Unit 3

9/12-13

Agenda

- Goal setting bellwork
- Overview of unit 3
- Start unit 3.1-3.3
    - Lab demo
    - Notes
    - Practice problems

Unit 2

- Even met both their goals
- Allergies?
- Cookies on Friday

Unit 3- Intermolecular forces

- Higher percentage on AP test
- More topics (13)
- We will break the unit into 4 subunits
    - 3.1-3.3 (Solids, liquids, gases)
    - 3.4-3.6 (Gases)
    - 3.7-3.10 (Solutions)
    - 3.11-3.13 (Light)

TOPIC 3.1: INTERMOLECULAR FORCES

Enduring Understanding	
SAP-5	Intermolecular forces can explain the physical properties of a material.
Learning Objective	
SAP-5A	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.

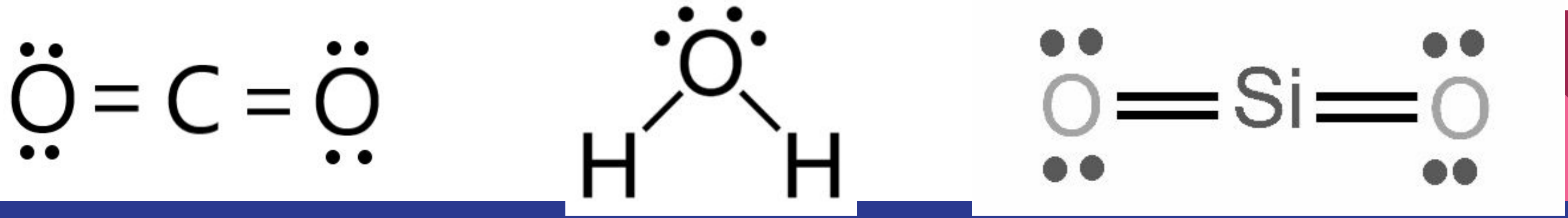
http://lab.concord.org/embeddable.html#interactives/sam/intermolecular-attractions/2-comparing-dipole-dipole-to-london-dispersion.json

IMFs

Intermolecular forces, IMFs, are the attractions BETWEEN atoms, ions or molecules that hold them together.

There are six different types of IMF. They differ in the arrangement of the electrons.

Intermolecular forces explain why carbon dioxide, CO_2, is a gas, water, H_2O is a liquid and silicon dioxide, SiO_2, is a solid, even though they are all only made from three atoms.



IMFs

To determine which type of IMF a substance will exhibit you need to look at the types of atoms that are present in the sample.

Increasing strength →

IMF	**London Dispersion**	**Dipole-Dipole**	**Hydrogen Bonding (a strong dipole-dipole)**	**Ionic**	**Metallic**	**Network Covalent**
Occurs Between	All Substances	Polar Molecules Asymmetrical molecules	Polar Molecules with H attached to F, O, or N	Metals with non-metals	Only metals	KNOW THESE: ● $C_{(graphite)}$ ● $C_{(diamond)}$ ● SiO_2 ● SiC

LDFs

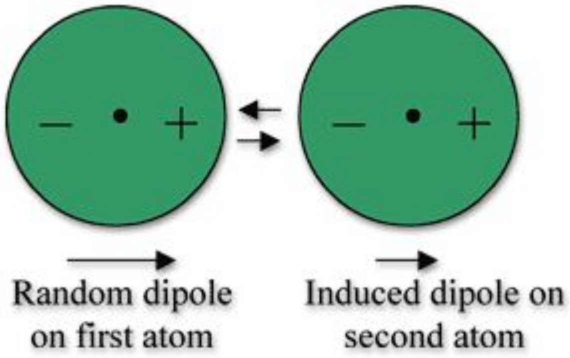
London dispersion forces occur in all substances.

They are the only type of IMF present in nonpolar molecules.

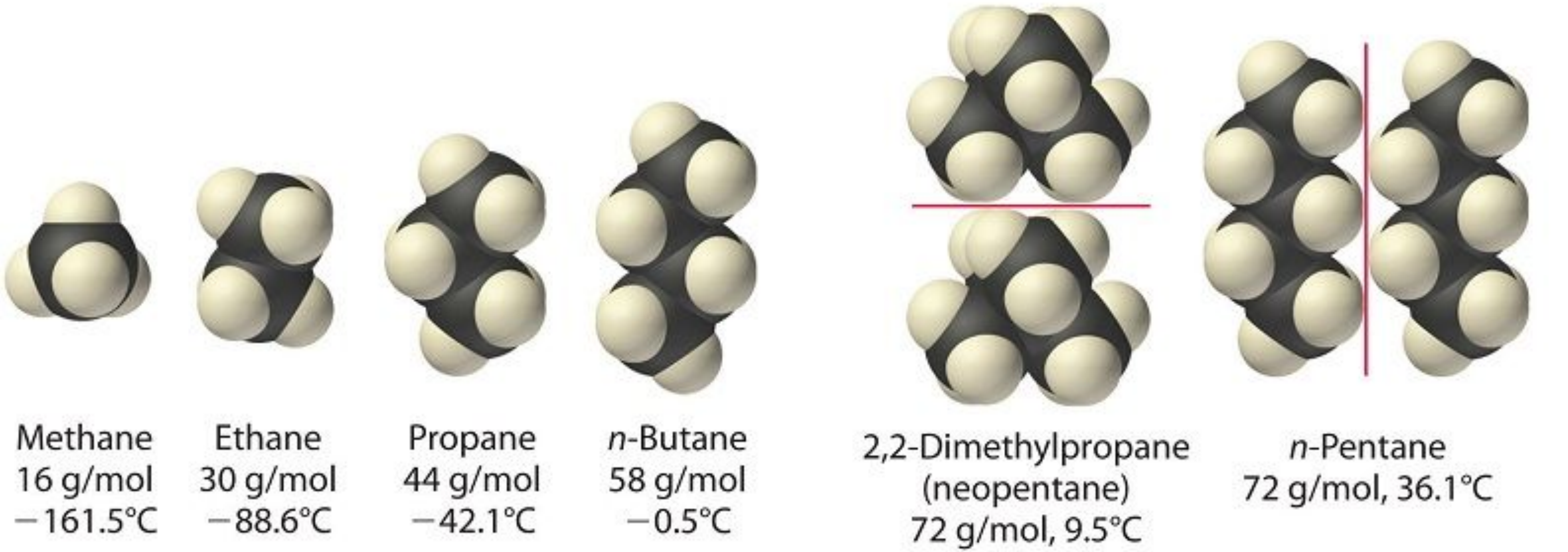
LDFs are caused by the motion of electrons. As the electrons move around with the atom there is a chance that they are unequally distributed. When this happens it causes the atom or molecule to have a **temporary dipole**.

The temporary dipole in one atom or molecule induces a dipole in the next atom or molecule and they become attracted to each other.

Short lived - LDF is a weak force.



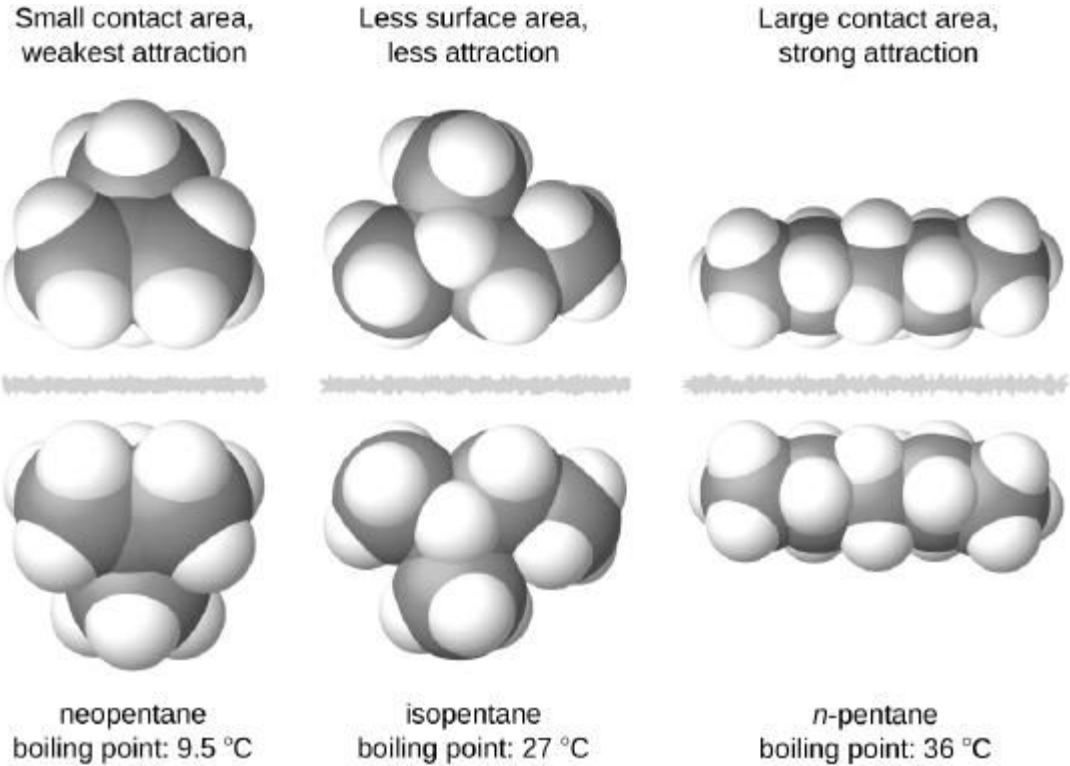
Strength of LDF is dependent on



(a) Increasing mass and boiling point

(b) Increasing surface area and boiling point

Strength of LDF is dependent on



Strength of LDF is dependent on

Mass

Surface area

The ability of an atom to form a temporary or induced dipole is known as its **polarizability**.

LDFs and Polarizability

- London dispersion forces are stronger when an atom or molecule has more electrons.
- Large molecules will have stronger LDF because the increase in electrons increases their polarizability.
- Long, thin molecules have more surface area and are more polarizable while short, round molecules have less surface area and are less polarizable.
- pi bonding, the presence of double or triple bonds, increases LDF strength.

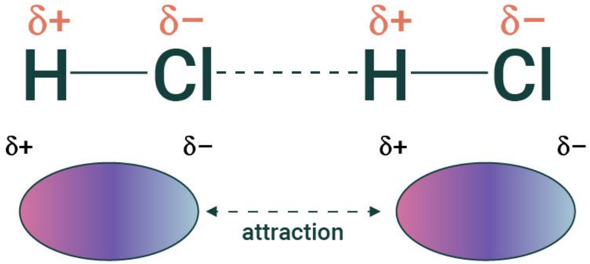
The more polarizable the molecule, the stronger the LDF will be.

Dipole-Dipole

Dipole-dipole interactions occur between molecules that have a permanent dipole (polar molecules.)

Dipole-dipole interactions are similar to LDFs, except that the **dipoles are permanent** and so the attractions are generally greater. The partially positive end of one molecule is attracted to the partially negative end of a different molecule, and so on throughout the system.

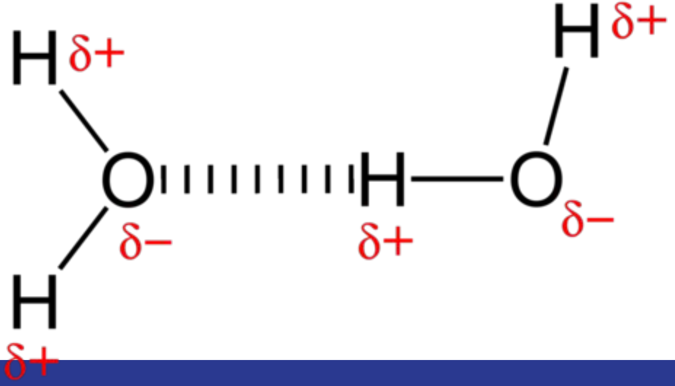
The more polar the molecules, the greater the attraction.



Hydrogen Bonding - not a bond

Hydrogen bonding is an IMF that occurs in molecules with Hydrogen attached directly to **F, O, or N**. (Hydrogen bonding is FON!)

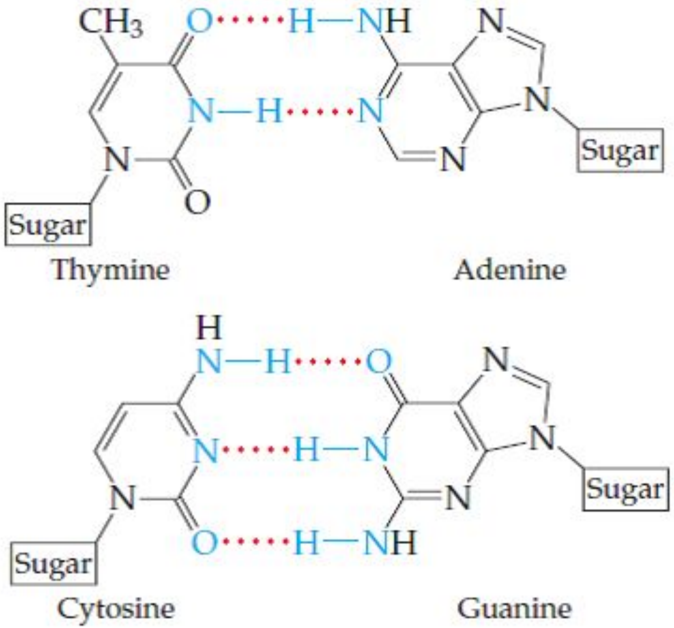
This difference in electronegativity creates a large molecular dipole, and this dipole causes a strong attraction to the dipole on the molecules surrounding it.



Throwback to BIO

In DNA, the base pairs GC and AT link together via hydrogen bonds. This phenomenon occurs in many biological molecules.

GC has 3 sites of H-bonding, while AT has 2 sites of H-bonding. This is why DNA sequences with more Cytosine and Guanine “melt” at a higher temperature.



Lab

Intermolecular Forces and Heat of Evaporation

Background

In today's lab we will examine the attractive forces that hold molecules together and the disruptive forces that break them apart. The forces between molecules that hold molecules together are called Intermolecular Forces (IMF) and are comprised of London dispersion forces (LDF), dipolar forces, and hydrogen bonding (H-bonding). The forces that break molecules apart are related to the temperature of the object. You will explore the effect of polarity on the strength of IMF and how the IMF is related to physical properties such as melting point, boiling point, state of matter, and viscosity of liquids.

All material is held together by attractive forces but there is always some disruptive force present that can break it apart. When an object is a solid at a given temperature it means that the attractive forces must be greater than the disruptive forces. When something is a gas at any temperature it means that the disruptive forces much be much greater than the attractive forces. Finally when the forces of disruption and attraction are on about the same level the substance will exist as a liquid. There are only two ways that elements are held together: sharing of electrons (as in covalent bonding) or by plus to minus charge or partial charge interactions (as in ionic compounds or intermolecular forces).

In this lab, we will be using the process of *evaporative cooling* to see if the amounts of predicted intermolecular forces in various liquids correlate with a trend in boiling/evaporation point.

Agenda

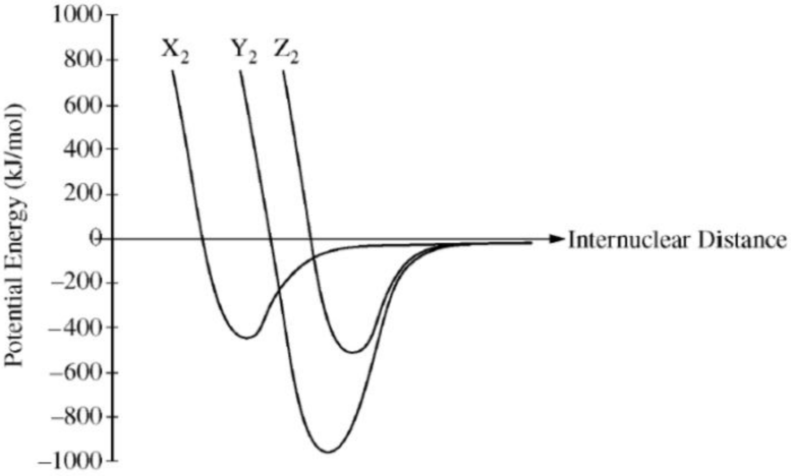
- Bellwork
- Review Unit 2 Test
- Set class goal
- 3.1 Practice questions
    - Each pair will answer one question and present the question to the whole class
- 3.2 Notes and practice questions

Bellwork

Unit 2 Test Review

- Look over your unit 2 test
- The whiteboards have commonly missed questions and solutions
- Fill out your test reflection sheet

Unit 2 Most missed MCQ



The potential energy as a function of internuclear distance for three diatomic molecules, X_2, Y_2, and Z_2, is shown in the graph above. Based on the data in the graph, which of the following correctly identifies the diatomic molecules, X_2, Y_2, and Z_2? (2.2)

	X_2	Y_2	Z_2
A)	H_2	N_2	O_2
B)	H_2	O_2	N_2
C)	N_2	O_2	H_2
D)	O_2	H_2	N_2

Test Review

Ionic, Metallic, and Network Covalent `(Not *technically* IMFs)`

IONIC

- occur between metal and nonmetal atoms
- electrostatic attraction
- stronger when the charges are larger and the ions are smaller (Coulomb's Law)

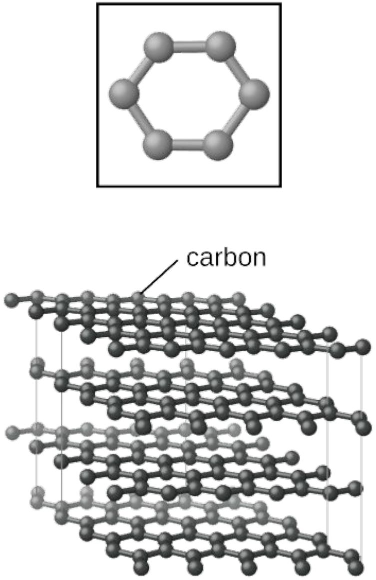
METALLIC

- between metal atoms
- Pure substance = one type of metal
- Alloy = different types of metals
- The metallic attractions are due to multiple metallic positive cores being attracted to a delocalized sea of valence electrons.
- Stronger when there are smaller metallic positive cores and when there are more valence electrons.

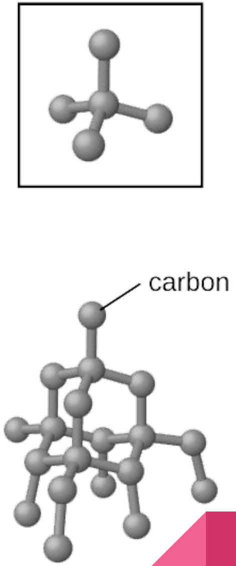
Ionic, Metallic, and Network Covalent `(Not *technically* IMFs)`

COVALENT NETWORK

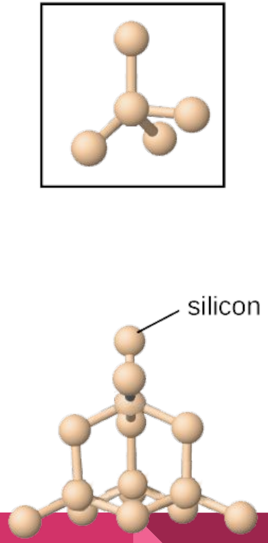
- between carbon atoms, silicon atoms, and carbon with oxygen or silicon, and silicon with oxygen.
- multiple covalent bonds forming a large macromolecule.
- stronger when there are more covalent bonds formed.



graphite



diamond



silicon

Properties of Substances Due to IMFs

Property	**Effect of increasing IMF**
Boiling/melting/freezing/etc. point	increases
Enthalpy of fusion/vaporization etc.	increases
Vapor pressure	decreases
Viscosity	increases
Surface Tension	increases
Lattice Energy	increases
Solubility (this depends a lot on the solvent and solute polarity)	decreases

Together

1. Pentane, C_5H_{12} and Octane C_8H_{18}, are both hydrocarbons.
    a. Identify the type of intermolecular forces present.

    b. Predict the substance that would have a higher vapor pressure and justify your claim.

You Do, We Review

2. The boiling point of HF is 293K whereas the boiling point of F_2 is 85K.
    a. Determine the type(s) of intermolecular forces present.
    b. Explain the difference in boiling points based on their intermolecular forces.

