UNIT 7: GASES AND SOLUTION

Chapter 12: Gases

12.1: Characteristics of 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 are fluids (they can be poured or flowed because they have little or no intermolecular forces).**
3. **Gases can be compressed much more easily than liquids and solids.**
4. **Gases have much lower densities (g/L) than liquids and solids (g/mL).**
5. **Two or more different gases will mix readily and thoroughly when placed in the same container to form a gaseous mixture.**

Pressure: - the amount of force per unit of area [1 $N/m^2 = 1$ Pascal (Pa)], measured in **kiloPascal (kPa), mm Hg**, **torr**, **atmospheric pressure (atm)**.
- in a pressurized container, a pressure can be felt as the particles pushed on the inside wall of the container.

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.

Vacuum

Column of Mercury

760 mm

Atmospheric Pressure

Mercury filled Dish
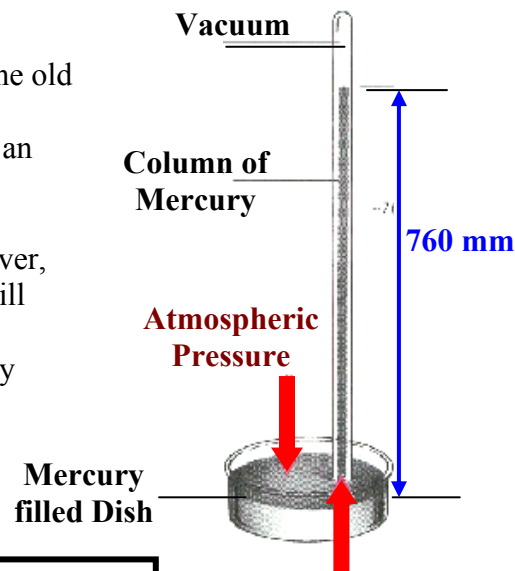
Units of Pressure

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 mm Hg = **525 torr** (1 mm Hg = 1 torr)

$525 \text{ mm Hg} \times \dfrac{1 \text{ atm}}{760 \text{ mm Hg}} =$ **0.691 atm**

$525 \text{ mm Hg} \times \dfrac{101.325 \text{ kPa}}{760 \text{ mm Hg}} =$ **70.0 kPa**

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

$350. \text{ kPa} \times \dfrac{1 \text{ atm}}{101.325 \text{ kPa}} =$ **3.45 atm**

$350. \text{ kPa} \times \dfrac{760 \text{ torr}}{101.325 \text{ kPa}} = 2625.2 \text{ torr} =$ **2.63×10^3 torr**

$2.63 \times 10^3 \text{ torr} =$ **2.63×10^3 mm Hg**
(1 mm Hg = 1 torr)

The Kinetic Molecular Theory of Gas

1. **All gaseous particles are so small that their volumes are essentially zero** especially compared to the amount of space between them.

2. **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.

3. There are no attracting or repelling forces between the particles (again due to the large distances between them).

4. **Temperature, express in Kelvin, is the average kinetic energy of the gas particles. (*This is also the definition of temperature*.)**

> ### Assignment
> **12.1 pg. 421 #1 to 3 (Practice); pg. 422 #1 to 12**

12.2: 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)**.
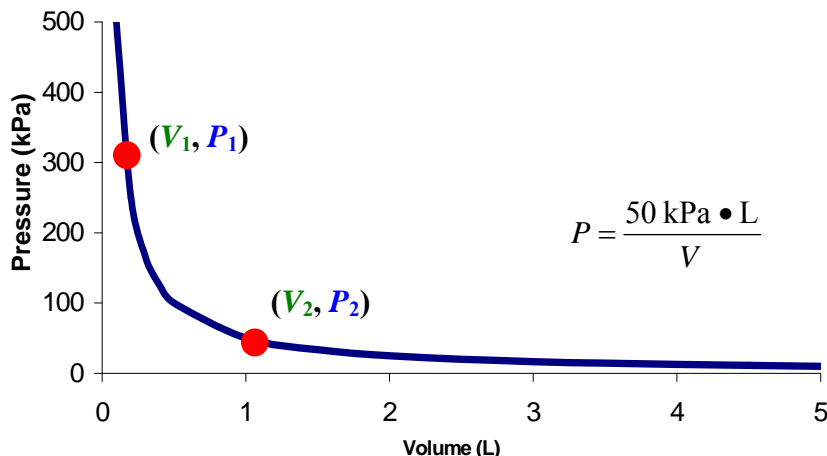
Laws that Relate Gas Variables:

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

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

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

k = constant of proportion

Volume ↑   Pressure ↓

Boyle's Law (Constant Temperature)

(V_1, P_1)

(V_2, P_2)

$$P = \frac{50 \text{ kPa} \bullet \text{L}}{V}$$

Boyle's Law

$$P_1V_1 = P_2V_2$$

P_1 = Pressure at Initial Condition V_1 = Volume at Initial Condition
P_2 = Pressure at Final Condition V_2 = 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?

$P_1 = 101.325 \text{ kPa}$

$P_2 = ?$

$V_1 = 2.50 \text{ L}$

$V_2 = 6.25 \text{ L}$

$$P_1V_1 = P_2V_2$$

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

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

As Volume ↑, Pressure ↓
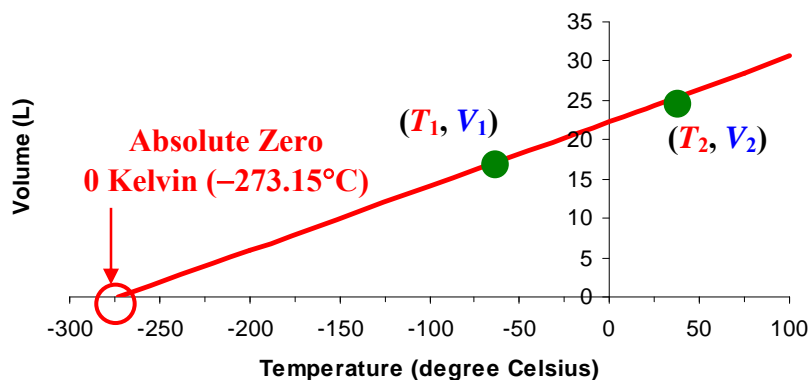
$P_2 = 40.5 \text{ kPa}$

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

$$V \propto T$$

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

k = constant of proportion

Temperature ↑   Volume ↑

Charles's Law (Constant Pressure)

(T_1, V_1)

**Absolute Zero
0 Kelvin (−273.15°C)**

(T_2, V_2)

Charles's Law

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

T_1 = Temperature at Initial Condition (in K) V_1 = Volume at Initial Condition
T_2 = Temperature at Final Condition (in K) V_2 = Volume at Final Condition

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

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

(Change °C to K)

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

$$\frac{V_1 T_2}{T_1} = V_2 \qquad V_2 = \frac{(3.25\,\text{L})(278.15\,\text{K})}{(298.15\,\text{K})}$$

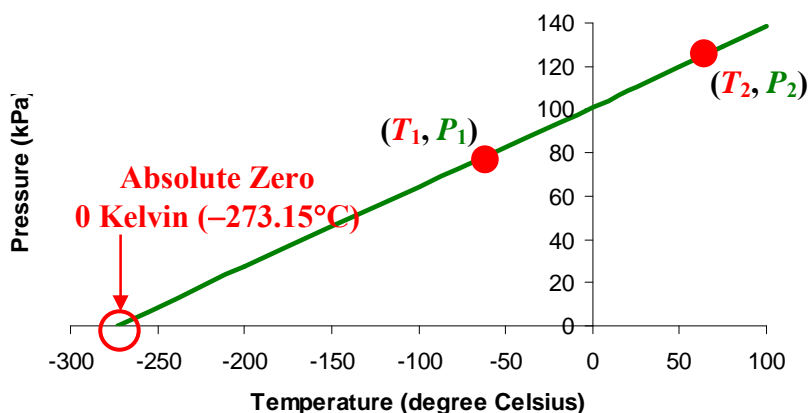**As Temp ↓, Volume ↓** $V_2 = 3.03$ **L**

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

$P \propto T$

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

k = **constant of proportion**

Temperature ↑   Pressure ↑

Gay-Lussac's Law (Constant Volume)

Gay-Lussac's Law

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

T_1 = Temperature at Initial Condition (in K) P_1 = Pressure at Initial Condition
T_2 = Temperature at Final Condition (in K) P_2 = Pressure at Final Condition

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

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

(Change °C to K)
(P can be in atm)

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

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

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

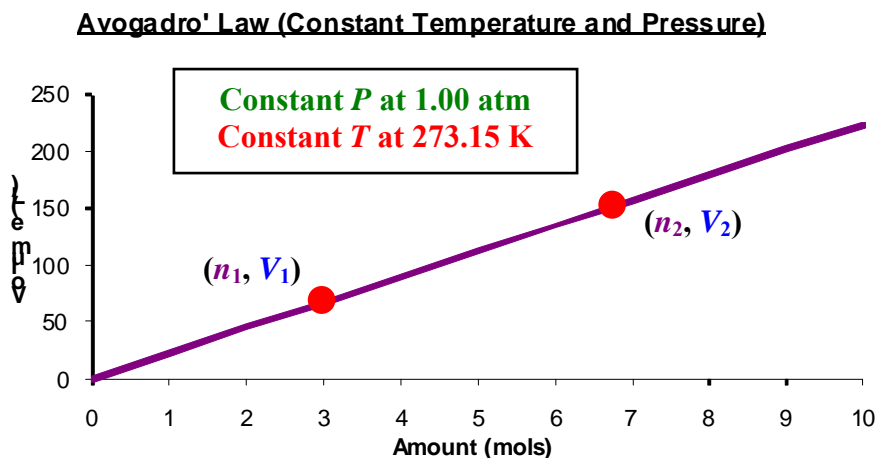**As Temp ↑, Pressure ↑**

$P_2 = 10.4$ **atm**

4. **Avogadro's Law**: - at constant pressure and temperature, **volume is <u>directly proportional</u> 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)

Constant P at 1.00 atm
Constant T at 273.15 K

(n_1, V_1)
(n_2, V_2)

Volume (L)
Amount (mols)

Avogadro's Law

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

n_1 = Amount of Moles at Initial Condition V_1 = Volume at Initial Condition
n_2 = Amount of Moles at Final 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.
   a. Write a balanced equation for this reaction.
   b. Assuming a complete reaction, what is the volume of phosphorus trichloride produced?

a. $P_{4(s)} + 6\ Cl_{2(g)} \rightarrow 4\ PCl_{3(g)}$

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

$n_2 = 9.00 \text{ mol } Cl_2 \times \dfrac{4 \text{ mol } PCl_3}{6 \text{ mol } Cl_2}$

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

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

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

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

As Moles ↓, Volume ↓

$V_2 = 149$ 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 $R = \dfrac{PV}{nT}$ from the Ideal Gas Law)

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

P_1 = Pressure at Initial Condition P_2 = Pressure at Final Condition
V_1 = Volume at Initial Condition V_2 = Volume at Final Condition
T_1 = Temperature at Initial Condition (in K) T_2 = Temperature at Final Condition (in K)
n_1 = Amount of Moles at Initial Condition n_2 = Amount of Moles at Final Condition

Example 5: 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)

$V_1 = 57.2$ L

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

$P_1 = 0.9210526316$ atm

$T_1 = 25.0°C = 298.15$ K

$V_2 = ?$

$P_2 = 0.365$ atm

$T_2 = -45.0°C = 228.15$ K

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

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

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

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

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

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

Example 6: 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?

