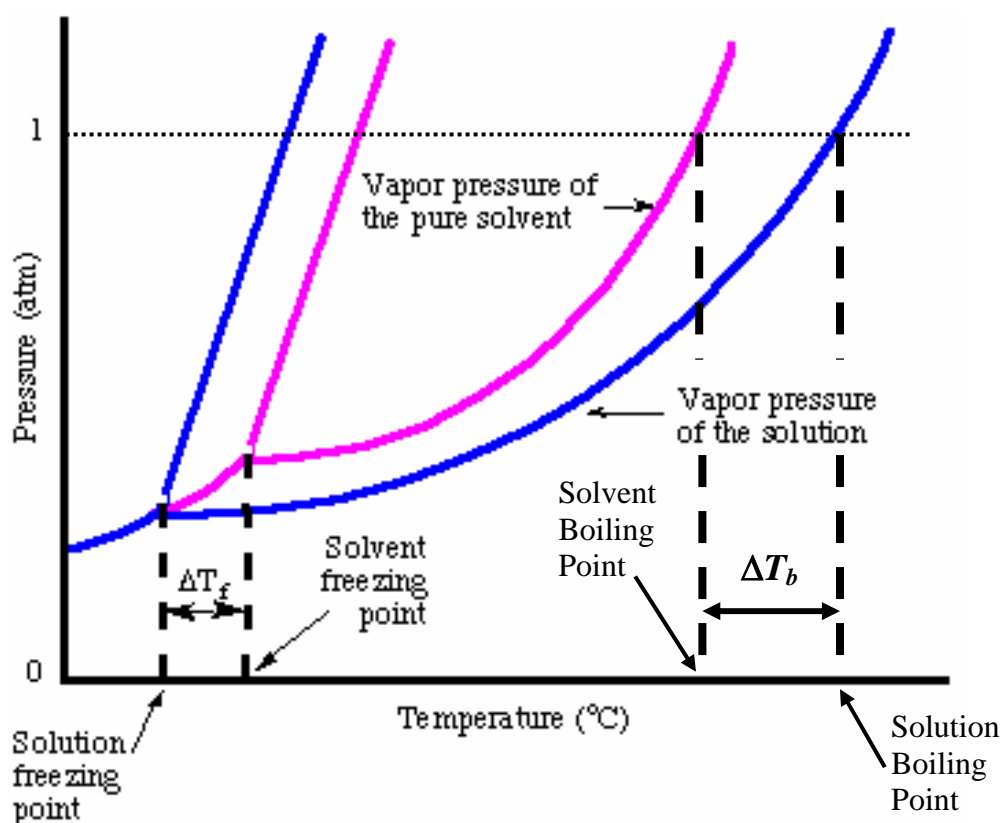
Boiling-Point Elevation: - the presence of solute in a solution raises the boiling point of a pure solvent due to the extra intermolecular forces between them (more energy is needed to boil the solution as vapour pressure is lowered).
 - the amount of temperature elevation to boil is directly proportional 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 Depression 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 ($^{\circ}\text{C}$)

ΔT_f = Change in Freezing Point Depression ($^{\circ}\text{C}$)

K_b = Molal Boiling-Point Constant ($^{\circ}\text{C} \cdot \text{kg/mol}$)

K_f = Molal Freezing-Point Constant ($^{\circ}\text{C} \cdot \text{kg/mol}$)

$\text{Molality}_{\text{solute}}$ = Molality of Solute (mol/kg of solvent)

- Example 1:** Antifreeze, ethylene glycol ($\text{CH}_2\text{OHCH}_2\text{OH}$), is commonly used to prevent water from freezing in the engine in cold temperature as well as overheating. The maximum temperature a radiator can reach is $120.^\circ\text{C}$. 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 $0.51^\circ\text{C} \cdot \text{kg/mol}$ and $1.86^\circ\text{C} \cdot \text{kg/mol}$,
- determine the volume of antifreeze needed to add to 10.0 L of water to sustain the maximum radiator temperature.
 - what is the freezing point of the solution?