TOPIC 3.2: PROPERTIES OF SOLIDS

Enduring Understanding	
SAP-5	Intermolecular forces can explain the physical properties of a material.
Learning Objective	
SAP-5B	Explain the relationship among the macroscopic properties of a substance, the particulate-level structure of the substance, and the interactions between these particles.

TOPIC 3.3: SOLIDS, LIQUIDS, & GASES

Enduring Understanding	
SAP-6	Matter exists in three states: Solid, Liquid, and Gas, and their differences are influenced by variances in spacing and motion of the molecules.
Learning Objective	
SAP-6A	Represent the differences between solid, liquid, and gas phases using a particulate-level model.

Stations

1: 3.2 Video 2

2: 3.2 Read through notes in packet

3: 3.2 Practice problems in packet

4: 3.3 Video

5: 3.3 Read through notes in packet

6: 3.3 Practice problems in packet

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Solids

Many properties of liquids and solids are determined by the strengths and types of intermolecular forces present.

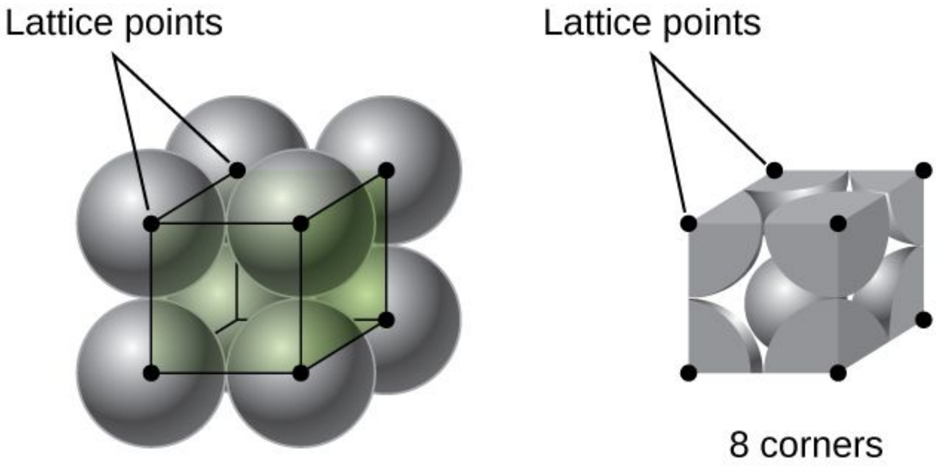
There are 4 basic types of solids:

- Ionic Solids
- Covalent Network Solids
- Molecular Solids
- Metallic Solids

Lattice

A lattice is a 3-dimensional system of points.

Lattice points are the positions of the centers of the components of a solid (atoms, ions, or molecules).



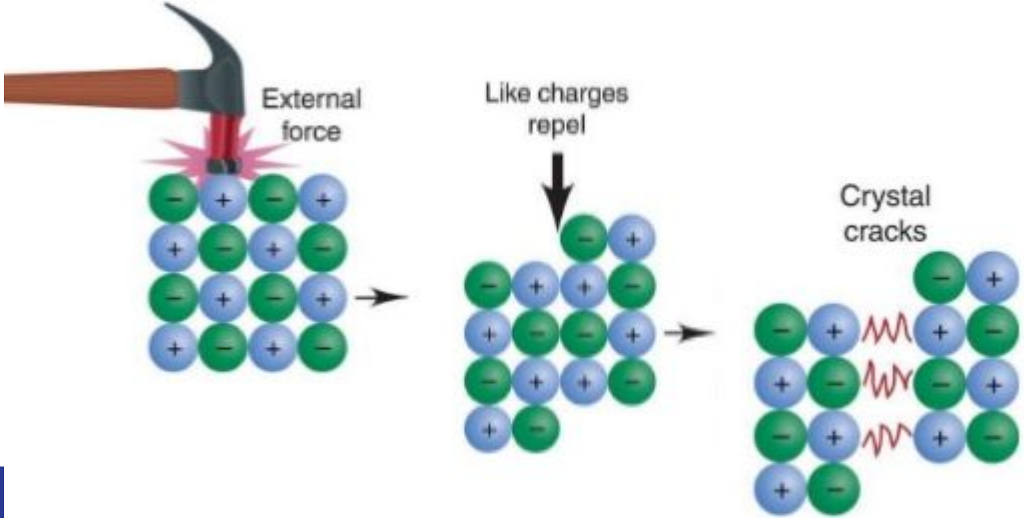
Simple cubic lattice cell

Ionic Solids

- Ions at the points of the lattice that describes the structure of the solid.
- Due to strong interactions between ions, ionic solids tend to have **low vapor pressures**, **high melting points**, and **high boiling points**.
- Electrostatic attraction between the positive and negative ions can be described by Coulomb's Law.
- According to Coulomb's Law, smaller ions &/or ions with higher charges will have stronger attractions between the ions resulting in higher lattice energy values.

Ionic Solids

- **Brittle** due to the repulsion of like charges caused when one layer slides across another layer.
- **Conduct electricity when the ions are mobile**, as when the ionic solid is melted (molten) or dissolved in a solvent.

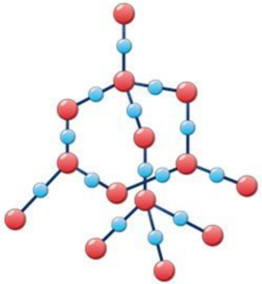


Network Covalent Solids

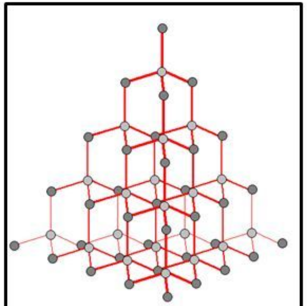
- Covalent network solids have atoms at the lattice points with strong covalent bonds
- These are only formed from nonmetals: elemental (e.g., diamond, graphite) or binary compounds of two nonmetals (e.g., silicon dioxide and silicon carbide).
- Example of lattice points in covalent network solids:
  - Carbon (diamond & graphite)
  - silicon dioxide (SiO_2)
  - silicon carbide (SiC)

Network Covalent Solids

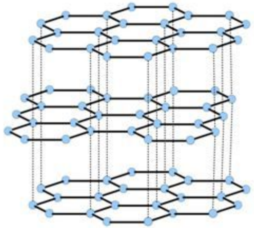
- Covalent network solids are typically characterized by hardness, strength, and high melting points.
- Three-dimensional network solids are rigid and brittle because the covalent bond angles are fixed.



Silicon dioxide (silica)



Diamond



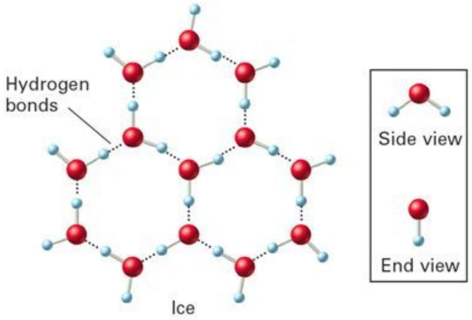
Graphite

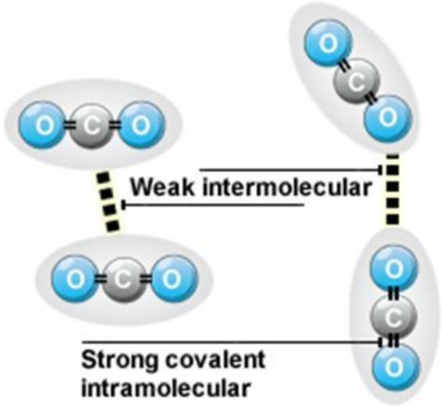
Molecular Solids

- Composed of distinct, individual units of covalently-bonded molecules attracted to each other through relatively weak intermolecular forces.
- Molecules at lattice points
- The molecules used to form molecular solids are composed of nonmetals atoms covalently bonded together.
- Generally have a low melting points
- They do not conduct electricity because their valence electrons are tightly held within the covalent bonds and the lone pairs of each constituent molecule.

Molecular Solids

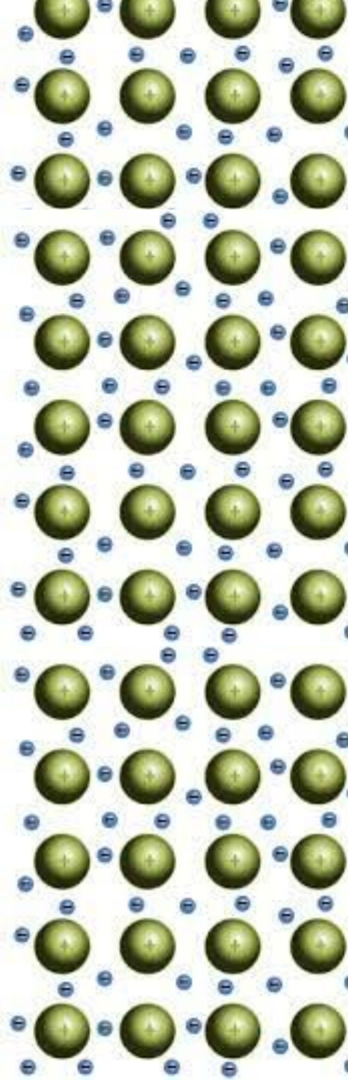
- Sometimes composed of very large molecules or polymers.
- Examples:
    - Water (ice) - H_2O
    - Carbon Dioxide (dry ice) - CO_2


H_2O - ice


Dry Ice - CO_2

Metallic Solids

- Metallic solids consist of metallic crystals with spherical metal atoms packed together and bonded to each other equally in all directions.
- A metallic bond is considered to be a close packed lattice of positive cores surrounded by a “sea” of moving electrons.
- The movement of the electrons cause metals to be good conductors of heat and electricity
- The close packed positive cores allow for metals to be malleable and ductile, due to the ease with which the metal ions can rearrange their structure.

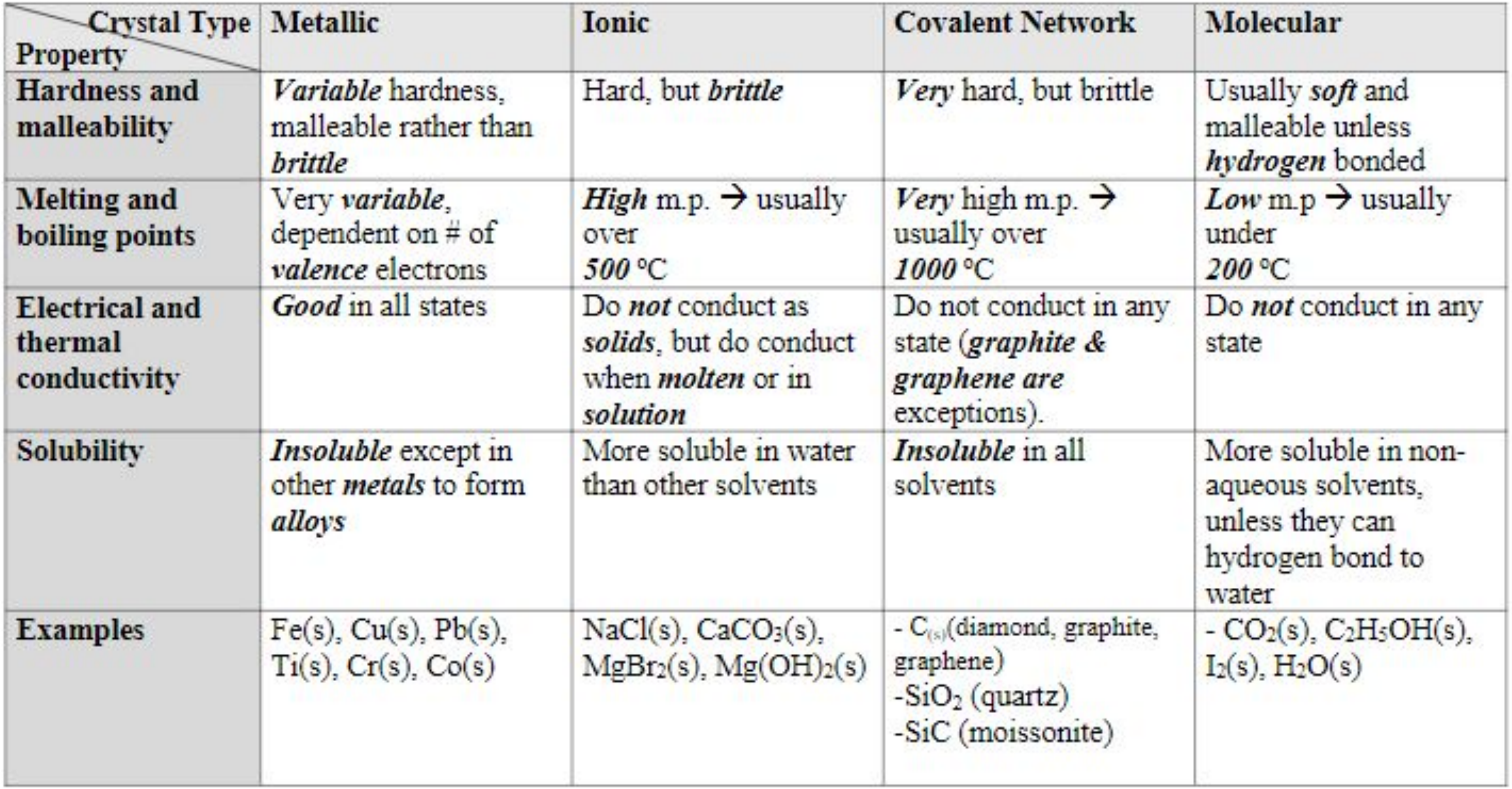


Alloys - mixture of Metals

Alloys keep a “sea” of electrons so they are still conductive.

- Substitutional Alloy:
  - Since the atoms have similar radii, the density of the alloy falls between the density values of the two metals
  - Alloy remains malleable and ductile
  - Example: 14k gold (Ag atoms replace some of the Au atoms)
- Interstitial Alloy:
  - Since the atoms have different radii, the interstitial atoms tend to make the lattice more dense
  - Decreased malleability and ductility
  - Example: steel (C atoms fit between Fe atoms)

Summary of Properties



Crystal Type **Property**	**Metallic**	**Ionic**	**Covalent Network**	**Molecular**
Hardness and malleability	***Variable*** hardness, malleable rather than ***brittle***	Hard, but ***brittle***	***Very*** hard, but brittle	Usually ***soft*** and malleable unless ***hydrogen*** bonded
Melting and boiling points	Very ***variable***, dependent on # of ***valence*** electrons	***High*** m.p. → usually over ***500*** °C	***Very*** high m.p. → usually over ***1000*** °C	***Low*** m.p → usually under ***200*** °C
Electrical and thermal conductivity	***Good*** in all states	Do ***not*** conduct as ***solids***, but do conduct when ***molten*** or in ***solution***	Do not conduct in any state (***graphite & graphene are*** exceptions).	Do ***not*** conduct in any state
Solubility	***Insoluble*** except in other ***metals*** to form ***alloys***	More soluble in water than other solvents	***Insoluble*** in all solvents	More soluble in non-aqueous solvents, unless they can hydrogen bond to water
Examples	Fe(s), Cu(s), Pb(s), Ti(s), Cr(s), Co(s)	NaCl(s), $CaCO_3$(s), $MgBr_2$(s), $Mg(OH)_2$(s)	- $C_{(s)}$(diamond, graphite, graphene) -SiO_2 (quartz) -SiC (moissonite)	- CO_2(s), C_2H_5OH(s), I_2(s), H_2O(s)

Together

1. CO_2 melts at -78°C and SiO_2 melts at 1,650°C.
   a. Identify the type of solid each substance forms.

   b. Justify the difference in melting points.

You Do, We Review

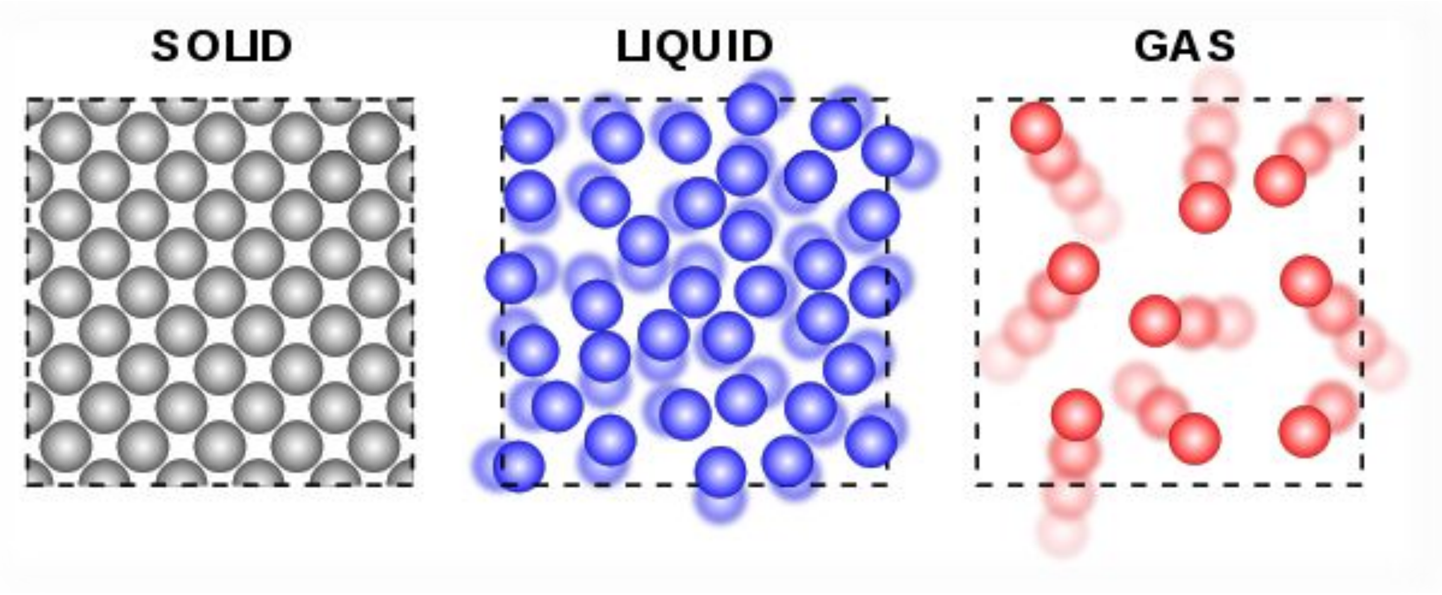
2. Classify Cl_2, Ni, BN, and FeS as ionic, molecular, covalent network, or metallic solids and arrange them in order of increasing melting points.

TOPIC 3.3: SOLIDS, LIQUIDS, & GASES

Enduring Understanding	
SAP-6	Matter exists in three states: Solid, Liquid, and Gas, and their differences are influenced by variances in spacing and motion of the molecules.
Learning Objective	
SAP-6A	Represent the differences between solid, liquid, and gas phases using a particulate-level model.

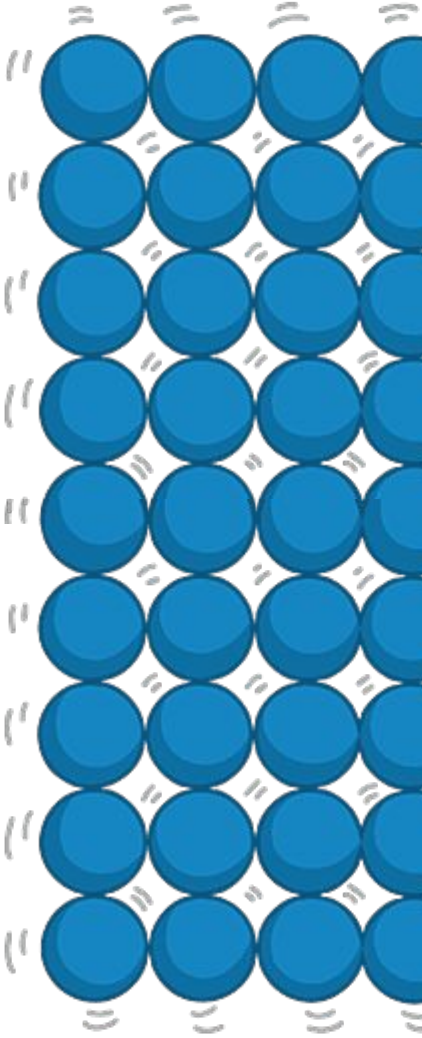
Particulate Diagrams

Particulate diagrams can be used to show the different properties between solids, liquids, and gases.