$n_1 = 0.852$ mol $n_2 = 0.750$ mol

$T_1 = 23.0°C = 296.15$ K

$T_2 = 148.0°C = 421.15$ K

$P_1 = 752$ mm Hg

$P_2 = 2854$ mm Hg

$V_1 = 1.00$ L $V_2 = ?$

$\Delta V = V_2 - V_1 = ?$

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

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

$$V_2 = 0.330\ L$$

$$\Delta V = V_2 - V_1 = 0.330\ L - 1.000\ L$$

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

Example 7: 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)

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

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

The new volume will be three times the original volume.

Assignment
12.2 pg. 425 #1 to 4 (Practice); pg. 428 #1 to 4 (Practice);
pg. 431 #1 to 3 (Practice); pg. 432 #1 to 12

12.3: Molecular Composition of Gases

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

Ideal Gas Law
$$PV = nRT$$

P = **Pressure (kPa or atm)** V = **Volume (L)**
n = **Amount of Gas (mol)** T = **Temperature (K)**

R = **Gas Constant = 0.0821** $\dfrac{\text{L} \bullet \text{atm}}{\text{K} \bullet \text{mol}}$ or   **8.314** $\dfrac{\text{L} \bullet \text{kPa}}{\text{K} \bullet \text{mol}}$

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. L
T = 15.0°C = 288.15 K
P = 32.0 atm
R = 0.0821 (L $\bullet$ atm)/(K $\bullet$ mol)
m = ?
n = ? (need to find n first)

$$PV = nRT$$
$$\frac{PV}{RT} = n$$
$$n = \frac{(32.0 \,\text{atm})(200. \,\text{L})}{(0.0821 \frac{\text{L}\bullet\text{atm}}{\text{K}\bullet\text{mol}})(288.15 \,\text{K})}$$
$$n = 270.5317195 \text{ mol}$$

For propane, C_3H_8, M = 44.11 g/mol

$m = nM$
m = (270.5317195 mol)(44.11 g/mol)
m = 11933.15415 g
$m = 1.19 \times 10^4$ g = 11.9 kg

Standard Mole-Volume Relationships:

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

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

STP = 22.4 L/mol @ 0.00°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
$n = 5.00 \,\text{L} \times \dfrac{1 \,\text{mol}}{22.4 \,\text{L}}$ $n = 0.223$ **mol**

b. SATP

SATP = 24.8 L/mol
$n = 5.00 \,\text{L} \times \dfrac{1 \,\text{mol}}{24.8 \,\text{L}}$ $n = 0.202$ **mol**

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

a. STP

$n = \dfrac{m}{M} = \dfrac{3.50 \,\text{g}}{28.02 \,\text{g/mol}} = 0.124910778 \text{ mol}$
STP = 22.4 L/mol
V = (0. 124910778 mol)(22.4 L/mol) $V = 2.80$ **L**

b. SATP

$n = \dfrac{m}{M} = \dfrac{3.50 \,\text{g}}{28.02 \,\text{g/mol}} = 0. 124910778 \text{ mol}$
SATP = 24.8 L/mol
V = (0. 124910778 mol)(24.8 L/mol) $V = 3.10$ **L**
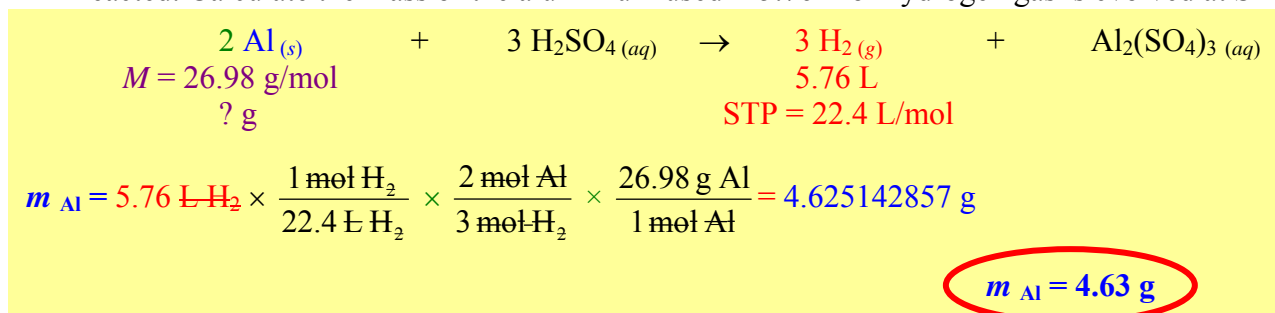
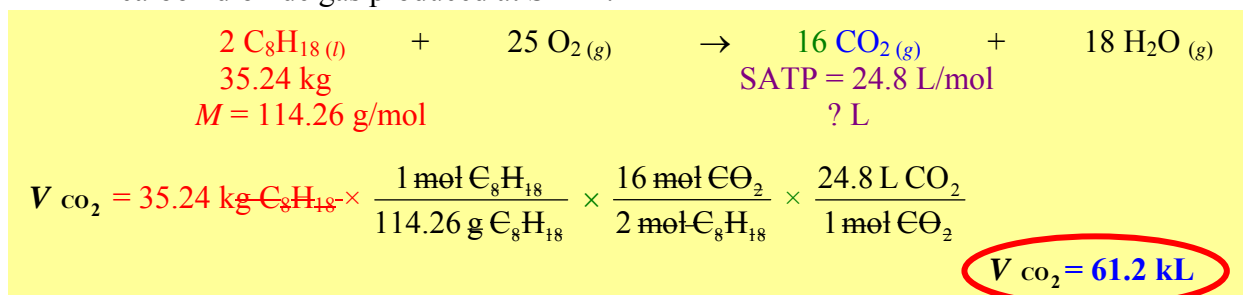**Gaseous Stoichiometry Procedure** (Ideal Gas, STP, or SATP)

1. Predict the products and balance the chemical equation.

2. Put all the information given under the appropriate chemicals.

3. Find the moles of the given chemical: $\left(n = \dfrac{m}{M} \right)$ or $\left(n = \dfrac{PV}{RT} \right)$ or

    (at STP: n = Volume × $\dfrac{1 \, \text{mol}}{22.4 \, \text{L}}$; at SATP: n = Volume × $\dfrac{1 \, \text{mol}}{24.8 \, \text{L}}$).

4. **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.

5. Find the mole of the required chemical using mole ratio:

    $\left(\textbf{mol of require} = \textbf{mol of given} \times \dfrac{\textbf{require coefficient}}{\textbf{given coefficient}} \right)$

6. Convert mole of the required chemical to its mass, concentration or volume equivalence: ($m = nM$) or or ($PV = nRT$) or (at STP: $V = n \times 22.4$ L/mol; at SATP = $n \times 24.8$ L/mol).

Example 4: 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.

$$2 \, Al_{(s)} \quad + \quad 3 \, H_2SO_{4 \, (aq)} \quad \rightarrow \quad 3 \, H_{2 \, (g)} \quad + \quad Al_2(SO_4)_{3 \, (aq)}$$

$M = 26.98$ g/mol $\qquad\qquad\qquad$ 5.76 L

? g $\qquad\qquad\qquad\qquad\qquad$ STP = 22.4 L/mol

$$m_{Al} = 5.76 \, \cancel{L \, H_2} \times \frac{1 \, \cancel{\text{mol } H_2}}{22.4 \, \cancel{L \, H_2}} \times \frac{2 \, \cancel{\text{mol } Al}}{3 \, \cancel{\text{mol } H_2}} \times \frac{26.98 \, g \, Al}{1 \, \cancel{\text{mol } Al}} = 4.625142857 \, g$$

$$\boxed{m_{Al} = 4.63 \, g}$$

Example 5: 35.24 kg of liquid octane (C_8H_{18}) is burned under excess oxygen. Determine the volume of carbon dioxide gas produced at SATP.
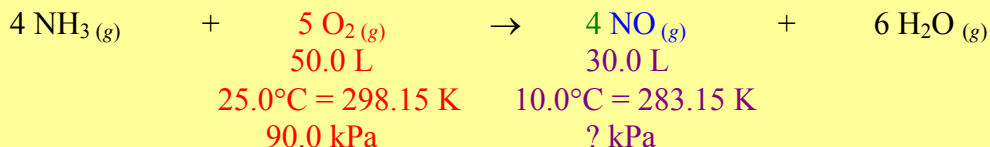
$$2 \, C_8H_{18 \, (l)} \quad + \quad 25 \, O_{2 \, (g)} \quad \rightarrow \quad 16 \, CO_{2 \, (g)} \quad + \quad 18 \, H_2O_{(g)}$$

35.24 kg $\qquad\qquad\qquad\qquad\qquad\qquad$ SATP = 24.8 L/mol

$M = 114.26$ g/mol $\qquad\qquad\qquad\qquad$? L

$$V_{CO_2} = 35.24 \, \cancel{kg \, C_8H_{18}} \times \frac{1 \, \cancel{\text{mol } C_8H_{18}}}{114.26 \, \cancel{g \, C_8H_{18}}} \times \frac{16 \, \cancel{\text{mol } CO_2}}{2 \, \cancel{\text{mol } C_8H_{18}}} \times \frac{24.8 \, L \, CO_2}{1 \, \cancel{\text{mol } CO_2}}$$

$$\boxed{V_{CO_2} = 61.2 \, kL}$$

Example 6: Ammonia is reacted with oxygen to form nitrogen monoxide and water vapour.

