

AP Exam Prep AP Chemistry Exam Prep Course

Fall 2024, Chapter 3 Notes



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3. Intermolecular Forces and Properties

3.1 Intermolecular Forces

3.1.1

Intramolecular vs Intermolecular Forces

Intramolecular Forces: A force acting between atoms **within a molecule**. These forces are relatively **strong**.

Examples: covalent bonding, ionic bonding

Intermolecular Forces: Attractive forces (electrostatic in nature) acting **between molecules**, atoms, ions. These forces are relatively **weak**.

- Hydrogen bonding
- Dipole Dipole
- London Dispersion Forces/Van der Waals forces

 **WIZE TIP**

"Intra" means within while "inter" means between!

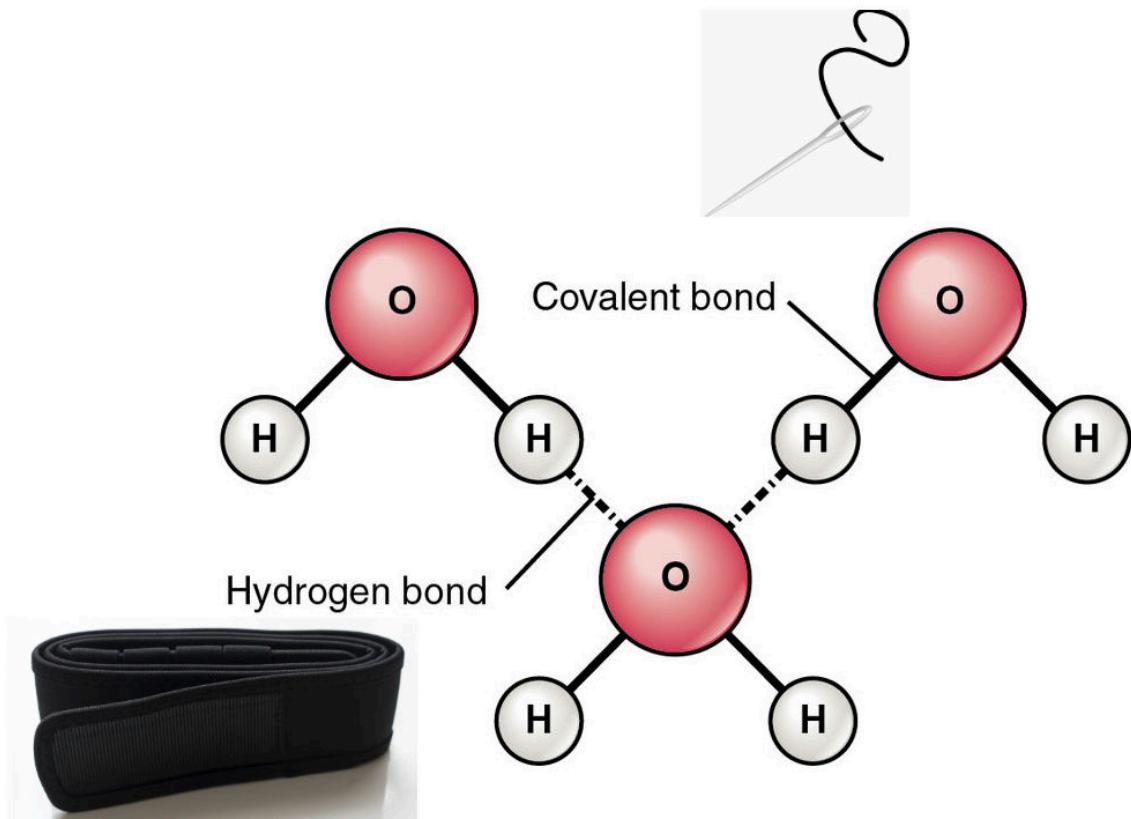


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Does the bond that is "sewn" with a needle represent an intramolecular or intermolecular bond/force?

Does the band that is attached by a velcro strap represent an intramolecular or intermolecular bond/force?

What do you think will happen if we try to pull the water molecules apart? Which forces would break first?

Intermolecular Forces

Intermolecular forces define physical properties of compounds (boiling points, melting points, solubility, vapour pressure, viscosities etc)

Boiling and Melting Points

The stronger the intermolecular forces, the lower/higher: _____ the boiling and melting points

Intermolecular Forces

1. Hydrogen Bonding



WIZE CONCEPT

To form a H bond you Need:

A H from H-F, H-O, or H-N bond to interact with a lone pair of electrons on another F, O, or N atom in a second molecule.

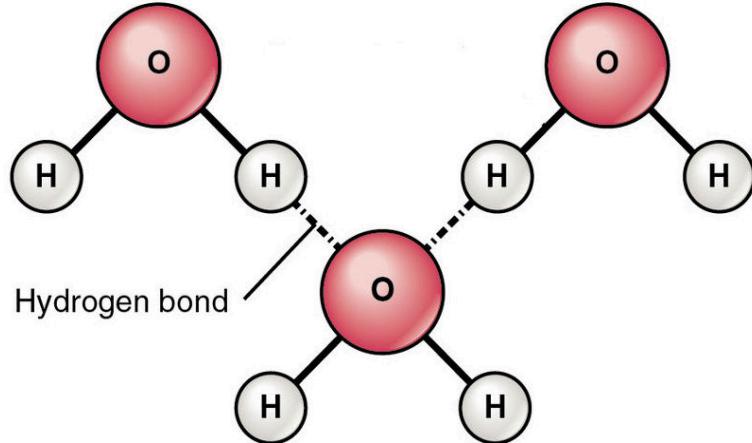
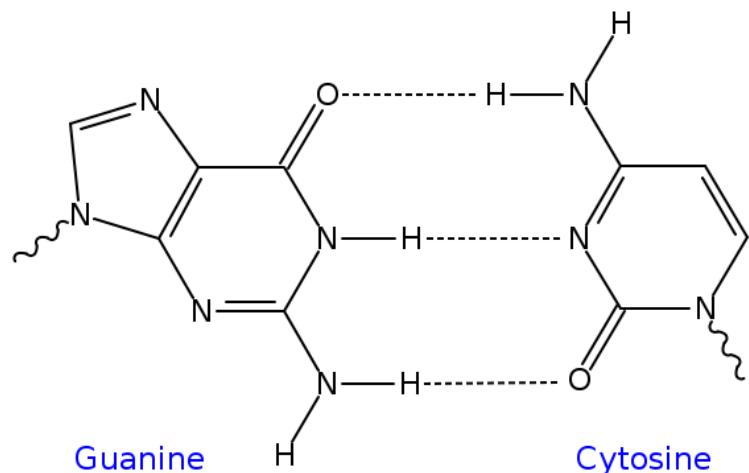


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 **WIZE CONCEPT**

Identifying features: Compound contains an O-H, N-H, or F-H bond.

Strength: The strongest intermolecular force.

Cause: The partial negative charge on the electronegative atom (N, O, or F) attracts the partial positive on the hydrogen atom that is bound to a highly electronegative atom (N, O, or F)

More H bonds=higher/lower: _____ bp

 **WIZE TIP**

Recall the order of EN:

F > O > N > Cl > Br > I > S > C ~ H

F, O, and N are the most electronegative elements. We also know that F is more EN than O, and O is more EN than N.

This is why a **hydrogen bond that involves F will be the strongest! O second strongest and N third strongest!**

2. Dipole-Dipole Interactions

This happens when the **positive end** of one **polar molecule** interacts with the **negative end** of another **polar molecule**!

! WATCH OUT!

Remember in order to be a **polar molecule** the molecule must have a net dipole!

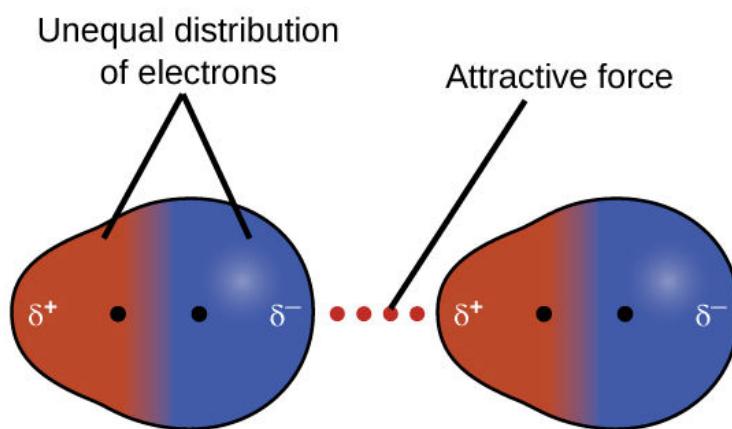


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3. London Dispersion Forces (LDF) (aka Van der Waals Forces)

These forces are involved in **interactions between non-polar parts of molecules**

This happens because a **temporary instantaneous dipole** can happen in a molecule as electrons move around. At a given moment they might not be evenly distributed, leading to a separation in charge. LDF is between 2 molecules with instantaneous dipoles.

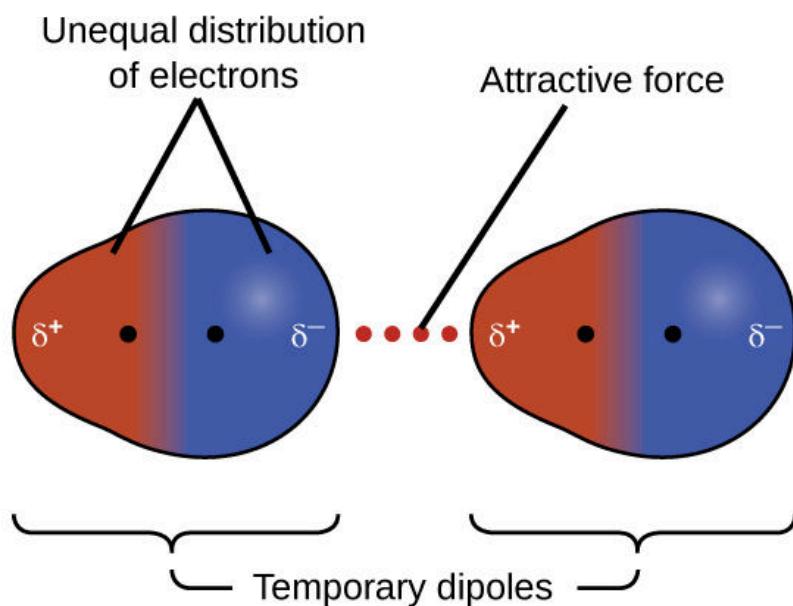


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WIZE CONCEPT

Identifying features: All compounds have these forces!

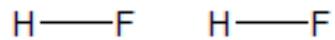
Strength: The weakest intermolecular force.

Cause: The electrons of one molecule being attracted to the nuclei of another.

Soluble in: non-polar or polar substances? _____ substances (Like Dissolves Like)

Example: Intermolecular Forces

Based on the interaction of the following two components, what is the strongest intermolecular force? What other forces can be found?



Example: Intermolecular Forces

What is the strongest intermolecular force that can be found for the following interaction? What other forces are present?

Xe and Ne

3.1.5

Practice: Strongest Intermolecular Force

Choose the strongest intermolecular forces in the following compounds.

Part 1



Ionic bonds

Hydrogen bonding

London Dispersion Forces

Dipole-Dipole Interactions

Practice: Strongest Intermolecular Force

Choose the strongest intermolecular forces in the following compounds.

Part 2



Ionic bonds

Hydrogen bonding

London Dispersion Forces

Dipole-Dipole Interactions

Practice: Strongest Intermolecular Force

Choose the strongest intermolecular forces in the following compounds.

Part 3



Ionic bonds

Hydrogen bonding

London Dispersion Forces

Dipole-Dipole Interactions

Practice: Strongest Intermolecular Force

Choose the strongest intermolecular forces in the following compounds.

Part 4



Ionic Bonds

Hydrogen Bonding

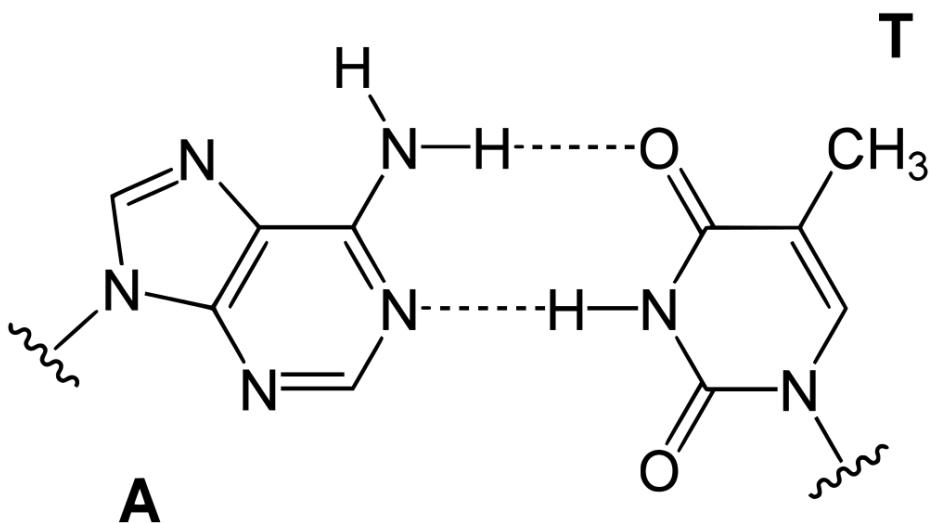
London Dispersion Forces

Dipole-Dipole Interactions

Practice: Strongest Intermolecular Force

Choose the strongest intermolecular forces in the following compounds.

Part 5



Ionic Bonds

Hydrogen Bonds

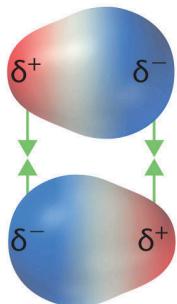
London Dispersion Forces

Dipole-Dipole Interactions

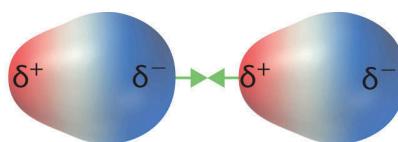
Practice: Strongest Intermolecular Force

Choose the strongest intermolecular forces in the following compounds.