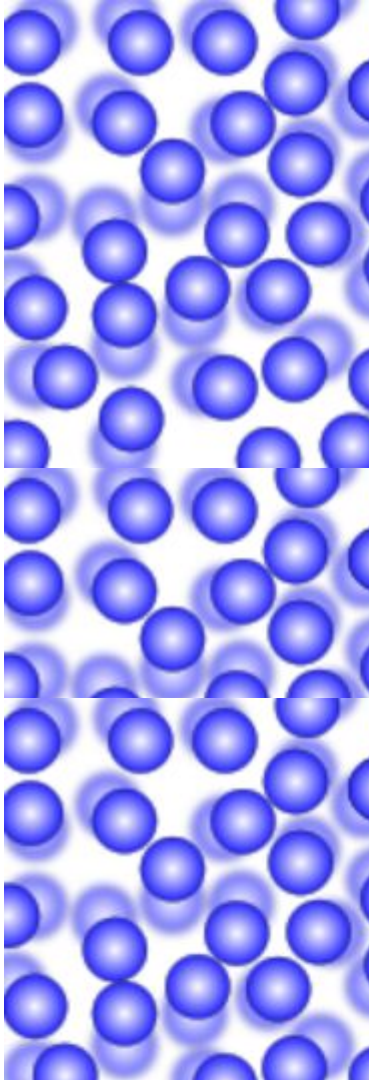


Solids

- Particles do not have enough kinetic energy to move freely, only vibrating.
- Not much space between particles because the IMFs hold the particles in place.
- Rigid, fixed volume and shape
- Solids cannot be compressed
- Location of the particles can either be crystalline or amorphous.
  - Crystalline solids have a repeating three-dimensional structure.
  - Amorphous solids the particles are disordered.
  - In both cases, the molecules do not have enough kinetic energy to move freely.



LIQUIDS



- particles are still close together
- molecules have gained enough kinetic energy for them to slide and roll around
- liquids take the shape of their container
- liquids cannot be compressed
- Solids and liquids of the same substance are often approximately the same size because the particles are close together in both states.
- The temperature range where a substance is a liquid is determined by the strength of the intermolecular forces
- The temperature range where a substance is a liquid can be compared by analyzing the strength of these forces.

Gases

- Particles have gained enough kinetic energy to overcome the intermolecular forces holding the particles together.
- Particles are far apart and moving quickly.
- Gases can be easily compressed
- Take the shape of their container
- No regular arrangement of the particles.
- The frequency and strength of collisions between particles and the container is dependent on the temperature, pressure, and volume.

TOGETHER

1. Which of the following is the best explanation for the fact that most liquids cannot be easily compressed:
   a. Liquid molecules are in constant motion
   b. Liquid molecules are relatively close together
   c. Liquid molecules have varying densities
   d. Liquid molecules have a fixed volume
   e. Liquid molecules move slower as temperature decreases

You do, We Review

2. Draw a particle diagram to represent water in a puddle as it starts to evaporate.

Unit 3 Part 2

Agenda

- Goal setting bellwork
- Phenomena
- PhET lab
- 3.4 Notes and practice problems

TOPIC 3.4: IDEAL GAS LAW

Enduring Understanding

SAP-7	Gas properties are explained macroscopically - using the relationships among pressure, volume, temperature, moles, gas constant - and molecularly by the motion of the gas.

Learning Objective

SAP-7A	Explain the relationship between the macroscopic properties of a sample of gas or mixture of gases using the ideal gas law.

GAS LAWS

Gas properties are explained macroscopically - using the relationships among pressure, volume, temperature, moles, and the gas constant.

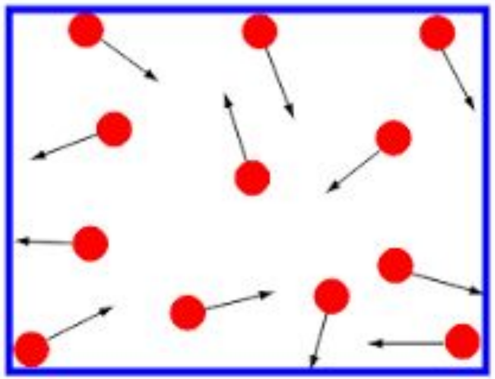
P = pressure in atm/kPa/torr/mmHg etc.

V = volume in L

n = moles of gas

R = gas constant

T = temperature **in Kelvin** (Celcius + 273.15)



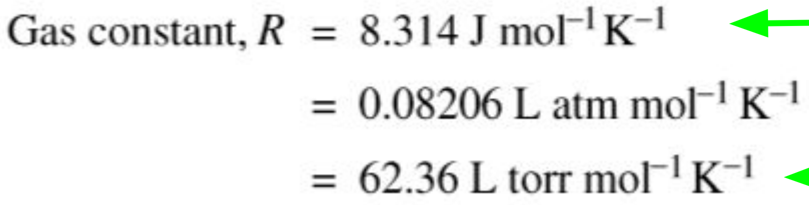
Gas Laws

Law	Relationship	Formula
Boyle's Law	At constant T and n, P and V are inversely proportional.	$P_1V_1 = P_2V_2$
Charles' Law	At constant P and n, V and T are proportional.	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$
Gay-Lussac's Law	At constant V and n, P and T are proportional.	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$
Combined Gas Law	At constant n, relates P, V, and T	$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
Avogadro's Law	At constant T and P, n is proportional to V	$\frac{V_1}{n_1} = \frac{V_2}{n_2}$

Ideal Gas Law

$$PV = nRT$$

There are several R values listed on the equation sheet, they differ in the units for pressure.



Gas constant, $R = 8.314 \text{ J mol}^{-1}\text{K}^{-1}$

$= 0.08206 \text{ L atm mol}^{-1}\text{ K}^{-1}$

$= 62.36 \text{ L torr mol}^{-1}\text{K}^{-1}$

J = L·kPa

When P is in atm

When P is in torr or mmHg

Ideal Gas Law

$$PV = nRT$$

The ideal gas formula can be used to solve for molar mass or density of a gas.

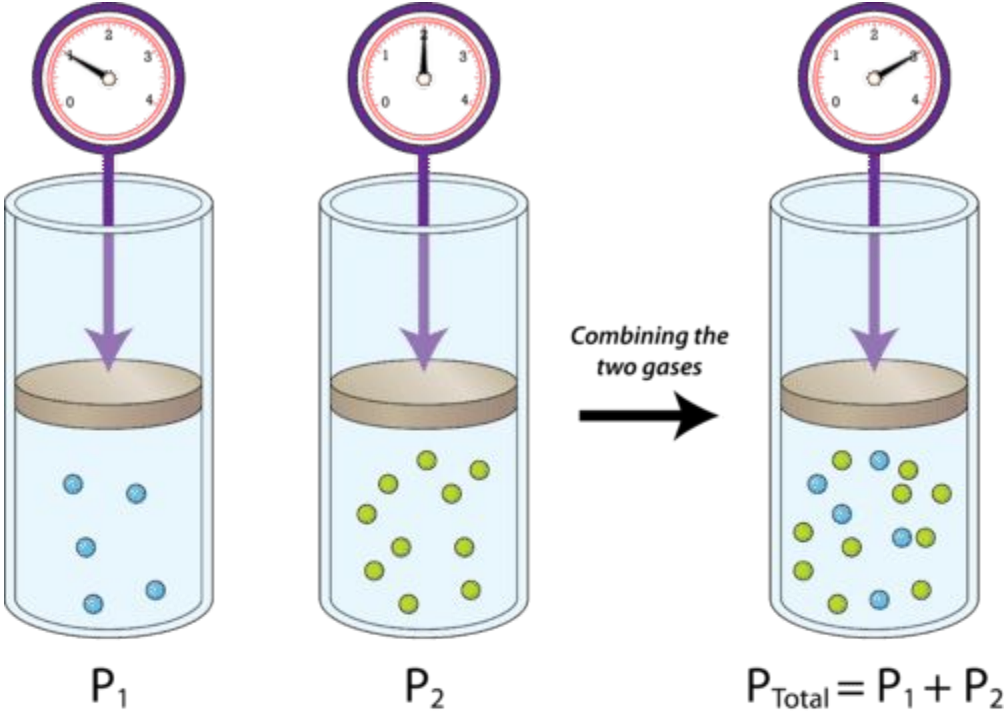
$$\text{Molar Mass (g/mol)} = \frac{\text{Density (g/L) x R x Kelvin Temperature}}{\text{Pressure}}$$

DALTON'S LAW OF PARTIAL PRESSURE

According to Dalton's Law of Partial Pressure, the sum of all the partial pressures of each gas in a mixture of gasses is equal to the total pressure.

$$P = P_A + P_B + P_C \ldots$$

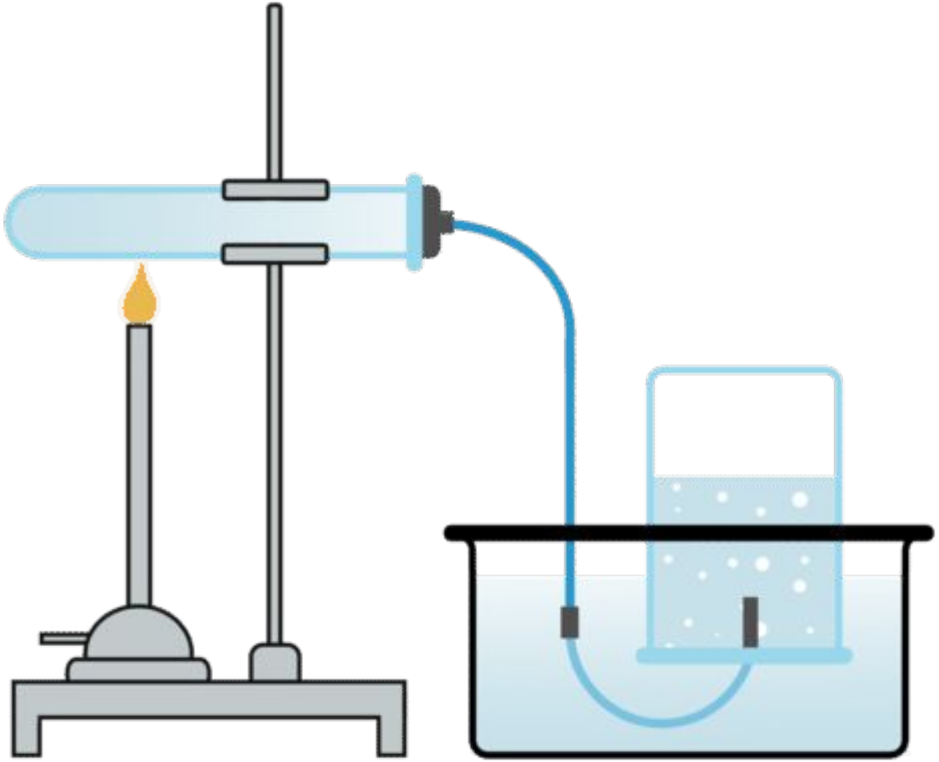
where A, B, and C are different gases.



Volume and temperature are constant

DALTON'S LAW OF PARTIAL PRESSURE

This is often used when gases are collected "over water" as shown in the image:



As the gas is produced, the water is displaced and the water in the inverted vessel will empty while the gas is collected.

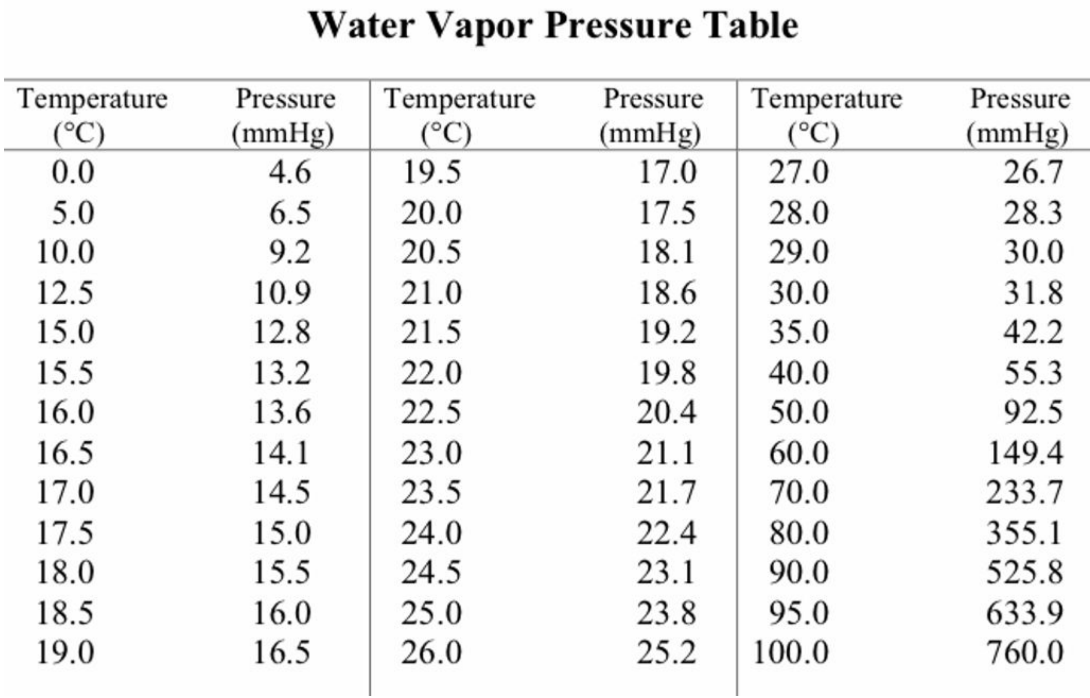
DALTON'S LAW OF PARTIAL PRESSURE

This method allows for the gas to be measured and minimizes the amount of room air that contaminates the sample.

As the gas travels through the water it will pick up water vapor which contributes to the gas produced.

The amount of water vapor that is picked up is a function of the temperature of the gas and can be subtracted using Dalton's law of partial pressures.

When the water is subtracted out the gas is described as "dry."



Water Vapor Pressure Table

Temperature (°C)	Pressure (mmHg)	Temperature (°C)	Pressure (mmHg)	Temperature (°C)	Pressure (mmHg)
0.0	4.6	19.5	17.0	27.0	26.7
5.0	6.5	20.0	17.5	28.0	28.3
10.0	9.2	20.5	18.1	29.0	30.0
12.5	10.9	21.0	18.6	30.0	31.8
15.0	12.8	21.5	19.2	35.0	42.2
15.5	13.2	22.0	19.8	40.0	55.3
16.0	13.6	22.5	20.4	50.0	92.5
16.5	14.1	23.0	21.1	60.0	149.4
17.0	14.5	23.5	21.7	70.0	233.7
17.5	15.0	24.0	22.4	80.0	355.1
18.0	15.5	24.5	23.1	90.0	525.8
18.5	16.0	25.0	23.8	95.0	633.9
19.0	16.5	26.0	25.2	100.0	760.0

MOLE FRACTIONS

Mole fraction:

$$\chi_A = \textbf{moles A/total moles}$$

If a mixture is 3.0 mol O_2 and 4.0 mol H_2, the mole fraction of O_2 = 3.0 moles/(3.0+4.0 moles) = 0.43

Partial pressure of gas A can be calculated:

$$\text{partial pressure}_A = \chi_A * \text{total pressure}$$

YOU DO, WE REVIEW

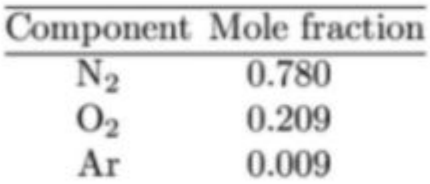
1. $$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

When heated strongly, solid calcium carbonate decomposes to produce solid calcium oxide and carbon dioxide gas, as represented by the equation above. A sample of $CaCO_3(s)$ is placed in a rigid 35 L reaction vessel from which all the air has been evacuated. The vessel is heated to 437°C at which time the pressure of $CO_2(g)$ in the vessel is constant at 1.00atm

Calculate the number of grams of $CaCO_3(g)$ that reacted to produce the carbon dioxide gas.

YOU DO, WE REVIEW

2. A basketball is left outside in winter when the temperature is -2.00 C, has a volume of 6.88 L and the pressure inside the basketball is 0.795 atm. How many moles of gas are in a basketball? What is the partial pressure of oxygen in the basketball?



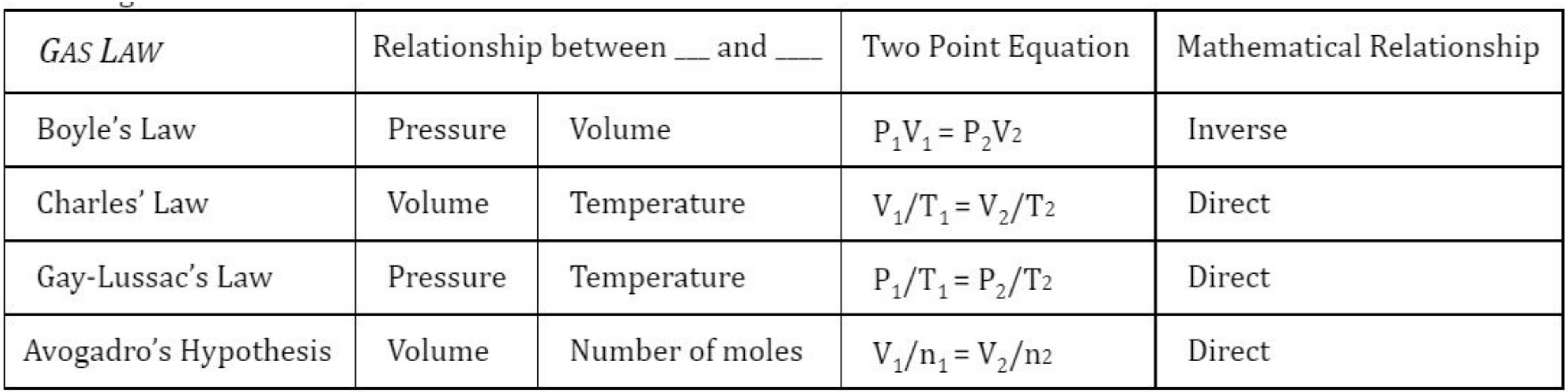
Component	Mole fraction
N_2	0.780
O_2	0.209
Ar	0.009

TOPIC 3.5: KINETIC MOLECULAR THEORY

Enduring Understanding	
SAP-7	Gas properties are explained macroscopically - using the relationships among pressure, volume, temperature, moles, gas constant - and molecularly by the motion of the gas.
Learning Objective	
SAP-7B	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

REVIEW OF GAS LAWS

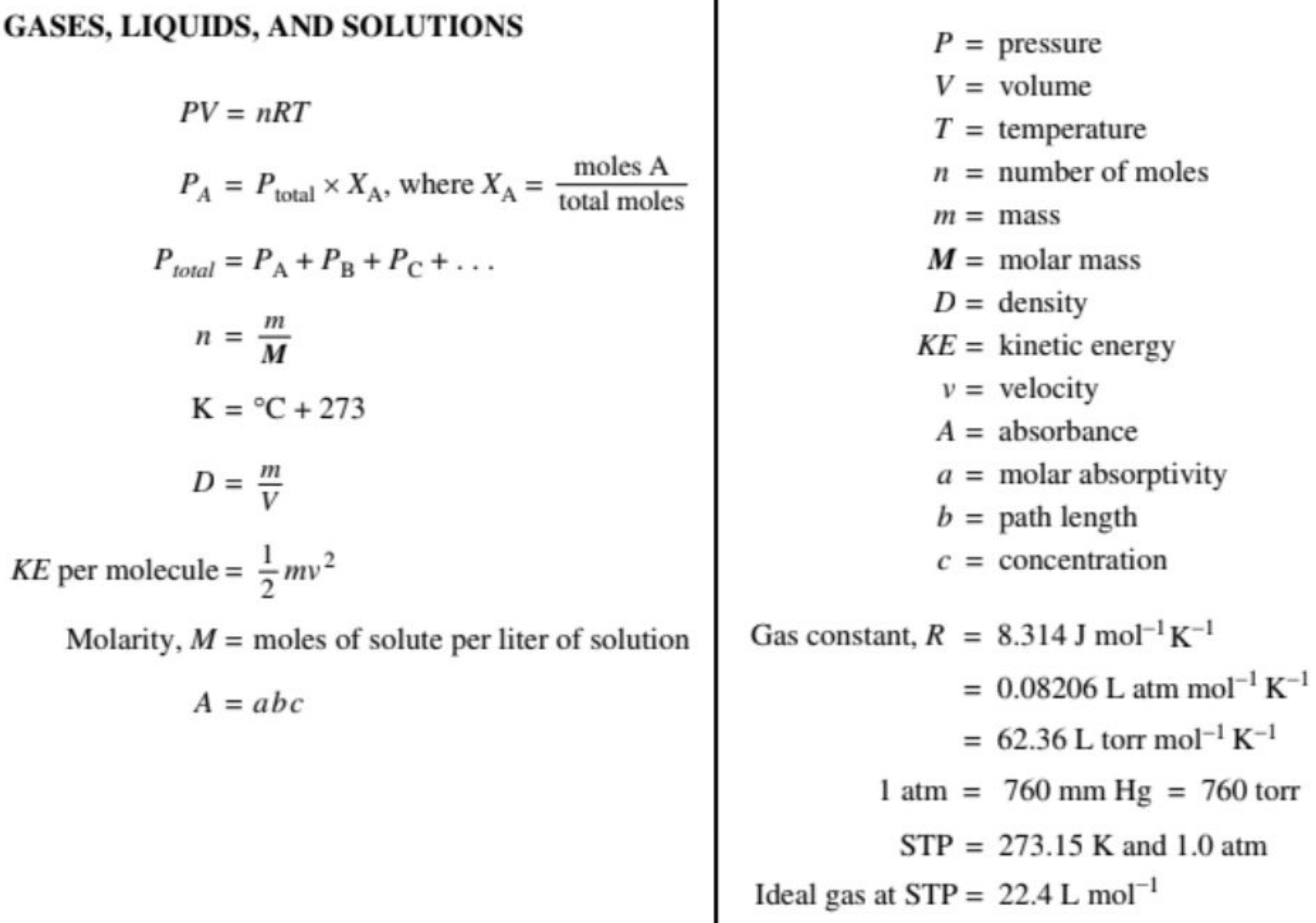
Pay attention to the last column.



