Unit 5 – Liquids, Solids, and Solutions

Objectives:

Intermolecular Forces (IMFs)

- Describe hydrogen bonding, London Dispersion Forces, and Dipole-Dipole Forces
- Identify the type of IMF based on the substance
- Relate the IMF to the physical properties of the substance

Properties of Water

- Describe the structure of a water molecule
- Explain how the physical properties of water are determined by the structure of water

Solutions

- Distinguish solutions from suspensions and colloids
- Identify and explain physical and chemical factors that affect solubility
- Define concentration terms of molarity, molality, and percent by volume.
- Perform calculations using molarity (M) and molality (*m*)

Ions in Aqueous Solutions and Colligative Properties

- Describe the dissociation of ionic compounds and the ionization of some molecular compounds when they dissolve in water
- Define electrolyte
- Distinguish between strong and weak electrolytes
- Write a net ionic equation using the solubility rules
- Perform calculations with different solution concentrations such as molarity, mass percent, molality, and mole fraction.
- Discuss effects of temperature, pressure, and structure on solubility.

- Use a solubility chart to determine if a solution is unsaturated, saturated, or supersaturated.
- Define colligative properties and use to calculate boiling point elevation, freezing point depression, and vapor pressure lowering
- Use colligative properties to determine the molar mass of a solute

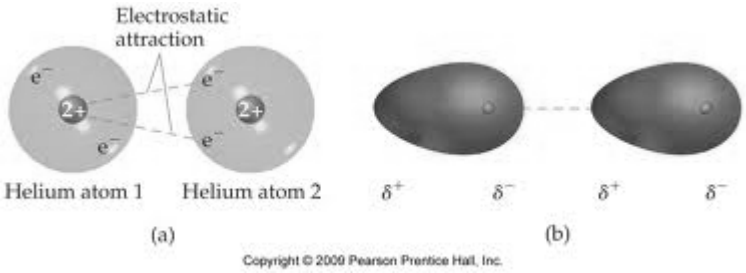
Vocabulary

- Capillary action
- Surface tension
- Soluble
- Insoluble
- Precipitate
- Precipitation
- Aqueous
- Solubility
- Solution
- Electrolyte
- Nonelectrolyte
- Solute
- Solvent
- Solvation
- Effervescence
- Henry's Law
- Hydration
- Immiscible
- Miscible
- Saturated
- Unsaturated
- Supersaturated
- Solution equilibrium
- Concentration
- Molality
- Molarity
- Dissociation
- Ionization
- Net ionic equation
- Spectator ion
- Strong electrolyte
- Weak electrolyte
- Colligative property
- Boiling point elevation
- Freezing point depression
- Molal boiling point constant (k_b)
- Molal freezing point constant (k_f)
- Vapor pressure lowering
- Hydrogen Bonding
- London Dispersion Forces
- Dipole-Dipole Forces
- Van der Waals Forces

Lesson 5.1: Don't Flip your Lid and Intermolecular Forces

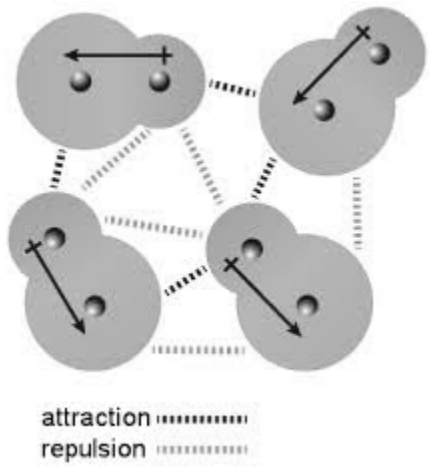
Comparing Intermolecular Forces

The forces that hold one molecule to another molecule are referred to as *intermolecular forces (IMFs)*. IntrAmolecular forces are the bonds (ionic or covalent) that hold atoms together within a molecule. These forces arise from unequal distributions of the electrons in the molecule and the electrostatic attraction between oppositely charged portions of molecules.



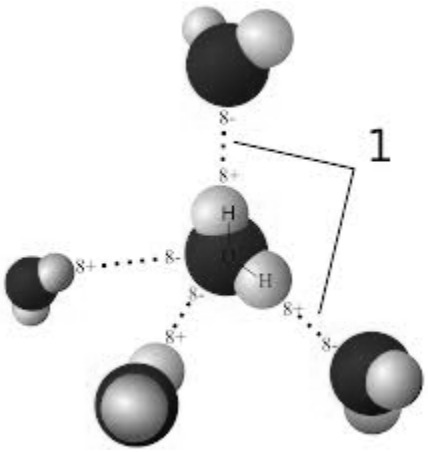
Van der Waals forces are a function of the number of electrons in a given molecular and how tightly those electrons are held. The term Van der Waals force is a generic umbrella term used to describe IMFs. Let us assume that the molecule involved is nonpolar. A good example would be O_2. Pretend that the molecule is alone in the universe. If that were the case, the electrons in the molecule would be perfectly symmetrical. However, the molecule is not really alone. It is surrounded by other molecules that are constantly colliding with it. When these collisions occur, the electron cloud around the molecule is distorted. This produces a momentary induced dipole within the molecule. The amount of distortion of the electron

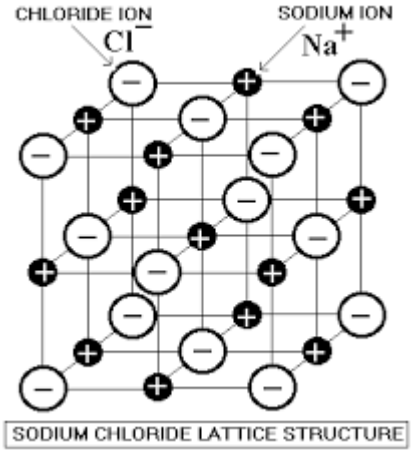
could is referred to as polariziability. Since the molecule now has a positive side and a negative side, it can be attracted to the other molecules. This attractive force is called a London dispersion force. Since all molecules have electrons, all molecules have London forces. These forces range from 5 – 40 kJ/mol.



Some molecules are naturally polar; therefore, in addition to dispersion forces, they can also have a permanent dipole which attracts other polar molecules (either induced or permanent). This is called a dipole-dipole force.

A third type of intermolecular force is *hydrogen bonding*. When hydrogen is covalently bonded to a small electronegative atom like nitrogen, oxygen or fluorine, the electron cloud on the hydrogen is very distorted and pulled toward the electronegative atom. Since hydrogen has no inner core electrons, the positive nuclear charge is somewhat exposed. This sets up the potential for a reasonably strong attraction between this hydrogen and an electronegative atom in another molecule. Hydrogen bonds are significant in determining such factors as the high boiling point of water, solubility of acetone in water, and the shape and structure of proteins and DNA.