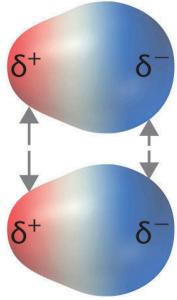
Part 6



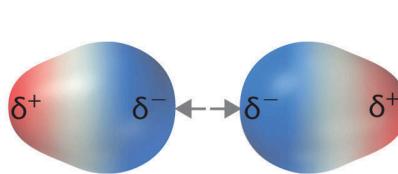
(a) Attraction



(b) Attraction



(c) Repulsion



(d) Repulsion

Ionic Bonds



Hydrogen Bonds



London Dispersion Forces



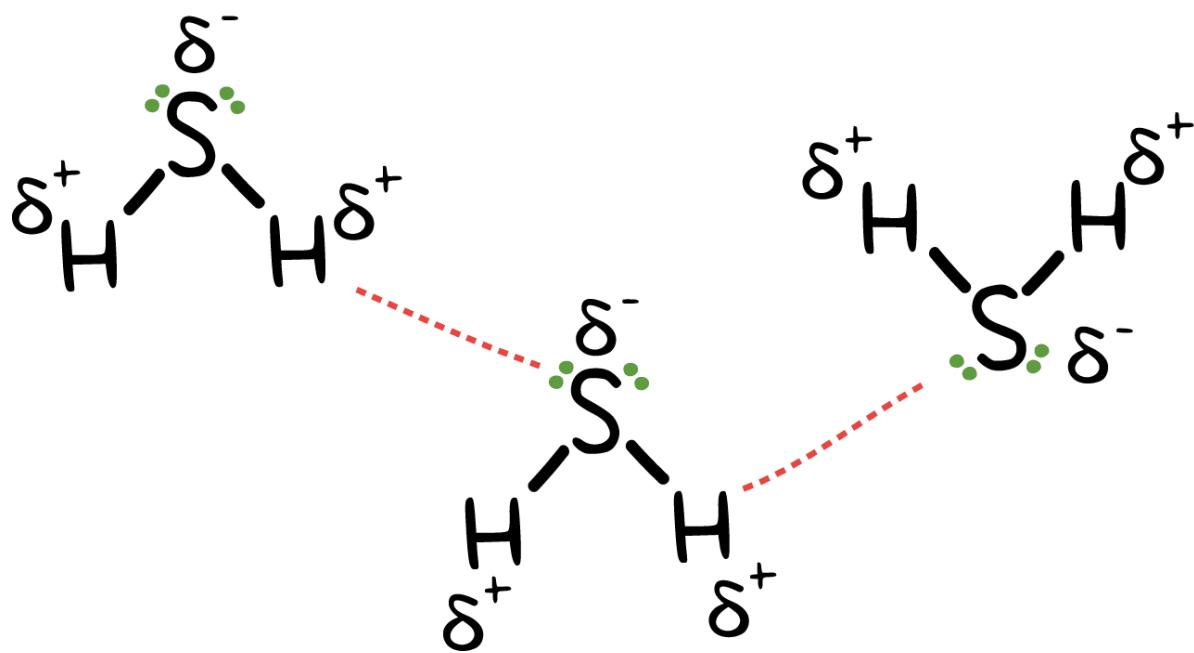
Dipole-Dipole Interactions



Practice: Strongest Intermolecular Force

Choose the strongest intermolecular forces in the following compounds.

Part 7



Ionic Bonds

Hydrogen Bonds

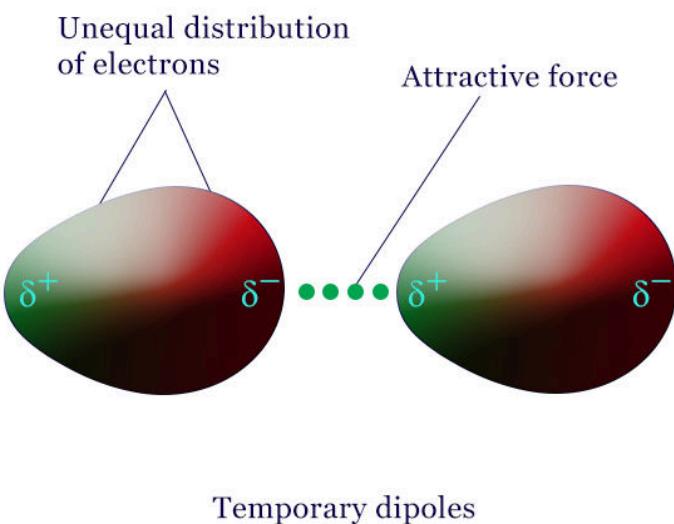
London Dispersion Forces

Dipole-Dipole Interactions

Practice: Strongest Intermolecular Force

Choose the strongest intermolecular forces in the following compounds.

Part 8



Ionic Bonds

Hydrogen Bonds

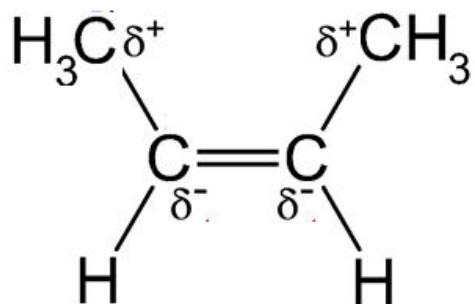
London Dispersion Forces

Dipole-Dipole Interactions

Practice: Strongest Intermolecular Force

Choose the strongest intermolecular forces in the following compounds.

Part 9



Ionic Bonds



Hydrogen Bonds



London Dispersion Forces



Dipole-Dipole Interactions



3.2

Other Intermolecular Forces (more detailed)

3.2.1

Types of Intermolecular Forces

Ion - Ion

- Oppositely charged ions held together by electrostatic attraction
- aka Ionic bonding (Ionic bonding is an intramolecular bond between a non-metal and a metal)
- Strongest intermolecular force

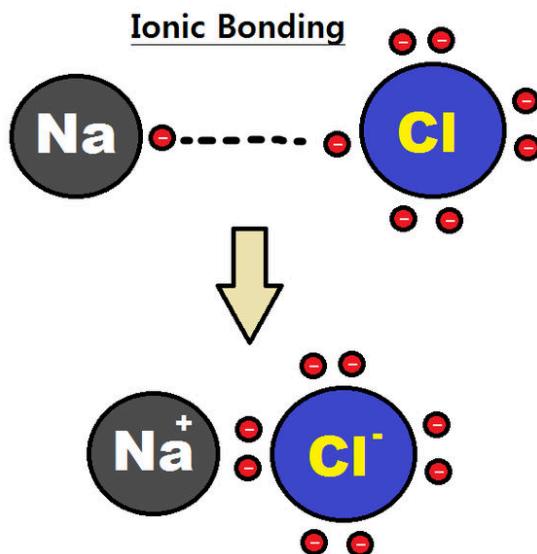
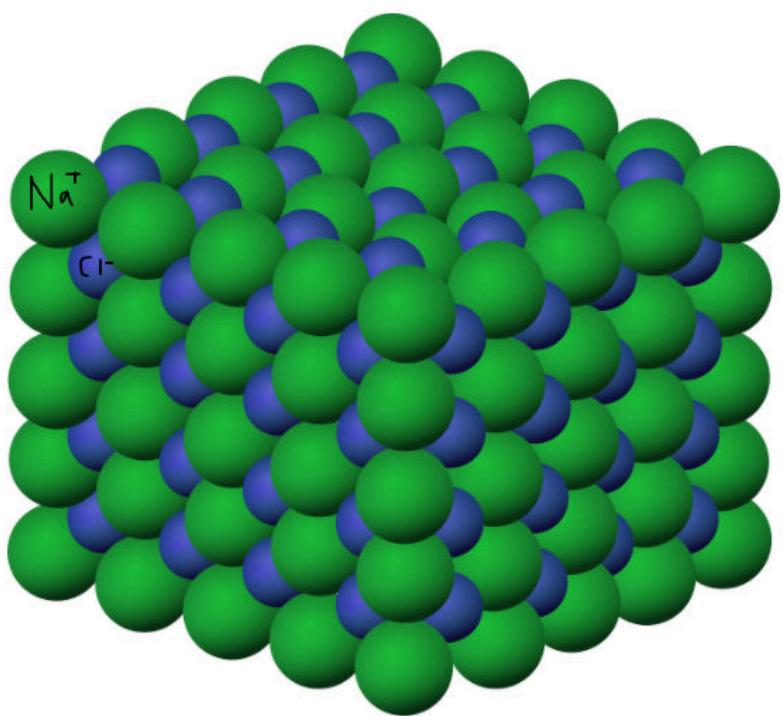
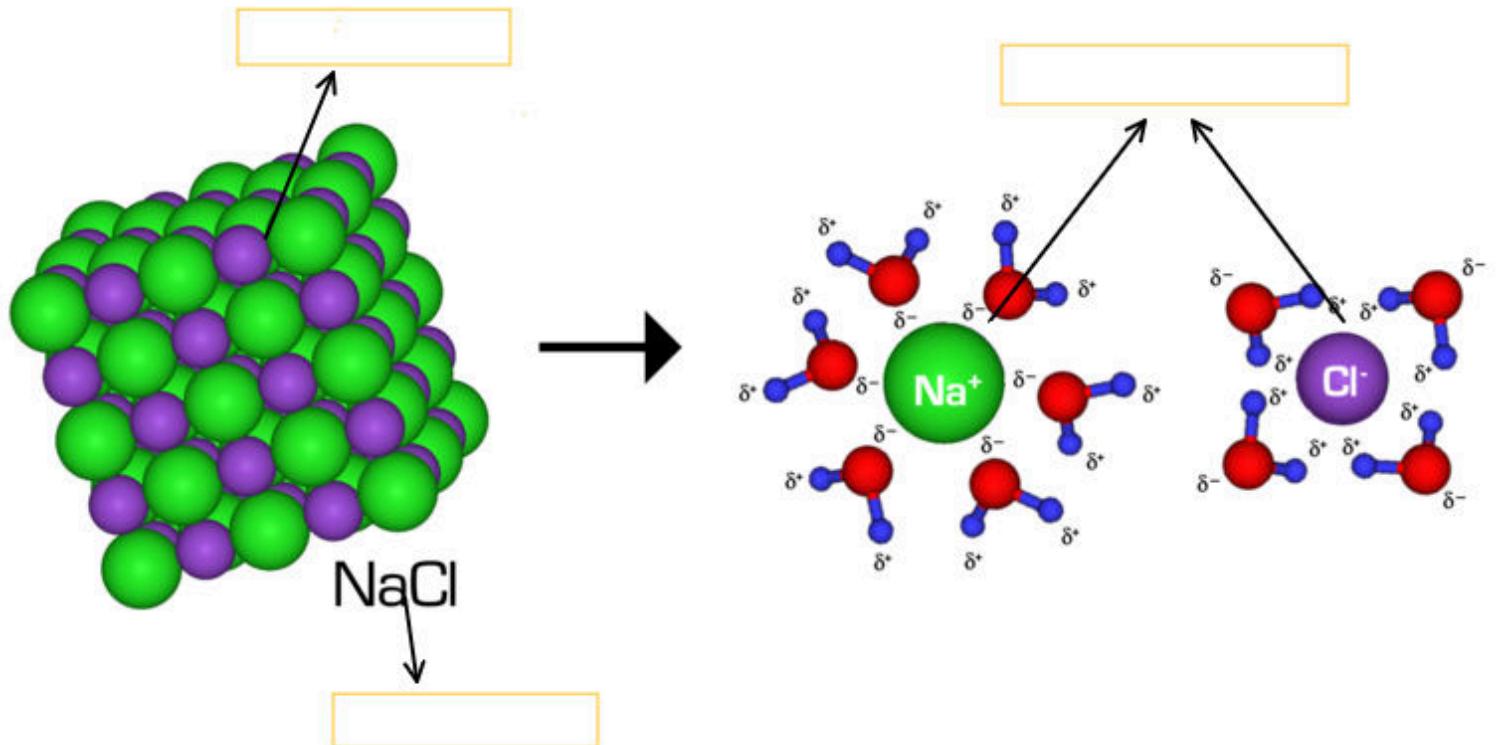
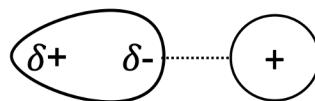


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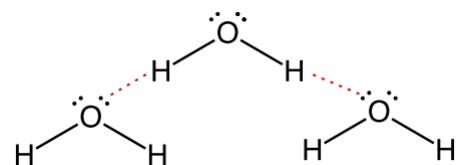
Ion - Dipole

- Involves an **ion** interacting with a **polar molecule** (molecule with a net dipole)
 - A **positively charged ion** will interact with the **$\delta-$ side of the polar molecule**
 - A **negatively charged ion** will interact with the **$\delta+$ side of the polar molecule**



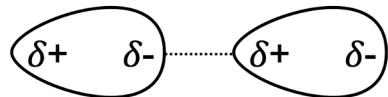
Hydrogen Bonding

- Hydrogen bonded to N, O, or F is attracted to lone pairs of electrons on other N, O or F



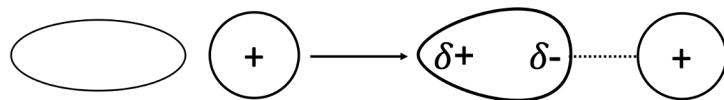
Dipole - Dipole

- Interaction between **two polar molecules**
- Opposite dipoles attract ($\delta-$ and $\delta+$)



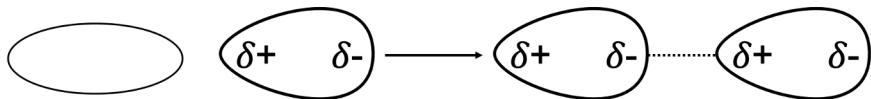
Ion - Induced Dipole:

- **Ion** causes a **temporary dipole in non-polar species**
- In the example below, the positively charged ion makes the non-polar species shift its electrons towards the positive charge, creating an uneven charge distribution in the non-polar species



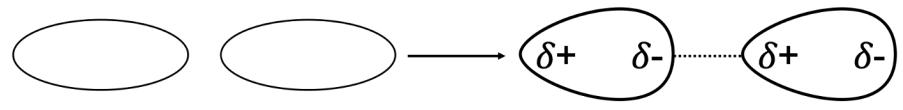
Dipole - Induced Dipole

- **Polar molecule** causes a **temporary dipole in non-polar species**
- Similar to ion-induced dipole, except that in dipole-induced dipole it is a polar molecule instead of an ion that induces a temporary dipole in the non-polar species
- Weaker than ion-induced dipole forces



Induced Dipole - Induced Dipole

- Molecules interact randomly and distort each other's electron clouds, causing temporary dipoles
- **Electrons randomly move** around and at one point, **more electrons may be on one side than the other** in a non-polar molecule, creating a **temporary dipole**. That can happen with another non-polar molecule as well and the two can interact
 - We call this random movement of electrons **polarizability**
- aka **London dispersion forces** (all molecules have this)



3.2.2

Intermolecular Forces

Intermolecular forces = forces between molecules

Force	Nature of Attraction	Example	Relative Strength	Functional Groups Commonly Involved
Ion-ion (AKA ionic)	• Attraction between positively-charged and negatively-charged ion	$\text{Na}^+ \cdot \text{Cl}$ (Table salt!)	Very strong	Charged molecules, ions
Ion-dipole	• Interaction with a charged ion and a molecule with a dipole		Strong	A molecule with a strong dipole (e.g. H_2O) interacting with an ion (Na^+ , Ca^+ , K^+ , Br^- , Cl^- , ...)
Hydrogen bond	• A strong type of dipole-dipole interaction between an acidic hydrogen and a lone pair		Medium	Alcohols (CH_3OH) Carboxylic acids (COOH) Amines (NH_2)
Dipole-dipole	• Two molecules with strong dipoles are attracted to each other	$\text{H}-\text{Cl} \cdots \text{H}-\text{Cl}$	Weak	Compounds that contain a polar bond
Dispersion (AKA London Forces)	• All compounds have these • Attraction of weak dipoles that form at any given time when molecules are close • The primary intermolecular force in gases		Very weak	Alkanes (CH_3) Alkenes (CH_2CH_2) Benzene Molecules with zero net dipole (I_2 , F_2 , ...)



