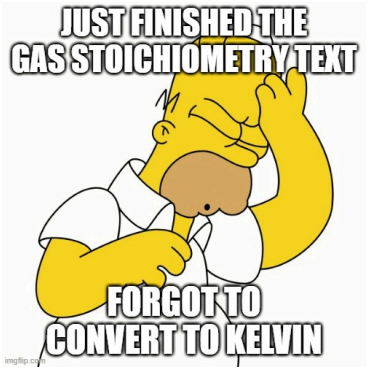
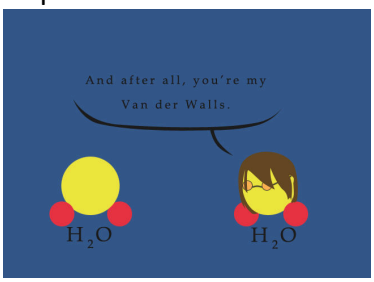
AP Matter Class Packet Unit 3

AP Learning Objectives:

- Explain the relationship between the chemical structures of molecules and the relative strength of their intermolecular forces when: a. The molecules are of the same chemical species. b. The molecules are of two different chemical species. (3.1)
- Explain the relationship among the macroscopic properties of a substance, the particulate-level structure of the substance, and the interactions between these particles.(3.2)
- Represent the differences between solid, liquid, and gas phases using a particulate level model.(3.3)
- Explain the relationship between the macroscopic properties of a sample of gas or mixture of gases using the ideal gas law. (3.4)
- Explain the relationship between the motion of particles and the macroscopic properties of gases with: a. The kinetic molecular theory (KMT). b. A particulate model. c. A graphical representation.(3.5)
- Explain the relationship among non-ideal behaviors of gases, interparticle forces, and/or volumes.(3.6)
- Calculate the number of solute particles, volume, or molarity of solutions.(3.7)
- Using particulate models for mixtures: a. Represent interactions between components. b. Represent concentrations of components. (3.8)
- Explain the relationship between the solubility of ionic and molecular compounds in aqueous and nonaqueous solvents, and the intermolecular interactions between particles. (3.9)
- Explain the relationship between the solubility of ionic and molecular compounds in aqueous and nonaqueous solvents, and the intermolecular interactions between particles. (3.10)
- Explain the relationship between a region of the electromagnetic spectrum and the types of molecular or electronic transitions associated with that region. (3.11)
- Explain the properties of an absorbed or emitted photon in relationship to an electronic transition in an atom or molecule.(3.12)
- Explain the amount of light absorbed by a solution of molecules or ions in relationship to the concentration, path length, and molar absorptivity(3.13)
- Represent changes in matter with a balanced chemical or net ionic equation: a. For physical changes. b. For given information about the identity of the reactants and/or product. c. For ions in a given chemical reaction. (4.2)
- Explain the relationship between macroscopic characteristics and bond interactions for: a. Chemical processes. b. Physical processes.(4.4)











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Ms. Drury

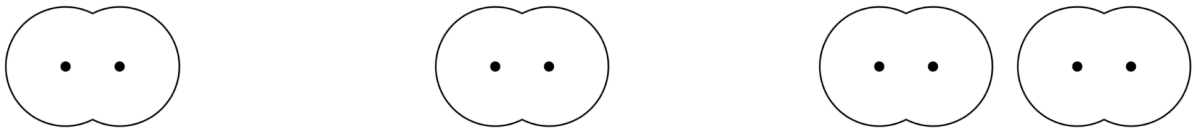
Bonds and IMF Lab

Guiding Question: Can we rank the intermolecular forces of attraction based on strength?

Procedure and Data:

1. London Dispersion Forces: Watch the clip showing how the electrons of a molecule can influence another.
   a. Complete the models that mimic the short clip you watched of a molecule of H_2 when it approaches another molecule of H_2. Add electrons and charges.

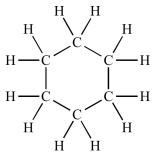
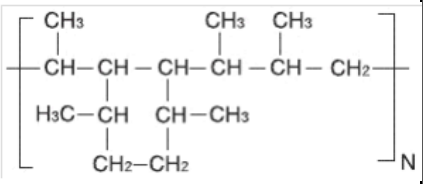
Far apart **Close together**



   b. Is the movement of the electrons temporary or permanent?

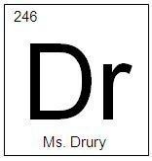
   c. Is this a strong or weak attraction?

   d. Observe the substances and complete the table and following questions.

	methane (CH_4)	cyclohexane (C_6H_{12})	petroleum resin ($C_{20}H_{30}O_2$)
Phase			
Boiling Point	-161.6°C	80.8°C	200.0°C
Lewis diagram			
Total Number of e-			

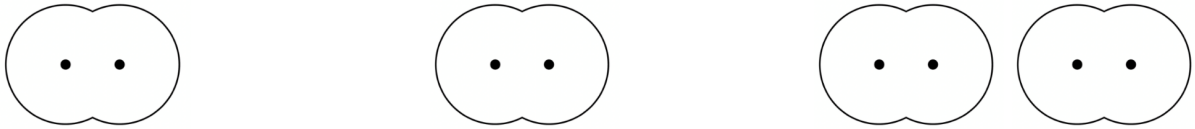
   e. State what type of bonds the three substances have in common (ionic, metallic, polar covalent or nonpolar covalent).

   f. Using the short clip and observations, try to explain why each of these molecules are in a different phase in terms of “intermolecular forces of attraction” and “polarizable electrons”.



2. <u>Dipole Dipole Forces:</u> Watch the clip showing how the electrons of a molecule can influence another.
   a. Complete the models that mimic the clip you watched of a molecule of HCl when it approaches another molecule of HCl. Add electrons and charges.

Far apart **Close together**



   b. Is the movement of the electrons temporary or permanent?

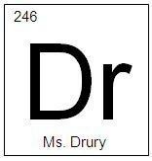
   c. Is this a stronger or weaker attraction than London Dispersion Forces?

   d. Observe the substances and complete the table and following questions.

	Br_2	ICl
Phase		
Lewis diagram		
Total Number of electrons		
Electronegativity differences		
Type of bond		

   e. ICl and Br_2 have similar numbers of electrons but are different phases at room temperature. What other properties might account for these differences?

   f. Using the short clip and observations, try to explain why each of these molecules are in different phases at room temperature in terms of "intermolecular forces of attraction" and "electronegativity differences."



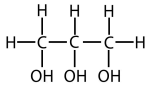
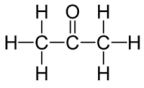
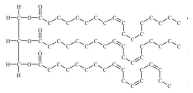
3. Hydrogen Bonds: Watch the clip showing how the electrons of a molecule can influence another. Then, play with some molecular models of water.
   a. Draw models that mimic the clip you watched of a molecule of H_2O when it approaches another molecule of H_2O. Add charges.

Far apart **Close together**

   b. Is the movement of the electrons temporary or permanent?

   c. Is this attraction more similar to London Dispersion Forces and Dipole Dipole Forces?

   d. Place a few drops of each liquid on the construction paper in order. Record observations about the rate of evaporation of each substance and complete the table and questions.

	Glycerol ($C_3H_8O_3$)	H_2O	Acetone (CH_3COCH_3)	Oil
Evaporation Observations				
Boiling Point	290°C		56°C	~300°C
Lewis diagram *with lines of symmetry*				
Total Number of e^-s				A lot
Electronegativity differences				0
Polar or Nonpolar Molecule			Nonpolar	
Type of force(s)			LDF	

   e. Compare the rate of evaporation and boiling points of the following and use the terms "intermolecular forces of attraction" and either "number of polarizable electrons", "symmetry", or "electronegativity differences" to explain the observations.
      i. Glycerol and water

      ii. Water and acetone

iii. Acetone and oil

f. Water and oil might seem like a discrepancy but it is not. Why might oil's LDF be so much stronger than water's strong hydrogen bonds?

g. Based on the "like dissolves like" rule, which substances should mix well together?

4. Dipole Induced Dipole Forces: Watch the clip showing how the electrons of a molecule can influence another.
   a. Mix the following solutions and record observations and complete the table.

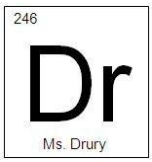
	Observations
Glycerol and water	
Oil and water	
Acetone and water	

   b. Explain your observations using the types of intermolecular forces in the mixture.
      i. Glycerol and water

      ii. Oil and water

   c. Acetone and water might seem like they shouldn't mix. However, water's strong hydrogen bonds influence the highly polarizable electrons of acetone. Draw a model to show how these two substances may interact and mix including electrons and charges.

Far apart **Close together**



5. Ion-Dipole: Watch the clip showing how the charges of a molecule can influence a salt.
   a. Observe the models constructed that represent ionic solids. Complete the table:

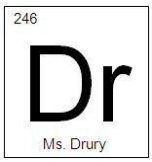
	Macroscopic Drawing of the Models	Lewis diagram of one particle
NaCl		
KI		

   b. Compare the two salts:
      i. Using the atomic radii and charges of the ions present in NaCl and KI, determine which salt has a stronger lattice energy.

      ii. Which salt will have a higher melting point?

   c. Using the NaCl sample, separate the ions as if the sample was heated. Place water molecules around the ions. Draw a model to represent how the water orientates itself around the ions.

   d. Which salt, NaCl or KI, will have a stronger attraction to water particles and dissolve better?

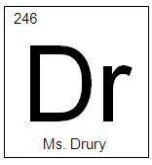


6. Network Solids:
   a. Observe the samples and models that represent network solids. Complete the table:

	Observations of the sample	Macroscopic Drawing of the Models	Lewis diagram of one particle of C	Melting point
$C_{diamond}$				3550°C
$C_{graphite}$				3652°C
SiO_2				1710°C

   b. Are melting points or boiling points generally higher values for one substance? Think about the values for water.

   c. The melting points of these are considerably higher than boiling points of other samples shown earlier in the lab. What does that imply about the strength of these network solids?



7. Metallic Bonds: Watch the clip showing how the electrons of a metal behave.
   a. Observe and test the conductivity of the samples. Complete the table:

	Observations of the sample	Observations of the models	Conductivity (Yes/No)	Type of bonds and/or forces	Melting point
Sugar $C_6H_{12}O_{6(s)}$					186°C
Sugar $C_6H_{12}O_{6(aq)}$					186°C
NaCl(s)					801°C
$NaCl_{(aq)}$					801°C
Cu(s)					1085°C

   b. Which substance never conducts in any phase?
      i. What bonds and forces does the substance have?

      ii. Why doesn't it conduct (what is it missing)?

   c. Which substance conducts only in the liquid or aqueous phases?
      i. What bonds and forces does the substance have?

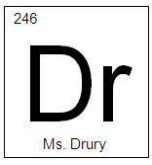
      ii. Why doesn't it conduct in the solid phase? Draw a model to describe your answer.

      iii. Why do the liquid/aqueous particles conduct? Draw a model to describe your answer.

   d. Which substance conducts in the solid phase?
      i. What bonds does it have?

      ii. Why does it conduct in the solid phase? Draw a model to describe your answer.

   e. Compare the structures of NaCl and Cu in the models. Which of the substances are malleable?



Conclusions:

1. Rank the forces and bonds in order from weakest to strongest attraction.

2. Draw 4 molecules of HF at cold temperatures and hot temperatures:

   a. What happened to the space between particles?

   b. What happened to the attractions between molecules?

   c. What happened to the bonds between atoms?

3. Summarize what the strength of each force depends on:
   a. LDF:

   b. Dipole Dipole:

   c. Dipole Induced Dipole:

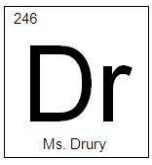
   d. Hydrogen Bonds:

   e. Ion-Dipole:

   f. Network Solids:

   g. Metallic Bonds:





IMF AP FRQs

1.

Substance	Melting Point, ºC
H_2	-259
C_3H_8	-190
HF	-92
CsI	621
LiF	870
SiC	>2,000

   a. Discuss how the trend in the melting points of the substances tabulated above can be explained in terms of the types of attractive forces and/or bonds in these substances.

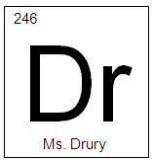
   b. For any pairs of substances that have the same kind(s) of attractive forces and/or bonds, discuss the factors that cause variations in the strengths of the forces and/or bonds.