$$4\ NH_{3\ (g)} + 5\ O_{2\ (g)} \rightarrow 4\ NO_{(g)} + 6\ H_2O_{(g)}$$
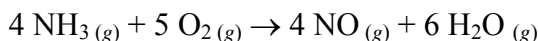
If 50.0 L of oxygen at 90.0 kPa at 25.0°C were allowed to react with excess ammonia, what would be the pressure of nitrogen monoxide in a collector vessel measuring 30.0 L at a temperature of 10.0°C?
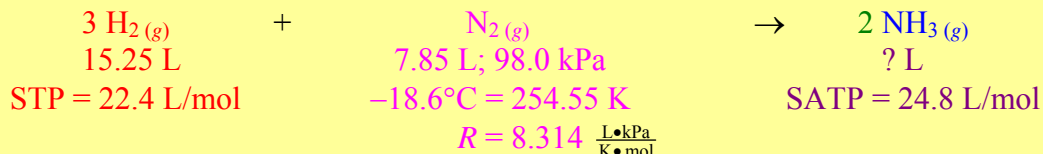
$4\ NH_{3\ (g)}$	+	$5\ O_{2\ (g)}$	$\rightarrow$	$4\ NO_{(g)}$	+	$6\ H_2O_{(g)}$
		50.0 L		30.0 L		
		25.0°C = 298.15 K		10.0°C = 283.15 K		
		90.0 kPa		? kPa		

Using Ideal Gas Law to find the moles of O_2 first ($n = \frac{PV}{RT}$). Once we calculate the moles of NO using the mole ratio, we can convert n_{NO} to P_{NO} ($P = \frac{nRT}{V}$).

$$P_{NO} = \frac{(90.0\ \text{kPa } O_2)(50.0\ \text{L } O_2)}{\left(8.314\ \frac{\text{kPa } O_2 \bullet \text{L } O_2}{\text{mol } O_2 \bullet \text{K } O_2}\right)(298.15\ \text{K } O_2)} \times \frac{4\ \text{mol NO}}{5\ \text{mol } O_2} \times \frac{\left(8.314\ \frac{\text{kPa NO} \bullet \text{L NO}}{\text{mol NO} \bullet \text{K NO}}\right)(283.15\ \text{K NO})}{30.0\ \text{L NO}} \quad \boxed{P_{NO} = 114\ \text{kPa}}$$

If we view the coefficients of NO and O_2 as n_1 and n_2, then we can use the Combined Gas Law instead of Stoichiometry.

$n_1 = 5\ \text{mol } O_2$ $n_2 = 4\ \text{mol NO}$
$T_1 = 296.15\ \text{K } O_2$ $T_2 = 283.15\ \text{K NO}$
$V_1 = 50.0\ \text{L } O_2$ $V_2 = 30.0\ \text{L NO}$
$P_1 = 90.0\ \text{kPa } O_2$ $P_2 = ?\ \text{kPa NO}$

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

$$P_2 = \frac{(90.0\ \text{kPa})(50.0\ \text{L})(4\ \text{mol})(283.15\ \text{K})}{(5\ \text{mol})(296.15\ \text{K})(30.0\ \text{L})} \quad \boxed{P_{NO} = 114\ \text{kPa}}$$

Example 7: 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.

$3\ H_{2\ (g)}$	+	$N_{2\ (g)}$	$\rightarrow$	$2\ NH_{3\ (g)}$
15.25 L		7.85 L; 98.0 kPa		? L
STP = 22.4 L/mol		−18.6°C = 254.55 K		SATP = 24.8 L/mol
		$R = 8.314\ \frac{\text{L} \bullet \text{kPa}}{\text{K} \bullet \text{mol}}$		

Since there is enough information to determine the moles of two reactants (quantities of both reactants are given), we need to find the volume of the product from each of these reactant before labelling which reactant is limiting.

① $V_{NH_3} = 15.25\ \text{L } H_2 \times \frac{1\ \text{mol } H_2}{22.4\ \text{L } H_2} \times \frac{2\ \text{mol } NH_3}{3\ \text{mol } H_2} \times \frac{24.8\ \text{L } NH_3}{1\ \text{mol } NH_3} = \boxed{\textbf{11.3 L } NH_3}$

<div align="right">(smaller product volume)</div>

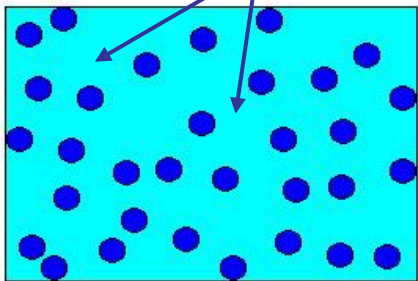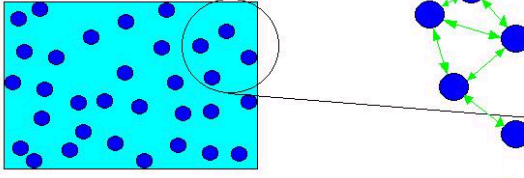Using Ideal Gas Law to find the moles of N_2 first ($n = \frac{PV}{RT}$)

$$V_{NH_3} = \frac{(98.0\ \text{kPa})(7.85\ \text{L } N_2)}{\left(8.314\ \frac{\text{kPa} \bullet \text{L}}{\text{mol} \bullet \text{K}}\right)(254.55\ \text{K})} \times \frac{2\ \text{mol } NH_3}{1\ \text{mol } N_2} \times \frac{24.8\ \text{L } NH_3}{1\ \text{mol } NH_3} = 18.0299\ldots\ \text{L } NH_3$$

Since H_2 gives a smaller calculated product volume, <u>**H_2 is the limiting reactant; N_2 is the excess**</u>.

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	Real Gas
Low Pressure	High Pressure
High Temperature	Low Temperature

V_{actual} = **Volume of Spaces between Molecules**

At low pressure, there are a lot of spaces between molecules. Therefore,
$V_{actual} \approx V_{container}$

At high pressure, there is a lot less space between molecules and
$V_{actual} < V_{container}$

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.**

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

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

Dalton's Law of Partial Pressures

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

$$\frac{n_{Total}RT}{V} = \frac{n_1 RT}{V} + \frac{n_2 RT}{V} + \frac{n_3 RT}{V} + \dots$$

(Using Ideal Gas Law $P = \dfrac{nRT}{V}$**)**

(R, T, and V are common to all components in gas mixture)

$$n_{Total}\left(\frac{RT}{V}\right) = \left(\frac{RT}{V}\right)(n_1 + n_2 + n_3 + \dots)$$

(Take out Common Factor $\dfrac{RT}{V}$ and cancel both sides)

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

Example 8: Write the equations for partial pressures and mole components for air, which is composed of N_2, O_2, Ar, CO_2, and other trace gases (http://www.mistupid.com/chemistry/aircomp.htm).

$$P_{total} = P_{N_2} + P_{O_2} + P_{Ar} + P_{CO_2} + P_{trace\,gases}$$

$$n_{total} = n_{N_2} + n_{O_2} + n_{Ar} + n_{CO_2} + n_{trace\,gases}$$

Example 9: Write the equations for partial pressures for air, which is composed of N_2, O_2, CO_2, Ar, and other trace gases (78.084%, 20.947%, 0.033%, 0.934%, 0.002% by volume respectively). Determine the partial pressures of these gases for 1 atm of air in kPa.

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

$P_{N_2} = (0.780\ 84)(101.325\ \text{kPa})$ $P_{N_2} = \textbf{79.115 kPa}$

$P_{O_2} = (0.209\ 47)(101.325\ \text{kPa})$ $P_{O_2} = \textbf{21.225 kPa}$

$P_{CO_2} = (0.000\ 33)(101.325\ \text{kPa})$ $P_{CO_2} = \textbf{0.033 kPa}$

$P_{Ar} = (0.009\ 34)(101.325\ \text{kPa})$ $P_{Ar} = \textbf{0.946 kPa}$

$P_{trace\ gases} = (0.00\ 002)(101.325\ \text{kPa})$ $P_{trace\ gases} = \textbf{0.002 kPa}$

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

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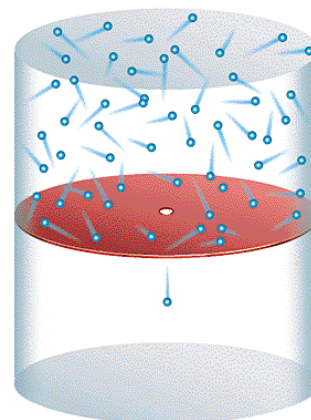
Graham's Law of Effusion: - the effusion rate of a gas is proportional to the square root of its molar mass.

Graham's Law of Effusion

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

$M = \textbf{Molar Mass in g/mol or kg/mol}$

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

Example 10: Calculate the ratio of effusion rates between $NH_{3\ (g)}$ and $HCl_{(g)}$ under the same temperature.

$M_{HCl} = 36.46$ g/mol
$M_{NH_3} = 17.04$ g/mol

$\dfrac{r_{NH_3}}{r_{HCl}} = ?$

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

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

NH_3 effuse about 1.5 times faster than HCl.

Example 11: Determine the effusion rate F_2 through a porous barrier if the effusion rate of N_2 through the same barrier is 274 mL/min.

r_{N_2} = 274 mL/min

M_{F_2} = 38.00 g/mol

M_{N_2} = 28.02 g/mol

r_{F_2} = ?

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

Since F_2 is bigger than N_2, r_{F_2} is less than r_{N_2} r_{F_2} = 235 mL/min

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 <u>Rate of Diffusion is Greater than the Rate of Effusion</u> is when the <u>gas is traveled FROM an area of High Pressure to an area of very Low Pressure (near vacuum)</u>** – like a breach on the hull of a spacecraft or a plane.

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

Assignment
12.3 pg. 435 #1 to 4 (Practice); pg. 438 #1 to 4 (Practice);
pg. 442 #1 to 3 (Practice); pg. 442 #1 to 14
Chapter 12 Review pg. 445−447 #1, 5 to 11, 13, 14, 16, 18, 21,
23, 25 to 28, 30 to 32, 35, 36, 37 to 53 (odd),
54 to 61, 63 to 65

Chapter 13: Solutions

13.1: What is Solution?

Solution: - a homogeneous mixture of two or more substances even at the microscopic level.

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.

Note: - For solution with molecular solutes (such as sugar), all molecules are separated from each other within the solvent. For solution with ionic solutes (such as salt), all ionic units are separated into ions within the solvent.

Different Types of Solutions in Various Phases:

Solute	Solvent	Solution	Examples
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)

Suspension: - a mixture that appears as homogeneous solution at first while the fine particles are being stirred in the solvent, but eventually the two components separate over time.

    Examples: Fine Clay in Water, some Salad Dressings with similar colour of Vinegar and Oil.