3.2.3

Example: Recognizing Intermolecular Forces

Choose the strongest type of interaction two of the same molecule pictured to the left as Lewis structures, will have with one another

	London Dispersion Forces	Dipole – Dipole	Hydrogen Bonding	Dipole – Induced Dipole	Ion - Dipole

Example: Strongest Interaction Between Pairs of Molecules

Predict what the strongest type of interaction will be between the following pairs of molecules.

a) CH_4 and H_2O

b) NH_3 and HF

c) NaCl and H_2O

d) H_2 and Cl_2

3.2.5

Practice: Ranking Intermolecular Interactions

Based on the three examples below, rank the strength of the intermolecular interaction from highest to lowest.

- A) NH₃ with NH₃ B) CH₄ with Ne C) Ca²⁺ with H₂O

A > B > C

B > A > C

C > A > B

A > C > B

C > B > A

Practice: Intermolecular Forces

A student mixes HF with PCl_3 at 0°C and forms a liquid mixture. What types of intermolecular forces are present between the two components of this liquid mixture, HF and PCl_3 ?

Ion-Dipole, Hydrogen Bonding and Dispersion Forces

Hydrogen Bonding

Dipole Dipole Forces and Dispersion Forces

Ion-Dipole Forces and Dispersion forces

Hydrogen Bonding and Dispersion forces

3.3 Physical Properties

3.3.1

Intermolecular Forces and Boiling Point

Intermolecular forces define physical properties of compounds (boiling points, melting points, solubility, vapour pressure, viscosities etc)

Boiling Points/Melting Points

In general, if a molecule has the stronger intermolecular forces, would its boiling point be higher or lower?

When trying to determine which molecule has the higher boiling point, we can consider their intermolecular forces.

Example #1: Which of the following molecules has the higher boiling point?



 **WIZE TIP**

The **stronger the intermolecular forces, the higher the boiling point.**

(Highest boiling point) Intramolecular bonding (covalent/ionic) > Hydrogen bonding > Dipole-dipole forces > London dispersion forces **(Lowest boiling point)**

Example #2: Which of the following has the highest bp?



When trying to determine which molecule has the higher bp, what happens when two compounds have the same types of intermolecular forces?

Example #3: Which of the following has the highest bp?



The strongest intermolecular forces that the molecule on the left has are:

The strongest intermolecular forces that the molecule on the right has are:

 **WIZE CONCEPT**

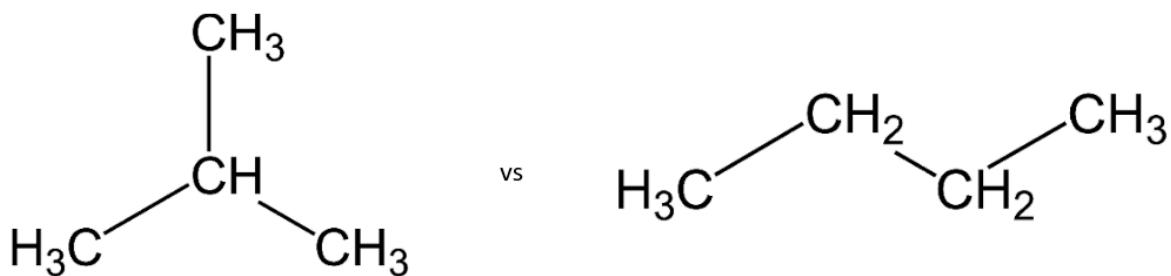
If two compounds have identical intermolecular forces, the larger molecule will have a higher mp/bp due to increased London dispersion forces forces

Higher MW=More london dispersion forces

Therefore, the molecule on the (left/right) _____ has the stronger intermolecular forces and (lower/higher) _____ boiling point!

When trying to determine which molecule has the higher bp, what happens when both compounds have the same number of carbon atoms and the same forces?

Example #4: Which of the following has the higher bp?



The strongest intermolecular forces that the molecule on the left has are:

The strongest intermolecular forces that the molecule on the right has are:

The molecule on the left has _____ C atoms and the molecule on the right has _____ C atoms.

Analogy: Tree trunk vs branches



Other Physical Properties and Intermolecular Forces

Changes of Phase

Solids have stronger intermolecular forces than liquids and gases.

Solids > Liquids > Gases

Solubility

When one molecule is dissolved in another molecule, typically if the two species have similar types of intermolecular forces, they will mix (be soluble); if not, they will tend not to mix very well (be insoluble).

In general, "**like dissolves like**"

- Polar molecules are soluble in other polar molecules (but not in non-polar molecules)
- Non-polar molecules are soluble in other non-polar molecules (but not in polar molecules)

Example: Phospholipid bilayer of a cell

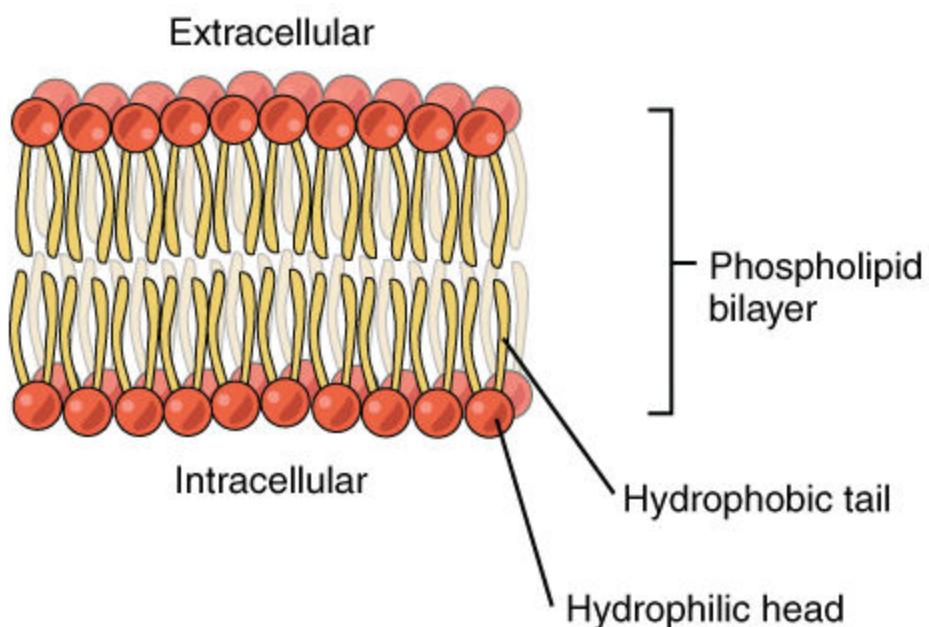


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Hydrophilic heads are "water-loving" and contain polar molecules
Hydrophobic tails are "water-fearing" and contain non-polar molecules

Vapor Pressure

Vapor pressure: pressure exerted by the gaseous phase of a liquid that evaporated from the exposed surface of a liquid

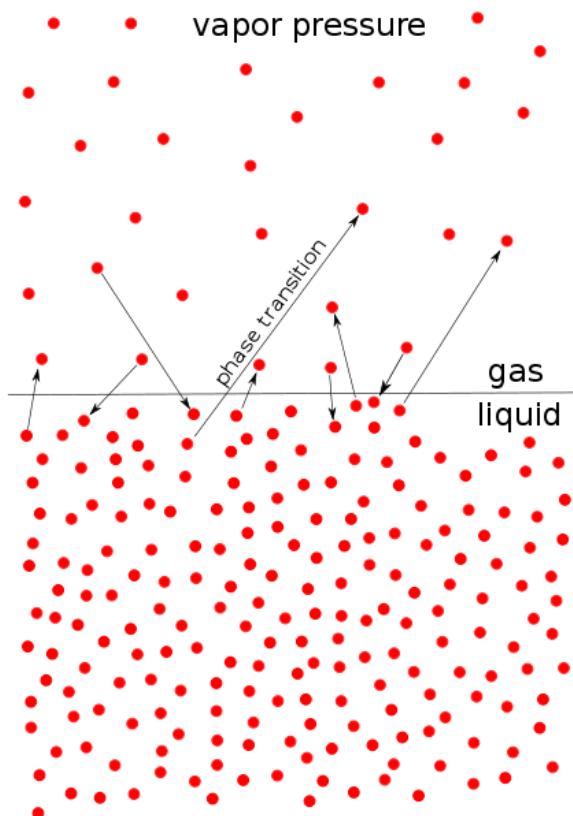


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The **lower the intermolecular forces** the higher or lower the vapor pressure?

We call liquids with high vapor pressures **volatile**.

The last thing to note is temperature's effect on vapor pressure:

- As temperature increases, do intermolecular forces increase or decrease?
- As temperature increases, does the vapor pressure increase or decrease?

Example: Highest Melting Point

Determine which of the following pairs of molecules will have the higher melting point.

a) I_2 or Br_2

b) CO_2 or NO_2^-

c) H_2O or H_2S

d) $NaCl$ or NH_3

Example: Volatility

Which of the following would be expected to be the most volatile?

- A) H₂O at 273 K
- B) NH₃ at 373 K
- C) H₂O at 373 K

Practice: Boiling Point and Size

The boiling point of hydrocarbons increases as the hydrocarbon gets more carbon atoms added to it. Methane CH₄ (-161.5°C), ethane C₂H₆ (-89°C), propane C₃H₈ (-42°C) This can best be explained by:

More hydrogen bonds are present in propane than methane which increases the intermolecular forces

The larger number of atoms leads to a higher weight which makes it harder to form vapour

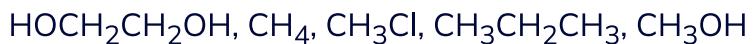
The larger mass of the hydrocarbons leads to stronger dispersion forces which are the strongest intermolecular forces in hydrocarbons

Propane is more flammable so its the least volatile

None of the above

Practice: Ranking Boiling Points

Please rank the following compounds from lowest boiling point to highest boiling point.



A) $\text{CH}_3\text{Cl} < \text{CH}_4 < \text{CH}_3\text{CH}_2\text{CH}_3 < \text{CH}_3\text{OH} < \text{HOCH}_2\text{CH}_2\text{OH}$

B) $\text{CH}_4 < \text{CH}_3\text{CH}_2\text{CH}_3 < \text{CH}_3\text{Cl} < \text{HOCH}_2\text{CH}_2\text{OH} < \text{CH}_3\text{OH}$

C) $\text{CH}_4 < \text{CH}_3\text{CH}_2\text{CH}_3 < \text{CH}_3\text{Cl} < \text{CH}_3\text{OH} < \text{HOCH}_2\text{CH}_2\text{OH}$

D) $\text{CH}_4 < \text{CH}_3\text{Cl} < \text{CH}_3\text{CH}_2\text{CH}_3 < \text{CH}_3\text{OH} < \text{HOCH}_2\text{CH}_2\text{OH}$

3.4 Types of Solids

3.4.1

Types of Solids

There are two major types of solids: crystalline and amorphous

- In an **amorphous solid**, the particles are randomly distributed.
- In **crystalline solids**, particles are uniformly arranged.
 - Crystalline solids can be further classified into **molecular, ionic, metallic and covalent solids**.