GAS LAW	Relationship between ___ and ___		Two Point Equation	Mathematical Relationship
Boyle's Law	Pressure	Volume	$P_1V_1 = P_2V_2$	Inverse
Charles' Law	Volume	Temperature	$V_1/T_1 = V_2/T_2$	Direct
Gay-Lussac's Law	Pressure	Temperature	$P_1/T_1 = P_2/T_2$	Direct
Avogadro's Hypothesis	Volume	Number of moles	$V_1/n_1 = V_2/n_2$	Direct

These laws are all represented in the ideal gas law: $PV=nRT$

Given:



GASES, LIQUIDS, AND SOLUTIONS

$$PV = nRT$$

$$P_A = P_{total} \times X_A, \text{ where } X_A = \frac{\text{moles A}}{\text{total moles}}$$

$$P_{total} = P_A + P_B + P_C + \ldots$$

$$n = \frac{m}{\mathbf{M}}$$

$$\text{K} = {}^\circ\text{C} + 273$$

$$D = \frac{m}{V}$$

$$KE \text{ per molecule} = \frac{1}{2}mv^2$$

Molarity, M = moles of solute per liter of solution

$$A = abc$$

P = pressure
V = volume
T = temperature
n = number of moles
m = mass
$\mathbf{M}$ = molar mass
D = density
KE = kinetic energy
v = velocity
A = absorbance
a = molar absorptivity
b = path length
c = concentration

Gas constant, $R = 8.314 \text{ J mol}^{-1} \text{K}^{-1}$
$= 0.08206 \text{ L atm mol}^{-1} \text{K}^{-1}$
$= 62.36 \text{ L torr mol}^{-1} \text{K}^{-1}$
1 atm = 760 mm Hg = 760 torr
STP = 273.15 K and 1.0 atm
Ideal gas at STP = 22.4 L mol^{-1}

KINETIC MOLECULAR THEORY (KMT)

KMT is a theoretical model that describes the nature of ideal gases. This theory is based on the following postulates, or assumptions:

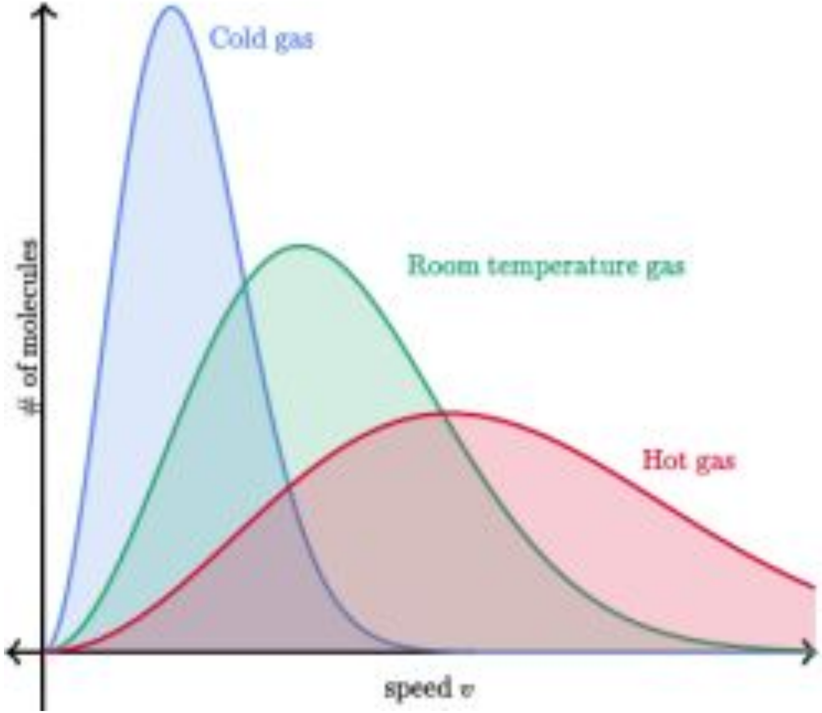
1. The volume of gas particles can be ignored because they are so small compared with the distances between the particles.
2. Gas particles are in constant, random motions. They move in straight lines until they bump into each other or the walls of the container.
3. The particles are assumed to have no attractive or repulsive forces between them, in other words, the effect of IMFs can be ignored.
4. The average kinetic energy of a sample of a gas is proportional to the Kelvin temperature of the gas. ($KE = \frac{1}{2} mv^2$)

Maxwell-Boltzmann Distribution

A Maxwell-Boltzmann distribution shows the distribution of the kinetic energies of particles at a given temperature.

All 3 curves represent the same number of gas particles.

By increasing the temperature of a gas the distribution of the particle shifts to the right.



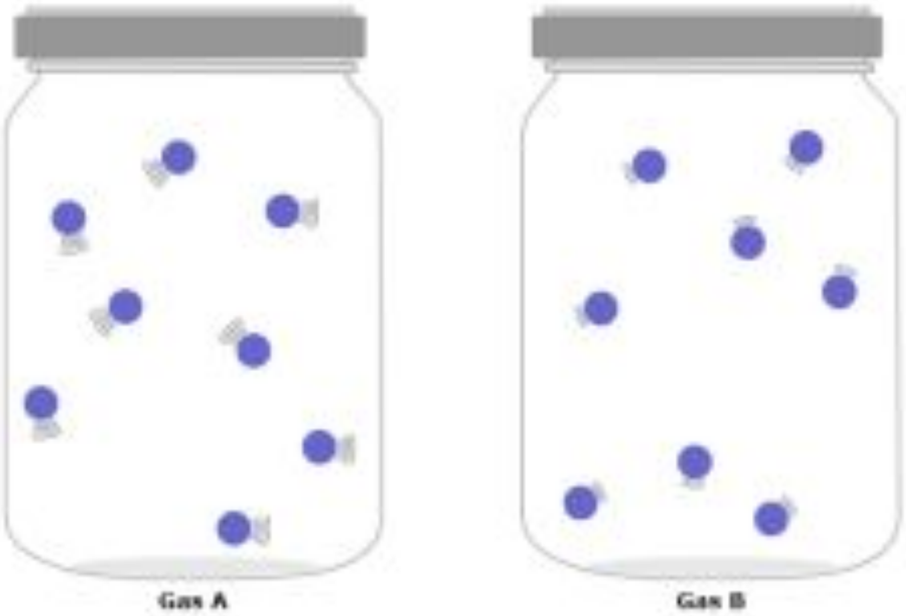
Kinetic Energy (Unit = Joules)

The temperature of the first gas is greater than the temperature of the second gas.

The mass is the same, but the velocity differs, this results in a difference in the kinetic energy.

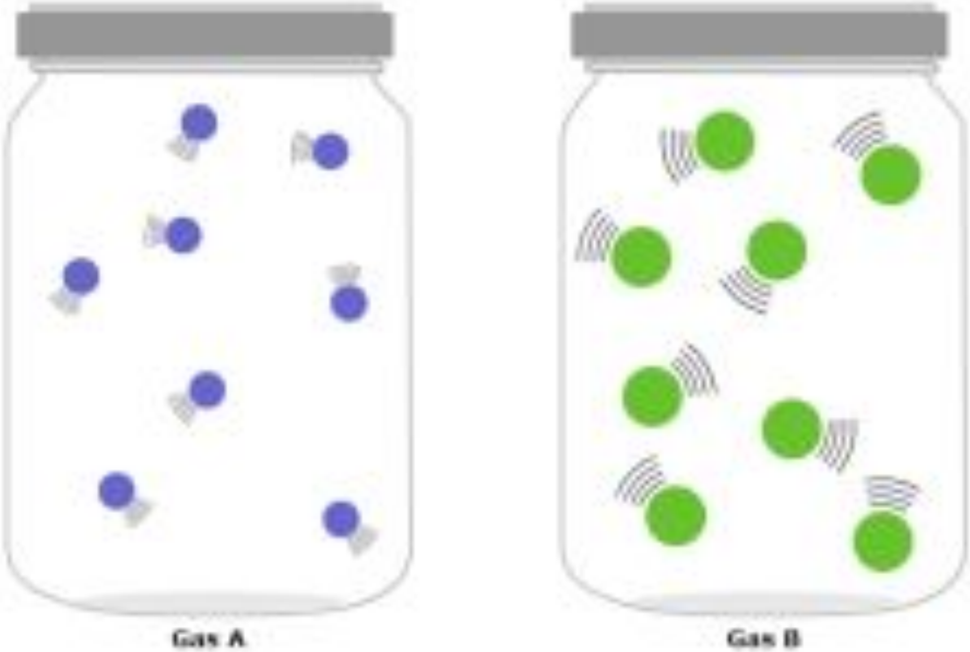
$$KE = \frac{1}{2} mv^2$$

Which gas has the greater kinetic energy?



SAME VELOCITY, DIFFERENT MASSES

Which gas has a greater kinetic energy?

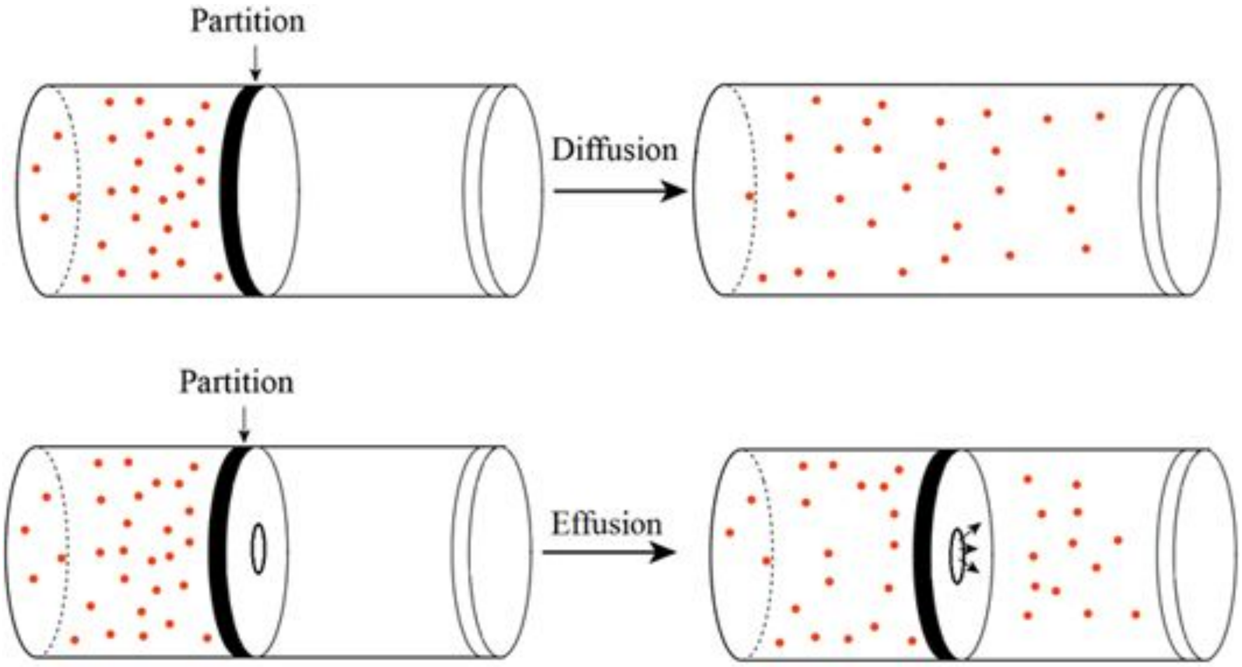


EFFUSION & DIFFUSION

Rate of effusion (the ability to diffuse through a small hole) is related to the mass of the particles. Grahams’ Law describes this relationship:

$$\frac{Rate\ 1}{Rate\ 2} = \sqrt{\frac{Molar\ Mass\ 2}{Molar\ Mass\ 1}}$$

EFFUSION AND DIFFUSION



Pictorial description of diffusion and effusion of gases

Together

Given a sample of H_2, HCl and Cl_2 at the same temperature, place the particles in order of decreasing average particle speed.

YOU DO, WE REVIEW

Consider an equimolar mixture of gases, H_2, N_2 and F_2, placed into a container with a pinhole opening. After some time has passed, place the gases in order of increasing amount *remaining in the container*.

TOPIC 3.6: DEVIATION FROM IDEAL GAS LAW

Enduring Understanding	
SAP-7	Gas properties are explained macroscopically - using the relationships among pressure, volume, temperature, moles, gas constant - and molecularly by the motion of the gas.
Learning Objective	
SAP-7.C	Explain the relationship among non-ideal behaviors of gases, interparticle forces, and/or volumes.

IDEAL GASES

Gases behave ideally under ordinary conditions – high temperatures and low pressure.

Kinetic Molecular Theory, KMT, assumes that under these conditions, volumes of gas molecules themselves are insignificant and that gases do not experience any intermolecular attractions.

Under high temperatures they are moving quickly, and under low pressure they are distant from each other so the gas particles cannot experience forces of attraction to one another.

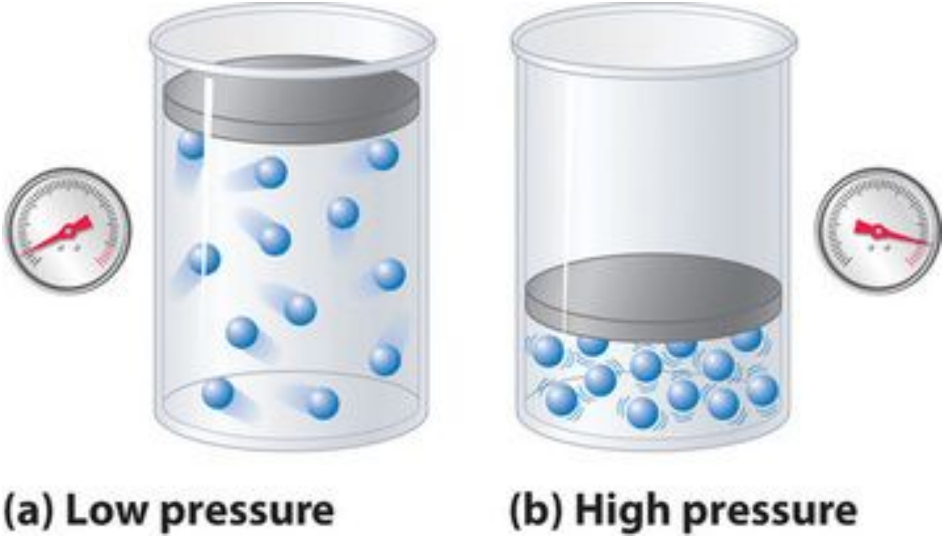
Real Gases

However, all real gases **do not behave ideally** when at high pressures (P > 5 atm) and at low temperatures.

- As pressure is increased the particles are pushed closer together which results in the **volume of the gas molecules becoming significant**.
- As the temperature decreases the particles are moving slower and the **intermolecular attractions become significant**.

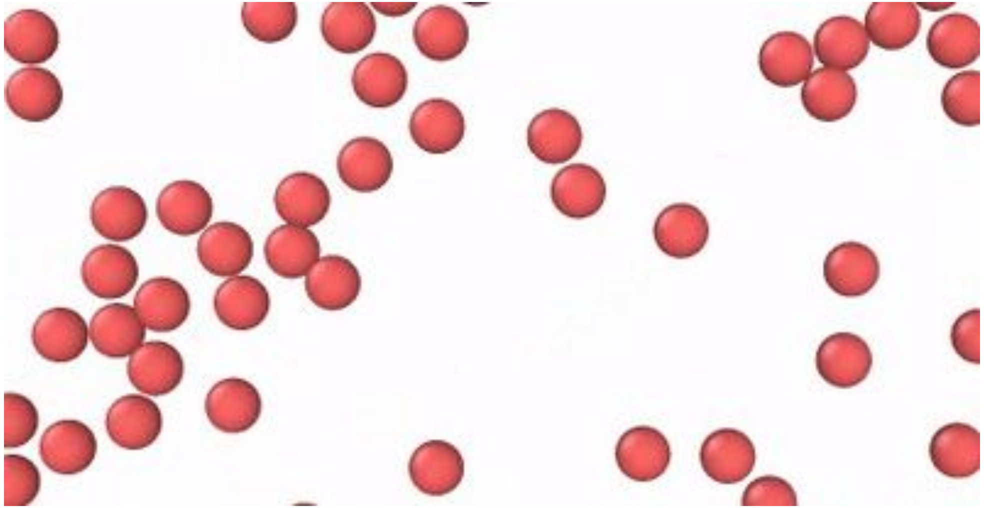
High Pressure/Low Volume

- molecules occupy a significant part of the volume of the container (cannot be considered negligible)
- gas particles closer together (cannot ignore IMFs)



LOW TEMPERATURE

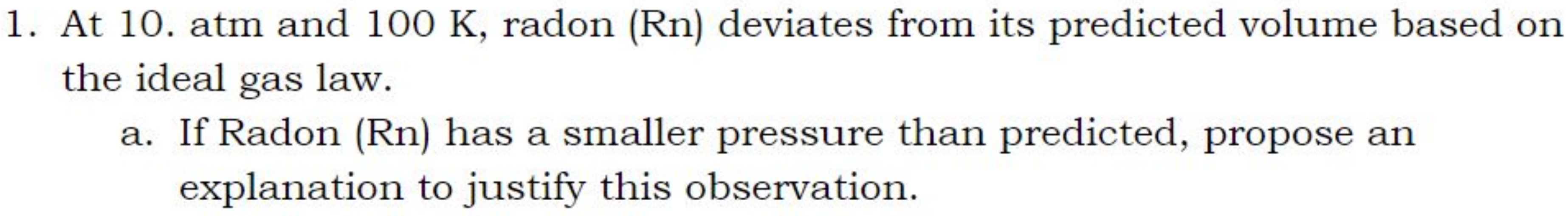
- Molecules move slower and have less energetic collisions
- They “feel” the attractive forces and “clump” together more
- collide with the walls of the container less, thereby decreasing the pressure of the gas



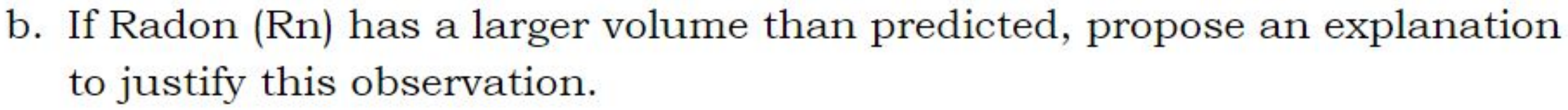
Summary

- High pressure/low volume increases the significance of molecular volumes and forces molecules closer together increasing IMFs.
- Molecules move slower at low temperatures, increasing intermolecular attractions between them.
- Nonzero molecular volume makes the actual **volume greater than predicted**.
- Intermolecular attractions make the **pressure less than predicted.**

YOU DO, WE REVIEW



1. At 10. atm and 100 K, radon (Rn) deviates from its predicted volume based on the ideal gas law.
    a. If Radon (Rn) has a smaller pressure than predicted, propose an explanation to justify this observation.