2. Using principles of chemical bonding and/or intermolecular forces, explain each of the following.
   a. Xenon has a higher boiling point than neon has.

   b. Solid copper is an excellent conductor of electricity, but solid copper chloride is not.

   c. SiO_2 melts at a very high temperature, while CO_2 is a gas at room temperature, even though Si and C are in the same chemical family.

   d. Molecules of NF_3 are polar, but those of BF_3 are not.



3. Explain each of the following observations in terms of the electronic structure and/or bonding of the compounds involved.

a. At ordinary conditions, HF (normal boiling point = 20ºC) is a liquid, whereas HCl (normal boiling point = -114ºC) is a gas.

b. Molecules of AsF_3 are polar, whereas molecules of AsF_5 are nonpolar.

c. The N-O bonds in the NO_2^- ion are equal in length, whereas they are unequal in HNO_2.

d. For sulfur, the fluorides SF_2, SF_4, and SF_6 are known to exist, whereas for oxygen only OF_2 is known to exist.

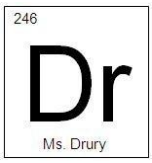
4. Account for each of the following observations about pairs of substances. In your answers, use appropriate principles of chemical bonding and/or intermolecular forces. In each part, your answer must include references to <u>both</u> substances.

   a. Even though NH_3 and CH_4 have similar molecular masses, NH_3 has a much higher normal boiling point (-33°C) than CH_4 (-164°C).

   b. At 25°C and 1.0 atm, ethane (C_2H_6) is a gas and hexane (C_6H_{14}) is a liquid.

   c. SiC melts at a much higher temperature (1,410°C) than Cl_2 (-101°C).

   d. MgO melts at a much higher temperature (2,852°C) than NaF (993°C).



5. Explain each of the following in terms of atomic and molecular structures and/or intermolecular forces.
   a. Solid K conducts an electric current, whereas solid KNO_3 does not.

   b. $SbCl_3$ has a measurable dipole moment, whereas $SbCl_5$ does not.

   c. The normal boiling point of CCl_4 is 77ºC, whereas that of CBr_4 is 190ºC.

   d. NaI(s) is soluble in water, whereas I_2(s) has a solubility of only 0.03 gram per 100 grams of water.

AP Chemistry: IMF Multiple Choice

Use the following answers for questions 1-2.

(A) A network solid with covalent bonding
(B) A molecular solid with zero dipole moment
(C) A molecular solid with hydrogen bonding
(D) An ionic solid
(E) A metallic solid

1. Solid ethyl alcohol, C_2H_5OH
2. Silicon dioxide, SiO_2

Hydrogen Halide	Normal Boiling Point, °C
HF	19
HCl	−85
HBr	−67
HI	−35

3. The liquefied hydrogen halides have the normal boiling points given above. The relatively high boiling point of HF can be correctly explained by which of the following?
(A) HF gas is more ideal. (C) HF molecules have a smaller dipole moment.
(B) HF is the strongest acid. (D) HF is much less soluble in water.
(E) HF molecules tend to form hydrogen bonds.

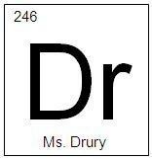
Use these answers for questions 4-5
(A) hydrogen bonding (B) hybridization (C) ionic bonding
(D) resonance (E) London dispersion forces

4. Is used to explain why iodine molecules are held together in the solid state
5. Is used to explain why the boiling point of HF is greater than the boiling point of HBr

6. The structural isomers C_2H_5OH and CH_3OCH_3 would be expected to have the same values for which of the following? (Assume ideal behavior.)
(A) Gaseous densities at the same temperature and pressure
(B) Vapor pressures at the same temperature (C) Boiling points
(D) Melting points (E) Heats of vaporization

7. CH_3CH_2OH boils at 78 °C and CH_3OCH_3 boils at − 24 °C, although both compounds have the same composition. This difference in boiling points may be attributed to a difference in…
(A) molecular mass (B) density (C) specific heat
(D) hydrogen bonding (E) heat of combustion

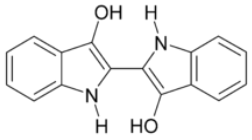
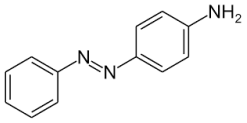
8. The boiling points of the elements helium, neon, argon, krypton, and xenon increase in that order. Which of the following statements accounts for this increase?
(A) The London (dispersion) forces increase.
(B) The hydrogen bonding increases.
(C) The dipole-dipole forces increase.
(D) The chemical reactivity increases.
(E) The number of nearest neighbors increases.



AP IMF Glitter Activity

Guiding Question: What intermolecular forces of attraction (IMF) exist between compounds?

Pre-lab: Draw the Lewis Structures below and provide the IMF between molecules of themselves.

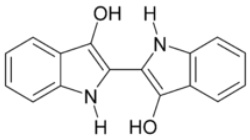
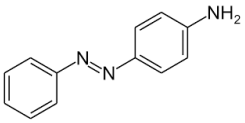
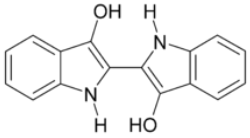
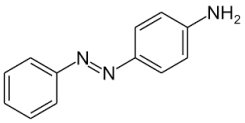
Acetone (CH_3COCH_3)	H_2O	Blue Dye 	Yellow Dye 
IMF:	IMF:	IMF:	IMF:

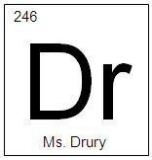
Procedure and Questions:

1. Observe the demonstration of water and glitter being stirred by a stir bar. Record observations.

2. When the teacher added soap to the mixture what happens? Record observations and explain what happened using IMFs.

3. Observe the 4 containers of mixtures presented by your teacher. Record observations below.

	Macroscopic Observations: Draw what you see in each container.	**Model of interactions and explanations:** Sketch an interaction between the acetone and blue dye and explain what IMFs are present.
Water + glitter (blue dye)		
Water + yellow dye		
Acetone + glitter (blue dye)		
Acetone + yellow dye		



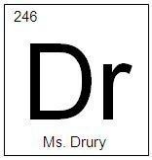
4. Explain why one color dye is more vividly mixed with water and the other dye is more vividly mixed with acetone using IMFs.

5. Pour 20 mL of the most vivid yellow solvent into a large test tube. Add 30 ml of the most vivid blue solvent on top. Pour gently so as to not mix the solvents. Record observations in the initial section below.

Initial Macroscopic Observations: Draw what you see and label what must be present in each layer BEFORE shaking.	**Final Macroscopic Observations:** Draw what you see and label what must be present in each layer AFTER shaking.	**Explain** why you think this happened using IMFs.

6. Now put a stopper in your test tube and shake the contents vigorously. Then put it back in your test tube rack and let it stand for a moment. Record observations in the final section above.
7. Explain why you think this happened using IMFs in the last column.
8. Weight 7 g of NaCl and add it to the test tube. Stopper the test tube and shake vigorously, until you see that the salt is starting to dissolve. Then put it back in your test tube rack and let it stand for a moment. Record observations below.

Macroscopic Observations: Draw what you see in the test tube and label what must be present in each layer.	**Interactions and explanations:** Explain any interactions between the 4 substances and explain what IMFs are present.



KMT Review

1. Kinetic Molecular Theory (KMT) describes how ideal gases behave. What does ideal mean? (For example, think about your ideal day off.)

2. Based on the demonstration performed by your teacher, describe how "ideal" gas particles behave below.

Motion	**Collisions**
Volume	**Attractions**

3. Draw models of gas particles in both low and high temperatures. Which seems more ideal?

low temperature	high temperature

4. Draw models of gas particles in both low and high external pressures. Which seems more ideal?

low pressure	high pressure

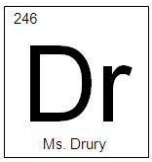
5. Under what conditions of Temperature and external Pressure can you get a real gas to behave the MOST like an ideal gas?

6. Under what conditions of Temperature and external Pressure can you get a real gas to behave the LEAST like an ideal gas?

7. Summarize your thoughts by completing the statement below:

   Kinetic Molecular Theory is the study of real gases (such as water vapor, oxygen, and helium) which may behave ideally when they are held at _____ temperatures and _____ pressures.

8. What could happen to the gas if it is held at extreme cold temperatures and high pressures?



Combined Gas Law Review

1. If the temperature of a 50mL sample of a gas is changed from 200K to 400K under constant pressure, what is the new volume of the gas?

2. The volume of a gas is 204mL when the pressure is 925kPa. At constant temperature, what is the final pressure if the volume increases to 306ml?

3. A 1.53L sample of sulfur (IV) oxide at a pressure of 5.60kPa. If the pressure is changed to 15.0kPa at constant temperature, what will be the new volume of the gas? Assume temperature is constant.

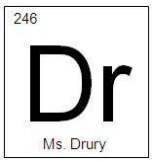
4. An aerosol spray can with a volume of 456mL contains 3.18g of propane gas as a propellant. If the can is at 23C, and 0.50atm, what volume would the propane occupy at STP?

5. A gas has a volume of 50. mL at a temperature of 10.0 K and a pressure of 760. kPa. What will be the new volume when the temperature is changed to 20.0 K and the pressure is changed to 380. kPa?

6. The volume of a sample of a gas at 273 K is 100.0 L. If the volume is decreased to 50.0 L at constant pressure, what will be the new temperature of the gas?

7. A gas has a volume of 2.00 L at 323 K and 3.00 atm. What will be the new volume if the temperature is changed to 273 K and the pressure is changed to 1 atm?

8. A gas occupies a volume of 500. mL at a pressure of 380. kPa and a temperature of 298 K. At what temperature will the gas occupy a volume of 250. mL and have a pressure of 760. kPa?



Graham's Law of Diffusion Review

1. At STP, which gas diffuses at the faster rate?
   A) H_2 B) N_2 C) CO_2 D) NH_3

2. Under the same conditions of temperature and pressure, which gas will diffuse at the *slowest* rate?
   A) He B) Ne C) Ar D) Rn

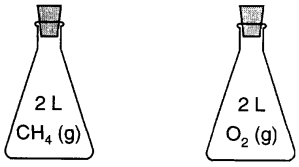
3. Which gas would diffuse most rapidly under the same conditions of temperature and pressure?
   A) gas *A*, molecular mass = 4
   B) gas *C*, molecular mass = 36
   C) gas *B*, molecular mass = 16
   D) gas *D*, molecular mass = 49

4. Arrange the following gas in order of increasing average molecular speed at 25 C: He, O_2, CO_2, H_2O.

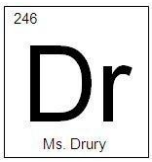
Avogadro's Law Review

1. A sample of oxygen gas is sealed in container X. A sample of hydrogen gas is sealed in container Z. Both samples have the same volume, temperature, and pressure. Which statement is true?
   A) Container X contains more gas molecules than container Z.
   B) Container X contains fewer gas molecules than container Z.
   C) Containers X and Z both contain the same number of gas molecules.
   D) Containers X and Z both contain the same mass of gas.

2. At the same temperature and pressure, 1.0 liter of CO(g) and 1.0 liter of CO2(g) have
   A) equal masses and the same number of molecules
   B) different masses and a different number of molecules
   C) equal volumes and the same number of molecules
   D) different volumes and a different number of molecules



3. Each stoppered flask to the right contains 2 liters of a gas at STP. Each gas sample has the same
   A) Density B) mass C) number of molecules D) number of atoms



Dalton's Law of Partial Pressures Review

1. Draw models that represent:
   a. one puff of helium gas (5 particles) in a 1L container at 273K and 1atm
   b. two puffs of neon gas in a 1L container at 273K and 2atm
   c. three puffs of gas: one of helium, two of neon at 273K. This is a combination of the containers.

a	b	c

   d. What do you think the pressure is inside container c? Explain your reasoning.