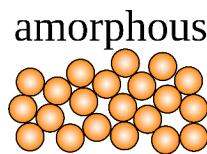
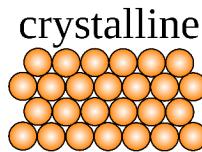
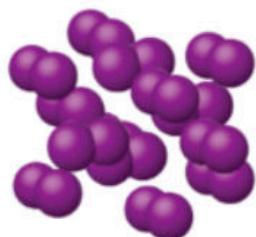
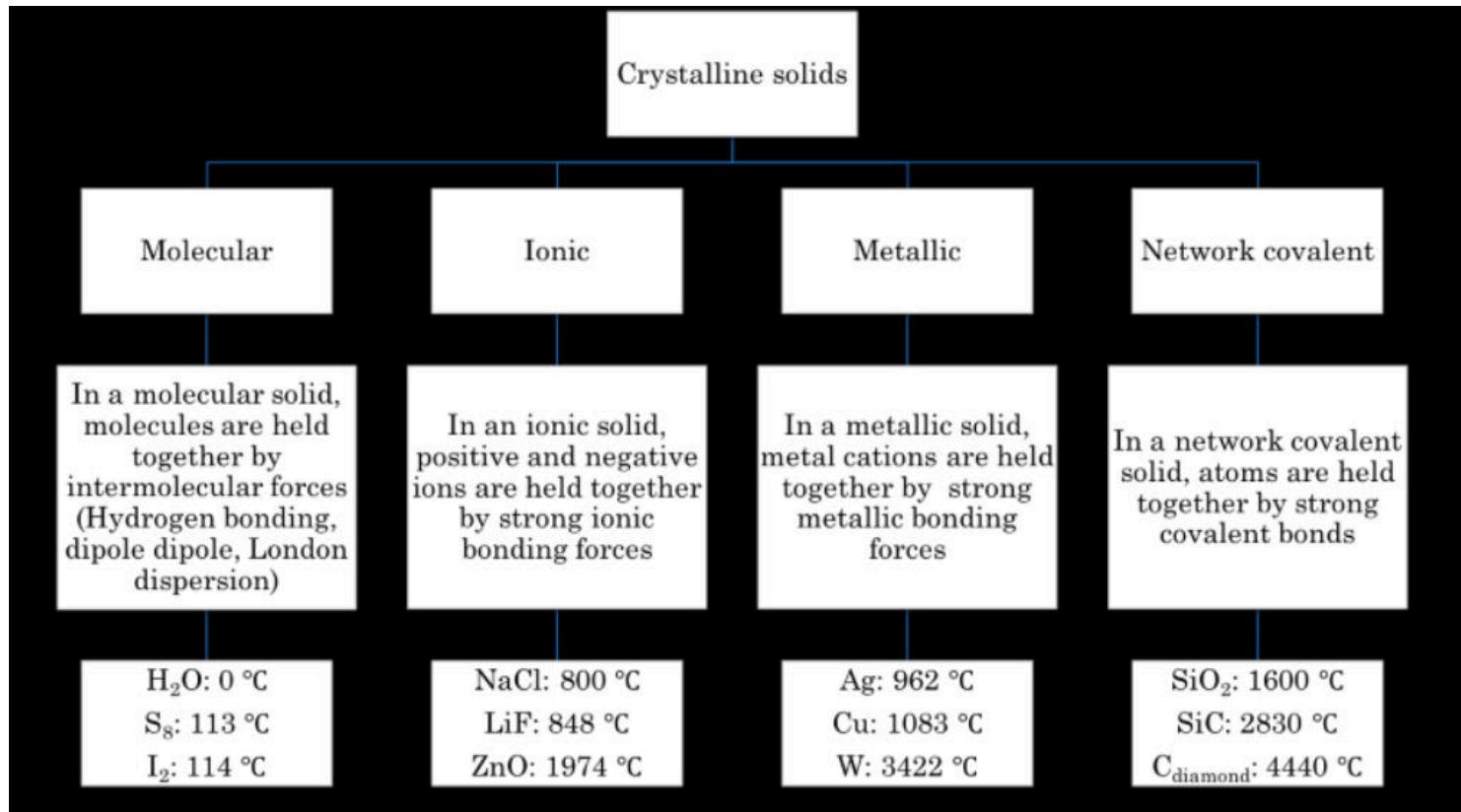
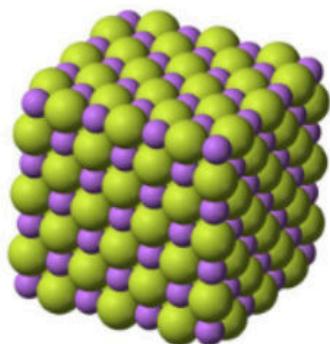


Photo by cDang / CC BY

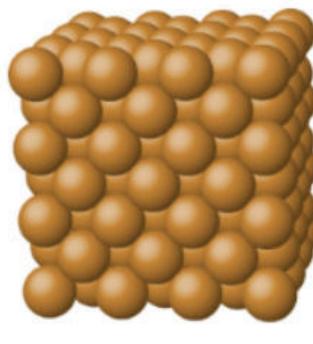
Crystalline Solids



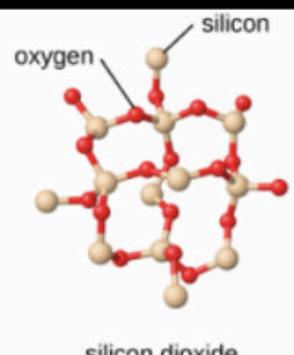
Iodine



LiF



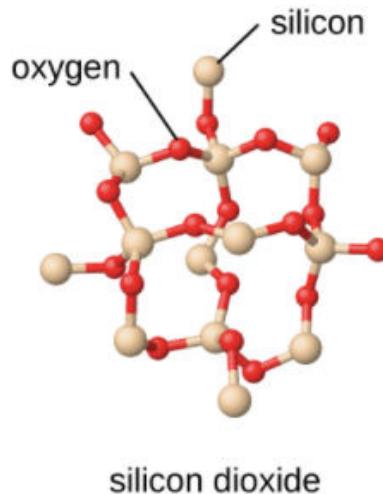
Cu



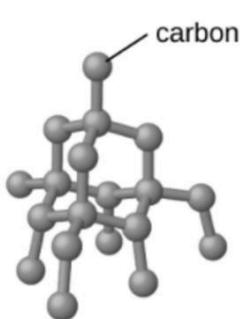
silicon dioxide

First, third, and last photo by Openstax / CC BY

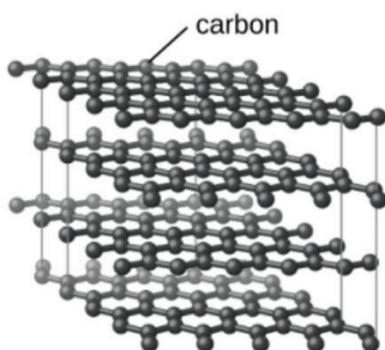
Network Covalent Solids



- **Network Covalent solids:** atoms are held together by strong covalent bonds and they have very high melting points.
Example: C-diamond and C-graphite.
 - They are **allotropes:** same chemical composition but different structural arrangement.
- Diamond (C) are a 3D array of atoms with tetrahedral geometry, and therefore are sp³ hybridized. These materials are very hard and have high melting points.
- Graphite consists of carbon atoms arranged in sheets of planar hexagonal rings, and therefore are sp² hybridized. Weak London dispersion forces hold the layers of graphite together. Graphite is a good lubricant.

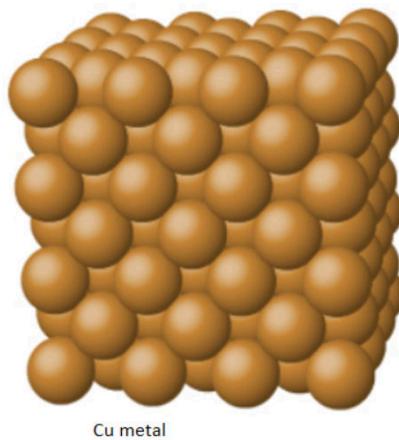


diamond



graphite

Metallic Solids



Cu metal

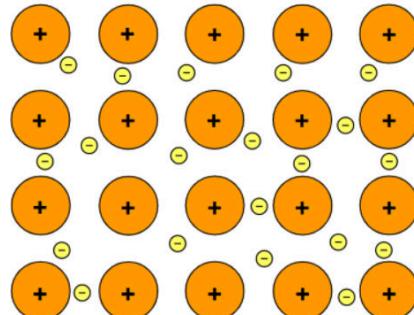
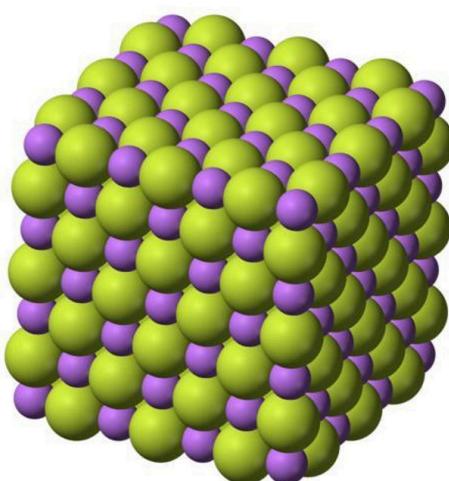


Photo by Muskid / CC BY

- Metal atoms held together by metallic bonds; sea of **delocalized electrons** free to move throughout the solid.
- High electrical and thermal conductivity, malleable and ductile
- Melting point is different for each metal.
 - The **greater the delocalization of electrons and charge of + cation**, the stronger interactions and **higher melting point**.

Example: Aluminium (3 electrons) MP = 660°C whilst Sodium (1 electron) MP = 98°C

Ionic Solids



Example: LiF(s)

- Solids made of ions held together by **ionic bonds** (electrostatic interactions between negatively charged ions and cations).

Example: NaCl (sodium chloride - table salt)

- Ionic solids are hard, brittle, and have **high melting points**. When **dissolved in water they conduct an electrical current**. If the ionic solid has not dissolved, the solid is a poor conductor of electricity.
- The **larger the charge between the two ions** and the **smaller the bond distance**, the **greater the ionic bond strength**

Types of Solids and Their Properties

Types of Crystalline Solids and Their Properties				
Type of Solid	Type of Particles	Type of Attractions	Properties	Examples
ionic	ions	ionic bonds	hard, brittle, conducts electricity as a liquid but not as a solid, high to very high melting points	NaCl, Al ₂ O ₃
metallic	atoms of electropositive elements	metallic bonds	shiny, malleable, ductile, conducts heat and electricity well, variable hardness and melting temperature	Cu, Fe, Ti, Pb, U
covalent network	atoms of electronegative elements	covalent bonds	very hard, not conductive, very high melting points	C (diamond), SiO ₂ , SiC
molecular	molecules (or atoms)	IMFs	variable hardness, variable brittleness, not conductive, low melting points	H ₂ O, CO ₂ , I ₂ , C ₁₂ H ₂₂ O ₁₁

Photo by OpenStax / CC BY

Example: Crystalline Solids

Identify the following crystalline solids.

A. NaCl

B. Au

Example: Types of Solids

Will network covalent solids melt at a higher or lower temperature than molecular solids? Why?

Practice: Solids

Which identification is INCORRECT for the following substances as solids?

A) SiC is a covalent-network solid

B) Li_2CO_3 is an ionic solid

C) SiO_2 is a covalent-network solid

D) Diamond is a molecular solid

3.5

Phases of Matter

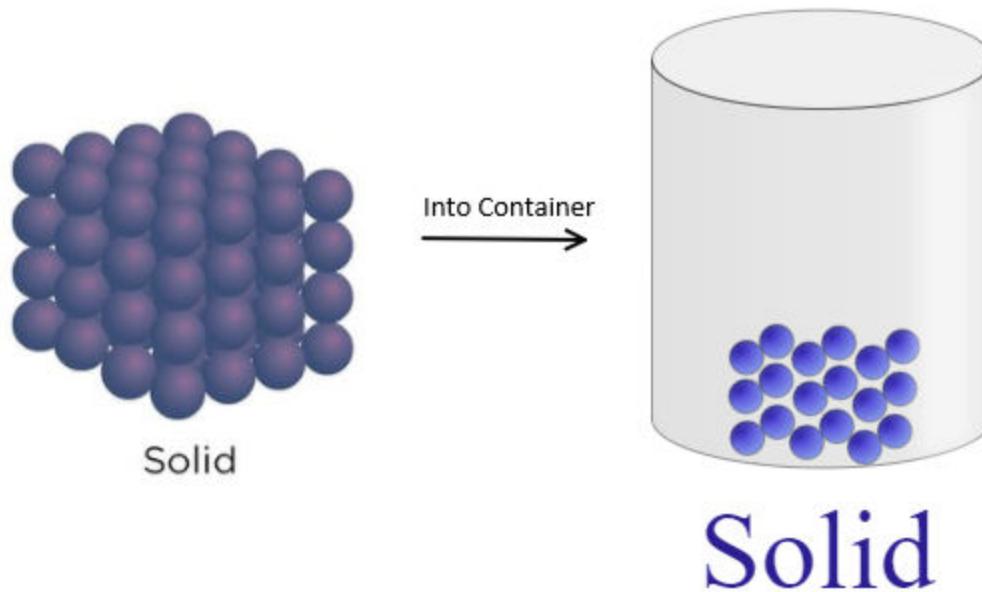
3.5.1

Phases of Matter

There are **3 phases of matter**: solid, liquid, and gas. We need to have a general understanding of each one to understand how a phase change occurs. For example, for the phase change of a solid → liquid, energy is required to break apart some of the strong intermolecular forces that solids have!

Solid

- Don't fill the volume or take the shape of their container (have a fixed volume and shape)
- Hard
- Are not fluid like liquids and gases (don't flow unless extreme shearing/stretching forces)
- Has the **strongest intermolecular forces**

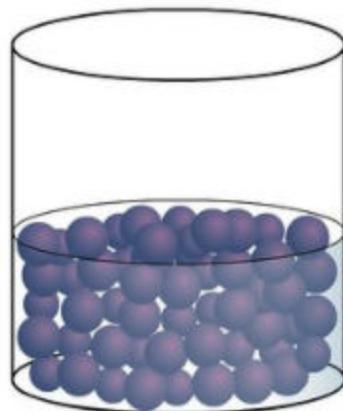


Left photo by Siyavula Education / CC BY

Right photo by Yelod - Wikipedia Commons / CC BY

Liquid

- Don't fill the volume of the container, but take the shape of their container
 - Fixed volume, not shape since if you transfer the liquid to a different container the volume of liquid will remain the same but the shape of the liquid will change depending on the shape of the container
- Fluid
- Has the **weaker intermolecular forces**

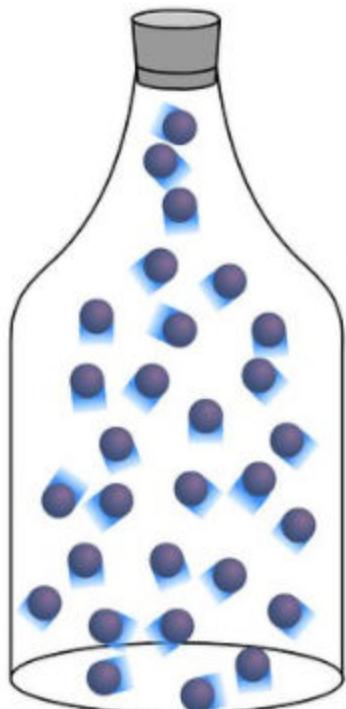


Liquid

Photo by Siyavula Education / CC BY

Gas

- Fill the volume and take the shape of their container
 - Volume and shape is not fixed, both can change when you change the container that is holding the gas!
- Fluid
- Has the **weakest intermolecular forces**
 - Note how the molecules are very spread apart from each other. This is because the intermolecular forces between them are very weak!



Gas

Photo by Siyavula Education / CC BY

3.6 Ideal Gas Law

3.6.1

The Ideal Gas Law

The equation below **describes ideal gases**. Real gases act differently, but act most similarly to ideal gases at **high temperatures and low pressures!**

$$PV = nRT$$

P is **pressure** measured in kPa (or maybe atm Torr, mmHg...)

V is **volume** measured in **L** or **m^3**

n is **number of moles** of the gas

R is the **ideal gas constant** (will be either **8.314 J/mol K** OR **0.082 L atm/mol K**)

T is **temperature** measured in **K**

 **WIZE TIP**

Pressure=Force/unit area=N/m²=1 Pa

How to convert different units of pressure:

$$101\ 300\ Pa = 101.3\ kPa = 1\ bar = 1\ atm = 760\ mmHg = 760\ torr$$

We mentioned that R (the ideal gas constant) could be either 8.314J/mol K or 0.082 L atm/mol K. **How do we know which value to use?**

! WATCH OUT!