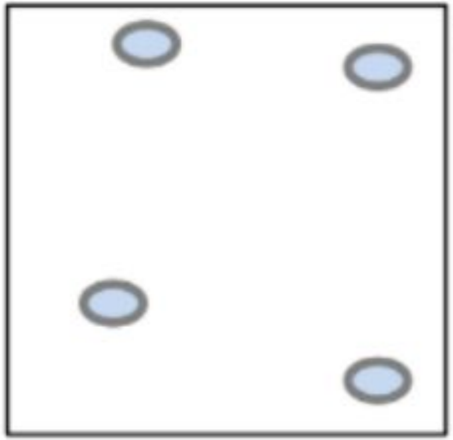


    b. If Radon (Rn) has a larger volume than predicted, propose an explanation to justify this observation.

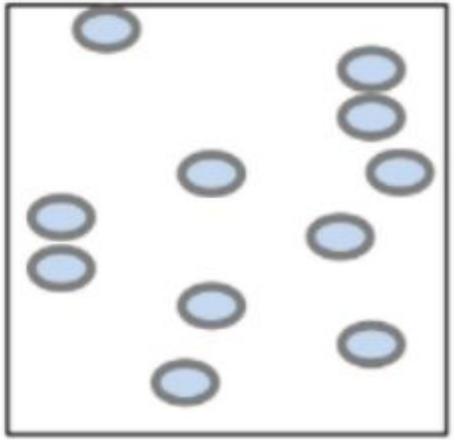
YOU DO, WE REVIEW

2. Sample 1 and Sample 2 represent the same gas at 298K. Identify the gas that is more likely to deviate from ideality and justify your selection.

Sample 1:



Sample 2:



BELLWORK

1. Draw the Lewis dot structure for H_3O^+
2. What molecular geometry does this molecule have?
3. Bond angles?
4. What type of intermolecular forces would two of these molecules have?

Unit 3 Part 3

Agenda

- Goal setting bellwork
- Schedule
- Talk about project
- Notes and practice problems on 3.7-3.9

Schedule

- Today 3.7-3.9
- Tuesday/Wednesday: 3.10, lab (project grade for lab section), 3.11
- Thursday/Friday: 3.11-3.13
- Monday: Review
- Tuesday/Wednesday: Test, lab after test
- Thursday/Friday: Project presentations

Project

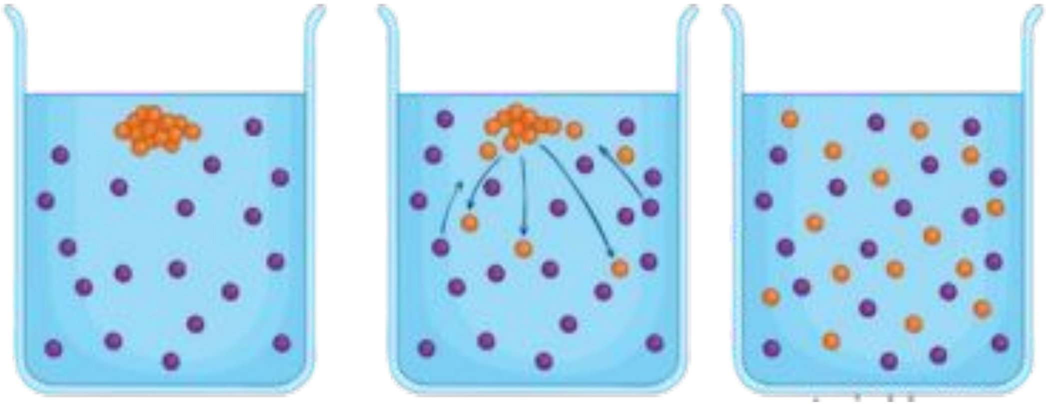
- Every student will prepare a 5-7 minute review lesson on one topic
    - The student facing materials (practice problems, slides, note sheet) will be due 10/2
    - Extra credit will be given if they are completed by next class
    - The presentations will be 10/5 and 10/6

TOPIC 3.7: SOLUTIONS & MIXTURES

Enduring Understanding	
SPQ-3	Interactions between intermolecular forces influence the solubility and separation of mixtures.
Learning Objective	
SPQ-3.A	Calculate the number of solute particles, volume, or molarity of solutions.

Solutions

Solutions are **homogeneous mixtures** that are composed of a **solute** and a **solvent**. They can be gases, liquids, or solids. For a liquid solution the solute can be a gas, liquid or a solid.



SOLUTIONS

The composition of a solution can be expressed in terms of the number of moles of the solute, per liter of solution, otherwise known as the **molarity**.

$$M = n_{solute}/L_{solution}$$

In relatively low molarity solutions, there are small numbers of solute particles compared to the number of solvent particles.

In relatively high molarity solutions there are relatively large numbers of solute particles compared to the number of solvent particles.

DILUTION

Often, solutions are prepared by adding water to (diluting) more concentrated ones.

Dilution Calculations:

1. Calculate the number of moles present in the final, diluted solution, by applying moles = (concentration) (volume).
2. Calculate the volume of the starting, more concentrated solution that supplies this number of moles by applying moles = (concentration) (volume).
3. The volume of water that must be added to the concentrated solution is simply the difference between the volume of the final, diluted solution and the volume of the concentrated solution.

DILUTION EXAMPLE

Calculate the volume of water that must be added to prepare 2.0 L of 3.0 M KOH from a stock solution that has a concentration of 8.0 mol/L.

1. (2.0 L)(3 mol/L) = 6.0 mols of KOH
2. What volume of stock solution contains 6.0 mols of KOH?

$$\frac{6.0\ \text{mol}}{8.0\ \text{mol/L}} = 0.75\ \text{L}$$

3. By taking 0.75 L of the stock solution and adding 1.25 L of water to make the solution up to 2.00 L, the final, diluted solution, will have a concentration (molarity) = 6.0 mol/2.0 L = 3.0 mol/L or 3.0 M

IONIC COMPOUNDS

Ionic compounds dissociate into ions when they form solutions.

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

$CaCl_2\ (s) \rightarrow Ca^{2+}\ (aq) + \underline{\mathbf{2}}Cl^{-}\ (aq)$

Because of this a 1 molar solution of calcium chloride will contain 1 mole of calcium ions but **two** moles of chloride ions, therefore the chloride ions will have twice the molarity of the solution.

YOU DO, WE REVIEW

Calculate the volume of 1.25M lithium chloride solution that can be made with 13.3g of lithium chloride?

What would be the concentration of chloride ion, [Cl^{1-}] in a solution that is made by dissolving 10.0g zinc chloride, $ZnCl_2$, in a solution that has a final volume of 500. mL?

TOPIC 3.8: REPRESENTATIONS OF SOLUTIONS

Enduring Understanding

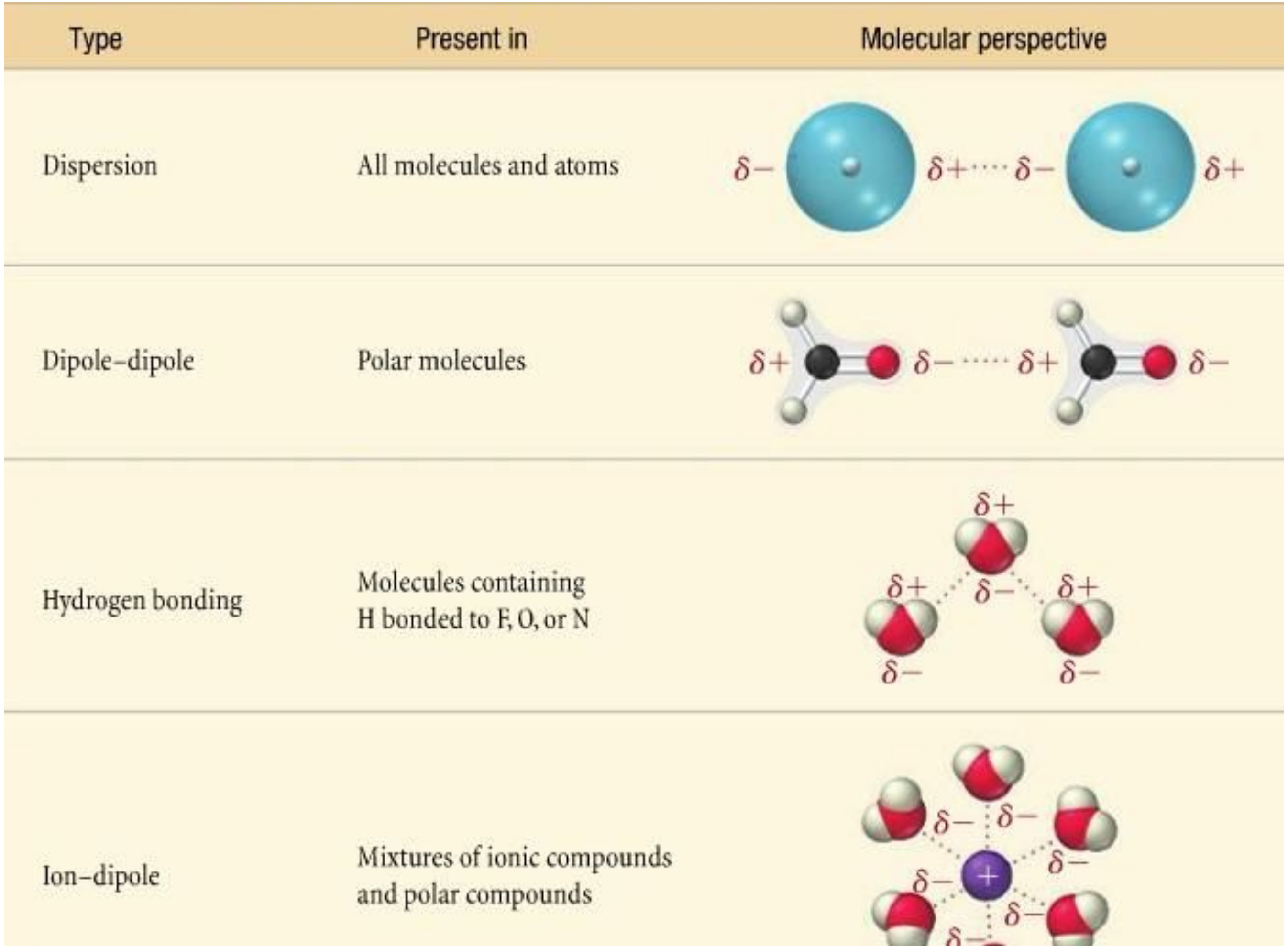
SPQ-3	Interactions between intermolecular forces influence the solubility and separation of mixtures.

Learning Objective

SPQ-3.B	Using particulate models for mixtures: a. Represent interactions between components b. Represent concentrations of components

INTERMOLECULAR FORCES

Solids and liquids exhibit all types of intermolecular forces, as summarized in the particulate drawings right:



Type	Present in	Molecular perspective
Dispersion	All molecules and atoms	
Dipole–dipole	Polar molecules	
Hydrogen bonding	Molecules containing H bonded to F, O, or N	
Ion–dipole	Mixtures of ionic compounds and polar compounds	

MISCIBILITY

Intermolecular forces exist between solute particles, between solvent particles, and between solute and solvent particles.

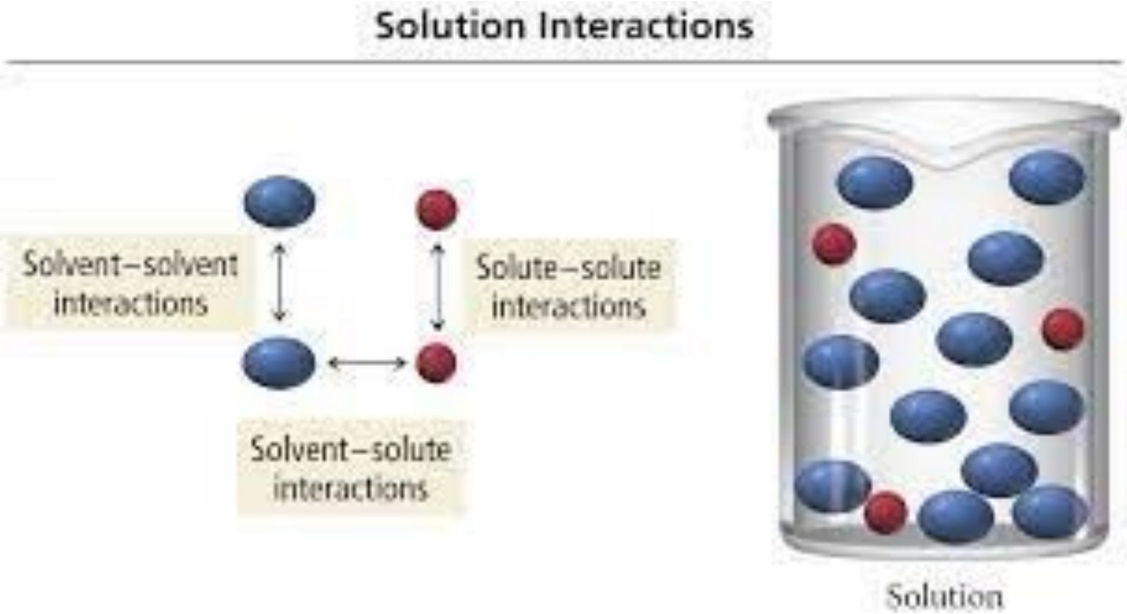
If the **solvent-solvent** and **solute-solute** attractions are strong compared to the solvent-solute interactions, a solution will not form.

Miscibility describes the ability of two substances to mix without separating. Similar materials are miscible in each other, or like dissolves like. Polar solvents dissolve polar solutes, and nonpolar solvents dissolve nonpolar solutes.

SOLUTION INTERACTIONS

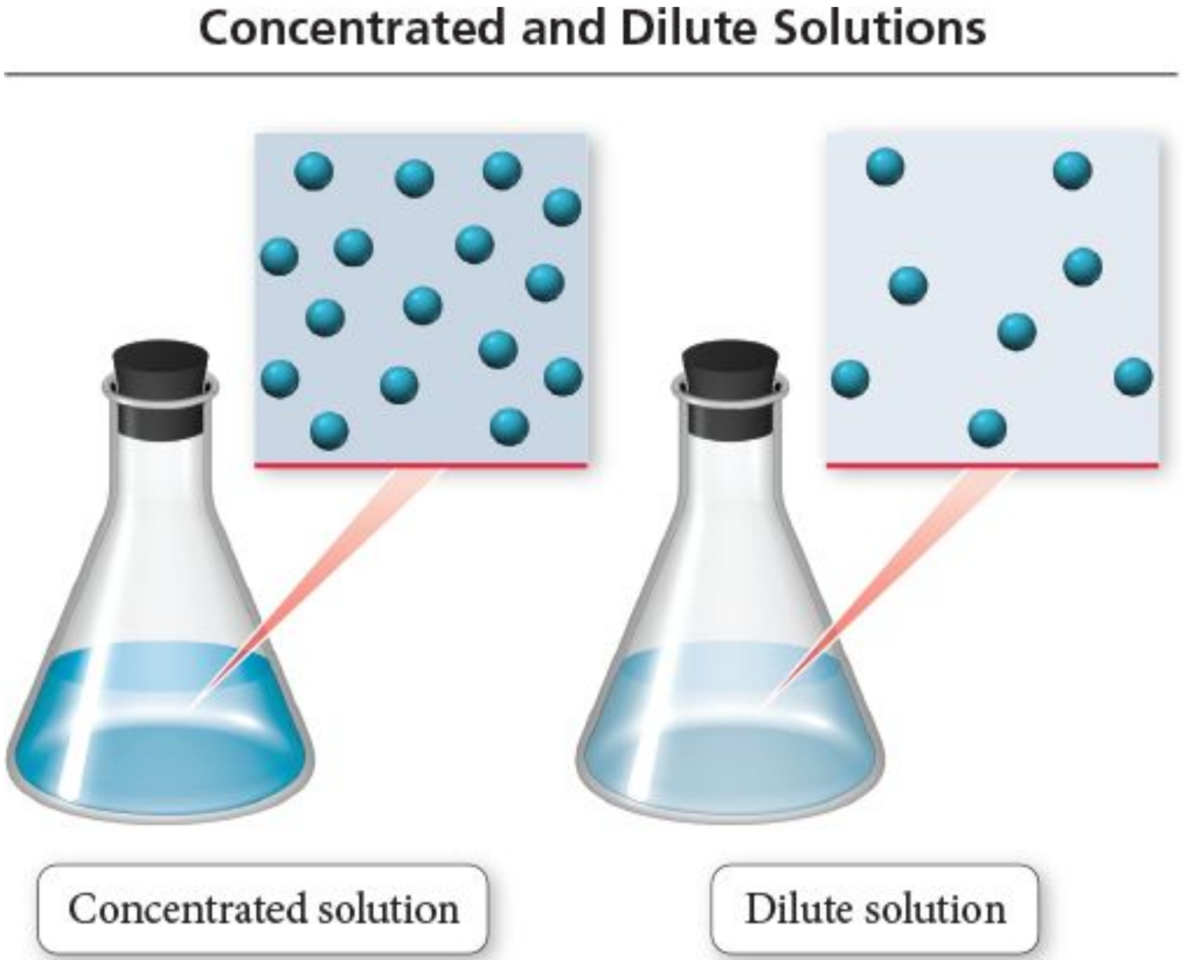
A solution will form if the solvent-solute interactions (attraction) is greater than the solvent-solvent and solute-solute interactions.

A solution will also form if the solvent-solute interactions are relatively equal to the solvent solvent and solute-solute interactions.



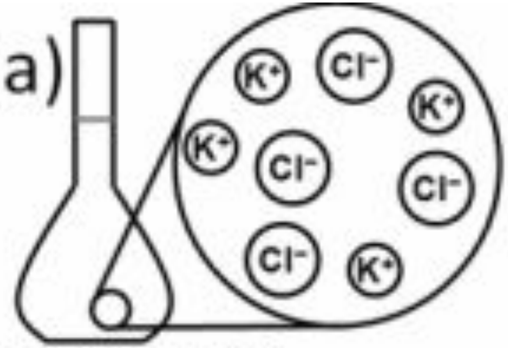
Concentration can be represented in particulate diagrams simply by showing relative amounts of solute dissolved in solution.

The more particles, the higher the concentration.



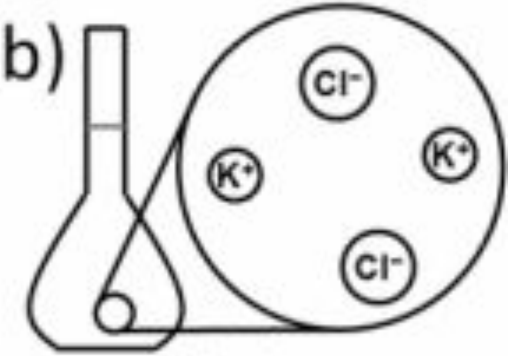
When particulate drawings are compared to one another, calculate the ratio of individual particles to determine the difference in concentration (M) between the solutions.