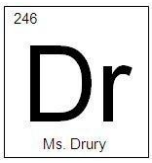
2. What is the pressure of a mixture of CO_2, SO_2, and H_2O gases, if each gas has a partial pressure of 200mmHg?
   a. 200mm b. 400mm c. 600mm d. 800mm

3. A flask contains a mixture of $N_2(g)$ and $O_2(g)$ at STP. If the partial pressure exerted by the $N_2(g)$ is 300mmHg, the partial pressure of the $O_2(g)$ is
   a. 300mm b. 360mm c. 460mm d. 600mm

4. Gas samples *A*, *B,* and *C* are contained in a system at STP. The partial pressure of sample *A* is 0.25atm and the partial pressure of sample *B* is 0.35atm. What is the partial pressure of sample *C*?
   a. 0.40atm b. 0.25atm c. 0.35atm d. 1.00atm

5. The partial pressures of gases *A*, *B*, and *C* in a mixture are 0.750 atmosphere, 0.250 atmosphere, and 1.25 atmospheres, respectively. What is the total pressure of the gas mixture in mmHg?
   a. 2.25mm b. 228mm c. 1710mm d. 2000mm

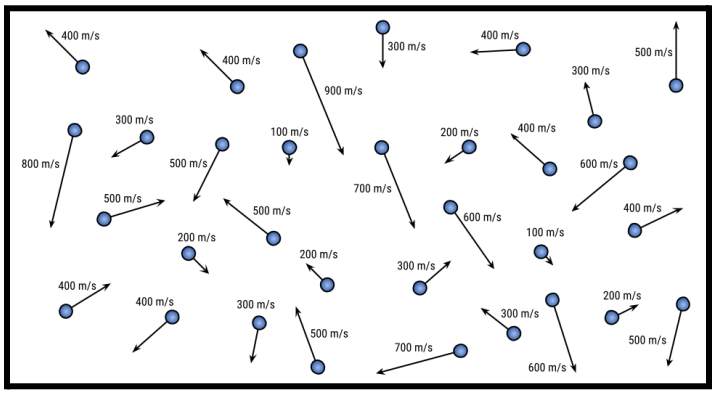
6. A mixture of gases contains 0.75mol nitrogen, 0.30mol oxygen, and 0.15 mol of carbon dioxide. If the total pressure is 2.3atm, what are the partial pressures?



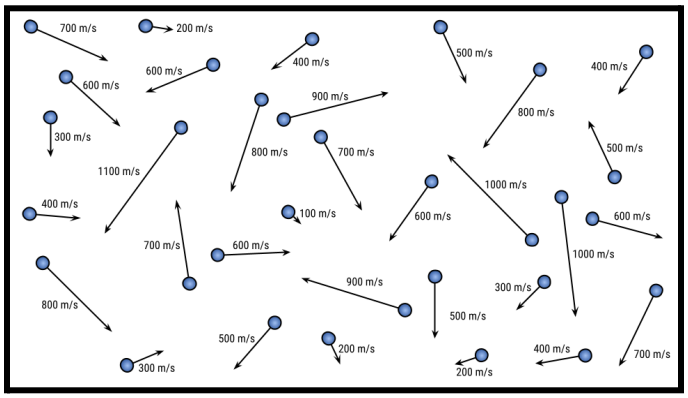
Maxwell-Boltzmann Distributions

Model 1 — Graphing Particle Speeds at Different Temperatures: The diagram below on the left shows a collection of gas particles in a rigid container. The speed of each particle is shown. The diagram to the right shows the same sample of particles heated using a lab burner.

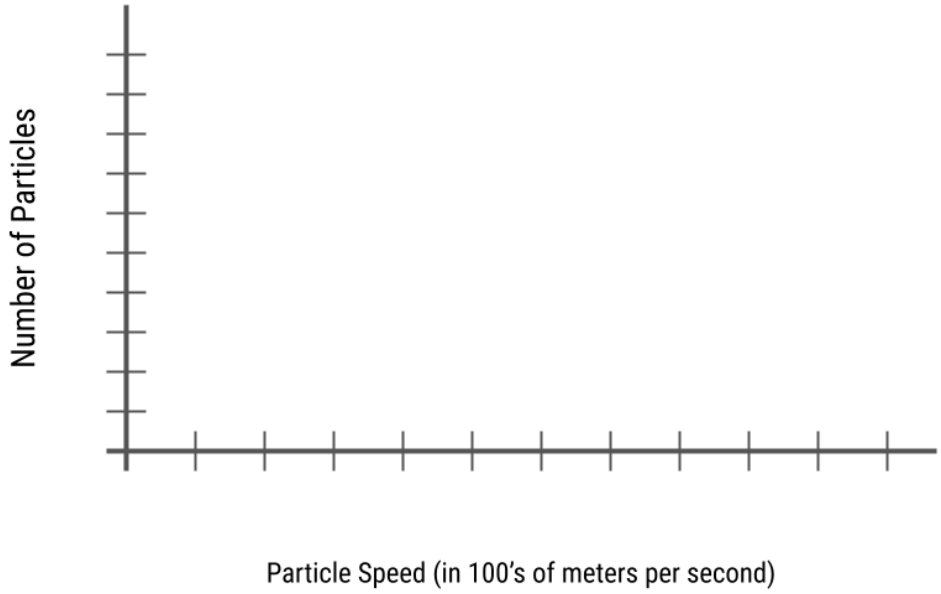
COLD



HOT



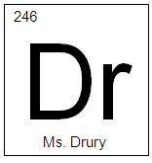
1. Create a line graph that shows the number of particles moving at each particle speed for the cold gas. Then also plot a curved line for the hot gas speeds.



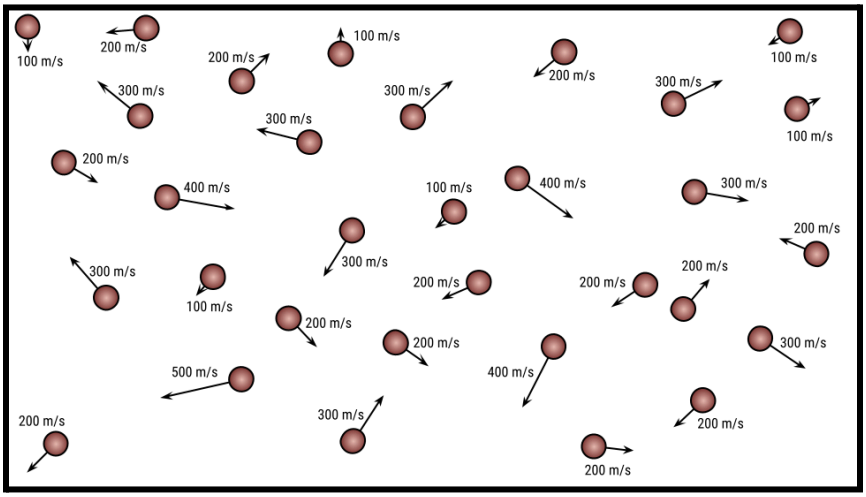
2. What do you notice about the particle speeds after being heated? Are there still slow particles?

3. What do you notice about how the graph has shifted after heating? What happened to the peak height and location?

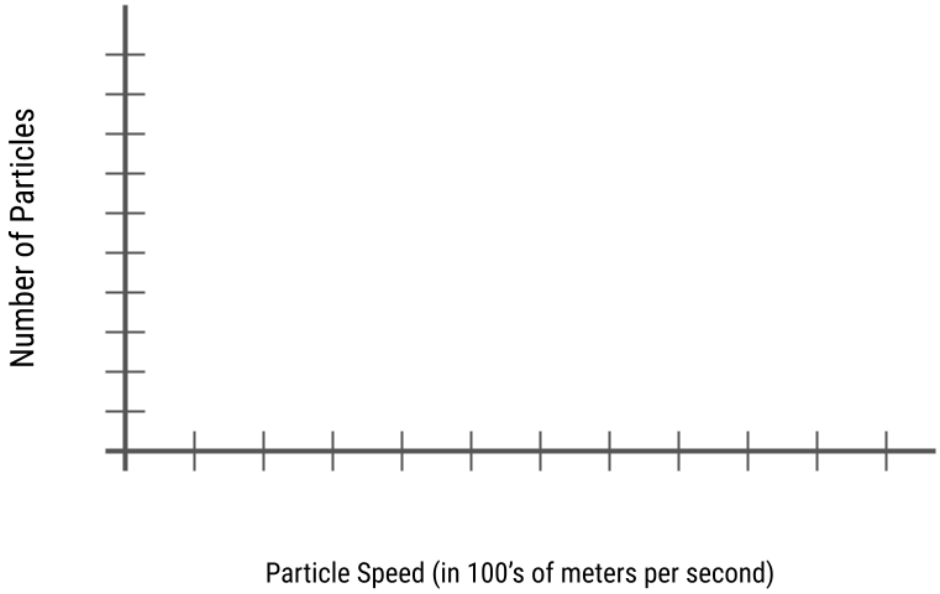
4. Are there more, less, or the same total number of particles in the hot and cold samples?



Model 2 — Graphing Particle Speeds for Particles of Different Mass: Here is another sample of gas particles. Each particle has four times the mass of the particles from Model 1. However, the temperature of this sample of particles is actually nearly exactly the same as the cold particles in Model 1.

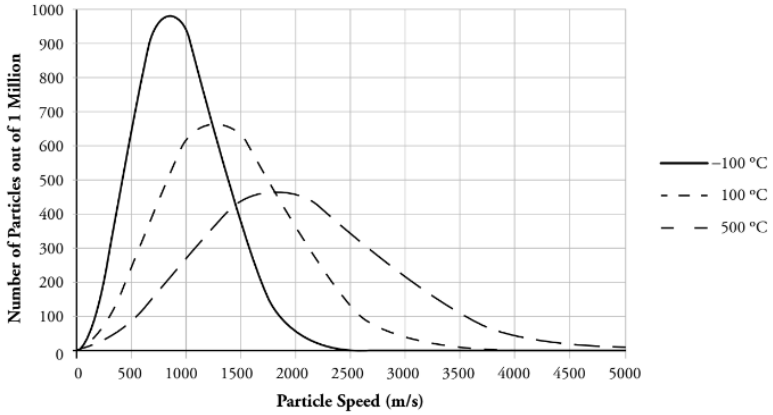


5. Create two Maxwell-Boltzmann distributions on the graph below, one for the less-massive cold particles (Model 1) and one for the more-massive (still cold) particles (Model 2). Label the curves "less massive" and "more massive."



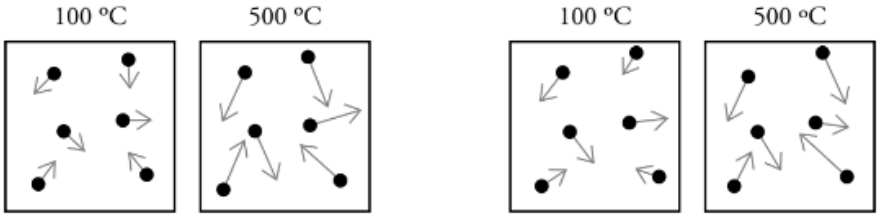
6. Keeping in mind that temperature is a measure of the average kinetic energy of a collection of particles, and that $\mathbf{KE_{average} = \frac{1}{2} mv^2}$, explain how these particles could have the same temperature of the particles in Model 1 even though their speeds are much slower.

Model 3: Helium at Different Temperatures

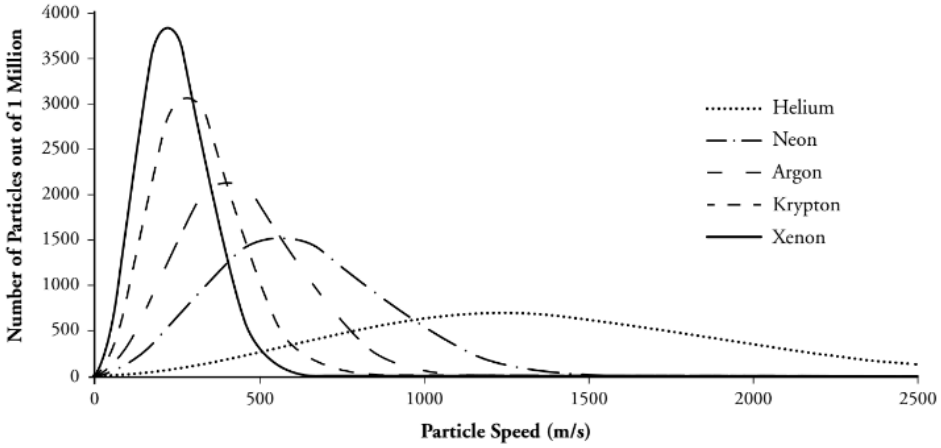


7. Label the peaks in the graph in Model 3 above with what they represent. Which peak hhas the largest grouping of particles moving the same speed? Is that a fast or slow speed compared to other atoms i the sample? Why are some peaks “flatter”?

8. Circle the set of drawings below that more accurately illustrates the difference in particle speeds for particles of helium at 100 °C and 500 °C? Justify your reasoning. Hint: The longer the arrow, the faster the particle is moving.