Using the correct value for R is more important than you might think! Students often make the mistake of using the wrong value of R on an exam and they get the question wrong. Let's see how we can easily prevent that :)

The ideal gas constant appears frequently in chemistry, for this reason, it comes in many forms depending on the context of its use and the **units involved**:

Common Values for R
8.3145 kPa · L / mol · K
8.3145 N · m / mol · K
8.3145 J / mol · K
8.3145 kg · m ² /s ² mol · K
0.08206 atm · L / mol · K

Example:

Suppose the variables in the ideal gas law equation had the following units:

$$P(kPa) \times V(L) = n(mol) \times R(?) \times T(K)$$

What should R be in this case? (These terms need to combine correctly to ensure that this equation makes sense!)

3.6.2

Example: Ideal Gas Law

1.1 mols of Argon are stored in a 2.0 L container kept at 10.0 °C. What is the pressure of this gas?
(ans. in kPa)

Example: Ideal Gas Law

- a) A 4 L cylinder containing 3 moles of He(g) arrives from a chemical supply company. Given that the lab you are working in has a room temperature of 298 K, calculate the pressure inside the flask.

- b) At what temperature will the cylinder have an internal pressure of 9 atm?

- c) Your lab mate steals some of the gas from your cylinder for a reaction he is doing. He fills a 1 L bulb with 1 atm of your gas at 298 K. If the cylinder is at room temperature, what is the internal pressure of the cylinder once the gas has been removed?

3.6.4

Practice: Stoichiometry and the Ideal Gas Law

The combustion of methanol (CH_3OH) is shown below. If 16g of methanol is burned in oxygen, what volume of CO_2 is produced if the pressure is 101.3kPa, $T=25^\circ\text{C}$, and $R=8.314 \text{ kPa L/mol K}$? Round to the nearest whole number.



6L



12L



24L



27L



3.7 Gas Laws

3.7.1

Gas Laws

Contained within the ideal gas law are a series of simpler relationships describing how pairs of gas properties change when others are held constant. We will go through each equation and understand how to get each equation without having to memorize them! 🤓

Name	Conditions	Show Your Work to Find Equation	Equation	Are the variables (inversely/directly) proportional
Boyle's Law	Constant T & n			
Charles' Law	Constant P & n			
Gay-Lussac's (aka Amonton's) Law	Constant V & n			
Avogadro's Law	Constant P & T			
Combined Gas Law	Constant n			

Example: Gas Laws

A sample of argon is loaded into a 3 L sealed vessel and at room temperature the initial pressure is 2.6 atm. The vessel is then loaded into an autoclave and heated to 580 °C. What is the final pressure inside the vessel?

Practice: Gas Laws

If the pressure inside a container is kept constant and the container is heated, what would happen to the volume of the gas?

volume of the gas expands

volume of the gas contracts

volume of the gas expands then contracts

more information would be required to answer this question

3.7.4

Practice: Gas Laws

A 15 L sample of N₂ gas held at 328 K is compressed such that the pressure rises from 1.0 atm to 7.8 atm. What is the final volume of the gas?

0.02L



2L



8L



13L



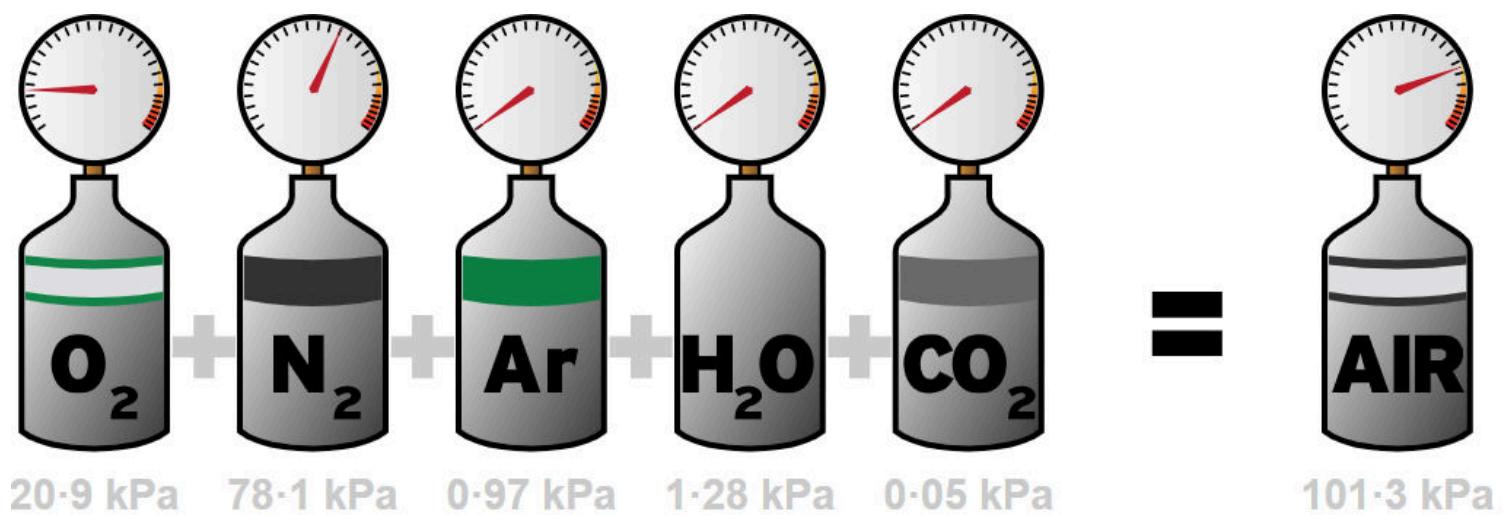
3.8

Gas Mixtures & Partial Pressures

3.8.1

Gas Mixtures & Partial Pressures

When dealing with mixtures of gases, the total pressure of the gases is equal to the sum of the **partial pressure** of each component.



Dalton's Law of Partial Pressures:

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

P_{Total} is the single **measured pressure of the mixture**

P_1 is the **partial pressure of gas 1**, and so on

Partial pressure: is the pressure that would be exerted by one of the gases in the mixture if it occupied the same volume on its own.

Note: The total pressure depends on the total number of moles (not on the identity of the gases)

$$PV = nRT$$

$$P_{total} = \frac{n_{total}RT}{V}$$

We can also relate the partial pressure of an individual gas (ex. Gas "A") to the total pressure by the **mole fraction (X)** of that gas.

The **mole fraction (X)** is a **proportion** of that gas.

$$P_A = \frac{n_A}{n_{Total}} P_{Total} = X_A P_{Total}$$

Example: Partial Pressures

If we have a container with 75% nitrogen, 20% oxygen, and 0.04% carbon dioxide, the total pressure in the container is measured to be 760 Torr (recall this is equal to 1 atm!). What is the partial pressure of oxygen gas? (Ans in torr)

3.8.3

Practice: Partial Pressures

A vessel contains He, Ne, and Ar gas. There's twice as much He as Ne and half as much Ar as He. The total pressure in the vessel is 400torr. What is the partial pressure of Ar(g)?

75 torr

100 torr

175 torr

225 torr

3.8.4

Practice: Partial Pressures

Two gases were mixed together in a 5 L vessel at 243 K to create a total pressure of 9.3 atm. If gas 1 has 2 moles, what is the partial pressure of gas 2? ($R = 0.08206 \text{ L}^*\text{atm}/\text{mole}^*\text{K}$)

0.7 atm



1.3 atm



3.1 atm



5.4 atm



3.9

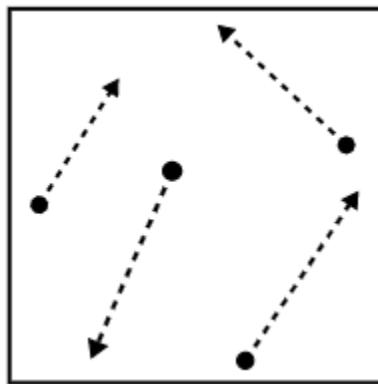
Kinetic Molecular Theory of Gases

3.9.1

Kinetic Molecular Theory (KMT)

This theory describes the behaviour of **ideal gases**.

In normal conditions we can apply this to real gases as well...(more on the conditions required for this later)



The KMT operates under the following assumptions/simplifications:

 WIZE CONCEPT

1. Ideal gases are in **constant, random motion** and move in straight lines
2. The size of gas particles is negligible compared to the distance between particles. As a result of this, the **volume of gas particles is considered to be 0**.
3. There is constant motion of particles and **collisions** of gas particles (with other gas particles and with the walls of the container) **are completely elastic** so that the total kinetic energy is conserved before and after a collision.
4. There are **no intermolecular forces** between gas particles (they don't attract or repel each other).
5. **Temperature is directly proportional to the average kinetic energy** of the gas particles.

- If we increase temperature, do you think the gas molecules would collide with container walls more/less often? _____ often.
- And with an increase in temperature would the gas particles collide at slower or greater speeds?
_____ speeds

! WATCH OUT!

T should always be in Kelvins (K) for chemistry calculations :)

$$T(\text{in K}) = T(\text{in } ^\circ\text{C}) + 273$$

i WIZE TIP

You can just assume we are dealing with an ideal gas unless stated otherwise in a particular exam question!

Note: to measure pressure($P=F/A$), we need to take into consideration the **number of collisions and force of those collisions between gas particles and the walls of the container**

This can help us understand WHY the KMT says what it does for IDEAL gases:

- If gas particles had volume they would be quite large and would have more/less: _____ intermolecular forces/attractions to one another. This would mean that they would interact with the walls of the container more/less: _____ and result in a decrease/increase: _____ in pressure.
- If collisions weren't elastic, then over time gas particles would hit the walls of the container with more/less: _____ force and increase/decrease: _____ the pressure.
- We will explain this concept more when we look at real vs ideal gases!

The KMT for ideal gases helps to keep conditions in the container stable!

3.10 Kinetic Energy, Temperature, and Speed of Gas Particles

3.10.1

Relationships Between Kinetic Energy, Temperature, and Speeds!

Just focus on **understanding**  the following lesson.

Scenario #1:

We have two different gases at the **same temperature** in two different containers. What can we say about the **average kinetic energy** of the two gases? (are they the same, different, if different how?)

Hint: This was the part of the kinetic molecular theory!

We can make sense of this using the following equation:

$$KE_{avg} = \frac{3}{2}RT$$

KE_{avg} is the average kinetic energy of a mole of gas particles

R here is 8.314 J/mol K

How are KE_{avg} and T related?

! WATCH OUT!

The molar masses of the gases are not important when it comes to the average kinetic energy so if you're ever asked about the average kinetic energy, you just need to think about how it relates to temperature!

Scenario #2:

We have two different gases at the same temperature in two different containers. What can we say about the speeds of the two gases? (are they the same, different, if different how?)

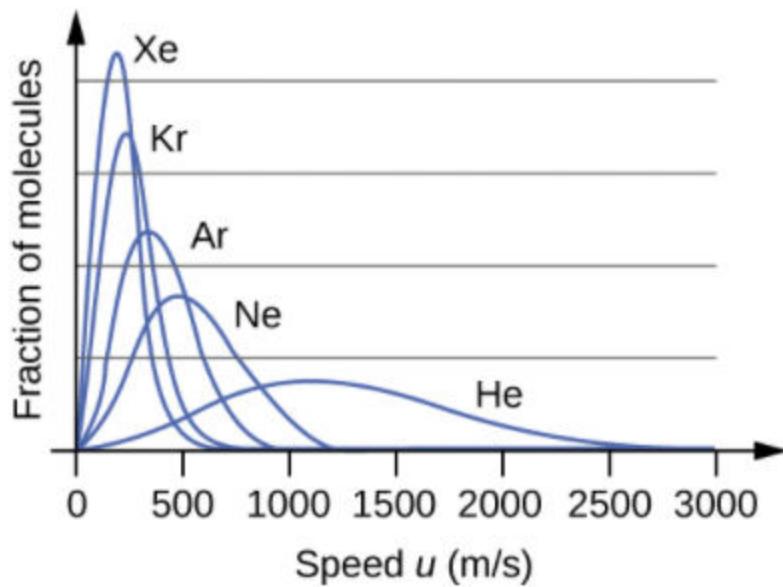


Photo by CNS OpenStax / CC BY

Therefore, a heavier gas will move at a (faster/slower) _____ average speeds than a lighter gas at the same temperature.

Note: The lighter the gas molecule, the more (broad/narrow) _____ the speed distribution

Scenario #3:

If we just consider one gas in a container, **all of the gas particles in the container don't move at the exact same speed!**

There is actually a distribution of speeds.

If temperature gets increased, then the average speed (increases/decreases) _____ and the average kinetic energy (increases/decreases) _____ and the spread in the graph below (increases/decreases) _____.

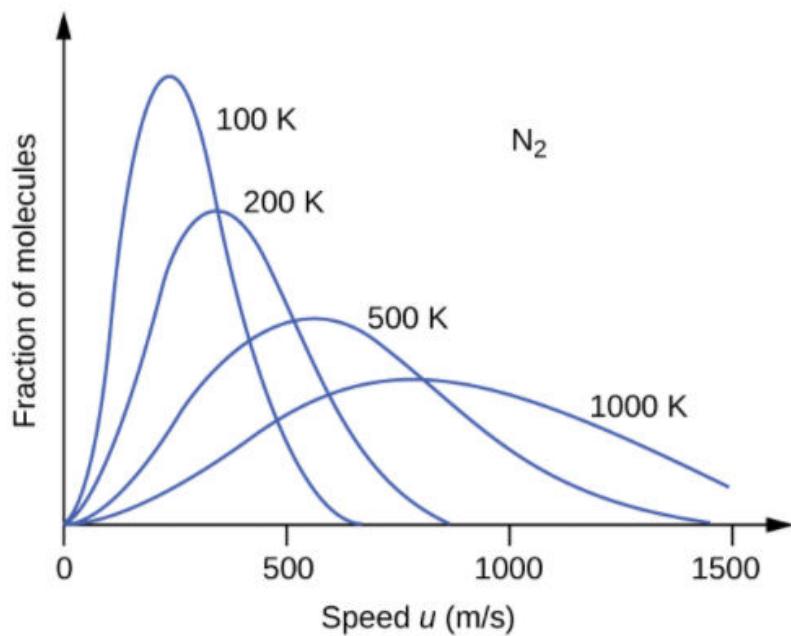
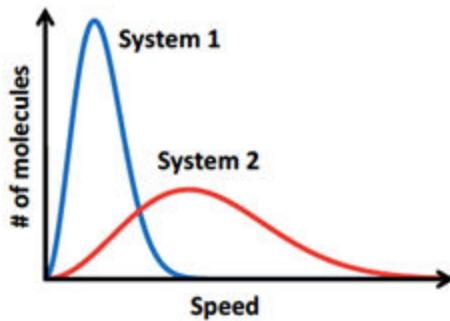


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Example: Applications of Kinetic Molecular Theory

Consider the following two curves showing distributions of the number of molecules vs. their molecular speeds for two different gas samples, System 1 and System 2. If the area under both curves is equal, provide two possible scenarios which would result in these observed differences. Complete the table below to describe each of your two scenarios.