Ionic bonds are formed between charged particles. Ionic compounds do not form molecules, but rather are help together in a crystal by electrostatic attraction. These electrostatic attractions are strong and are omnidirectional.

PURPOSE

In this laboratory activity you will determine the relative melting points of four substances, paraffin, NaCl (table salt), $C_6H_{12}O_6$ (glucose) and iodine (I_2). Paraffin is a medium size nonpolar molecule, NaCl is an ionic compound. Glucose is a medium size polar molecule and Iodine is a large nonpolar molecule.

MATERIALS

Bunsen burner
Ring stand
Tin Can Lid

Solid samples of the following: sodium chloride (NaCl), glucose ($C_6H_{12}O_6$), paraffin, iodine (I_2).

SAFETY ALERT!!! Wear your goggles. Do not inhale the iodine vapors. Allow the can lid to cool before dispersal

PROCEDURE

1. In the space marked HYPOTHESIS on your student answer page, write a statement predicting in which order the salt, paraffin, iodine and glucose will melt.
2. Set up your ring stand. Place the can lid on the ring stand
3. Transfer a small amount of each of the four samples into one of the four areas in a groove of the outer side of the can lid. You only need a very small sample – just enough to be able to see the compound. The teacher will transfer the crystal of iodine.
4. Turn the hotplate on and observe. You will need to record the relative order of melting. As soon as three of the compounds melt, turn off the hot plate.
5. Allow the can lid to cool before using tongs to dispose of it in a trash can or as your teacher directs.
6. Answer the CONCLUSIONS questions on your student answer page.

HYPOTHESIS:

OBSERVATIONS: Record the relative order of melting for the four compounds.

ANALYSIS:
From what you know about intermolecular forces, explain the relative order of the melting points.

CONCLUSIONS QUESTIONS:

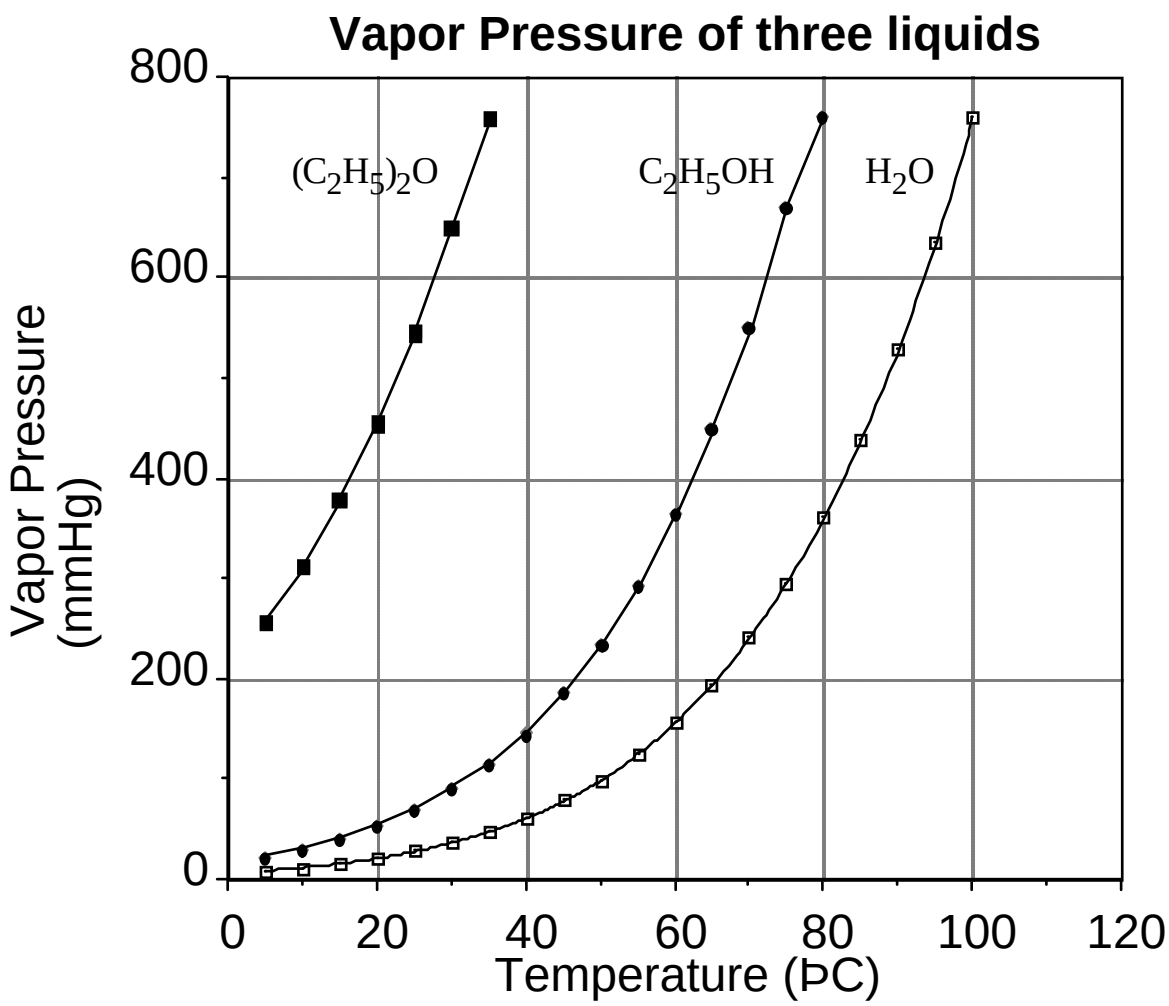
1. Glucose, $C_6H_2O_6$, and sucrose, $C_{12}O_{22}O_{11}$, are both sugars with similar polarities. Which do you think would have the higher melting point? Justify your answer.

2. Hydrogen sulfide is a gas at room temperature, while water is a liquid, yet hydrogen sulfide has more electrons that water. Explain this anomaly.

3. Consider the halogens at room temperature and 1 atmosphere of pressure. Why are fluorine and chlorine gases at room temperature, while bromine is a liquid and iodine is a solid?

There are three basic intermolecular forces that hold covalent molecules together as either solids or liquids. To change the phase of a substance, the IMFs must be overcome. The amount of energy required for the phase change is dependent on the substance and the strength (type) of the IMF. Typically substances that contain hydrogen bonding will require more energy for the phase change to occur. Substances with LDFs will typically require the least amount of energy.

Vapor pressure is the pressure due to particles of a substance in the vapor phase above its liquid in a closed container at a given temperature. The weaker the forces holding the liquid together, the higher the vapor pressure of the liquid will be.