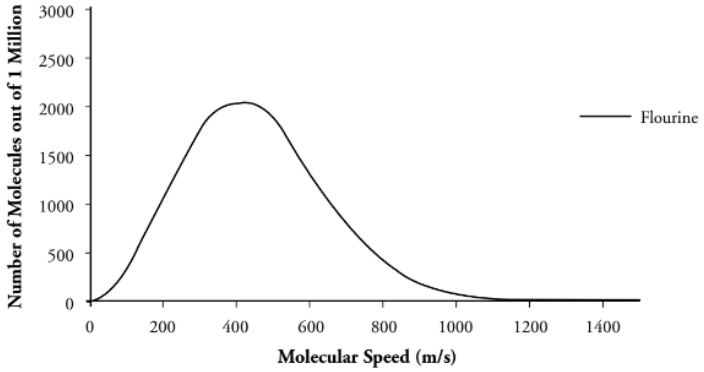


Model 4: Noble Gases at 100C

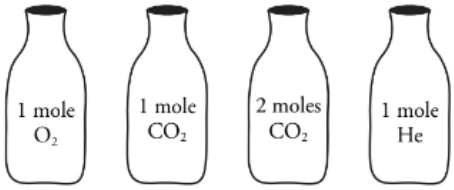


9. In a complete sentence, describe the relationship between the molar mass of particles in a sample and the average particle speed at a given temperature.

10. Consider the distribution curve below for fluorine gas (F_2) at 100 °C. Sketch an approximate distribution curve for chlorine gas (Cl_2) at 100 °C on the same graph.



11. A student is presented with four bottles containing different gases. All samples are at the same temperature.



a. Which gas sample has the fastest average particle speed?

b. Which gas sample has the highest average kinetic energy?

Molar Volume of a Gas Lab

Guiding Question: What is the volume of a mole of gas at STP?

Background Information: You will react magnesium with excess hydrochloric acid to form hydrogen gas. This gas will be fixed for STP values and you will determine your own molar volume. A eudiometer will be used to measure the gas collected over water. You must remember to correct the classroom pressure with the vapor pressure of the water.

Materials: Magnesium, 3M Hydrochloric acid, copper thread, eudiometer, stopper with hole, beaker of water.

Pre-Lab Questions:

1. Write the molecular equation of hydrochloric acid reacting with magnesium metal.

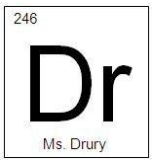
2. What is the limiting reagent in the lab? Explain.

3. You are using a piece of copper wire to hold the magnesium in place as it reacts with the acid. Does the copper interfere with the reaction? Explain.

4. What happens to the magnesium chloride that is produced in the reaction?

Procedure:

1. Obtain and measure the length of the Mg ribbon.
2. Obtain a piece of copper thread about 15 cm long. Tie the thread to the Mg with extra Cu hanging.
3. Obtain 10mL of hydrochloric acid and carefully pour it into the eudiometer.
4. Tilt the eudiometer and carefully pour water into the tube without mixing with the acid until full.
5. Lower the Mg ribbon into the tube and cap it in place with a stopper.
6. Set up a ring stand to hold the eudiometer and have a beaker of water ready.
7. With your finger over the stopper hole, lower the eudiometer into the water and clamp it to the ring stand. The top of the eudiometer should be submerged and almost touching the bottom of the beaker.
8. Let the reaction continue and take observations until the reaction stops. At that time, tap the eudiometer to dislodge stuck bubbles. Record the volume of Hydrogen produced.
9. Record the room temperature, room pressure, water vapor pressure and mass of 1m Mg.
10. Repeat for trial 2.



Data:

	Trial one	Trial two	Average of Trials
Length of Mg			
Mass of 100 cm of Mg			
Volume of H_2 in tube			
Room Temperature			
Room Pressure			
(Water) Vapor pressure			

Calculations:

1. Calculate the mass of Magnesium used because of the length of magnesium.

2. Calculate the moles of hydrogen gas that should be created using the mass of magnesium used.

3. Calculate the Pressure of the hydrogen gas only using the pressure of the room and water vapor.

4. Calculate the volume of hydrogen collected at STP.

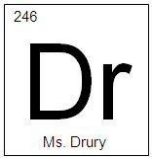
Claim: Answer the guiding question: What is the volume of a mole of gas at STP?

Evidence:

Reasoning: Provide necessary scientific concepts that support your claim and evidence. Calculate your percent error with 22.4L as the accepted value. Explain how errors in the lab lead to your %.

°C	torr (mmHg)	°C	torr (mmHg)	°C	torr (mmHg)	°C	torr (mmHg)
0	4.6	19	16.5	26	25.2	60	149.4
5	6.5	20	17.5	27	26.7	70	233.7
10	9.2	21	18.7	28	28.3	80	355.1
15	12.8	22	19.8	29	30.0	90	525.8
16	13.6	23	21.1	30	31.8	100	760.0
17	14.5	24	22.4	40	55.3	105	906.1
18	15.5	25	23.8	50	92.5	110	1074.6





Ideal Gas Law

1) A 7.8g piece of solid carbon dioxide is placed in a 4.0L container at 27 degrees Celsius. What is the pressure in atmospheres in the container after all the carbon dioxide vaporizes?

2) Automobile air bags inflate when sodium azide (NaN_3) decomposes into its elements. A sample of 97.5g of sodium azide is added to an air bag that will inflate to a volume of 0.30L and heat up to 100.°C.
   a) Write and balance the decomposition reaction of sodium azide.

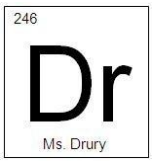
   b) Calculate the moles of nitrogen produced.

   c) Calculate the pressure exerted on the bag by the nitrogen gas produced in the reaction.

3) Using the Haber process to produce ammonia gas, you can obtain the greatest yield of ammonia at 500.°C and 250. atmospheres. A sample of 21.0g of nitrogen was reacted with excess hydrogen gas.
   a) Write and balance the production of ammonia, NH_3, from its elements.

   b) Calculate the moles of ammonia that were produced.

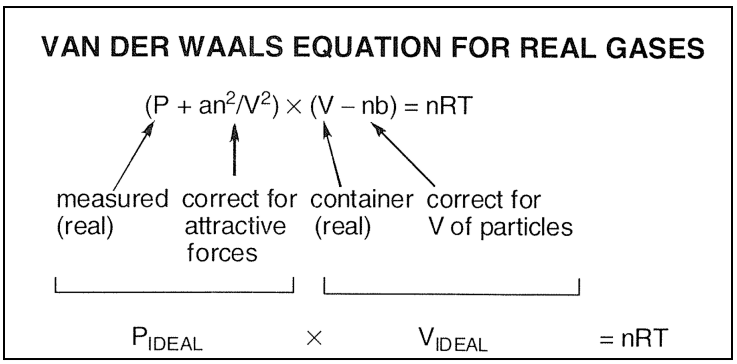
   c) What volume would ammonia occupy if 21.0g of nitrogen is reacted with excess hydrogen?



4) The Hindenburg exploded in 1937. It held 2.0×10^5 m^3 of hydrogen gas at 23C and standard pressure.
   a) Calculate the volume of hydrogen gas produced, recalling that $1cm^3 = 1mL$.

   b) Calculate the mass of hydrogen that burned in the Hindenburg explosion.

5) A scuba diver's tank contains 0.29kg of oxygen compressed into a volume of 2.3L. What is the pressure in the tank at 9.0°C?



6) The Real gas correction equation is given.
   a) Why would a correction to the pressure need to be made for real gases? → What happens to the pressure if the gases have IMFs?

   b) Why would a correction to the volume need to be made for real gases? → What happens to the volume calculation given that gases actually have volume?

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Ideal Gas Law with Density

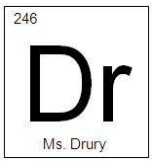
1. What is the density of nitrogen gas at 1.50atm and 298K?

2. What is the density of NO_2 gas at 0.970 atm and 35C?

3. Which gas is most dense at 1.00atm and 298K? CO_2, N_2O, or Cl_2?

4. A 44.8L balloon is filled with Helium with a pressure of 1atm at 25C. What is the density of the gas?

5. An AP student decided to bribe me with cookies for a better grade. He allowed the yeast to ferment so the dough would rise. If 200.0g of glucose (sugar) is used at standard pressure and 30C what is the volume of carbon dioxide gas produced? What is the density?



AP Gas Problems

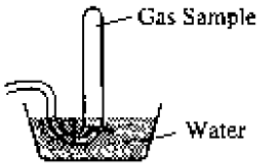
1. Observations about real gases can be explained at the molecular level according to the kinetic molecular theory of gases and ideas about intermolecular forces. Explain how each of the following observations can be interpreted according to these concepts, including how the observation supports the correctness of these theories.

(a) When a gas-filled balloon is cooled, it shrinks in volume; this occurs no matter what gas is originally placed in the balloon.

(b) When the balloon described in (a) is cooled further, the volume does not become zero; rather, the gas becomes a liquid or solid.

(c) When NH_3 gas is introduced at one end of a long tube while HCl gas is introduced simultaneously at the other end, a ring of white ammonium chloride is observed to form in the tube after a few minutes. This ring is closer to the HCl end of the tube than the NH_3 end.

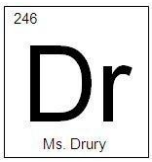
2. A student collected a sample of hydrogen gas by the displacement of water as shown by the diagram above. The relevant data are given in the following table.



GAS SAMPLE DATA	
Volume of sample	90.0 mL
Temperature	25°C
Atmospheric Pressure	745 mm Hg
Equilibrium Vapor Pressure of H_2O (25°C)	23.8 mm Hg

(a) Calculate the number of moles of hydrogen gas collected.

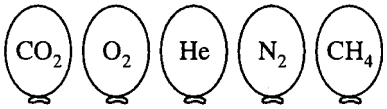
(b) Calculate the number of molecules of water vapor in the sample of gas.



(c) Explain the ratio of the average speed of the hydrogen molecules to the average speed of the water vapor molecules in the sample.

(d) Which of the two gases, H_2 or H_2O, deviates more from ideal behavior? Explain your answer.

3.



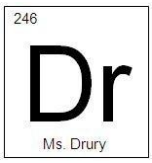
Represented above are five identical balloons, each filled to the same volume at 25°C and 1.0 atmosphere pressure with the pure gases indicated.

(a) Which balloon contains the greatest mass of gas? Explain.

(b) Compare the average kinetic energies of the gas molecules in the balloons. Explain.

(c) Which balloon contains the gas that would be expected to deviate most from the behavior of an ideal gas? Explain.

(d) Twelve hours after being filled, all the balloons have decreased in size. Predict which balloon will be the smallest. Explain your reasoning.



Gases AP Multiple Choice

1. When a sample of oxygen gas in a closed container of constant volume is heated until its absolute temperature is doubled, which of the following is also doubled?
(A) The density of the gas (B) The pressure of the gas (C) The average velocity of the gas
(D) The number of molecules per cm^3 (E) The potential energy of the molecules

2. The density of an unknown gas is 4.20 grams per liter at 3.00 atmospheres pressure and 127 °C. What is the molecular weight of this gas? (R = 0.0821 L-atm / mole-K)
(A) 14.6 (B) 46.0 (C) 88.0 (D) 94.1 (E) 138

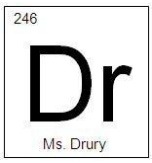
3. Equal masses of three different ideal gases, X, Y, and Z, are mixed in a sealed rigid container. If the temperature of the system remains constant, which of the following statements about the partial pressure of gas X is correct?
(A) It is equal to 1/3 the total pressure
(B) It depends on the intermolecular forces of attraction between molecules of X, Y, and Z.
(C) It depends on the relative molecular masses of X, Y, and Z.
(D) It depends on the average distance traveled between molecular collisions.
(E) It can be calculated with knowledge only of the volume of the container.

4. Two flexible containers for gases are at the same temperature and pressure. One holds 0.50 grams of hydrogen and the other holds 8.0 grams of oxygen. Which of the following statements regarding these gas samples is FALSE?
(A) The volume of the hydrogen container is the same as the volume of the oxygen container.
(B) The number of molecules in the hydrogen container is the same as the number of molecules in the oxygen container.
(C) The density of the hydrogen sample is less than that of the oxygen sample.
(D) The average kinetic energy of the hydrogen molecules is the same as the average kinetic energy of the oxygen molecules.
(E) The average speed of the hydrogen molecules is the same as the average speed of the oxygen molecules.