	Scenario 1	Scenario 2
What single property could be different between System 1 and System 2?		
For the property you chose, which system has a higher value?		
Sketch what the kinetic energy distribution would look like for System 2.		

3.10.3

Practice: Root Mean Squared Speed

Which of the following would have the lowest u_{rms} ?

A) NO₂ at 273 K

B) NO₂ at 453 K

C) O₂ at 273 K

Practice: Kinetic Energy of Gases

Which of the following have the same average kinetic energy?

- I He at 297 K
- II He at 351 K
- III H₂ at 297 K
- IV O₂ at 212 K
- V SO₂ at 297 K

I and II

II and IV

I, III, and V

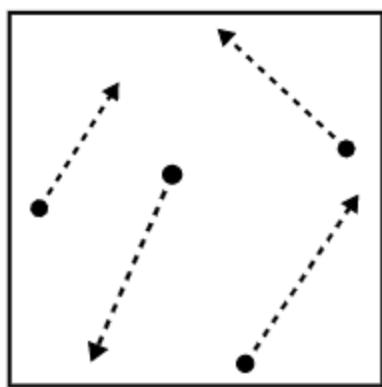
all of the above

3.11 Real vs Ideal Gases

3.11.1

Real Gases and the Van der Waals Equation

Ideal Gases



i WIZE TIP

Ideal gases have **negligible volume**.

And there are **no intermolecular forces** between gas particles.

Note: $PV=nRT$ is the **IDEAL gas law** and describes **IDEAL gases**

Real Gases

WIZE TIP

Real gases **have volume**.

And they **have intermolecular forces** (attraction between gas molecules).

Real gases are most like ideal gases when temperature is high (well above the condensation temp for that gas), pressure is low (usually less than 1 atm)

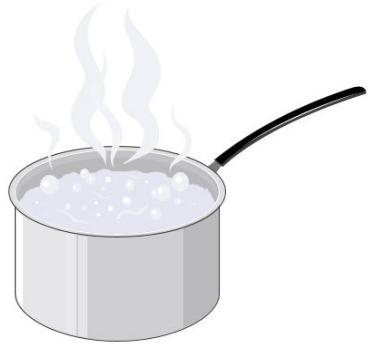


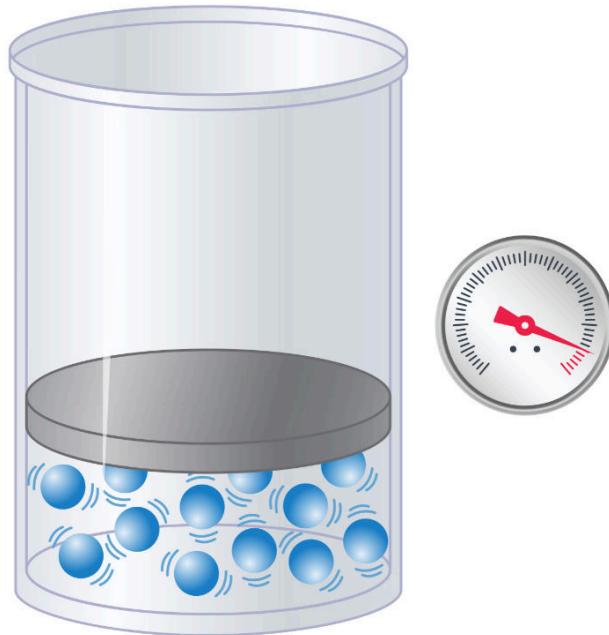
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WIZE CONCEPT

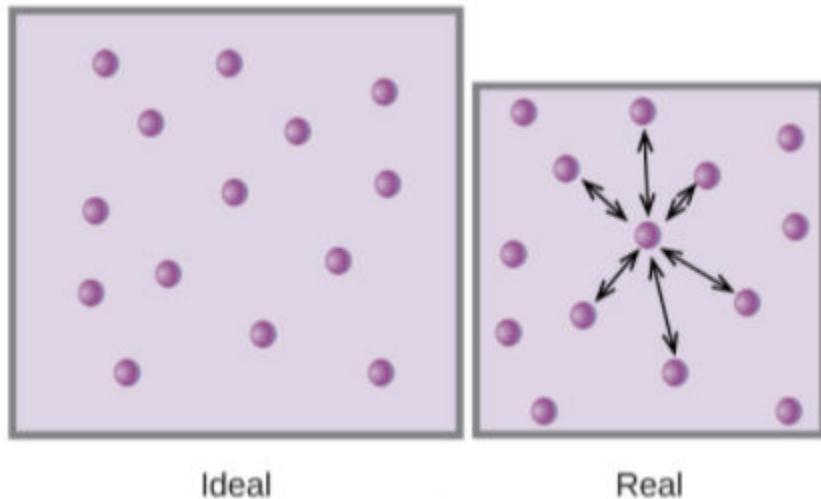
Real gases deviate more from ideal gases when temperature is low (near bp) and pressure is high ($\sim >10\text{ atm}$)

- With a lower temperature and higher pressure just think that we are going more towards liquids than gases
- High pressures help to compress and push molecules closer together, resulting in less distance between molecules. As a result, the volume of the gas particles becomes more significant.
 - Both of these factors (low temp and high pressure) cause gas molecules to interact more with each other (have intermolecular forces) and interact less with the walls of the container (decreasing pressure).



High pressure

Ideal vs Real Gas in a Constant Pressure Container (Volume of the Container Can Change)



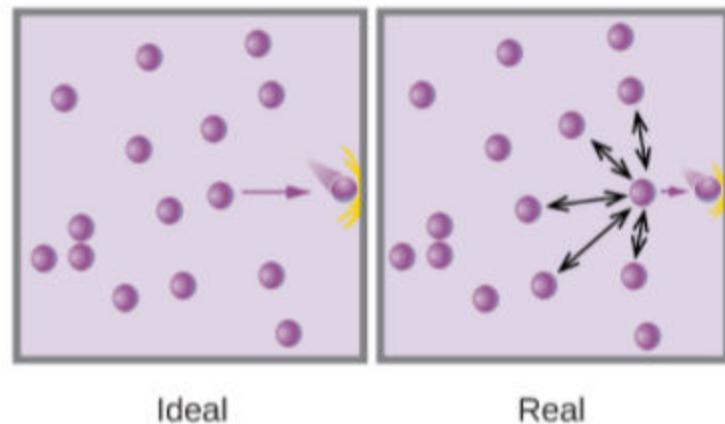
Ideal Gas

- Has no intermolecular forces between gas particles
- Has no volume
- Container remains unchanged

Real Gas

- Real gases have a significant volume!
- Since they **have volume**, real gas particles want to interact with other gas particles (**intermolecular attractions**)
- This interaction causes the particles to come closer together and the **volume of the container decreases**

Ideal vs Real Gas in a Constant Volume Container (Pressure in the Container Can Change)



Ideal Gas

- Ideal gases have no intermolecular forces

Real Gas

- Real gases **have intermolecular forces**
- The intermolecular forces between real gas particles causes real gas particles to spend (more/less) _____ time close to other gas particles compared to ideal gases
- As a result, real gas particles hit the walls of the container (more/less) _____ and with (more/less) _____ force
- The pressure inside the container for a real gas is therefore (higher/lower) _____ than the pressure inside the container for an ideal gas!

WIZE CONCEPT

Real gases have intermolecular attractions and this results in the pressure of the container decreasing.

Van der Waals Equation

ideal gases $\rightarrow PV = nRT$

$$\boxed{\text{real gases} \rightarrow \left(P + \frac{a}{V^2} \right) (V - nb) = nRT}$$

For real gases, we need to adjust the ideal gas law. The adjusted version is called the **Van der Waals equation**.

Real gas particles are attracted to each other (they have intermolecular attractions).

- In the Van der Waals equation, to "adjust" the equation to be more like ideal gases there is a constant, "a"
 - **a=attractive force correction**
 - **a increases with increasing intermolecular attraction between molecules**

Real gas particles also have volume.

- In the Van der Waals equation, to "adjust" the equation to be more like ideal gases there is a constant, "b"
 - **b=bulkiness correction**
 - **b increases with molecular weight** of gas molecule (more volume)

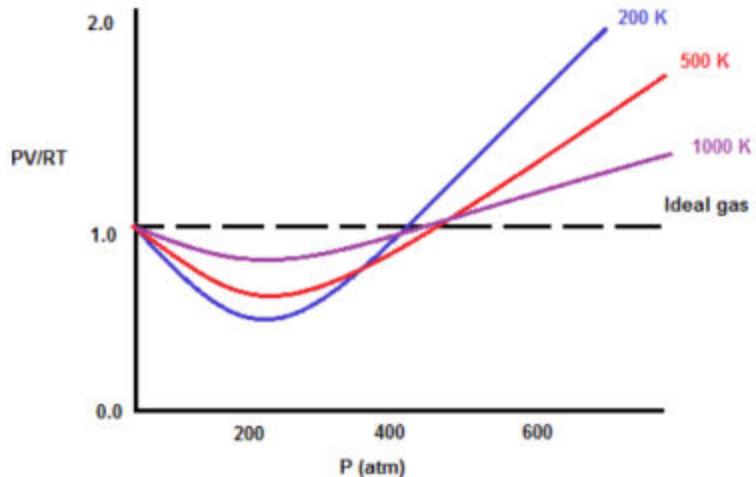
Note: a and b are Van der Waals constants and experimentally determined parameters that are different for every gas.

Real Gases vs Ideal Gases & The Compressibility Factor

Review: There are two main differences between real gases and ideal gases: (**you should know this!!)

Review: Real gases are more likely to deviate from ideal behaviour under what conditions? (**you should know this!!)

At _____ pressure and _____ temperature real gases deviate most from ideal gas behaviour.



Let's interpret the graph based on our knowledge!

- $PV=nRT$
- $PV/RT=n$ so the y axis is n and on the graph we have ideal gas labelled as 1 mol
- In lower temperatures did we say we would see more/less of a deviation from an ideal gas?
 - This is why we see the 200K line curve farthest away from the horizontal line representing an ideal gas!

Compressibility Factor (Z)

- We can calculate how close a real gas behaves like an ideal gas by measuring the **compressibility factor, Z**.
- For an **ideal gas, $Z=1$**
- In this equation V is the molar volume (V/n)

$$Z = \frac{PV}{RT}$$

3.11.3

Practice: Deviating from Ideal Gas Behavior

Out of the following, which gas would deviate the most from ideal behavior at a low temperature and high pressure?

Cl_2

CO_2

H_2

N_2

3.12 Solutions and Molarity

3.12.1

Classifying Matter

Matter refers to either a solid, liquid, or gas.

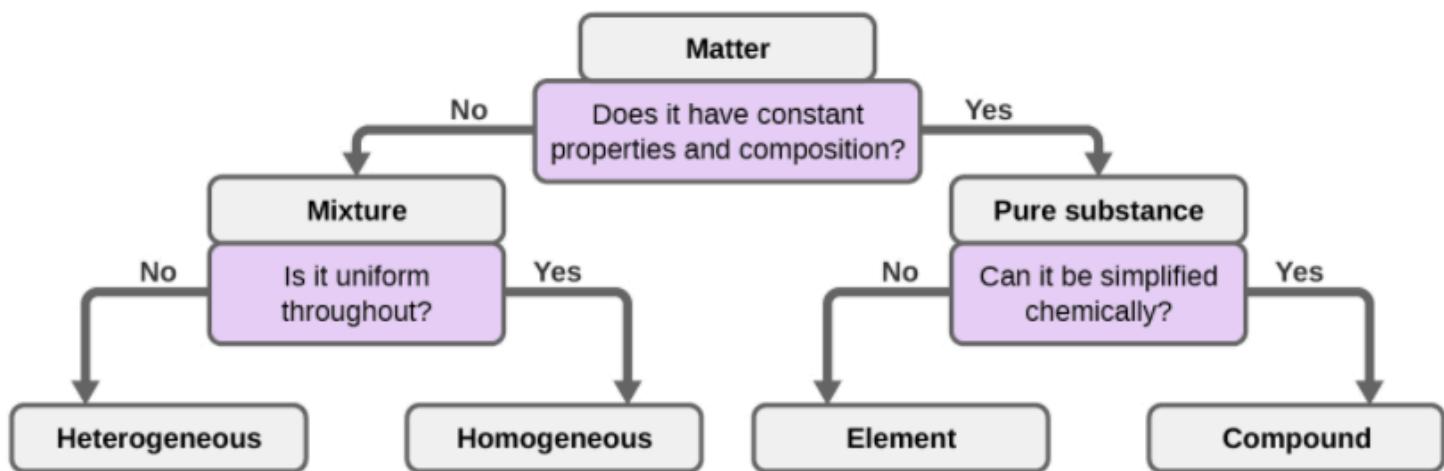


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We can either have **heterogeneous matter** or **homogeneous matter**:

- **Heterogeneous matter:** non-uniform composition
- **Homogeneous matter:** uniform in composition
 - Heterogeneous matter can be separated into homogeneous matter
 - Sometimes you can only tell if something is homogeneous or heterogeneous matter by looking at it under the microscope.
Example: if we just look at blood with our eyes, it looks like it is made of one thing, but if we look at it under a microscope, we can see that there are many components-white blood cells, red blood cells, etc so blood is actually

Example:

Is this a heterogeneous mixture or a homogeneous mixture?



Homogeneous matter can be either a **solution/mixture** or **pure substance**:

- **Solution:** is a homogeneous mixture because it looks like it is uniform in composition, but it is really made of many different things

Examples: tea! (has water and tea mixed together)

- **Pure substance:** cannot be further purified
 - Since a solution is made of more than one component, you can separate it into pure substances (these methods are not perfect!)

To separate a solution to try and get something in it's pure form you could:

- Use **filtration-separate based on size** (bigger components are held back in a net, and smaller components able to filter through)
- Use **distillation-separates based on volatility** (or boiling point)
 - liquids that are most volatile (or have the lowest bps) are separated out first
- Use **chromatography**-a mixture of components can be separated as the mobile phase moves up the plate.