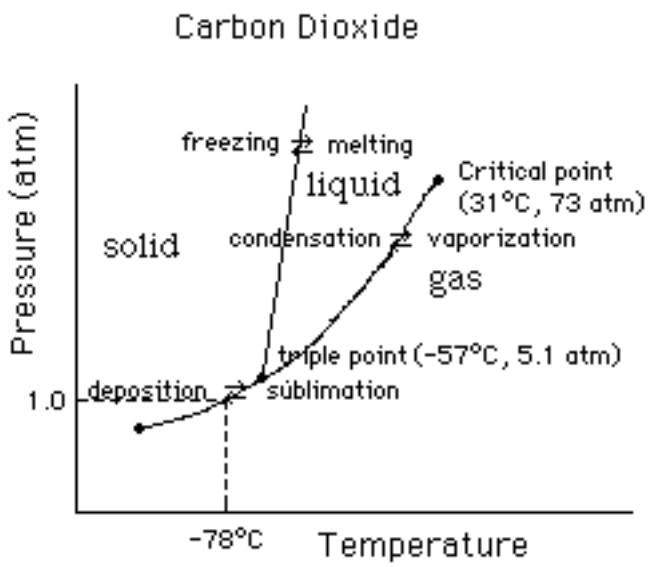


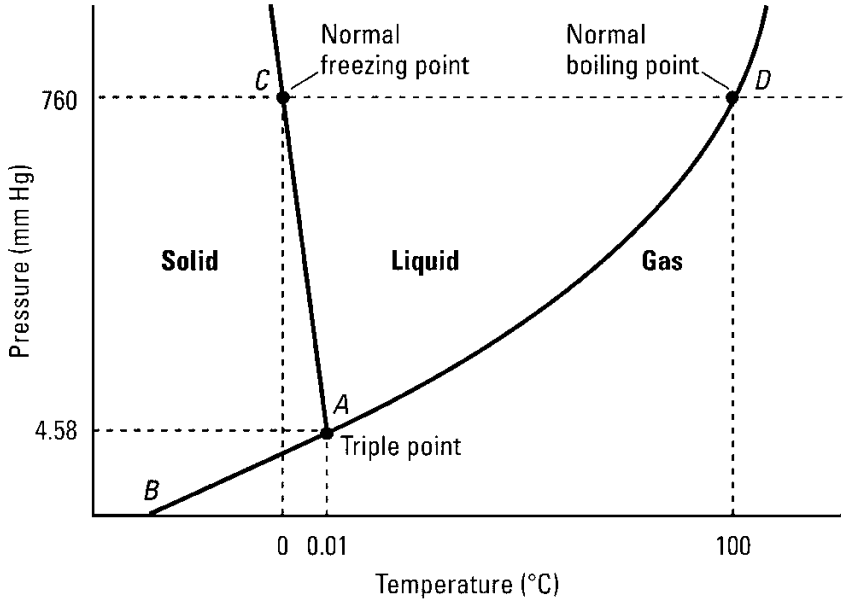
Boiling point is the temperature at which the vapor pressure of a liquid equals the atmospheric pressure. The normal boiling point is the temperature at which the vapor pressure of the liquid equals 1 atmosphere.

The molar heat of vaporization (ΔH_{vap}) is the amount of heat required to change one mole of a pure liquid into one mole of a pure gas. The molar heat of fusion (ΔH_{fus}) is the amount of heat required to change one mole of a pure substance from a solid to a liquid. What generalization can be made between IMFs and boiling temperature and boiling temperature?

Phase Diagrams

Phase diagrams are a way of graphically representing the relationship of the three states of a pure substance. Below is an example of a phase diagram for carbon dioxide.

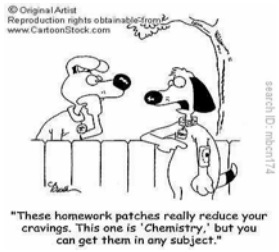




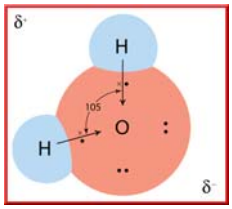
Above is the phase diagram for water. One very important point to notice is the negative slope of the solid liquid interface. This shows that the density of ice is actually less than the density of water - a very unusual property of water that makes much of life possible.

The phase of substance is not only dependent on the IMF present but also on the physical conditions, temperature and pressure.

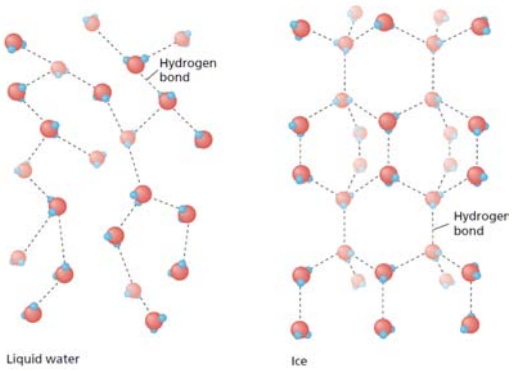
Solutions- Day 1



Water

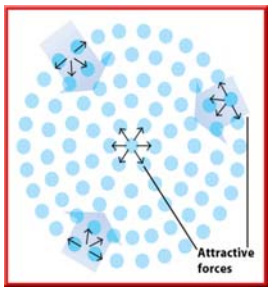


- A group of water molecules will orient at the molecular scale because of electrical forces.
- This effect is especially great at low temperatures.
- The water molecules move less rapidly, and they are able to be drawn closer together by dipole-dipole attractions.
- The volume of the water decreases because the molecules pull together.



Surface Tension

- A water molecule forms a drop because of **surface tension**, which is the force needed to overcome intermolecular forces and break through the surface of a liquid or spread the liquid out.
- The higher the surface tension of a liquid is, the more resistant the liquid is to having its surface broken.
- The net inward force makes the surface of the drop contract



Capillary Action

- the rising of liquids in narrow tubes, also known as capillarity
- Capillarity results from the competition between the IMFs between the molecules of liquid and the attractive forces between the liquid and the tube that contains it.

Like Dissolves Like

- Polar dissolves polar
- Non polar dissolves non polar

Definitions

- A solution is a homogeneous mixture
- A solute is dissolved in a solvent.
  - solute is the substance being dissolved
  - solvent is the liquid in which the solute is dissolved
  - an *aqueous* solution has water as solvent

Dissolution of Solid Solute (Solvation)

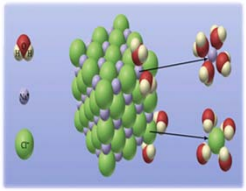
What are the driving forces which cause solutes to dissolve to form solutions?

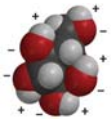
1. Ionic solutes dissolve by dissociation into their ions.