(a)



0.20 M KCl (*aq*)

(b)



0.10 M KCl (*aq*)

Together

Which solvent would be most effective dissolving the following solutes, hexane (C_6H_{14}) or water? Identify the dominant type of intermolecular force present between solute and solvent particles and draw a particulate drawing showing where these interactions occur.

1. Ethanol (C_2H_5OH)
2. Potassium chloride (KCl)

Agenda

- Bellwork
- Lab
- AP Daily video 3.7 & 15 minutes for practice problems
- 3.9 and 3.10 Notes and practice problems

Bellwork

List all 7 diatomic elements

PBO

SWBAT use the <u>Ideal Gas Law</u> and <u>Dalton's Law of Partial Pressure</u> IOT experimentally determine the <u>molar mass</u> of butane.

Lab- Determination of molar mass of butane

- Work in partners
- Complete the prelab questions now
- Read the procedure
- Record the atmospheric pressure as 761.49 mmHg

TOPIC 3.9: SEPARATION OF SOLUTIONS

Enduring Understanding

SPQ-3	Interactions between intermolecular forces influence the solubility and separation of mixtures.

Learning Objective

SPQ-3.C	Explain the relationship between the solubility of ionic and molecular compounds in aqueous and nonaqueous solvents and the intermolecular interactions between particles.

CHROMATOGRAPHY

Chromatography provides a method of separating a mixture of solutions based upon polarity differences. The polarity differences are sometimes considered solubility as well.

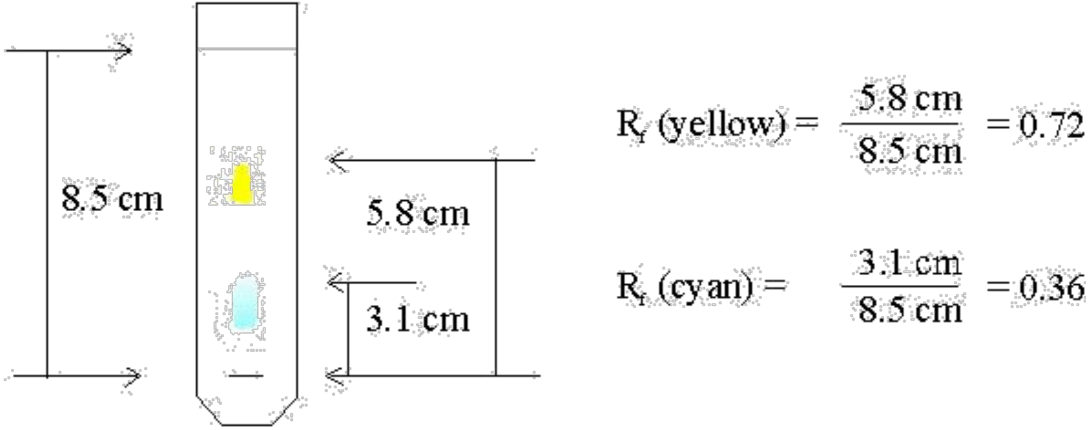
There are three types of chromatography.

1. Paper chromatography
2. Thin layer chromatography
3. Column Chromatography

Paper Chromatography (most common on the AP Test)

1. A line is drawn (typically in pencil) near the end of the strip of paper.
2. A drop of the sample to be separated is placed on that line.
3. The paper is placed in an (ideally) sealed container with a shallow layer of solvent where the bottom of the paper is touching the solvent, but the line is not.
4. Time passes and the solvent and dyes will separate.
5. Remove the paper before the solvent reaches the top of the paper.
6. Measure the height of the leading edge of the solvent and measure the leading edge of each dye.

PAPER CHROMATOGRAPHY



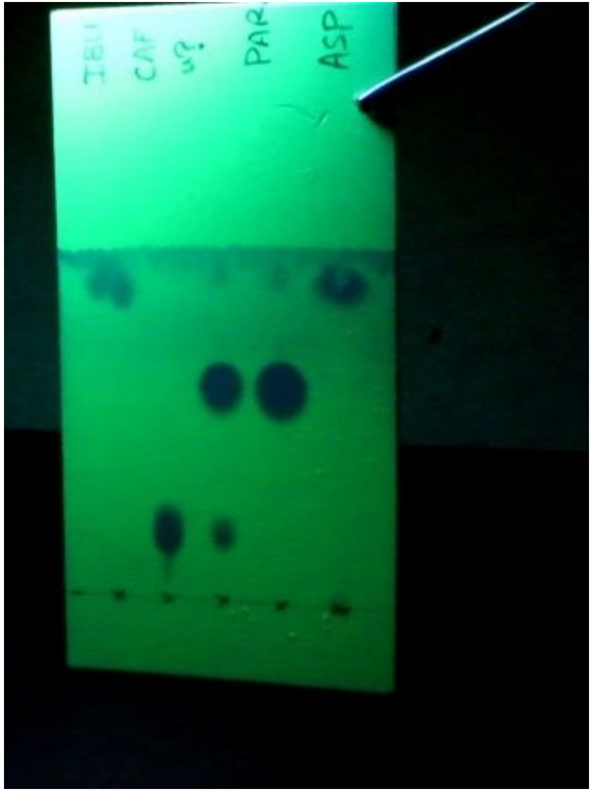
- The more similar in **polarity** the sample is to the solvent the farther it will travel.
- Polar water will have polar samples travel far, and nonpolar samples travel a short distance.
- Nonpolar benzene will have polar samples travel a short distance, and nonpolar sample travel far.
- Identification of a sample is possible by comparing R_f values- NOT COMPARING DISTANCES.
- The stationary phase is the paper and the mobile phase is the solvent.

Thin Layer Chromatography

- Thin layer chromatography is extremely similar to paper chromatography except the separation occurs on a thin layer on a sheet of plastic.
- The layer is typically silica or alumina which are extremely **polar**.
- Thin layer chromatography can be used to separate some samples that are not colored to the naked eye.
    - UV light is used with a solvent that will fluoresce under UV light to see the substances.
- This method is frequently used with amino acids

THIN LAYER CHROMATOGRAPHY

- The more nonpolar the sample is, the farther it will travel.
- The more polar the sample is, the shorter it will travel.
- The leading edge must still be identified and not allowed to reach the top.
- Identification of a sample is possible by comparing R_f values- NOT COMPARING DISTANCES.



COLUMN CHROMATOGRAPHY

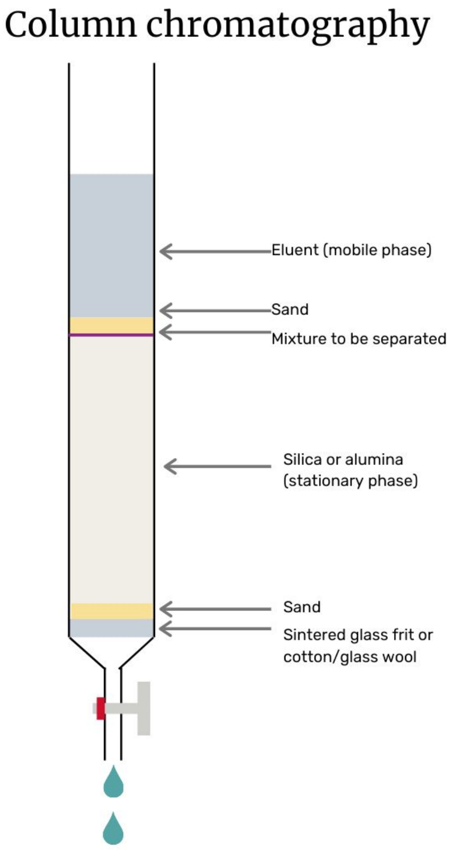
Methodology for column chromatography

1. Steel wool is placed at the bottom of a burette to prevent the gel from escaping.

2. A burette is filled with very polar silica or alumina gel.

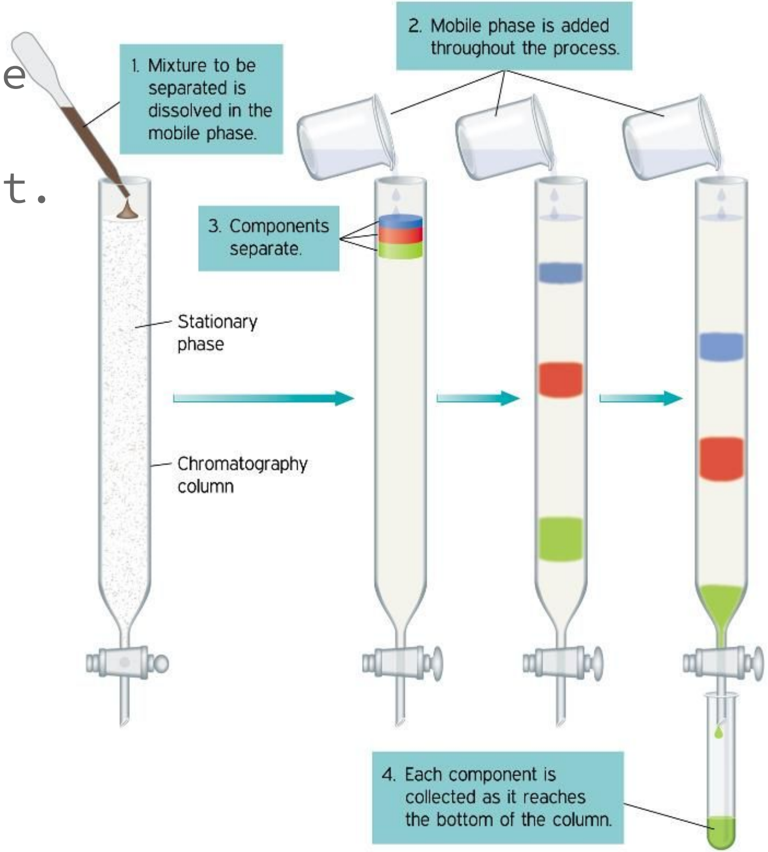
3. The mixture that is to be separated is placed at the top of the burette and is flushed time and again with nonpolar solvent. (pour solvent, wait and watch, repeat)

4. The mixture separates and each phase is collected in a beaker or flask one at a time.



COLUMN CHROMATOGRAPHY ANALYSIS

- The most polar parts of the mixture will travel the slowest and the least polar will travel the fastest.
- Once you have separated one part, you may use a new solvent to speed up the movement of the remaining part(s).
- **This is meant for separation more than analysis.**
- Once separated, you could use a different method of chromatography to identify the component.

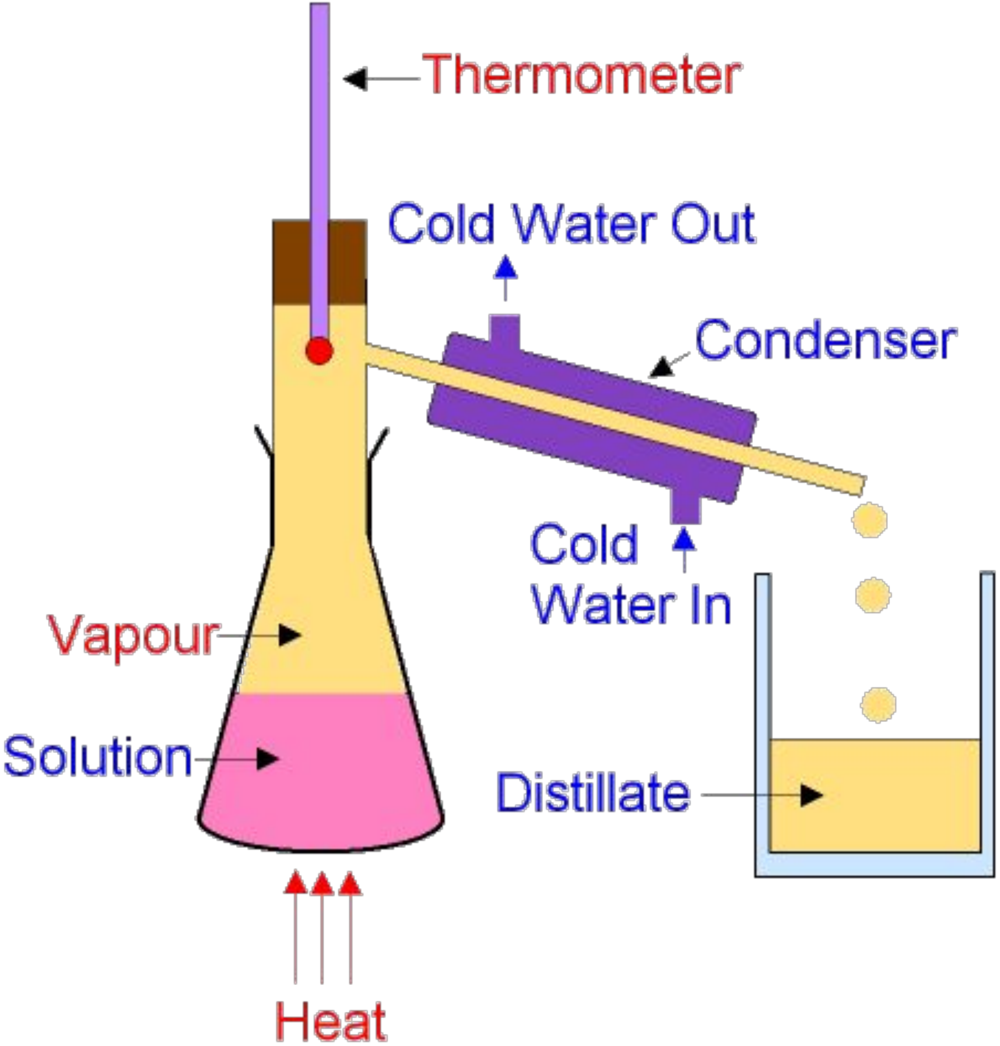


DISTILLATION

Distillation is another method of physically separating mixtures.

Distillation separates substances based on differences in boiling point (and therefore intermolecular forces).

Shown is a simple distillation apparatus.



DISTILLATION

Applications of distillation include:

- Distillation of fermented beverages
- Desalination of salt water
- Separation of crude oil into fuels and other petroleum products (this often occur through fractional distillation)

*Typically, if the boiling points are close together, fractional distillation is a better technique to separate substances. Fractional distillation allows the vapor to condense and revaporize several times, leading to a better purification.

TOGETHER

1) Which of the following is a most appropriate reason to perform column chromatography over paper chromatography?

a. The sample is a mixture of strictly polar particles

b. The sample is a mixture of strictly nonpolar particles

c. The goal is to separate and recover the parts of the mixture

d. The densities of the component parts of the mixture differ

TOPIC 3.10: SOLUBILITY

Enduring Understanding	
SPQ-3	Interactions between intermolecular forces influence the solubility and separation of mixtures.
Learning Objective	
SPQ-3.C	Explain the relationship between the solubility of ionic and molecular compounds in aqueous and nonaqueous solvents and the intermolecular interactions between particles.

SOLUBILITY

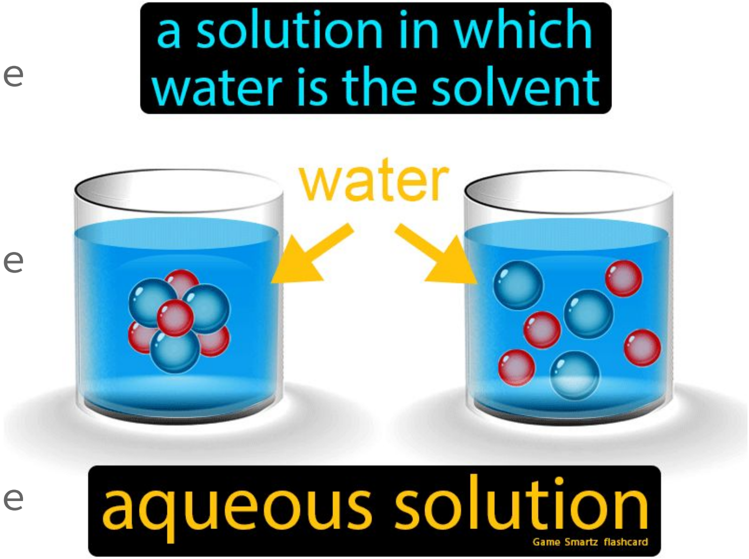
- Solubility is the extent to which a solute will dissolve into a solvent to form a solution.
- If a lot of the solute can dissolve it it described as highly soluble.
- If very little dissolves it is described as slightly (sparingly) soluble.
- If none can dissolve it is insoluble.

AQUEOUS SOLUTION

Water is a common solvent because it has the ability to dissolve many different substances.

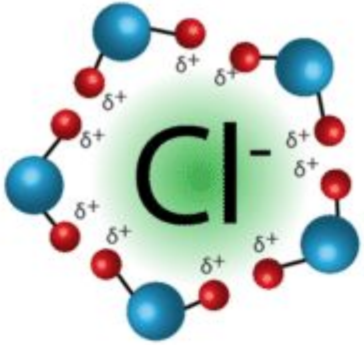
When a solution has water as the solvent the solution is referred to as an **aqueous solution.**

Water is a good solvent because it is a highly polar molecule. It has 2 sets of lone pairs on the highly electronegative oxygen atom meaning that the electrons are “pulled” towards the oxygen atom leaving the hydrogens with a partial positive charge.

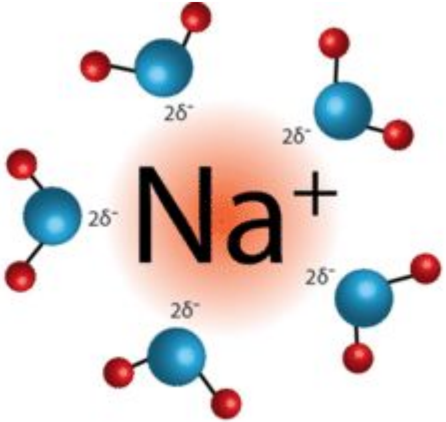


HYDRATION

When an ionic solid is dissolved into water, the ions become hydrated.



Slightly positive hydrogen are attracted to chlorine anions



Slightly negative oxygen are attracted to sodium cations

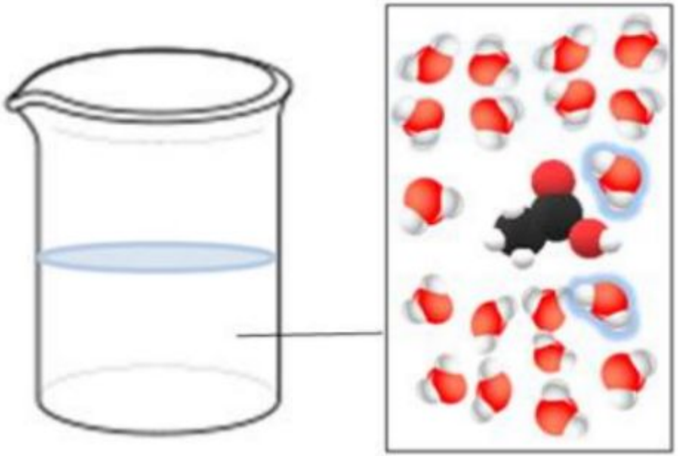
Hydration is the process where the water molecules surround the ions to dissolve them.

When water molecules surround the ions they orient themselves so that the partially positive ends of the water molecule are closest to the negative ions and the partially negative end of the water molecule is closest to the positive ion.

This attraction is referred to as an ion-dipole.

Dissolution of polar Solids

Water can also dissolve non-ionic substances as well, as shown in this image. The water is attracted to the polar part of the acetic acid, CH_3COOH, molecule. The -O-H on the acetic acid is capable of forming a hydrogen bond with water. This attraction allows the acetic acid to dissolve in water.



Notice that the acetic acid molecule does not break apart like the ionic compound.

Like Dissolves Like

Not all substances can dissolve in water.

The memory aid for determining solubility is “like dissolves like.”

Substances with similar IMFs to the solvent will dissolve but substances with IMFs that differ from the solvent are unlikely to dissolve.

Polar substances will dissolve into polar solvents and nonpolar substances will dissolve into nonpolar solvents.