Filtration:



Photo by Suman6395 / CC BY

Distillation:

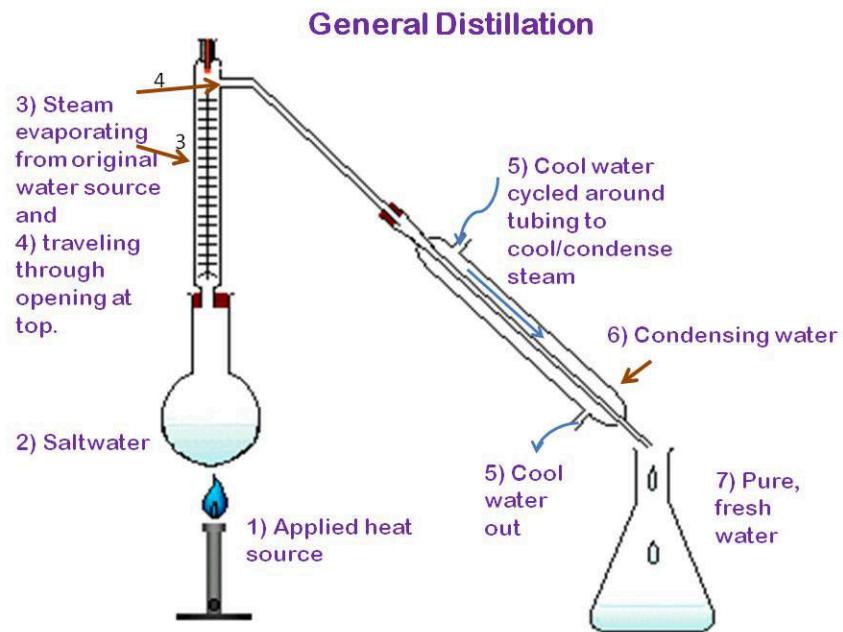


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Chromatography:

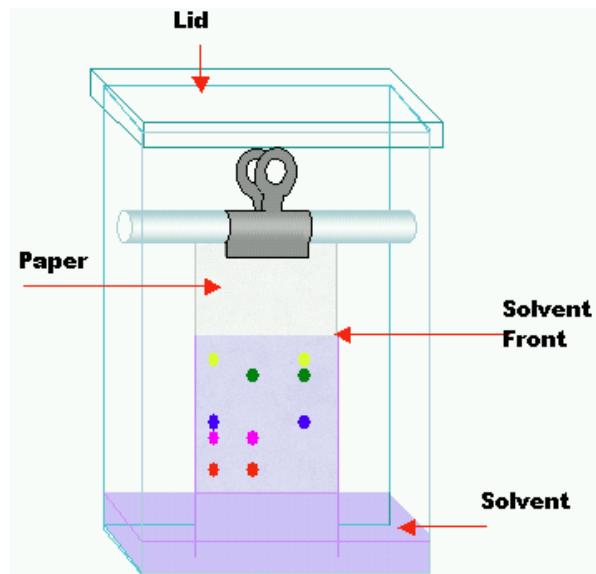


Photo by Theresa Knott / CC BY

Finally, **pure substances** can be either **elements** or **compounds**

- **Elements:** can't be subdivided into anything else
- **Compounds:** are made up of elements
 - We can react elements together to form compounds (ex. $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g})$)
 - We can separate compounds to form elements (one way to do this is electrolysis)

3.12.2

What is the molarity (M) of potassium ions in a solution prepared by mixing 100 mL of 0.250 M potassium bicarbonate with 400 mL of 0.400 potassium sulfate?

Answer

3.13 Electromagnetic Spectrum & Spectroscopy

3.13.1

Electromagnetic Spectrum

Up until this point in history, scientists were unsure if light shared the properties of waves or particles...there was some evidence for both!

The discovery of the electromagnetic spectrum showed the different wavelengths and frequencies that electromagnetic radiation could exist at. Since the different **types of radiation in this spectrum behave as waves**, and visible light is a part of this spectrum, it became clear to scientists that **light acts as waves!**

Types of Electromagnetic Radiation

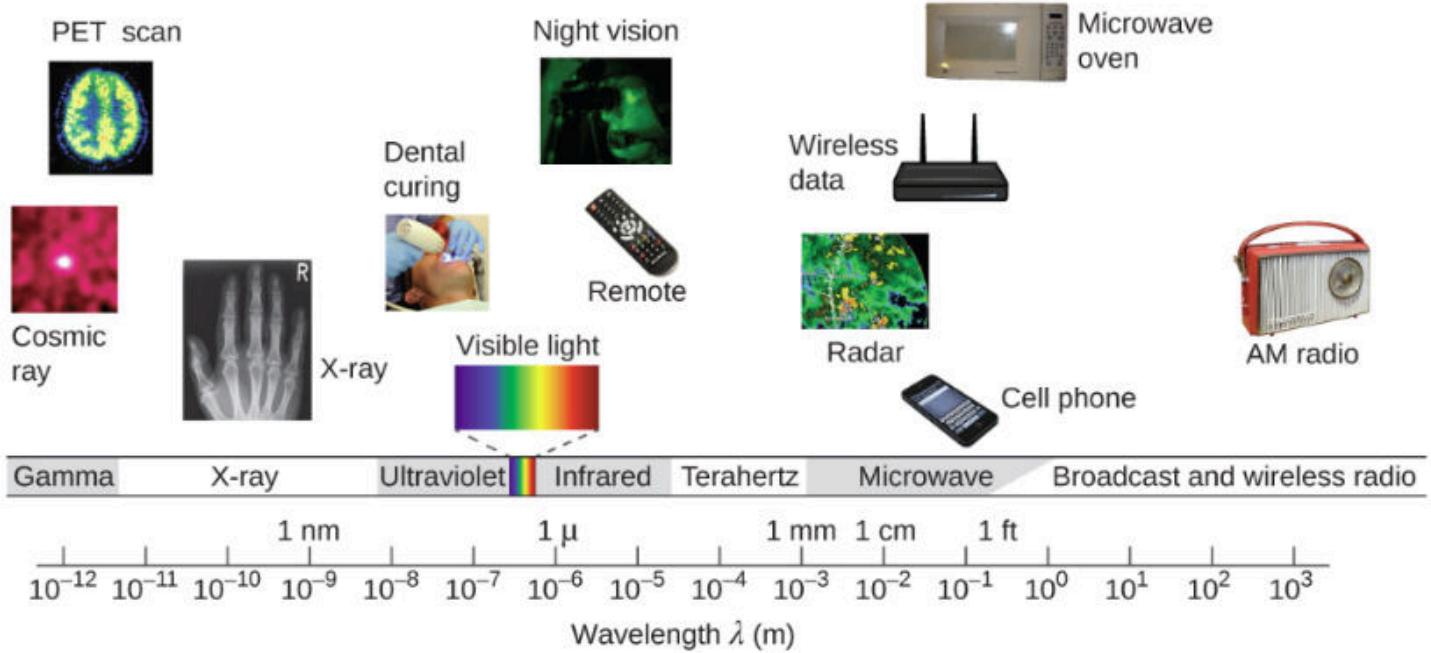


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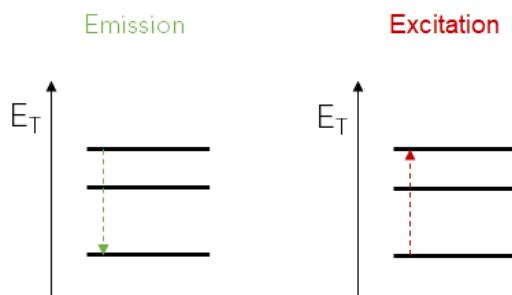
- In the diagram above, the electromagnetic radiation is organized from the types of radiation that are **highest in energy (gamma rays)** to the types of radiation that are **lowest in energy (radiowaves)**
 - You might recognize a lot of common household objects on the side of the spectrum with lower amounts of energy!

3.13.2

Spectroscopy and Transitions

- Remember that transitions in the UV-Visible range are _____ transitions.
- While IR spectroscopy involves _____ transitions.

UV-Vis



- A colored dye has a maximum absorbance at 500nm. How do the kinetic and potential energies of the electrons change when light is absorbed at 500nm?

Light is absorbed (electron going from lower to higher energy level) so PE _____ and KE _____ because the electron is in the higher energy level (moving slower).

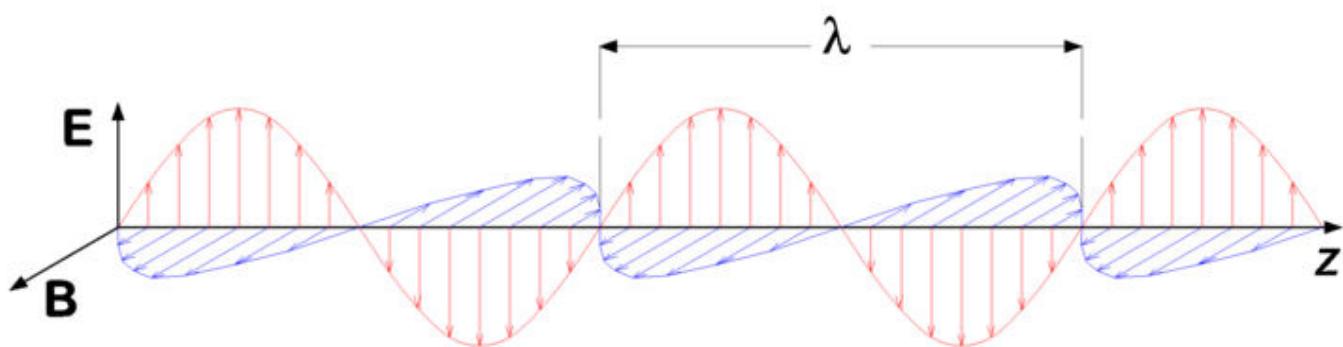
3.14 Photon Equations

3.14.1

Photon Equations

Electromagnetic Radiation Waves

- Light (or electromagnetic radiation) behaves as a wave
- A wave is a self-propagating transverse oscillation of electric and magnetic fields



where E is the electric field and B is the magnetic field

- To describe waves, we can use the following terms:
 - Amplitude (A): the height of a wave crest or depth of a trough
 - Wavelength (λ): The distance a wave travels in one cycle (m)
 - Frequency (v): cycles per second (s^{-1} or units **Hz**)
 - Energy (E): units are Joules (**J**)

Energy of a Photon Equations

$$E = \frac{hc}{\lambda} = hv$$

where ***h*** is Planck's constant: 6.626×10^{-34} Js
and ***c*** is the speed of light: 3.0×10^8 m/s

 **WIZE TIP**

You should memorize the value for the speed of light (*c*) = 3.0×10^8 m/s as that value is not always provided on exams.

Photon Equations (Cntd.)

Key Relationships Between Variables in Photon Equations

$$E = \frac{hc}{\lambda}$$

- Are the E and λ of a photon directly or inversely proportional? _____ proportional

$$E = hv$$

- Are the E and v of a photon directly or inversely proportional? _____ proportional

$$\frac{hc}{\lambda} = hv$$

- Are the wavelength and frequency of a photon directly or inversely proportional?
_____ proportional

Summary

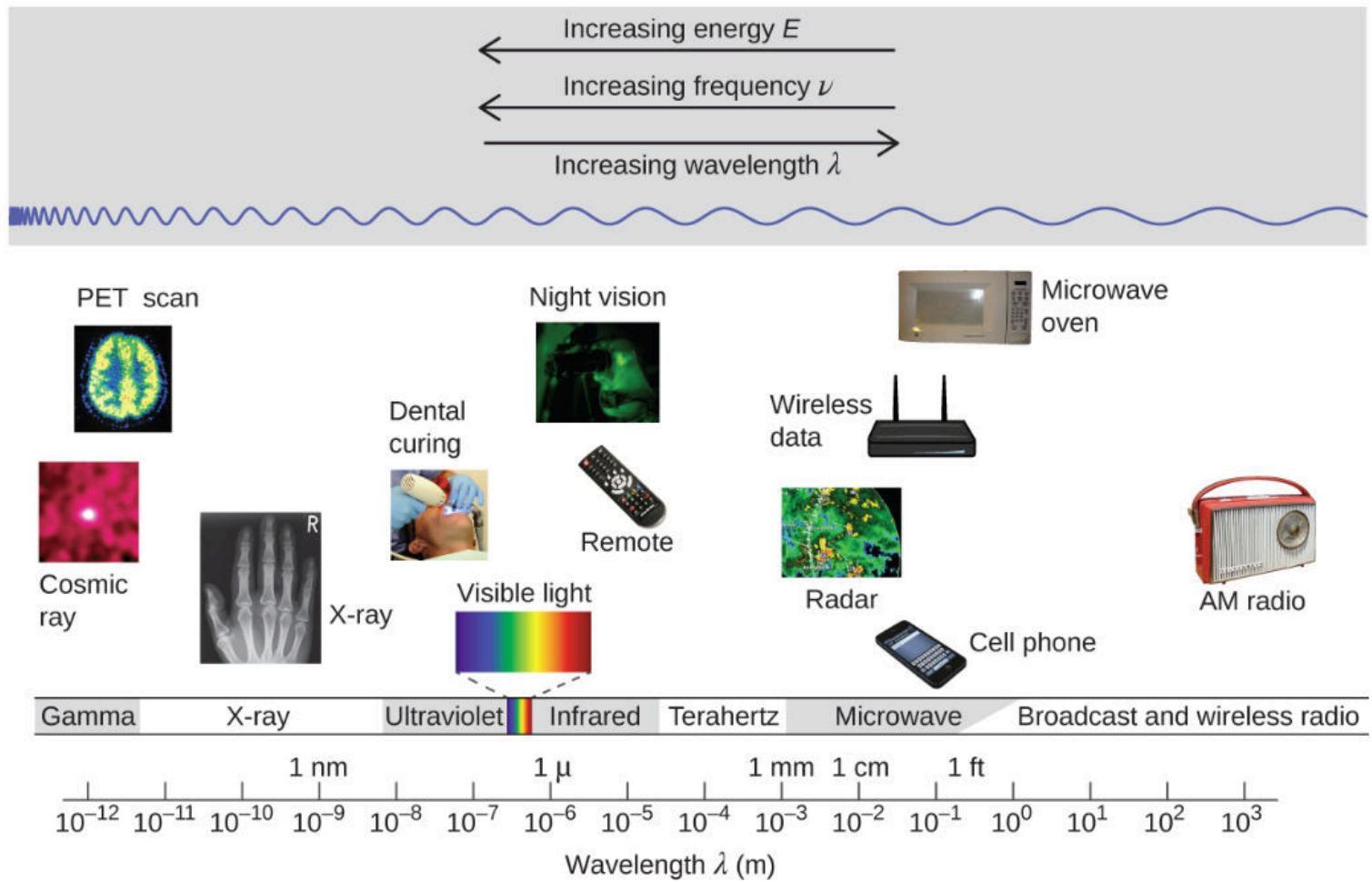


Photo by Rice University/ CC BY

Example

Example: Calculating Energy of a Photon Given Wavelength

A photon of light was found to emit light at 723 nm. How much energy is associated with this light?
 $h=6.626 \times 10^{-34} \text{ Js}$

! WATCH OUT!