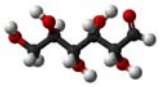
Video of dissociation:

http://www.youtube.com/watch?v=EBfGcTAJF4o

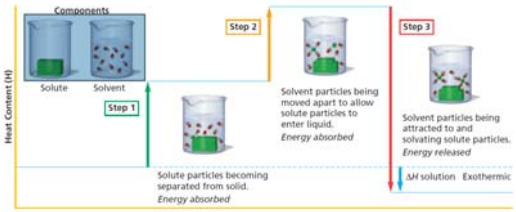
2. Covalent solutes dissolve by H-bonding to water or by Van der Waals (dispersion forces)







What has to happen for dissolving to take place



The net amount of heat energy absorbed or released when a specific amount of solute dissolves in a solvent is the ***heat of solution.***

Solution and Concentration

- 3 ways of expressing concentration

–**Molarity(M):** = moles solute / Liter solution

–Mass percent = (mass solute / mass of solution) x 100%

–Molality* (m) = moles solute / kg solvent

*** Note that molality is the only concentration unit in which denominator contains only solvent information rather than solution.**

- Chemists never apply the terms *strong* and *weak* to solution concentrations always use concentrated or dilute.

% Concentration

- $\% \text{ (m/m)} = \frac{\text{mass solute}}{\text{mass solution}} \times 100$
- $\% \text{ (m/v)} = \frac{\text{mass solute}}{\text{volume solution}} \times 100$
- $\% \text{ (v/v)} = \frac{\text{volume solute}}{\text{volume solution}} \times 100$

Example Problems

1) How many grams of beryllium chloride are needed to make 125 mL of a 0.0500 M beryllium chloride solution?

2) How many grams of beryllium chloride would you need to add to 125 mL of water to make a 0.0500 molal solution?

3) What will the volume (in L) of a 0.500 M solution be if it contains 25.0 grams of calcium hydroxide?

4) How many grams of ammonia are present in 5.00 L of a 0.0500 M solution?

5) If 3.50 g of $CoCl_2$ is dissolved in 100. mL solvent, what is concentration of the solution in % mass (m/m)? (Assuming the density of the solvent is 1.00 g/mL)

6) Explain how to make at least 1.00 liter of a 1.25 molal ammonium hydroxide solution.

7) Explain how to make 1.00 liters of a 1.25 molar ammonium hydroxide solution.

Dilution

-When a solution is diluted, solvent is added to lower its concentration.
-The amount of solute remains constant before and after the dilution:
-Therefore moles BEFORE = moles AFTER

$$M_1V_1 = M_2V_2$$

M= concentration in molarity
v= volume

Example Problems

1) If 25.0 mL of water is added to 125 mL of a 0.150 M NaOH solution, what will the molarity of the diluted solution be?

2) Given 345 mL of a 1.50 M NaCl solution. If the solution is boiled so its volume decreases 95.0 mL, what will the molarity of the solution be?

3) How much water would have to be added to 500. L of a 2.40 M KCl solution to make it a 1.00 M solution?

Solutions Day 2

How to describe a solution

- A saturated solution contains the maximum amount of solute per amount of the solution under the given conditions - no more solute is able to dissolve.
  - A saturated solution represents an equilibrium: the rate of dissolving is equal to the rate of crystallization. The salt (solute) continues to dissolve, but crystallizes at the same rate so that there "appears" to be nothing happening.
- A unsaturated solution contains less dissolved solute dissolved than the maximum that could be dissolved.
- A supersaturated solution more dissolved solute than the usual maximum amount and are unstable.

Factors Affecting Solubility Liquids and Gases

1. Nature of Solute / Solvent. - Like dissolves like (IMF) need to have polar with polar or non polar with non polar

2. Temperature -

i) Solids/Liquids- Solubility increases with Temperature (usually)
Increase K.E. increases motion and collision between solute / solvent.

ii) gas - Solubility decreases with Temperature
Increase K.E. result in gas escaping to atmosphere.

3. Pressure Factor -

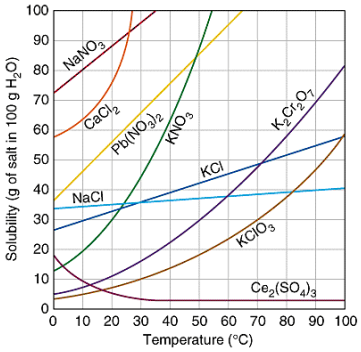
i) Solids/Liquids - Very little effect
Solids and Liquids are already close together, extra pressure will not increase solubility.

ii) gas - Solubility increases with Pressure.
Increase pressure squeezes gas solute into solvent.

Solubilities of Ionic Solids vs Temperature

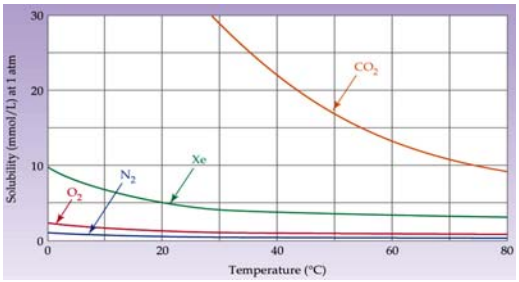
Solubilities of several ionic solid as a function of temperature. MOST salts have greater solubility in hot water. Heat of solution of most ionic solids is positive endothermic.

A few salts have negative heat of solution, (exothermic process) and they become less soluble with increasing temperature.



Temperature & the Solubility of Gases

The solubility of gases DECREASES at higher temperatures



Henry's Law
The effect of partial pressure on solubility of gases

At pressure of few atmosphere or less, solubility of gas solute follows Henry Law which states that the amount of solute gas dissolved in solution is directly proportional to the amount of pressure above the solution.