5. When the actual gas volume is greater than the volume predicted by the ideal gas law, the explanation lies in the fact that the ideal gas law does NOT include a factor for molecular…
(A) volume (B) mass (C) velocity (D) attractions (E) shape

6. A compound is heated to produce a gas whose molecular weight is to be determined. The gas is collected by displacing water in a water-filled flask inverted in water. Which of the following is needed to calculate the molecular weight of the gas, but does NOT need to be measured during the experiment?
(A) Mass of the compound used in the experiment (B) Temperature of the water in the trough
(C) Vapor pressure of the water (D) Barometric pressure
(E) Volume of water displaced from the flask

7. A sample of 9.00 grams of aluminum metal is added to an excess of hydrochloric acid. The volume of hydrogen gas produced at standard temperature and pressure is…
(A) 22.4 liters (B) 11.2 liters (C) 7.46 liters (D) 5.60 liters (E) 3.74 liters



8. A gaseous mixture containing 7.0 moles of nitrogen, 2.5 moles of oxygen, and 0.50 mole of helium exerts a total pressure of 0.90 atmospheres. What is the partial pressure of the nitrogen?

(A) 0.13 atm (B) 0.27 atm (C) 0.63 atm (D) 0.90 atm (E) 6.3 atm

9. A sample of 0.010 moles of oxygen gas is confined at 127 °C and 0.80 atmospheres. What would be the pressure of this sample at 27 °C and the same volume?

(A) 0.10 atm (B) 0.20 atm (C) 0.60 atm (D) 0.80 atm (E) 1.1 atm

10. Hydrogen gas is collected over water at 24 °C. The total pressure of the sample is 755 millimeters of mercury. At 24 °C, the vapor pressure of water is 22 millimeters of mercury. What is the partial pressure of the hydrogen gas?

(A) 22 mm Hg (B) 733 mm Hg (C) 755 mm Hg (D) 760 mm Hg (E) 777 mm Hg

11. A 2.00-liter sample of nitrogen gas at 27 °C and 600. millimeters of mercury is heated until it occupies a volume of 5.00 liters. If the pressure remains unchanged, the final temperature of the gas is…

(A) 68 °C (B) 120 °C (C) 477 °C (D) 677 °C (E) 950. °C

12. $2\ K + 2\ H_2O \rightarrow 2\ K^+ + 2\ OH^- + H_2$

When 0.400 moles of potassium reacts with excess water at standard temperature and pressure as shown in the equation above, the volume of hydrogen gas produced is…

(A) 1.12 liters (B) 2.24 liters (C) 3.36 liters (D) 4.48 liters (E) 6.72 liters

13. A sample of 0.0100 moles of oxygen gas is confined at 37 °C and 0.216 atmospheres. What would be the pressure of this sample at 15 °C and the same volume?

(A) 0.0876 atm (B) 0.175 atm (C) 0.201 atm (D) 0.233 atm (E) 0.533 atm

14. Which of the following gases deviates most from ideal behavior?

A) SO_2 B) Ne C) CH_4 D) N_2 E) H_2

15. A hydrocarbon gas with an empirical formula CH_2 has a density of 1.88 grams per liter at 0 °C and 1.00 atmospheres. A possible formula for the hydrocarbon is…

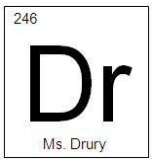
(A) CH_2 (B) C_2H_4 (C) C_3H_6 (D) C_4H_8 (E) C_5H_{10}

16. A sample of 3.0 grams of an ideal gas at 121°C and 1.0 atmospheres pressure has a volume of 1.5 Liters. Which of the following expressions is correct for the molar mass of the gas? (R = 0.082 L-atm / mole-K)

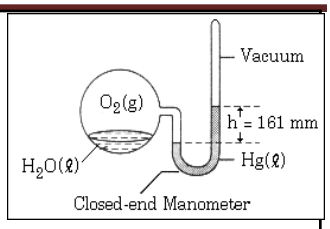
(A) [(0.082)(394)] / [(3.0)(1.0)(1.5)]
(B) [(l.0)(l.5)] / [(3.0)(0.082)(394)]
(C) [(0.082)(1.0)(1.5)] / [(3.0)(394)]
(D) [(3.0)(0.082)(394)] / [(1.0)(1.5)]
(E) [(3.0)(0.082)(1.5)] / (1.0)(394)]

17. Samples of F_2 gas and Xe gas are mixed in a container of fixed volume. The initial partial pressure of the F_2 gas is 8.0 atmospheres and that of the Xe gas is 1.7 atmospheres. When all of the Xe gas reacted, forming a solid compound, the pressure of the unreacted F_2 gas was 4.6 atmospheres. The temperature remained constant. What is the formula of the compound?

(A) XeF (B) XeF_3 (C) XeF_4 (D) XeF_6 (E) XeF_8



18. The system shown in the picture is at equilibrium at 28°C.
At this temperature, the vapor pressure of water is 28 millimeters of mercury. The partial pressure of $O_{2(g)}$ in the system is…



(A) 28 mm Hg (B) 56 mm Hg (C) 133 mm Hg
(D) 161 mm Hg (E) 189 mm Hg

19. A sample of an ideal gas is cooled from 50.0 °C to 25.0 °C in a sealed container of constant volume. Which of the following values for the gas will decrease?
I. The average molecular mass of the gas
II. The average distance between the molecules
III. The average speed of the molecules
(A) I only (B) II only (C) III only (D) I and III (E) II and III

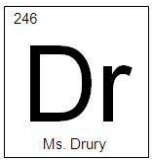
20. At 25 °C, a sample of NH_3 (molar mass 17 grams) effuses at the rate of 0.050 moles per minute. Under the same conditions, which of the following gases effuses at approximately one-half that rate?
(A) O_2 (molar mass 32 grams) (B) He (molar mass 4.0 grams) (C) CO_2 (molar mass 44 grams)
(D) Cl_2 (molar mass 71 grams) (E) CH_4 (molar mass 16 grams)

	Pressure of $O_{2\,(g)}$ above $H_2O_{(l)}$ (atm)	Temperature of $H_2O_{(l)}$ °(C)
(A)	5	80
(B)	5	20
(C)	1	80
(D)	1	20
(E)	0.5	20

21. Under which of the following sets of conditions could the most $O_{2\,(g)}$ be dissolved in $H_2O_{(l)}$?

22. A hot-air balloon, shown at the right, rises. Which of the following is the best explanation?
(A) The pressure on the walls of the balloon increases with increasing temperature.
(B) The difference in temperature between the air inside and outside the balloon produces convection currents.
(C) The cooler air outside the balloon pushes in on the walls of the balloon.
(D) The rate of diffusion of cooler air is less than that of warmer air.
(E) The air density inside the balloon is less than that of the surrounding air.

23. A rigid metal tank contains oxygen gas. Which of the following applies to the gas in the tank when additional oxygen is added at constant temperature?
(A) The volume of the gas increase. (B) The pressure of the gas decreases.
(C) The average speed of the gas molecules remains the same.
(D) The total number of gas molecules remains the same.
(E) The average distance between the gas molecules increases.



24. $W_{(g)} + X_{(g)} \rightleftharpoons Y_{(g)} + Z_{(g)}$

Gas W and X react in a closed, rigid vessel to form Gas Y and Z according to the equation above. The initial pressure of $W_{(g)}$ is 1.20 atm and that of $X_{(g)}$ is 1.60 atm. No $Y_{(g)}$ or $Z_{(g)}$ is initially present. The experiment is carried out at constant temperature. What is the partial pressure of $Z_{(g)}$ when the partial pressure of $W_{(g)}$ has decreased to 1.0 atm?

(A) 0.20 atm (B) 0.40 atm (C) 1.0 atm (D) 1.2 atm (E) 1.4 atm

25. $NH_4NO_{3(s)} \rightarrow N_2O_{(g)} + 2\ H_2O_{(g)}$ (Container Volume = 1.0 L)

A 0.03 mol sample of $NH_4NO_{3(s)}$ decomposes completely according to the balanced equation above. The total pressure in the flask measured at 400 K is closest to which of the following? (R = 0.08 L-atm / mole-K)

(A) 3 atm (B) 1 atm (C) 0.5 atm (D) 0.1 atm (E) 0.03 atm

26. Equal numbers of moles of $He_{(g)}$, $Ar_{(g)}$, and $Ne_{(g)}$ are placed in a glass vessel at room temperature. If the vessel has a pinhole-sized leak, which of the following will be true regarding the relative values of the partial pressures of the gases remaining in the vessel after some of the gas mixture has effused?

(A) $P_{He} < P_{Ne} < P_{Ar}$ (B) $P_{He} < P_{Ar} < P_{Ne}$ (C) $P_{Ne} < P_{Ar} < P_{He}$

(D) $P_{Ar} < P_{He} < P_{Ne}$ (E) $P_{He} = P_{Ar} = P_{Ne}$

Questions 27-29 refer to the following gases at 0°C and 1 atm.

(A) Ne
(B) Xe
(C) O_2
(D) CO
(E) NO

27. Has an average atomic or molecular speed closest to that of N_2 molecules at 0°C and 1 atm.
28. Has the greatest density.
29. Has the greatest rate of effusion through a pinhole.

30. A flask contains 0.25 mol of $SO_{2(g)}$, 0.50 mole of $CH_{4(g)}$, and 0.50 mole of $O_{2(g)}$. The total pressure of the gases in the flask is 800 mm Hg. What is the parital pressure of the $SO_{2(g)}$ in the flask?

(A) 800 mm Hg (B) 600 mm Hg (C) 250 mm Hg
(D) 200 mm Hg (E) 160 mm Hg

31. A 2 L container will hold about 4 g of which of the following gases at 0°C and 1 atm?

(A) SO_2 (B) N_2 (C) CO_2 (D) C_4H_8 (E) NH_3

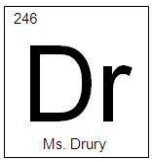
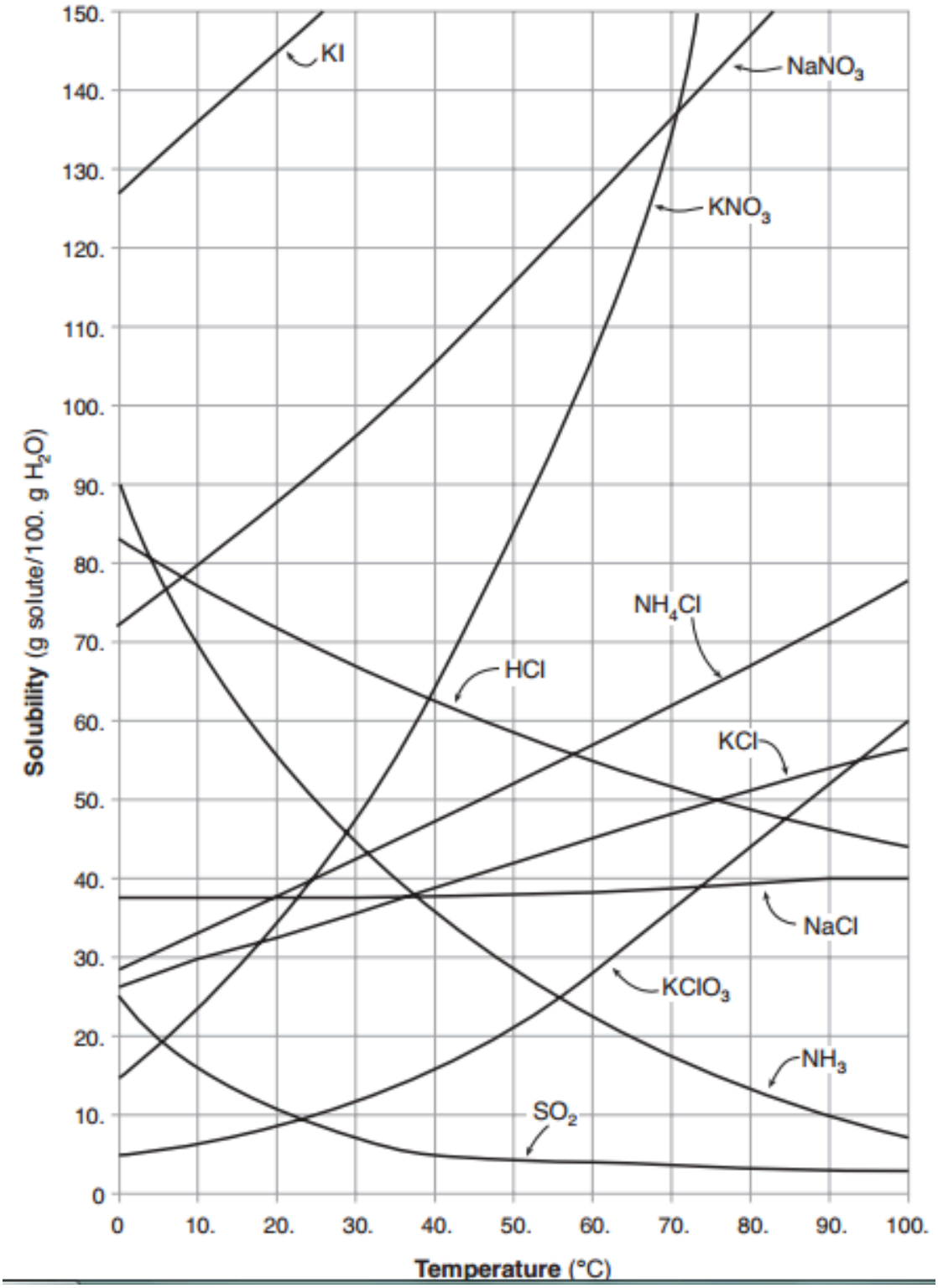


Table G
Solubility Curves at Standard Pressure



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Dr
Ms. Drury

Solutions Review

Solubility Rules: Compounds containing Alkali metals, NH_4^+, $C_2H_3O_2^-$, and NO_3^- are always soluble.