a.

$$\Delta T_b = 120.^\circ\text{C} - 100^\circ\text{C} = 20.^\circ\text{C}$$

$$K_b = 0.51^\circ\text{C} \cdot \text{kg/mol}$$

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

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

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

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

$$n_{\text{antifreeze}} = ? \quad m_{\text{antifreeze}} = ?$$

$$V_{\text{antifreeze}} = ?$$

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

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

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

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

$$V_{\text{antifreeze}} = 22 \text{ L}$$

b.

$$K_f = 1.86^\circ\text{C} \cdot \text{kg/mol}$$

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

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

$$\Delta T_f = ?$$

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

$$\Delta T_f = (1.86^\circ\text{C} \cdot \text{kg/mol}) \left(\frac{392.1568627 \text{ mol}}{10.0 \text{ kg}} \right) = 72.94117646^\circ\text{C}$$

$$\text{New Freezing Point} = 0^\circ\text{C} - 72.94117646^\circ\text{C}$$

$$T_{\text{f(soln)}} = -72.9^\circ\text{C}$$

- Example 2:** 20.0 g of a newly synthesized enzyme can lower the freezing point of 100.0 g CCl_4 by 4.70°C . Given that K_f for CCl_4 is $30.0^\circ\text{C} \cdot \text{kg/mol}$, calculate the molar mass of this new enzyme.

$$\Delta T_f = 4.70^\circ\text{C}$$

$$K_f = 30.0^\circ\text{C} \cdot \text{kg/mol}$$

$$m_{\text{CCl}_4} = 100.0 \text{ g} = 0.1000 \text{ kg}$$

$$m_{\text{enzyme}} = 20.0 \text{ g}$$

$$n_{\text{enzyme}} = ?$$

$$M_{\text{enzyme}} = ?$$

$$\Delta T_f = K_f \times \text{Molality}_{\text{enzyme}} = K_f \times \frac{n_{\text{enzyme}}}{\text{kg of CCl}_4}$$

$$n_{\text{enzyme}} = \frac{\Delta T_f (\text{kg of CCl}_4)}{K_f} = \frac{(4.70^\circ\text{C})(0.1000 \text{ kg})}{(30.0^\circ\text{C} \cdot \text{kg/mol})} = 0.0156666667 \text{ mol}$$

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

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

Colligative Properties of Electrolytic Solutions

Because ionic solutes form electrolytes when they dissociate in solutions, we must account for the total moles of solute ions form.

van't Hoff Factor (i): - is a factor that relates the total number of moles of 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

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

Note: Moles of Ionic Solutes = Total Number of Ions Dissociated

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

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

Example 3: Determine the freezing point of a saturated salt solution (35.7 g / 100 g H₂O) if the K_f for pure water is 1.86 °C • kg/mol.

$$K_f = 1.86^\circ\text{C} \cdot \text{kg/mol}$$

$$m_{\text{water}} = 100 \text{ g} = 0.100 \text{ kg}$$

$$n_{\text{NaCl}} = \frac{35.7 \text{ g}}{58.44 \text{ g/mol}} = 0.610882 \text{ mol}$$

$$i = 2 \text{ (1 mol Na}^+ \text{ and 1 mol Cl}^- \text{)}$$

$$\Delta T_f = ?$$

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

$$\Delta T_f = (2)(1.86^\circ\text{C} \cdot \text{kg/mol}) \left(\frac{0.6108829569 \text{ mol}}{0.100 \text{ kg}} \right) = 22.724846^\circ\text{C}$$

$$\text{New Freezing Point} = 0^\circ\text{C} - 22.724846^\circ\text{C} \quad T_{f(\text{soln})} = -22.7^\circ\text{C}$$

Assignment

**13.6 pg. 449–452 #38, 41, 55, 56, 58 to 60,
62, 67, 70, 71, 73 to 75, 77, 96**