Colloid: - a mixture consisting of tiny particles that are intermediate in size between those in solutions and those in suspensions.
- the particles have similar charges and hence, do not tend to collect into larger sizes and collect to the bottom like suspensions do.
- However, they are not considered solutions because the particles did not break down at a molecular level.

    Examples: Milk (each particle has proteins and fat globules called whey)
              Blood (platelets, plasma, red and white blood cells – can be separated by centrifuging)

Methods of Separating Mixtures:

1. 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.

2. Precipitation: - the process in which dissolved solute comes out of the solution and form small particles over a relative short period of time.
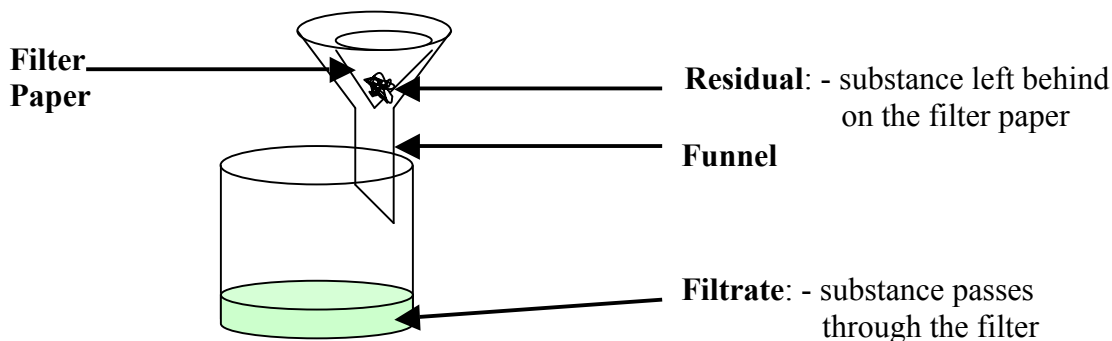
(Left) Lead iodide precipitate from the reaction between lead chlorate and potassium iodide

(Right) Home-grown alum crystals

3. Filtration: - using a filter and a funnel, a mechanical (heterogeneous) mixture consists of liquids and solids can be separated. (The solid solutes must be bigger than the holes in the filter paper.)

Filter Paper

Residual: - substance left behind on the filter paper

Funnel

Filtrate: - substance passes through the filter

4. Distillation: - an aqueous solution that consists of a liquid solute can be separated by evaporation of the substance with a lower boiling point followed by condensation.

5. Centrifuge: - using high speed spinning motion (centripetal force), a suspension and even some colloid can be separated into their components.
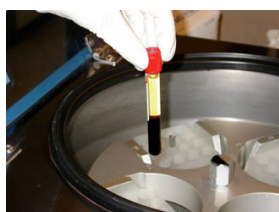
(left) Blood processing from a centrifuge. The top layer is plasma while the bottom layer is red blood cell. (right) A Distillation apparatus

thermometer

water outlet

distilling flask

condenser

mixture

cold water inlet   distillate
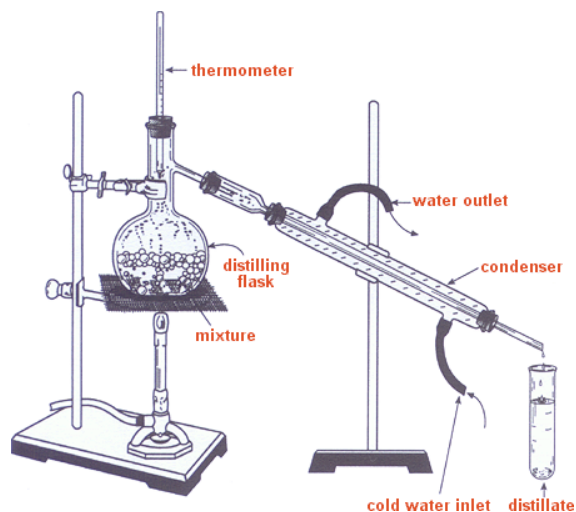
6. Chromatography: - a collective term that describes separation techniques that utilizes passing a mixture dissolves in a "mobile phase" (the solvent that moves in a definite direction) through a stationary phase.
- since chemicals in a mixture has small (or large) differences in polarity, they will spend different amount of time as they move with the "mobile phase" through a stationary phase. These differences in rates separate various components of a mixture.

Polar Solvent

Most Polar Component travels faster up the paper

Assignment
13.1 pg. 459 #1 to 12
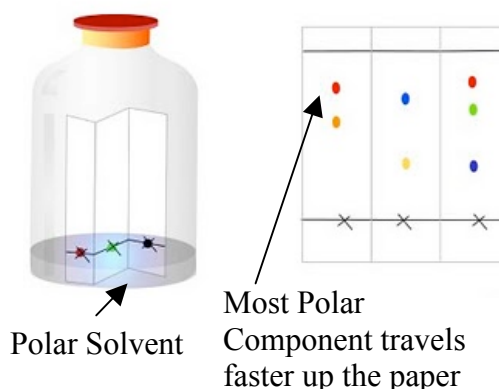
13.2: Concentration and Molarity

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

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Parts per Million

$$\text{ppm} = \frac{m_{solute} \text{ (mg)}}{V_{solution} \text{ (L)}} = \frac{\text{g of solute}}{1 \text{ million g of solvent}}$$

</div>

Example 1: 4.52×10^{-2} g of sulfate ion is dissolved in 500. mL (also equal to 500. g) of water. Calculate the concentration of sulfate ion in ppm.

4.52×10^{-2} g of SO_4^{2-} = 0.0452 g SO_4^{2-} = 45.2 mg SO_4^{2-}

$$\text{ppm} = \frac{m_{solute} \text{ (mg)}}{V_{solution} \text{ (L)}} = \frac{45.2 \text{ mg}}{0.500 \text{ L}} = \boxed{\textbf{90.4 ppm } SO_4^{2-}}$$

OR $\text{ppm} = \frac{\text{g of solute}}{1 \text{ million g of solvent}} = \frac{4.52 \times 10^{-2} \text{ g}}{500 \text{ g}} \times \frac{1,000,000 \text{ parts}}{1 \text{ million}} = \boxed{\textbf{90.4 ppm } SO_4^{2-}}$

Molarity (Molar Concentration): - moles of solute per Litre of <u>solution</u> (M = mol/L).

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Molarity (Concentration)

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

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

Concentration Notation: $[Na^+]$ means concentration of Na^+ ion

</div>

Example 2: 3.46 g of copper (II) nitrate is dissolved in 250.0 mL of water. Calculate the molarity of the solution formed.

$m = 3.46$ g $Cu(NO_3)_2$

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

$n = 0.0184464467$ mol $Cu(NO_3)_2$

$V = 250.0$ mL $= 0.2500$ L

$C = ?$

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

$\boxed{\textbf{C = 0.0738 mol/L or 73.8 mmol/L}}$

Example 3: Determine the mass of sodium dichromate needed for 500.0 mL of 0.0300 M.

$V = 500.0$ mL $= 0.5000$ L
$C = 0.0300$ mol/L
$M = 261.98$ g/mol $Na_2Cr_2O_7$
$n = ?$ $m = ?$

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

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

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

$\textbf{n = 0.0150 mol}$

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

$\boxed{\textbf{m = 3.93 g}}$

Example 4: The Palmense Mineral Water from the city of Fermo in Italy has a sodium hydrogen carbonate concentration of 5.207 mmol/L. What volume of this mineral water will contain 4.00 mg of sodium hydrogen carbonate?

$m = 4.00$ mg $NaHCO_3$

$n = \dfrac{4.00 \text{ mg}}{84.01 \text{ g/mol}}$

$n = 0.0476133794$ mmol $NaHCO_3$

$C = 5.207$ mmol/L

$V = ?$

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

$V = \dfrac{n}{C} = \dfrac{0.0476133794 \text{ mmol}}{5.207 \text{ mmol}/L}$

$V = 0.00914$ **L or 9.14 mL**

Solution Preparation from a Solid Solute:

1. Calculate the mass of the solute required using **($n = CV$) and ($m = nM$)**.
2. Measure out the calculated mass carefully on a balance.
3. Completely dissolve the solute with about **_half_** the volume of the solvent.
4. Transfer the solution into a volumetric flask of the proper size (the same as the volume of the solution needed). Wash the beaker and the stirring rod with more solvent and carefully transfer the washed fluid into the volumetric flask.
5. Cap the volumetric flask, invert and shake a few times to ensure through mixing.

A Volumetric Flask and filling the solvent's meniscus to the line on its neck.

Solution Preparation Video:
http://www.youtube.com/watch?v=cckAwavEKA0

Solution Stoichiometry Procedure

1. Predict the products and balance the chemical equation.

2. Put all the information given under the appropriate chemicals.

3. **Check if the problem is limiting reactant question** (if you are given enough information to find the moles of two chemicals). You might have to do the following steps twice starting with different reactant.

4. Find the moles of the given chemical: $\left(n = \frac{m}{M}\right)$ or **($n = CV$)** or relevant gas quantity to mole conversion.