Nonpolar substances **do not** dissolve into polar solvents and polar substances **do not** dissolve into nonpolar solvents **because there is not enough energy released** in the last step to negate the energy needed for the first two steps.

Steps

There are three steps that must occur in order for a solution to form.

1) The solute particles must separate (energy required)

2) The solvent particles must separate (energy required)

3) The solute and solvent particles must come back together (energy released)

The amount of energy needed for each step depends on the IMFs of the materials. The stronger the IMFs are the more energy will be involved.

MISCIBILITY

When two substances can mix together they are said to be miscible, like vinegar and water.

When two substances cannot mix together they are said to be immiscible, like oil and water.

Together

For each of the following substances, determine the type(s) of IMF present and then decide if the substance will dissolve better in hexanes (C_6H_{14}) or water (H_2O).

a. NaCl

b. C_3H_8

c. CO_2

d. CH_2O

YOU DO, WE REVIEW

Draw a molecular level diagram that shows what happens when the following reaction takes place.

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

Be sure to show the solutions before and after the reaction takes place and include a key.

Agenda

- Bellwork
- 3.11
- 3.12
- 3.13
- Progress check

BELLWORK:

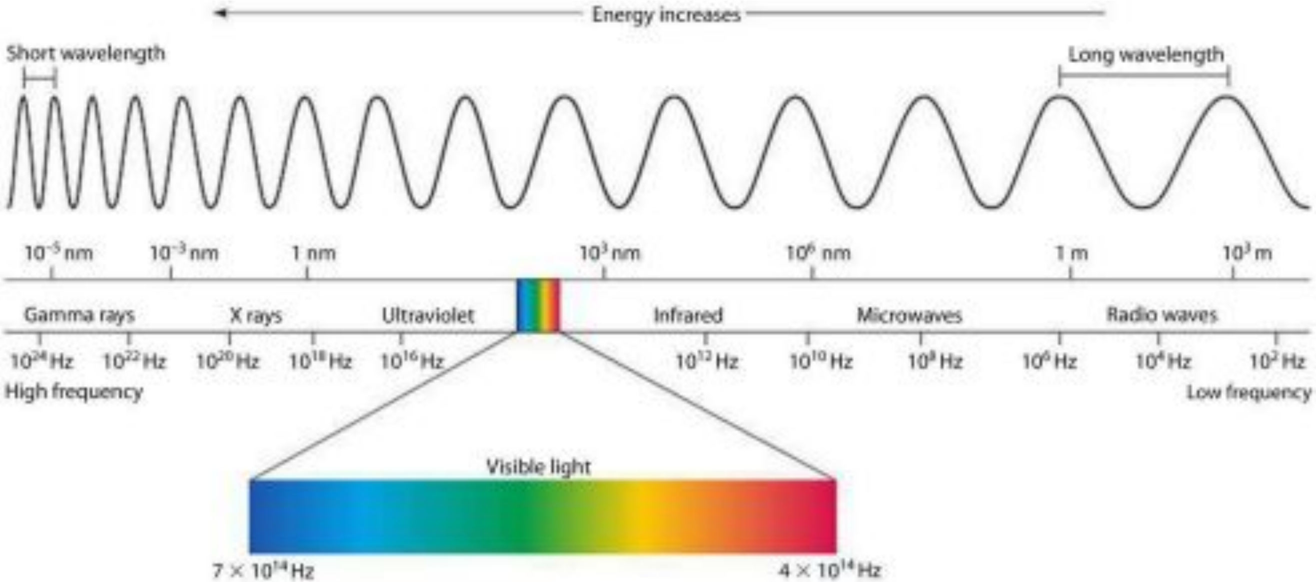
1. What is the strongest intermolecular force between NH_3 and H_2O?
2. Draw a particle diagram showing this IMF.

TOPIC 3.11: SPECTROSCOPY & THE ELECTROMAGNETIC SPECTRUM

Enduring Understanding	
SAP-8	Spectroscopy can determine the structure and concentration in a mixture of a chemical species.
Learning Objective	
SAP-8A	Explain the relationship between a region of the electromagnetic spectrum and the types of molecular or electronic transitions associated with that region.

Spectroscopy

Spectroscopy is the study of how matter interacts with electromagnetic radiation. Electromagnetic radiation is classified into different types based on the wavelength (or frequency) of the radiation.

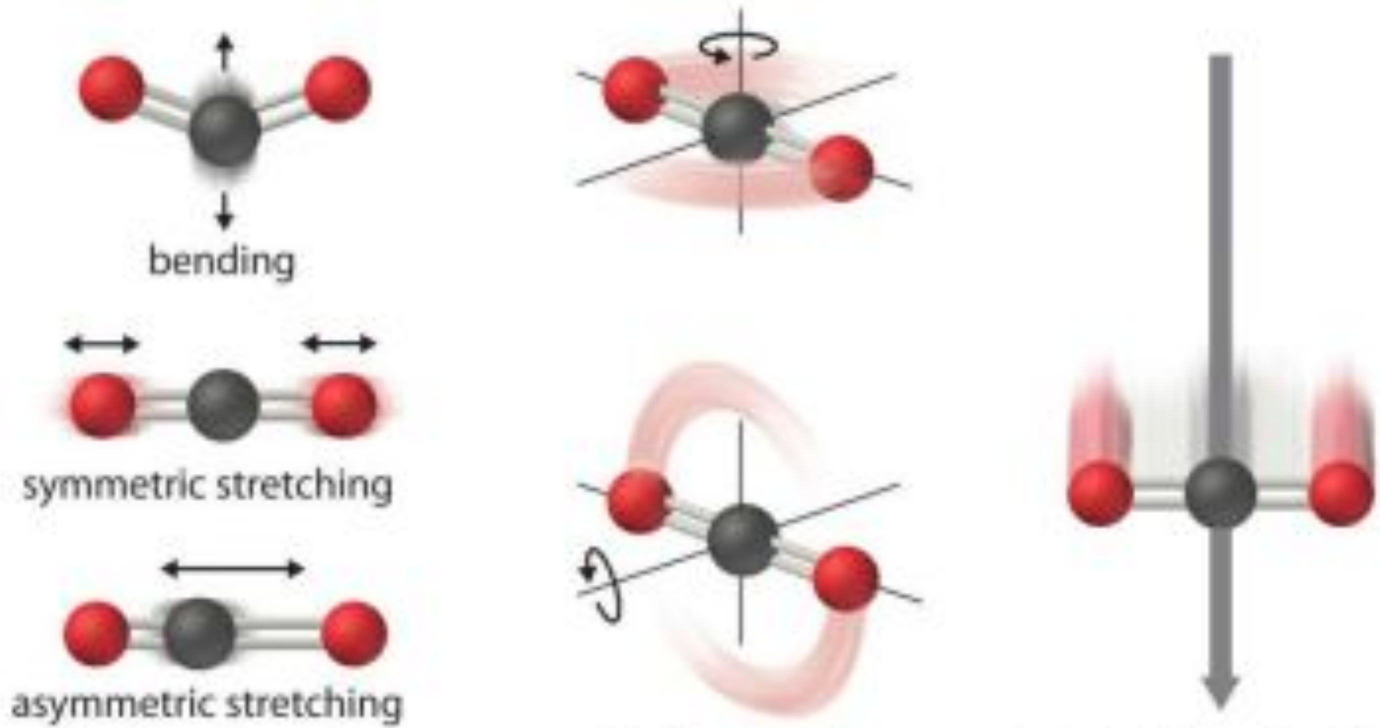


MOLECULAR MOVEMENT

Molecules can move in different ways.

- Translational - movement from one place to another
- Rotational - spinning in place
- Vibrational - moving in place - the movement of the atoms within the molecule relative to one another, such as **bending**, **stretching** and other internal movements within the molecule.

Molecular Movement



vibrational motion **rotational motion** **translational motion**

MICROWAVE ROTATIONAL SPECTROSCOPY

Microwave radiation is lower in energy than visible light. When the microwave radiation strikes the particles it causes them to rotate. This is due to the interaction of the dipole of the molecules interacting with the electromagnetic field of the microwave photons.

Key Point: Microwave radiation is associated with transitions in molecular **rotational** levels.

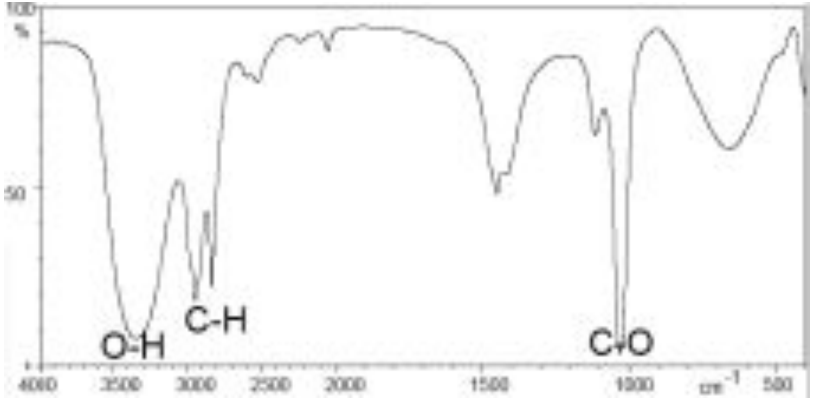
INFRARED VIBRATIONAL SPECTROSCOPY

Infrared spectroscopy is higher in energy than microwaves, but less than visible light. Infrared (IR) spectroscopy measures the vibration of the atoms and allows the determination of the functional groups present in a molecule. It analyzes absorption, emission and reflection. Generally, lighter atoms and stronger bonds tend to vibrate at higher frequencies.

Key Point: Infrared radiation is associated with transitions in molecular vibrational levels.

INFRARED VIBRATIONAL SPECTROSCOPY

Methanol IR Spectrum:



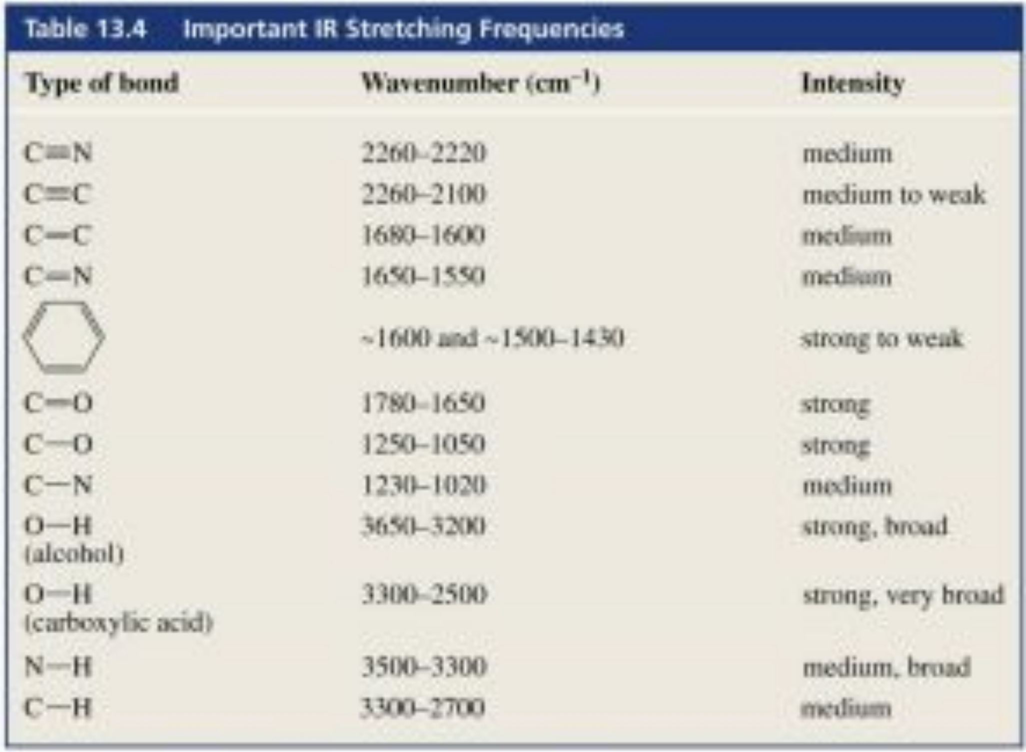


Table 13.4 Important IR Stretching Frequencies

Type of bond	Wavenumber (cm^{-1})	Intensity
C≡N	2260–2220	medium
C≡C	2260–2100	medium to weak
C=C	1680–1600	medium
C=N	1650–1550	medium
	~1600 and ~1500–1430	strong to weak
C=O	1780–1650	strong
C—O	1250–1050	strong
C—N	1230–1020	medium
O—H (alcohol)	3650–3200	strong, broad
O—H (carboxylic acid)	3300–2500	strong, very broad
N—H	3500–3300	medium, broad
C—H	3300–2700	medium

ULTRAVIOLET - VISIBLE LIGHT SPECTROSCOPY (UV-VIS)

Visible light spectroscopy is concerned with the part of the electromagnetic spectrum that we can see; the colors in the rainbow, the continuum from red to violet, from low to high energy, ~400-700 nm.

The ultraviolet portion of the electromagnetic spectrum is higher in energy as the wavelengths are shorter (~50-400 nm) and therefore the frequency is higher.

The interaction of these wavelengths of electromagnetic spectrum and molecules or atoms can allow the determination of the concentration of the molecules in the solution. The photons of colored light are absorbed by a compound causing electrons to move from their ground state to a higher energy excited state.

UV-VIS SPECTROSCOPY

A spectrophotometer measures the amount of light leaving (transmitted from) a sample and comparing it to the amount of light that entered the sample. We can find the amount of light that was absorbed by the sample. By knowing the amount of light absorbed, one can determine the concentration of a colored substance within a solution.

This is explained further in 3.13 Beer-Lambert Law.

Key Point: Ultraviolet/visible radiation is associated with transitions in electronic energy levels.

TOGETHER

Two samples of copper (II) sulfate were analyzed using visible light spectroscopy. Sample One absorbed twice as much light as Sample Two.

a. What can you deduce about the concentrations of the two solutions?

b. What must be true of the solutions in order to get results from this type of spectroscopy?

c. Describe the type of changes that occur within the molecule during the process of UV-Vis spectroscopy.

YOU DO, WE REVIEW

A. What type of spectroscopy would be responsible for the change from $1s^2 2s^2 2p^5$ to $1^{s2} 2s^2 2p^4 3s^1$?

B. What type of spectroscopy would be responsible for the bending of the O-H bonds in water?

TOPIC 3.12: THE PHOTOELECTRIC EFFECT

Enduring Understanding	
SAP-8	Spectroscopy can determine the structure and concentration in a mixture of a chemical species.
Learning Objective	
SAP-8B	Explain the properties of an absorbed or emitted photon in relationship to an electronic transition in an atom or molecule.

Photoelectric Effect

When light shines on the surface of metal, electrons can be ejected from the surface, this is known as the **photoelectric effect**.

The photoelectric effect is based on light having wave-like and particle-like properties, as predicted by Albert Einstein. Light is composed of photons. We measure the energy of a photon with:

$$E = h\nu$$

E = energy

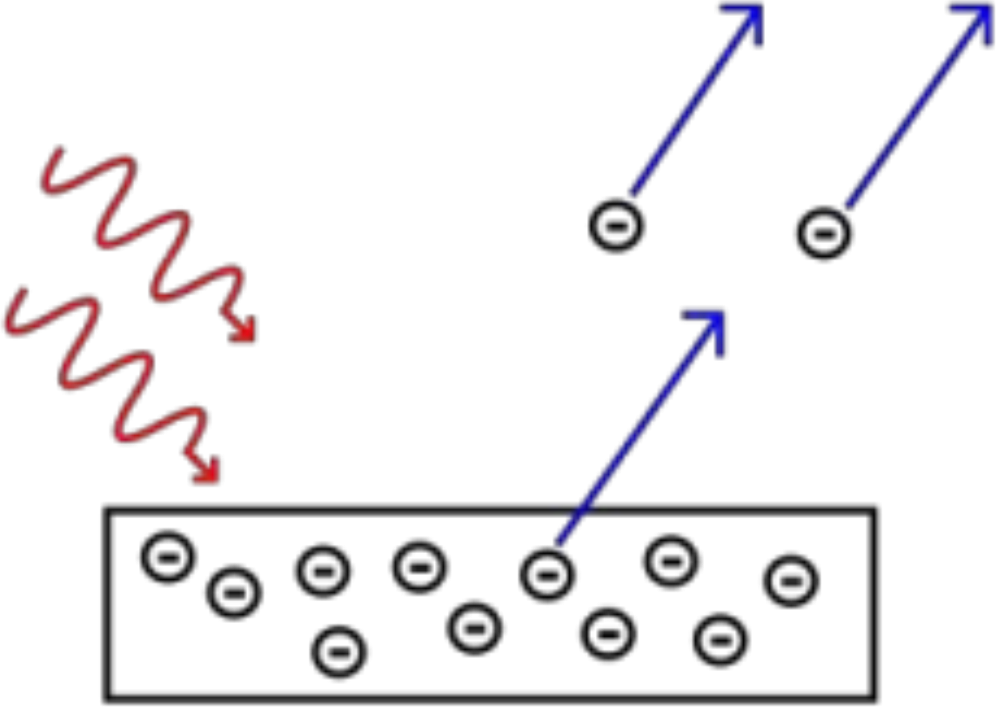
h = Planck's constant

v = frequency

Energy in (as light), electrons out

By measuring the amount of energy needed to remove the electrons we can deduce how tightly the electrons are being held in the atom, this is known as the **binding energy**.

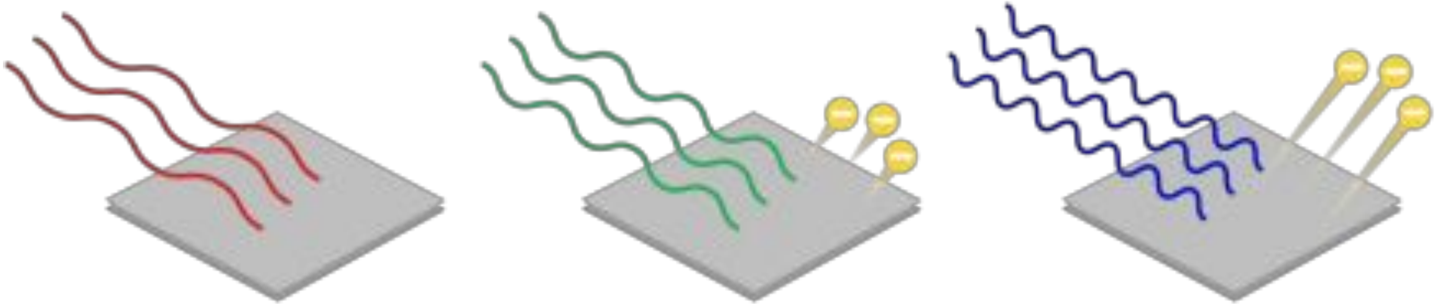
Greater energy values indicate the electrons that are closest to the nucleus or that the nucleus is a higher charge.



THRESHOLD FREQUENCY, v_0

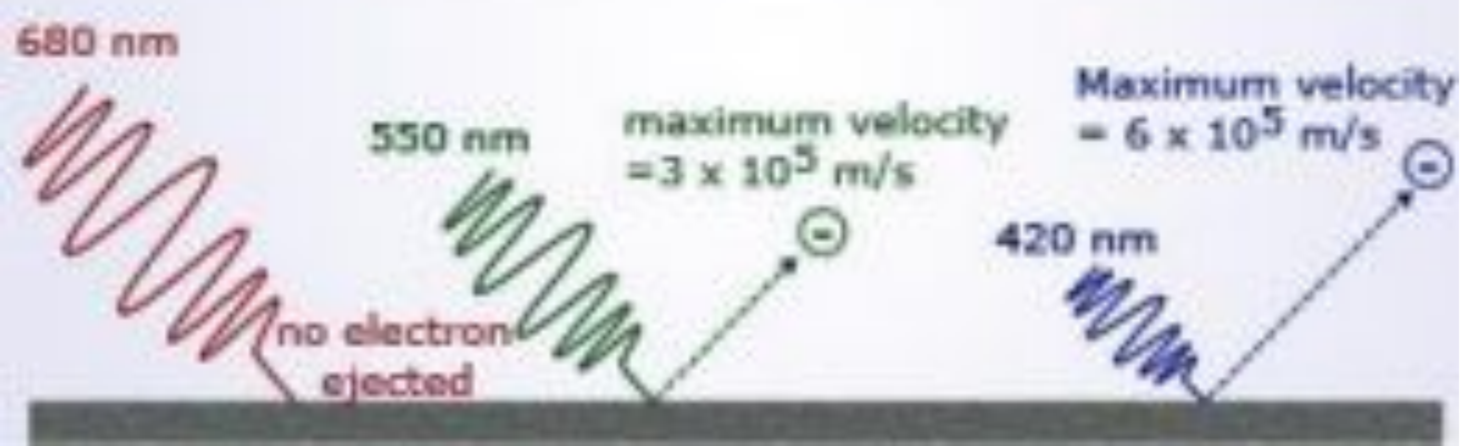
The red light does not have enough energy to remove electrons. The amount of energy needed to remove an electron is measured by the threshold frequency, v_o.