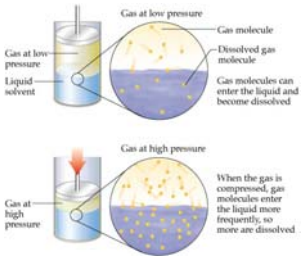
Example of Henry's Law:

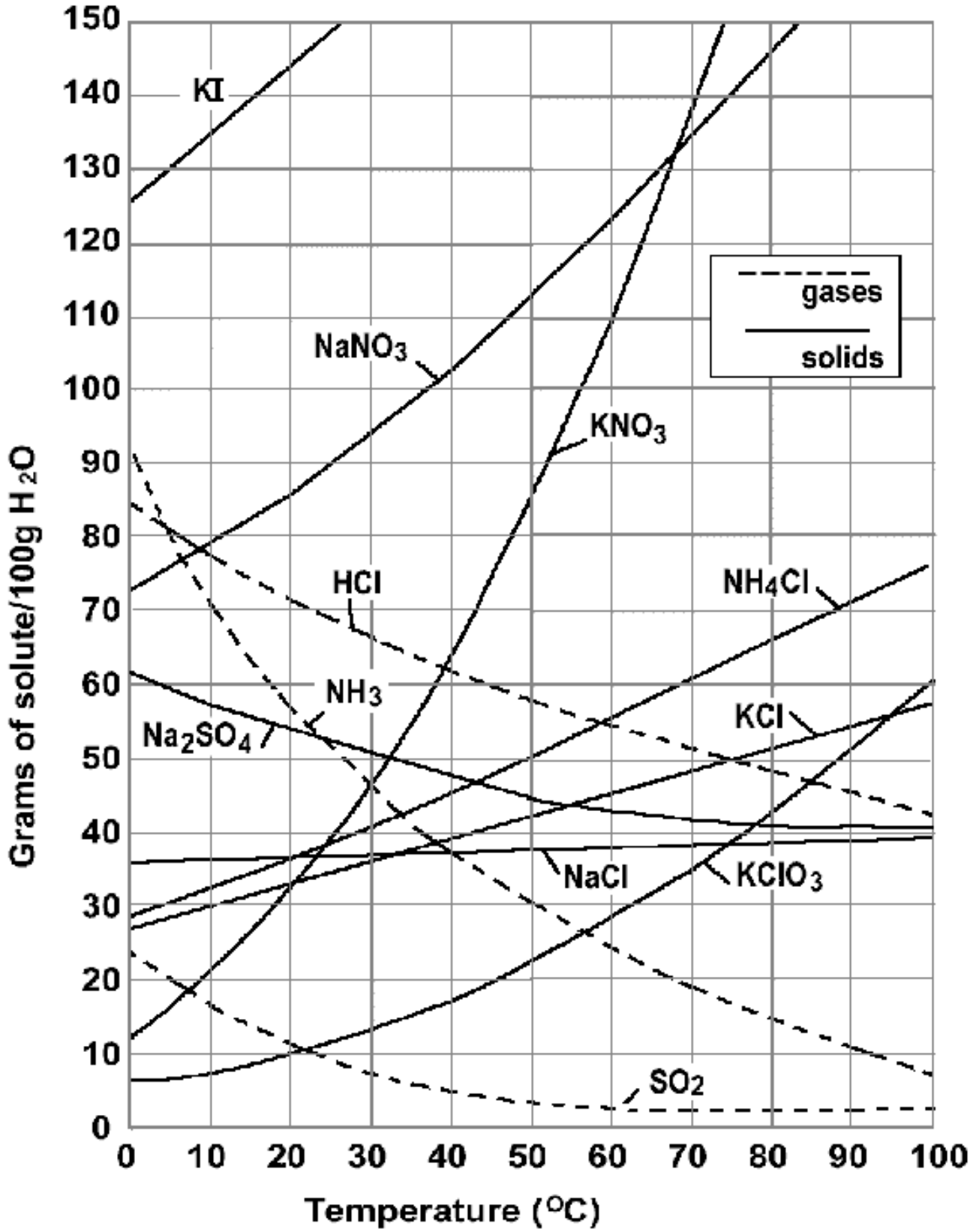
-Soft drinks contain "carbonated water" – water with dissolved carbon dioxide gas.

-The drinks are bottled with a CO2 pressure greater than 1 atm.

-When the bottle is opened, the pressure of CO2 decreases and the solubility of CO2 also decreases, according to Henry's Law.

-Therefore, bubbles of CO2 escape from solution.




150
140
130
120
110
100
90
80
70
60
50
40
30
20
10
0
Grams of solute/100g H2O
0
10
20
30
40
50
60
70
80
90
100
Temperature (°C)
KI
NaNO3
KNO3
HCl
NH4Cl
NH3
Na2SO4
KCl
NaCl
KClO3
SO2
gases
solids

Solubility Curves

Based on the solubility below, decide whether each of the following is A: unsaturated, B: saturated, C: supersaturated, or whether D: not enough information is given. * assume it's dissolved *

1) 50 g KCl in 100 g of water at 90°C. ____

2) 50 g KCl in 100 g of water at 60°C. ____

3) 50 g KNO_3 in 100 g of water at 60°C. ____

4) 50 g KNO_3 in 25 g of water at 60°C. ____

5) 65 g KNO_3 in 50 g of water at 70°C. ____

6) 25 g KNO_3 in 100 g of water. ____

7) 25 g NaCl in 100 g of water. ____

8) 40 g of KCl in 100 g of water at 20°C. ____

9) How many grams of KCl can dissolve in 100.0 g of water at 65°C? _____

10) What temperature would be required to get 85 g of KNO_3 to dissolve in 100.0 g of water?

<u>SHOW ALL WORK FOR THE FOLLOWING</u>

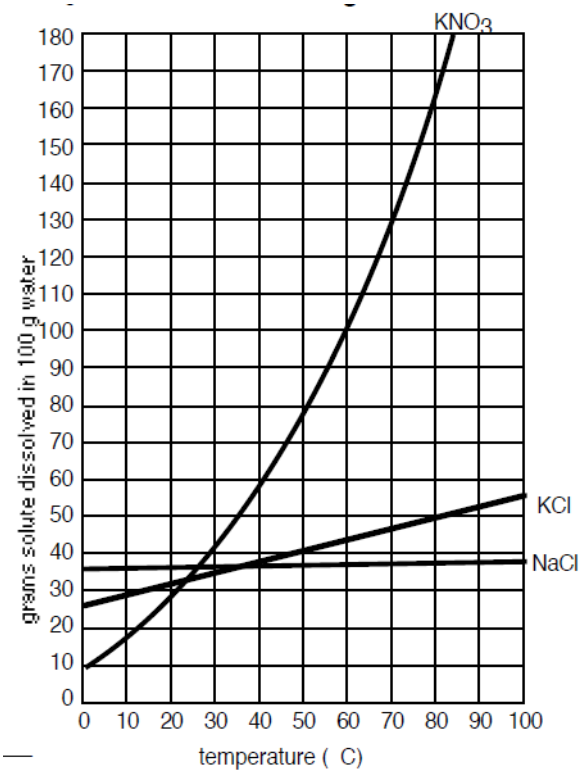
11) How many grams of KNO_3 can be dissolved in 50.0g of water at 50.0°C? _____

12) What mass of KCl can be dissolved in 200.0 g of water at 15.0°C? _____

13) How much KNO_3 can be dissolved in 14.3 g of water at 69.0°C? _____

14) How many grams of water will it take to dissolve 28.0 g NaCl at 60.0°C? _____

15) How much water is needed to dissolve 46.6 g of KNO3 at 52°C? _____



16) What temperature would be required to get 51.0 g of KCl to dissolve in 156 g of water?