5. Find the mole of the required chemical using mole ratio:

$$\left(\textbf{mol of require} = \textbf{mol of given} \times \frac{\textbf{require coefficient}}{\textbf{given coefficient}}\right)$$

6. Convert mole of the required chemical to its mass, concentration or volume equivalence: **($m = nM$)** or $\left(C = \frac{n}{V}\right)$ or $\left(V = \frac{nRT}{P}\right)$ or **(at STP: $V = n \times 22.4$ L/mol; at SATP $= n \times 24.8$ L/mol).**
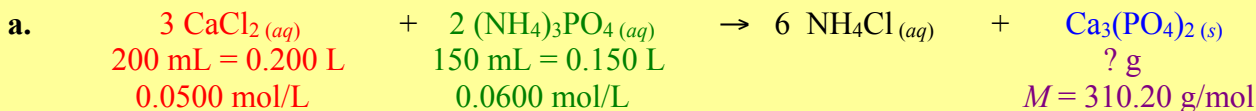
Example 5: What volume of 1.75 M of HBr $_{(aq)}$ would be needed to react completely with 28.4 g of
Na$_2$CO$_{3\ (s)}$ to produce water, carbon dioxide and NaBr $_{(aq)}$?

$$2\ \text{HBr}_{(aq)}\ \ +\ \ \text{Na}_2\text{CO}_{3\ (aq)}\ \longrightarrow\ \ \text{H}_2\text{O}_{(l)}\ +\ \ \text{CO}_{2\ (g)}\ +\ 2\ \text{NaBr}_{(aq)}$$

C = 1.75 mol/L   28.4 g
? L M = 105.99 g/mol

$$V_{\text{HBr}} = 28.4\ \cancel{\text{g Na}_2\text{CO}_3} \times \frac{1\ \cancel{\text{mol Na}_2\text{CO}_3}}{105.99\ \cancel{\text{g Na}_2\text{CO}_3}} \times \frac{2\ \cancel{\text{mol HBr}}}{1\ \cancel{\text{mol Na}_2\text{CO}_3}} \times \frac{1\ \text{L HBr}}{1.75\ \cancel{\text{mol HBr}}} = 0.3062283504\ \text{L}$$

$$\boxed{V_{\text{HBr}} = 0.306\ \text{L} = 306\ \text{mL}}$$
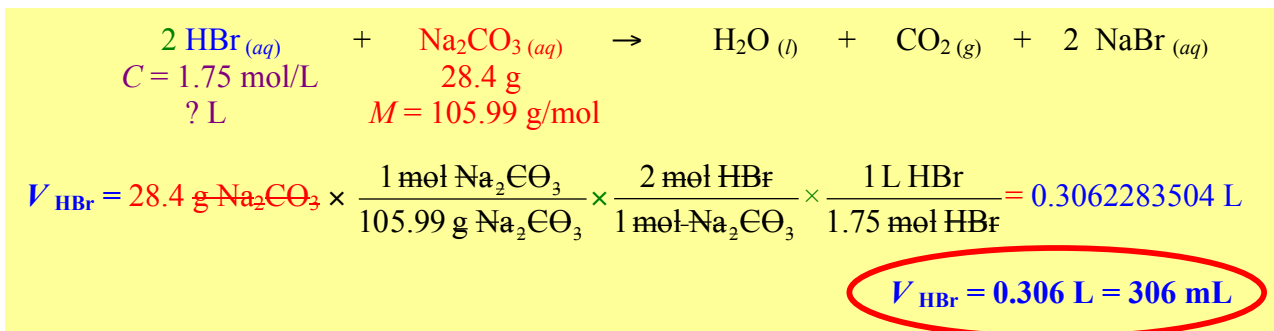
Example 6: 200. mL of 0.0500 M of calcium chloride is reacted with 150. mL of 0.0600 M of ammonium
phosphate.
a. Determine the mass of the precipitate formed in this reaction.
b. If the experimental mass of the precipitate is 1.28 g, calculate the % error. How can you
interpret this result?

a. $3\ \text{CaCl}_{2\ (aq)}$ $+\ \ 2\ (\text{NH}_4)_3\text{PO}_{4\ (aq)}$ $\longrightarrow\ 6\ \text{NH}_4\text{Cl}_{(aq)}$ $+$ $\text{Ca}_3(\text{PO}_4)_{2\ (s)}$

200 mL = 0.200 L   150 mL = 0.150 L   ? g
0.0500 mol/L   0.0600 mol/L M = 310.20 g/mol

**Since there is enough information to determine the moles of two reactants (quantities of both
reactants are given), we need to find the mass of the precipitate from each of these reactant
before labelling which reactant is limiting.**

$$m_{\text{Ca}_3(\text{PO}_4)_2} = 0.200\ \cancel{\text{L CaCl}_2} \times \frac{0.0500\ \cancel{\text{mol CaCl}_2}}{1\ \cancel{\text{L CaCl}_2}} \times \frac{1\ \cancel{\text{mol Ca}_3(\text{PO}_4)_2}}{3\ \cancel{\text{mol CaCl}_2}} \times \frac{310.20\ \text{g Ca}_3(\text{PO}_4)_2}{1\ \cancel{\text{mol Ca}_3(\text{PO}_4)_2}}$$

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

$$m_{\text{Ca}_3(\text{PO}_4)_2} = 0.150\ \cancel{\text{L (NH}_4)_3\text{PO}_4} \times \frac{0.0600\ \cancel{\text{mol (NH}_4)_3\text{PO}_4}}{1\ \cancel{\text{L (NH}_4)_3\text{PO}_4}} \times \frac{1\ \cancel{\text{mol Ca}_3(\text{PO}_4)_2}}{2\ \cancel{\text{mol (NH}_4)_3\text{PO}_4}} \times \frac{310.20\ \text{g Ca}_3(\text{PO}_4)_2}{1\ \cancel{\text{mol Ca}_3(\text{PO}_4)_2}}$$

$$m_{\text{Ca}_3(\text{PO}_4)_2} = 1.40\ \text{g Ca}_3(\text{PO}_4)_2$$

Since CaCl$_2$ gives a smaller calculated precipitate mass, **CaCl$_2$ is the limiting reactant;**
(NH$_4$)$_3$PO$_4$ is the excess.

b.   % error = $\dfrac{|\text{Experimental} - \text{Theoretical}|}{\text{Theoretical}} \times 100\% = \dfrac{|1.28\ \text{g} - 1.03\ \text{g}|}{1.03\ \text{g}} \times 100\%$ $\boxed{\textbf{% error = 24.3\%}}$

**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).**

Assignment
13.2 pg. 461 #1 to 7 (Practice); pg. 465 #1 to 7 (Practice);
pg. 467 #1 to 3 (Practice); ~~pg. 467 #1 to 14~~

13.3: Solubility and the Dissolving Process

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

Dilution

$$C_1V_1 = C_2V_2$$

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

Example 1: Concentrated hydrochloric acid comes in 17.4 M. What is the volume of concentrated $HCl_{(aq)}$ needed to obtain 250. mL of 1.50 M of $HCl_{(aq)}$?

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

$C_1V_1 = C_2V_2$

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

$$\boxed{V_1 = 21.6 \text{ mL}}$$

Example 2: Determine the volume of water needed to dilute 30.0 mL of 0.500 M $CuSO_{4\,(aq)}$ to 0.0750 M.

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

$C_1V_1 = C_2V_2$

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

$$V_2 = 200. \text{ mL}$$

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

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

$$\boxed{V_{water} = 170. \text{ mL}}$$

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

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.

Reason that Water is a Common Solvent:

1. **Polar Molecule**: - dissolves many ionic compounds due to its ability to attract cations and anions
            (electrolytes).
            - *Note:* **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.

   Dissociation: - the process when an ionic solute breaks apart into its cation and anion during the
            dissolving process.

      Example: Sodium Chloride (Soluble Ionic Compounds)

$$NaCl_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + Cl^-_{(aq)}$$
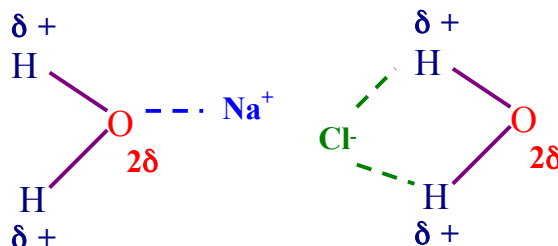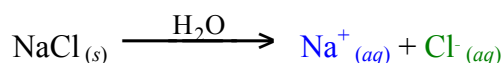
2. **Strong O-H Hydrogen Bond**: - dissolves many molecular compounds that have O-H hydrogen bonds.

Example: $CH_3OH_{(l)}$ (Methanol)
(molecular solute with covalent bonds)

Strong O-H Hydrogen Bonds
(intermolecular force)
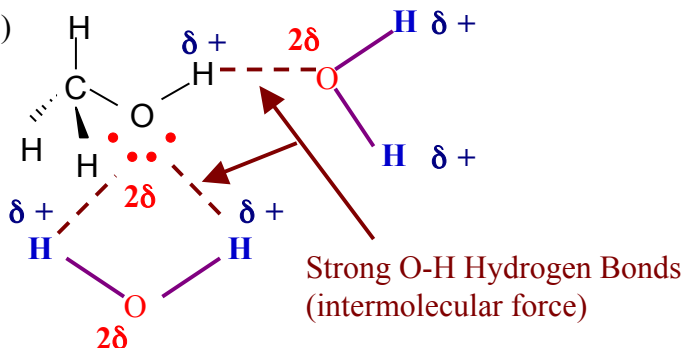
"Like-Dissolves-Like": - a general solubility principle that states polar solvents tend to dissolve polar
solutes; non-polar solvents tend to dissolve non-polar solutes.

Examples: Vitamin C is very soluble in water (polar solvent) because its structure has many –OH bonds.
Vitamin A is soluble in fat (non-polar solvent) because its structure is largely non-polar
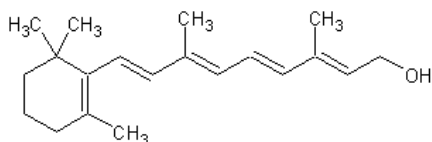
Molecular Structure of Vitamin A
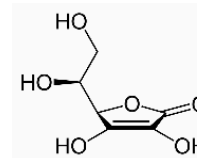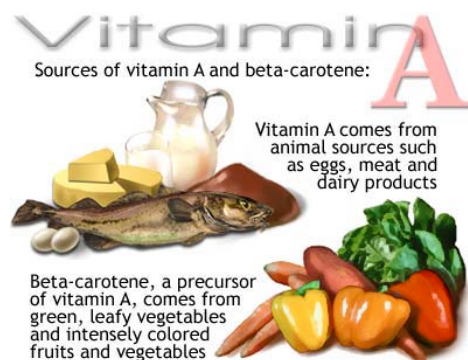
Molecular Structure of Vitamin C

General Descriptions of Solutions with Various Solubilities:

1. **Miscible**: - when two liquids can dissolve each other completely with infinite solubility.
              (Example: Vinegar-Water Solution – Both have polar bonds.)

2. **Partially Miscible**: - when two liquids are slightly soluble in each other.
              (Example: Anaesthetic Diethyl ether-Water Solution – weak polar versus strong polar bonds.)

3. **Immiscible**: - when two liquids cannot dissolve at all, which result in a mechanical or heterogeneous mixture.
              (Example: Italian Salad Dressing – Vinegar and Oil: polar versus non-polar bonds.)

General Descriptions of Solutions at Various Levels of Concentration:

1. **Unsaturated Solution**: - when a solvent can still dissolve more solute at a particle temperature.

2. **Saturated Solution**: - when a solvent cannot dissolve anymore solute.
   - the rate of solvation is the same as the rate of crystallization (at equilibrium).

3. **Supersaturated Solution**: - when a saturated solution is heated and more solute is dissolved, which is then followed by a slow cooling without any recrystallization.
   - the density and the viscosity of a supersaturated solution is higher than a saturated solution. (Example: Sugar Syrup – Supersaturated Sugar Water)

Soluble: - when a solute can dissolve in a solvent.