If a frequency higher than the threshold frequency is used the electrons will be ejected with more kinetic energy.

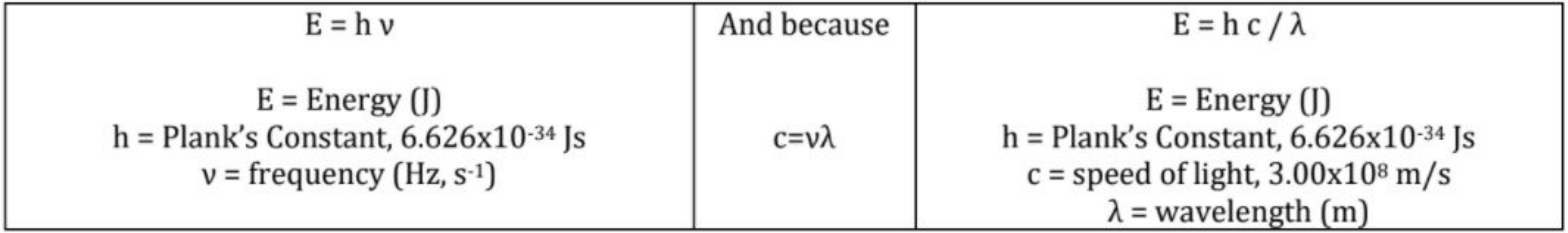


Threshold Wavelength, λ_{max}

The amount of energy needed to eject an electron can also be thought of in terms of the longest wavelength that will still eject an electron this is called the threshold wavelength λ_{max}, and any waves that are shorter than the max will have enough energy to eject electrons.



Energy is proportional to frequency, but inversely proportional to the wavelength.



$E = h\nu$	And because	$E = hc/\lambda$
E = Energy (J) h = Plank's Constant, 6.626×10^{-34} Js ν = frequency (Hz, s^{-1})	$c=\nu\lambda$	E = Energy (J) h = Plank's Constant, 6.626×10^{-34} Js c = speed of light, 3.00×10^{8} m/s λ = wavelength (m)

Wavelength is often measured in nanometers or 1x 10 $^{-9}$ meters.

SUMMARY OF THE PHOTOELECTRIC EFFECT

- The energy of the photons (light) must meet the minimum energy. This is given in terms of being a higher frequency than the threshold frequency, v_o, or having a shorter wavelength than the λ_{max}. If the minimum energy is not met, electrons will not be emitted.
- If the minimum energy is met to eject electrons, the kinetic energy of the photoelectrons is dependent on the energy of the light used to eject the electrons.

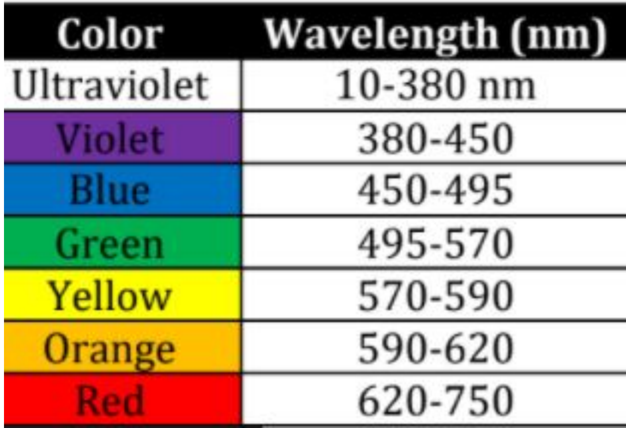
TOGETHER

When a metal surface is exposed to light with increasing frequency and energy of photons, electrons first begin to be ejected from the metal when the energy of the photons is 3.3×10^{-19} J.

a. What is the frequency of light with energy 3.3×10^{-19} J?

b. Calculate the wavelength for light with energy of 3.3×10^{-19}J.

c. What is the color of this light?



Color	Wavelength (nm)
Ultraviolet	10-380 nm
Violet	380-450
Blue	450-495
Green	495-570
Yellow	570-590
Orange	590-620
Red	620-750

You Do, We Review

$CuSO_4$ solutions absorb light at a wavelength of 635 nm. Calculate the approximate energy of one photon of this light.

TOPIC 3.13: BEER-LAMBERT LAW

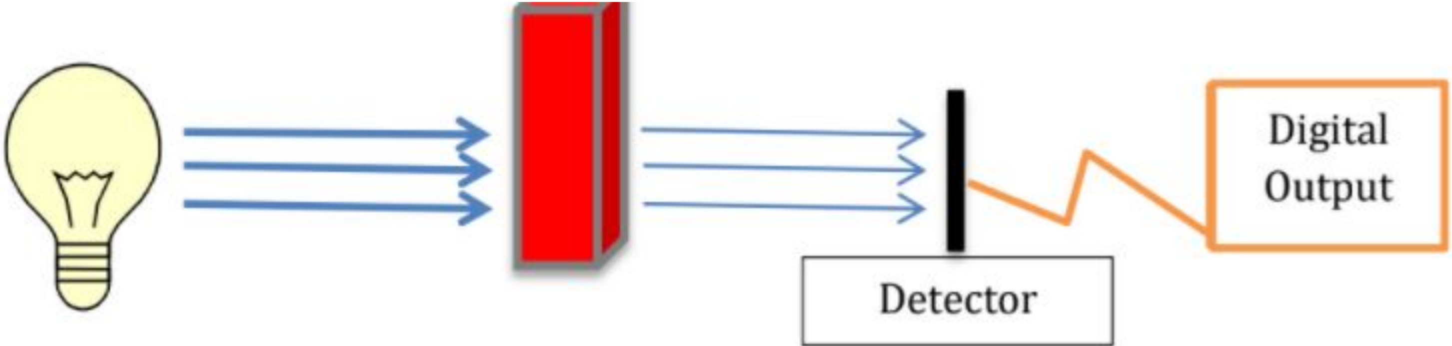
Enduring Understanding	
SAP-8	Spectroscopy can determine the structure and concentration in a mixture of a chemical species.
Learning Objective	
SAP-8C	Explain the amount of light absorbed by a solution of molecules or ions in relationship to the concentration, path length, and molar absorptivity.

ABSORBANCE

We can measure the concentration of a colored substance is by seeing how much light can be passed through it or how much light is stopped by the solution.

The measure of the light that is stopped by or absorbed by a solution is referred to as the absorbance.

A spectrophotometer or a colorimeter can be used to measure the absorbance. Below is a simplified schematic.

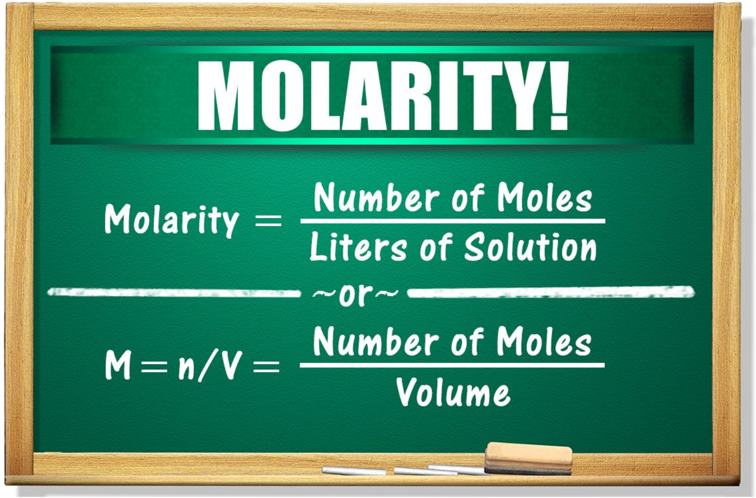


MOLARITY

The more concentrated the solution is, the less light can pass through the solution. Concentration is a measure of the amount of particles that are distributed in a solution.

You can think of it as how densely packed the particles are.

Concentration is measured in a lot of different ways but chemists often use **molarity**.



MOLARITY!

$$\text{Molarity} = \frac{\text{Number of Moles}}{\text{Liters of Solution}}$$

~or~

$$M = n/V = \frac{\text{Number of Moles}}{\text{Volume}}$$

Buzzle.com

Decreasing Molarity of Red solution

MOLARITY EXAMPLE

If I have 2 Liters of a 0.25 M solution, how many moles of solute are there?

Molarity Example

If I have 2 Liters of a 0.25 M solution, how many moles of solute are there?

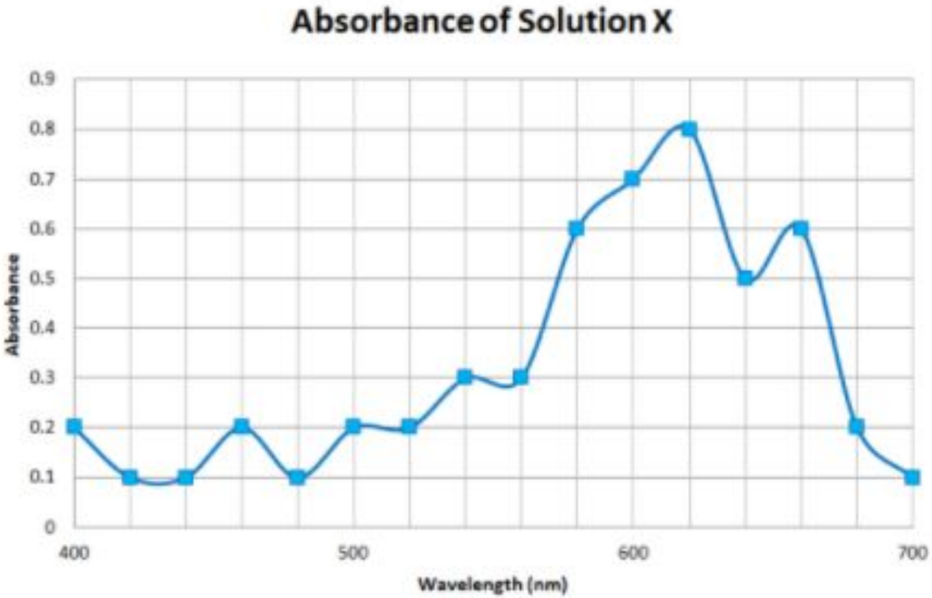
$$M = \frac{\text{moles of solute}}{\text{Liters of solution}} = \frac{\text{mol}}{\text{L}}$$

$$\text{mol} = M \cdot L = (0.25)(2) = 0.5\ \text{mol}$$

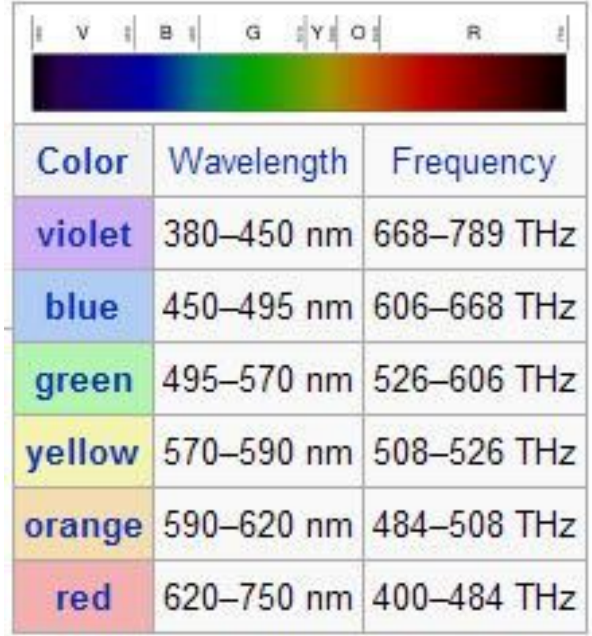
USING ABSORBANCE TO MEASURE CONCENTRATION

1. Determine the best wavelength to use to measure your sample. To do this you will place your sample into a spectrophotometer and generate a graph of absorbance vs. wavelength.

   *The best choice for a wavelength is one where the absorbance is close to 1.



Complementary Colors



Color	Wavelength	Frequency
violet	380–450 nm	668–789 THz
blue	450–495 nm	606–668 THz
green	495–570 nm	526–606 THz
yellow	570–590 nm	508–526 THz
orange	590–620 nm	484–508 THz
red	620–750 nm	400–484 THz

Different colors of solutions absorb different colors (or wavelengths) of light.

If we are using a red solution, we should choose to use green light with a wavelength of 495-570 nm to measure it. Red and green are complementary colors.

USING ABSORBANCE TO MEASURE CONCENTRATION

2. Create a calibration graph with standard solution concentrations and their corresponding absorbance value. This graph should be linear, and the slope of the line will give you the molar absorptivity constant. The more concentrated a solution is the more light is absorbed by the solution.

Beer-Lambert Law: $A = \varepsilon\ b\ c$

A = Absorbance

ε = molar absorptivity constant (this is the slope of the graph) (1/M*cm)

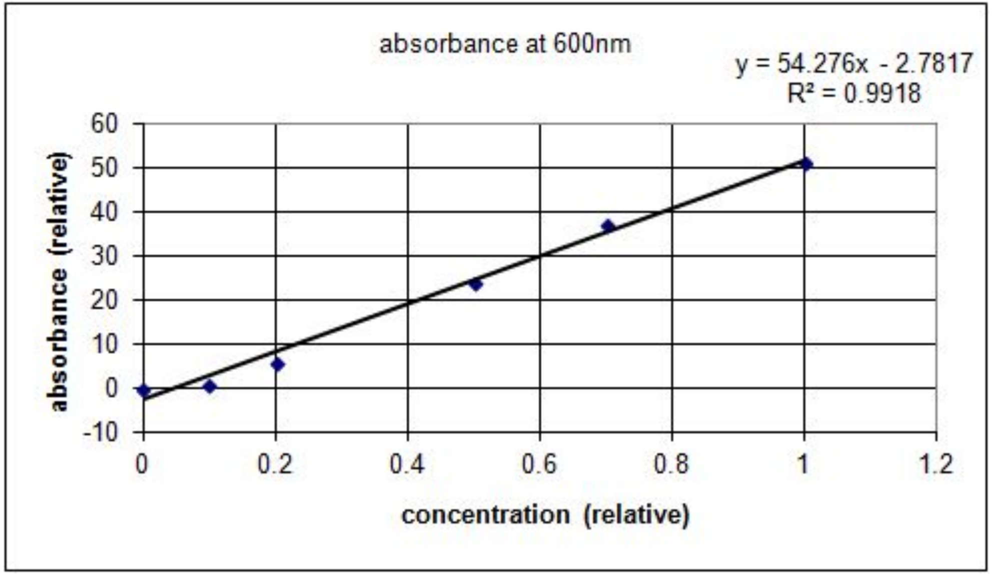
b = path length (usually 1 cm)

c = concentration (Molarity, M)

Using Absorbance to measure concentration

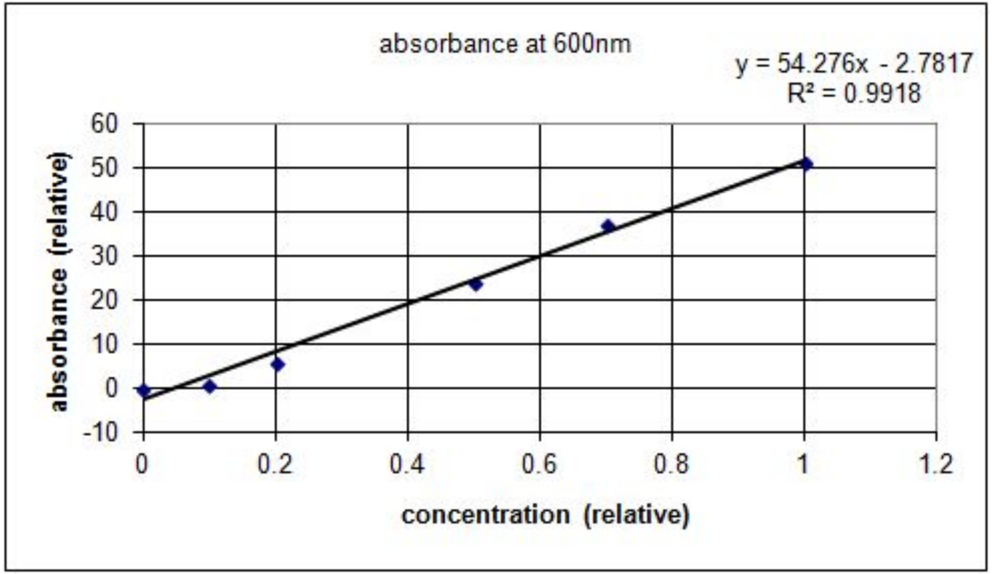
When you calibrate the spectrophotometer you should start with a **blank** - your solvent only, usually water.

By calibrating the spectrophotometer with the blank it will remove the data from the solvent and therefore only outputs the absorbance of the solute. (It is analogous to using the tare or zero button on a digital scale.)



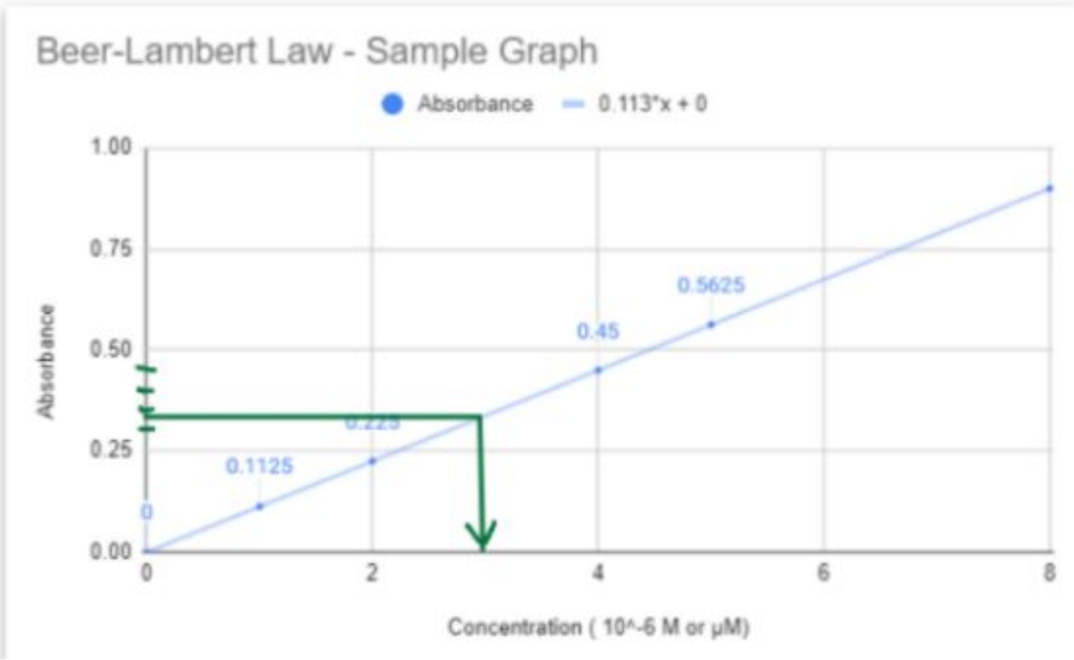
Once you have your graph

You can measure the absorbance for an unknown sample and determine the concentration from the equation for the line.



TOGETHER

1. Based on the provided calibration graph, if the absorbance of an unknown concentration of the solution was 0.34, what was the molarity of the solution?



You Do, We Review

#1 on your practice problems