17) What is the % KCl in a solution that is saturated at 61°C? _____

18) What temperature is required to make a 50.0% KNO_3 solution? _____

19) What temperature is required to make a 63.0% KNO_3 solution? _____

20) Based on what you've learned in class about soda & fish, do gases behave the same as or different than solids when it comes to solubility & temperature? What do you think a solubility graph of gases would look like?

Ans (iro+5): A, A, A, B, B, C, C, C, D, 18, 19, 25, 31, 39, 39, 46, 54, 57, 60, 60, 68, 76, 82
Units (iro+1): %, g, g, g, g, g, g, °C, °C, °C , °C, °C

Molarity

Determine the concentration (molarity) for each of the solutions:

a) 3.0 mol sugar dissolved in 2.0 L of solution. ___

b) 0.40 mol NaCl dis. in 10.0 L of soln. ____

c) 0.030 mol KNO_3 dis. in 50.0 mL of soln. _____

d) 350 g KNO_3 dis. in 5.0 L of soln. _____

e) 6.45 g of Na_2SO_4 dis in 250 mL of soln. ___

f) 465 mg KF of dis. in 0.054 L of soln. _____

2. How many moles of sugar are needed to make 2.5 L of 1.4 M sugar solution? Ans: _____

3. How many moles of NaBr are needed to make 150 mL of 3.0 M NaBr solution? Ans: _____

4. How many grams of $NaNO_2$ are needed to make 3.5 L of 0.50 M NaNO2 solution? Ans: _____

5. How many grams of K_2CO_3 are needed to make 300.0 mL of 1.25 M K_2CO_3 solution? Ans: _____

6. What volume of 0.25 M sugar solution can be made using 4.0 moles sugar? Ans: _____

7. How many mL of 2.50 M Na_3PO_4 solution can be made using 1.8 g of Na_3PO_4? Ans: _____

8. 65.0 mL of K_3PO_4 solution are evaporated, and 1.54 g of solid K_3PO_4 are recovered. What was the molarity of the original solution? Ans:_____

Ans (IRO +1): 0.040 0.112 0.15 0.18 0.45 0.60 0.69 1.5 2.85 3.5 4.4 16 51.8 120

Units: (IRO + 1): moles, moles, g, g, g, L, mL, M, M, M, M, M, M, M

Dilutions

1. Determine the concentrations for each of the following mixtures:

a) equal volumes of 3.0 M KCl & water: ___

b) equal volumes of 3.0M KCl & 7.0 M KCl: ____

c) one vol. of 8.0 M KCl & one vol. water: ___

d) one vol. of 6.0 M KCl & two vol's water: ____

e) one vol. water & two vol's of 6.0M KCl: ___

f) one vol. of 5.0 M KCl & 4 vol's of water: ____

g) one vol. of 2.5 M KCl & 9 vol's water: ____

h) one vol. of 2.5 M KCl & 99 vol's water: ____

2. To make orange juice from frozen concentrate, one usually mixes the can of concentrate with three cans of water. This dilutes the concentrate to _____ (what fraction?) its original concentration.

3. Use the dilution equation to determine the concentrations of the following mixtures...

a) 45 L of 3.6 M KCl & 71 L of water:

Ans: _____

b) 215 mL of 2.8 M KCl & 47 mL water:

Ans: _____

c) 45 mL of 3.6 M KCl & 71 mL of 6.2 M KCl:

Ans: _____

d) 83 mL of 2.0 M KCl & 25 mL of water:

Ans: _____

e) 38 mL of 6.0 M KCl dil. to a tot vol of 100 mL:

Ans: _____

4. To what total volume must 26.0 mL of 4.80 M KCl be diluted to reduce its concentration to...

a) ... 2.10 M

Ans: _____

b) ... 0.480 M

Ans: _____

Ans (IRO+1): 0.025 0.25 1/4 1.0 1.4 1.5 1.5 2.0 2.0 2.3 2.3 3.6 4.0 4.0 5.0 5.2 59.4 125 260.Units (IRO): M M M M M M M M M M M M M M M M M mL mL

(Warning: one of the questions on this page is impossible... When you find it, explain why it's impossible!)5. What volume of water must be added to 35 mL of 2.6 M KCl to reduce its concentration to...

a) ... 1.2 M

Ans: _____

b) ... 0.26 M

Ans: _____

6. What volume of 2.5 M KCl must be added to 37 mL of 6.0 M KCl to make the total concentration of:

a) ... 1.5 M

Ans: _____

b) ... 4.2 M

Ans: _____

7. What volume of 2.5 M KCl must be added to 37 mL of water to make the total concentration 1.8M?

Ans: _____

8. You mix 32 mL of 4.5 M KCl, 56 mL of 6.2 M KCl and some water, and the total concentration comes out to be 1.7 M. How much water must have been added?

Ans: _____

9. Sketch a volumetric flask and explain precisely how you would use a 500.0 mL volumetric flask to make some 1.500 M $NaNO_3$ solution. (You have available some 2.000 M $NaNO_3$ solution and whatever other lab equipment you need) **How much 2.000 M solution is needed?**

10. You need to make up some 5.0 M KCl solution but all you have is 125 mL of 3.0 M KCl.

Explain what to do to make up the 5.0 M solution. How much 5.0 M KCl will you get? Show calculations: (hint-calculate how much water to evaporate)

Ans (IRO+1): 39 41 75 95 172 201 315 375 Units(IRO) mL mL mL mL mL mL mL mL mL

Electrolytes, Precipitates and Balancing Net Ionic Equations

Electrolytes: substances that form ions in aqueous solutions (dissolved in water). Positive and negative ions carry current (conduct electricity) in an aqueous solution. Usually these are soluble ionic salts, strong acids (completely dissociate), and strong bases (completely dissociate). Substances that will not dissociate in solution are insoluble and nonelectrolytes (cannot conduct electricity). The following table is a guide to whether a solution is soluble or not.

SOLUBILITY RULES: *memorize!!!*

1. **Most alkali metal salts AND** NH_4^+ salts ARE ***soluble***
2. Cl^-, Br^-, I^- are ***soluble***, *except for Ag^+, Hg_2^{+2}, Pb^{+2}
3. F^- are ***soluble,*** *except for IIA metals
4. NO_3^-, ClO_3^-, ClO_4^-, and CH_3COO^- are ***soluble***
5. SO_4^{-2} are ***soluble,*** except for Ca^{2+}, Sr^{+2}, Ba^{+2}, Ag^+, Pb^{+2}, Hg_2^{2+}
6. CO_3^{-2}, PO_4^{-3}, $C_2O_4^{-2}$, CrO_4^{-2}, S^{-2}, OH^-, and O^{-2} are ***INSOLUBLE*** (rule 1 takes priority!)*

It can be assumed that ionic cmpds. that dissolve in water are strong electrolytes and are therefore soluble.