1. Which of the following must be soluble? NaCl, $PbBr_2$ K_3PO_4 $MgCO_3$ calcium hydroxide, copper (II) nitrate

2. State whether each of the following solutions is *saturated, unsaturated, or supersaturated*.
   (a) 80 g $NaNO_3$ in 100 g H_2O at 10ºC _____
   (b) 75 g $NaNO_3$ in 100 g H_2O at 10ºC _____
   (c) 90 g $NaNO_3$ in 100 g H_2O at 10ºC _____
   (d) 5 g $KClO_3$ in 100 g H_2O at 5ºC _____
   (e) 35 g $NaNO_3$ in 50 g H_2O at 10ºC _____
   (f) 5 g $KClO_3$ in 50 g H_2O at 5ºC _____
   (g) 5 g $KClO_3$ in 200 g H_2O at 5ºC _____
   (h) 30 g NH_4Cl in 200 g H_2O at 10ºC _____

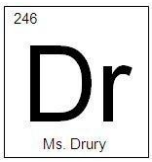
3. Identify how many MORE grams of each solute must be added to 100 g of water to form a saturated solution at that temperature.

Grams Solute per 100 g H_2O	Solute Added to make Saturated
a. 35 g KNO_3 at 40ºC	
b. 50 g NH_3 at 10ºC	
c. 15 g KCl at 75ºC	
d. 95 g KI at 15ºC	

4. Identify how many grams of each solute will crystallize/precipitate/settle. Assume all solutions are saturated and in 100 grams of H_2O.

Amount cooled	Amount Precipitated
a. KNO_3 (aq) is cooled from 70ºC to 40ºC	
b. NH_4Cl (aq) is cooled from 90ºC to 20ºC	
c. KCl (aq) is cooled from 55ºC to 30ºC	
d. KI (aq) is cooled from 20ºC to 5ºC	

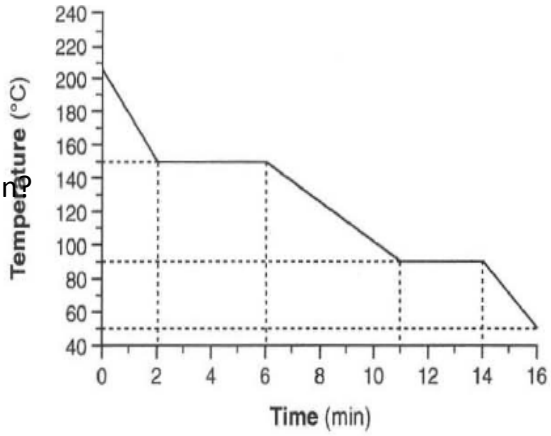
5. Rank the following solids in order from least to most soluble in 100 g H_2O at 50ºC :
   NH_4Cl, $NaNO_3$, $KClO_3$, KNO_3



Phase Changes Review

1. Label the line segments with their phase(s).
2. What is this substance's melting point? _____
3. What is this substance's boiling point? _____
4. Does this represent an endothermic or exothermic reaction?
5. Heat is being released at 60.0 kilojoules per minute.

How much heat is released when the substance freezes?



6. Label the point with the most kinetic energy with a star.

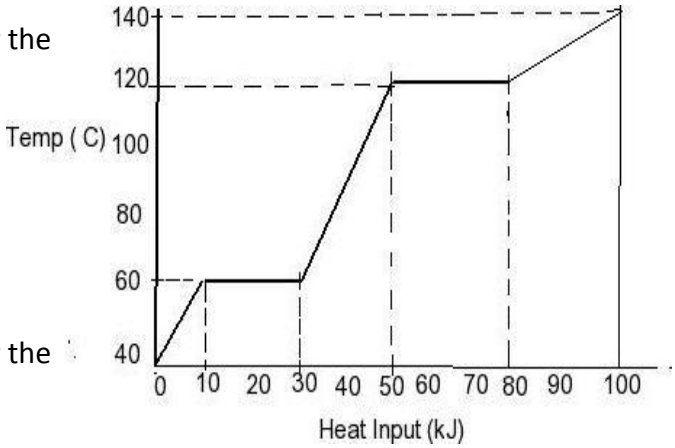
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7. Draw six particles of this substance as it looks for the first line segment in the box below.

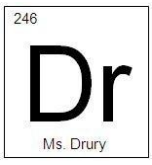


8. Draw six particles of this substance as it looks for the last line segment in the box below.





9. At which point is the potential energy the highest? Label it with a star.
10. What is the boiling point of this substance? _____
11. What is the melting point of this substance? _____
12. What would you expect the graph to do if the substance continued to be heated?
13. Which takes longer and why: melting ice or vaporizing water?



Concentration

A solution is a homogeneous mixture. Even though concentrations of solutions can be presented in a variety of different ways in chemistry, concentrations can be broken into 2 big groups: those that depend on mass only, and those that depend on both mass and volume. Since volume of an aqueous solution is temperature dependent, concentrations that depend on volume are only used when temperature is constant.

Density (d):	$\frac{\text{Mass solution}}{\text{Volume solution}}$	Mole Fraction (x) = $\frac{\text{Moles Part}}{\text{Total moles}}$
Molarity (M):	$\frac{\text{Moles solute}}{\text{Volume solution (L)}}$	Mass percent = $\frac{\text{Mass part}}{\text{Total mass}}$ x 100%
Molality (m):	$\frac{\text{Moles solute}}{\text{Mass solvent (kg)}}$	
Parts per Million (ppm):	$\frac{\text{mass of part}}{\text{Total mass whole}}$ x 1,000,000	

Fill in each box in the column with "**X**" if that measurement is needed to calculate the concentration.

Density	Molarity	Mole fraction	Molality	Mass percent	Parts per Million	***Measurement needed***
						Mass solute
						Mass solvent
						Volume solution

Which of these concentrations would be temperature dependent?

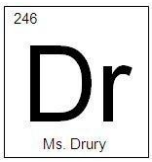
This question will ask you to calculate the concentration of an aqueous solution of **concentrated** HCl in a variety of the concentration units outlined above using the information given in the question.

1. You will need to work with a specific volume of the concentrated HCl solution, so choose any volume you want and record it here_____.

2. Concentrated hydrochloric acid solution has a density of 1.19 g/mL. Calculate the mass of this solution for the sample volume you chose in #1.

3. Concentrated hydrochloric acid solution is 37.2 % by mass HCl. Calculate:
   a. The mass of HCl in this sample.

   b. The moles of HCl in this sample.



c. The mass of the solvent in this sample.

d. The moles of solvent in this sample.

4. Calculate the following for a sample of the concentrated HCl solution:
   a. Molarity of HCl

   b. Molality of HCl

   c. χ_{HCl}

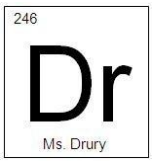
   d. χ_{H2O}

   e. parts per million HCl

5. A solution of Na_2CO_3 is prepared by dissolving 25.0 g in water to prepare a solution with a total volume of 350.0 mL. Calculate the concentration, in M, of each of the following.
   a. Na_2CO_3

   b. Na^+

   c. CO_3^{-2}



6. Explain why the procedure shown below for preparing 350.0 mL of a solution of $Na_2CO_3(aq)$ is incorrect.
   (1) Add 25.0 g $Na_2CO_3(s)$ to a 350.0 mL volumetric flask
   (2) Add 350.0 mL distilled water to the flask.
   (3) Swirl the solution until all of the solid dissolves completely.

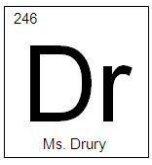
7. A student needs to prepare 500.0 mL of 0.800 M $CuSO_4(aq)$.
   a. Calculate the mass of $CuSO_4(s)$ needed to prepare this solution.

   b. The most common form of $CuSO_4$ found in the laboratory is not anhydrous $CuSO_4$, but rather the hydrate compound, copper(II) sulfate pentahydrate, which has the chemical formula $CuSO_4 \bullet 5H_2O$. Calculate the mass of $CuSO_4 \bullet 5H_2O(s)$ needed to prepare 500.0 mL of 0.600 M.

8. A student needs to prepare 200.0 mL of 0.750 M KOH(aq). Calculate the volume of 3.00 M KOH(aq)needed to prepare this solution. Describe the dilution procedure.

9. The solubility of $Ca(OH)_2$ in water is equal to 0.0118 M at room temperature. Calculate the minimum volume of distilled water (in milliliters) that must be added to a sample of 0.300g $Ca(OH)_2(s)$ in order to dissolve the solid completely.

10. Draw particle diagrams to represent the following:
    a. 1M NaCl in 1L container
    b. 1M NaCl in 2L container
    c. 2M NaCl in 1 liter container
    d. 1M $MgCl_2$ in 1 liter container



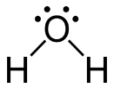
AP Chromatography Lab

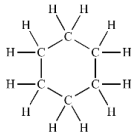
Guiding Question: How does polarity play a role in the separation of mixtures by Chromatography? (Are all black markers made of the same dye?)

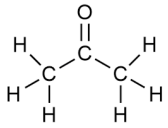
Materials: Chromatography paper, beakers, rulers, dyes, water, acetone, cyclohexane.

Procedure: You are given markers of unknown polarity, beakers, a solvent, a pencil, and chromatography paper.

1. Using a pencil mark a line across the paper, **about 2 cm from the bottom**.
2. Using pencil, mark an x, **1cm above the line**. Place one small dot of dye 1 from a marker on the x.
3. Add a small amount of solvent to the beaker, only enough to fill the bottom of the beaker, **with a height of the line you made in procedure 1 (hold the paper up to the beaker)**.
4. Carefully place the marked paper in the beaker and allow the solvent to move up the paper.
   NOTE: If the marker color (solute) immediately touches the solvent and begins to dissolve before the solute/solvent starts to move up the paper you must start over.
5. Allow the solvent to migrate until the solute color stops moving.
6. Remove the paper and measure the distance the solvent moved (from the origin line) and the distance the solute moved (from origin x). Record color observations. Calculate the retention factor:
   Rf = distance from origin x to center of solute spot / distance from origin line to solvent top
7. Repeat steps 1-6 with marker 2.
8. On the whiteboard draw and label your results. Be sure to include labels for the type of **solvent**, the name of the **solute**, and **Rf value**. Visit other groups to record their results.

Water 

cyclohexane 

acetone 

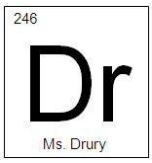
	Dye 1	Dye 2
Water observations and Rf values		
Cyclohexane observations and Rf values		
Acetone observations and Rf values		

Post Lab Questions:

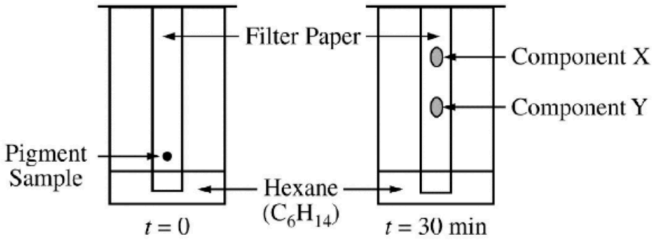
1. Identify the polarity of each of the solvents in this lab:
   a. Water b. Cyclohexane c. Acetone

2. Using rules of intermolecular forces of attraction, identify the polarities of the dyes and explain how you came to your claim:

3. Why did both dyes dissolve in acetone?



4. The diagram shown is a representation of a paper chromatography experiment.



   a. The mobile phase is which is classified as polar or nonpolar? Explain.