Insoluble: - when a solute does not dissolve in a solvent at all. (polar and non-polar substances)

Slightly Soluble: - when a solute has a very low solubility (some ionic substances – see chart below).

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^- $OOCCOO^{2-}$	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)	$RbClO_4$ $CsClO_4$ $AgCH_3COO$ $Hg_2(CH_3COO)_2$	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:** $Co(IO_3)_2$ $Fe_2(OOCCOO)_3$ are soluble	most	most

**Infinitely Soluble**: - when a solute has no saturation point in a solvent.

Examples: Ethanol (a form of alcohol) can dissolve infinitely in water (90% grain alcohol).
All gases can mixed infinitely with each other.

(Right) Methanol is very soluble in water because both molecules have O-H bonds. Therefore, hydrogen bonds are created between the molecules. The strong intermolecular forces ensure complete mixing at the molecular level.
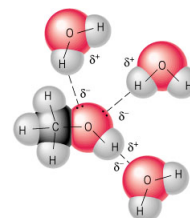
Precipitate: - solid solute reappears when the solution reached saturation point.
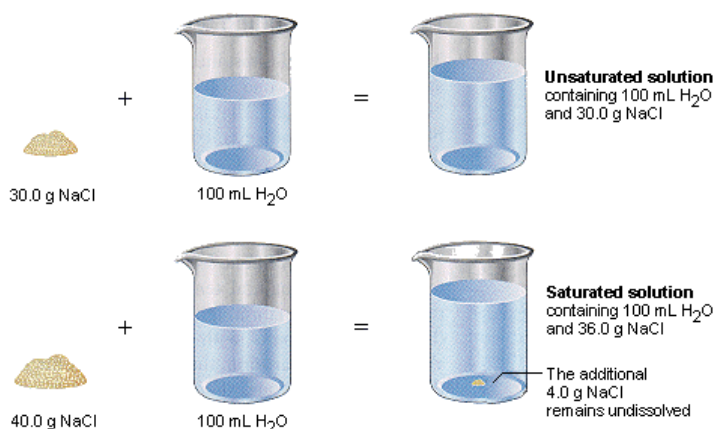
(Left) Salt has a solubility of 36 g / 100 g of H_2O at 25°C. When 30.0 g NaCl is dissolved in 100 g of water, all of it dissolves. However, when 40.0 g of salt is dissolved in 100 mL of water, only 36 g is allowed to dissolve and 4 g is left over as precipitate.

**Factors affecting Solubility**:

1. Solubility verus Temperature:

- as **temperature increases**, solubility of **some solid solutes generally increases.**
  (exceptions are sodium sulfate and cerium sulfate).
- when temperature increases, the solvent molecules move faster and the amount of space between them increases as well. As this expansion proceeds, there is more room for the solid solutes to get in between the solvent molecules. Hence, solubilities of solid solutes increase with temperature.
- 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 ↓**
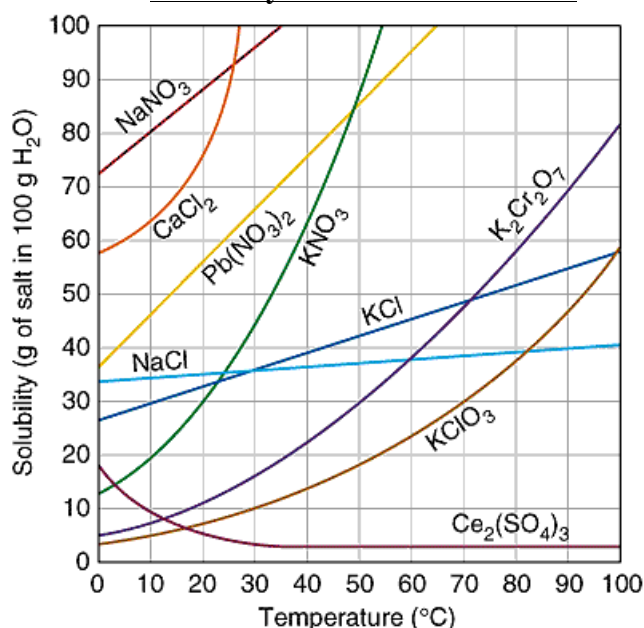
Solubility of some Solid Solutes   **Solubility of some Gaseous Solutes**
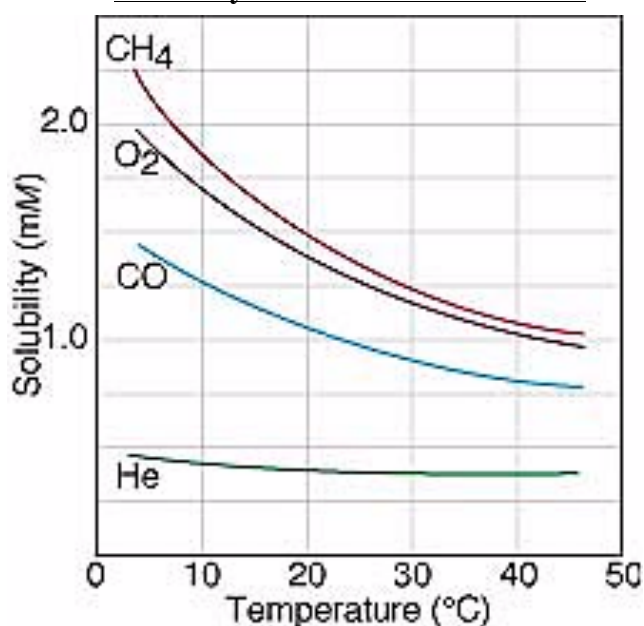
Example 3: A maximum 6.70 g of carbon dioxide gas is dissolved in 2000 g of water at 0°C under normal atmosphere pressure (101.3 kPa). Calculate the solubility of carbon dioxide gas in g / 100 g of H_2O.

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

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

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

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

$$x = 0.335 \text{ g } CO_2 \quad \boxed{\text{Solubility} = 0.335 \text{ g } CO_2 \text{ / 100 g of water}}$$

Example 4: The solubility of sucrose (cane sugar) is 230.9 g / 100 g of H_2O. Determine the amount of solvent needed to dissolve 3.00 kg of cane sugar.

$$\text{Solubility} = \frac{230.9 \text{ g sucrose}}{100 \text{ g } H_2O} \times \frac{3000 \text{ g sucrose}}{x \text{ g } H_2O} \quad \xleftarrow{} \begin{array}{l} \textbf{3.00 kg = 3000 g} \\ \text{(cross-multiply)} \end{array}$$

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

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

$$x = 1299.263751 \text{ g water} \quad \boxed{x = 1.30 \text{ kg of } H_2O}$$

Example 5: Using the graph of the solid solutes solubilities on the previous page, calculate the amount of precipitate when a 300 mL saturated $K_2Cr_2O_{7 \, (aq)}$ is cooled from 90°C to 50°C.

From the graph, the solubility of $K_2Cr_2O_7$ is **70 g / 100 g of H_2O at 90°C** and **30 g / 100 g of H_2O at 50°C.**

For 300 mL of solution (300 g of H_2O used), the maximum solute amounts at each temperature are:

90°C: $\quad \dfrac{70 \text{ g } K_2Cr_2O_7}{100 \text{ g } H_2O} = \dfrac{x \text{ g } K_2Cr_2O_7}{300 \text{ g } H_2O}$ $\qquad\qquad$ 50°C: $\quad \dfrac{30 \text{ g } K_2Cr_2O_7}{100 \text{ g } H_2O} = \dfrac{y \text{ g } K_2Cr_2O_7}{300 \text{ g } H_2O}$

$\qquad\qquad$ **x = 210 g at 90°C** $\qquad\qquad\qquad\qquad\qquad\qquad\qquad$ **y = 90 g at 50°C**

Mass Precipitated = 210 g – 90 g $\qquad\qquad\qquad$ **Mass Precipitated = 120 g of $K_2Cr_2O_7$**

2. 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.

$$\boxed{\textbf{Pressure} \uparrow \qquad \textbf{Gas Solute Solubility} \uparrow}$$

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

Example: Most commericial carbonated beverages (soda) are sold under pressurized containers

3. **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).

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Agitation ↑   **Solid Solute Solubility ↑**
**Size of Solute ↑**   **Rate of Solubility ↓**
"Like Dissolves Like"

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Assignment
13.3 ~~**pg. 477 #1 to 10 and 12**~~;
Worksheet: Solubility Calculations

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Worksheet: Solubility of Ionic Compounds and Solubility Calculations

1. Predict if the following compounds are soluble in water.

a. barium nitrate	**b.** ammonium carbonate	**c.** lead (II) chloride	**d.** sodium chlorate
e. iron (II) bromide	**f.** mercury (I) perchlorate	**g.** potassium phosphate	**h.** magnesium sulfite
i. nickel (II) sulfate	**j.** aluminum chloride	**k.** calcium acetate	**l.** copper (II) sulfide

2. Use the solubility graph below to answer the following questions.
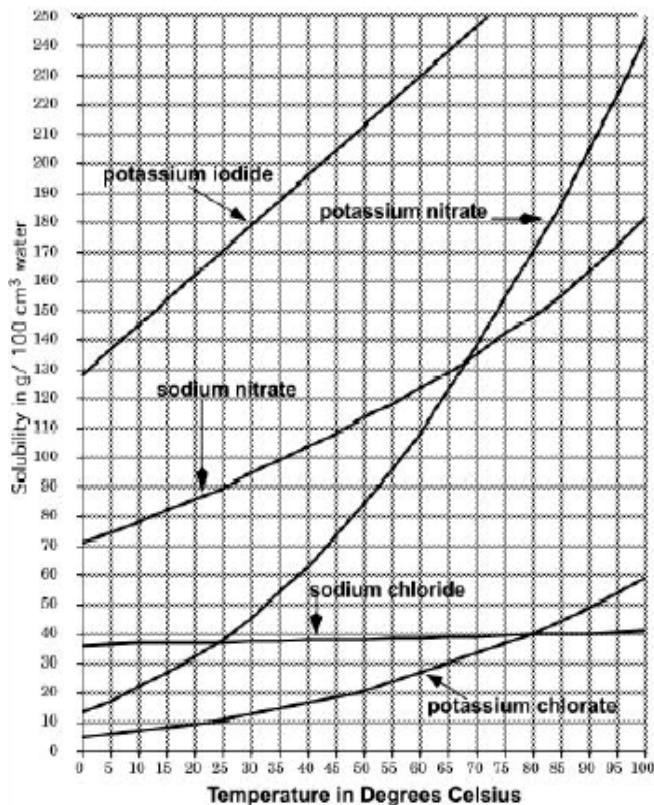
a. Which substance is most soluble at 60ºC?

b. Which two substances have the same solubility at 80ºC?

c. Which substance's solubility changes the most from 0ºC to 100ºC?

d. Which substance's solubility changes the least from 0ºC to 100ºC?

e. What is the solubility of KNO_3 at 90ºC?

f. At what temperature does KI have a solubility of 150 g / 100 cm^3 water?

g. You have a solution of $NaNO_3$ containing 140 g at 65ºC. Is the solution saturated, unsaturated, or supersaturated?

h. You have a 50 mL solution of $KClO_3$ containing 4 g at 65ºC. How many additional grams of solute must be added to it, to make the solution saturated?

i. A solution of KI at 70ºC contains 200 g of dissolved solute in 100 mL water. The solution is allowed to cool. At what new temperature would crystals begin to start forming ?