Don't forget that for the photon equations **E must be in joules (J) and λ (wavelength) must be in m (meters) not nm (nanometers)!**

- If you are given a wavelength value in nm, you would have to convert it to m before plugging it into an equation.

Remember: $1\text{nm} = 1 \times 10^{-9}\text{m}$

Example: Calculating Frequency of a Photon Given Energy

Calculate the frequency of a photon with an energy of 236 kJ.

$$h=6.626 \times 10^{-34} \text{ Js}$$

3.14.5

Practice: Different Types of Electromagnetic Radiation

Rank the following types of electromagnetic radiation from lowest to highest energy.

Radio Waves, Infrared Radiation, X-rays, Microwaves, Visible Light, UV Radiation

X-rays < Infrared Radiation < Microwaves < UV Radiation < Visible Light < Radio Waves

Radio Waves < Microwaves < Infrared Radiation < Visible Light < UV Radiation < X-rays

Radio Waves < Infrared Radiation < Microwaves < Visible Light < UV Radiation < X-rays

Radio Waves < Infrared Radiation < Visible Light < Microwaves < UV Radiation < X-rays

3.14.6

Practice: Calculating Frequency of a Photon Given Wavelength

A photon of light has a wavelength of 700 nm. What is its frequency?

$4.29 \times 10^{-11} \text{ s}^{-1}$

$6.72 \times 10^{12} \text{ s}^{-1}$

$4.29 \times 10^{14} \text{ s}^{-1}$

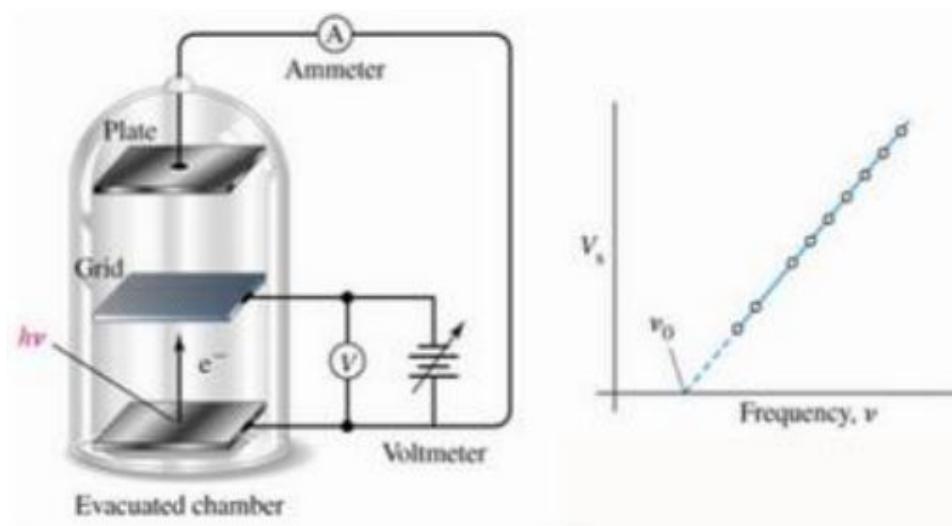
$7.23 \times 10^{16} \text{ s}^{-1}$

3.15 The Photoelectric Effect

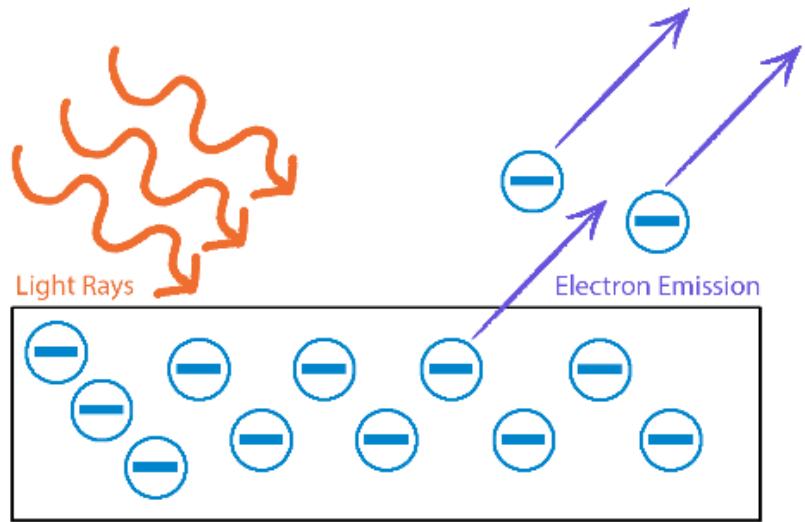
3.15.1

The Photoelectric Effect

In 1905, Einstein conducted his **photoelectric effect experiment**.

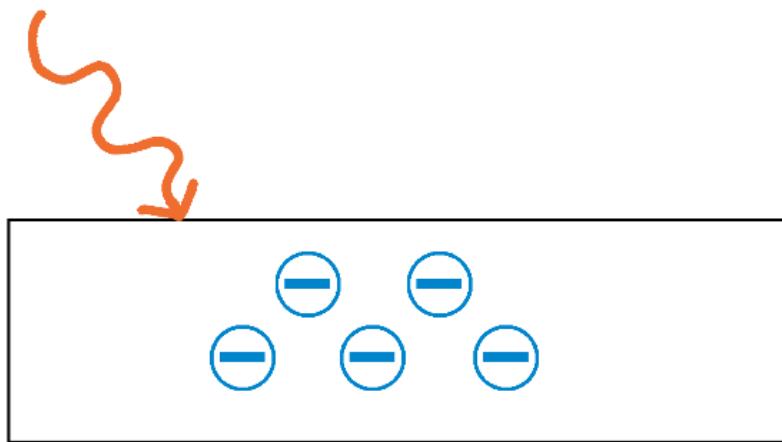


- The apparatus is made of a metal grid in which light of different intensities and frequencies was shown onto it.
- The ammeter (A) measures current (the **number of electrons emitted**).
- The voltmeter (V) measures the stopping potential (the **energy of the electrons emitted**).

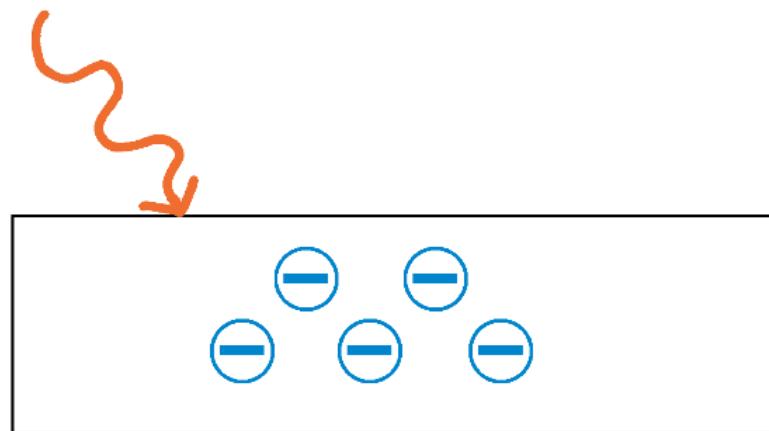


Experimental Findings

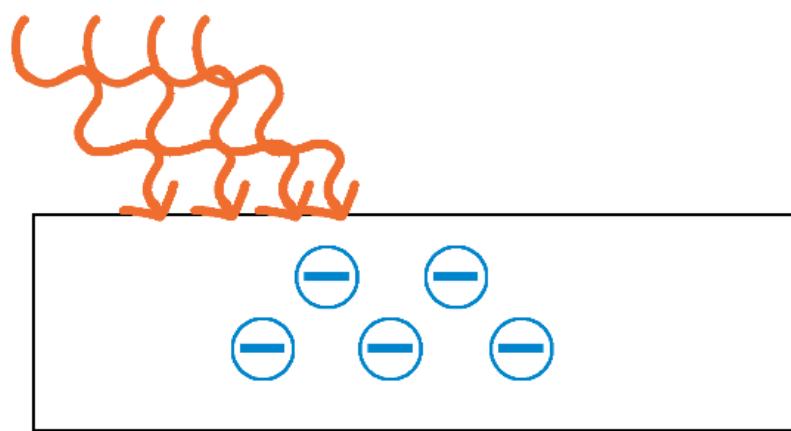
1) 1 photon with **sufficient energy** to eject an electron:



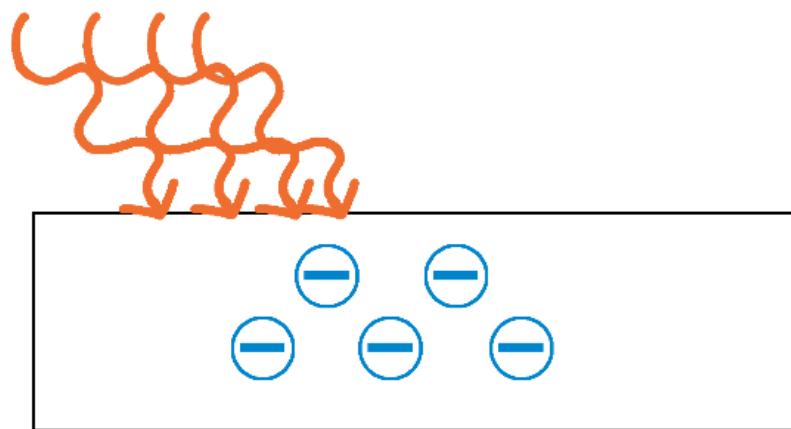
2) 1 photon with **insufficient energy** to eject an electron



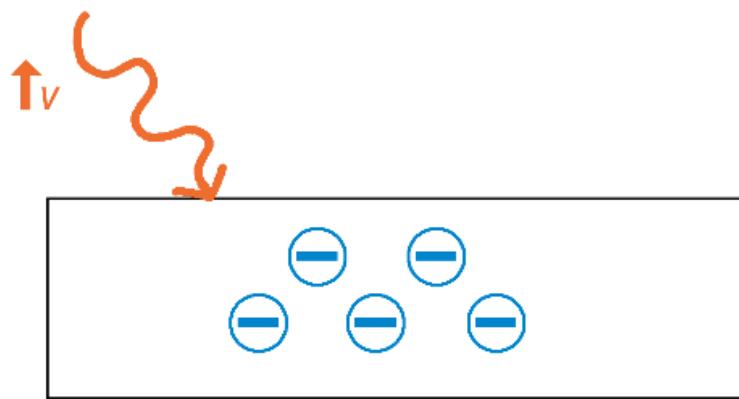
3) Multiple photons hitting the metal, **individually they have insufficient energy** to eject an electron but if you add up the their energy they have more energy than required to eject an electron



4) Multiple photons, each with sufficient energy needed to eject an electron hitting the metal



5) A photon with an **increased frequency** than the photons with minimum sufficient energy to eject an electron



Summary of the Findings:

- Electrons must be ejected by 1:1, photon : electron collisions
- Increased intensity of the light (more photons for a given part of the light beam), increased the number of electrons emitted (if each of the photons had sufficient energy to eject electrons)
- Increased frequency of the light, increased the energy of the electrons emitted

Before this experiment, people thought of light only as WAVES, but Einstein showed that light could also be thought of as particles (photons).

- When particle A hits particle B, particle A transfers its momentum to particle B. This is not the case for waves!
- The photoelectric effect showed that light can behave like particles since photons of light can hit a solid substance and cause electrons to be emitted

3.16 Beer-Lambert Law

3.16.1

Beer-Lambert Law

The Beer-Lambert law applies to any absorption spectroscopy but it is most typically used in UV-vis absorption spectroscopy.

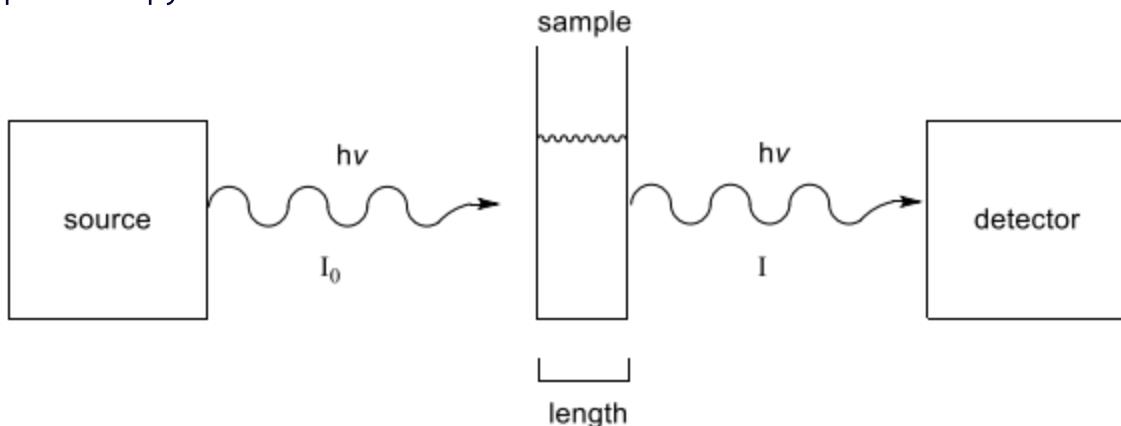


Figure 3.2.1 Schematic of a UV-vis absorption spectrometer

- The Beer-Lambert law is shown below. Where A is absorbance, T is transmittance, I_0 is the intensity of the radiation before it hits the sample, I is the intensity after passing through the sample, ϵ is the molar absorptivity constant, c is the concentration of the compound and l is the length of the sample cell.

$$A = \log_{10} \left(\frac{I_0}{I} \right) = \epsilon cl$$

- ϵ , the molar absorptivity is a constant but it is unique for each compound. The larger the molar absorptivity value the more strongly coloured that compound is.

3.16.2

A sample of ferrocene absorbs half the incident light at 500nm in a 1.0 cm cell. If the same solution is transferred to a 5.0 cm cell what percentage of light will be absorbed?

1. Set up an expression for each of the two cases described in the question

2. The sample has not changed so ϵc will be the same for both cases. We can substitute ϵc from the first case into the second giving

3. Solve for I

4. How much light was absorbed?