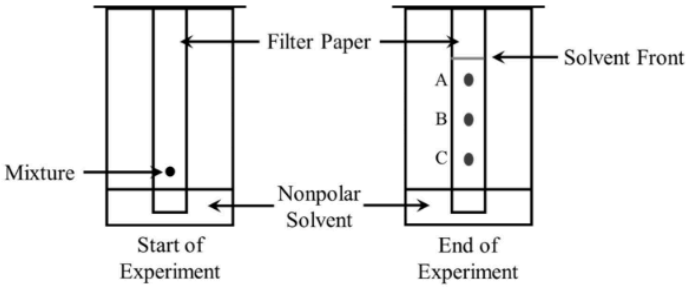
   b. The stationary phase in this experiment is the paper itself, which happens to be moderately polar. Which component of the pigment sample, X or Y, traveled a farther distance up the paper in this experiment?

   c. Substances with similar intermolecular interactions tend to be miscible or soluble in one another. Therefore which one of the following is a reasonable conclusion based on the results of this experiment?

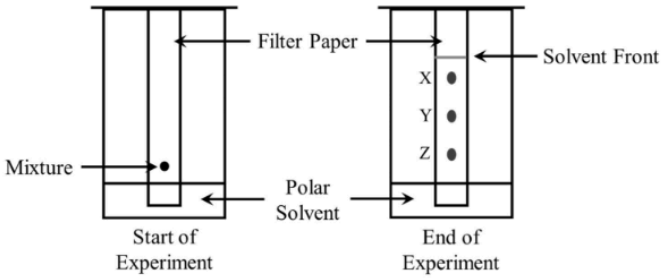
      X is more polar than Y.

      Y is more polar than X.

5. The solvent front is the mark on the chromatography paper that indicates how far the solvent traveled up the paper. The component of the mixture that travels the farthest distance up the paper will be the one that is the most soluble in the solvent (mobile phase).

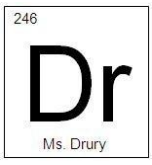


   a. Based on the information above, which substance, A, B, or C, is the most polar component of the mixture?

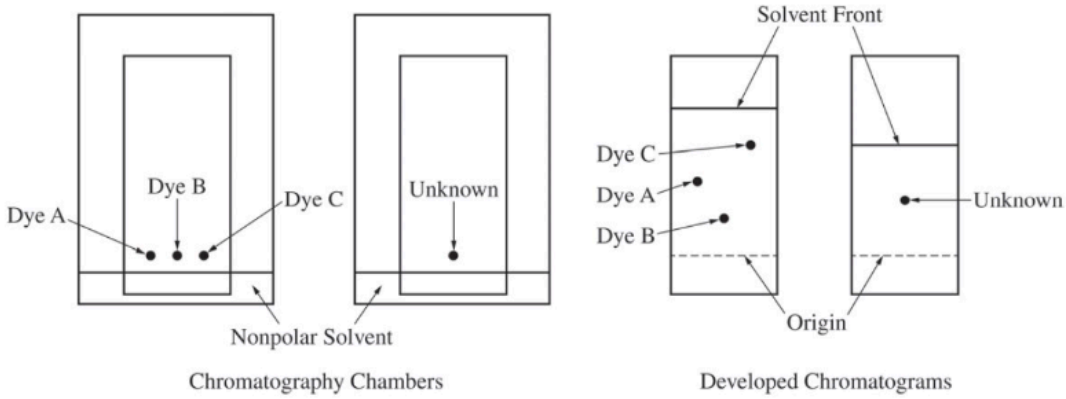


   b. Based on the information above, which substance, X, Y, or Z, is the most polar component of the mixture?

   c. What can you conclude about the solubility of a substance that does not travel up the paper at all during a chromatography experiment?



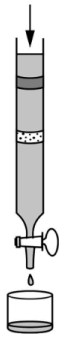
Separating Mixtures



1. A student investigates various dyes using paper chromatography. The student has samples of three pure dyes, labeled A, B, and C, and an unknown sample that contains one of the three dyes. The student prepares the chromatography chambers shown above on the left by putting a drop of each dye at the indicated position on the chromatography paper (a polar material) and standing the paper in a nonpolar solvent. The developed chromatograms are shown above on the right.
    a. Which dye (A, B, or C) is the least polar? Which dye is the most polar? Justify your answers in terms of the interactions between the dyes and the solvent or between the dyes and the paper.

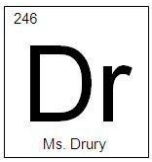
    b. Which dye is present in the unknown sample? Justify your answer.

2. Suppose that you have a mixture of dyes or pigments, and you need to separate the mixture into separate components and collect a substantial sample of each component. You could use column chromatography to do this. The mixture of dyes would be loaded onto a column that usually contains a solid such as silica gel (a form of hydrated SiO_2). A solvent runs through the column and serves as the mobile phase. The mixture travels downward through the column, and each component of the mixture should have different strengths of their attractions for the mobile phase and the stationary phase. Each dye should pass through the column at a different rate, and this allows for collection of each dye separately.

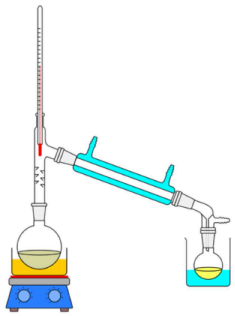


    a. If the mobile phase is polar, which type of dye should come out of the column first?

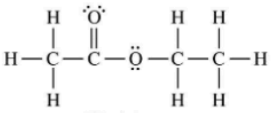
    b. If the mobile phase is non polar, which type of dye should come out of the column first?



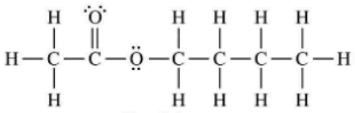
3. The diagram shows a distillation setup. If a mixture of ethanol (BP = 78°C) and water (BP = 100°C) is gently heated, which component of the mixture, ethanol or water, will appear in the collection flask first? Justify your answer.



4. A mixture containing equal numbers of moles of ethyl acetate and butyl acetate was separated using distillation. Based on the diagrams shown above, which of the following identifies the substance that would be initially present in higher concentration in the distillate and correctly explains why that occurs?

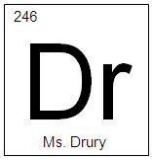


Ethyl Acetate



Butyl Acetate

a. Ethyl acetate, because it has fewer C–C bonds to break
b. Ethyl acetate, because it has a shorter carbon chain and weaker London dispersion forces
c. Butyl acetate, because it has more C–C bonds to break
d. Butyl acetate, because it has a longer carbon chain and weaker dipole-dipole attraction

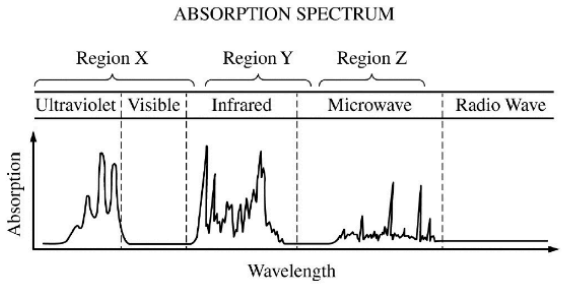


Spectroscopy

Spectroscopy is the study of how electromagnetic radiation interacts with a sample of matter. A sample of matter can absorb energy, and it can also emit energy. Different forms of radiation will have different effects on different samples of matter. Use this simulation to summarize effects of ER on molecules. Then answer the questions that follow. https://phet.colorado.edu/en/simulations/molecules-and-light

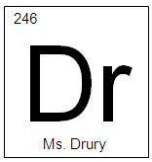
Radiation	Relative Energy	Effect on Molecule
microwave	low	
infrared	medium	
visible and UV	high	

1. N_2 molecules absorb ultraviolet light but not visible light. I_2 molecules absorb both visible and ultraviolet light. Which of the following statements explains the observations?
   a. More energy is required to make N_2 molecules vibrate than is required to make I_2 molecules vibrate.
   b. More energy is required to remove an electron from an I_2 molecule than is required to remove an electron from a N_2 molecule.
   c. Visible light does not produce transitions between electronic energy levels in the N_2 molecule but does produce transitions in the I_2 molecule.
   d. The molecular mass of I_2 is greater than the molecular mass of N_2



2. The diagram above represents the absorption spectrum for a pure molecular substance. Which of the following correctly indicates the type of transition observed for the substance in each of the regions of the absorption spectrum?

	Region X	Region Y	Region Z
a.	Molecular vibration	Molecular rotation	Electronic transition
b.	Electronic transition	Molecular rotation	Molecular vibration
c.	Molecular rotation	Molecular vibration	Electronic transition
d.	Electronic transition	Molecular vibration	Molecular rotation



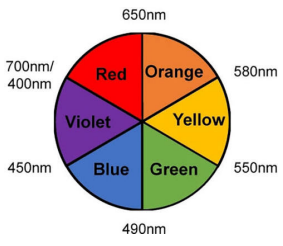
3. Which of the following best explains what happens as photons of visible light are absorbed by dye molecules?
    a. Certain electrons in the dye molecule move to a higher energy level, with the difference in energy between the lower and higher energy levels being the same as the energy of the absorbed photons.
    b. Certain chemical bonds in the dye molecules begin to bend and stretch, with the difference in energy between the lower and higher vibrational states being the same as the energy of the absorbed photons.
    c. The dye molecules begin to rotate faster in certain modes, with the difference in energy between the lower and higher rotational states being the same as the energy of the absorbed photons.
    d. Certain covalent bonds in the dye molecules begin to break and reform, with the bond energies of the bonds being the same as the energy of the absorbed photons.

In spectroscopy, a sample is exposed to a source of electromagnetic radiation. This radiation causes the sample to be affected in some way, involving both the absorption and the emission of energy. The changes in energy are processed through a detector, producing a spectrum. The energy absorbed or emitted from a sample can be described in terms of photons. A photon refers to a particle of radiation that has a specific wavelength, frequency, and energy. Recall the following calculations:

4. The Cl−Cl bond has a bond energy of 242 kJ mol^{-1}.
    a. Calculate the amount of energy, in joules, needed to break a single Cl−Cl bond.

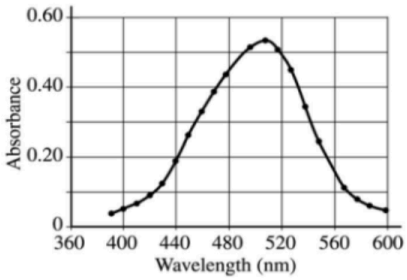
    b. Calculate the longest wavelength of light, in meters, that can supply the energy per photon necessary to break the Cl−Cl bond.

Instruments such as spectrophotometers or colorimeters are designed to measure the absorbance of a a particular sample. Many substances are able to absorb light. When a substance absorbs light, electrons in the ground state become excited and move up to higher energy levels. For many transition metal ions, electrons absorb photons of light in the visible portion of the electromagnetic spectrum. That is why many transition metal ions appear colored. The color wheel below illustrates that the perceived color has a complementary relationship with the color of visible light that is absorbed. For example, Cu^{2+} (aq) appears blue to our eyes because the solution has a strong absorbance in the orange portion of the visible spectrum (around 600 nm). A spectrophotometer or colorimeter is an instrument that consists of a light source and a diffraction grating that separates the light into different wavelengths. The light passes through a cuvette,(small test tube) used to hold the sample that is being analyzed. After the light passes through the sample, a detector measures how much light passes through the sample. Most spectrophotometers are



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designed to measure absorbance in the visible or the ultraviolet region of the electromagnetic spectrum. A spectrophotometer can be used to scan the sample over a wide range of wavelength values. In the spectrum shown, the maximum value for absorbance is at a wavelength of approximately 500 nm. When using a spectrophotometer to analyze a sample, it is important to set the instrument to a particular wavelength in which there is a relatively high absorbance value. Based on the absorbance spectrum shown above, any wavelength that is between 400 nm and 600 nm would be an appropriate setting for the instrument. Before inserting a sample into the spectrophotometer, a blank sample of water is inserted to calibrate the instrument. The absorbance value for the blank sample is set to zero. When the samples are analyzed in the instrument, the absorbance values are compared to that of the blank sample. The cuvette should be filled with enough solution so that the beam of light passes entirely through the solution and not through empty air space. A standard cuvette normally has a diameter (path length) of 1.0 cm. Cuvettes are normally wiped with a lint-free tissue to remove any moisture or fingerprints before they are placed in the instrument. Any smudges, dirt, or fingerprints on the transparent sides of the cuvette will scatter or block the light, preventing it from reaching the detector. This will result in a measured value for absorbance that will be too high.