3. Calculate the volume of water needed to completely dissolve 75 g of KNO_3 at 60ºC (solubility of KNO_3 at 60°C is 108 g / 100 g H_2O).

4. Find the maximum mass of $Pb(NO_3)_2$ that can be dissolved into 125 g of water at 50ºC (solubility of lead (II) nitrate at 50°C is 85 g / 100 g H_2O).

5. What volume of water is required to completely dissolve 95 g of $Pb(NO_3)_2$ at 50ºC?

6. Determine the maximum mass of NH_4Cl that can be dissolved into 175 g of water at 70ºC (solubility of ammonium chloride at 70°C is 60 g / 100 g H_2O).

7. What is the maximum mass of Cs_2SO_4 that can be dissolved into 500 g of water at 20ºC? (solubility of Cs_2SO_4 at 20°C is 180 g / 100 g H_2O)

8. Determine the volume of water needed to completely dissolve 85.0 g of K_2CrO_4 at 40ºC. (solubility of potassium chromate at 40°C is 64 g / 100 g H_2O).

9. A 250 mL of 0.10 M potassium acetate solution is diluted to a volume of 750 mL, what will the concentration of the final solution?

10. A lab technician needs to make a 0.500 M solution of HNO_3 from concentrated 12.0 M HNO_3. If the volume of the dilute needs to be 500 mL, then how many mL of the concentrated does she need?

11. It is necessary to make a 0.500 M solution of HBr from 500.0 mL of a 3.00 M solution of HBr. What is the volume of the new solution?

12. What is the molarity of a solution which has a volume of 1500 mL if it was obtained by diluting 250.0 mL of a 6.0 M solution of H_2SO_4?

13. If an uncovered 750 mL of 0.50 M sodium bromide solution was left on a window sill and 150 mL of the solvent evaporates, what will the new concentration of the sodium bromide solution be?

14. A student added 560 mL of water to a 340 mL of a 0.5 M lithium iodide solution. Calculate the final concentration of the solution.

Answers:

1. **Soluble:** a, b, d, e, f, g, i, j, k ; **Insoluble:** c, h, l

2	**a.** KI	**b.** NaCl & $KClO_3$	**c.** KNO_3	**d.** NaCl	**e.** 205 g KNO_3 / 100 g H_2O at 90°C
	f. 13°C	**g.** supersaturated	**h.** 11 g	**i.** 43°C	

3. 69.4 g H_2O	4. 106.25 g $Pb(NO_3)_2$	5. 112 g H_2O
6. 105 g NH_4Cl	7. 900 g Cs_2SO_4	8. 133 g H_2O
9. 0.0333 M potassium acetate	10. 20.8 mL of concentrated $HNO_{3\ (aq)}$	11. 3000 mL
12. 1.0 M $H_2SO_{4\ (aq)}$	13. 0.625 M sodium bromide	14. 0.189 M $LiI_{(aq)}$

13.4: Physical Properties of Solutions

Conductivity: - refers to a substance's relative degree of electrical charges and their ability to move about.
             - the more movable charged particles a substance has, the higher its conductivity.
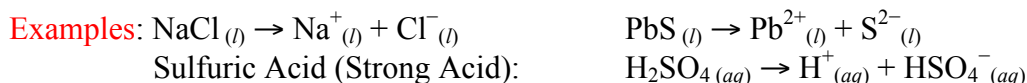
Electrolytes: - ions which have <u>dissolved completely</u> in water and allow the conduction of electricity.

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

    i. All ionic compounds containing Na^+, NH_4^+, NO_3^-, CH_3COO^-, ClO_3^-, or ClO_4^-.
        (except $RbClO_4$, $CsClO_4$, $AgCH_3COO$, $Hg_2(CH_3COO)_2$)

        Examples: $NaCl_{(s)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)}$ $(NH_4)_2CO_{3\,(s)} \rightarrow 2\,NH_4^+_{(aq)} + CO_3^{2-}_{(aq)}$

    ii. Strong Acids: $HClO_{4\,(aq)}$, $HI_{(aq)}$, $HBr_{(aq)}$, $HCl_{(aq)}$, $H_2SO_{4\,(aq)}$, and $HNO_{3\,(aq)}$
    iii. Strong Bases: $NH_4OH_{(aq)}$, $LiOH_{(aq)}$, $NaOH_{(aq)}$, $KOH_{(aq)}$, $Ca(OH)_{2\,(aq)}$, $Sr(OH)_{2\,(aq)}$, $Ba(OH)_{2\,(aq)}$
    iv. Molten Ionic Compounds: - when ionic compounds melt, they break up into ions.

        Examples: $NaCl_{(l)} \rightarrow Na^+_{(l)} + Cl^-_{(l)}$ $PbS_{(l)} \rightarrow Pb^{2+}_{(l)} + S^{2-}_{(l)}$
                      Sulfuric Acid (Strong Acid): $H_2SO_{4\,(aq)} \rightarrow H^+_{(aq)} + HSO_4^-_{(aq)}$

Note: Tap water conducts electricity eventhough pure $H_2O_{(l)}$ is overall neutral. This is because tap water is not pure but contain small amounts of dissolved ions (like CO_3^{2-}, Ca^{2+}, Mg^{2+}, Na^+, and K^+). When a person contact even pure water, the salt and ions on the skin will dissolved. Thus, making water conductive. Therefore, it is never a good idea to place electrical devices near water.
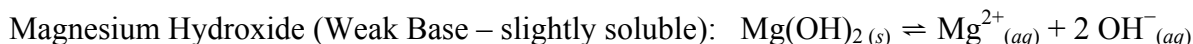
b. Weak Electrolytes: - ionic compounds that **dissociate partially** into their ions and **conduct electricity poorly**. (Check Solubility Table in the next section.)

    i. Some Ionic Compounds: $AgCl_{(s)}$, $PbCl_{2\,(s)}$, $Hg_2Cl_{2\,(s)}$, $HgCl_{2\,(s)}$, and $CuCl_{(s)}$
    ii. Weak Acids: $HF_{(aq)}$, $HCH_3COO_{(aq)}$, $H_2SO_{3\,(aq)}$, and other acids.
    iii. Weak Bases: $Mg(OH)_{2\,(aq)}$, $Al(OH)_{3\,(aq)}$, $NH_{3\,(aq)}$, 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.

    Examples: Silver (I) Chloride (Slightly Soluble Ionic Compound): $AgCl_{(s)} \rightleftharpoons Ag^+_{(aq)} + Cl^-_{(aq)}$

               Hydrofluoric Acid (Weak Acid): $HF_{(aq)} \rightleftharpoons H^+_{(aq)} + F^-_{(aq)}$

               Magnesium Hydroxide (Weak Base – slightly soluble): $Mg(OH)_{2\,(s)} \rightleftharpoons Mg^{2+}_{(aq)} + 2\,OH^-_{(aq)}$

(Right) A diagram showing a saturated calcium fluoride solution. The Ca^{2+} and F^- ions recombined to form the precipitate at the bottom of the container, a process know as crystallization. However, solid calcium fluoride breaks up into ions and are dissolved back into the solution at the same rate as the crystallization. Hence, it appears if the precipitate stays the same.
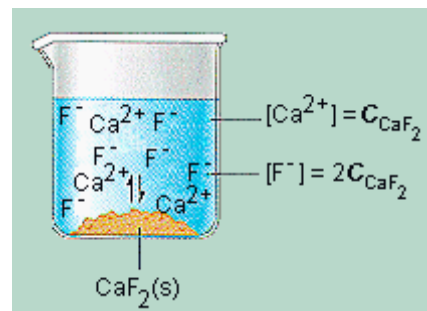
c. **Non-Electrolytes**: - soluble molecular compounds and some ionic compounds that **do not dissociate** in the solvent. **Non-electrolytes do not conduct any electricity at all**.

Examples: $CH_3OH_{(aq)}$, $C_{12}H_{22}O_{11\ (aq)}$ and $C_6H_{12}O_{6\ (aq)}$

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

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

$$[\text{solute}] \uparrow \quad P_{soln} \downarrow$$

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

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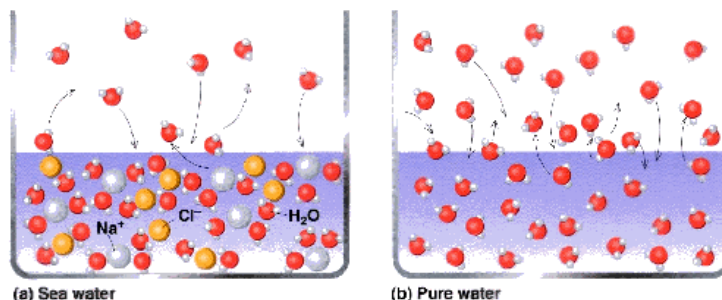
Colligative Properties: - changes in physically properties of a pure substance as it is mixed with a solute.
- these include freezing-point depression, boiling-point elevation and osmotic pressure.

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).

Note: because **dissolved ionic solutes** can generate two ions or more, at the same molal concentration (mol of solute per kg of solvent), they **have higher boiling point elevations**.