*hydroxides of Ca^{2+}, Sr^{+2}, Ba^{+2} are soluble

Heavy metal

Identify the following compounds as either electrolytes or nonelectrolytes. Also list the ions they will form in solution if applicable. Assume all are aqueous.

1) LiCl

2) $(NH_4)_2S$

3) Ammonium Chlorate

4) $Ca(OH)_2$

5) $Mg(OH)_2$

6) Silver Chloride

Precipitates: An insoluble solid that emerges from an aqueous solution. The emergence of the insoluble solid from solution is called precipitation.

Precipitates can form when two soluble salts react in solution to form one or more insoluble products. The insoluble product separates from the liquid and is called a precipitate.

Precipitates can also form when the temperature of a solution is lowered. The lower temperature lower temperature reduces the solubility of the salt resulting in the formation of a precipitate.

Spectator Ions: Ions in a solution that do not participate in a chemical reaction.

Balancing Net Ionic Equations: typically these are double replacement/displacement reaction types. Net ionic equations are used to show which chemical species are actively reacting and eliminates spectator ions. The reactions we will be considering will take place in $H_2O(l)$.

Below is an example of how these are done.

EXAMPLE: $KCl_{(aq)} + Pb(NO_3)_{2(aq)} \rightarrow$

1. **a. Take <u>only one</u> of the first cation(s) and match it with <u>one</u> of the second anion(s). (Write the cation first)**

   b. Take <u>only one</u> of the second cation(s) and match it with <u>one</u> of the first anion(s). (Write the cation first)

   $$KCl_{(aq)} + Pb(NO_3)_{2(aq)} \rightarrow KNO_3 + PbCl$$

2. Correct the formulas of the products based on the charges of the ions. (Note the subscript change on Cl)

   $$KCl_{(aq)} + Pb(NO_3)_{2(aq)} \rightarrow KNO_3 + PbCl_2$$

3. Balance the equation

   $$\underline{2}\ KCl_{(aq)} + Pb(NO_3)_{2(aq)} \rightarrow \underline{2}\ KNO_3 + PbCl_2$$

4. Consult the solubility rules and assign the correct state symbol. This should agree with any observations concerning the formation of a precipitate which gets the symbol (*s*). If water is formed, water is a molecule; it does not ionize to any significant extent. It is given *(l)*. Note the state variables on the product side. Why does KNO_3 get (aq) and $PbCl_2$ get (s).

   $$2\ KCl\ (aq) + Pb(NO_3)_2\ (aq) \rightarrow 2\ KNO_3\ (aq) + PbCl_2\ (s)$$

5. Write the Total Ionic Equation. All compounds that are aqueous (*aq*) break up into individual cations and anions

   $$2\ K^+(aq) + 2Cl^-(aq) + Pb^{2+}(aq) + 2NO_3^-\ (aq) \rightarrow 2\ K^+\ (aq) + 2\ NO_3^-\ (aq) + PbCl_2\ (s)$$

6. Eliminate spectator ions. Spectator ions are in the same form on each side of the equation arrow.

   $$2\ \cancel{K^+_{(aq)}} + 2Cl^-_{(aq)} + Pb^{2+}_{(aq)} + 2\cancel{NO_3^-}_{(aq)} \rightarrow 2\ \cancel{K^+_{(aq)}} + 2\ \cancel{NO_3^-}_{(aq)} + PbCl_{2(s)}$$

7. Write the Net Ionic Equation. The convention is to write the cation first followed by the anion on the "reactants" side. Don't forget that chemical equations are written using the lowest common coefficients (including net ionic equations). If all ions cancel each other out then it is NR.

   $$Pb^{2+}_{(aq)} + 2Cl^-_{(aq)} \rightarrow + PbCl_{2(s)}$$

Class work Net Ionic Equations. Write and Balance the net ionic equation for the following reactions.

1.. K_3PO_4 (aq) + $Al(NO_3)_3$ (aq) →

2. BeI_2 (aq) + Cu_2SO_4 (aq) →

3. $Ni(NO_3)_3$ (aq) + KBr (aq) →

4. cobalt(III) bromide + potassium sulfide →

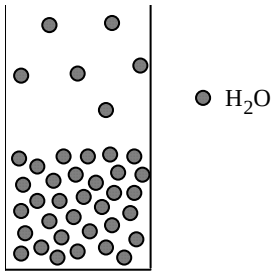
5. barium nitrate + ammonium phosphate →

Colligative Properties

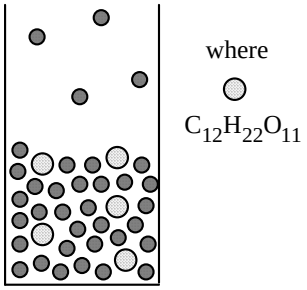
Colligative properties are properties that depend on the <u>number of molecules or ions</u> of solute present, and not on what the particles are (as long as they are not volatile). Properties of solutions that are colligative properties,

1) Vapor Pressure lowering
2) Boiling Point elevation
3) Freezing Point depression

The following diagram shows why a solute will affect the colligative properties



Before addition of the solute: The vapor above a pure liquid is shown on the right. The equilibrium vapor pressure is a result of the presence of molecules in the vapor phase above the molecules in the liquid phase. The figure below shows the effect on the vapor pressure of the solution upon addition of a nonvolatile solute.



After addition of the solute: The addition of a nonvolatile solute decreases the vapor pressure due to the liquid in the vapor phase. This occurs because the presence of solute particles on or near the surface decreases the ability of the solvent particles to escape into the vapor phase. Because proportionally fewer solvent particles are on the surface, the vapor pressure drops.

The above leads to Raoults law:

$P_{solution} = \chi_{solvent} \cdot P_{solvent}$

Where;

$P_{solution}$ is the vapor pressure of the solvent above the solution at the particular temperature

$\chi_{solvent}$ is the mole fraction of the solvent in the solution $\chi_{solvent} = \frac{mol_{solvent}}{mol_{solvent} + mol_{solute}}$

$P_{solvent}$ is the vapor pressure of the pure solvent at the particular temperature.