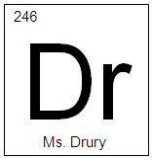


5. What is a cuvette and what is it used for?

6. Why is the cuvette first filled with water?

7. If a solution is blue, why can't a wavelength of 460nm be used to measure the sample?

8. What does pathlength mean?



Solutions Webquest Activity

1. Open a browser and go to https://phet.colorado.edu/
2. Go to the "Beer's Law Lab" simulation under the "Chemistry simulations". Click on "Beer's Law."
3. Choose the Cobalt (II) Nitrate, $Co(NO_3)_2$ solution to start and set the concentration to 0 mM.
4. Use the ruler to make sure the container has a "path length" of 1.0cm.
5. The wavelength of the light on the left should be at a preset wavelength of 549nm and if you click the red button to turn it on the color will shine green through your solution.
6. Record the transmittance and absorbance reported by the meter on the right.

   T= A=

7. Change the concentration of the Cobalt (II) Nitrate to 100mM. Record the transmittance and absorbance and color change.

   T= A=

8. Using your previous answers, what do you think transmittance and absorbance means in this simulation?

9. Hypothesize what you think will happen to the transmittance and absorbance if you change the concentration to 400 mM. Then change the concentration and record the actual values.
   a. Did the transmittance increase, decrease, or remain the same?

   b. Did the absorbance increase, decrease, or remain the same?

   c. What did you notice about the color?

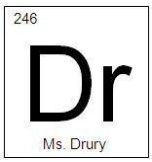
10. Choose another salt, Nickel (II) chloride, $NiCl_2$, from the drop down list. What happened to the wavelength and color of light coming from the device on the left?

11. At 0 mM, are the values different from cobalt (II) nitrate?

12. At 100mM, are the values different from cobalt (II) nitrate?

13. Play around with more salts and concentrations.
    a. What is the relationship between concentration and absorbance?

    b. What is the relationship between concentration and transmittance?



c. Record the color of the solution as well as the wavelength and color of light used for each solution.

Solution	Color of solution	color of light	wavelength light
$Co(NO_3)_2$			
$CoCl_2$			
$K_2Cr_2O_7$			
K_2CrO_4			
$NiCl_2$			
$CuSO_4$			
$KMnO_4$			

d. What do you notice about the color of the solution versus the color of the light used?

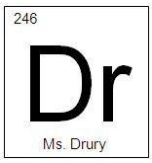
14. Reset the simulation to 100mM of $NiCl_2$.Use the ruler to change the width of the container. This is called the path length of the light through the solution. What is the relationship between pathlength and **absorbance**?

15. The Beer's Law formula is $A=\varepsilon bc$. Use your reference table to define each letter.

16. Choose copper sulfate, $CuSO_4$ at 100mM. Create a path length of 1cm. Record the absorbance. Calculate the molar absorptivity (with units) of copper sulfate using the concentration and path length.

17. Use Beer's law formula to calculate the Absorbance of light for $CuSO_4$ at 200mM. Show work. Then verify with the simulation.

18. Use Beer's law formula to calculate the absorbance of $CuSO_4$ at 150mM. Show work. Then verify with the simulation.



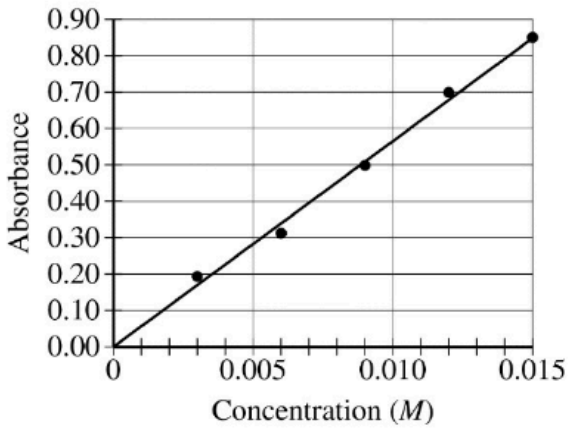
19. What values are constant for a solution in Beer's Law?

20. Will the values you selected above be constant for a new solute?

21. In a lab, a student used the formula below to calculate various absorbance values and concentrations. Explain how this formula is applicable as long as the solute stays the same?

$$\frac{A_1}{C_1} = \frac{A_2}{C_2}$$

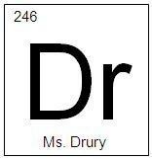
22. An ore containing cobalt and some inert materials is analyzed in a laboratory. A spectrophotometer is used to determine the mass percent of cobalt present in the ore sample. Solutions of known concentration of $Co^{2+}(aq)$ are prepared, and the absorbance of each solution is measured at the wavelength of optimum absorbance. The data are used to create a calibration plot, which is shown above. A sample of the ore with a mass of 0.630 g is completely dissolved in concentrated $HNO_3(aq)$. The mixture is diluted with water to a final volume of 50.00 mL. Assume that all of the cobalt present in the ore sample is converted into $Co^{2+}(aq)$. The absorbance of a sample of this solution is measured as 0.74.



a. What is the concentration of $Co^{2+}(aq)$ in this solution?

b. Calculate the number of moles of $Co^{2+}(aq)$ in the 50.00 mL solution.

c. Calculate the mass percent of Co in the 0.630 g sample of the ore.



Solutions AP Multiple Choice

1. When 70. milliliters of 3.0-molar Na_2CO_3 is added to 30. milliliters of 1.0-molar $NaHCO_3$ the resulting concentration of Na^+ is…
(A) 2.0 M (B) 2.4 M (C) 4.0 M (D) 4.5 M (E) 7.0 M

2. A student needs to prepare 2.00 liters of 0.100-molar KIO_3. The proper procedure is to…

(A) 42.8 grams of KIO_3 and add 2.00 kilograms of H_2O

(B) 42.8 grams of KIO_3 and add H_2O until the final homogeneous solution has a volume of 2.00 liters

(C) 21.4 grams of KIO_3 and add H_2O until the final homogeneous solution has a volume of 2.00 liters

(D) 42.8 grams of KIO_3 and add 2.00 liters of H_2O

(E) 21.4 grams of KIO_3 and add 2.00 liters of H_2O

3. The weight of H_2SO_4 (molecular weight 98.1) in 50.0 milliliters of a 6.00-molar solution is…
(A) 3.10 grams (B) 12.0 grams (C) 29.4 grams (D) 294 grams (E) 300. grams

4. How many milliliters of 11.6-molar HCl must be diluted to obtain 1.0 liter of 3.0-molar HCl?
(A) 3.9 mL (B) 35 mL (C) 260 mL (D) 1,000 mL (E) 3,900 mL

5. Which of the following does NOT behave as an electrolyte when it is dissolved in water?
(A) CH_3OH (B) K_2CO_3 (C) NH_4Br (D) HI (E) Sodium acetate, CH_3COONa

6. Given that a solution is 5 percent sucrose by mass, what additional information is necessary to calculate the molarity of the solution?
I. The density of water II. The density of the solution III. The molar mass of sucrose
(A) I only (B) II only (C) III only (D) I and III (E) II and III

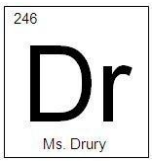
7. If 87 grams of K_2SO_4 (molar mass 174 grams) is dissolved in enough water to make 250 milliliters of solution, what are the concentrations of the potassium and the sulfate ions?

	$[K^+]$	$[SO_4^{2-}]$
(A)	0.020 M	0.020 M
(B)	1.0 M	2.0 M
(C)	2.0 M	1.0 M
(D)	2.0 M	2.0 M
(E)	4.0 M	2.0 M

8. A yellow precipitate forms when 0.5 M $NaI_{(aq)}$ is added to a 0.5 M solution of which of the following ions?
(A) $Pb^{2+}_{(aq)}$ (B) $Ba^{2+}_{(aq)}$ (C) $PO_4^{3-}{}_{(aq)}$ (D) $SO_4^{2-}{}_{(aq)}$ (E) $OH^-_{(aq)}$

9. When 100 mL of 1.0 M Na_3PO_4 is mixed with 100 mL of 1.0 M $AgNO_3$, a yellow precipitate forms and $[Ag^+]$ becomes negligibly small. Which of the following is a correct listing of the ions remaining in solution in order of increasing concentration?
(A) $[PO_4^{3-}] < [NO_3^-] < [Na^+]$ (B) $[PO_4^{3-}] < [Na^+] < [NO_3^-]$ (C) $[NO_3^-] < [PO_4^{3-}] < [Na^+]$
(D) $[Na^+] < [NO_3^-] < [PO_4^{3-}]$ (E) $[Na^+] < [PO_4^{3-}] < [NO_3^-]$

10. The volume of distilled water that should be added to 10.0 mL of 6.00 M $HCl_{(aq)}$ in order to prepare a 0.500 M $HCl_{(aq)}$ solution is approximately…
(A) 50.0 mL (B) 60.0 mL (C) 100. mL (D) 110. mL (E) 120. mL



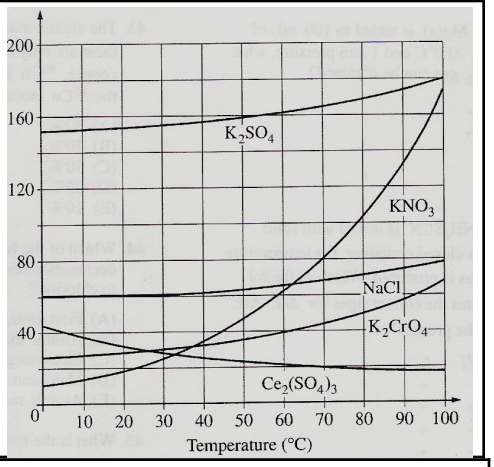
11. Substances X and Y that were in a solution were separated using the technique of fractional crystallization. This fractional crystallization is possible because substances X and Y have different…
(A) boiling points (B) melting points (C) densities (D) crystal colors (E) solubilities

12. What is the mole fraction of ethanol, C_2H_5OH, in an aqueous solution that is 46 percent ethanol by mass? (The molar mass of C_2H_5OH is 46 g; the molar mass of H_2O is 18 g.)
(A) 0.25 (B) 0.46 (C) 0.54 (D) 0.67 (E) 0.75

13. Approximately what mass of $CuSO_4 \cdot 5H_2O$ (250 g mol^{-1}) is required to prepare 250 mL of 0.10 M copper(II) sulfate solution?
(A) 4.0 g (B) 6.2 g (C) 34 g (D) 85 g (E) 140 g

14. A 0.10 M aqueous solution of sodium sulfate, Na_2SO_4, is a better conductor of electricity than a 0.10 M aqueous solution of sodium chloride, NaCl. Which of the following best explains this observation?
(A) Na_2SO_4 is more soluble in water than NaCl is.
(B) Na_2SO_4 has a higher molar mass than NaCl has.
(C) To prepare a given volume of 0.10 M solution, the mass of Na_2SO_4 needed is more than twice the mass of NaCl needed.
(D) More moles of ions are present in a given volume of 0.10 M Na_2SO_4 than in the same volume of 0.10 M NaCl.
(E) The degree of dissociation of Na_2SO_4 in solution is significantly greater than that of NaCl.

15. On the basis of the solubility curves shown, the greatest percentage of which compound can be recovered by cooling a saturated solution of that compound from 90°C to 30°C ?
(A) NaCl (B) KNO_3
(C) K_2CrO_4 (D) K_2SO_4 (E) $Ce_2(SO_4)_3$



16. If 200. mL of 0.60 M $MgCl_{2(aq)}$ is added to 400. mL of distilled water, what is the concentration of $Mg^{2+}_{(aq)}$ in the resulting solution? (Assume volume are additive.)
(A) 0.20 M (B) 0.30 M (C) 0.40 M (D) 0.60 M (E) 1.2 M

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