Examples: At 5.00 mol of the following solutes in 1 kg of water,

$$C_{12}H_{22}O_{11\ (s)} \rightarrow C_{12}H_{22}O_{11\ (aq)} \qquad \text{New } T_b = 102.6°C$$
$$NaCl_{(s)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)} \qquad \text{New } T_b = 105.2°C$$
$$MgCl_{2\ (s)} \rightarrow Mg^{2+}_{(aq)} + 2\ Cl^-_{(aq)} \quad \text{New } T_b = 107.8°C$$
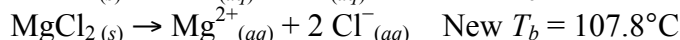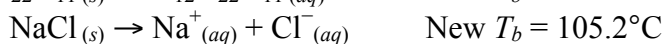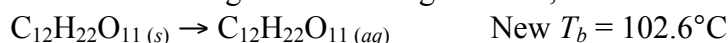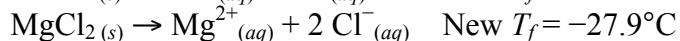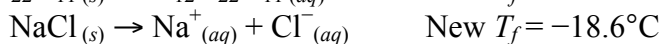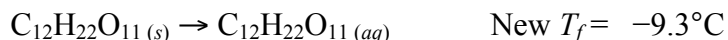
Freezing-Point Depression: - the presence of solute decreases the vapour pressure of the solvent. Since ice has a higher vapour pressure than that of water in the solution, the freezing point has effectively been lowered. (In order for ice to form, vapour pressure of the ice has to be lowered than that of water).

- because extra particles are in the solvent, they interrupt the crystallization process as the solvent moves from the liquid phase to solid phase. Hence, the temperature has to lower (all the particles have to move much slower), in order for the crystals to form.

Note: due to the fact that **dissolved ionic solutes** can generate two ions or more, at the same molal concentration (mol of solute per kg of solvent), they **have lower boiling point depressions**.
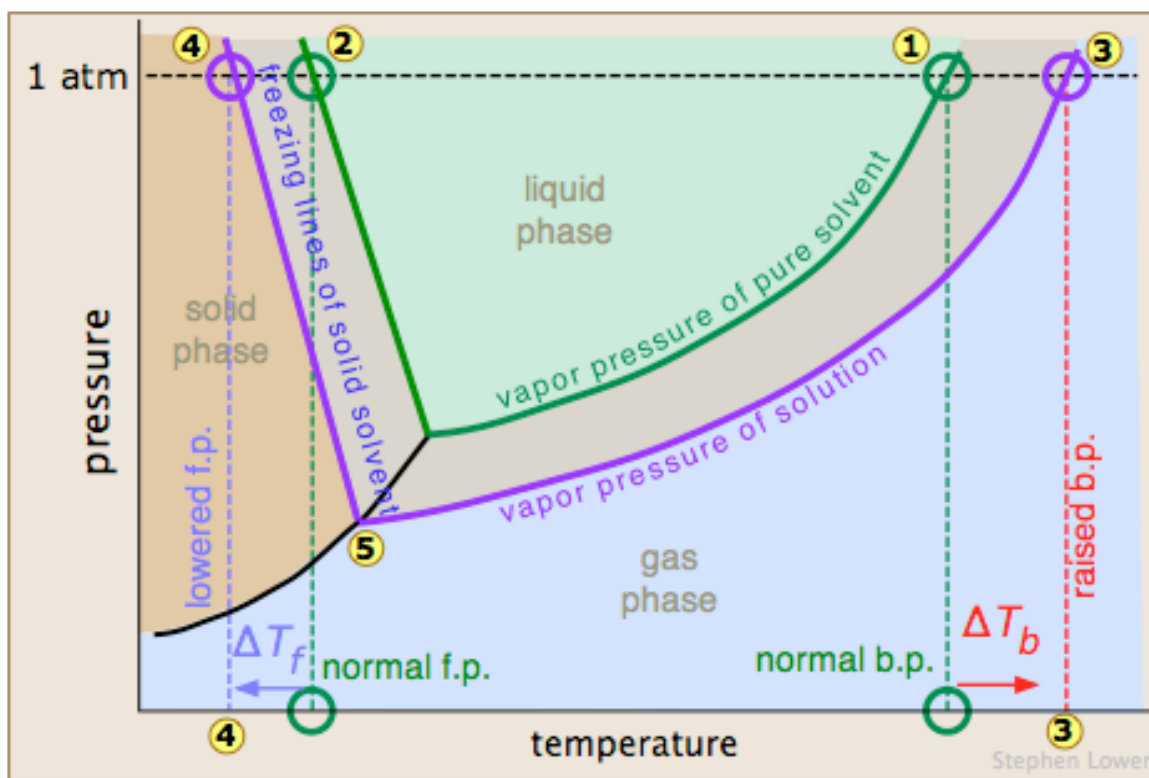
Examples: At 5.00 mol of the following solutes in 1 kg of water,

$$C_{12}H_{22}O_{11\ (s)} \rightarrow C_{12}H_{22}O_{11\ (aq)} \qquad \text{New } T_f = \ -9.3°C$$
$$NaCl_{(s)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)} \qquad \text{New } T_f = -18.6°C$$
$$MgCl_{2\ (s)} \rightarrow Mg^{2+}_{(aq)} + 2\ Cl^-_{(aq)} \quad \text{New } T_f = -27.9°C$$

Boiling Point Elevation and Freezing Point Depression Animation at:
http://www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/flashfiles/propOfSoln/colligative.html

Effects on the Phase Digram when a Solute is added to a Pure Substance

Grime: - a combination of dirt and grease, which is also non-polar.
- normally, non-polar solvents like turpentine or trichloroethane is used to remove grime.

Surfactant: - a substance that concentrates at the interface between two phases, either the solid-liquid, liquid-liquid, or gas-liquid phase.

Emulsion: - a colloid-sized droplets suspended in a liquid in which they would ordinary be insoluble unless stabilized by an **emulsifying agent**.

Soap: - a molecule containing non-polar end that can attract grime and a polar end that can be washed away with water.
- soap is a surfactant and an emulsifying agent because it concentrates on a boundary layer between two phases and it allows oil and water to form an emulsion.

(Left) A Soap Molecule which has a polar end and a non-polar end.

(Right) As soap molecules encounter grease, their nonpolar ends attract to the grime. This action literally traps the grease particle. The polar ends of the soap molecules faces the outside, and are easily attracted to the water, which can be washed away.
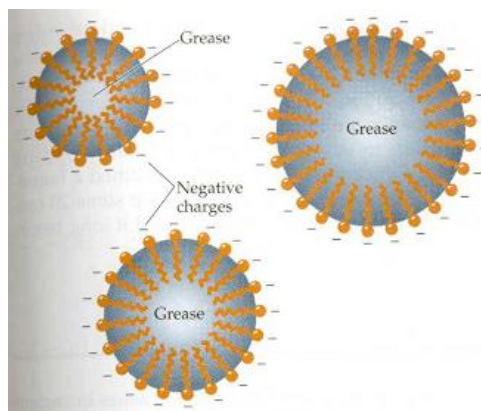
Making of Soap: - soap are made by adding sodium hydroxide to animal fat to produce three fatty acid soap molecules along with glycerol.
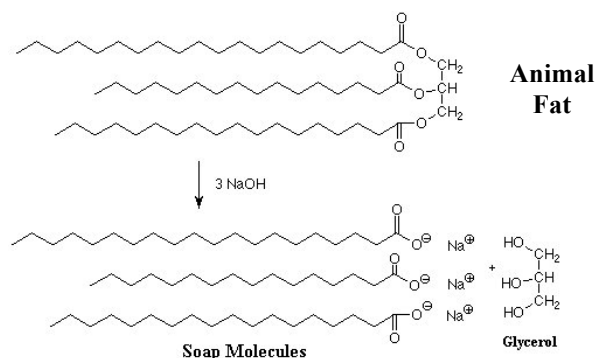

Detergents: - synthetic soap such as sodium lauryl sulfate that is used in toothpastes and sodium dodecyl benzenesulfonate found in dishwashing fluids.
- they are biodegradable and cheap to produce.

Hard Water: - water that contains calcium and magnesium ions (Ca^{2+} and Mg^{2+}).
- these ions can easily combined with other anion such as carbonate ion (CO_3^{2-}) to form insoluble substances that clog pipes.
- they also reduce the effectiveness of soap and detergent as they attract the polar end of the soap and detergent molecules.

Water Softener: - a device that removes the magnesium and calcium ion utilizing ion-exchange with resin.
- the resin is originally attracted to sodium ion (Na^+) from the salt (brine) solution (from water softener salt).
- as hard water goes through the resin, the calcium and magnesium ions displace the sodium ions of the resin.
- at specific time at night, the resin is set to "regenerated" with fresh brine solution, calcium and magnesium ions on the resin are now washed off by new sodium ions.
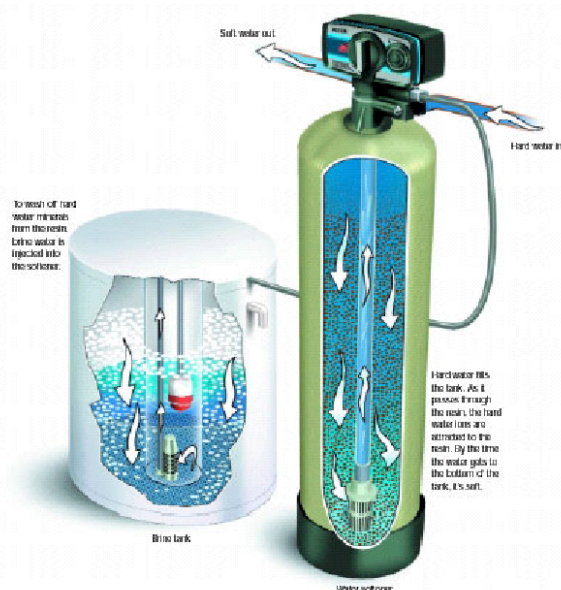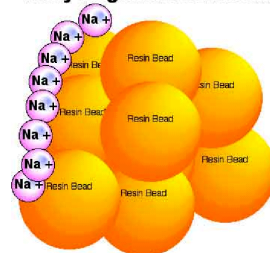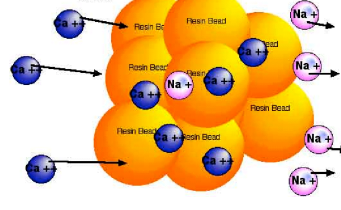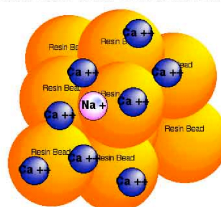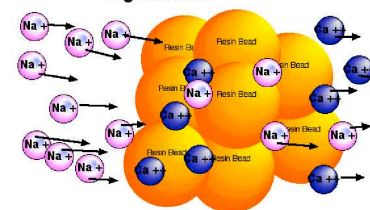

Assignment

13.4 pg. 486 #1 to 15;
Chapter 13 Review (optional) pg. 488−491 #1, 5 to 8, 12, 13, 17, 18, 20 to 24, 27 to 34, 37 to 51 (odd), 52 to 54, 67 to 72