The other colligative properties are freezing point depression and boiling point elevation.

$$\Delta T_f = ik_f m \qquad \text{or} \qquad \Delta T_b = ik_b m$$

where:

ΔT_f or ΔT_b is the change in freezing point or the change in boiling point.

k_f or k_b are the freezing point and boiling point constants (solvent dependent)

m is the molality of the particles in the solution $\left(\frac{\text{mol solute}}{\text{kg solvent}}\right)$

i the Van't Hoff factor and is the number of ions of solute in solution

example: if the substance is a nonelectrolyte - i = 1

NaCl – i = 2

Practice Problems

1. Calculate the expected vapor pressure of at 25°C for a solution prepared by dissolving 97.4 grams of common table sugar (MM 342 g/mol) in 453 mL of water. Vapor pressure of water at 25°C is 23.8 mmHg (Ans 23.49 mmHg).

2. Glucose is often used as a form of nourishment to hospital patients through intravenous feeding. Calculate the new boiling point and freezing point for a 0.70 m solution of glucose, $C_6H_{12}O_6$.

3. Calcium chloride is a common drying agent and is very exothermic when dissolved in water. If 47.50 grams of calcium chloride is dissolved in 395 mL of water, calculate the boiling point and freezing point of the solution that is formed.

4. Moth balls are often composed of pure naphthalene, $C_{10}H_8$. The normal freezing point of naphthalene is 80.2°C. When 5.25 grams of an unknown nonvolatile substance was mixed with 35.00 grams of naphthalene, the freezing point was found to be 78.8°C. Calculate the molecular weight of the unknown substance. The k_f for naphthalene is 6.9°C/*m*.

5. Which solution freezes at a lower temperature, 0.20 *m* sodium sulfate or 0.20 *m* magnesium sulfate? Explain.

Unsaturated, Saturated, Supersaturated Lab

Follow the procedure below, and record all observations. Ignore the letters in ( ).

1) Obtain a clean, unscratched test tube. Add 2.0 mL of water (A). Then add 4.00 g of $NaC_2H_3O_2$ ("sodium acetate," which we will abbreviate SA from here on), but don't shake yet (B). Use a perm. pen to mark the top of the undissolved SA level on the side of test tube. Stopper and mix for 3 sec, tilt to get any undissolved crystals off the sides, and note any changes (C).

Obs:_____

2) Re-mark the top of the undissolved solute level. Mix for another 5 sec and observe the changes, including feeling the test tube (D).

Obs: _____

3) Repeat step #2 until no more change occurs (E).

Obs: _____

4) REMOVE STOPPER! Heat test tube (by placing in hot water) for 90 sec (while waiting, weigh out another 0.10 g of SA for step #5). Remove test tube from heat, stopper & mix for 10 sec (F).

Obs: _____

5) While still hot, add the 0.10 additional grams of SA, stopper & mix for 10 sec (G)

Obs: _____

6) Add an additional 4.00 g of SA, stopper & mix for 10 sec (H).

Obs: _____

7) **REMOVE STOPPER!** Heat for 2 min (while waiting, weigh out 1.00 g of SA for step #8), then remove from heat, stopper and mix for 10 sec (I). Obs: _____

8) Add the 1.00 g of SA & mix (J). Obs: _____

9) Reheat until all crystals have dissolved (K), stopper and mix, and then cool in cold water for 50-60 sec (L), (If recrystallization occurs during cooling, reheat to redissolve it, then re-cool it.)

* Then add 1 crystal SA & observe (M). Obs: _____

10) (bonus) Reheat until all crystals have dissolved and then an additional 30 sec (N), make sure your test tube rim is ULTRA-clean, and cool in water for 50-60 sec (O). Place a crystal or two on a clean petri dish lid. Then, carefully, drop-by-drop, pour your solution out onto the crystal. Observe what happens (P). Advice: Don't allow the growing pillar to come too close to the mouth of the test tube.
Obs: _____

QUESTIONS:

1. Consider each of the points throughout the procedure indicated by the letters (A-P) and decide whether at each particular moment, the test tube contained a solution that was unsat, sat. or supersat.

   Briefly justify your answers. The first one is done for you.

   A unsat _it's pure water…no solute____

   B ____ _____

   C ____ _____

   D ____ _____

   E ____ _____

   F ____ _____

   G ____ _____

   H ____ _____

   I ____ _____

   J ____ _____

   K ____ _____

   L ____ _____

   M ____ _____

   N ____ _____

   O ____ _____

   P ____ _____

2. If you were handed a solution and told to determine whether it was unsaturated, saturated or supersaturated, explain what you would do and what you would expect to see for each of three possible cases.

   unsaturated: _____

   saturated: _____

   supersaturated: _____

3. A solution has some undissolved crystals sitting on the bottom. Could it be...

   unsaturated? Y / N Explain: _____

   saturated? Y / N Explain: _____

   supersat.? Y / N Explain: _____

4. Use the solubility curves in the packet to explain **precisely**, step-by-step, how you would go about making a **supersaturated** KNO_3 solution. State precisely how many grams of water, how many grams of KNO3 and what temperatures you would use.

Preparing Solutions Lab:

Background: In chemistry, spectrophotometry is the quantitative measurement of the reflection or transmission properties of a material as a function of wavelength. Spectrophotometry involves the use of a spectrophotometer. A spectrophotometer is a photometer that can measure intensity in absorbance and transmittance. In lab you will prepare solutions of Copper (II) Sulfate Pentahydrate and measure their absorbance at a wavelength of 635 nm.

Fill in the data table below from the board:

Concentration (M)	**Absorbance at 635 nm**
0.019 M	
0.038 M	
0.075 M	
0.150 M	

1) Use the data table above to generate a linear regression equation in the format y=mx +b where concentration is x and absorbance is y.

2) Use the molarity equation, a 50.0 mL volumetric flask, distilled water, and a balance to prepare 50.0 mL of a 0.130 M $CuSO_4$ solution. Describe below how to make this solution.

3) Use the dilution equation to determine how to prepare a 50.0 mL of a 0.060 M solution using your 0.130 M solution, a graduated cylinder, and a volumetric flask. Describe how to make this solution below.

4) Use a 50.0 mL volumetric flask, distilled water, and a balance to prepare 50.0 mL of a 0.030 M $CuSO_4$ solution. Describe below how to make this solution.

5) Use your linear regression equation to predict the absorbance of your prepared solutions.

6) Use a spectrophotometer to measure the actual absorbance of each of your prepared solutions at 635 nm. Record this data in the table below.

Concentration (M)	**Predicted Absorbance**	**Actual Absorbance**	**Percent Error**

UNKNOWN SAMPLE ABSORBANCE _____

PREDICTED MOLARITY OF UNKNOWN _____

Conclusion Questions:

1) Molarity is considered a temperature dependent concentration unit while molality is considered a temperature independent concentration unit. Why is this true?

2) What mass (in grams) of calcium hydroxide is needed to make 500.0 mL of a 2.40 M solution?

3) To what volume should 200. mL of an 8.00 M NaOH solution be diluted so that the final molarity of the solution is 1.50 M?